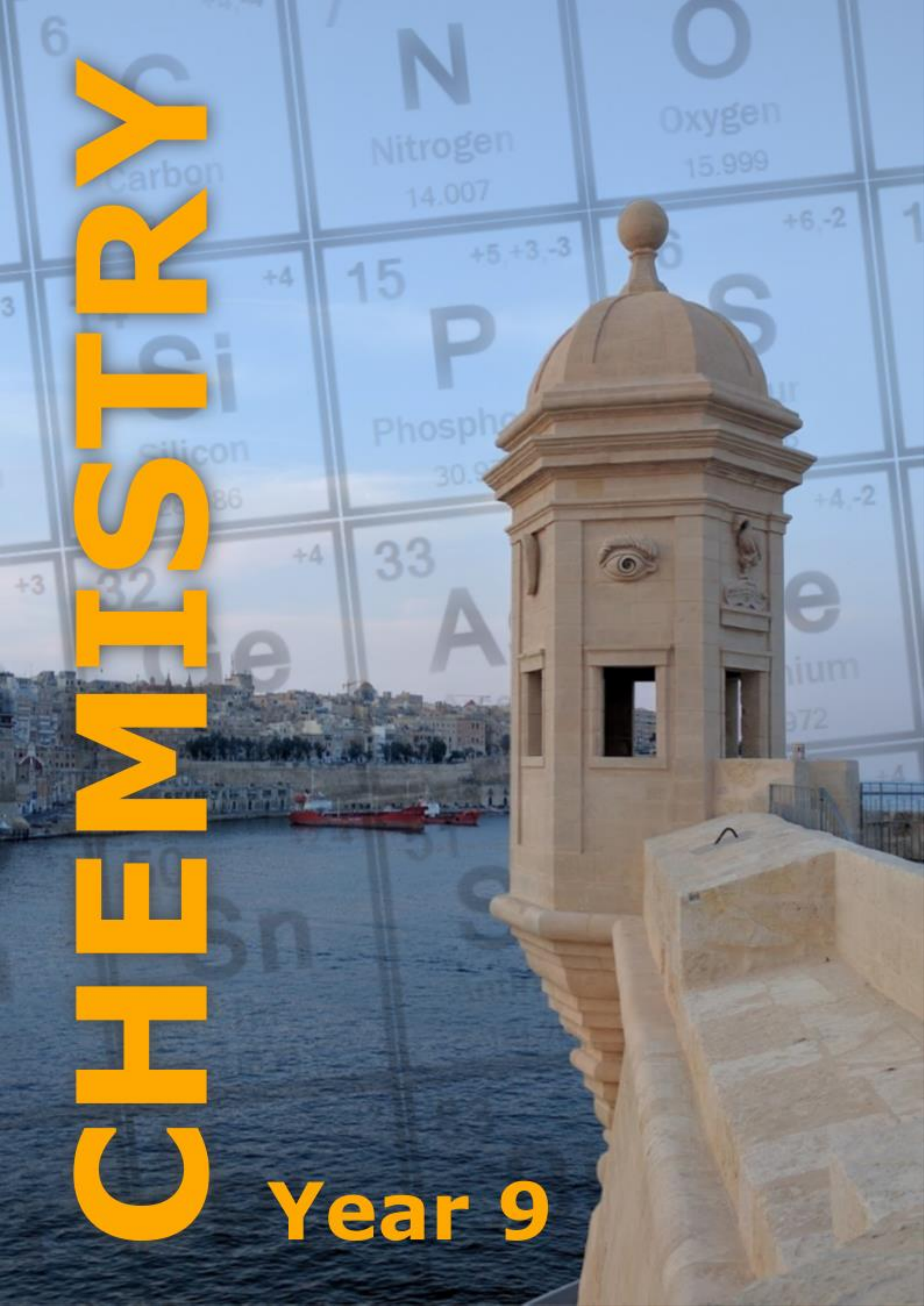
CHEMISTRY

Year 9



This textbook will be available in digital format (PDF) only. It is intended to be used by secondary school students and teachers in Malta and Gozo. It leads students to sit for the SEC06 Chemistry examination from 2025 onwards.

Version history

Version 1.0 distributed in January 2022.

Version 1.1 distributed in August 2023.

- Contains updates related to changes made to the SEC06 Chemistry syllabus.
- Some areas have been rearranged and/or reworded to reflect feedback received from teachers.

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How to use this e-book

This e-book is closely linked to the SEC Chemistry Learning Outcomes Framework syllabus which can be downloaded at: https://www.um.edu.mt/matsec/syllabi/secsyllabi2025

Each chapter represents the work expected in one learning outcome. Each chapter, except for chapter one, starts with a graphic that indicates the subject focus, which the learning outcome is part of.

Chapters are split into sections that deal with groups of related assessment criteria. These assessment criteria are declared so that students can focus on what needs to be learned.

Hyperlinks link to content that is online and serve to further develop students' knowledge. Online content is not examinable however students are encouraged to follow these links to further their understanding of the role of chemistry in the real world.

An internet connection is required to access video clips, animations, and online content. When using the digital version of this e-book, QR codes can be clicked to access the online resources.

When using this e-book on a computer, use the navigation feature provided by your PDF viewer.

Throughout this e-book you will find one of five types of textboxes shown below:

This textbox shows **exercises** related to content that students can work on their own or with their teachers.

This textbox shows **snippets** of information that are **NOT** meant to be studied for your examination, but which give extra information related to discoverers, discoveries, and further relevant content.

This textbox shows a QR code which links to an **online video clip**. Click on the QR code when using the digital version to access the resource.

This textbox shows a QR code which links to an **animation**. Click on the QR code when using the digital version to access the resource.

This textbox shows a QR code which links to **further reading**. Click on the QR code when using the digital version to access the resource.

Also,

Definitions and **important information** are presented on a blue background.

Contents

1 Understanding how chemistry works and is communicated (LO 1)

1.1 Understanding the scientific process

By the end of this section, you should be able to:

- State that scientific knowledge changes with new evidence/ observations/ experiments. (1.1*a*)
- Distinguish between *a* fact, *a* hypothesis, and *a* theory. (1.2*a*)
- Discuss briefly the meaning of science in terms of its healthy scepticism, aimed objectivity, and the value of physical (observable/ measurable) evidence. (1.3*a*)

1.1.1 The nature of scientific knowledge

People once thought that the Earth was flat. This was the dominant view before the 2nd century BC. From the evidence collected over the centuries this idea changed and we know that the Earth is a planet. Scientific knowledge is never absolute, it can change when new discoveries are made. So, one will wonder, what is science? How is scientific knowledge produced?[1]

Science consists of **a body of knowledge** and the **process** by which that knowledge is produced.

We tend to think that science is a body of knowledge since we are presented with many different facts e.g., the Earth is round, electrons are negatively charged, etc. However, it is important to be aware how scientific knowledge is generated. This is known as the **scientific process** which is a way of building knowledge and making predictions about the world in such a way that they are **testable**. The question of whether Earth is flat or round, could be put to the test. This could be studied through multiple lines of research, and the evidence is evaluated to determine whether the evidence supports a round or flat planet. Different scientific disciplines typically use different methods and approaches to investigate the natural world, but **objective testing** lies at the core of scientific inquiry for all scientists.

As scientists analyse and interpret their data, they generate hypotheses and theories which help explain their results and place them in context of the larger body of scientific knowledge. These different kinds of explanations are tested by scientists through additional experiments, observations, modelling, and theoretical studies. Thus, the body of scientific knowledge builds on previous ideas and is constantly growing. It is deliberately shared with colleagues through the process of **peer review**, where scientists comment on each other's work, and then through publication in the scientific literature, where it can be evaluated and integrated into the body of scientific knowledge by the larger community. Scientists' beliefs, knowledge and experiences influence their observations and interpretations. Observations are never neutral but are always interpreted from within theoretical frameworks which introduce an element of subjectivity. Scientists attempt to be "open minded and apply mechanisms such as peer review and data triangulation to improve objectivity" (Chen, 2006, p. 806).

One of the hallmarks of scientific knowledge is that it is subject to change, as new data are collected, and reinterpretations of existing data are made. Scientific knowledge, therefore,

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[1] Adapted from https://www.visionlearning.com/en/library/Process-of-Science/49/The-Nature-of-Scientific-Knowledge/185

although reliable, does not lead to absolute truth and is subject to change in the light of new evidence or interpretations or social and cultural changes. Changes can take place in two ways:

- **evolutionary** when theories evolve with new evidence or interpretations,
- **revolutionary** when theories are completely replaced by new theories following new evidence or interpretation.

Major theories, which are supported by multiple lines of evidence are rarely completely changed, but new data and tested explanations add nuance and detail.

Scientific knowledge helps people understand and answer questions about the world around us. Science is built on several facts conceived through several theories. Theories are generated through evidence, observation, and experimentation. This is the nature of scientific knowledge. If the theory supports the evidence collected through observation and experimentation, and is also able to make predictions, then facts and knowledge are retained. If not, new experiments have to be constructed. The observations and evidence collected through these new experiments may change the nature of scientific knowledge. For this reason, scientific knowledge is dynamic and is constantly developing. Generation of scientific knowledge also involves human creativity at all stages of investigations; prior to, during collection and interpretation of data, and during generation of conclusions. Scientists also use imagination together with logical reasoning to create scientific knowledge.

Finally, science is a human endeavour which is affected by the culture and society it is accomplished in. Cultural values as well as power structures such as religion, political and economic factors, influence how science is conducted and interpreted. Such properties are exhibited through public funding for scientific research.

The following table shows the differences between the terms: fact, hypothesis, and theory.

Term	Meaning	Example
Fact	A phenomenon that has been observed to be true and has been repeatedly confirmed to be so.	Mercury is the only metallic element that is liquid at standard temperature and pressure.
Hypothesis	A tentative explanation about an observation. A hypothesis triggers off a series of experiments/investigations.	Soluble solids would dissolve faster if the temperature of the water is increased.
Theory	An explanation of why a phenomenon occurs, which has been substantially confirmed by facts and hypotheses.	The Kinetic Theory of Matter

Scientists continuously utilize critical thinking and reasoning skills in order to evaluate existing evidence. This is referred to as **healthy scepticism**. This dynamic, logical approach to science enhances the validity of scientific results and deepens understanding. Scientific knowledge is built through an accumulation of observations and experimentation. For this reason, it is healthy to adopt an objective approach in science. When planning and conducting experiments, gathering observable and measurable evidence, reporting observations and results, and communicating experimental outcomes; scientists need to do so within a frame of mind free from bias and personal interest.

The scientific method.

1.2 Health and safety

By the end of this section, you should be able to:

- Follow health and safety regulations. (1.1e)
- State health and safety considerations. (1.2e)
- Evaluate an experiment in terms of health and safety. (1.3e)

Studying chemistry involves a good amount of laboratory work where glassware, fire, electricity, and chemical substances are used. Always follow the instructions provided by your teacher while performing experiments in the laboratory.

1.2.1 General rules

When you are in a chemistry laboratory, follow the rules below:

- Always wear eye protection.
- Always wear a laboratory coat.
- Wash your hands regularly.
- Keep your work area clear and tidy.
- Never eat or drink in the laboratory.
- Check the SDS (**S**afety **D**ata **S**heet also known as MSDS that is **M**aterial **S**afety **D**ata **S**heet) of chemicals before starting work and take the necessary precautions.
- Clean up all spillages immediately.
- Keep long hair tied back.
- Wear appropriate clothing that does not leave skin exposed (e.g. shorts, skirts, sandals).
- Label bottles and other containers clearly and correctly. Do not write directly on to bottles where the ink can run or be easily removed, e.g. by solvent.
- Keep walkways clear.
- Check glassware for damage, cracks, and broken edges before use.
- Make sure all equipment is in good working order and fit for use.

(List adapted from: rsc.org)

1.2.2 Glassware

Glassware refers to any equipment that is used in the laboratory which is made of glass. Laboratory equipment is made of glass as it is transparent, resists heating and easy to clean. However, glass is brittle and can break if not handled properly. The problem with glass is that as it breaks, it forms sharp edges that can penetrate one’s skin causing serious injury.

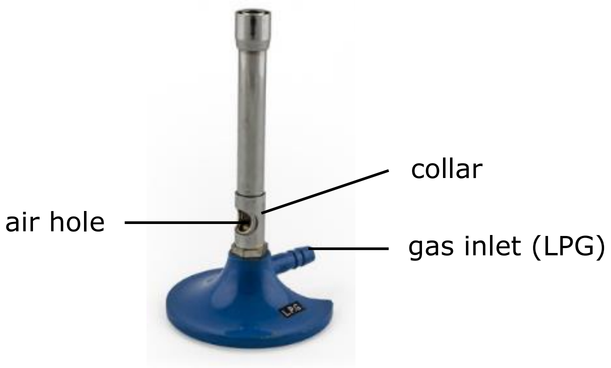
In the event that glassware is broken, do NOT handle the broken pieces bare handed. Inform your teacher that glassware has been broken.

Your teacher will take note of the equipment broken, wear protective gloves to handle the larger pieces and then use a brush and dustpan to remove the smaller pieces. The broken glassware will then be placed in a container marked accordingly to prevent injury to other people.

1.2.3 Fire

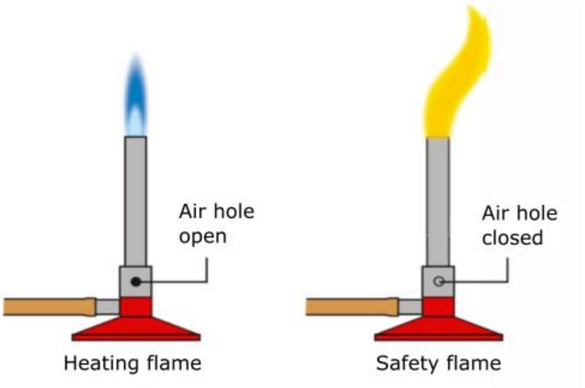
On some occasions, chemical reactions need to be heated so that a chemical reaction can proceed at a desired rate. In a chemistry laboratory this is often done using a Bunsen burner.

Since a Bunsen burner produces a naked flame, it is important to handle this equipment with care. Wear eye protection. People with long hair must pull it up to prevent their hair from catching fire. Avoid leaning over the bench while working especially if a Bunsen burner is in use!



The Bunsen burner is designed such that it has a collar that one can rotate to control the amount of air that is mixed with the LPG by controlling the size of the air hole. This is a safety feature that is inbuilt so that when the burner is not in use the air hole is closed to produce a yellow flame, which is more visible than a clean blue flame.

When the air hole is opened, a blue Bunsen burner flame is produced that must be used while heating since it produces the hottest flame and does not produce soot, which makes glassware black.



Some chemical substances are said to be **flammable** or **inflammable**. These two words both mean that a substance can catch fire easily. Examples include ethanol (present in alcoholic beverages such as vodka) and propanone (also known as acetone and present in nail polish remover).

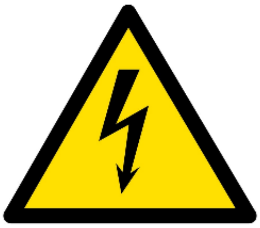
Flammable substances should never be heated with a Bunsen burner flame as they will catch fire. These substances are marked with a flammable sign to inform the user of its fire risk. However, they can be heated either using a heating mantle or using a water bath.

Heating mantle

Water bath

How to use the Bunsen burner.

1.2.4 Electricity



There will be instances when the use of electricity to perform experiments will be required. Electricity comes in two forms, either alternating current (AC) or direct current (DC).

The electricity supply in the laboratory and at your home is supplied via three pin sockets which is of the AC type. It provides considerable power which is why it needs to be used with care.

Usually, the type of electricity that is used by students in the laboratory is of the DC type which is provided by a special transformer that is connected to the AC laboratory power supply to provide a maximum of 20 V (Volt) DC.

The red socket is the positive outlet while the black socket is the negative outlet. These power supplies provide safe DC power. However, caution must be always taken.

1.2.5 Fume cupboard

A fume cupboard such as the one on the right, is an enclosed chamber that is open from one side only. It has a powerful pump that sucks air from the fume cupboard and passes it through a special filter to remove any toxic gases.

Fume cupboards are used widely in laboratories and are designed to capture and remove air-borne hazardous substances generated during laboratory experiments (e.g., gases, vapours, aerosols, and particulates/dust).

Experiments that involve toxic reactants and/or products are carried out in the fume cupboard.

1. State why the following laboratory rules need to be followed:
   a. Always wear eye protection.
   b. Always wear a laboratory coat.
   c. Never eat or drink in the laboratory.
   d. Keep long hair tied back.
   e. Wear appropriate clothing that does not leave skin exposed (e.g. shorts, skirts, sandals).

2. State the position that the collar of a Bunsen burner must be in to produce a blue flame.

3. Give an alternative name for the yellow flame produced by a Bunsen burner. Explain why the flame turns yellow when the Bunsen burner collar is closed.

4. Name the two types of electricity that are available. Which of them is used during science experiments in the lab?

1.2.6 Chemical substances

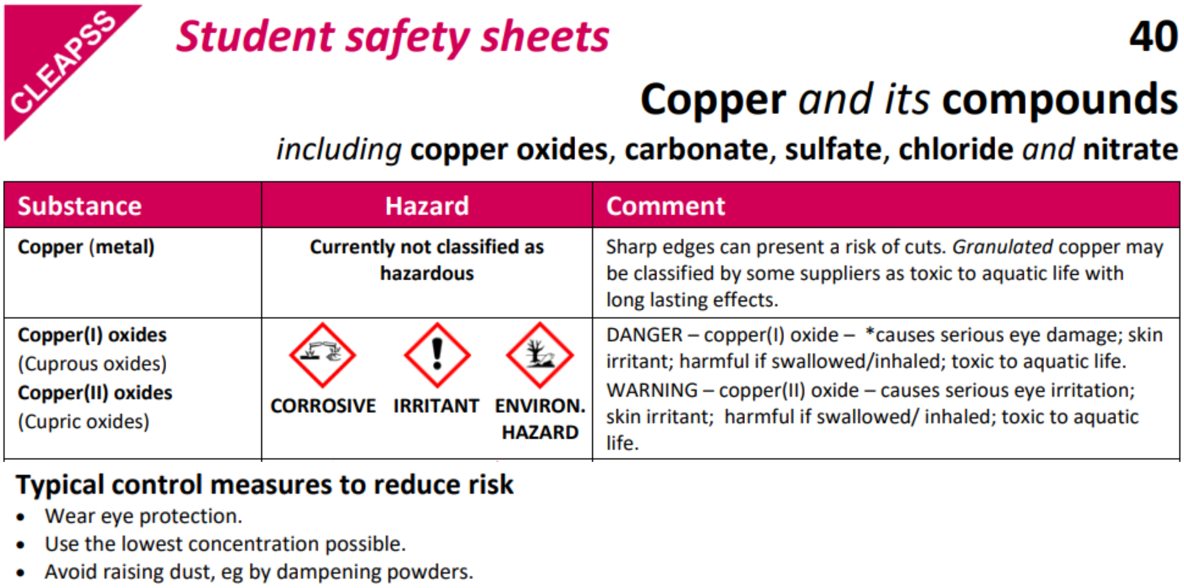
Some chemical substances are inert, while others have properties that might cause harm if inhaled, swallowed, or allowed to touch one's skin. This does not mean that as junior chemists you should avoid using these substances. Being informed of their potential hazardous properties puts one in a better position to take the necessary precautions to perform experiments safely.

For this reason, an international system of signs is in place to provide information of the hazards chemical substances might carry. You will find these symbols on the original packaging of materials and on laboratory reagent bottles. The table below shows the new and old hazard symbols.

Hazard	New Symbol	Old symbol
Corrosive		
Acute toxicity		
Serious health hazard		
Health hazard		
Flammable		
Oxidising		
Explosive		
Gas under pressure		No symbol available.
Hazard to the environment		

All chemical substances have a **Safety Data Sheet** (SDS). This document can easily be accessed online by typing the name of the substance followed by SDS. SDS documents provide all the information required by chemists so they are informed of the hazard chemicals may pose.

Knowing about the hazards in advance, makes working with chemicals that much safer as one would know how to handle these chemicals safely. All SDS documents have the same sixteen section format so that chemists all over the world know what to expect when reading these documents. All chemistry labs are required to have an easily accessible folder containing SDSs for all chemicals that are used in that laboratory. Meanwhile, your teacher will provide you with a simplified, student friendly version of these SDSs that you can use during your laboratory work. Below is an excerpt of one of these SDSs.



CLEAPSS

Student safety sheets **40**

Copper *and its* **compounds**

including **copper oxides, carbonate, sulfate, chloride** *and* **nitrate**

Substance	Hazard	Comment
Copper (metal)	**Currently not classified as hazardous**	Sharp edges can present a risk of cuts. *Granulated* copper may be classified by some suppliers as toxic to aquatic life with long lasting effects.
Copper(I) oxides (Cuprous oxides) **Copper(II) oxides** (Cupric oxides)	**CORROSIVE** **IRRITANT** **ENVIRON. HAZARD**	DANGER – copper(I) oxide – *causes serious eye damage; skin irritant; harmful if swallowed/inhaled; toxic to aquatic life. WARNING – copper(II) oxide – causes serious eye irritation; skin irritant; harmful if swallowed/ inhaled; toxic to aquatic life.

Typical control measures to reduce risk

- Wear eye protection.
- Use the lowest concentration possible.
- Avoid raising dust, eg by dampening powders.

1. Give the name of the following hazard symbols:

2. Use the SDS for the following chemicals to complete the table below.

Chemical	Hazard	Safety precaution
Magnesium ribbon		
Hydrochloric acid (1.0 mol dm^{-3})		
Copper(II) sulfate crystals		

Crash Course Chemistry: Lab Techniques and Safety

1.3 Performing experiments

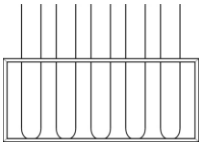
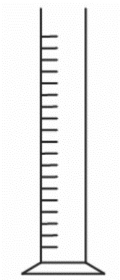
By the end of this section, you should be able to:

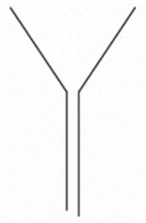
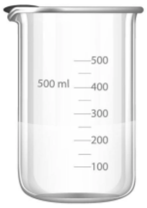
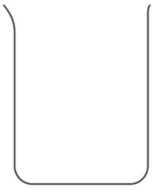
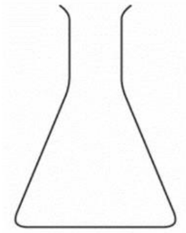
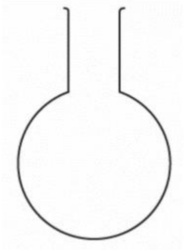
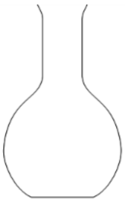
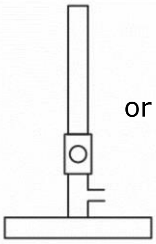
- Carry out, with supervision/ with limited supervision/ with no direct supervision, a written procedure for an experiment. (1.1h, 1.2h, 1.3h)

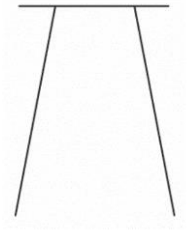
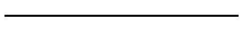
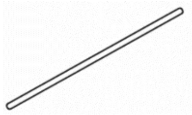
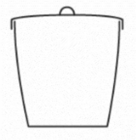
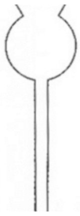
To carry out experiments safely and accurately, the instructions given in your experiment task sheet must be read and followed carefully. Failing to do so could result in your experiment not working properly or even worse, causing injury to you and others.

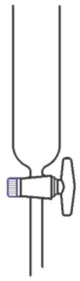
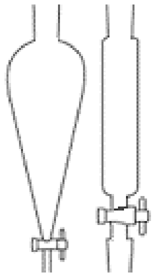
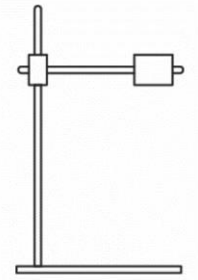
1.3.1 Laboratory apparatus

The first step towards performing experiments safely is to be familiar with the equipment being used. Knowing their names and what they are used for, is a must. This will allow you to follow the instructions in your experiment without misinterpretation. The table below shows the names and pictures of equipment that are commonly used during experiments. The table also includes how the equipment should be drawn in your diagrams, as well as what it is used for.

Name	Used for…	Picture	Diagram
Test tube / boiling tube	A test tube is used to mix chemical substances on a small scale while a boiling tube is used to heat substances to perform chemical reactions.		
Test tube rack	Holding test tubes.		
Test tube holder	Holding a test tube while it is being heated on a flame to avoid burning yourself.		Not usually shown in diagrams
Test tube cleaner	Cleaning test tubes.		Not usually shown in diagrams
Measuring cylinder	Measuring a specific volume of liquid.		

Name	Used for…	Picture	Diagram
Glass funnel	Pouring a liquid from one container into another to avoid spillage or used during filtration.		
Spatula	Moving solid substances from their container to a test tube, beaker, or flask.		Not usually shown in diagrams
Beaker	Holding and/or heating liquid substances.		
Conical flask	Holding and reacting liquid substances at room temperature. It is ideal for swirling its contents as it avoids spilling. DO NOT HEAT AS IT MAY CRACK.		
Round bottomed flask	Heating liquids/ powder on a Bunsen flame.		
Flat bottomed flask	Holding and reacting liquid substances at room temperature. DO NOT HEAT AS IT MAY CRACK.		
Bunsen burner	Heating containers. It includes a safety feature to produce a safety flame. It can be represented by the following symbol:		 or ↑ heat

Name	Used for…	Picture	Diagram
Tripod	Holding a container above a heat source such as a Bunsen burner. DO NOT TOUCH WHEN HOT!		
Crucible tongs	Handling hot objects such as crucibles as well as wire gauzes etc.		Not usually shown in diagrams
Wire gauze	Placed between a container and a tripod. It helps to distribute heat evenly. DO NOT TOUCH WHEN HOT!		
Glass rod	Stirring substances normally in a beaker.		
Watch glass	Holding a solid substance temporarily. It is not meant to be heated.		
Evaporating dish/basin	Heating solutions of salts in a process called evaporation to dryness.		
Crucible with lid	Heating solid substances to very high temperature. DO NOT TOUCH WHEN HOT!		
Pipe clay triangle	Holding a crucible on a tripod. DO NOT TOUCH WHEN HOT!		
Thistle funnel	Transferring a liquid into a reaction vessel. It must be fitted such that its leg is underneath the level of the reacting liquid.		

Name	Used for…	Picture	Diagram
Dropping funnel	Transferring a liquid into a reaction vessel and allowing control over the amount added.		
Separating / tap funnel	Separating immiscible liquids.		
Stand and clamp	Holding equipment at a height and angle as required.		
Thermometer	Measuring temperature.		

1.4 Report writing

By the end of this section, you should be able to:

- Structure a laboratory report in sections (1.2t).
- Write a scientific report for an experiment carried out (1.3t).

Practical work is important in Chemistry. Through experiments you will develop experimental skills and techniques such as handling apparatus, performing tests or procedures, identifying variables to alter or control, conducting observations and measurements, and tabulating data. Furthermore, during data processing you can plot graphs, work out calculations, look for patterns and trends, analyse and interpret data observed, draw conclusions and link to scientific knowledge, principles, and theory. Conducting experiments will help you to get a feel of the phenomena such as they can make the connections between observing concrete evidence and the more abstract ideas or theories.

The following template, published by SEC, shows the sections and respective notes that should be included in an experiment report. Third person past tense should be used in a report.

Write up of an experiment	
Date	Write the date when the experiment was carried out in the lab.
Title	The title indicates the links to particular assessment criteria as outlined in the curriculum.
Aim	The purpose of the experiment is clearly stated.
Apparatus	A list of apparatus and materials/chemicals used during the experiment.
Diagram	Clear diagram/s of the experimental setup are to be drawn and labelled in pencil. Diagrams should not be too small nor too large.
Procedure	This section will be given to the students.
Variables (if applicable)	A variables grid should be included specifying the independent variable (the one which is changed during the experiment) and the dependent variable (the one which is measured for a change in the independent variable). Other variables which are kept constant, to ensure fair testing, should be included.
Precautions	A list of precautions taken to improve the accuracy of the experiment. Each precaution needs to be supported with reason/s explaining why such precautions are taken.
Safety Considerations (if applicable)	A short paragraph/list that identifies safety considerations associated with the preparation and implementation of the experiments to prevent any accidents
Results and Observations	Depending on the nature and type of the experiment: • Observations can be written in paragraphs or in tables. • Observations are to be written in a sequential order as noted during the different stages of the experiment. • Numerical results should be tabulated. ○ Write the name of the measurement and its units in the column headers of the table of results. ○ Repeated readings should be taken when possible and recorded in the table. ○ Numerical values should be given to the same number of significant figures appropriate to the measuring device.
Processing data	• Graphs are a pictorial way of looking at a table of results. Patterns can be observed, and anomalous results can be identified. • Line graphs should include at least 5 data points. • Suitable scales should be chosen which makes it easy to plot data. • At least $^2/_3$ of the graph paper should be used. • Each axis should be labelled with the name and unit of the quantity being plotted. • The data points should be clearly marked, and the points are joined to have a line of best fit or a smooth curve.

	• The line must go through the origin for quantities which are directly proportional. • Data can also be presented in the form of bar graphs. • Gradient of line graphs are calculated, and answers are given with the appropriate units. • Show all steps in the calculations. • In working calculations, the answer should have the same number of significant figures as the measurements used in the calculation. • Avoid excessive rounding especially in calculations involving moles.
Discussion and Conclusion	Include the following points as applicable to the nature of the experiment. • A summary of the findings of the experiments and relate them clearly to the aim of the experiment. • A discussion of any patterns or trends in the data. • State any relationships discovered or confirmed between variables being tested in the experiment. • Compare numerical results with known values from data books and suggest any reasons for any differences. • A complete analysis or interpretation of observations noted in the experiment (including balanced chemical equations and ionic equations where appropriate). • Draw a conclusion based on experimental evidence and relate it to scientific knowledge, laws, and theory.
Evaluation	• Identify and comment on any sources of error in the experiment. • Discuss any difficulties encountered in carrying out the experiment and any precautions taken to achieve accuracy. • Suggest way/s of improving the experimental set-up and or results. • Suggest any other experiments which can be done to support the conclusions.

Source: https://www.um.edu.mt/__data/assets/pdf_file/0007/509191/SEC062025Resources.pdf

The following sample experiment report shows how a report should be written.

SAMPLE EXPERIMENT REPORT

Date: Write the date when the experiment was carried out.

Title: The effect of surface area on the rate of a reaction.

Aim: To determine the effect of changing surface area on the rate of a chemical reaction.

Apparatus and materials:

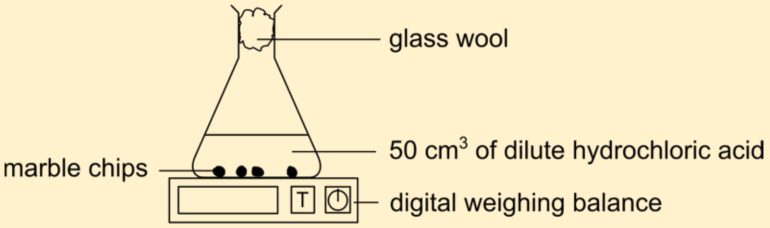
Measuring cylinder, conical flask, electronic balance, weighing boat, stopwatch, glass wool, pestle and mortar, calcium carbonate, dilute hydrochloric acid.

SAMPLE EXPERIMENT REPORT – Continued

Procedure:

1. Measure 50 cm^3 of dilute hydrochloric acid using a measuring cylinder.
2. Measure 1 g of crushed calcium carbonate and place in conical flask.
3. Pour the dilute acid in the conical flask and quickly close with a glass wool.
4. Switch on the stopwatch and measure the mass of flask and its contents at 1-minute intervals for 10 minutes.
5. Repeat the above steps using 1 g of larger lumps of calcium carbonate.

Diagram:



Variables:

Independent variable	Surface area of calcium carbonate
Dependent variable	Rate of loss of CO_2
Controlled variables	Mass of calcium carbonate, volume and concentration of acid, Temperature

Precautions:

- Volume of acid in the measuring cylinder was read at eye level to prevent parallax errors.
- The following variables: volume and concentration of acid, mass of calcium carbonate, and temperature, were kept constant for fair testing.
- A glass wool was used to allow carbon dioxide to escape but prevents the acid spray from escaping.

Safety considerations:

Chemical	Hazard	Safety precautions
hydrochloric acid 1.0 mol dm^{-3}	Not hazardous	Dilute acid may still cause harm to the eyes or the skin. Wear eye protection and use low volume and concentration.
calcium carbonate	Not hazardous	Wear eye protection and avoid raising dust.

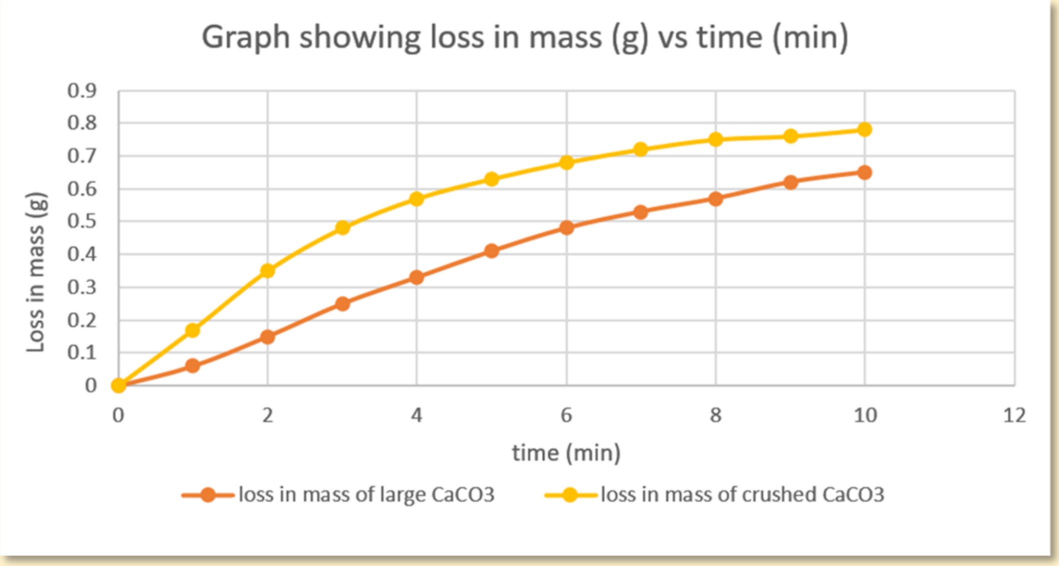
Results and observations:

When crushed calcium carbonate reacted with an acid faster effervescence took place. On the other hand, when larger pieces of calcium carbonate were added to the acid, slower effervescence was noted. In both experiments effervescence was fastest at the beginning. With time the rate of effervescence decreased in both reactions.

SAMPLE EXPERIMENT REPORT – Continued

Time (min)	Mass of container with large pieces of Ca CO_3 (g)	Mass of carbon dioxide (g)	Mass of container with small pieces of Ca CO_3 (g)	Mass of carbon dioxide (g)
0	55.25	0	55.24	0
1	55.19	0.06	54.07	0.17
2	55.10	0.15	54.89	0.35
3	55.00	0.25	54.76	0.48
4	54.92	0.33	54.67	0.57
5	54.84	0.41	54.61	0.63
6	54.77	0.48	54.56	0.68
7	54.72	0.53	54.52	0.72
8	54.68	0.57	54.49	0.75
9	54.63	0.62	54.48	0.76
10	54.60	0.65	54.46	0.78

Processing data:



Discussion and Conclusion:

During this experiment calcium carbonate reacted with dilute hydrochloric acid to form calcium chloride, water and carbon dioxide as seen in the following equation:

calcium carbonate + hydrochloric acid → calcium chloride + water + carbon dioxide

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

When calcium carbonate reacted with an acid, effervescence was observed since carbon dioxide gas was released. The gas escaped through the glass wool and the flask, and its contents became lighter. The rate of this reaction was followed by measuring the loss of carbon dioxide at regular time intervals for 10 minutes.

SAMPLE EXPERIMENT REPORT – Continued

The aim of this experiment was to determine whether the surface area of a solid reagent affects the rate of a reaction. From the graph it can be concluded that the reaction of crushed calcium carbonate with acid was faster than the reaction of large calcium carbonate due to a steeper gradient. This experiment showed that by increasing the surface area of calcium carbonate the rate of reaction increased.

From the graph it can also be concluded that initially in the first two minutes both reactions were fast. Then from the 3rd and 9th minute both reactions slowed down. On reaching the 10th minute both reactions were almost over since the effervescence slowed down considerably.

This change in rate of reaction can be explained using the collision theory. Initially there are many successful collisions between the reactants resulting in a fast reaction. As the reagents are used up there are less successful collisions, hence the rate of reaction slows down. Moreover, by increasing the surface area of the solid reagent there are more exposed particles reacting with acid, leading to more successful collisions producing a faster reaction with crushed calcium carbonate. When larger pieces of calcium carbonate are used, the acid hits the outer layer of the calcium carbonate resulting in a slower reaction.

Evaluation:

The experiment produced a good set of readings to determine whether surface area affects the rate of reaction. However, the experiment could be improved by modifying the setup as shown in Figure 1. The calcium carbonate is placed in a small crucible in the conical flask filled with acid. The advantage of using this setup is that the reagents are quickly mixed, and the stopwatch can be started immediately. Secondly the experiment could run for a longer period of time such that the last three readings remain unchanged showing that the reaction is over.

This experiment could also be carried out by collecting the gas in a gas syringe and noting the volume of gas at regular time intervals as shown in Figure 2.

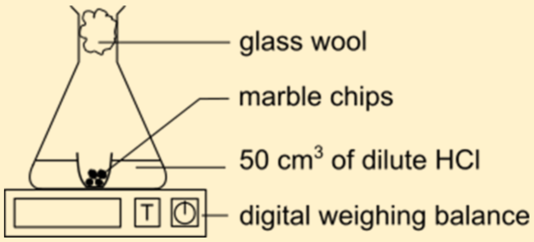


Figure 1

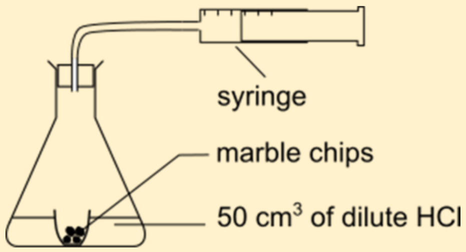


Figure 2

Besides changing the surface area, the rate of reaction of calcium carbonate and acid can be altered by changing the concentration of the acid, or by changing the temperature at which the reaction takes place.

1.4.1 Aim of the experiment

By the end of this section, you should be able to:

- State the aim/s of an experiment / investigation. (1.1d)

The aim of the experiment is stated at the beginning of the lab report. It should be concise and indicates what you are about to do in the experiment.

1.4.2 Drawing and labelling diagrams

By the end of this section, you should be able to:

- Label given diagrams (1.1k)
- Draw labelled diagrams from given apparatus (1.2k).
- Draw labelled diagrams of apparatus used during experiment/ investigations (1.3k)

Diagrams are an important aspect of the laboratory report since they give a pictorial representation of the setup used during the experiment / investigation. If the experiment involves several separate setups, then a diagram for each setup is expected. Section 1.3.1 shows how laboratory apparatus is to be drawn in the report. A science stencil is useful to draw neat diagrams of the appropriate size.

Follow the rules below when drawing and labelling diagrams:

1. Use a sharp pencil for the diagrams and labelling.

   Do not use a pen, coloured pencil, or blunt pencil.

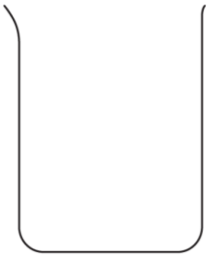
✓

✗

✗

2. Draw objects in two dimensions.

   Do not draw in 3D.



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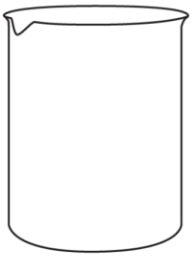

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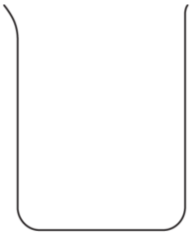
3. Draw clean single lines.

   Do not draw sketchy lines.



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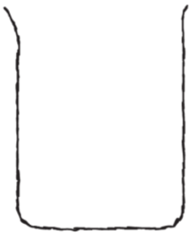

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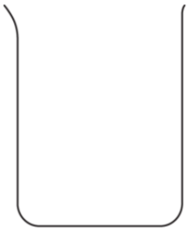
4. Do not close off openings of glassware.



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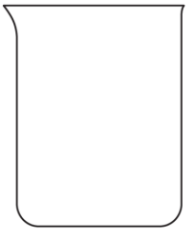

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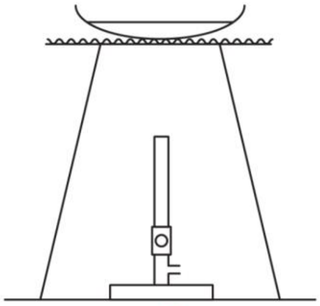
5. Use a ruler to draw straight lines.

   Do not draw lines freehand.



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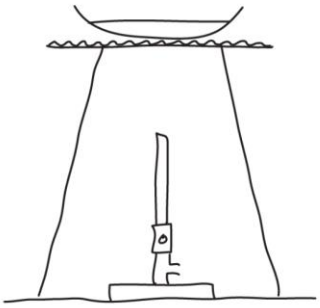

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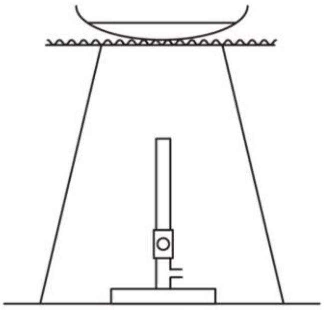
6. Do not shade or colour in.

   They might look nicer coloured in, but scientific diagrams need to be clear and simple.



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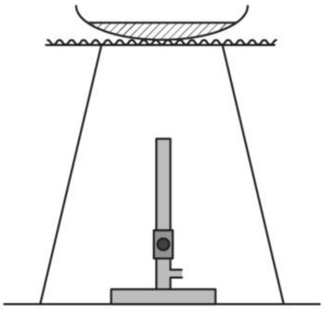

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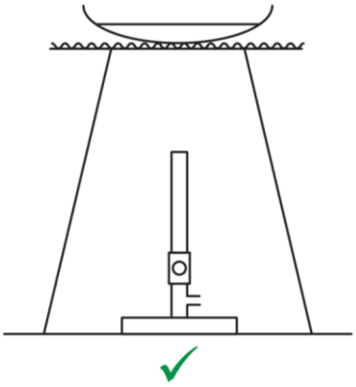
7. Do not float objects.

   If objects are in contact with each other, show it in the diagram.



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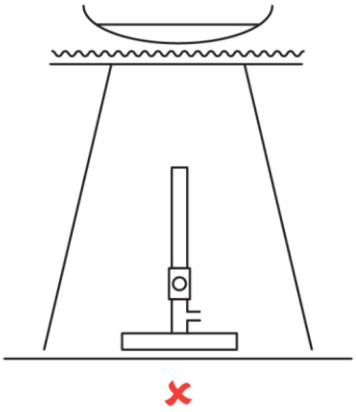

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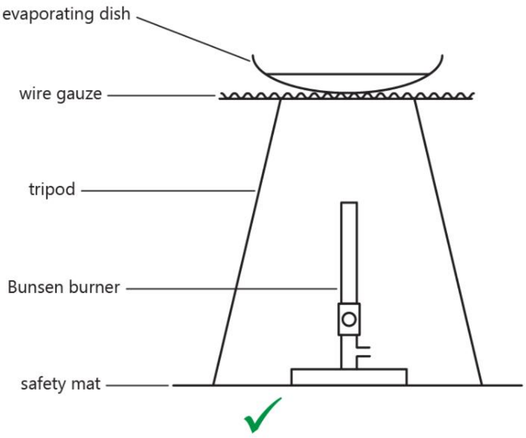
8. Label objects with straight lines.

   Do not use arrow heads. Avoid crossing over lines. Write labels in a horizontal manner.



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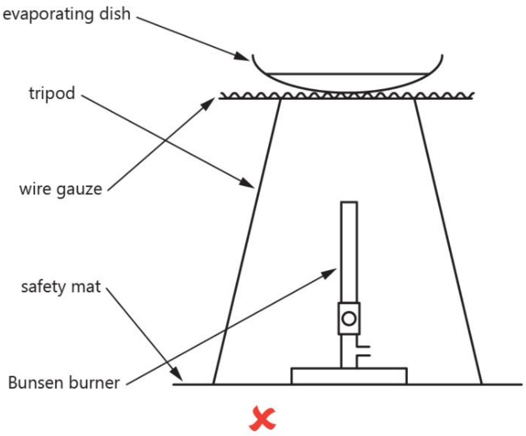


✗

(Rules and diagrams in this section adapted from:
https://www.goodscience.com.au/year-7-chemistry/scientific-diagrams/)

Lines with arrowheads should only be used to show movement such as the flow of a liquid. The size of a diagram must neither be too small nor too large. As a rule of thumb, the height of a diagram should not exceed 8 to 10 cm.

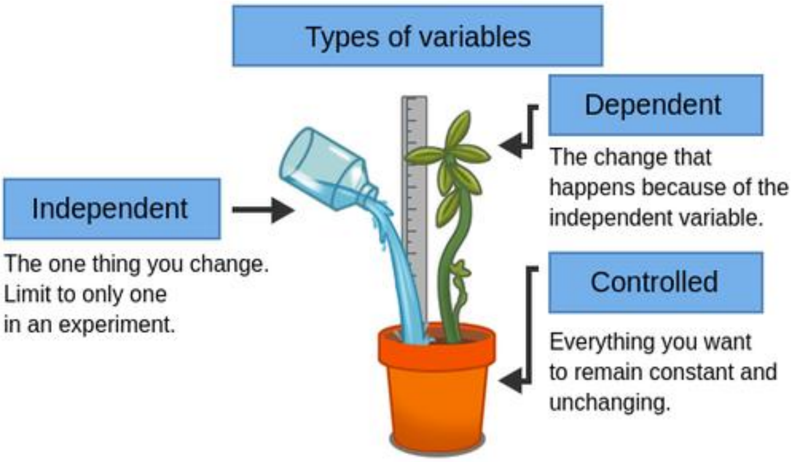
1.4.3 Identifying variables in an experiment

By the end of this section, you should be able to:

- Identify variables in an experiment (1.1c)
- Identify dependent, independent, and controlled variables. (1.3c)

Some experiments are carried out to determine a cause-and-effect relationship.

For example, does watering a plant (cause), improve growth (effect)?



Source: https://theory.labster.com/experimental_variables/

There are many factors that could be altered, and these are known as **variables**.

A variable is any factor, trait, or condition that can exist in differing amounts or types.

An experiment usually has 3 types of variables: **independent, dependent,** and **controlled**.

- The **independent variable** is the factor whose values are changed and controlled by the scientist to study the cause and effect this has on the values of the dependent variable. In other words, it is the quantity the scientist is changing. To ensure a fair test, a good experiment has only one independent variable.
- The **dependent variable** is the factor that is being investigated in an experiment and it changes in response to the value of the independent variable. In other words, it is the quantity being measured after the independent variable is changed.
- The **controlled variables** are conditions which must be controlled and kept constant throughout the experiment so that they do not interfere with the dependent variable. When the experiment is repeated, the controlled variables must be the same.

In the above experiment the variables will be as follows:

Independent variable (the variable that is changed)	Volume of water.
Dependent variable (the variable that is measured)	Length/growth of plant.
Controlled variables (the variables that are kept the same)	Type of soil, intensity of light, same kind of plant, ambient temperature, etc.

1.4.3.1 Fair and objective testing in science

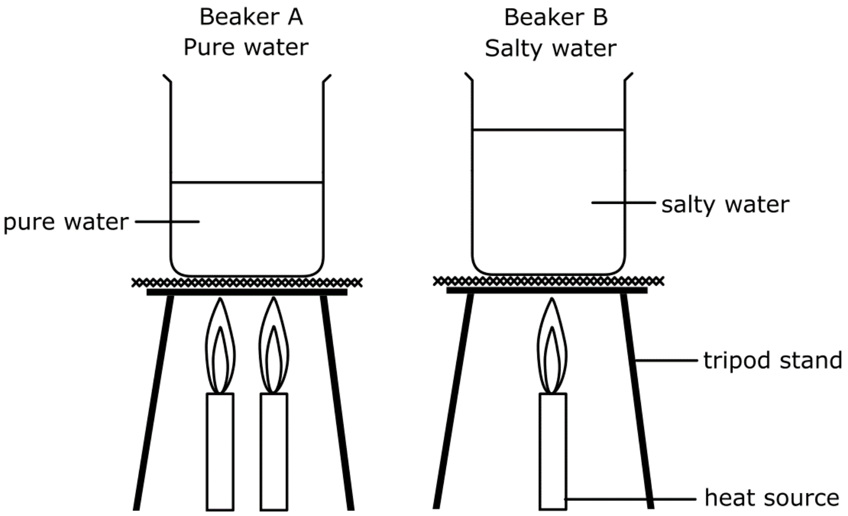
By the end of this section, you should be able to:

- State the importance of fair (objective) testing in science. (1.1b)
- Discuss the importance of fair (objective) testing in science. (1.2b)
- Evaluate an experiment in terms of its objectivity. (1.3b)

While performing science experiments, it is important to experiment fairly. This is done by changing only one factor (variable) at a time while all other conditions are kept constant (the same). The outcome of this variable on the experiment is observed and noted. Experimenting fairly helps increase the validity and reliability of the results obtained.

For example, consider the following experiment:

Experiment:
To determine whether pure and salty water would take the same time to boil.



One should note that there are several aspects shown in the setup that render this experiment useless since the experiments are not being done fairly. These aspects include:

- the volume of liquid used is not the same, and
- the number of candles used is not the same.

When evaluating and applying experimental results, one needs to be aware of how objective the scientist was when planning, conducting, and reporting the experiment. Being objective implies that throughout the scientific process:

- the scientist was free from all bias, interest, and manipulations,
- the scientist considered all possible viewpoints when conducting the experiment,
- the experimental conclusions are an honest reflection of truthful facts that actually resulted from the experiment.

1. Consider the following 2 experiments:

Exp 1: Measuring the volume of oxygen gas in 3 minutes from hydrogen peroxide solution after addition of **LARGE potato pieces** with a total mass of 5 g.

Exp 2: Measuring the volume of oxygen gas in 3 minutes from hydrogen peroxide solution after addition of **SMALL potato pieces** with a total mass of 5 g.

a. Identify the dependent and independent variables.

b. Identify the controlled variables to have a fair test.

2. Identify the variable being controlled in the following pairs of experiments.

Examples of Paired Experiments		Controlled Variable
1.	• Reaction between 2 g magnesium ribbon and 20 ml dilute hydrochloric acid. • Reaction between 4 g magnesium ribbon and 20 ml dilute hydrochloric acid.	
2.	• Reaction between 2 g magnesium ribbon and 20 ml dilute hydrochloric acid. • Reaction between 2 g magnesium powder and 20 ml dilute hydrochloric acid.	
3.	• Reaction between 2 g magnesium powder and 20 ml dilute hydrochloric acid. • Reaction between 2 g magnesium powder and 20 ml concentrated hydrochloric acid.	

3. Identify the independent and dependent variables in the following experiments.

Experiment	Independent Variable	Dependent Variable
Investigating the rate of production of oxygen gas when using hydrogen peroxide solutions of different concentrations.		
Investigating the rate of production of carbon dioxide gas when reacting dilute hydrochloric acid with powdered and lumpy calcium carbonate.		
Investigating the rise in temperature when reacting magnesium powder with dilute and concentrated hydrochloric acid.		
Investigating the mass of silver produced when exposing silver chloride to lights of various intensities.		
Investigating the effect on rhythm of the heart produced by varying potassium levels in the blood.		

1.4.4 Precautions and safety considerations

By the end of this section, you should be able to:

- Identify precautions for a given experiment/ investigation (1.2f).
- Justify precautions for a given experiment/ investigation (1.3f).

Before carrying out an experiment, the procedure must be read and understood. If there are steps that you do not understand, make sure to ask your teacher to clarify your misunderstanding. After reading the procedure, outline the precautions that need to be taken to improve the **accuracy** and **reliability** of the experiment. You also need to explain why each precaution needs to be taken. In a report, precautions are written in the third person past tense.

Prior to an experiment, consult SDS documents for chemicals so that health and safety considerations are given due importance. Safety precautions related to glassware, electricity and fire should also be considered. Remember, that no matter how excited you may be to start your experiment; safety comes first for your sake and that of others.

1.4.5 Recording observations and measurements

By the end of this section, you should be able to:

- Complete observations/ measurements in a given table for an experiment. (1.1i)
- Record all observations/ measurements in a given table for an experiment. (1.2i)
- Record observations/ measurements appropriately for an experiment/investigation. (1.3i, 1.2j)
- Determine which observations/ measurements are to be measured for an investigation. (1.3j)

A written procedure will surely provide instances where you will need to measure quantities. Usually, a table is used to write and present the information that is measured during an experiment in an orderly manner.

The first row would include the headings of the quantities that are being measured as well as calculated quantities where applicable. These headings also include the unit of measurement in brackets to avoid writing the unit of measurement with every measurement taken. It is important to write readings with the same number of decimal places. The independent variable usually goes into the first column and the dependent variable goes into the second column. At times you may decide to repeat the experiment to obtain more reliable data. The table of results would look like this.

Independent variable (unit)	**Dependent variable (unit)**			
	1st trail	2nd trial	3rd trial	Mean (average)

Apart from the measurement of quantities one needs to keep a watchful eye to what is happening during an experiment such as changes in state, colour, texture, odours, liberation of a gas (effervescence), emission or absorption of heat, etc. Any of these observable changes must be noted and included in your laboratory report.

1.4.6 Plotting graphs

By the end of this section, you should be able to:

- Plot a single series of data on given axes (1.1m).
- Plot a single series of data (1.2m).
- Plot multiple series of data on the same axes (1.3m). Not examinable in Year 9

During an experiment that involves measuring quantities, the information obtained is usually presented in a table. To better understand the relationship between variables, a graph would need to be plotted to obtain a visual perspective of the data obtained during the experiment.

Plotting a graph involves some skills that need to be mastered as listed below:

1. Draw the x and y axes on your graph paper.
2. Label each axis with the variable it represents, including the units.
3. Put the independent variable on the x-axis and the dependent variable on the y-axis.
4. Identify the range of data you must include for each variable such that the graph would occupy at least $^{2}/_{3}$ of the graph area.
5. Decide how many units every line on the graph represents for each of your variables.
6. Plot your data by using a plus (**+**) notation. Do not use an "**×**" to plot points.
7. Connect the data points with a best straight line or smooth curve depending on the general shape of the graph plotted.
8. Steps 5 and 6 should be repeated when plotting several series of data on the same graph. Use different symbols for the data points of each graph. Label each graph accordingly.
9. Write the graph's title at the top of the graph.

(List adapted from: https://www.wikihow.com/Make-a-Line-Graph)

1.4.7 Reading and interpreting graphs

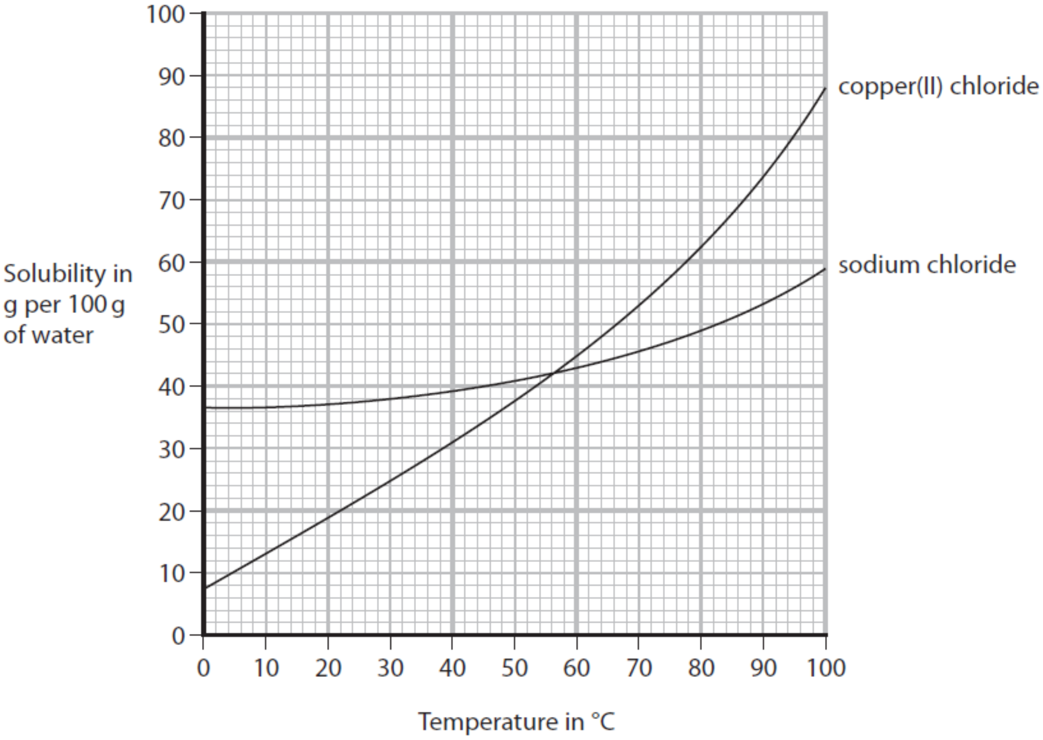
By the end of this section, you should be able to:

- Read values from simple graphical representations (1.1l).
- Interpret graphical representations containing single series of data (1.2l).
- Interpret multiple series of data plotted on the same axis (1.3l).

Reading and interpreting graphs are two different skills that one needs to develop while studying Chemistry. Consider the solubility curves shown on the next page which show the solubilities of copper(II) chloride and sodium chloride at different temperatures.

Reading a graph means that you should be able to give the values for the x and y axes of a given point on a graph. For example, at 44 °C, 40 g of sodium chloride dissolve in 100 g of water whereas at the same temperature, 34 g of copper(II) chloride dissolve in 100 g of water. At 56 °C both salts have the same solubility that is 42 g in 100 g of water.

Interpreting a graph means that when you look at the plotted graph, you should be able to tell how solubility varies with temperature. For example, the solubility of sodium chloride increases steadily with increasing temperature, but the solubility of copper(II) chloride rapidly increases with higher temperature.



1.4.8 Sketching graphs

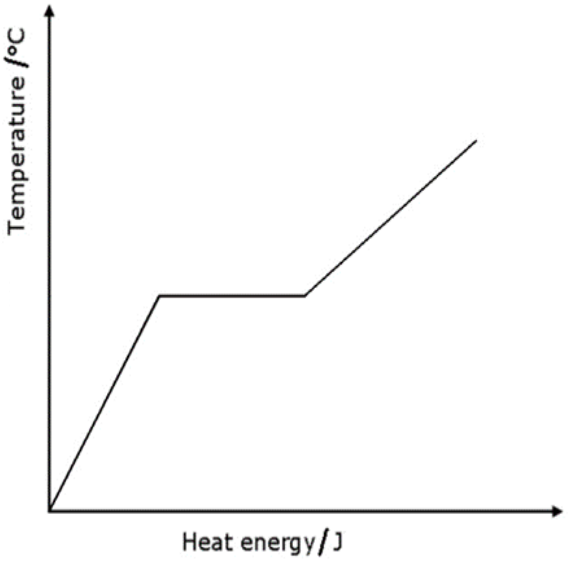
By the end of this section, you should be able to:

- Interpret situations by sketching a graph (1.2n)
- Interpret situations by sketching graphs in relation to existing plotted graphs (1.3n) Not examinable in Year 9

Sketching graphs means that you should be able to draw a graph without having specific coordinates. However, since you would know in advance what the graph should look like, you would be able to sketch the graph.

The graph shown here shows a sketch of ice being heated until it melts and reaches room temperature. Note how the sketch shows the rise in temperature of the ice until it reaches the melting point, which is shown by the horizontal line and then the rise in temperature of the water is shown by the second slope.

Sketches may also be drawn on existing plotted graphs to show how the plotted graph would change if a variable were changed in a particular way.



1.4.9 Discussing results and drawing conclusions

By the end of this section, you should be able to:

- Draw conclusions from an experiment (1.2o)
- Draw conclusions from an experiment by relating it to scientific knowledge, laws, and theory (1.3o)

After performing an experiment and processing data, you should be able to conclude your experiment by:

- Writing a summary of the findings of the experiments and relating them clearly to the aim of the experiment.
- Writing a discussion of any patterns or trends in the data.
- Stating any relationships discovered or confirmed between variables being tested in the experiment.
- Comparing numerical results with known values from data books and suggesting any reasons for any differences.
- Making a complete analysis or interpretation of observations noted in the experiment (including balanced chemical equations and ionic equations where appropriate).
- Drawing a conclusion based on experimental evidence and relating it to scientific knowledge, laws, and theory.

1.4.10 Evaluating a laboratory report

By the end of this section, you should be able to:

- Identify sources of error to suggest improvements (1.2p)
- Evaluate an experimental procedure and results by suggesting improvements (1.3p).

Your report should end with an evaluation where you should be able to:

- Identify and comment on any sources of error in the experiment.
- Discuss any anomalous results and try to explain why they happened.
- Discuss any difficulties encountered in carrying out the experiment and any precautions taken to achieve accuracy.
- Suggest way/s of improving the experimental setup and/or results.
- Suggest any other experiments which can be done to support the conclusions.

1.5 Investigative experiments

By the end of this section, you should be able to:

- Predict what might happen in an experiment/ investigation. (1.2g)
- Justify prediction/s made for an experiment/ investigation. (1.3g)
- Plan an experiment to solve a given problem with supervision. (1.2q)
- Plan an experiment to solve a given problem without direct supervision. (1.3q)
- Carry out an experiment to solve a given problem with supervision. (1.2r)
- Carry out an experiment to solve a given problem without direct supervision. (1.3r)

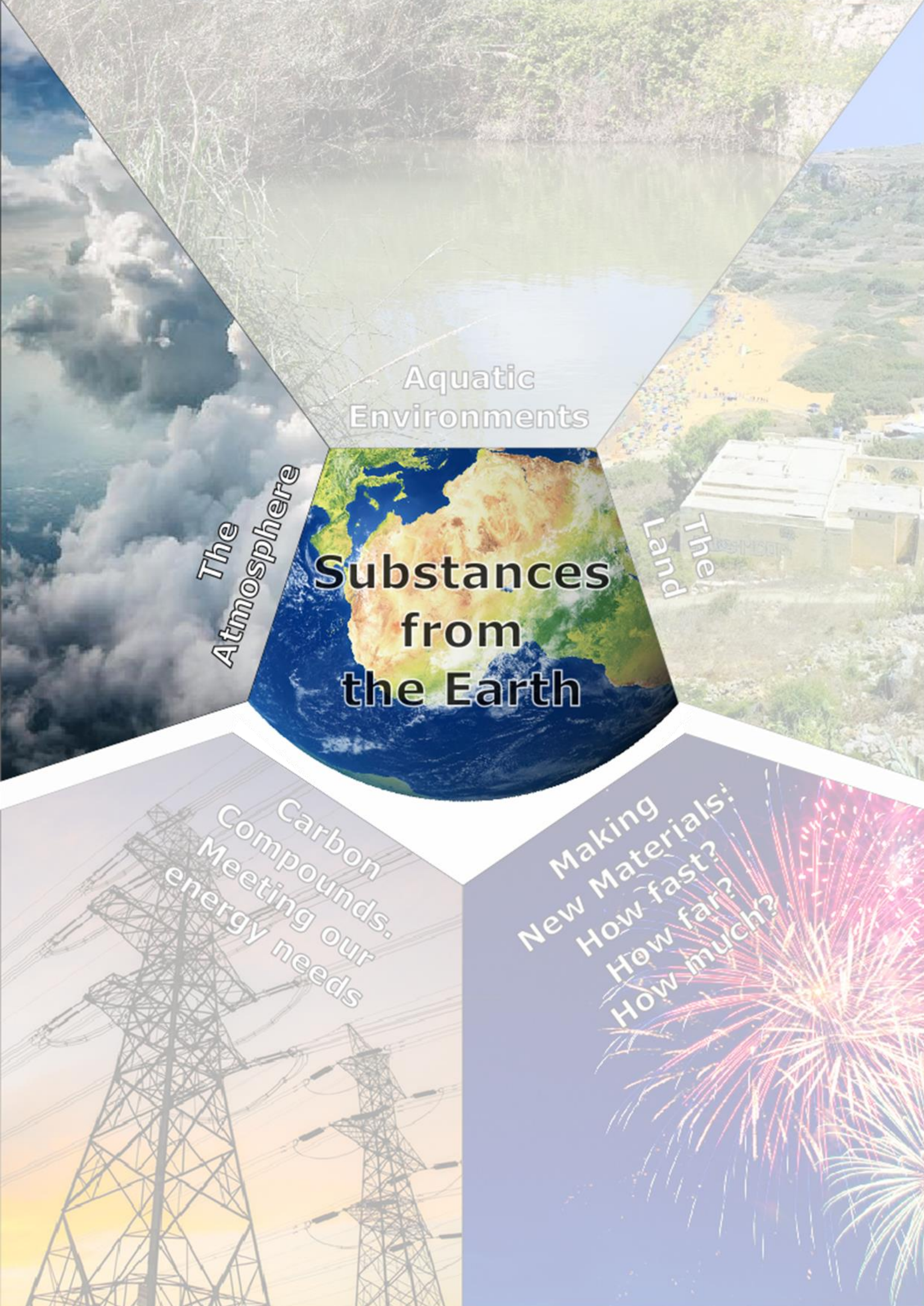
An investigative experiment differs from a normal experiment in that you would be given a problem that needs to be solved. You will need to think of a solution and design an experiment to test your solution. Your plan (investigation outline) needs to be presented to your teacher so that s/he could assess your plan for health and safety issues only. When the plan is approved by your teacher, you will then execute your plan and perform the experiment that you proposed. Sometimes, when your plan may not be adequate, you will need to adjust it accordingly while you are performing the experiment. These changes need to be noted in the experimental procedure as well as in the evaluation section.

The following template, as published by SEC, show the sections and respective notes that should be included in an investigation report.

Write up of an investigation	
Investigation outline	This section should contain an outline of the procedure that will be devised in the investigation together with scientific theory required to understand the investigation. The plan should be concise and written in the future tense. This section should include: • The title. • A short statement of the problem to be investigated. • The aim of the investigation. • A brief description of the scientific procedure. • A list of materials and apparatus. • Any pre-experiment work. • A variables grid may be presented to highlight all the variables in the investigation (where applicable). Variables should be identified as independent and dependent variables. Other significant/relevant variables should be noted including the way they are controlled for results to be more reliable. • Any background theory/research where applicable is given. • The hypothesis section (where applicable) should give an outline of what may happen and why. (Note: Students are to be made aware that no marks will be lost if the hypothesis is disproved.)
Precautions and safety considerations	This section should include: • Any precautions taken to achieve a more accurate result and improve the outcome of the investigation. • Safety considerations associated with the preparation and implementation of the investigation to prevent any accidents.

Procedure Followed	This section should include: A detailed account of the procedure followed. All the steps involved to perform the experiment including any modifications made to the plan and any additional materials and apparatus used should be stated. The method should include measurements used, diagrams, and photos, where applicable. Note: • Results should not be included in this section. • Third person past tense should be used. • Any concentrations, measurements, amounts, times, and temperatures should be quantified. • The procedure should be written in such a way that an independent person could repeat the experiment without referring to the person writing the report.
Results and observations	This section should include: • All observations and/or measurements should be presented in an organised form. • Any calculated data should be presented showing all steps. • Graphical representations should be used to display data when possible. Note: • Tables may be the best way of presenting data. • Tables should have headings and units. • An adequate number of readings should be taken especially if a graph has to be plotted. • Results should not be interpreted in this section. • Third person past tense should be used to describe any observations.
Discussion and Conclusion	This section should include: A brief summary of the aim of the investigation. • A summary of the most important findings including trends and patterns emerging from analysis of the results. • An explanation why calculations were used if any, and their link to the investigation. • A very brief description stating whether the investigation has supported/falsified the hypothesis. • A description and an explanation of how the results relate to the expectations based on laws, theories, relationships, patterns, and models studied. • This section should be concluded by a closure of all findings.
Evaluation and references	This section should include: • A list of procedural/sources of errors that may have affected the result. • A list of improvements and any other experiments which can be done to support the conclusions. All sources cited in the text should be listed in full. A basic format should be used when listing the sources.

Source: https://www.um.edu.mt/__data/assets/pdf_file/0007/509191/SEC062025Resources.pdf



Substances from the Earth

The Atmosphere

Aquatic Environments

The Land

Carbon Compounds. Meeting our energy needs

Making New Materials: How fast? How far? How much?

2 Gases found in air and their lab preparation (LO 2)

2.1 Gases in the Earth’s atmosphere

By the end of this section, you should be able to:

- Identify the gases that make up the air naturally and those that may be added by humans. (E.g., nitrogen, oxygen, carbon dioxide, water vapour, noble gases, carbon monoxide, sulfur dioxide, nitrogen oxides and ozone.) (2.1a)
- State the approximate percentage of nitrogen, oxygen, carbon dioxide and noble gases in dry, unpolluted air. (2.2a)

2.1.1 Composition of air

The composition of air varies from place to place because air is a mixture of gases.

A **mixture** is made up of two or more substances mixed together, that can be separated by physical means.

The composition by volume of a typical sample of air is as follows:

	GAS	PERCENTAGE VOLUME
Clean dry air	Nitrogen	78 %
	Oxygen	21 %
	Carbon dioxide	0.03 %
	Noble gases * (Argon, neon, helium, krypton, xenon)	Approximately 1 %
Humidity	Water vapour	Variable
Air pollution	Impurities or pollutants (e.g., sulfur dioxide (SO_2), oxides of nitrogen (NO and NO_2), carbon monoxide (CO), ozone (O_3), etc.)	Variable

* *The most abundant of the noble gases is argon which amounts to 0.9%.*

The composition of air is mainly affected by the following factors:

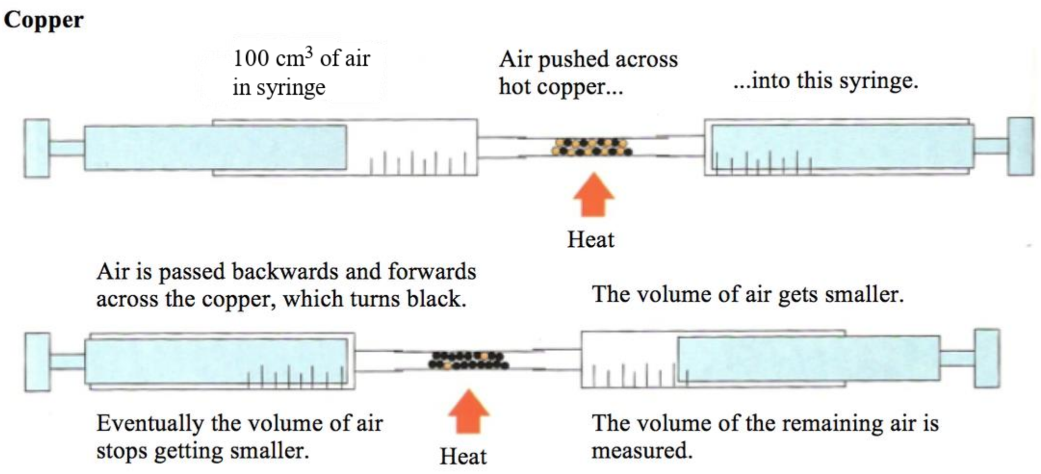
- **respiration** by all living things, in which oxygen is taken in and carbon dioxide is given out to the atmosphere.
- **photosynthesis** (the way plants build up their food using sunlight) in which carbon dioxide is taken in and oxygen is given out to the atmosphere.
- **decay** (when living things die and are consumed by microorganisms) in which carbon dioxide is released.
- **volcanic activity** in which water vapour, carbon dioxide and other impurities (sulfur dioxide and hydrogen sulfide) are released when a volcano erupts.
- **human activity** involving burning of fossil fuels, in which carbon dioxide and water vapour are produced as well as other pollutant gases such as sulfur dioxide and oxides of nitrogen.

2.1.2 Measuring the percentage of oxygen in air

By the end of this section, you should be able to:

- Determine experimentally the percentage oxygen in air. (2.3a)

During combustion reactions, substances react with oxygen to form oxides. The following experiment can be used to measure the percentage of oxygen found in air.



1. Place an excess of red-brown copper turnings in a hard glass tube connected to two gas syringes one on each side.
2. Check that the apparatus is airtight by pushing the air from one syringe to the other for several times.
3. Pass 100 cm^3 of air over the heated copper turnings by pushing the plunger to and fro until the volume in the gas syringes does not change.
4. Leave the syringes to cool to room temperature before taking any readings since air expands on heating and a higher value would be obtained if read at a high temperature.

Volume of air at the start (cm^3)	**Volume of air at the end (cm^3)**	**Volume of oxygen used (cm^3)**
100	79	21

Therefore, the percentage of oxygen gas in the air is:

$$\frac{21}{100} \times 100 = 21\%$$

In this experiment the reactive part of air, i.e. oxygen, is reacting with copper to form black copper(II) oxide, as shown in the equation below.

$$\text{copper} + \text{oxygen} \rightarrow \text{copper(II) oxide}$$

$$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$$

Measuring the percentage of oxygen in air.

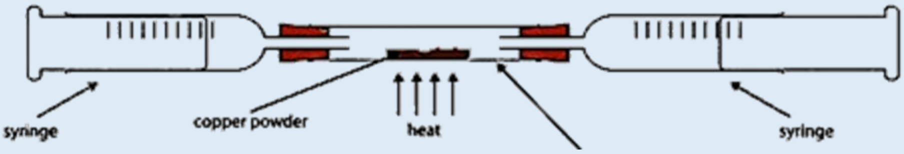
1. Fill in the table below with the composition by volume of a typical sample of air.

Gas	**Percentage volume**
Nitrogen	
Oxygen	
Carbon dioxide	
Noble gases (argon, neon, helium, krypton, xenon)	
Water vapour	
Impurities or pollutants (e.g., NO_x, SO_2, etc)	

2. Indicate whether the following statements are true or false. Correct the false statements.

a. The air is mainly a mixture of two gases – hydrogen and oxygen. T / F
b. Component gases of clean air are carbon dioxide, water vapour, noble gases, and sulfur dioxide. T / F
c. The composition of air varies from place to place. T / F
d. The most abundant of the noble gases in air is neon. T / F
e. Burning fossil fuels contribute to the amount of carbon dioxide, sulfur dioxide and oxides of nitrogen in the atmosphere. T / F
f. Oxygen takes up one fourth of air. T / F

3. The apparatus shown below is used to estimate the proportion of oxygen in the atmosphere. Copper is present in excess. A volume of dry air (90 cm^3) is passed backwards and forwards over heated copper until no further change in volume takes place. The apparatus is then allowed to cool down to room temperature and the final volume reading is taken. Some typical results are shown below.



Volume of gas before passing over heated copper = 90.0 cm^3
Volume of gas after passing over heated copper = 70.7 cm^3

During the experiment the copper slowly turned black.

a. Explain why it is important that copper is present in excess quantity.
b. Explain why the apparatus is allowed to cool back to room temperature before the final volume reading is taken.
c. Using the information given above, calculate the reduction in volume in the gas syringes at the end of the experiment.
d. Explain why there is a change in volume.
e. Calculate the percentage reduction in volume.
f. Would you expect the copper to have increased or decreased in mass during the experiment? Explain your answer.
g. Write a word equation for the reaction taking place.
h. Give the name of the main residual gas at the end of the experiment.

2.1.3 Properties and uses of the gases in dry air

By the end of this section, you should be able to:

- Describe the properties of nitrogen, oxygen, carbon dioxide and noble gases. (2.1b)
- Relate the properties of nitrogen, oxygen, carbon dioxide and noble gases to their uses. (2.2b)

2.1.3.1 Nitrogen – properties and uses

Nitrogen is the most abundant gas in air. The following include some of nitrogen's properties.

- It is a colourless and odourless gas.
- It is a neutral gas (pH 7).
- It is slightly soluble in water.
- It has a similar density to air. (N_2 = 1.25 g/l; air = 1.29 g/l at STP)
- It is very unreactive compared with oxygen.
- Nitrogen also combines with oxygen at high temperatures to form oxides: nitrogen monoxide (NO) and nitrogen dioxide (NO_2). This reaction occurs naturally in the air due to lightning during thunderstorms, as well as inside running car engines and power station furnaces.

Nitrogen is used in the following situations.

- Liquid nitrogen is very cold. It boils at -196 °C. It is used to quick-freeze food in food factories. Liquid nitrogen is often used as a refrigerant. It is used for storing sperm, ova, and other cells for medical research and reproductive technologies. It is also used in hospitals to store tissue samples.
- Due to its lack of reactivity, nitrogen gas is flushed through food packing to remove oxygen and keep the food fresh for a longer period of time.
- To produce ammonia when combined with hydrogen under special conditions.

Source: https://www.thoughtco.com/liquid-nitrogen-facts-608504

Further reading on facts and properties of liquid nitrogen.

Daniel Rutherford and the 'property' of nitrogen

The Scottish chemist, Daniel Rutherford, isolated nitrogen gas from air in 1772. He was a student of Joseph Black, who discovered the presence of CO_2 in air. Rutherford carried out experiments that removed the oxygen and carbon dioxide present in air. He noticed that living things could not survive in the remaining gas. This gas was later named nitrogen.

2.1.3.2 Oxygen – properties and uses

Oxygen has the following properties.

- It is a colourless and odourless gas.
- It is a neutral gas.
- It is slightly soluble in water; in fact it supports aquatic life.
- It is slightly denser than air. (O_2 = 1.43 g/l; air = 1.29 g/l at STP)
- It is a reactive element, combining vigorously with many metals (e.g., sodium, magnesium, copper) and non-metals (e.g., carbon, sulfur). It supports combustion.

Oxygen is used in the following situations.

- Required during combustion reactions.
- As an aid to breathing where the natural supply of oxygen is insufficient. For example, in high-altitude flying or climbing, and when anaesthetics are administered to a patient.
- In the oxyacetylene (oxygen-ethyne) torch, which can be used for welding and for cutting very thick steel plates. The temperature of the flame reaches about 3200 °C.
- For making steel.
- In sewage treatment. If it were not for this treatment, many rivers and lakes would be spoiled by sewage.

Further reading on facts about oxygen.

Properties of oxygen - John Mayow

John Mayow, born in England in 1640, was one of the first scientists who identified the nature of air. Through his work, he established that air was not composed of a single element, but of different substances. Mayow showed that combustion was supported by only a part of air and this same portion was responsible for respiration. Mayow, wrote on the function of oxygen in respiration almost a hundred years before oxygen was discovered.

2.1.3.3 Carbon dioxide – properties and uses

Carbon dioxide has the following properties.

- It is a colourless, odourless gas.
- It is an acidic gas because it is slightly soluble in water, forming carbonic acid, H_2CO_3.
- It is denser than air (CO_2 = 1.96 g/l; air = 1.29 g/l at STP).
- Things will not burn in it (it does not support combustion).
- It is not poisonous but does not support respiration.

Carbon dioxide is used in the following situations.

- It is taken up by plants during the process of photosynthesis.
- It is used in the production of sparkling mineral water, fizzy drink, and beers.
- It is used in fire extinguishers.
- It is used as a coolant.

Carbon dioxide in the form of dry ice is used in the winemaking industry. Dry ice is solid carbon dioxide, and it is used to cool down grapes to prevent fermentation. CO_2 sublimes at -78 °C.

Sublimation is a process where a substance changes from a solid to a gas without going through the liquid phase.

2.1.3.4 Noble gases – properties and uses

The noble gases are helium, neon, argon, krypton, xenon, and radon. Their properties include:

- they are colourless and odourless gases.
- they are monoatomic, as they exist as single atoms. E.g., helium – He; neon – Ne.
- they are very unreactive because their outer electron shell is full.
- become denser on going down the group (helium and neon are less dense than air but argon, krypton and xenon are denser than air).
- their boiling points increase on going down the group.

The following noble gases are used in the following situations.

- **Helium** is used in weather balloons and in mixtures with oxygen for deep sea diving tanks.
- **Argon** was used to fill filament light bulbs because the hot filament won't react with it.
- **Neon** is used extensively in advertising signs, in lasers, and produces a bright red colour.
- **Krypton** and **xenon** are used in lighthouse lamps and lasers.

Discovery of the noble gases

Sir William Ramsey was a Scottish chemist, born in 1852, who discovered the first noble gas, for which he won a Noble prize. Ramsey was able to isolate a very small, yet heavy component of air, and reported it to be 'astonishingly' unreactive. He named this gas, Argon, which in Greek means idle or lazy. With the help of other scientists, he also discovered neon, krypton, and xenon, and later isolated helium and radon.

Sir William Ramsey is the scientist responsible of adding the group of Noble gases to the Periodic Table.

Infographic about Gases for Scuba Diving

2.2 Elements, compounds, and mixtures

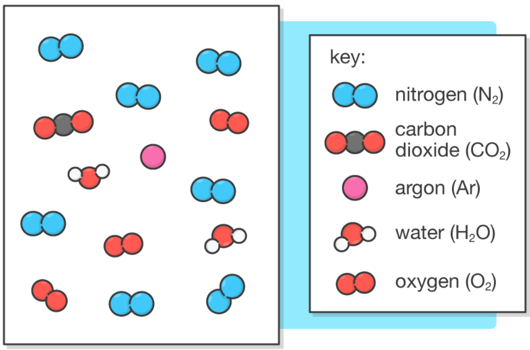
By the end of this section, you should be able to:

- Distinguish between elements, compounds, and mixtures. *(E.g. using gases in air.)* (2.1c)
- Explain the difference between elements, compounds, and mixtures. (2.2c)

2.2.1 Mixtures

Air is a mixture of gases, some are **elements** (e.g., oxygen, nitrogen, helium) while others are **compounds** (e.g., water vapour and carbon dioxide). A mixture is made up of two or more substances that are mixed together that can be separated by physical means.

Compostion of Air



Sea water, milk and soil are examples of mixtures.

- The composition of different substances in a mixture can vary.
- No chemical change takes place when a mixture is formed.
- Mixtures can be easily separated by physical means.
- The properties of the components are those of the individual substances.

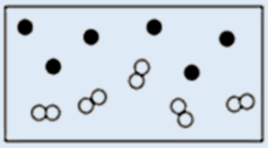
Solutions are also mixtures for e.g., a sugar solution is made of dissolved sugar in water.

1. Underline which of the following substances is a mixture.

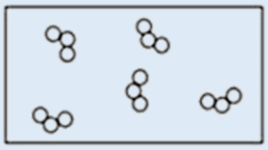
nitrogen milk table salt seawater steel

soil air copper(II) sulfate solution steam orange juice

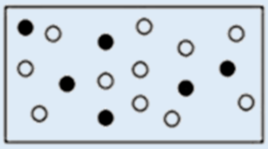
2. Identify which diagram represent the particles in a mixture.



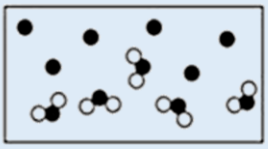
A



B



C

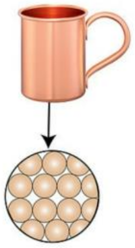


D

2.2.2 Elements

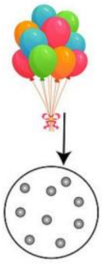
An **element** is a pure substance which cannot be further divided into simpler substances by chemical methods.

Elements contain only **one type** of atom. Here are three examples of elements:



Copper is an element made up of copper atoms only.

Carbon is an element made up of carbon atoms only.

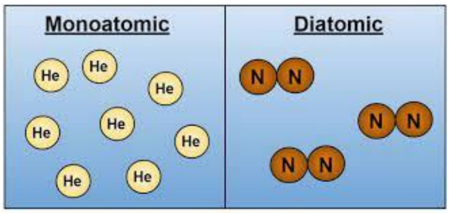


Helium is an element made up of helium atoms only.

There are many other elements that we make use of everyday like copper in wires, carbon (in the form of graphite) in pencils and aluminium in window frames. To date, there are 118 chemical elements; 98 elements occur naturally and range from reactive gases like fluorine and chlorine to unreactive elements like gold and platinum. Twenty elements have been artificially made by scientists like plutonium, curium and oganesson. The latter element was discovered in 2002! These elements are all found in the **Periodic Table of the Elements**.

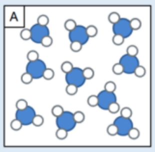
A **physical property** is any characteristic of a substance that can be measured, for example density, melting and boiling point. The elements have different properties and can be classified according to their properties. For instance, elements can be classified as **metals** or **non-metals**. Some are known as **metalloids** which means they have properties of metals and non-metals.

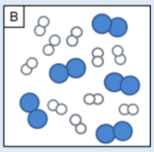
Elements can be monoatomic or diatomic.

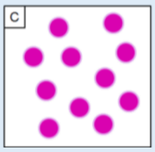


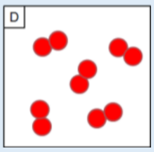
- Noble gases are **monoatomic** (made up of 1 atom).
- Seven elements are **diatomic** (made of 2 atoms) which include hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, and iodine.

1. What is an element?
2. Give two examples of elements found in air.
3. Tick which diagrams represent the particles in an element. Explain your answer.



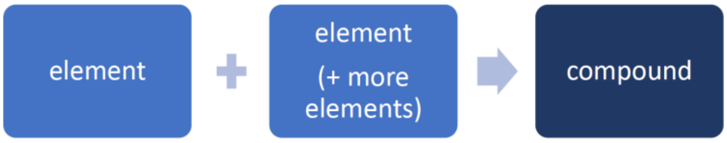






2.2.3 Compounds

A **compound** is a pure substance which is made of two or more elements that are chemically combined together.



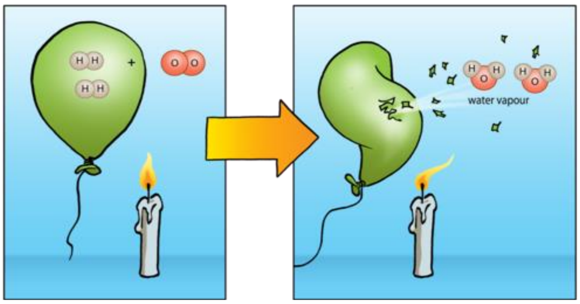
The elements in a **compound** are not just mixed; they are joined together by **chemical bonds**.

Examples of compounds found in air include water vapour and carbon dioxide. Many other compounds can be found around us in the ground such as calcium carbonate (found in limestone) and aquatic environments such as sodium chloride (common salt).

Look at the following example where hydrogen and oxygen are reacting to produce water. This reaction is initiated in the presence of a flame.

hydrogen + oxygen → water

This is a **chemical change** since a new substance is formed from a chemical reaction.



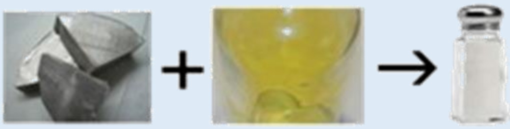
Source: https://www.quora.com/What-does-hydrogen-give-off-when-burned

Compounds have different chemical and physical properties from the elements of which they are made of. The table below shows that water has different properties from hydrogen and oxygen.

Elements		Compound
hydrogen	**oxygen**	**water**
• Hydrogen is the lightest gas in the universe. • It is a very flammable gas.	• Oxygen is slightly denser than air. • It is required during combustion reactions.	• Water vapour is the compound formed as hydrogen reacts with oxygen. The water vapour produced condenses to form liquid water. • Water is used to put out fire.

Sodium, a very reactive metal, reacts with chlorine which is a greenish yellow gas to form white crystals of sodium chloride, commonly known as table salt.

sodium + chlorine → sodium chloride



a. Research one property of: sodium, chlorine, and sodium chloride.

b. Explain how the properties of the compound differ from that of the constituent elements.

2.3 The Periodic Table of the Elements

By the end of this section, you should be able to:

- Use a Periodic Table to find information about elements. *(Including an online Periodic Table.)* (2.1d)

In the early 19th century various new elements were being discovered and people were studying them for any similarities or patterns. Various people suggested different ways of how to group the elements. Dimitry Mendeleev arranged the elements in order of increasing atomic weights in a table nowadays known as the Periodic Table of the Elements.

Elements are represented by chemical symbols that have been established through international agreement. Some elements were discovered by ancient civilisations like silver, iron, gold, lead, and mercury, others have been discovered along the years.

Dimitri Mendeleev

A Russian chemist, Dimitri Mendeleev (born in 1834), devised a table of elements where elements were placed next to each other in order of their atomic weight to form rows. Those elements that had similar chemical properties were placed under each other to form columns. In Mendeleev's time it was still unknown that atoms are made of other particles (that is protons, neutrons, and electrons).

In this Periodic Table, Mendeleev left empty spaces for elements that were still undiscovered. In time, these gaps were filled with newly found elements, whose properties were previously predicted due to their position in the Periodic Table. The metallic radioactive element Md, of atomic number 101, was named Mendelevium after this scientist.

PERIODIC TABLE OF THE ELEMENTS

1	2											3	4	5	6	7	0
							1 **H** Hydrogen 1										4 **He** Helium 2
7 **Li** Lithium 3	9 **Be** Beryllium 4											11 **B** Boron 5	12 **C** Carbon 6	14 **N** Nitrogen 7	16 **O** Oxygen 8	19 **F** Fluorine 9	20 **Ne** Neon 10
23 **Na** Sodium 11	24 **Mg** Magnesium 12											27 **Al** Aluminium 13	28 **Si** Silicon 14	31 **P** Phosphorus 15	32 **S** Sulfur 16	35.5 **Cl** Chlorine 17	40 **Ar** Argon 18
39 **K** Potassium 19	40 **Ca** Calcium 20	45 **Sc** Scandium 21	48 **Ti** Titanium 22	51 **V** Vanadium 23	52 **Cr** Chromium 24	55 **Mn** Manganese 25	56 **Fe** Iron 26	59 **Co** Cobalt 27	59 **Ni** Nickel 28	63.5 **Cu** Copper 29	65 **Zn** Zinc 30	70 **Ga** Gallium 31	73 **Ge** Germanium 32	75 **As** Arsenic 33	79 **Se** Selenium 34	80 **Br** Bromine 35	84 **Kr** Krypton 36
85 **Rb** Rubidium 37	88 **Sr** Strontium 38	89 **Y** Yttrium 39	91 **Zr** Zirconium 40	93 **Nb** Niobium 41	96 **Mo** Molybdenum 42	99 **Tc** Technetium 43	101 **Ru** Ruthenium 44	103 **Rh** Rhodium 45	106 **Pd** Palladium 46	108 **Ag** Silver 47	112 **Cd** Cadmium 48	115 **In** Indium 49	119 **Sn** Tin 50	122 **Sb** Antimony 51	128 **Te** Tellurium 52	127 **I** Iodine 53	131 **Xe** Xenon 54
133 **Cs** Caesium 55	137 **Ba** Barium 56	139 **La** Lanthanum 57	178 **Hf** Hafnium 72	181 **Ta** Tantalum 73	184 **W** Tungsten 74	186 **Re** Rhenium 75	190 **Os** Osmium 76	192 **Ir** Iridium 77	195 **Pt** Platinum 78	197 **Au** Gold 79	201 **Hg** Mercury 80	204 **Tl** Thallium 81	207 **Pb** Lead 82	209 **Bi** Bismuth 83	210 **Po** Polonium 84	210 **At** Astatine 85	222 **Rn** Radon 86

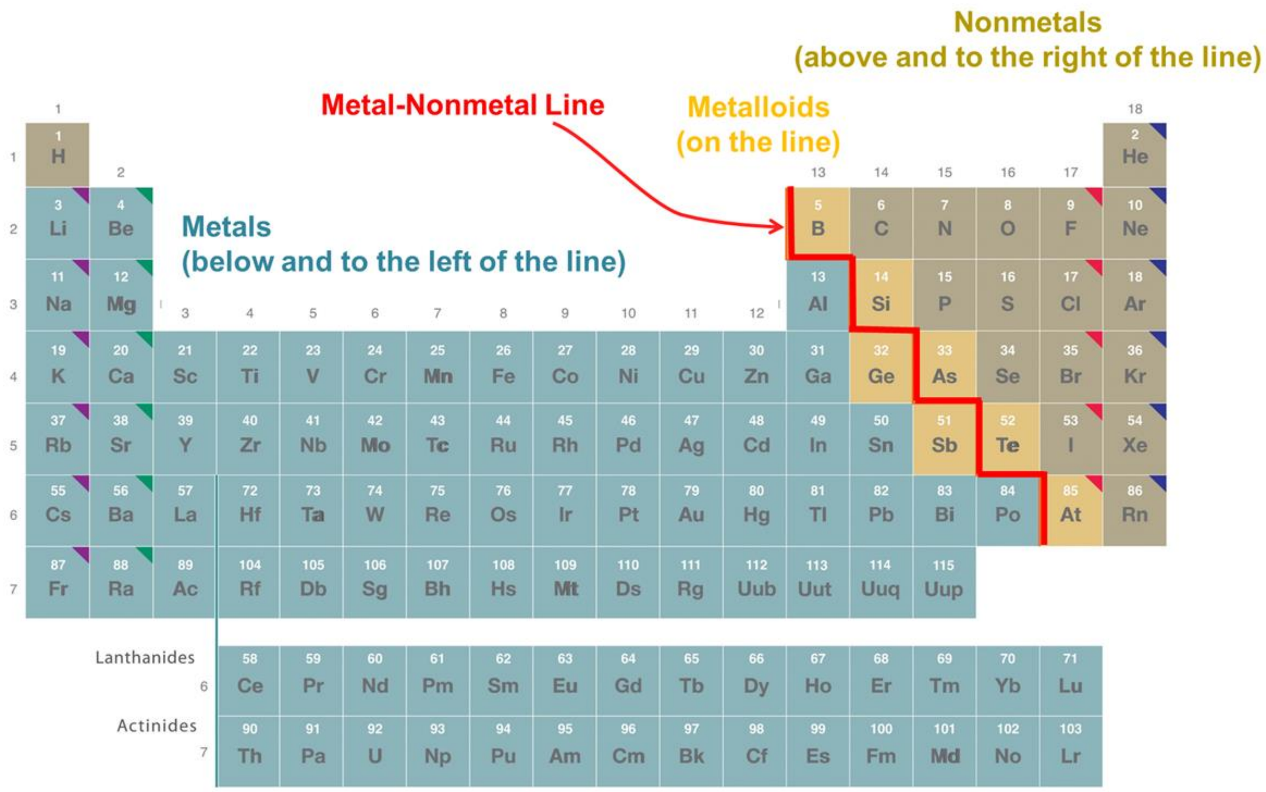
Key:	a **X** Y b	*relative atomic mass* ***SYMBOL*** ***Name*** *atomic number*

Nowadays, the Periodic Table shows the arrangement of the elements in order of **increasing atomic number**, with elements having similar properties positioned in the same vertical column, known as a **group**. The Periodic Table on page 39 is the one that you will be provided with during annual and SEC examinations. A larger copy is provided in section 5.

There are three types of elements: **metals, non-metals,** and **metalloids**.

- metals are situated to the left-hand side of the Periodic Table,
- non-metals are situated to the right-hand side,
- metalloids are sandwiched in between the metals and non-metals and have the properties of both metals and non-metals e.g., silicon is a shiny brittle black solid and is a semi-conductor.

The following picture shows which elements are metals, non-metals, and metalloids.



Source: https://chem.libretexts.org

The vertical columns are called **groups** and the horizontal rows are called **periods.**

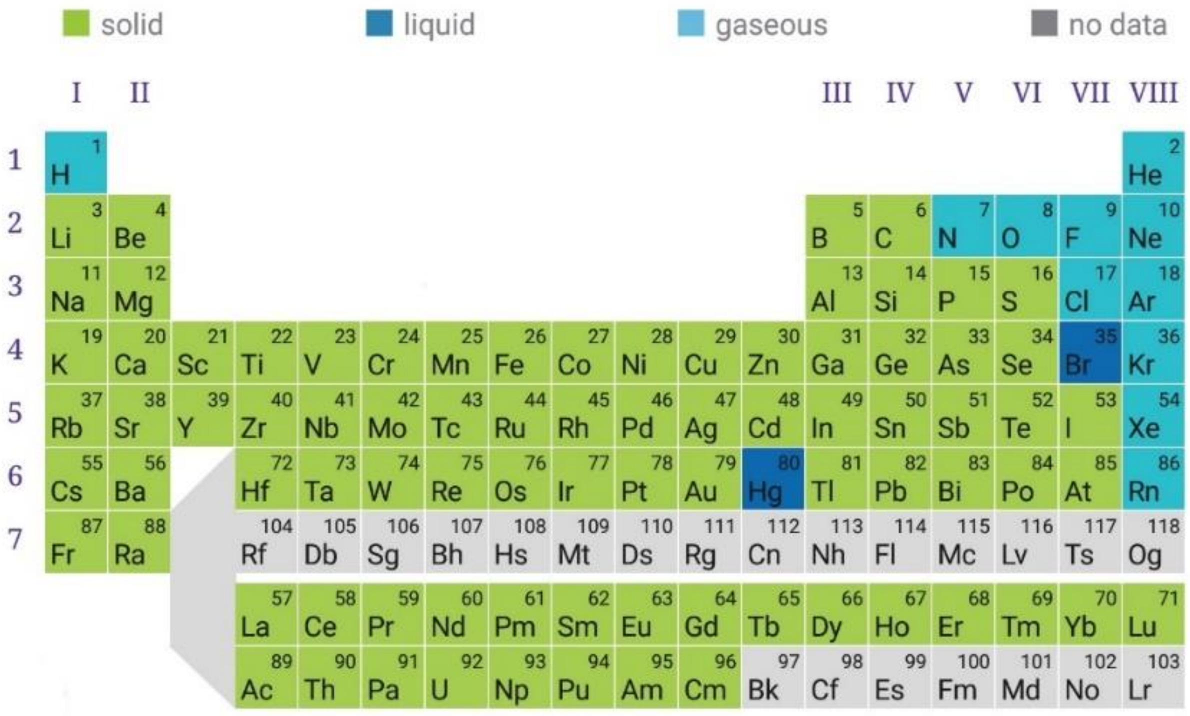
- The elements in a group have similar chemical properties due to their similar atomic structure.
- The properties of the elements in the same period change when going from left to right such that their nature changes from being metallic to non-metallic.

Crash Course Chemistry: The Periodic Table

Chemicool website including detailed information about all the elements.

Elements exist in different states of matter at room temperature (normally taken as 25 °C). The states of matter are solid, liquid, and gas. All metals are solid at room temperature except for mercury which is a liquid. Non-metals exist as gases or solids depending on the size of their particles except for bromine which is a liquid at room temperature.

The following are links to online Periodic Tables: Periodic Table 1 and Periodic Tables 2. There are also apps which can be downloaded onto a smart phone or tablet which are very useful to look up information on the elements such as the ones offered by Merck and the Royal Society of Chemistry.



🟩 solid
🟦 liquid
🟦 gaseous
⬛ no data

	I	II											III	IV	V	VI	VII	VIII
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og
				57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
				89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Screenshot from Merck app

2.3.1 Chemical naming and symbols

Chemists use symbols that are a kind of shorthand system for writing the names of elements and compounds. Each element has a particular one or two-letter symbol. The first letter of a symbol must be written in upper case, and if there is a second letter, it must be written in lower case e.g., magnesium (Mg). A complete set of symbols for the elements is found in a Periodic Table. The table below shows the name and symbols of the first 18 elements.

Element	Symbol	Element	Symbol
Hydrogen	H	Neon	Ne
Helium	He	Sodium	Na
Lithium	Li	Magnesium	Mg
Beryllium	Be	Aluminium	Al
Boron	B	Silicon	Si
Carbon	C	Phosphorus	P
Nitrogen	N	Sulfur	S
Oxygen	O	Chlorine	Cl
Fluorine	F	Argon	Ar

One might note that not all symbols tally with their names. The following table lists such elements and their symbols as well as their Latin names.

Element	Symbol	Latin Name
Sodium	Na	Natrium
Potassium	K	Kalium
Iron	Fe	Ferrum
Copper	Cu	Cuprum
Silver	Ag	Argentum
Tin	Sn	Stannum
Antimony	Sb	Stibium
Tungsten	W	Wolfram
Gold	Au	Aurum
Mercury	Hg	Hydrargyrum
Lead	Pb	Plumbum

1. There are more metals than non-metals in the Periodic Table. Write which of the following element is a metal, non-metal or metalloid.

Element	**Metal, metalloid, or non-metal?**
Carbon	
Hydrogen	
Oxygen	
Silicon	
Helium	
Magnesium	
Aluminium	
Germanium	
Neon	

2. Using a Periodic Table:

Write the symbols of the following elements:

lithium	calcium	iron	magnesium	astatine	boron	sulfur

Write the names of the following elements:

H	O	N	K	Na	Cu	Ag

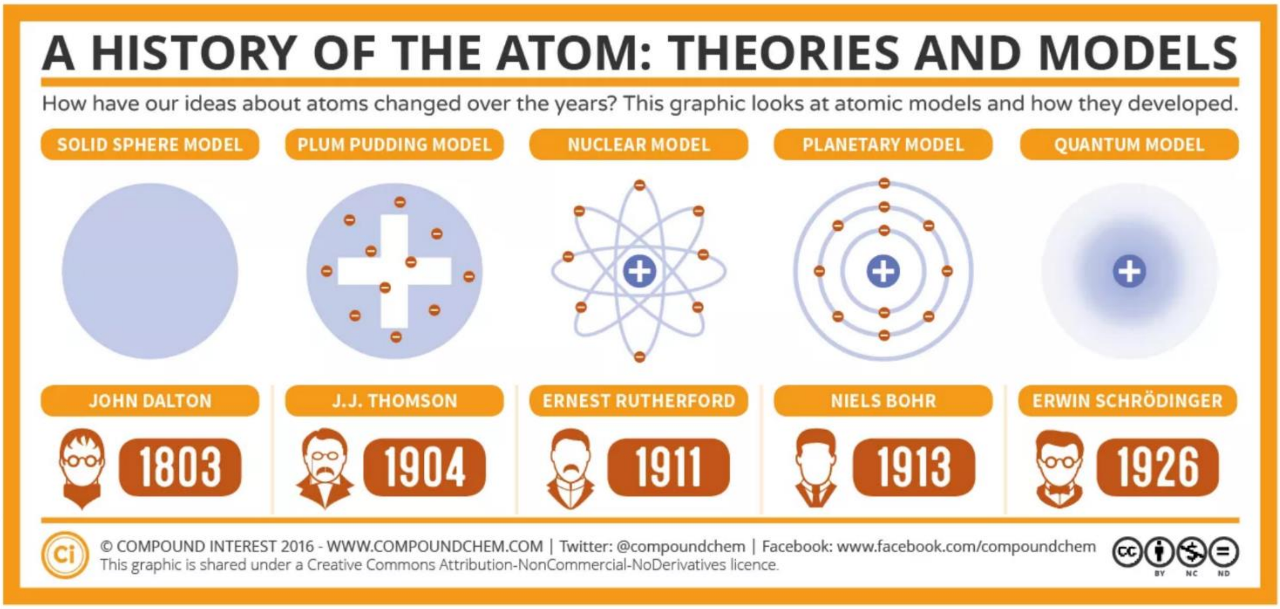
3. Use an online Periodic Table to find some information about gold and aluminium.
   a. When were these elements discovered?
   b. Give two uses of each element.

2.4 The nuclear model - protons, neutrons, and electrons

By the end of this section, you should be able to:

- Use a Periodic Table to describe and/or model atoms showing differences between atoms. *(E.g. subatomic particles - protons, neutrons, and electrons; atomic number, mass number, isotopes, and relative atomic mass.)* (2.2d)
- Calculate relative atomic mass from isotopic data. (2.3d)
- Determine the electron configuration of the first 18 elements in relation to their position in the Periodic Table. (2.2e)

Scientists have often wondered what matter is made of. One of the Greek philosophers, Democritus (460–370 BC), proposed that everything is composed of atoms. However, his ideas were not so popular at the time. Atoms were identified for the first time in 1958 using an electron microscope.



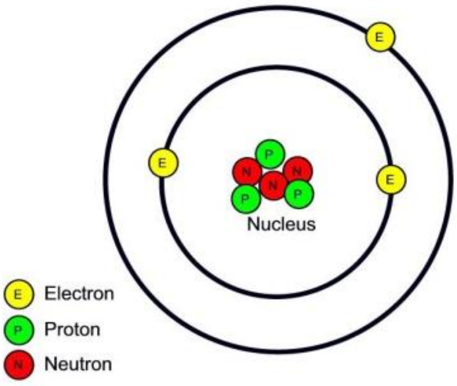
A theory of atomic structure is known as the **Nuclear Model**. This model suggests that the atom consists of:

- a very small and extremely dense region called the nucleus; and
- a cloud of negative electrically charged particles surrounding the nucleus at a distance.

The nucleus is very small when compared to the rest of the whole atom. For example, if the size of the atom were to be the size of a football stadium, the nucleus would be the size of a hazelnut on the kick-off spot.

This model also suggests that the atom consists of three smaller **sub-atomic** particles which are the **protons**, **neutrons,** and **electrons**.

These particles are extremely small and light, such that their mass is impractical to be measured in grams; rather, **atomic mass units** (or **amu**) are used instead.



Protons: are found in the nucleus of every atom and carry one positive charge each (i.e., its charge is +1). Its mass is approximately taken as 1 amu.

Neutrons: are uncharged, have approximately the same mass of protons and form part of the nuclei of all atoms, except that of hydrogen, $^{1}_{1}H$ since it has no neutrons.

Electrons: each carries one negative charge (i.e., has a charge of -1) and has a mass of about 1/1837 of the proton; because of this it is taken to have a mass of 0 as its mass is negligible.

Electrons are found moving at very high speed around the nucleus. Electrons exist in different shells, each shell having a different amount of energy increasing on moving further away from the nucleus. The electrons are held within the atom by an electrostatic force of attraction towards the positive charge of the protons in the nucleus.

Properties of sub-atomic particles

Sub-Atomic Particle	Charge	Mass (amu)	Found in
proton	+1	approx. 1	nucleus
neutron	nil	1	nucleus
electron	-1	negligible	electron shells

Discovery of Protons and Neutrons

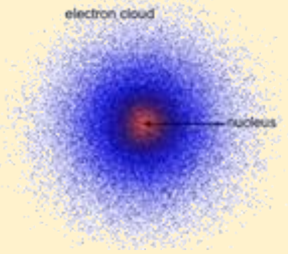
Ernest Rutherford, a physicist from New Zealand, was the scientist who discovered the nucleus within the atom. Through his work he found that the nucleus could be broken down into smaller particles.

In 1911, the Rutherford model of the atom was formulated, and it showed that most of the mass of the atom was concentrated within the nucleus. Moreover, he also noted that the nucleus had a positive charge and occupied only a very small space at the centre, and it was orbited by electrons with low mass.

Discovery of the electron – J.J. Thomson

The first subatomic particle to be discovered was the electron, in 1897, by the English physicist, Sir Joseph John Thomson. Before 1897, atoms were believed to be the smallest indivisible particle that existed. Using a cathode ray tube, Thomson was able to show the existence of subatomic particles that were around 1800 times lighter than the lightest atom (that of hydrogen). Later these were named electrons and were identified as the particles which flow through a metal that carries electric current.

Electron orbitals



At Intermediate and Advanced level Chemistry you will learn that electrons are found in a region of space called an orbital. It is very difficult to give the exact position of the electron within this electron cloud.

2.4.1 Atomic number and mass number

To describe the number of protons, neutrons, and electrons in an atom, scientists use two numbers: the **atomic number (Z)** and the **mass number (A)**.

$${}^{A}_{Z}X$$

This shorthand notation gives us information about the atomic number and mass number of an atom. The symbol of the element is written as (X). At the top, left-hand corner of the symbol we write the mass number (A) and at the bottom left-hand corner of the symbol we write the atomic number (Z).

2.4.1.1 Atomic number

The **atomic number** of an atom is the **number of protons** in the nucleus and is given the symbol **Z**.

atomic number (Z) = number of protons

Each element has its own atomic number, and no two elements have the same atomic number. For example, the element with an atomic number of 26 is iron. If you look at the Periodic Table, you will note that each element has a different atomic number! Therefore, the number of protons in an atom or the atomic number gives identity to the atom.

Complete the table by writing the number of protons for each element. The first two have been done for you.

symbol	${}^{1}_{1}H$	${}^{4}_{2}He$	${}^{7}_{3}Li$	${}^{9}_{4}Be$	${}^{11}_{5}B$	${}^{12}_{6}C$	${}^{14}_{7}N$	${}^{16}_{8}O$
No of protons	1	2						

Since the **atom** is neutral and has no overall electric charge, the number of protons (positively charged) is equal to the number of electrons (negatively charged).

The **number of electrons** of an atom is equal to the **atomic number**.

Complete the table by writing the number of protons and electrons for each element. The first two have been done for you.

Symbol	${}^{1}_{1}H$	${}^{4}_{2}He$	${}^{7}_{3}Li$	${}^{9}_{4}Be$	${}^{11}_{5}B$	${}^{12}_{6}C$	${}^{14}_{7}N$	${}^{16}_{8}O$
No of protons	1	2						
No of electrons	1	2						

Crash Course Chemistry: The Nucleus

2.4.1.2 Mass number

The **mass number** of an atom is the **number of protons and neutrons** in the nucleus and is given the symbol **A.** This is also called the nucleon number.

mass number (A) = number of protons (Z) + number of neutrons (N)

The number of neutrons can be obtained by subtracting the atomic number, Z from the mass number, A.

number of neutrons (N) = mass number (A) - number of protons (Z)

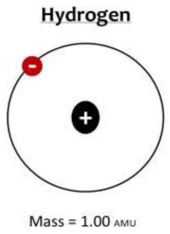
The atomic number and mass number of an element are shown in the table below.

Element	${}^{A}_{Z}\mathrm{X}$	Atomic Number	Mass Number
hydrogen	${}^{1}_{1}\mathrm{H}$	1	1
helium	${}^{4}_{2}\mathrm{He}$	2	4
lithium	${}^{7}_{3}\mathrm{Li}$	3	7
beryllium	${}^{9}_{4}\mathrm{Be}$	4	9
boron	${}^{11}_{5}\mathrm{B}$	5	11
carbon	${}^{12}_{6}\mathrm{C}$	6	12
nitrogen	${}^{14}_{7}\mathrm{N}$	7	14
oxygen	${}^{16}_{8}\mathrm{O}$	8	16
fluorine	${}^{19}_{9}\mathrm{F}$	9	19
neon	${}^{20}_{10}\mathrm{Ne}$	10	20

For example,

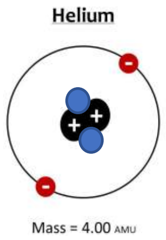
Hydrogen is ${}^{1}_{1}\mathbf{H}$.

- Hydrogen has only one proton in its nucleus.
- Since atoms have no net electric charge, i.e., overall, they are electrically neutral; a hydrogen atom would have one proton in its nucleus with one electron going around it.
- This hydrogen atom has no neutrons.



Helium is ${}^{4}_{2}\mathbf{He}$.

- The atomic number (Z) that is number of protons is 2.
- Since the number of electrons is equal to the number of protons, the number of electrons is 2.
- The mass number (A) is 4. The number of neutrons in an element can be calculated by subtracting Z from A: 4 - 2 = 2.



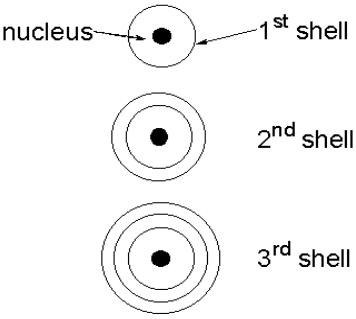
Work out the number of neutrons of the following elements.

Symbol	${}^{7}_{3}\mathrm{Li}$	${}^{9}_{4}\mathrm{Be}$	${}^{11}_{5}\mathrm{B}$	${}^{12}_{6}\mathrm{C}$	${}^{14}_{7}\mathrm{N}$	${}^{16}_{8}\mathrm{O}$	${}^{19}_{9}\mathrm{F}$	${}^{20}_{10}\mathrm{Ne}$
No of neutrons								

2.4.2 Electrons in shells

Electrons, moving at very high speed around the nucleus of an atom, possess varying quantities of energy. Electrons can be arranged in sets of energy levels (or shells). These shells can accept a limited number of electrons.

- The shell closer to the nucleus (or the first shell) can take up to 2 electrons.
- The second shell can take up to 8 electrons.
- While the third shell takes up to 8 electrons up to argon.

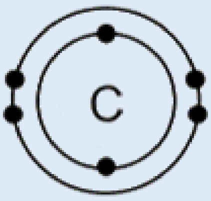


Larger atoms have other shells, but for now it is enough to consider these three shells. The way the electrons are arranged in an atom is called the **electron configuration** or **electron structure**.

Electron configuration for atoms of the first 18 elements

Element	Protons	Electrons	Electron configuration
hydrogen	1	1	1
helium	**2**	**2**	**2**
lithium	3	3	2, 1
beryllium	4	4	2, 2
boron	5	5	2, 3
carbon	6	6	2, 4
nitrogen	7	7	2, 5
oxygen	8	8	2, 6
fluorine	9	9	2, 7
neon	**10**	**10**	**2, 8**
sodium	11	11	2, 8, 1
magnesium	12	12	2, 8, 2
aluminium	13	13	2, 8, 3
silicon	14	14	2, 8, 4
phosphorus	15	15	2, 8, 5
sulfur	16	16	2, 8, 6
chlorine	17	17	2, 8, 7
argon	**18**	**18**	**2, 8, 8**

Draw diagrams to show the electron structure of the element of magnesium and chlorine using the information from the above table. (Hint: Draw the symbol of the element to represent the nucleus of the element. Then draw circles around the nucleus to represent the electron shells. In this example carbon has 6 electrons and its electronic configuration is 2,4. So 2 electrons go in the 1st electron shell and the remaining 4 electrons in the 2nd electron shell.



The third shell can take up to 18 electrons. This is true for elements whose atomic number is larger than that of calcium. This will be dealt with at Intermediate and Advanced level.

Niels Bohr and the electron configuration

Niels Bohr was a Danish physicist who contributed to the understanding of the atomic structure. In 1923 he proposed that the electron structure of an atom, could help in explaining the properties of the elements. He suggested that electrons orbit in shells, which were at fixed distances from the nucleus. It was later, around 1936, that this theory was validated (with some corrections), because of experiments involving atomic spectra.

Go-Lab: Atom and Ion Builder

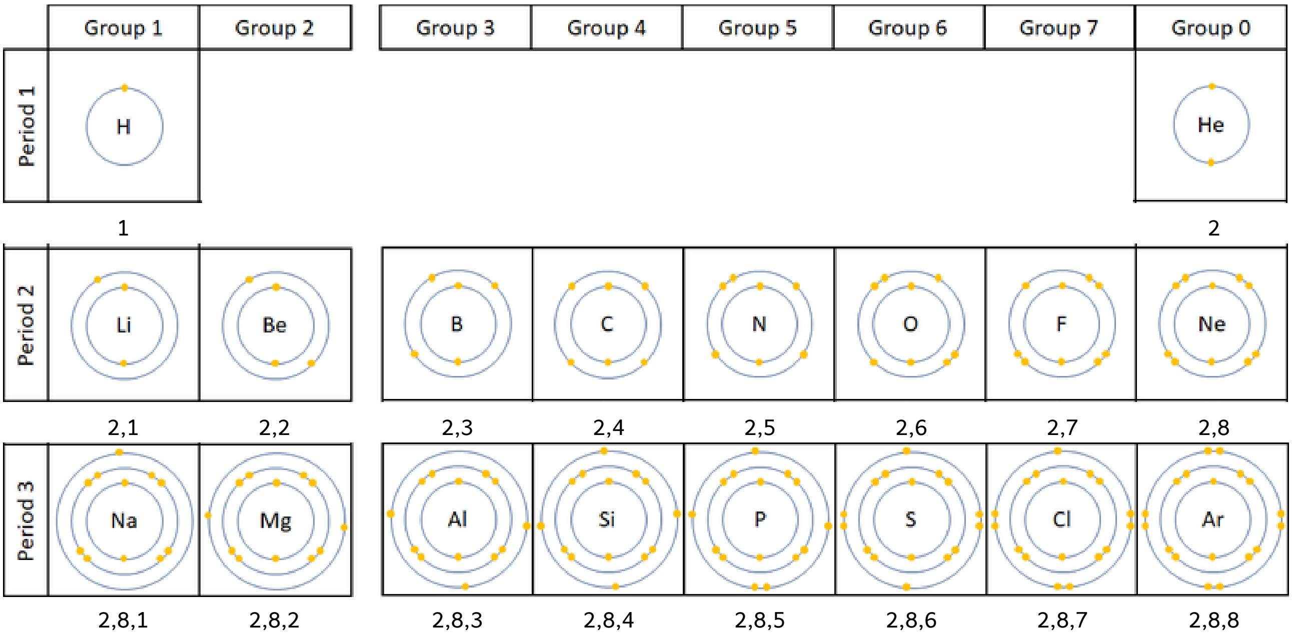
1. Complete the following paragraph:
   - a) Matter is made of smaller particles called ____________________.
   - b) These particles are made of smaller particles called ________________ particles.
   - c) There are three kinds of these smaller particles which are called _______________, ________________ and _________________.
   - d) In the nucleus there are __________________ and __________________.
   - e) The atomic number represents the number of _________________.
   - f) The mass number represents the number of ___________________________________.
   - g) The ______________ are found orbiting in shells.
   - h) The first shell can take ______________, the second shell can take _________________, the third shell can take ________________________.
   - i) An atom is neutral because it has ____.

2. Every element has an atomic number and a mass number.
   - a) Define atomic number.
   - b) Define mass number.
   - c) Use the Periodic Table to complete the missing information.

Element	**Atomic number**	**Electron configuration**	**Mass number**	**Number of neutrons**	**Group**
sodium					
fluorine					
magnesium					
oxygen					
nitrogen					
lithium					
sulfur					
hydrogen					
aluminium					
neon					

2.4.3 Relationship between electron structure, group, and period number

Look at the electronic structure of the first 18 elements. What patterns do you notice?



The rows are known as **periods** and in

- period 1 there are hydrogen and helium with 1 electron shell.
- period 2 there are the elements from lithium to neon with 2 electron shells.
- period 3 there are the elements from sodium to argon with 3 electron shells and so on.

Note that across a period, the outer electron shell is filling up till it becomes a full outer shell.

The **period number** shows the number of shells present in an atom.

The columns are known as **groups**.

Elements in the same group have the same number of electrons in their outer shell. E.g.: Elements in group 1 have 1 electron in their outer shell, group 2 have 2 electrons etc. Noble gases found in group 0 have a full outer shell. Since elements in the same group have the same number of electrons in their outer shell, they have similar chemical properties.

All elements in a group have the same number of electrons in their outermost shell.
The **group number** is the same as the number of electrons in the outer shell.

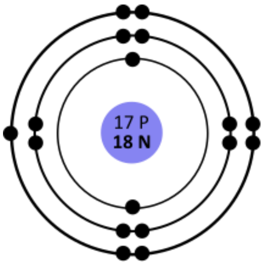
Note that in some Periodic Tables hydrogen is placed in group 1, but hydrogen does not have the same properties as the group 1 elements. Ideally it should be placed on its own.

1. An atom has an electronic configuration of 2, 8, 5.
   a) In which group is this element found? Explain.
   b) In which period is this element found? Explain.
   c) Name the element and state its symbol.
   d) Work out the number of protons, neutrons, and electrons of this element.

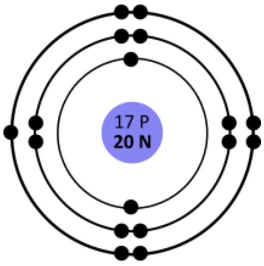
2.4.4 Isotopes

In some elements not all the atoms are identical. Whereas all the atoms of an element contain the same number of protons, the number of neutrons may vary, i.e., atoms of an element all have the same atomic number but may have a different mass number.

Isotopes are atoms of the same element that have the same number of protons but different number of neutrons.



chlorine - 35



chlorine - 37

For example, there are two isotopes of chlorine, one which has an atomic mass of 35 and the other which has an atomic mass of 37.

Isotopes	Number of protons	Number of electrons	Number of neutrons
$^{35}_{17}Cl$	17	17	18
$^{37}_{17}Cl$	17	17	20

Isotopes have the same chemical properties because they have the same electronic structure. Due to more neutrons the elements have different physical properties e.g., density.

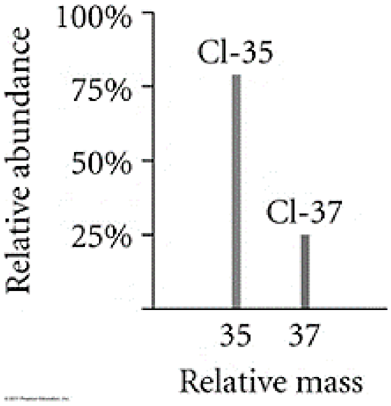
Some of the isotopes of certain atoms are unstable due to the extra number of neutrons. They are said to be radioactive. Radioactivity is a process by which atoms disintegrate to form other atoms. E.g., Uranium-235 is used in nuclear reactors. Radioactivity is dealt with in Physics.

Many elements have different isotopes. The table below provides some examples of isotopes of different elements.

Hydrogen	Carbon	Oxygen	Uranium
$^{1}_{1}H$ $^{2}_{1}H$ $^{3}_{1}H$	$^{12}_{6}C$ $^{13}_{6}C$ $^{14}_{6}C$	$^{16}_{8}O$ $^{17}_{8}O$ $^{18}_{8}O$	$^{235}_{92}U$ $^{238}_{92}U$

Isotopes were discovered using an instrument called a mass spectrometer (which will be dealt with at Advanced level). The detector in this instrument not only measures how many isotopes are present but also the **percentage abundance**. The relative abundance of an isotope is the proportion of one particular isotope in a mixture of isotopes found in nature.

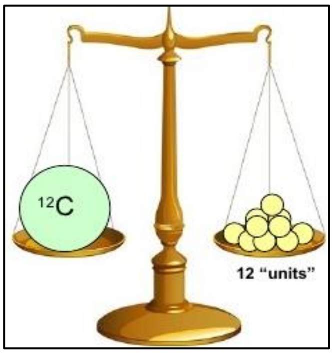
For example, in a sample of chlorine 75% of atoms are chlorine-35 while 25% of the atoms are chlorine-37.



2.4.5 Relative atomic mass

The average mass of a large number of atoms of an element is called the **relative atomic mass** (**RAM** or **A_r**) This quantity considers the percentage abundance of all the isotopes of an element which exist.

In 1961, the International Union of Pure and Applied Chemistry (IUPAC), recommended for a new standard to be adopted where **the mass of all atoms is compared to carbon-12**, which was taken to have a mass of 12 units or amu. The relative atomic mass of an element is the average mass of the naturally occurring atoms of an element on a scale where ^{12}C has a mass of exactly 12 units.



Source: https://www.tes.com/teaching-resource/relative-atomic-mass-and-relative-formula-mass-12219862

Relative atomic mass is defined as the average mass of isotopes of an element compared to $\frac{1}{12}$th of the mass of an atom of ^{12}C.

$$\text{RAM} = \frac{\text{average mass of isotopes of element}}{\frac{1}{12} \times \text{mass of 1 atom of carbon-12}}$$

$$\frac{1}{12} \text{ of the mass of 1 carbon-12 atom} = 1 \text{ amu}$$

The relative atomic mass of elements is calculated as follows:

E.g., chlorine has two isotopes:

- chlorine-35 (abundance = 75%),
- chlorine-37 (abundance = 25%).

This means that in nature, 75% of the chlorine atoms is the Cl-35 isotope and 25% is the Cl-37 isotope or that in nature chlorine-35 and chlorine-37 exist in a ratio of 3:1.

Relative atomic mass of chlorine:

$$= \frac{(75 \times 35) + (25 \times 37)}{100} = \mathbf{35.5} \quad \text{OR} \quad = \frac{(3 \times 35) + (1 \times 37)}{4} = \mathbf{35.5}$$

For compounds the term **relative molecular mass (RMM)** or **relative formula mass (RFM)** is used. RMM is used for covalently bonded substances and RFM is used for ionically bonded substances. RMM and RFM can be calculated by referring to the formula of the compound and the Periodic Table.

RMM of CO_2:
elements present 1 carbon and 2 oxygen
C: 1 x 12 = 12
O: 2 x 16 = 32
RMM of CO_2 = 12 + 32 = 44

RFM of NaCl:
elements present 1 sodium and 1 chlorine
Na: 1 x 23 = 23
Cl: 1 x 35.5 = 35.5
RFM of NaCl = 23 + 35.5 = 58.5

Note that RAM, RMM and RFM have no units since it is a comparison of masses.

Go-Lab: Isotopes and atomic mass

1. Use the Periodic Table to complete the following exercise:

Element	**Atomic Number**	**Mass Number**	**Number of Neutrons**
Carbon			
Oxygen			
Magnesium			

2. What are isotopes?

3. Considering ^{35}Cl and ^{37}Cl.
   a. State one similarity and one difference between the two isotopes.
   b. Explain why these two isotopes react in the same way.

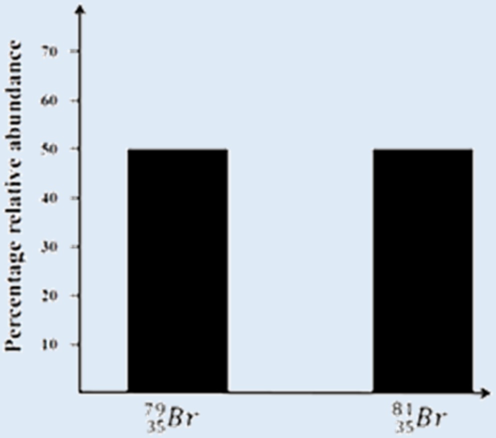
4. One gas present in air is neon and exists in two forms of isotopes: neon-20 and neon-22 in a 9 : 1 ratio. Use this information to calculate the relative atomic mass of neon in air.

5. Calculate the relative atomic mass of magnesium if it occurs naturally with the following percentage abundance of each isotope: 79% of magnesium-24, 10% of magnesium 25 and 11% of magnesium 26.

6. A sample of gallium is 60% ^{69}Ga atoms and 40% ^{71}Ga atoms. Calculate the relative atomic mass of this sample of gallium.

7. The relative atomic mass of boron is 10.81. Boron has two isotopes: ^{10}B and ^{11}B. Which of these two isotopes is more abundant? Explain your answer.

8. There are two of atoms of bromine. One type has a mass number of 79 and the other has a mass number of 81. The graph of the percentage relative abundance of these forms is shown below.
   a) From the graph what can you conclude about their percentage abundance?
   b) Use the graph of the relative abundance of the two forms to calculate the relative atomic mass of bromine.



9. Using the periodic table state the RAM of: a) K, b) Ne c) Be

10. Calculate the RMM of: a) O_2 b) H_2O c) $C_6H_{12}O_6$

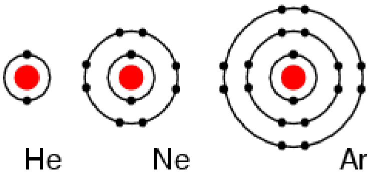
11. Calculate the RFM of: a) KBr b) $MgCl_2$ c) $CuSO_4$

2.4.6 Monoatomic elements

By the end of sections 2.4.6 and 2.4.7, you should be able to:

- Distinguish between gases which are monoatomic and others which are diatomic. *(Limited to noble gases, H_2, N_2, and O_2)* (2.1f)

Noble gases do not react readily with other elements to form compounds. They are practically unreactive because they have a full outer shell of electrons. These elements are found in Group 0 of the Periodic Table. Noble gases are **monoatomic**.



The chemical formula for a noble gas is written as He for helium, Ne for neon, Ar for argon etc.

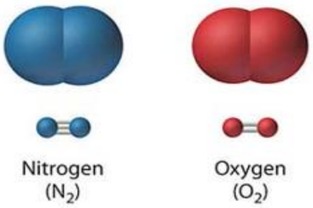
2.4.7 Formation of molecules

Atoms are the most stable when they acquire a full outer shell of electrons, also known as a noble gas configuration. Since noble gases have a full outer shell, they are stable and unreactive.

When non-metals which do not have a full outer shell of electrons react together, they share electrons to reach a noble gas configuration. A **molecule** is formed when a group of atoms are held together by covalent bonds.

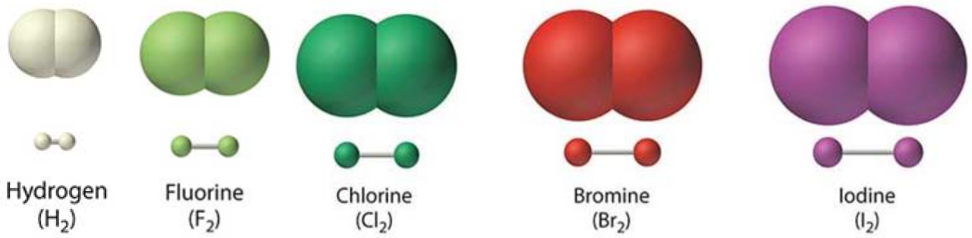
2.4.7.1 Elements

Gases in air like nitrogen and oxygen form **diatomic molecules.** These molecules are made of two atoms of the same element joined together by a bond known as the **covalent bond**. When non-metal atoms of the same element combine to form molecules, they are still elements, they are not compounds!



Covalent bonding happens **only between non-metal atoms**.

Hydrogen, and Group 7 elements fluorine, chlorine, bromine, and iodine all form diatomic molecules. Hydrogen, nitrogen, oxygen, fluorine, and chlorine are gases at room temperature and pressure (RTP). Bromine is a liquid while iodine is a solid at RTP. The formulae of these elements are written as H_2 for hydrogen, F_2 for fluorine, etc.

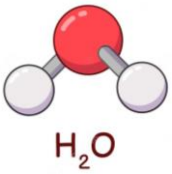


Other non-metal elements although they might exist in molecules of varying combinations, are always written with the name of their element. E.g., although sulfur exists as S_8 molecules, the chemical formula for sulfur is written as S.

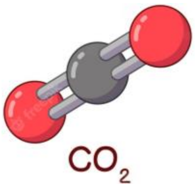
Diatomic Elements in the Periodic Table

2.4.7.2 Compounds

Atoms of one element may join with atoms of a different element to form a **new substance**. These new substances are called **compounds**.



Water is a compound that exists in the atmosphere in the form of water vapour. It exists in large quantities in liquid form in lakes, rivers, and the sea. It is found in the solid state as ice/snow at Earth's poles and mountain tips. Water is a compound made of oxygen and hydrogen. It consists of one atom of oxygen chemically combined with two atoms of hydrogen. Since this compound consists of non-metal atoms only, it consists of molecules. Its chemical formula is H_2O.



Carbon dioxide is another compound that is found in air. It consists of one carbon atom chemically combined to two oxygen atoms. Its chemical formula is CO_2.

Both water vapour and carbon dioxide are molecules whose atoms are covalently bonded.

2.4.8 Valency

Water has a chemical formula of H_2O. One thing that baffled scientists over the ages was how the ratio of atoms of different elements in a compound can be predicted. Why is the formula of water H_2O and not HO_2, or some other combination?

To determine how elements combine one first needs to work out the electron configuration of the elements as in section 2.4.2. The next step is to determine how many electrons can be lost, gained, or shared for the atom to reach a noble gas configuration, that is a full outer shell. Valency is a measure of the power of an element to combine with other elements. Elements combine chemically together when they gain, lose, or share electrons to achieve a noble gas configuration.

Valency of an element is the number of electrons its atoms lose, gain, or share to obtain a noble gas configuration.

Atoms join up with other atoms in one of two ways.

- By sharing electrons - in **covalent** bonding between non-metals only, or
- By gaining or losing electrons - in **ionic** (or **electrovalent**) bonding between metals and non-metals. Ionic bonding will be dealt with in the next chapter.

By using the Periodic Table, the valency of elements can be worked out.

Periodic Table group	Number of electrons in outer shell	Electrons lost, gained, or shared	Valency number
1	1	1 lost	1
2	2	2 lost	2
3	3	3 lost	3
4	4	4 gained or shared	4
5	5	3 gained or shared	3
6	6	2 gained or shared	2
7	7	1 gained or shared	1
0	8	none	0

2.5 Covalent bonding

By the end of this section, you should be able to:

- Explain how covalent bonds are formed. (2.3g)

Covalent bonding **exists between non-metal atoms**. The non-metal atoms **share electrons** and form **molecules**. The non-metal atoms achieve a complete outer electron shell of the nearest noble gas by sharing electrons.

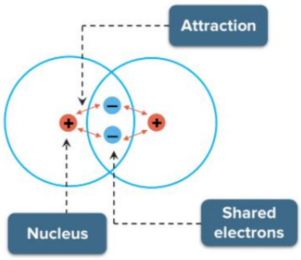
The number of electrons required for a non-metal atom to have a full outer shell is worked out by subtracting the number of electrons in outer shell from 8 (for the first 18 elements and except for hydrogen). This number shows the valency of the element.

The number of covalent bonds a non-metal can form is linked to its position in the Periodic Table and to the valency. The following table shows how one can work out the number of covalent bonds that can form for elements in Groups 4, 5, 6, and 7. Note that Groups 1, 2 and 3 are metals and do not form covalent bonds. Also, the noble gases are unreactive and do not usually form covalent bonds.

Group in the Periodic Table	4	5	6	7
Number of electrons in outer shell	4	5	6	7
Number of electrons required to have a full outer shell	4	3	2	1
Valency	4	3	2	1
Covalent bonds formed	4	3	2	1

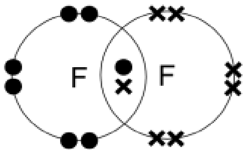
A covalent bond is a **very strong bond,** and it takes a large quantity of energy to break.

These bonds are called **intramolecular bonds** since these are bonds that are found within a molecule.

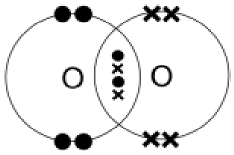


A **covalent bond** is the result of the attraction between the bonding pair of the electrons (negative charge) and the nuclei (positive charges) of the atoms involved.

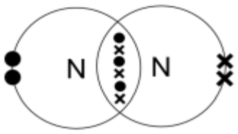
Covalent bonds come in three varieties: single, double, or triple covalent bonds.



- **A single covalent bond** forms when two atoms each contribute one electron to a shared pair of electrons. A single covalent bond is represented by a single line e.g., F-F.



- **A double covalent bond** forms when two atoms each contribute two electrons to a have 2 pairs of shared electrons. A double covalent bond is represented by a double line e.g., O=O.



- **A triple covalent bond** forms when two atoms each contribute three electrons to a have 3 pairs of shared electrons. A triple covalent bond is represented by a triple line e.g., N≡N.

2.5.1 Covalent bonding between atoms of the same non-metals

By the end of sections 2.5.1 and 2.5.2 you should be able to:

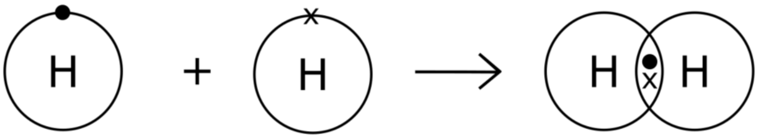
- Represent covalent bonds using dot and cross diagrams showing outer electron shells only. *(E.g. hydrogen, oxygen, nitrogen, chlorine, methane, water, carbon dioxide, ammonia, and hydrogen chloride.)* (2.3h)

This section shows dot-cross diagrams of elements that are made up of diatomic molecules, like hydrogen, fluorine, oxygen, and nitrogen. For covalent bonding you will be asked to draw dot-cross diagrams **showing outer electron shells only**.

2.5.1.1 Hydrogen

A hydrogen atom has only 1 electron. Its shell can hold 2 electrons to become full. When two hydrogen atoms get close enough their electron shells overlap and share electrons. A hydrogen molecule is formed with chemical formula **H_2**. There is a single covalent bond between the hydrogen atoms and is represented as H-H.

The following dot-cross diagram represents the bonding present in hydrogen. The dots and crosses represent electrons from different atoms.

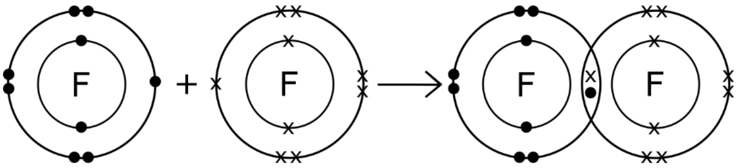


2.5.1.2 Fluorine

An atom of fluorine has an atomic number of 9 with an electron configuration of 2, 7.

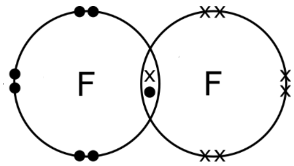
Each fluorine atom needs 1 electron to reach a noble gas configuration. A fluorine atom shares the unpaired electron in their outer shell with another fluorine atom to form a fluorine molecule.

The following dot-cross diagram shows how single covalent bonding happens. The outer shells of the two atoms overlap and one electron from each atom is shared such that both atoms achieve a noble gas configuration. Note that a single covalent bond contains two shared electrons one from each atom.



You will be asked to draw dot-cross diagrams showing outer electron shells only, so the first shell is omitted as shown below.

Chemical Formula: F_2
Displayed formula: F-F
The hyphen between the fluorine atoms shows that there is a single covalent bond between the fluorine atoms.



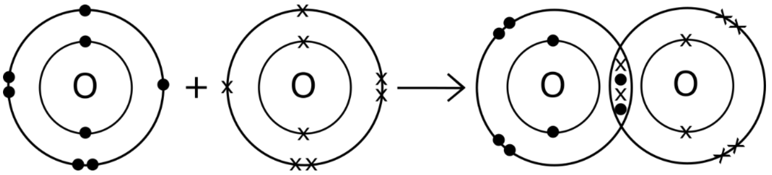
All the halogens, that is chlorine, bromine and iodine exhibit the same type of bonding like fluorine. They form diatomic molecules that is Cl_2, Br_2, and I_2.

2.5.1.3 Oxygen

Oxygen has 8 electrons with an electron configuration of 2, 6.

An atom of oxygen needs 2 electrons for it to have a noble gas configuration. Oxygen atoms join up in twos so that each oxygen atom acquires a noble gas configuration by sharing two of its outer shell electrons.

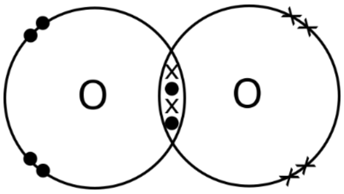
The diagram below shows how the outer shells of 2 oxygen atoms overlap to share two electrons each, and thus obtaining a full shell of electrons. A double bond is formed which contains 4 electrons, two from each atom.



Chemical Formula: O_2

Displayed formula O=O

A double covalent bond is present between oxygen atoms.

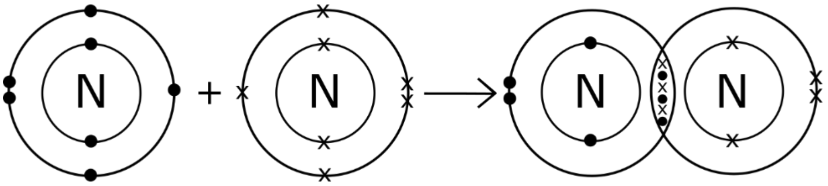


2.5.1.4 Nitrogen

Nitrogen has 7 electrons with an electron configuration of 2, 5.

Nitrogen needs 3 electrons to achieve a noble gas configuration.

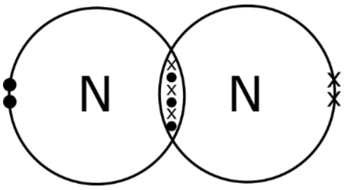
The following diagram shows how two atoms of nitrogen share three electrons each to form a nitrogen molecule. A triple covalent bond is formed which contains 6 shared electrons, 3 from each nitrogen atom.



Chemical Formula: N_2,

Displayed formula $N \equiv N$.

A triple covalent bond between the nitrogen atoms, makes nitrogen a fairly unreactive gas.



2.5.2 Covalent bonding between atoms of different non-metals

Covalent bonding can happen between different non-metal elements to form a variety of compounds. Here are the dot-cross diagrams of the following compounds: Carbon dioxide (CO_2), methane (CH_4), ammonia (NH_3), water (H_2O), and hydrogen chloride (HCl). These covalent compounds are all gases at RTP except for water which is a liquid.

There are different ways which you can use to draw the molecules. You can think in terms of how many electrons are required for each combining element to have a full shell or you can think in terms of valency. Both methods are presented below, select the one which you prefer.

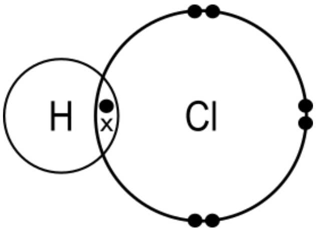
2.5.2.1 Hydrogen chloride

Hydrogen chloride (HCl) is a compound that is formed when hydrogen reacts with chlorine.

hydrogen + chlorine → hydrogen chloride

The dot-cross diagram below is a representation of the molecule of hydrogen chloride showing outer shell electrons only.

- Hydrogen with 1 electron in the outer shell needs to share one electron to gain an electronic configuration of a noble gas. Chlorine has 7 electrons in its outer shell, so it also needs to share one electron to gain a full outer shell. Both share the unpaired electron to form a single covalent bond.
- Both hydrogen and chlorine which have a valency of 1, share one electron each to form a single covalent bond.



2.5.2.2 Water

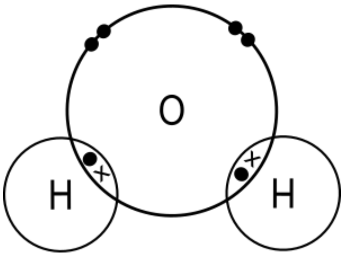
Water is another covalent substance which is found in the liquid state at room temperature and pressure.

Water is a compound that forms when hydrogen reacts with oxygen.

hydrogen + oxygen → water

The dot-cross diagram is a representation of the molecule of water showing outer shell electrons only.

- Oxygen has 6 electrons and requires another 2 to have a full outer shell. Hydrogen has 1 electron and requires another one, thus two hydrogen atoms share electrons with an oxygen atom.
- Since oxygen has a valency of 2 and hydrogen has a valency of 1, it takes 2 hydrogen atoms to combine with 1 oxygen atom.

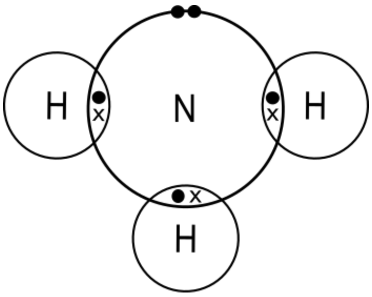


The resulting molecule has the chemical formula H_2O. Note that oxygen forms two single covalent bonds with hydrogen.

2.5.2.3 Ammonia

Nitrogen and hydrogen react together to form ammonia.

nitrogen + hydrogen → ammonia



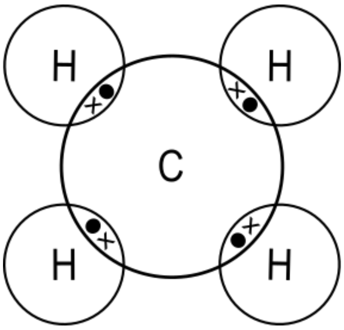
The dot-cross diagram is a representation of the ammonia molecule showing outer electron shells only.

- Nitrogen requires another 3 electrons to have a noble gas electronic configuration. Three hydrogen atoms can share an electron each to form the ammonia molecule.
- The valency of nitrogen is 3 while the valency of hydrogen is 1. It takes three hydrogen atoms to form three single covalent bonds with nitrogen.

2.5.2.4 Methane

Methane is a hydrocarbon that is it is made up of carbon and hydrogen. The dot-cross diagram is a representation of the methane molecule showing outer shell electrons only.

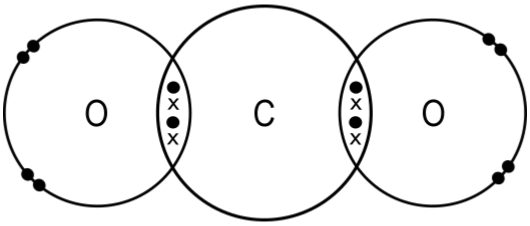
- Carbon requires 4 more electrons to achieve a full shell, thus 4 hydrogen atoms can share an electron each to form methane.
- Carbon has a valency of 4 while hydrogen has a valency of 1. It takes four hydrogen atoms to combine with 1 carbon atom. Carbon forms 4 single covalent bonds with 4 hydrogen atoms.



2.5.2.5 Carbon dioxide

Carbon reacts with oxygen to form carbon dioxide.

carbon + oxygen → carbon dioxide



The diagram is showing outer shell electrons only.

- Carbon requires another 4 electrons to achieve a full shell, whereas each oxygen atom requires 2 electrons. Each oxygen atom shares a pair of electrons with carbon to form two double bonds with carbon.
- Since the valency of carbon is 4 and the valency of oxygen is 2 then it takes two oxygen atoms to react with one carbon atom.

Note that carbon forms two double bonds with two oxygen atoms. You can tell since there are four electrons that are shared between each oxygen and carbon.

Gilbert N. Lewis

The notion of the sharing of electron pairs between atoms and chemical bonding, was first published in 1916, by Gilbert N. Lewis, an American physical chemist. He proposed that an atom shares enough electrons to form a full outer electron shell. Later, the term covalent bonding, was used in the article 'The Arrangement of Electrons in Atoms and Molecules', written by Irving Langmuir in 1919.

2.5.3 Chemical formulae of covalent compounds

When we refer to molecules of compounds, we do not write specific symbols for each and every compound, but we write **chemical formulae**. Taking water as an example, since one molecule of water consists of two atoms of hydrogen and one atom of oxygen, the chemical formula of water is H_2O. The reason behind chemical formulae having specific ratios is due to the fact that elements have particular valencies.

For example, in the case of water, you should recall that oxygen has a valency of 2 while hydrogen has a valency of 1. Since non-metals react by sharing electrons, oxygen shares two of its electrons with two hydrogen atoms that in turn share one electron each with oxygen. This means that both oxygen and hydrogen achieve a full outer shell of electrons. So, this is why the chemical formula of water is H_2O.

When dealing with names of molecular chemical formulae, there are a few rules listed below:

1. Some covalent molecules have common names, for example: water for dihydrogen oxide and ammonia for nitrogen trihydride.
2. You need to know which element to put first. The rule of thumb is to use the order in which the elements exist in the Periodic Table with the exception that hydrogen comes in between nitrogen and oxygen, for example: H_2O, HCl, HF, and HBr but NH_3 and CH_4.
3. Where multiples of an element exist within a molecule it is prefixed with mono- for 1, di- for 2, tri- for 3, tetra- for 4, penta- for 5, and hexa- for 6, etc. The **mono-** prefix is not used if only one kind of molecule exists. The mono- prefix is used when there are molecules that are similar, for example: CO_2 is carbon dioxide, however CO is carbon monoxide.
4. Note how the name of the element which is written at the back changes slightly such that it takes an **-ide** suffix.

1. Write the chemical formulae of the following covalent compounds or elements:

Name	Chemical formula	Name	Chemical formula
hydrogen chloride		oxygen	
ammonia		hydrogen sulfide	
methane		sulfur dioxide	
nitrogen		sulfur trioxide	

2. Give the names of the following chemical formulae.

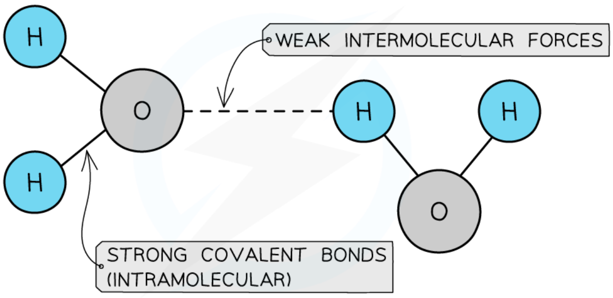
Chemical formula	Name	Chemical formula	Name
HBr		HI	
H_2O		NO	
CO		NO_2	
CO_2		CCl_4	

2.5.4 Properties of covalent compounds

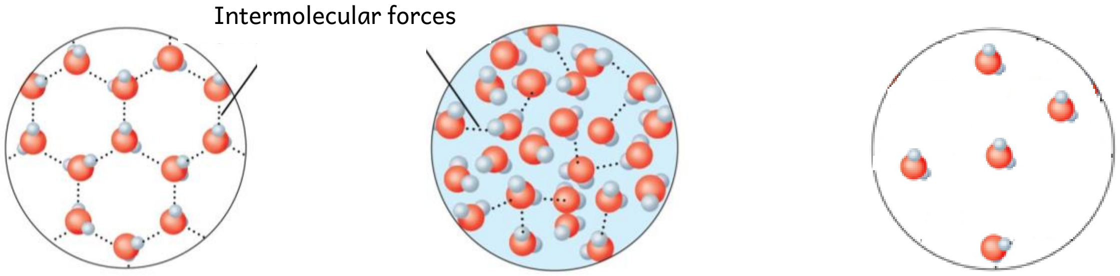
By the end of this section, you should be able to:

- Explain the properties of covalent substances for simple molecules. *(Limited to melting and boiling points, non-conduction of electricity.)* (2.3i)

Most non-metal elements exist as simple molecules with very weak forces of attraction between molecules. This feature leads non-metal elements to exist as gases or liquids at room temperature and pressure (RTP). Room temperature is taken as 25 °C and pressure as 1 atmosphere or 100 kPa. These covalent substances have a **simple molecular structure.** Simple molecular structures are made of molecules in which there are strong covalent bonds between the atoms within the molecule (intramolecular bonds) but have weak forces of attraction between the molecules (intermolecular bonds). The latter represented by dotted lines.



For example, there are weak forces of attraction between the water molecules. In ice there are many intermolecular forces of attraction between the molecules. In the liquid state, the particles are moving and so the weak forces of attraction are constantly breaking and reforming. When liquid water boils, it is these weak forces of attraction that break and in the gaseous form the water molecules are separate and move around at high speeds and in a random manner.



Ice:
many weak intermolecular forces

Liquid water:
intermolecular forces constantly break and reform

Water vapour:
no weak forces of attraction

Properties of simple covalent molecules:

- are gases or liquids at RTP,
- have low melting and boiling points at RTP. This is due to weak intermolecular forces of attraction which exist between simple molecules. They require low energy to break these bonds.
- do not dissolve in water but dissolve very easily in organic solvents,
- do not conduct electricity since covalent compounds have neither free electrons such as metals, nor free ions such as solutions of salts in water.

2.6 Representing chemical reactions

By the end of this section, you should be able to:

- Represent a chemical reaction using a word equation. (1.1s)
- Represent a chemical reaction using a balanced chemical equation. (1.2s)

2.6.1 Word equations

When substances react together, products are formed. Chemists may use word equations to represent what is happening during a chemical reaction.

For example, when calcium reacts with sulfur it produces calcium sulfide. This sentence can be represented by means of a **word equation**:

calcium + sulfur → calcium sulfide

Calcium and sulfur are called **reactants** while calcium sulfide is the **product** of this reaction. Note that reactants and products are separated by means of a single headed arrow pointing to the right (→) which shows that the reactants on the left-hand side of the arrow react to produce the products on the right-hand side of the arrow.

1. Write word equations for the following reactions:
    a. Nitrogen reacts with hydrogen to produce ammonia.
    b. Calcium reacts with chlorine to produce calcium chloride.
    c. Carbon reacts with oxygen to produce carbon dioxide.
    d. Hydrogen reacts with bromine to produce hydrogen bromide.
    e. Sulfur reacts with oxygen to form sulfur dioxide.

2.6.2 States of matter

Matter exists in one of three states of matter: **solid**, **liquid**, or **gas**. When writing chemical equations, normally it is required to include these state symbols with each and every reactant and product in the equation. If a reactant or product is dissolved in water its state of matter is said to be **aqueous**. You will find which substances are soluble or not when you learn the solubility rules in section 3.7.2.

Solids are represented with **(s)**, liquids with **(l)**, gases with **(g)** and aqueous solutions with **(aq)**.

The following are examples of word equations including state symbols:

hydrogen (g) + oxygen (g) → water (l)

carbon (s) + oxygen (g) → carbon dioxide (g)

methane (g) + oxygen (g) → carbon dioxide (g) + water (g)

2.6.3 Balanced chemical equations

Chemists represent chemical reactions by means of balanced chemical equations. This means that the word equations mentioned earlier need to be converted into symbols that all chemists can understand. This is done by following the steps below.

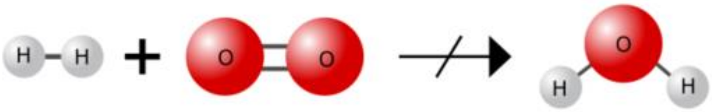
1. Make sure that the word equation includes all the reactants and products for a given reaction. For example:

   hydrogen + oxygen → water

2. Convert every reactant and product into its correct chemical formula. In this case:

   hydrogen: H_2, oxygen: O_2 and water: H_2O, so:

   $$H_2 + O_2 \nrightarrow H_2O$$

   Note that the number of particles of the reactants and products is not the same. This means that the equation is not balanced.

3. Check that the number of particles of each element on the left-hand side is equal to that on the right-hand side.

   Remember that one cannot change the formula of a compound. So, to balance the equation, only multiples of formulae can be used. The multiples (coefficients) are written in front of a chemical formula.

   a. To balance the above equation first count the number of atoms on each side of the equation:

   $$H_2 + O_2 \nrightarrow H_2O$$

   H:2 O:2 H:2; O:1

   b. Although the number of hydrogen atoms is equal on each side of the equation, the number of oxygen atoms is not, so place 2 in front of water.

   $$H_2 + O_2 \nrightarrow 2H_2O$$

   H:2 O:2 H:4; O:2

   c. Now the number of oxygen atoms are equal on each side of the equation, but the number of hydrogen atoms are not; so place 2 in front of hydrogen.

   $$2H_2 + O_2 \rightarrow 2H_2O$$

   H:4 O:2 H:4; O:2

   Now the number of hydrogen and oxygen atoms on both sides of the equation are equal and we say that the equation is balanced. So the balanced chemical equation is:

$2H_2 + O_2 \rightarrow 2H_2O$

In more complicated equations which involve several elements and/or compounds, always leave balancing of H and O for last. Balance elements which are least involved (that is present in the least number of compounds).

Let's consider another chemical reaction where methane (CH_4) also known as natural gas reacts with oxygen (O_2) in the air to produce carbon dioxide (CO_2) and water vapour (H_2O). The steps described above are shown in the table shown below.

1. Write a word equation.	methane + oxygen → carbon dioxide + water
2. Write down the correct formulae. From now onwards, chemical formulae MUST NEVER CHANGE.	$CH_4 + O_2 \rightarrow CO_2 + H_2O$
3. Count the number of atoms for each element on either side.	$CH_4 + O_2 \rightarrow CO_2 + H_2O$ C: 1 C: 1 H: 4 H: 2 O: 2 O: 3
4. Hint: Leave O & H for LAST! Check other elements which appear once. C is balanced. There is no need to make any changes.	$CH_4 + O_2 \rightarrow CO_2 + H_2O$ C: 1 C: 1 H: 4 H: 2 O: 2 O: 3
5. Hint: From O or H, start from the one that features in less formulae on either side of equation. • In this case, H is balanced next. • Add coefficient on right hand side to balance H. • Note that the number of oxygen atoms increased on the right-hand side.	$CH_4 + O_2 \rightarrow CO_2 + 2H_2O$ C: 1 C: 1 H: 4 H: 4 O: 2 O: 4
6. Note that only O is left unbalanced. Add coefficient on left hand side to balance O.	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ C: 1 C: 1 H: 4 H: 4 O: 4 O: 4

Balancing chemical equations

Jefferson Lab: Balancing Act!

Go-Lab: Balancing chemical equations

1. Balance the following chemical equations involving covalent substances.

a. $H_2 + I_2 \rightarrow HI$

b. $CO_2 + C \rightarrow CO$

c. $NO_2 \rightarrow N_2O_4$

d. $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$

e. $NH_3 + O_2 \rightarrow N_2 + H_2O$

f. $C_2H_5OH + O_2 \rightarrow CO_2 + H_2O$

g. $H_2S + SO_2 \rightarrow S + H_2O$

2. Write balanced chemical equations for the following word equations:

a. hydrogen + bromine → hydrogen bromide

b. carbon + oxygen → carbon dioxide

c. carbon + oxygen → carbon monoxide

d. sulfur + oxygen → sulfur dioxide

a. sulfur dioxide + oxygen → sulfur trioxide

b. nitrogen + oxygen → nitrogen monoxide

c. nitrogen monoxide + oxygen → nitrogen dioxide

d. nitrogen + hydrogen → ammonia

e. methane + oxygen → carbon dioxide + water

f. methane + oxygen → carbon monoxide + water

3. For the following reaction description, first write a word equation and then write a balanced equation including state symbols.

a. Hydrogen gas is reacted with chlorine gas to form hydrogen chloride gas.

b. Ethane gas (C_2H_6) is reacted with oxygen to form carbon dioxide gas and water vapour.

c. Carbon monoxide gas is oxidised (reacts with oxygen) to form carbon dioxide gas.

d. Hydrogen peroxide (H_2O_2) decomposing to form water and oxygen gas.

e. Ammonia gas decomposing into nitrogen and hydrogen gases.

Go-Lab: Reactants, products, and leftovers

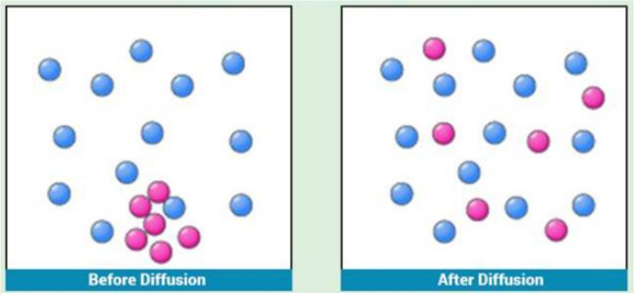
2.7 Diffusion and density of gases

By the end of this section, you should be able to:

- Explain that gases have different diffusion rates depending on their atomic or molecular mass. (2.2j)
- Explain why gases have different densities when measured under the same conditions of temperature and pressure. (2.3j)

2.7.1 Diffusion

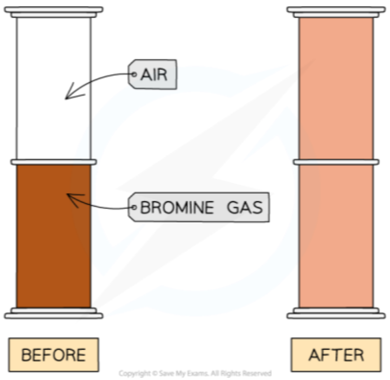
Have you ever opened a bottle of perfume? You can quickly smell the perfume all around the room. This happens because the particles of perfume leave the bottle and spread out through the air in the room. This phenomenon is known as diffusion.

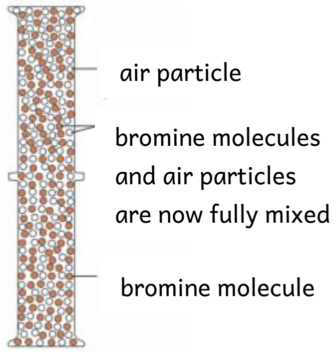


Source: byjus.com

Diffusion happens when the molecules move from a high concentration to a low concentration until all of the space is evenly occupied.

In the following experiment a few drops of bromine liquid are added to a gas jar. An inverted gas jar is quickly placed over the gas jar as shown in the diagram below. The liquid bromine quickly vaporises and the brown gas rises filling up both gas jars.



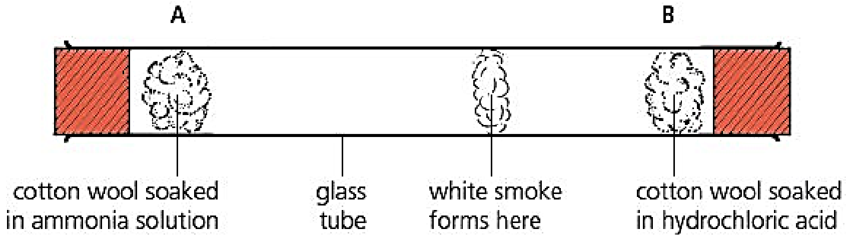


All gases diffuse to fill the space available. This happens because gas particles are in a continuous state of random motion. Bromine molecules mix among the air particles, and they spread from where they are more concentrated to where they are less concentrated until evenly mixed.

Diffusion of bromine through air.

Gases also diffuse at different rates. The rate at which molecules diffuse depends on factors such as temperature, pressure, and molecular mass. In the next experiment, the temperature and pressure are common to both gases under investigation and the only variable is the molecular mass of the gases concerned. This experiment is carried out in a fume cupboard.

A cotton wool is soaked in concentrated ammonia solution. This releases ammonia gas, NH_3. Another cotton wool is soaked in concentrated hydrochloric acid. This releases hydrogen chloride gas, HCl. Both cotton wools are inserted at opposite ends of the glass tube at the same time. After some time, a white ring forms which is closer to the cotton wool releasing hydrogen chloride gas. Why do you think this happens?



Both gases diffuse but their diffusion rate is not equal. This is due to their molecular mass.

RMM of NH_3 = 14 + (1 x 3) = 17 RMM of HCl = 1 + 35.5 = 36.5

Ammonia has a lower relative molecular mass than hydrogen chloride, so it travels faster since it is lighter. Hydrogen chloride is heavier and diffuses at a slower rate.

The **lower** the relative molecular mass of the gas **the faster** the gas will diffuse.

When these two gases meet, they form a white ring of ammonium chloride, NH_4Cl.

ammonia (g) + hydrogen chloride (g) → ammonium chloride (s)

$$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$$

If this experiment was carried out at a higher temperature, the particles gain heat energy, and they move faster. They collide with more energy and diffuse faster so the white ring of ammonium chloride will form faster at the same place, that is closer to the cotton wool soaked with hydrochloric acid.

Diffusion of gases

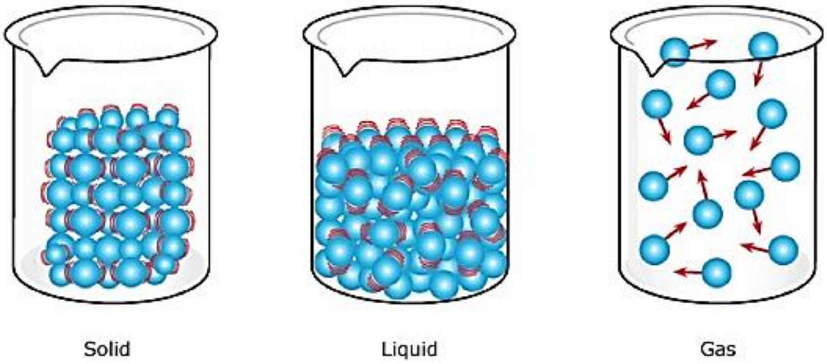
Go-Lab: Diffusion and temperature

Go-Lab: Diffusion and molecular mass

2.7.2 Density of gases

Gases, like liquids and solids, have density since they have mass and occupy space. Density is calculated using the equation:

$$\textbf{density} = \frac{\textbf{mass}}{\textbf{volume}}$$



The following table describes the differences between solids, liquids and gases.

Solids	Liquids	Gases
▪ Particles are closely packed in a regular structure. ▪ There are strong forces of attraction between particles and so particles can only vibrate about fixed positions. ▪ Due to this close packing solids have a high density compared to the liquid/gas of the same substance.	▪ Particles are not packed in regular arrangement due to weaker forces of attraction between particles. ▪ The particles are close, but they can move around each other. ▪ Liquids have a lower density compared to solids of the same substance.	▪ In gases the spaces between particles are huge. ▪ They can move at high speed in random manner since there are virtually no forces of attraction. ▪ Gases have the lower density compared to liquids and solids.

1. Work out the relative molecular mass of each gas and then place the following molecules in order of increasing density: SF_6, HCl, Cl_2, F_2, H_2, and N_2.

2. The average molecular mass of air is 28.97 amu. Which of the gases in question 1 are lighter or heavier than air?

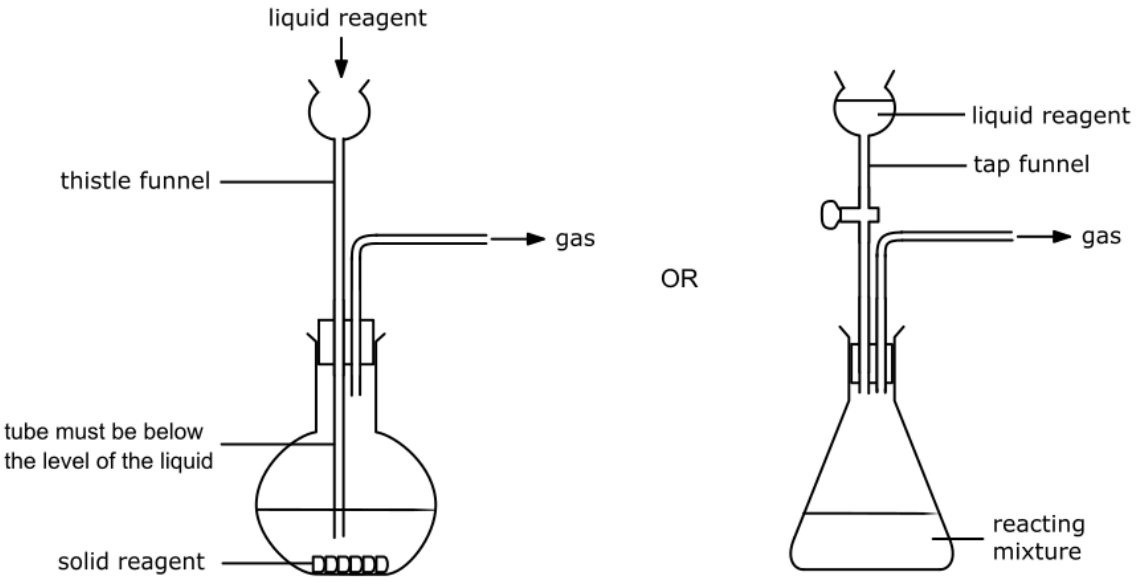
2.8 Preparing, collecting, and testing for gases

By the end of this section, you should be able to:

- Prepare gases safely. *(Limited to carbon dioxide by reacting acid with carbonates, oxygen from hydrogen peroxide, and hydrogen by reacting an acid with an appropriate metal.)* (2.2k)
- Prepare gases safely by selecting and assembling appropriate apparatus. *(Limited to carbon dioxide by reacting acid with carbonates, oxygen from hydrogen peroxide, and hydrogen by reacting an acid with an appropriate metal.)* (2.3k)
- Test the properties of gases following step by step instructions. *(Limited to carbon dioxide, hydrogen, and oxygen)* (2.2l)
- Test the properties of gases. *(Limited to carbon dioxide, hydrogen, and oxygen.)* (2.3l)
- Collect gases over water or in a gas syringe. (Limited to carbon dioxide, oxygen, and hydrogen.) (2.2m)
- Collect gases by upward or downward delivery. *(Limited to carbon dioxide, oxygen, and hydrogen. Reference to drying of gases is not required.)* (2.3m)

2.8.1 Preparing gases

Some chemical reactions produce a gas as a product of the reaction. The diagrams below show simple gas generator systems that are commonly used in laboratories to prepare the gas.



Two reagents are mixed in a conical, or in a flat-bottomed flask when the reaction takes place at room temperature. A round bottom flask is used when the reagents need to be heated to avoid breakages.

A thistle funnel or a tap funnel allows the liquid reagent to fall onto the other reagent in a controlled manner. The **thistle funnel** must dip below the level of reacting liquid to prevent escape of gas. The tap of the **tap funnel** must be closed upon adding the liquid reagent to prevent escape of gas.

The gas produced then exits through a delivery tube. The gas is then collected using further apparatus that is added to the reaction vessel as described in the next section. For toxic gases, the experiment setup needs to be in a fume cupboard.

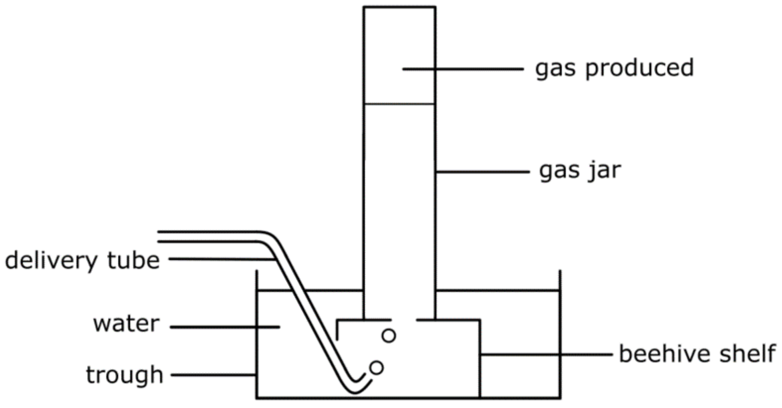
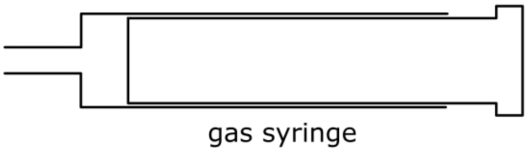
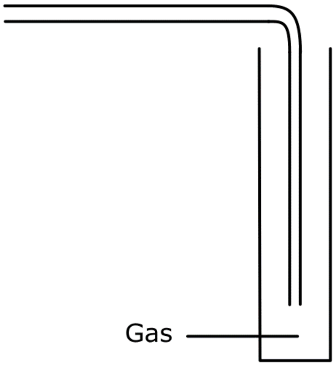
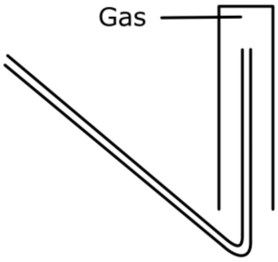
2.8.2 Collecting gases

Gases need to be collected safely so that it could be observed and tested as needed.

The method of collection of gases depends on the properties of the gas that is

- on the solubility of gas in water and
- on the density of the gas compared to air.

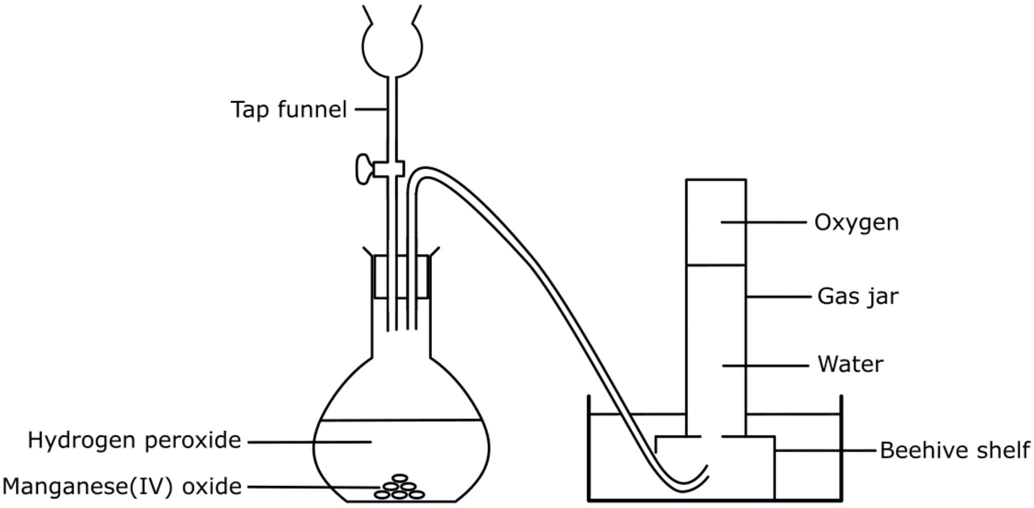
The table below shows the apparatus used to collect gases in the laboratory.

Setup used to collect gases	Description
	**Over water** Used to collect gases which are insoluble in water. E.g., hydrogen, and gases which are **relatively insoluble in water** such as oxygen.
	**Gas syringe** Used to collect any type of gas. Useful when the volume of gas needs to be measured especially during experiments where the rate of reaction needs to be measured.
	**Downward delivery** Used for gases which are **denser than air** e.g., carbon dioxide.
	**Upward delivery** Used for gases which are **less dense than air** e.g., hydrogen

2.8.2.1 Laboratory preparation of oxygen

Oxygen gas is prepared by the **catalytic decomposition** of hydrogen peroxide (H_2O_2).

Catalytic means that the reaction takes place in the presence of a catalyst. A **catalyst is a substance which speeds up a chemical reaction and is unchanged at the end of the process**. Decomposition means that one compound breaks down into two or more products during the reaction.



$$\text{hydrogen peroxide} \xrightarrow{\text{manganese(IV) oxide}} \text{water} + \text{oxygen}$$

$$2H_2O_2(l) \xrightarrow{MnO_2} 2H_2O(l) + O_2(g)$$

Hydrogen peroxide solution is added to manganese(IV) oxide, which catalyses the decomposition of the hydrogen peroxide. **Effervescence** (a lot of bubbles) is observed in the flat-bottomed flask. This reaction gives off heat and it is known as an **exothermic** reaction. In this case the oxygen produced is collected over water because oxygen is slightly soluble in water. This reaction takes place at room temperature.

Hydrogen peroxide solution decomposes very slowly on its own. This reaction is faster in the presence of light; therefore, the solution is stored in dark brown bottles. Catalase which is an enzyme present in fresh liver and uncooked potatoes, also catalyses this reaction.

1. Name another method by which oxygen can be collected.
2. Explain why it can be collected in this way.

Laboratory preparation of oxygen.

C. W. Scheele and oxygen

Oxygen was discovered by Swedish chemist, Carl Wilhelm Scheele in 1772. Scheele, unfortunately, was beaten to the punch by Joseph Priestly, an English chemist who, although discovered the oxygen molecule in 1774, published his findings three years before Scheele's. Scheele is still recognized as being the first man to even figure out what oxygen is.

2.8.2.2 Laboratory preparation of hydrogen

Hydrogen gas can be prepared from the reaction of a metal with an acid according to the following equation.

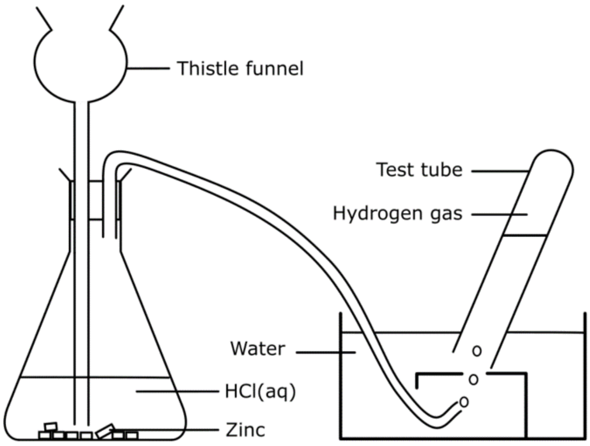
metal + acid → salt + hydrogen

Hydrochloric acid is commonly used as an acid in the laboratory preparation of hydrogen.

Most metals can be used to prepare hydrogen in the lab. Magnesium and zinc are usually used since they produce hydrogen at an acceptable rate which is neither too slow nor too fast.

Sodium and potassium metals cannot be used to prepare hydrogen in the lab because their reaction would be dangerous as these metals are highly reactive and will cause injury due to their explosive reaction with acids.

Copper, silver, and gold metals cannot be used as these metals do not react with acids.



Some pieces of zinc are placed in a conical flask. Dilute hydrochloric acid is added by means of a thistle funnel. There is effervescence and the gas produced is collected over water. Hydrogen is collected in small amounts since it is a highly flammable gas.

The chemical reaction involved is:

zinc + hydrochloric acid → zinc chloride + hydrogen

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

1. Name another method by which hydrogen can be collected.
2. Explain why it can be collected in this way.

Laboratory preparation of hydrogen.

Robert Boyle and hydrogen

In 1671, Robert Boyle discovered and described the reaction between iron filings and dilute acids, which results in the production of hydrogen gas. In 1766, Henry Cavendish was the first to recognize hydrogen gas as a discrete substance, by naming the gas from a metal-acid reaction "inflammable air".

2.8.2.3 Laboratory preparation of carbon dioxide

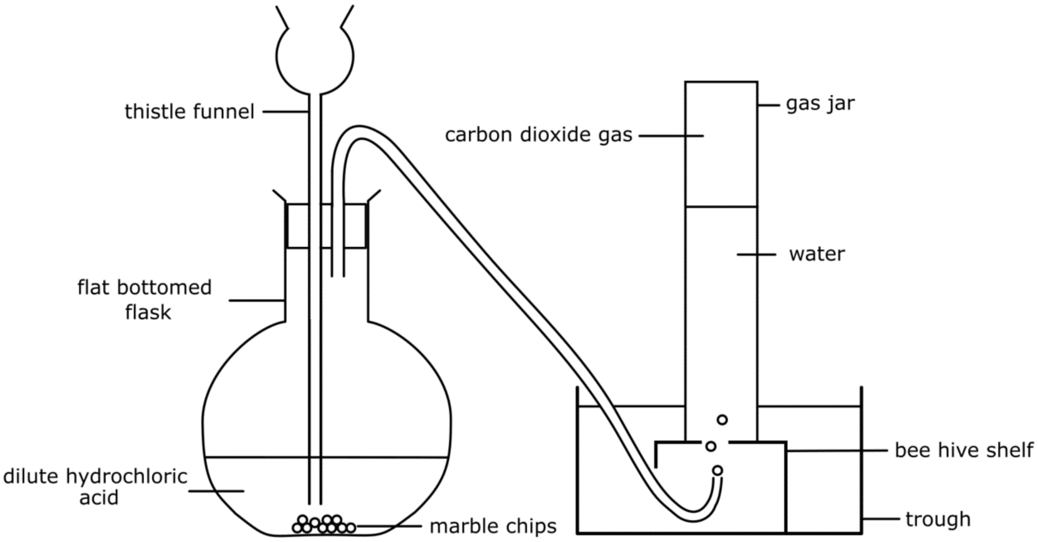
In the laboratory preparation of carbon dioxide, a metal carbonate and an acid can be used.

metal carbonate + acid → salt + water + carbon dioxide

Usually calcium carbonate, $CaCO_3$ and dilute hydrochloric acid, HCl are used. Calcium carbonate is the chemical name given to chalk, limestone, or marble. When calcium carbonate and dilute hydrochloric acid react together, they form calcium chloride, water, and carbon dioxide according to the equation below:

calcium carbonate +hydrochloric acid → calcium chloride + water + carbon dioxide

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



Since carbon dioxide is slightly soluble in water, the actual volume of carbon dioxide collected is less than the expected amount.

1. Name another method by which carbon dioxide can be collected.
2. Explain why it can be collected in this way.

Laboratory preparation of carbon dioxide.

Joseph Black and carbon dioxide.

The discovery of carbon dioxide by Joseph Black marked a new era of research on the respiratory gases. When he studied magnesium carbonate, he found that when this was heated or exposed to acid, a gas was evolved that he called "fixed air" because it had been combined with a solid material. He showed that the new gas extinguished a flame, that it could not support life, and that it was present in gas exhaled from the lung.

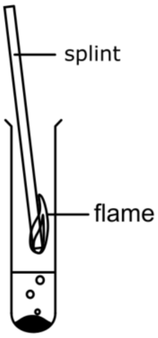
2.8.3 Testing for gases

The experiment which leads to the identification of an element or compound is called a **test**. Below are the tests for oxygen, hydrogen, and carbon dioxide.

2.8.3.1 Oxygen

In this test, a wooden splint is lit, allowed to burn for a few seconds, then blown out by mouth or by shaking. Whilst the tip is still glowing hot, the splint is introduced to the gas sample that has been trapped in a test tube.

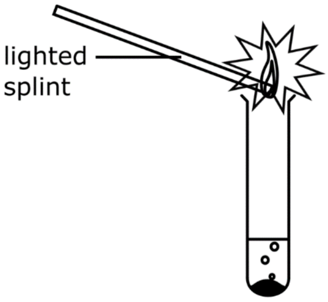
If the glowing splint relights, this indicates that the gas present is oxygen.



2.8.3.2 Hydrogen

A splint is lit and while it still burns it is held near the opening of a test tube containing a gas suspected to be hydrogen, then the stopper is removed to expose the burning splint to the gas.

If the gas is hydrogen, the mixture burns with a distinctive 'squeaky pop'.

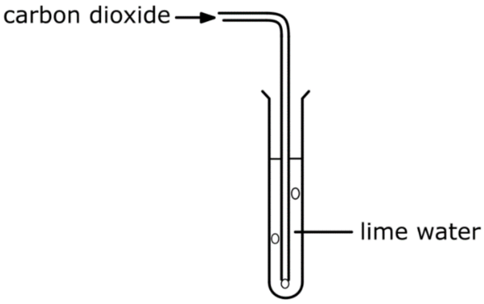


BBC article by Fergus Nicoll: Behind the wheel of a hydrogen-powered car.

2.8.3.3 Carbon dioxide

Carbon dioxide gas is bubbled through limewater, which is a solution of calcium hydroxide. If the limewater turns from colourless to milky, this indicates that this gas is indeed carbon dioxide.

The milky suspension is due to the formation of calcium carbonate which is insoluble in water as seen in the equation below:



carbon dioxide + calcium hydroxide → calcium carbonate + water

$$CO_2(g) + Ca(OH)_2(aq) \rightarrow CaCO_3(s) + H_2O(l)$$

Testing for carbon dioxide.

2.9 Atmospheric pollution

By the end of this section, you should be able to:

- Relate the emission of the pollutants present in air to human activities. *(Limited to carbon dioxide, carbon monoxide and soot.)* (2.1o)
- Describe how the amount of certain gases and particulates in the environment may increase due to combustion reactions. *(E.g. carbon dioxide due to complete combustion, carbon monoxide and soot due to incomplete combustion.)* (2.2o)
- Explain how the amount of certain gases and particulates in the environment may increase due to combustion reactions and natural causes. *(E.g. carbon dioxide, carbon monoxide, sulfur dioxide, nitrogen oxides and soot.)* (2.3o)
- Identify gases that contribute towards the greenhouse effect, ozone depletion and acid rain. *(Greenhouse gases: E.g. CO_2, CH_4 and water vapour. Ozone depletion: CFCs. Acid rain: E.g. SO_2 and NO_2.)* (2.1r)
- Explain environmental effects of pollutants. *(Such as greenhouse gases, CFCs, SO_2, NO_2 and particulates which include smog, soot, dust, and volcanic ash.)* (2.2r)
- Interpret data regarding environmental effects of some pollutants. *(Such as global warming, acid rain, effect of CFCs on ozone and particulates which include smog, soot, dust, and volcanic ash.)* (2.3r)
- Identify methods for reducing emission of pollutants into the atmosphere. *(E.g. use of renewable sources of energy.)* (2.1s)
- Describe methods for reducing emission of pollutants into the atmosphere. *(E.g. use of renewable sources of energy, banning or reduction of pollutants, better choice of non-renewable fuels.)* (2.2s)
- Discuss methods for reducing emission of pollutants into the atmosphere. *(E.g. use of renewable sources of energy, catalytic converters, and better choice of non-renewable fuels.)* (2.3s)

A substance present in air that can cause harm to humans and the environment is known as an air pollutant. Pollutants can be in the form of solid particles, liquid droplets, or gases. In addition, they may be natural or made by humans.

2.9.1 Greenhouse gases

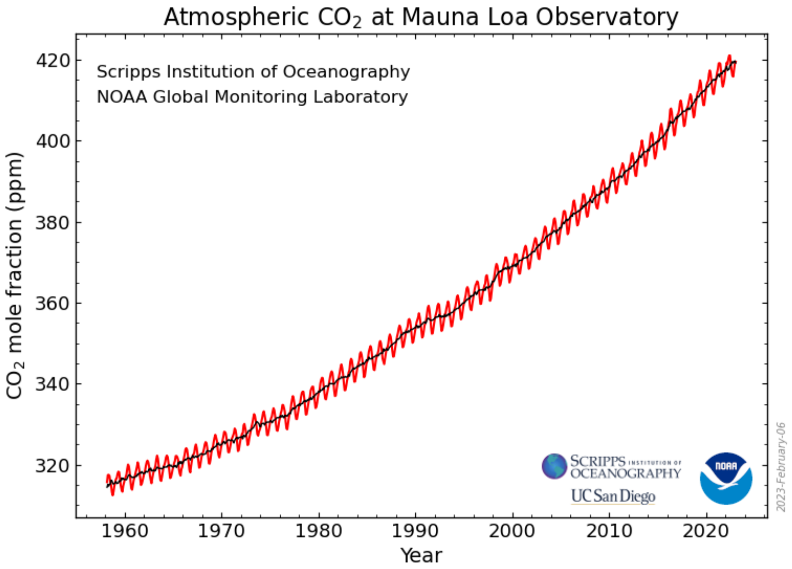
A **greenhouse gas** is one that absorbs heat in the atmosphere and prevents heat from escaping into space, keeping the Earth warmer.

Carbon dioxide (CO_2) is a colourless, odourless, non-toxic greenhouse gas. It is produced by many natural processes such as volcanic eruptions, respiration of animals and burning or decay of organic matter. Since the industrial revolution, the amount of CO_2 in the air has been increasing at an alarming rate due to the burning of fossil fuels. Fossil fuels, like oil and natural gas, which are primarily made of carbon and hydrogen react completely with oxygen in air during combustion to produce carbon dioxide and water vapour. This combustion reaction can be represented as follows:

$$\text{hydrocarbon} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water vapour}$$

The industrial production of cement is also a main contributor to the emission of carbon dioxide.

Scientists regularly measured the amount of CO_2 since 1958 in Hawaii as shown by the graph below. Analysis of the Mauna Loa record shows that the mean annual concentration of CO_2 has been increasing since then. The concentration of CO_2 in the atmosphere fluctuates between winter and summer because of seasonal variations in photosynthesis. Every year an extra 5000 million tonnes of CO_2 are produced into the air from burning fossil fuels. At the same time, tropical rainforests which absorb CO_2 are being destroyed. The oceans can dissolve some of the extra CO_2 but not all. The rest stays in the atmosphere, upsetting nature's balance.

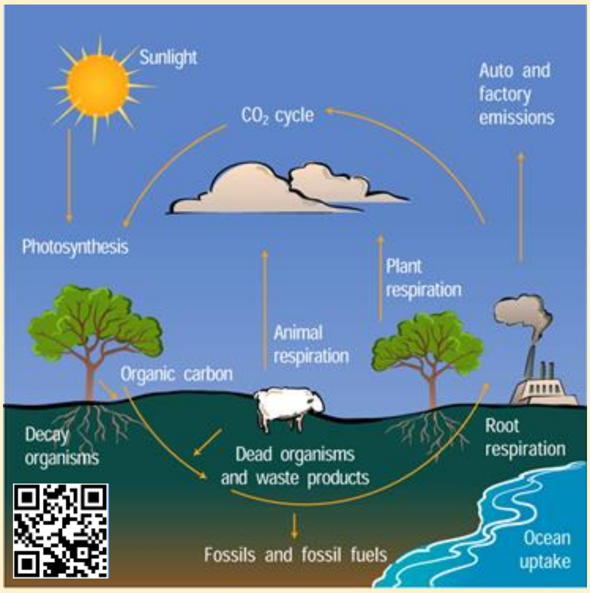


Other substances that are present in the atmosphere which are also greenhouse gases include water vapour, methane (CH_4), and CFCs. The level of CH_4 is rising due to an increase in animal farming, rice farming and landfill sites.

The carbon cycle

Carbon dioxide forms part of the carbon cycle. "The carbon cycle is the biogeochemical cycle by which carbon is exchanged among the biosphere, pedosphere, geosphere, hydrosphere, and atmosphere of the Earth. Carbon is the main component of biological compounds as well as a major component of many minerals such as limestone.

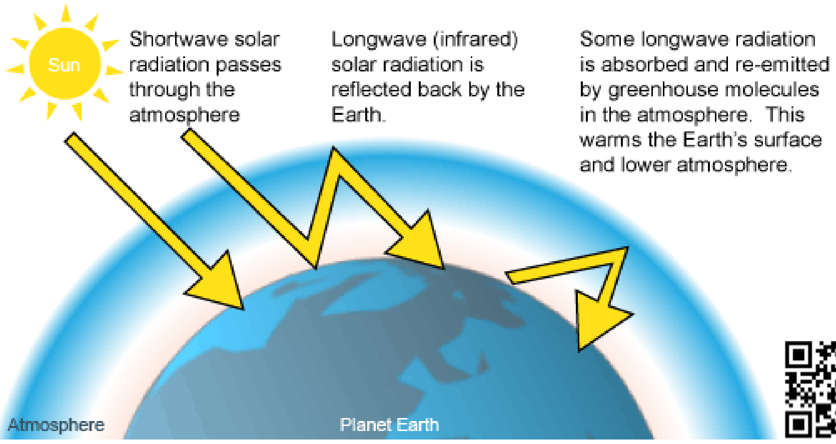
It describes the movement of carbon as it is recycled and reused throughout the biosphere, as well as long-term processes of carbon sequestration and release from carbon sinks.



The carbon cycle was discovered by Joseph Priestley and Antoine Lavoisier, and popularized by Humphry Davy."

2.9.1.1 The Greenhouse Effect and global warming

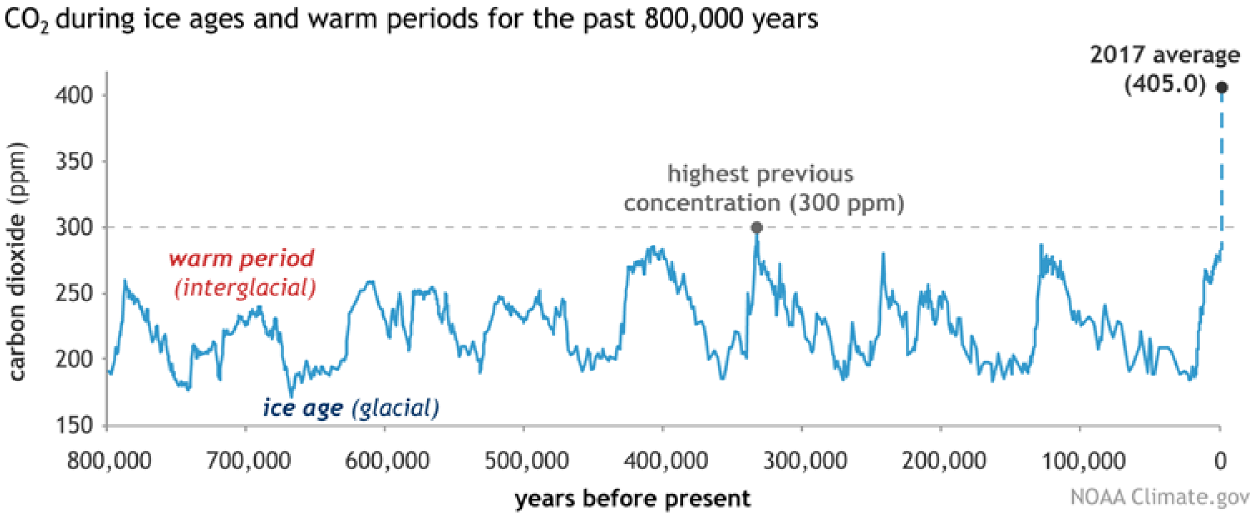
Without greenhouse gases our planet would be covered in ice with an average temperature 35 °C lower than at present. In fact, the temperature of our Earth has more to do with the greenhouse effect than with our proximity to the Sun!



Radiant energy from the Sun falls onto the Earth as short infra-red waves and warms it. The Earth radiates heat energy back into space as longer infra-red radiation. Unlike sunlight, infra-red radiation cannot travel freely through the air surrounding the Earth. Water vapour, carbon dioxide among other gases, absorb and re-emit some of the infra-red radiation. Since carbon dioxide and water vapour act like a greenhouse, their warming effect is called the greenhouse effect.

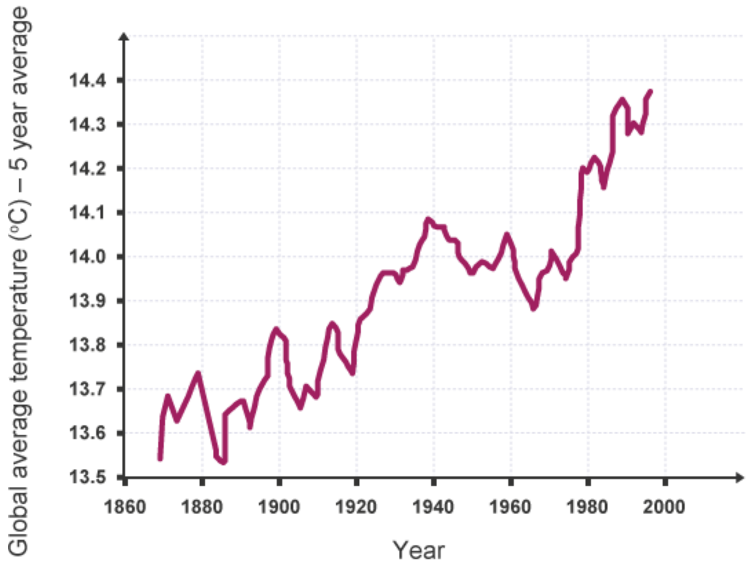
The greenhouse effect

The increase in greenhouse gases like CO_2 causes global warming. The graph below shows how the amount of CO_2 in the atmosphere has increased dramatically which leads to a greater greenhouse effect and consequently, to increased global warming.



NOAA Climate.gov
Data: NCEI

According to an article found at NASA's Earth Observatory, the surface of Earth has warmed by a little more than 1 °C since 1880. Two thirds of this increase happened since 1975.



Increasing global temperature causes:

1. Polar ice caps to melt and sea levels to rise.
2. Low-lying areas of land would submerge under the sea.
3. A change in climate and weather patterns all over the world. Extreme weather events such as hurricanes and cyclones would be more frequent. Tropical areas are experiencing more frequent and destructive storms and floods.
4. Some agricultural areas will cease to produce crops and become deserts due to climate change.
5. Some species that cannot adapt with a change in climate will die out.

Go-Lab: Sunlight, Infrared, $CO_{2,}$ and the Ground

2.9.1.2 Lowering the level of CO_2 – Combating climate change

Carbon dioxide emissions must be reduced by 70 - 80% in order to stabilise atmospheric carbon dioxide concentration and stop the increase in global temperature. This can be done by:

- Less consumption of fossil fuels.
- Using renewable sources of energy.
- Reduce reliance on cars.
- More energy efficient buildings.
- Improved power plant efficiency.
- Planting more trees.
- The Kyoto Protocol (1997) is the world's primary international agreement on combating global warming. 100 countries committed to reduce their emissions of CO_2 and five other greenhouse gases.
- The UN climate change conference to discuss the way forward on many issues including curbing CO_2 production.

2.9.2 Carbon monoxide

Carbon monoxide (CO) is a colourless, odourless, non-irritating but very poisonous gas. It is a product of incomplete combustion of fuel that is when natural gas, coal, or wood burns in a limited supply of air.

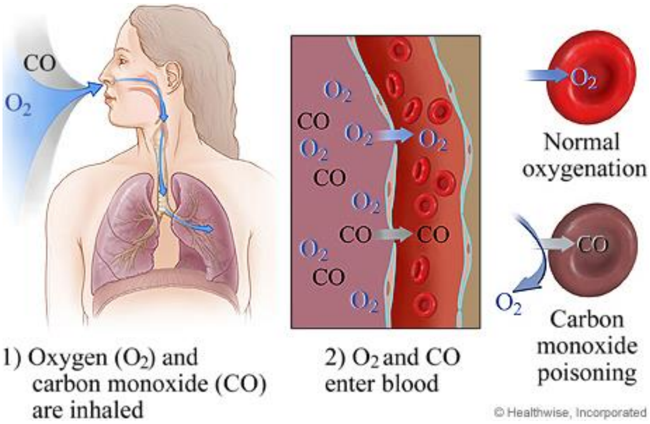
About 350 million tonnes are emitted into the atmosphere every year. Vehicular exhaust is a major source of carbon monoxide. No matter how efficient a car is (both petrol and diesel) at burning fuel, they all produce an amount of carbon monoxide. When cars are subjected to a vehicular road-worthiness test (VRT), one of the tests is to check that the engine is producing CO at an acceptable level. CO is produced from cigarette smoking as well.

For decades, canaries were used in coal mining as a warning system to alert miners of dangerous levels of carbon monoxide. These little birds sing most of the time. As long as the canary was singing, the miners knew it was safe to work. If the bird was silenced, killed because of the presence of poisonous gas, the miners would evacuate the mine until it could be made safer. Nowadays, digital carbon monoxide sensors are used instead.

2.9.2.1 Effect of Humans and the Environment

Carbon monoxide readily combines with haemoglobin in the red blood cells to form carboxyhaemoglobin, preventing the uptake of oxygen. The problem with CO is that it combines permanently with the red blood cells. Carbon monoxide does not allow for the usual transfer of oxygen and carbon dioxide and so we say that CO wastes blood.

A shortage of oxygen causes headache and dizziness and makes a person feel sluggish. If the level of carbon monoxide reaches 0.1% of the air, it causes death. Carbon monoxide poisoning and deaths have occurred in Malta many times. Always stay on the lookout not to allow burning fuels in a confined space that is not well ventilated.



2.9.2.2 Removal of carbon monoxide

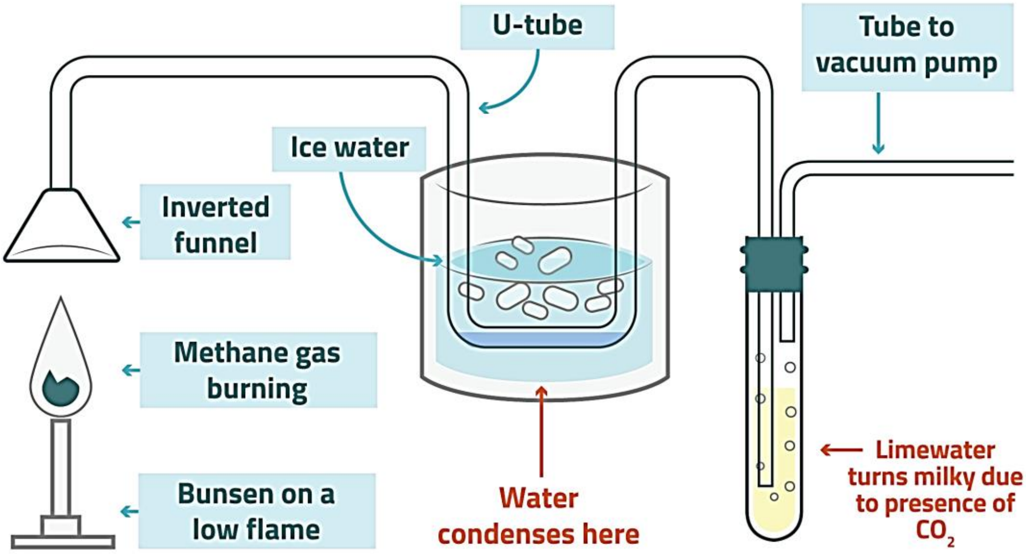
- Catalytic converters are fitted to the exhausts of many cars. The catalysts will oxidize carbon monoxide in the exhaust gases to carbon dioxide.
- New fuels may be used in the future. Some fuels e.g., alcohol burn more cleanly than hydrocarbons.
- Vehicle engines can be tuned to take in more air and produce only CO_2 and H_2O. Unfortunately, this increases the formation of oxides of nitrogen.

2.9.3 Complete and incomplete combustion

A **combustion reaction** is a chemical reaction in which a substance reacts rapidly with oxygen with the production of heat and light. The products of combustion of a hydrocarbon will depend on whether the combustion reaction takes place in a plentiful supply or limited supply of oxygen.

2.9.3.1 Complete Combustion

Fossil fuels are compounds made up of hydrogen and carbon, known as **hydrocarbons**. Methane (natural gas), petrol, diesel, fuel oil and wax are all hydrocarbons. In the following experiment you will be determining which products are given as methane is burning.



Source: Combustion: Learning about Air Quality (learnaboutair.com)

Methane is lit, and the air pump is switched on. The following observations are noted:

- A colourless liquid condenses in the U-tube. This can be tested using anhydrous copper(II) sulfate which turns from white to blue.
- Lime water turns milky (white precipitate) showing that carbon dioxide is produced.

Therefore:

hydrocarbon + oxygen → water vapour + carbon dioxide

When methane burns in a plentiful supply of oxygen

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

When the above products **only** are produced, the methane or other hydrocarbons undergo **complete combustion**.

Fuels undergo **complete combustion** when there is a plentiful supply of oxygen.

Fuels that burn with a clean blue flame, give the most heat energy possible.

Since, both carbon dioxide and water vapour are also present in the air, it would be a good idea to repeat this experiment without the fuel burning. It will take much longer for the indicators to show a change.

2.9.3.2 Incomplete Combustion

When a fuel does **not burn in a sufficient amount of oxygen**, it burns with a yellow flame, producing unburnt particulate carbon called **soot** and a colourless toxic gas called **carbon monoxide.** The fuel undergoes **incomplete combustion.**

E.g. methane burning in an insufficient amount of air.

$$2CH_4(g) + 3O_2(g) \rightarrow 2CO(g) + 4H_2O(g)$$

2.9.3.3 Comparison between CO_2 and CO, complete and incomplete combustion

Complete vs incomplete combustion

	Complete combustion	Incomplete Combustion
Amount of oxygen available	plentiful supply of oxygen	limited amount of oxygen
Colour of flame	clean blue flame	yellow sooty flame
Products evolved	carbon dioxide, water vapour	carbon monoxide, water vapour
Amount of energy released	most heat energy released	less heat energy released

Properties of carbon dioxide vs carbon monoxide

	Carbon dioxide	Carbon monoxide
Appearance	colourless	colourless
Odour	odourless	odourless
Effect on damp blue litmus	it turns damp blue litmus red (acidic oxide)	no effect on litmus (neutral oxide)
Bubbling through lime water	it turns lime water milky	no reaction with lime water
Bubbling through sodium hydroxide solution	reacts with sodium hydroxide (neutralisation reaction)	no reaction
Reducing agent	no	yes
Combustion	does nor burn, used to extinguish fires	burns with a bright blue flame to give CO_2
Toxicity	toxic at high concentrations	toxic at very low concentrations

Note that bubbling carbon dioxide and carbon monoxide through calcium hydroxide solution is the test used to distinguish between samples of the two gases.

Periodic graphics: The chemistry of wildfires

2.9.4 Particulates

Particulates, alternatively referred to as particulate matter (PM) or fine particles, are tiny particles of solid or liquid suspended in a gas. In contrast, aerosol refers to particles and the gas together.

Sources of particulate matter can be man-made or natural. Some particulates occur naturally, originating from volcanoes, dust storms, forest, and grassland fires, living vegetation, and sea spray. Human activities, such as the burning of fossil fuels in vehicles, power plants and various industrial processes such as mining also generate significant amounts of aerosols.

Averaged over the globe, aerosols made by humans, currently account for about 10% of the total amount of aerosols in our atmosphere. Increased levels of fine particles in the air are linked to health hazards such as heart disease, altered lung function and lung cancer.

2.9.4.1 Effect on humans and the environment

- Smoke increases the dangers of smog.
- Solid particles fall as grime on people, clothing, buildings, and plants.
- Particles darken the city by scattering light. Sunlight, which meets dust particles, is reflected back into space, and prevented from reaching the Earth. Some scientists believe that a fourfold increase in the amount of dust would make the Earth’s temperature fall by about 3 °C. This is known as global dimming. It would affect food production.
- Some particulates are also carcinogenic (cause cancer) when inhaled.

2.9.4.2 Removal of particulate matter; smoke, dust, and grit

Industries use a number of methods which include:

- Using sprays of water to wash out particles from their waste gases.
- Passing waste gases through filters.
- Electrostatic precipitators to remove dust particles from waste gases by electrostatic precipitation.

Smog

“Smog is a type of severe air pollution. The word "smog" was coined in the early 20th century and is a contraction of the words smoke and fog to refer to smoky fog, its opacity, and odour. The word was then intended to refer to what was sometimes known as pea soup fog, a serious problem in London from the 19th century to the mid-20th century.

This kind of visible air pollution is composed of nitrogen oxides, sulfur oxides, ozone, smoke, and other particulates. Smog is derived from coal combustion emissions, vehicular emissions, industrial emissions, forest and agricultural fires and photochemical reactions of these emissions.”

2.9.5 Sulfur dioxide

Sulfur dioxide is a chemical compound with the formula SO_2. It is produced by volcanoes and in various industrial processes. It is estimated that worldwide, over 150 million tonnes of SO_2 are emitted per year by burning coal and oil. Half of the output comes from burning of coal in power stations. All coal contains between 0.5 to 5% of sulfur. During combustion, sulfur is oxidised to sulfur dioxide as the fuel burns.

$$\text{sulfur} + \text{oxygen} \rightarrow \text{sulfur dioxide}$$

$$S(s) + O_2(g) \rightarrow SO_2(g)$$

Industrial smelters which produce metals from sulfides also release tonnes of SO_2 daily.

2.9.5.1 Effect of SO_2 on humans and the environment

Sulfur dioxide is a colourless gas with an irritating smell. Inhaling sulfur dioxide causes coughing, chest pains and shortness of breath. It is poisonous; at a level of 0.5% it will kill. Sulfur dioxide is thought to be one of the causes of bronchitis and lung diseases.

Sulfur dioxide emitted from industries and exhaust of motor vehicles, is carried away by air currents. It can form sulfurous and sulfuric acid. Eventually it falls as acid rain or acid snow, which may turn up hundreds of miles away from the source of the pollution.

$$\text{sulfur dioxide} + \text{water} \rightarrow \text{sulfurous acid}$$

$$SO_2(g) + H_2O(g) \rightarrow H_2SO_3(aq)$$

Sulfurous acid oxidises by reacting with oxygen in the air.

$$\text{sulfurous acid} + \text{oxygen} \rightarrow \text{sulfuric acid}$$

$$2H_2SO_3(aq) + O_2(g) \rightarrow 2H_2SO_4(aq)$$

The above two equations can be summarised by the following equation.

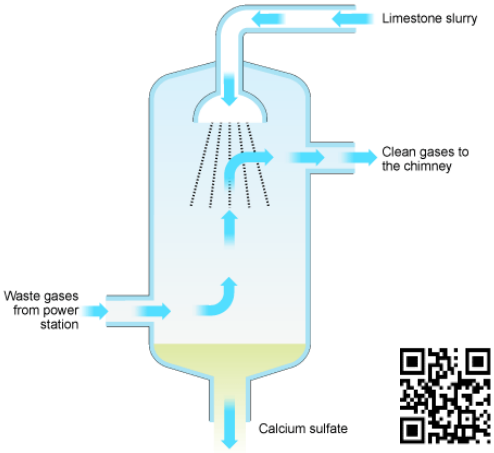
$$\text{sulfur dioxide} + \text{water} + \text{oxygen} \rightarrow \text{sulfuric acid}$$

$$2SO_2(g) + 2H_2O(l) + O_2(g) \rightarrow 2H_2SO_4(aq)$$

2.9.5.2 Lowering SO_2 levels

New power stations are fitted with a filter to stop sulfur dioxide escaping. Gases leaving the power station are passed through limestone ($CaCO_3$), which reacts with SO_2 to form calcium sulfate ($CaSO_4$). This process removes 95% of the acidic waste gases. Some of the filtering cost is recovered by selling calcium sulfate to make plaster for the building industry.

Low sulfur fuels are produced by crushing coal and washing it with a suitable solvent to reduce the sulfur content by 10 to 40%. The dirty solvent must be disposed of without creating pollution on land or in rivers.



2.9.6 Nitrogen oxides (NO_x)

Smog over Santiago in Chile.

Nitrogen monoxide (NO) is emitted from high temperature combustion such as in car engines and is also produced naturally during thunderstorms by electrical discharge. As soon as NO comes into contact with oxygen in the air, it oxidises to form **nitrogen dioxide** which is a dense brown gas. The picture to the right shows a brown haze above Santiago in Chile. This brown haze can sometimes also be observed if you are at Mdina looking at the densely populated part of Malta in the evening.

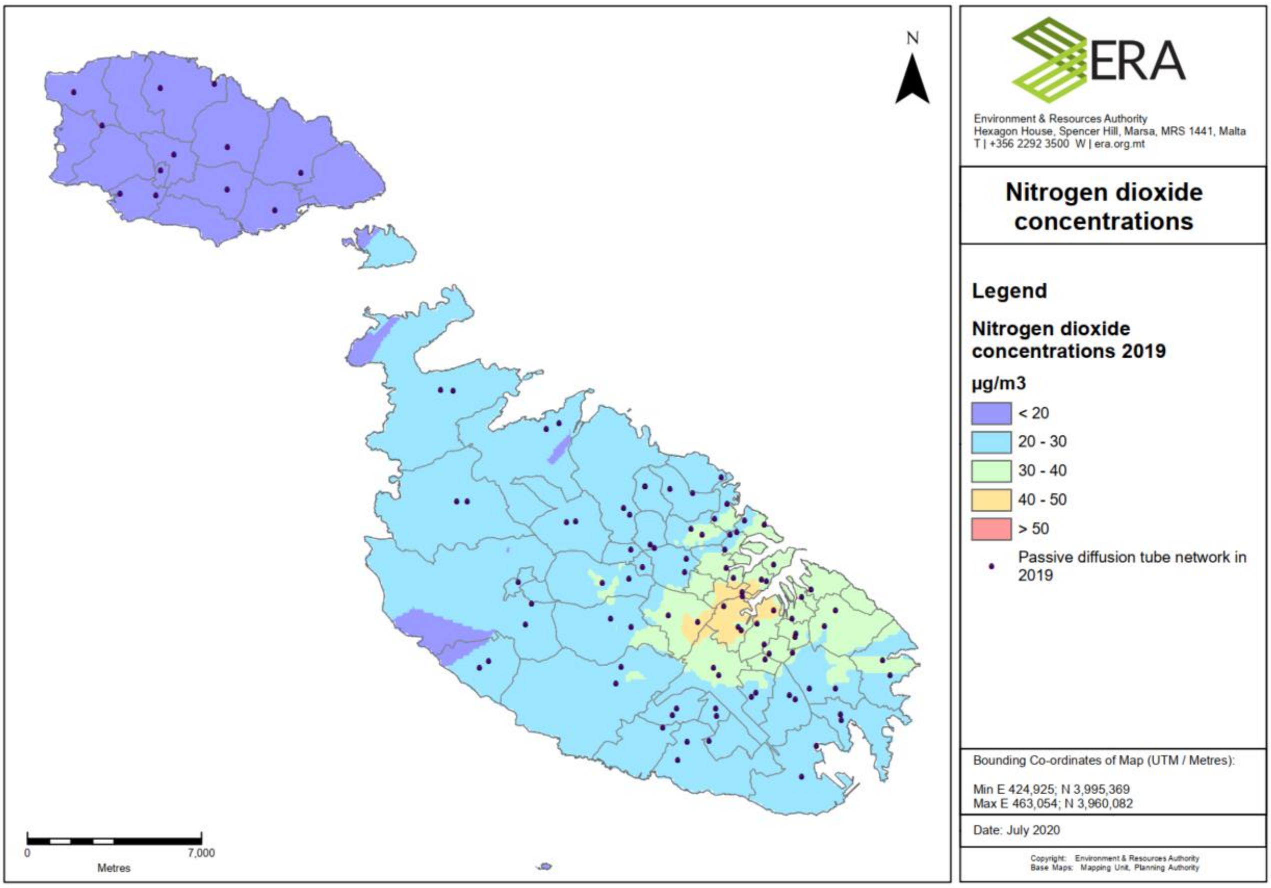
Nitrogen dioxide is a chemical compound with the formula NO_2. This reddish-brown toxic gas has a characteristic sharp, irritating odour. Nitrogen dioxide is one of the most prominent air pollutants. It is also one of the prominent gases that causes acid rain.

The release of oxides of nitrogen, known as **NO_x** (which includes NO, NO_2) is an indirect result of burning fossil fuels. During the burning of fuels in the internal combustion engine, air is drawn into the engine along with the fuel. The fuel is burned in oxygen, generating energy. Nitrogen is also present in the air. Combustion temperatures are high enough to make some nitrogen combine with oxygen. As a result, the gases nitrogen monoxide (NO) which is then oxidised to nitrogen dioxide (NO_2) are formed. These gases enter the air from exhaust of motor vehicles. They are also released from the chimneys of power stations and factories.

$$N_2(g) + O_2(g) \rightarrow 2NO(g)$$

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

Below is a chart of the distribution of NO_2 concentrations in Malta during 2019 according to data collected by ERA (Environment & Resources Authority).



2.9.6.1 Effect of NO_x on humans and the environment

Nitrogen monoxide (NO) is not a very dangerous gas and is pH neutral (neither acidic nor alkaline). However, it quickly reacts with air to form nitrogen dioxide (NO_2), which is highly toxic and irritates the breathing passages.

Nitrogen dioxide (NO_2) is acidic and reacts with oxygen and water to form a mixture of nitrous (HNO_2) and nitric acid (HNO_3), producing **acid rain**.

$$2NO_2(g) + H_2O(l) \rightarrow HNO_2(aq) + HNO_3(aq)$$

Nitrous acid (HNO_2) oxidises further by reacting with oxygen in the air to form nitric acid, HNO_3.

$$2HNO_2(aq) + O_2(g) \rightarrow 2HNO_3(aq)$$

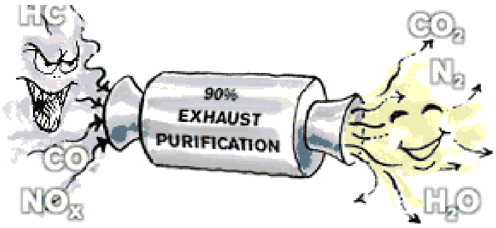
The above two equations can be summarised by the following equation.

$$\text{nitrogen dioxide} + \text{oxygen} + \text{water} \rightarrow \text{nitric acid}$$

$$4NO_2(g) + O_2(g) + 2H_2O(l) \rightarrow 4HNO_3(aq)$$

2.9.6.2 Lowering NO_x levels

Catalytic converters are fitted on petrol engines. A typical catalytic converter is made of platinum, rhodium and palladium supported on a honeycomb mesh. The hot exhaust gases pass over the catalytic surface and the harmful gases are converted into less harmful products. Catalytic converters are damaged by lead. They are only used in cars running on unleaded petrol.

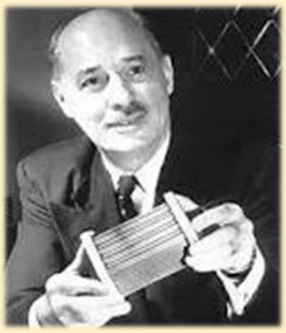


In the catalytic converter nitrogen monoxide reacts with carbon monoxide to form the non-toxic gases nitrogen and carbon dioxide.

$$2NO(g) + 2CO(g) \rightarrow N_2(g) + 2CO_2(g)$$

Oxidation catalysts are fitted on diesel engines to decrease the emissions of carbon monoxide and unburnt hydrocarbons. The oxidised products are carbon dioxide and water vapour.

History of catalytic converters



The catalytic converter was invented by Eugene Houdry, a French mechanical engineer and expert in catalytic oil refining who lived in the U.S. around 1950. When the results of early studies of smog in Los Angeles were published, Houdry became concerned about the role of smokestack exhaust and automobile exhaust in air pollution and founded a company, Oxy-Catalyst. Houdry first developed catalytic converters for smokestacks called cats for short.

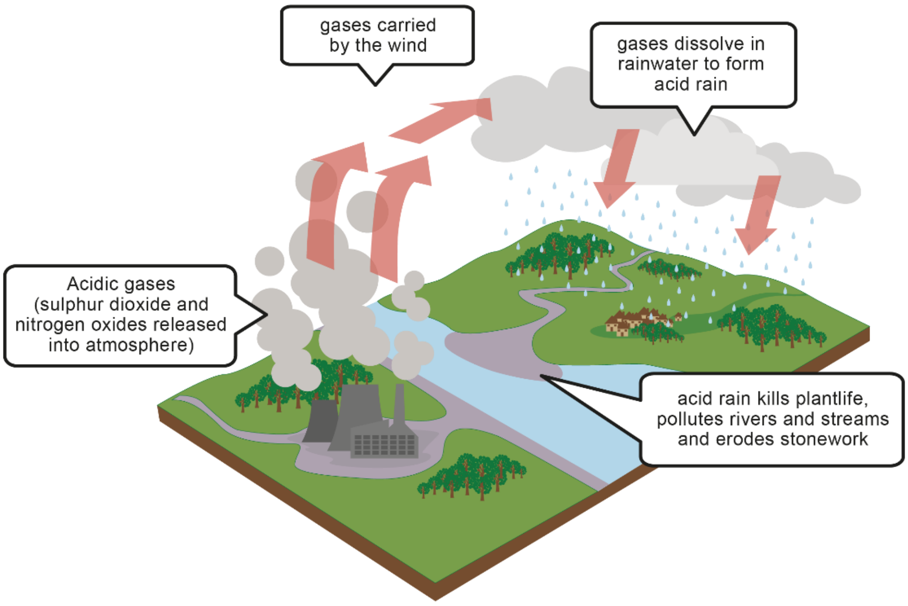
Then he developed catalytic converters for warehouse forklifts that used low grade non-leaded petrol. Then in the mid-1950s he began research to develop catalytic converters for petrol engines used on cars.

2.9.7 Acid rain

Rainwater is naturally weakly acidic with a pH of around 5.6, because carbon dioxide dissolves in water vapour to form carbonic acid, $H_2CO_3(aq)$. Rainwater is not considered as acid rain.

Acid rain contains stronger acids such as sulfuric and nitric acid.

Acid rain has a pH between 2.0 and 5.4.



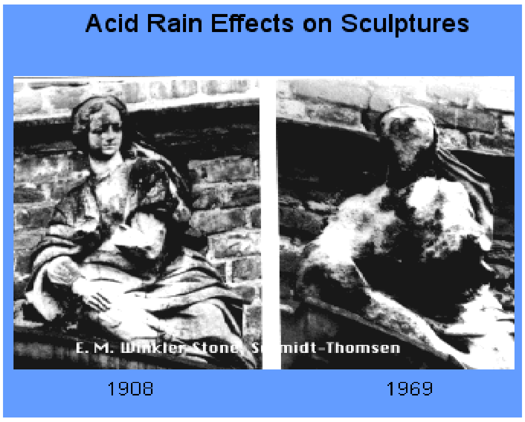
Source: http://data.allenai.org/tqa/effects_of_air_pollution_L_0048/

Acid rain that falls on land is absorbed by the soil. At first, the nitrates in the acid rain fertilise the soil and encourage the growth of plants. However, acid rain reacts with minerals, converting the metals in them into soluble salts. The rainwater containing these soluble salts of calcium, potassium, and aluminium and other metals trickle down through the soil into the subsoil where plant roots cannot reach them. In this way, salts are leaching out of the topsoil and crops are robbed of nutrients. One of the salts formed by acid rain is aluminium sulfate. This salt damages the roots of trees. The damaged roots are easily attacked by viruses and bacteria and the tree dies of a combination of malnutrition and disease.

The acidic rainwater trickles through the soil until it meets the rock. Then it travels along the layer of rock to emerge in lakes and rivers.

Lakes are more affected by acid rain than rivers. They become more and more acidic and concentrations of metal salts increase. Fish cannot live in acidic water. Aluminium compounds e.g., aluminium hydroxide is precipitated and deposited on the gills. The gills become clogged with mucus and the fish die. An acid lake is perfectly transparent because plants, plankton, insects, and other living things have perished.

Acid rain increases the rate of corrosion of metals and buildings containing limestone or marble.



The history of acid rain monitoring

Acid rain was first identified in 1872 in Sweden and studied in the U.S. beginning in the 1950s, when Professor F. Herbert Bormann of Dartmouth College, began taking his botany classes for field trips to the White Mountain National Forest in the early 1950s.

In 1955, he proposed to Forest Service scientist Robert Pierce that the area known as the Hubbard Brook watershed, just west of the village of West Thornton, New Hampshire be designated as an experimental forest. The first stream in the forest was fitted with monitoring devices in 1956.

For the first five or six years after its establishment the Hubbard Brook Experimental Forest (HBEF) was used for research in forestry but soon its value for study of the forest ecosystem was recognized and it attracted the interest of researchers from major universities. The work at Hubbard Brook has led researchers there and elsewhere to try to model or better understand the complex forest ecosystem, including its interaction with humans. HBEF research teams have succeeded in elucidating a number of vexing environmental problems, most notably the harmful effects of acid rain.

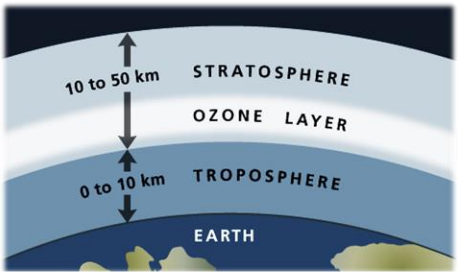
2.9.8 Chlorofluorocarbons and ozone

Chlorofluorocarbons (CFCs) are harmful to the ozone layer when they are emitted into the air. They are currently banned from use. CFCs were used as refrigerants and aerosols, however nowadays they have been replaced with ozone friendly substitutes.

2.9.8.1 The ozone layer

There is a layer of ozone (O_3) surrounding the Earth in the stratosphere. It is 5 km thick at 10 – 50 km from the Earth's surface.

The **ozone layer acts as a protective screen against excessive UV radiation reaching the Earth's surface**. Therefore, high-level ozone protects life from sun's UV rays!



Source: http://www.artinaid.com

O_3 is produced when energy from UV causes oxygen molecules to break into oxygen radicals,

$$O_2(g) \xrightarrow{UV} 2O^*(g)$$

and then the oxygen radicals combine with the oxygen molecules to form ozone.

$$O^*(g) + O_2(g) \rightarrow O_3(g)$$

At the same time ozone absorbs UV and breaks down into an oxygen radical (O^*) and an oxygen molecule (O_2). Over millions of years, a natural equilibrium has been established due to the continual formation and destruction of ozone.

$$O_3(g) \rightleftharpoons O^*(g) + O_2(g)$$

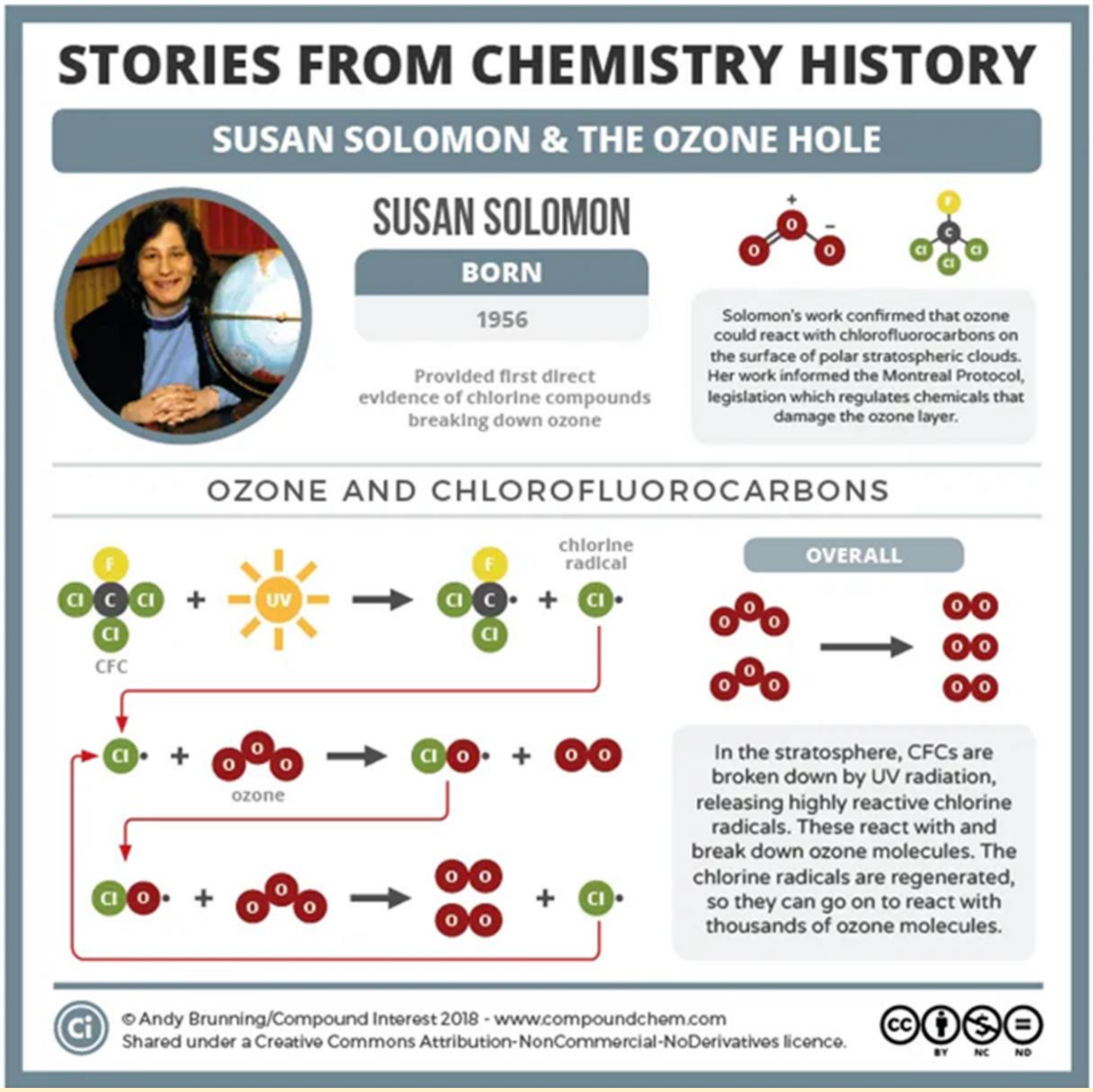
2.9.8.2 Chlorofluorocarbons & ozone depletion

CFCs include $CFCl_3$, CF_2Cl_2, $C_2Cl_2F_4$. From 1930s until late 1980s chemicals called CFCs were used extensively as refrigerants, spray propellants, air conditioning systems and in the production of plastic. They seemed ideal for such purposes because of their lack of reactivity, low flammability, and low toxicity. When the gases are released, they slowly rise into the stratosphere, a process that takes 20 to 30 years.

In the stratosphere, the CFC molecules are broken apart by U.V. radiation and release reactive chlorine atoms (chlorine radicals, Cl*). These atoms in turn react and destroy ozone molecules in a repeating cycle.

A single CFC molecule can destroy 100,000 ozone molecules. World-wide ozone levels have declined from 3% to 7% over the past few decades. Thus, the ozone layer is becoming thinner and ozone 'holes' have been discovered during the 1980s over the Arctic and Antarctic Region. It is thought that this "hole" started to form in the late 1970s. "Holes" refer to areas in the atmosphere where the concentration of ozone is very low when compared to other parts of the atmosphere.

CFCs are also considered to be **greenhouse gases** and contribute to global warming.



STORIES FROM CHEMISTRY HISTORY

SUSAN SOLOMON & THE OZONE HOLE

SUSAN SOLOMON

BORN

1956

Provided first direct evidence of chlorine compounds breaking down ozone

Solomon's work confirmed that ozone could react with chlorofluorocarbons on the surface of polar stratospheric clouds. Her work informed the Montreal Protocol, legislation which regulates chemicals that damage the ozone layer.

OZONE AND CHLOROFLUOROCARBONS

OVERALL

In the stratosphere, CFCs are broken down by UV radiation, releasing highly reactive chlorine radicals. These react with and break down ozone molecules. The chlorine radicals are regenerated, so they can go on to react with thousands of ozone molecules.

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2.9.8.3 CFC Ban

The Antarctic hole was discovered in 1985. The Montreal protocol finalised in 1987 is a global agreement to protect the stratospheric ozone by phasing out the production and consumption of ozone depleting substances. The Protocol is to date one rare treaty to achieve universal ratification. Aerosol cans now use a different propellant labelled as ozone friendly or use pump action cans.

Scientists announced that the depletion of the ozone layer is slowing down following the international ban on CFCs. Satellites and ground stations have confirmed that the stratospheric ozone depletion rate has slowed down significantly during the past decade. CFCs have extremely long atmospheric lifetimes, ranging from 50 to over 100 years, so the final recovery of the ozone layer is expected to take several lifetimes.

2.9.8.4 Depletion of the ozone layer and its effect on humans and the environment

Until very recently the thickness of the ozone layer in the stratosphere remained relatively constant. This situation has altered over the last 50 years as a result of chemicals (CFCs) released in the atmosphere. When the natural equilibrium is destroyed, the amount of ozone gets depleted, and more UV light reaches Earth's surface.

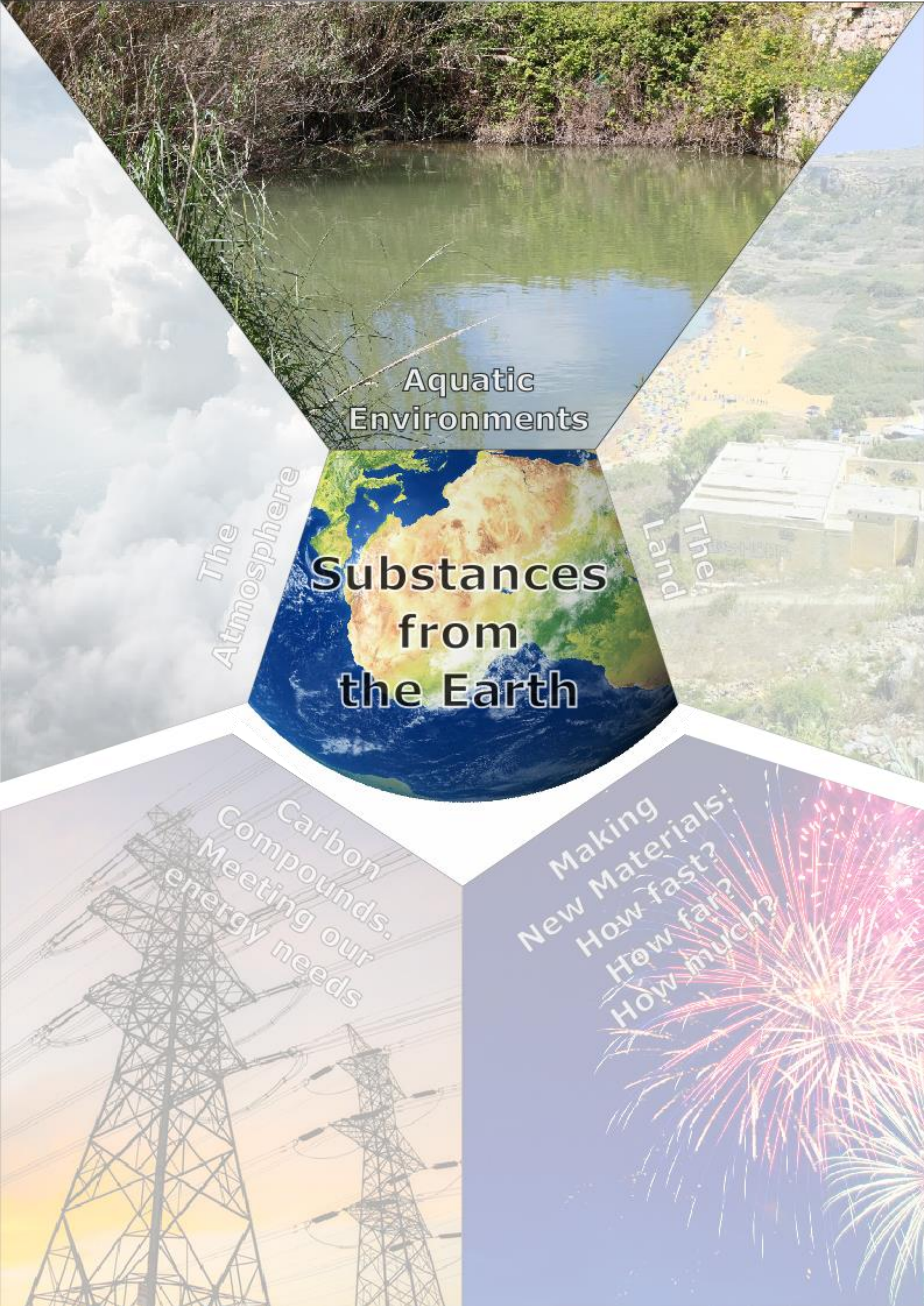
Long exposure to ultraviolet light can cause:

- skin cancer,
- additional cases of eye cataracts,
- weakens the immune system,
- damages crops,
- An excess of U.V. light kills phytoplankton, the minute plant life of the oceans which are the primary food on which life of an ocean depends.

Ozone can be either good or bad depending on where it is found.

- **Bad** – ozone near to the Earth's surface in the **troposphere** is an air pollutant with harmful effects on the respiratory systems of animals.
- **Good** – ozone in the upper atmosphere in the **stratosphere** protects living organisms by preventing harmful ultraviolet light from reaching the Earth's surface.

Air quality widget by ERA



Substances from the Earth

3 The solvent action of water including the impact of water hardness (LO 3)

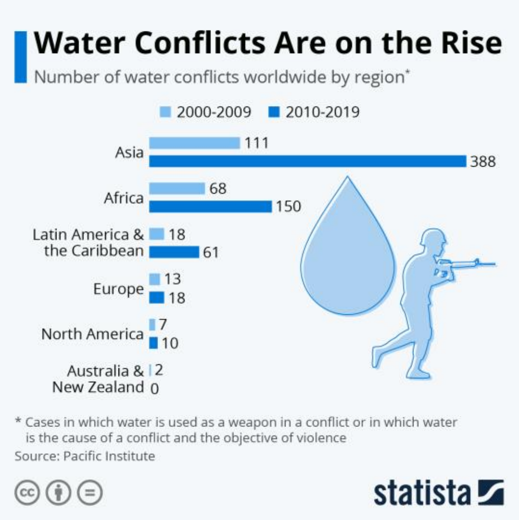
3.1 Water – a source of life and conflict

By the end of this section, you should be able to:

- Present ideas that water is a very precious resource in the world and a potential source of conflict. (3.2a)
- Relate ideas based on research about why water is a very precious resource in the world and a potential source of conflict. (3.3a)

All living things rely on water. Humans can survive up to 30 days without food but only a few days without water! Population growth and distribution have been linked to the availability of fresh water for centuries. Access to freshwater has a direct positive impact on people and communities leading to significant social, economic, and environmental benefits. Throughout history, people have settled next to water sources for drinking and to grow crops such as the Seine River in Paris to Lake Texcoco in Mexico City.

Water wars describes conflicts between countries, states, or groups over the fight to access water resources usually freshwater. The conflicts occur when the demand for potable water exceeds the supply or when allocation or control of water is disputed. Today, the world's exploding population and global warming are the two main factors that are igniting water wars. Water conflicts are on the rise particularly in Asia and in Africa as can be seen on the infographic provided by Statista. Conflicts over water are brewing:


Water Conflicts Are on the Rise

Number of water conflicts worldwide by region*

* Cases in which water is used as a weapon in a conflict or in which water is the cause of a conflict and the objective of violence

Source: Pacific Institute

- Between Egypt and Ethiopia over the latter's construction of the Great Ethiopian Renaissance Dam located in the headwaters of the Nile River. The Nile has been Egypt’s major water source since antiquity. Currently, Egypt is using more water than its internal, renewable resources can provide. Demand for water is projected to only grow due to rapid population growth and rising temperatures, and this will put strains on Egypt’s economy.
- India has been building dams that neighbouring Pakistan views as a threat. The long-standing conflict over water from the Kaveri River between the Indian states Karnataka and Tamil Nadu has recently resurfaced in the context of drier climate conditions. The implications are not only legal battles, but also violent protests following decisions to alter water distribution between the two states. Click on the link above to follow current issues related to this water-related issue.

- Turkey has been diverting water from the Tigris and Euphrates rivers, which affects Iraq and Syria.
- China has been on a "dam building spree" along the Mekong. China's dams have restricted flow and lead to severe drought in the lower Mekong basin, which is an important water and food source for Thailand, Lao PDR, Cambodia, Vietnam, and Myanmar.

3.2 Potable water sources in Malta

By the end of this section, you should be able to:

- Identify sources of potable water and their management in Malta. (3.1a)
- Name desalination techniques that can be used to create demineralised water from sea water. *(Limited to distillation and reverse osmosis.)* (3.1n)
- Describe how simple distillation and reverse osmosis are used to produce demineralised water from impure water. (3.2n)
- Evaluate desalination techniques that can be used to produce demineralised water from sea water. *(Limited to distillation and reverse osmosis.)* (3.3n)

Water needs to be of sufficient quality to be fit for human consumption. This kind of water is said to be **potable**. Although water is very common on Earth, only a small percentage is available as potable water. The majority of water exists as sea water which is too salty for humans and land-dwelling flora and fauna.

In countries where mountains reach a sufficient height that ice and snow form all year round at their tops, a steady flow of water forms into streams and rivers. This water is potable although in some cases it needs to be filtered and treated for pathogens by chlorination or exposure to ultraviolet light.

The quality of potable water in Malta is regulated by law (Subsidiary legislation 449.57 – WATER INTENDED FOR HUMAN CONSUMPTION REGULATIONS, dated: 16th January 2009) which states that water must meet the minimum requirements of dissolved chemical agents such as nitrates, nitrites, copper, lead, and be free from any micro-organisms and parasites.

Potable water in Malta is quite a problem since we have no lakes or rivers. Malta's semi-arid Mediterranean climate characterised by a general lack of rainfall especially during the summer period and high temperature results in low natural water availability. Potable water is obtained from ground water reservoirs which is replenished by rainwater. However, the amount of ground water is not enough for the ever-growing population of the Maltese islands and for the lack of rainfall.

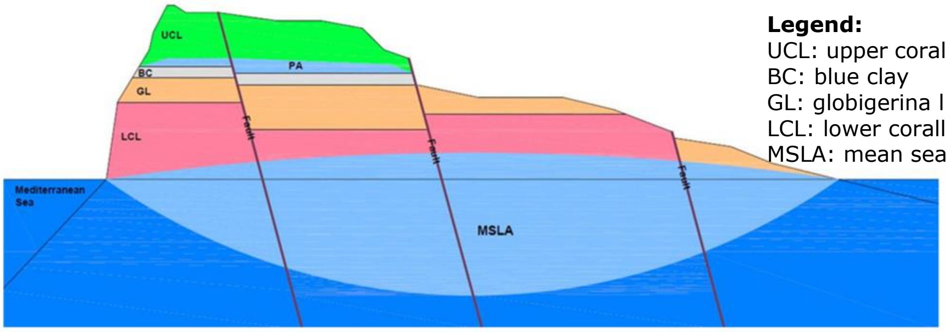
Indeed, water is a **scarce resource** in the Maltese islands.

Malta is the most water stressed European country and one of the top ten most water scarce countries in the world.

The following research article discusses Malta's water scarcity challenges.

3.2.1 Ground water

Groundwater is the only natural water resource which is available all year round. Potable water can be obtained through extraction of ground water by means of boreholes and underground rock cut tunnels called galleries, in pumping stations such as the one at Ta' Kandja which is situated in the limits of Siġġiewi. It is 97 m below ground and consist of a network of 6.2 km of galleries. Ground water from Ta' Kandja is obtained from the mean-sea level aquifer. The aquifer consists of a lens shape body of fresh water floating over the sea water. It is contained in the pores and fractures of the lower coralline limestone formation. In 2022 ground water accounted for approximately 36% or 12.7 million m^3 of the total potable water production.



Legend:
UCL: upper coralline limestone
BC: blue clay
GL: globigerina limestone
LCL: lower coralline limestone
MSLA: mean sea level aquifer

Source: https://www.mdpi.com/2073-4441/15/1/177

Over time the quality of ground water has deteriorated due to overextraction of water, becoming more saline. Furthermore, ground water is easily contaminated due to use of artificial fertilisers.

The use of boreholes, however, has been highly restricted due to the high salinity levels that are being recorded in groundwater. Replenishment of groundwater is also very limited due to the shortage of rainfall and surface runoff.

Source: Water Services Corporation

Ground water at Ta' Kandja mean-sea level aquifer.

There are two types of aquifers in the Maltese Islands:

1. the **perched aquifer** where rainwater seeps through the upper coralline limestone is collected over the impermeable blue clay layer and
2. the **mean sea level aquifer** which refers to a deep more extensive aquifer found in the lower coralline limestone at mean sea level.

The image on the right shows part of the Wignacourt aqueducts in Birkirkara which were structures used by the Knights of St John to deliver water from Rabat to Valletta way back in the 1600s. Even at that time, the growing population in Valletta could not be supported by groundwater alone and so the Knights' solution was to transport water from Rabat, where fresh water from the perched aquifer was available, to Valletta where it was needed.

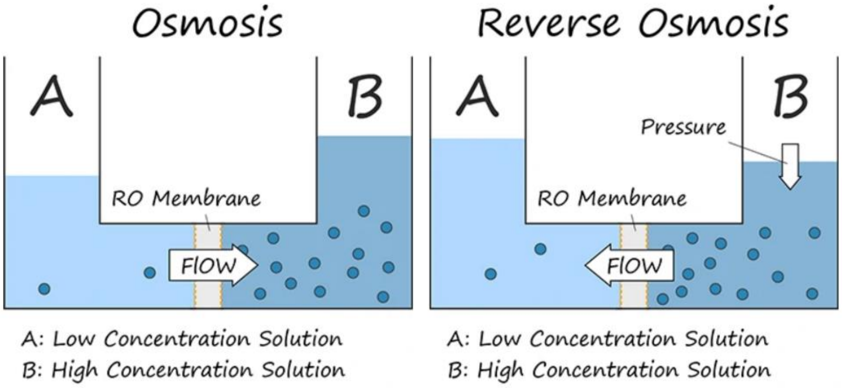
3.2.2 Reverse osmosis

The amount of groundwater available is not enough for the ever-growing population on the Maltese islands. To supplement the lack of groundwater, Malta has four **desalination** or **reverse osmosis** (RO) plants in Għar Lapsi, Pembroke, Ċirkewwa, and another one in Gozo. According to Water Services Corporation, 'the total amount of water produced from Reverse Osmosis plants in 2022 was 22,785,700 cubic meters, accounting for 64% of the total blend or an average of 62,425 cubic meters per day'. This proportion is set to increase as the demand for potable water increases and the amount of annual rainfall decreases.

At the RO plants, seawater that is extracted from deep shore wells drilled in Coralline Limestone. Sea water is filtered through cartridge filters and then salt is removed from sea water through a process known as reverse osmosis.

Osmosis is the passage of water from a solution with a LOW concentration of solute to a solution with a HIGH concentration of solute through a semi-permeable membrane.

Reverse osmosis is the reverse process where water from a solution with a high solute concentration (i.e., the brackish or sea water) passes to a solution with a lower solute concentration. Energy and high pressure are needed for this to occur.



Source: https://www.best-osmosis-systems.com/what-is-reverse-osmosis-membrane/

The potable water produced is chlorinated to kill any pathogens and remineralised by adding lime. Chlorination is a process that is used to purify water from any pathogens that may be present in water especially when the water is intended for human consumption. By the end of the process, silt, organic material, and salt is removed. Finally, water from this process is blended (mixed) with chlorinated ground water and is pumped into the distribution network.

Converting seawater into potable water is a very costly process, since a substantial quantity of electricity is required to operate the pumps that apply pressure on sea water through special filters. However, without this process Malta could never support the current population! Although Malta does not have the freshwater resources that other European countries have, we are surrounded by seawater that can be converted into freshwater using reverse osmosis albeit at a high cost. More so, in the light that other countries situated in arid regions which are further inland do not have this luxury and suffer droughts that extend over years on end, making life unbearable. This is another reason why we should keep our seas clean and use water carefully!

An article written by Marco Cremona on water myths and untruths in Malta.

In Gozo, at Ħondoq ir-Rummien, there was a derelict building which used to be a distillation plant. It used to supply fresh water from sea water to both Gozo and the north of Malta. It was no longer used for this purpose because there came a time where it was no longer economically feasible to run such an operation.

According to news reports, this building was being converted to a reverse osmosis plant to supply fresh water to Gozo.
On the 18th of November 2021, the new reverse osmosis plant at Ħondoq ir-Rummien in Gozo was inaugurated.

3.2.3 Desalination through distillation

Distillation is the process of removing any soluble salts from water. It converts brackish or sea water into pure water. The process involves heating the water to form steam and then the steam is cooled back into liquid form which is free from impurities. Distillation is further explained in section 3.8.2.2.

During Malta's history regarding the quest for the provision of fresh water to an ever-growing population, there was a time when the extraction of ground water alone was not meeting the needs of the population and so distillation of seawater was used to produce the amount of fresh water needed.
In Malta, a sea water distilling plant was constructed in Sliema in 1881 to provide water for Tigne Barracks, but this was a temporary solution.
In Gozo according to Water Services Corporation,

> "By 1966 the first seawater… distillers were producing 20,430 cubic metres of water per day. In 1971 the first distiller in Gozo began producing 2,270 cubic metres per day."

Eventually, distillation as a means of producing fresh water from seawater was replaced by reverse osmosis plants in the period between 1982 and 1988. This happened because reverse osmosis is more economically viable than distillation.

Distillation is a very energy-intensive process and since it is expensive it is not the most practical solution for large scale water treatment. The process of reverse osmosis (RO) requires less energy compared to distillation especially with advancements in membrane technology. RO systems can be easily scaled up or down according to water demand unlike distillation plants. Both desalination techniques are highly efficient at removing salts and minerals. Considering their environmental impact there is a high energy consumption in distillation leading to more greenhouse gas emissions, while throwing a high brine concentration into the sea may cause environmental problems in the sea next to industrial RO systems.

3.3 Properties of pure water

By the end of this section, you should be able to:

- Identify physical properties of pure water. (3.1b)
- State criteria of purity for water. *(Limited to melting point, boiling point, density, and conductivity.)* (3.2b)

All life depends on water. It is needed for the correct functioning of our bodies. Water is by far the most common natural compound on Earth. Water is used to generate electricity, in cooling, transporting things, to irrigate crops, cleaning and as a medium for reactions.

The **physical properties of pure water** are:

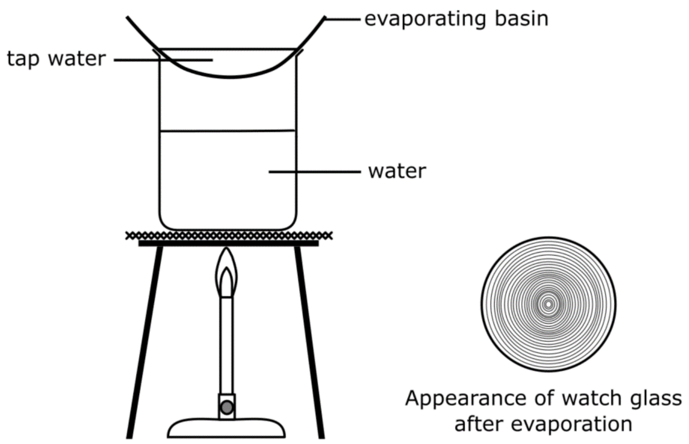
- It is odourless, tasteless, and colourless.
- It boils at 100 °C (at standard pressure that is 1 atmosphere) and freezes at 0 °C.
- The solid form of water (ice) is unusual because unlike other substances, it is less dense than the liquid form. This is why ice floats on water. The reason for this is that water expands upon freezing unlike other liquids. This causes pipes to burst in winter. In fact, its maximum density is 1 g/cm^3 at 4 °C.
- Water is a very good solvent.
- It is neutral to litmus.
- Pure water does not conduct electricity.

Pure water has nothing else dissolved in it. To check that a sample of water is pure you need to check its melting point and/ or boiling point.

Water is so good a solvent that it does not exist as pure water in nature. It has many dissolved substances in it. Water has often been described as a **universal solvent** because it dissolves most common substances, even if for some it only dissolves them slightly. The mixture that results when water dissolves soluble substances in it is called a **solution**. Water also has several gases dissolved in it, including oxygen which is important to marine and freshwater life.

Most covalently bonded substances are insoluble in water. However, some such as, glucose and sucrose are very soluble in water.

If some tap water is evaporated to dryness as shown in the diagram below, a residue of solid matter is left as the water evaporates.

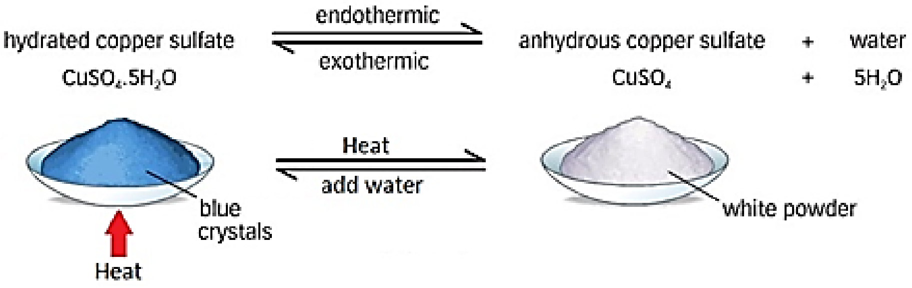


3.3.1 Hydrated salts

Hydrated salts are salts that contain water molecules in their structure. This water is referred to as **water of crystallisation**. The shape of the crystals is very much dependent on the presence of water of crystallisation. Some examples of hydrated salts include hydrated copper(II) sulfate, $CuSO_4.5H_2O$; hydrated iron(II) sulfate, $FeSO_4.7H_2O$ and hydrated cobalt(II) chloride, $CoCl_2.6H_2O$.

When hydrated salts are heated the water of crystallisation is removed to form an **anhydrous** salt. When water is added to the anhydrous salt it becomes hydrated. This is known as a **reversible reaction**.

$$\text{hydrated salt} \underset{\textit{add water}}{\overset{\textit{heat}}{\rightleftharpoons}} \text{anhydrous salt} + \text{water of crystallisation}$$



Source: https://www.tutormyself.com/edexcel-igcse-2017chem-318/

3.3.2 Testing for the presence of water

A few drops of an unknown liquid is added to white anhydrous copper(II) sulfate powder.

If the powder turns from white to blue, then the liquid is water.

$$\text{anhydrous copper(II) sulfate} + \text{water} \rightleftharpoons \text{hydrated copper(II) sulfate}$$

$$CuSO_4(s) + 5H_2O(l) \rightleftharpoons CuSO_4.5H_2O(s)$$

Another test can be conducted. A few drops of an unknown liquid is added to anhydrous cobalt(II) chloride crystals. If the crystals turn from blue to pink, then the liquid is water.

$$\text{anhydrous cobalt(II) chloride} + \text{water} \rightleftharpoons \text{hydrated cobalt(II) chloride}$$

$$CoCl_2(s) + 6H_2O(l) \rightleftharpoons CoCl_2.6H_2O(s)$$

Note that the above tests are only used to identify whether an unknown liquid is water. To find whether a given sample is pure water check for its melting and boiling point.


Anhydrous Cobalt(II) Chloride


Cobalt(II) Chloride Hexahydrate

- If the sample is pure water its melting point is 0°C and its boiling point is 100°C.
- If the sample is impure, it will have a higher boiling point and a lower melting point.

3.4 Mixtures

By the end of this section, you should be able to:

- Distinguish between solute, solvent, and solution. (3.1i)
- Distinguish between soluble and insoluble substances. (3.1j)

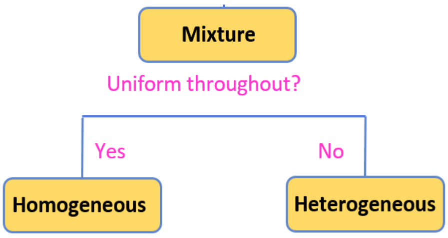
Sea water is an example of a mixture. The main solid dissolved in sea water is sodium chloride. Since sea water is the reservoir into which all soluble substances end up, (because rainwater eventually flows back to the sea), it contains a relatively high level of soluble solid impurities (3.6%). It would also contain calcium hydrogencarbonate because ground water eventually also makes its way to the sea. Other solids found in naturally occurring water include calcium sulfate, magnesium hydrogencarbonate, magnesium sulfate, silicates, nitrates, and ammonium salts. Gases such as oxygen, nitrogen and carbon dioxide will also be found in seawater.

3.4.1 Mixed states

Mixtures can be of two kinds: **homogenous** or **heterogenous**.

Homogenous mixtures are mixtures where the parts of the mixture are all in the **same phase or state of matter**. Rock salt is a homogenous mixture because its parts (salt and rock) are in the solid phase. Vodka is an example of a homogenous mixture because it is made of two liquids (ethanol and water). Air is a homogenous mixture because its parts are all in the gaseous phase.

Heterogenous mixtures are mixtures where its parts are in **different states of matter**. For example, a mixture of sand and water is a heterogenous mixture. A foam is a mixture of liquid and air. A gel is a mixture of a liquid dispersed in a solid mesh.




Sweet tea

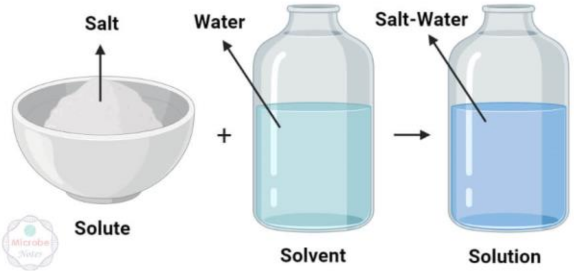

Ocean, wet sand

3.4.2 Solid/liquid mixtures - Soluble substances

Solutions are mixtures. Solids can “disappear” (or in scientific terms **dissolve**) in a liquid.

When this happens, we say that the solid is **soluble** in the liquid. When a solid dissolves, a mixture of the solid and the liquid is formed, e.g., when salt is dissolved in water. The mixture (of solid and liquid) is called a **solution**. Note that a solution in water is called an **aqueous** solution. The liquid which does the dissolving is called the **solvent**. The dissolved solid is called the **solute**.

If the solid does not dissolve in a liquid, we say that the solid is **insoluble** in that liquid. If the insoluble solid is present in a liquid as a fine powder, the mixture forms a **suspension**, where on standing the solid particles take a long time to settle at the bottom of the container.



Source: https://thechemistrynotes.com/solute-vs-solvent/

3.5 Obtaining table salt

By the end of this section, you should be able to:

- Describe how salt is produced from sea water. *(By evaporation and crystallisation.)* (3.1c)
- Explain how salt is produced from rock salt. *(By solution, filtration, evaporation, and crystallisation.)* (3.2c)
- Produce crystals of salt from rock salt. (3.3c)
- Compare size of crystals obtained from slow and fast crystallisation methods. (3.3d)

Common salt (sodium chloride) is an important substance. It is either extracted from sea water or by first mining rock salt from underground and then extracting salt from it.

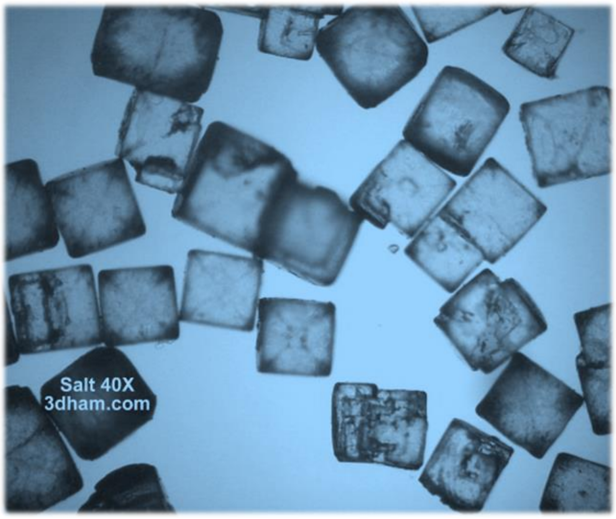
3.5.1 Production of salt from seawater in Malta


Facebook (2021):
Xwejni Salt Pans Leli tal-Melħ

The picture to the right shows a person collecting salt from salt pans. These can be found along the coast of both Malta and Gozo such as those found in Salina and Xwejni.

Salt pans are shallow basins dug in coastal rock which are filled with seawater. During the summer months when the sun is at its hottest, the rock becomes very hot, and water evaporates. This process is called **evaporation**. As the water evaporates, crystals of salt form in the salt pans. This process is called **crystallisation**.

These processes also happen naturally in rock pools on the coast where seawater is trapped during storms and then evaporates to leave a residue of salt as shown in the picture below. Keep your eyes open when you visit a rocky seaside during the summer months and look closely at the crystals in these rock pools.

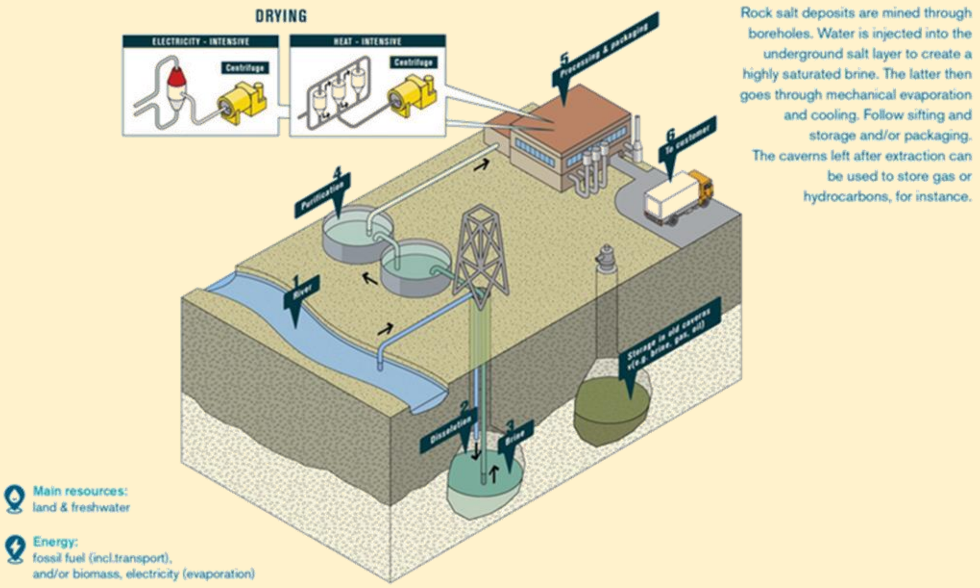


Since salt obtained this way forms over a matter of days, the salt crystals formed are large because they have time to form properly. One could note that crystals of salt (sodium chloride, NaCl) have a cubic structure.

3.5.2 Production of salt from rock salt

In Europe and other countries, many rock salt deposits were formed over 200 to 250 million years ago as a result of the evaporation of ancient seas. This includes a basin that stretched from the UK to Poland. However, this has eventually dried up and formed one of the layers in sedimentary rock. The preferred method of mining rock salt is known as solution mining. In this process, water is forced under pressure into a borehole drilled into an underground salt layer. The salt dissolves in the water to form brine. Brine is a concentrated solution of salt in water which is also used as a food preservative. The salt solution is then **filtered** to remove any solid insoluble impurities (pieces of rock). The solution then undergoes **evaporation** and **crystallisation** in order to obtain solid salt crystals.

Production of salt on an industrial scale



Rock salt deposits are mined through boreholes. Water is injected into the underground salt layer to create a highly saturated brine. The latter then goes through mechanical evaporation and cooling. Follow sifting and storage and/or packaging. The caverns left after extraction can be used to store gas or hydrocarbons, for instance.

Your teacher will give you a sample of rock salt. It is a mixture of rock and salt.

a. Predict what happens when a sample of rock salt is added to water.

b. Once you have produced a mixture of water and rock salt, discuss how you can separate:

  i. the insoluble component in rock salt.

  ii. salt from the mixture.

c. Discuss the precautions you need to take during the experiment.

3.5.3 Separation techniques

Filtration, evaporation, and crystallisation are separating techniques.

These techniques are used to separate substances that are mixed together. The kind of separating technique used depends on the properties of the substances mixed together. These techniques can be used in a laboratory to separate substances from each other so as to obtain pure forms of these substances.

3.5.3.1 Filtration

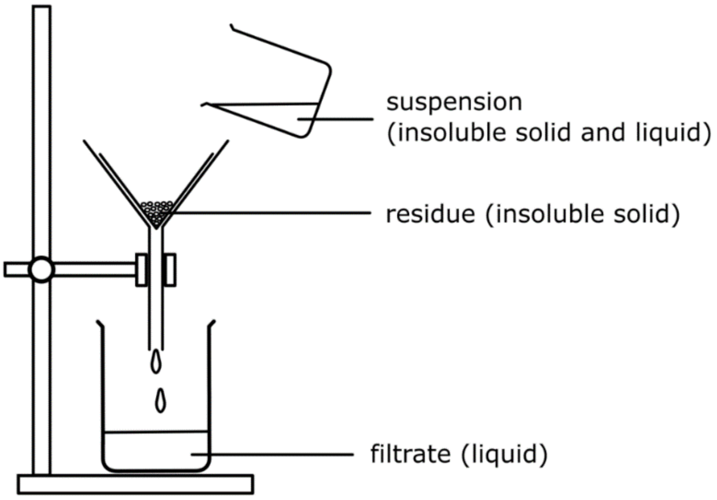
Filtration can be used to separate an insoluble solid from a liquid or aqueous solution.

The size of the particles of the insoluble solid has an effect on how the mixture acts on standing.

- If the particle size of the insoluble solid is relatively large as in the case of sand, no matter how vigorously one mixes the mixture of sand and water, the insoluble solid (sand) can be observed to fall to the bottom of the container very quickly.
- If on the other hand, the insoluble solid consists of very fine particles, then on mixing with a liquid, the insoluble solid takes a very long time to settle at the bottom of the container. When this happens, we say that we have a suspension.

A **suspension** is a mixture of fine insoluble solid particles suspended in a liquid.

The apparatus below is used to filter a mixture of sand and saline water.



The insoluble solid that remains on the filter paper is called the **residue**. The residue is washed with distilled water to remove any soluble impurities to ensure purity.

The liquid or aqueous solution that passes through the filter paper is called the **filtrate**.

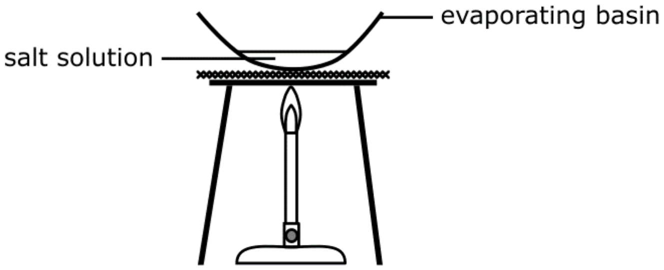
The following precautions are taken when conducting filtration. Explain why they are taken.

- A few drops of distilled water are added to the funnel before inserting the filter paper.
- The mixture is poured slowly during filtration.
- The residue is washed with distilled water.
- Sometimes two filter papers are used when filtering a fine powder from a suspension.

3.5.3.2 Evaporation to dryness

Evaporation to dryness is used to obtain a soluble solid from a liquid solvent.

The slower the rate of evaporation, the larger the crystals formed. This happens because they would have more time to form properly. This process can also be done by leaving the evaporating basin to stand on a windowsill so that the solvent (water) evaporates slowly and leaves behind large crystals of the solute (salt).



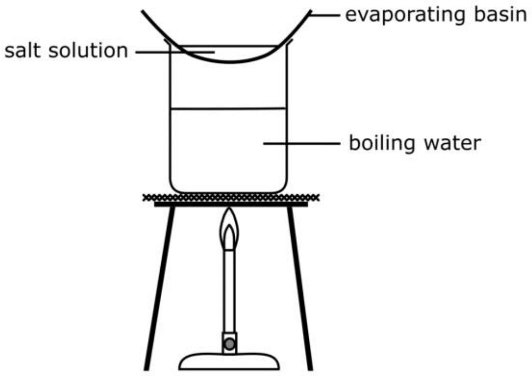
If the evaporating dish is heated until all the solvent evaporates then the salt crystals formed would be very small. This is known as evaporation to dryness.

Also, using a Bunsen burner as suggested by the diagram should only be used for salts that are thermally stable. **Thermally stable** compounds are those that do **not** decompose on heating.

Evaporation to constant mass.

3.5.3.3 Crystallisation

This method is used to obtain a **soluble** solid from the **liquid** solvent. A water bath (as in the diagram on the right) should be used. **Crystallisation** is the process of cooling a saturated solution so that the solubility of the solute decreases and the excess solute in the solution forms crystals within the solution. The remaining solution becomes super saturated.



If a solution is left to stand, the solvent (water) evaporates, and the same effect is achieved.

This method involves evaporating water at a slow controlled rate. It is best used when:

- **A hydrated salt is required**.

  For example, if crystals of copper(II) sulfate pentahydrate ($CuSO_4.5H_2O$) are required from solution this method must be used. Using evaporation to dryness instead, would drive off all the water leaving behind white anhydrous copper(II) sulfate.

- **The salt required is not thermally stable**.

  For example, if sodium hydrogencarbonate needs to be obtained from solution it will decompose to sodium carbonate if direct heating is used to evaporate all the water.

- **Large crystals are required**.

Evaporation techniques: direct heating vs crystallisation

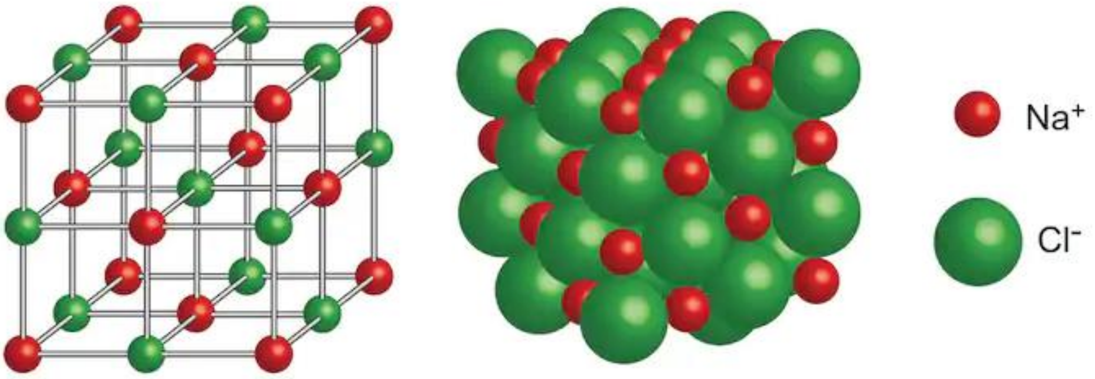
3.6 Salts and ionic bonding

By the end of this section, you should be able to:

- Explain that sea water contains dissolved charged ions that form crystals on evaporation. (3.1e)
- Identify which elements form positive ions and which form negative ions in relation to their position in the periodic table. (3.2e)
- Explain how ionic bonds lead to giant ionic structures. *(Structure limited to sodium chloride. Drawing of structure is not expected.)* (3.3e)
- Explain the properties of ionic compounds. *(Limited to solubility, melting/boiling points, and electrical conductivity in different states.)* (3.3f)
- Determine the electron configuration of ions of the first 18 elements (where applicable) in relation to their position in the Periodic Table. (3.2g)
- Draw dot and cross diagrams to represent ionic binary compounds showing all electron shells. *(Limited to the first 18 elements.)* (3.3g)

A **salt** is a chemical substance that is made of metal and non-metal ions which are chemically combined together and are solid at RTP (room temperature and pressure). The bond that exists between the metal and non-metal **ions** is very strong and is called an **ionic bond**.

Common salt or sodium chloride is an example of an ionic solid. Salts do not form molecules but form **giant ionic lattices** which are large structures of cations (positive ions) and anions (negative ions) held together by strong **electrostatic forces of attraction**. The diagram below shows the structure of sodium chloride.



shutterstock.com • 345228074

In reality the lines between the ions do not exist but they are shown for clarity. Although, in one grain of salt there are millions of cations and anions, the formula of a salt is given according to the ratio of cations with respect to anions. In the case of sodium chloride that ratio is 1:1 and so its formula is NaCl. NaCl is what makes table salt!

An aqueous solution of sodium chloride contains sodium and chloride ions dissolved in water. When a solution of sodium chloride is heated so that the water evaporates, the sodium and chloride ions start to reform the sodium chloride crystals. This holds for all salt solutions.

Animation of sodium chloride dissolving in water.

3.6.1 Ionic (electrovalent) bonding

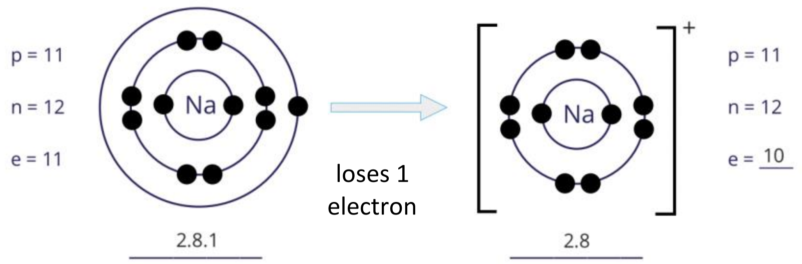
Metals can combine chemically with non-metals to form compounds. The type of chemical bond present when metals react with non-metals is **ionic** (or **electrovalent) bonding**. This means that metals lose one or more electrons while the non-metals gain these electrons. During the reaction between metals and non-metals, metals form positively charged metal ions (**cations**) while the non-metal atoms form negatively charged ions (**anions**).

Ions are charged particles since the number of protons is not equal to the number of electrons.

The **ionic bond** that forms between these ions
is an electrostatic force of attraction between oppositely charged ions.

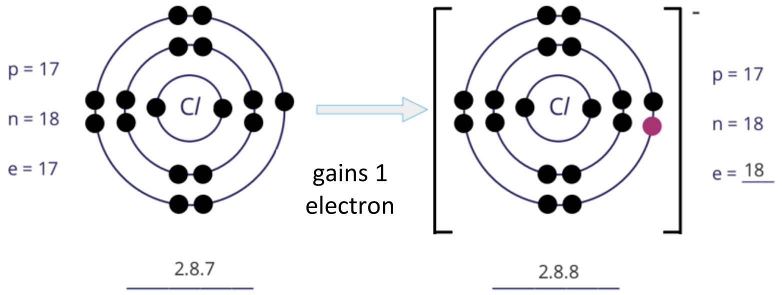
Note that ions have the electron configuration of one of the noble gases. For example:

- A sodium atom (Na) whose electron configuration is 2,8,1 loses one electron to form a **sodium ion (Na^+)**. Its electron configuration becomes 2,8 just like that of neon.
  A positive ion is formed since a Na^+ has 11 protons and 10 electrons.



Source: https://geniebook.com/tuition/secondary-3/chemistry/ionic-bonding

- A chlorine atom (Cl) whose electron configuration is 2,8,7 gains one electron to have a full outer shell to form a **chloride ion (Cl^-).** Its electron configuration becomes 2,8,8 just like that of argon. A negative ion is formed since Cl^- has 17 protons and 18 electrons.



Source: https://geniebook.com/tuition/secondary-3/chemistry/ionic-bonding

Ions are formed when the number of protons is not equal to the number of electrons.

For the following elements (i) write the electronic configuration of the atom (ii) write the symbol of the ion formed (iii) write the electronic configuration of the ion.

Element	Electronic configuration	Symbol of ion formed	Electronic configuration of ion
Li			
Mg			
O			
F			

3.6.2 Valency

Valency has already been discussed in section 2.4.8.

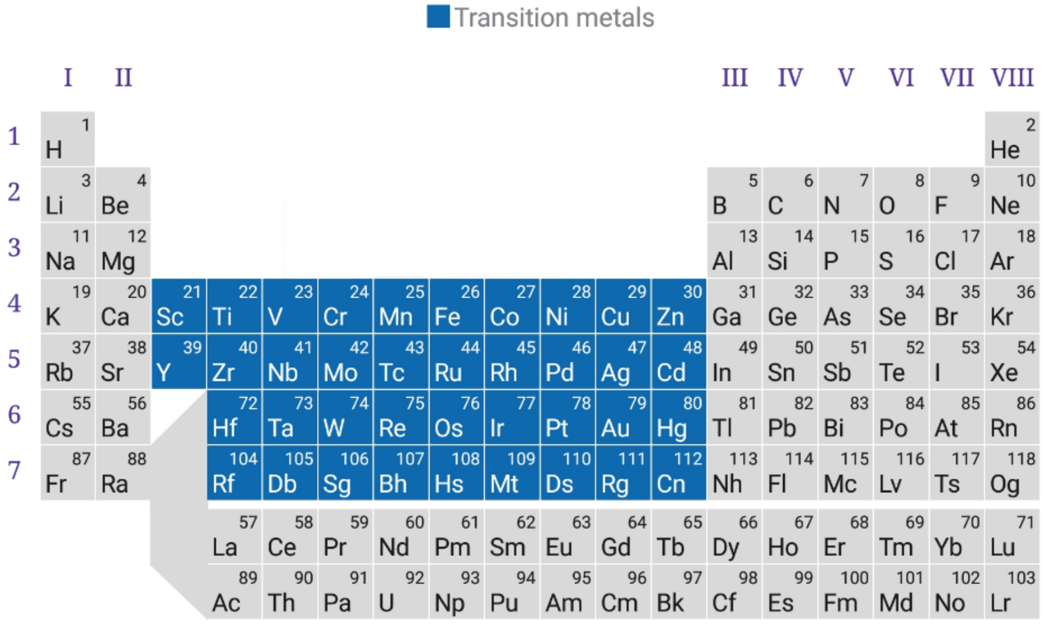
Meanwhile, one should note that as one goes across a period of the Periodic Table, valency increases up to group 4 and then decreases until group 0 is reached.

Periodic Table group	Number of electrons in outer shell	Loses or gains electrons	Charge on ion formed	Valency	Example
1	1	1 lost	1+	1	K^+
2	2	2 lost	2+	2	Mg^{2+}
3	3	3 lost	3+	3	Al^{3+}
4	4	4 gained or lost	4- or 4+	4	C^{4-} or Pb^{4+} *
5	5	3 gained	3-	3	N^{3-}
6	6	2 gained	2-	2	O^{2-}
7	7	1 gained	1-	1	Cl^-

* These examples refer to carbides and lead(IV) compounds which are not part of the syllabus.

The **transition metal block** is situated in between groups 2 and 3 in the Periodic Table.

The elements that are in the transition metal block have variable valencies, e.g., copper can have a valency of 1 or 2, iron 2 or 3.



The reason behind transition metals having variable valency will be dealt with if you take your studies in Chemistry to a higher level.

Other elements that are not in the transition metal block, also have multiple valencies. One such example is lead (Pb). Lead has two valencies, 2 and 4. Therefore lead forms two ions: Pb^{2+} and Pb^{4+}.

- State the valency of the following metals: sodium, calcium, aluminium, lithium, magnesium.
- Give two transition metals apart from copper and iron.

3.6.2.1 Drawing dot cross diagrams of ionic compounds

You will be asked to draw a dot-cross diagram showing all electron shells for binary compounds between metal and non-metal ions.

3.6.2.1.1 Sodium chloride

When sodium is burnt in chlorine, a white crystalline solid of sodium chloride, is formed.

sodium + chlorine → sodium chloride

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$

Let's take the electron configurations of sodium and chlorine.

Sodium atom: atomic number (Z) = 11. Therefore, its electron configuration is 2,8,1.

Chlorine atom: atomic number (Z) = 17. Therefore, its electron configuration is 2,8,7.

Sodium needs to **lose** one electron to have a full outer electron shell.

On the other hand, chlorine must **gain** one electron to have a full outer electron shell.

When sodium and chlorine react, an electron from the sodium atom is transferred to an atom of chlorine. Consequently, a sodium ion and a chloride ion are formed.

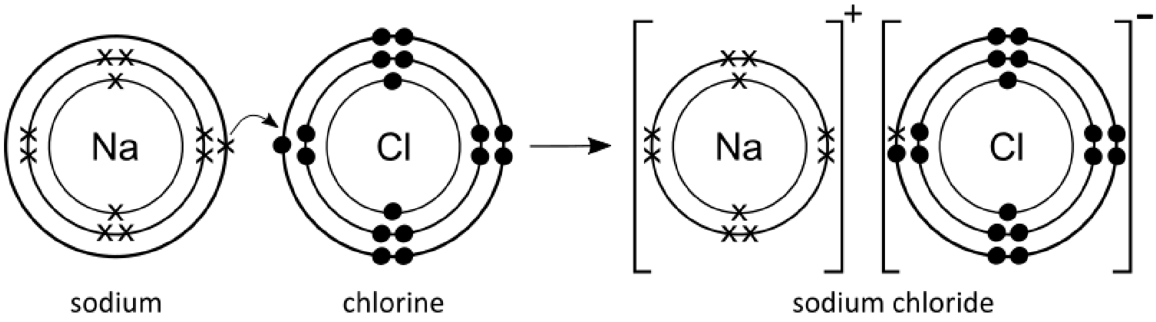
Sodium has lost 1 electron and now has a positive charge of 1+.

Chlorine has gained an electron and now has a negative charge of 1-.

Ions are written with their charge next to them, so sodium and chloride ions are written as: Na^+ and Cl^-.

Since both ions are oppositely charged, they are attracted towards each other and form a bond between them. This is the **ionic bond.**

Note that the diagram shows how sodium and chlorine react to produce sodium chloride.



Sometimes you are just required to draw the dot cross diagram of sodium chloride. In such a case you need to draw the ions enclosed within square brackets with the charge of the ions on the right top corner only. Failing to draw either the square brackets or the charge as indicated will be marked as incorrect.

3.6.2.1.2 Magnesium oxide

When magnesium is burnt in oxygen, a white ash of magnesium oxide is formed.

magnesium + oxygen → magnesium oxide

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

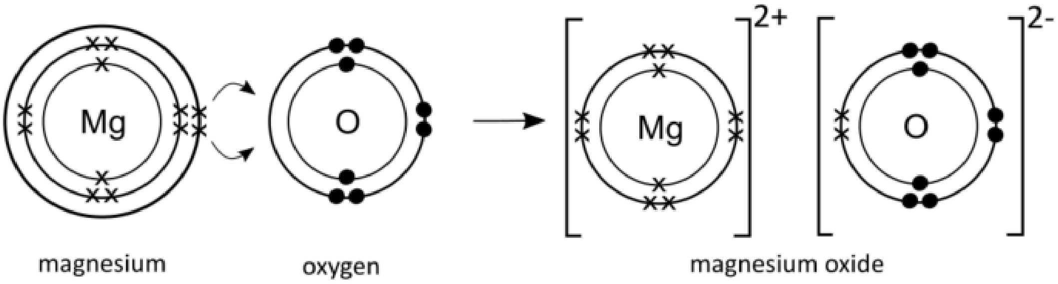
Let's take the electron configurations of magnesium and oxygen. The number of electrons that will be lost or gained to have a full outer shell are shown for each element.

Element	Atomic number	Electronic configuration	To obtain a full shell
Magnesium	12	2,8,2	loses 2 electrons
Oxygen	8	2,6	gains 2 electrons

When magnesium and oxygen react, 2 electrons from a magnesium atom are transferred to an atom of oxygen. Consequently, a magnesium ion Mg^{2+} and an oxide ion O^{2-} are formed.

Magnesium lost 2 electrons and now has a positive charge of 2+

Oxygen gained 2 electrons and now has a negative charge of 2-.



3.6.2.1.3 Magnesium chloride

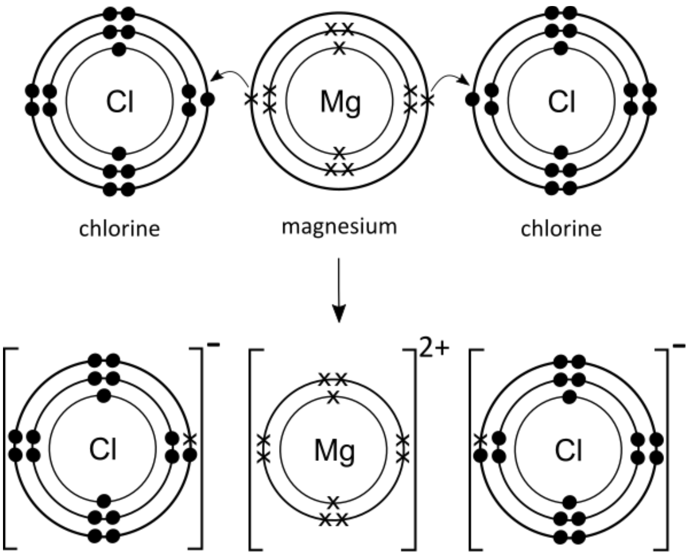
When magnesium burns in chlorine, magnesium chloride is produced.

magnesium + chlorine → magnesium chloride

$$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$$

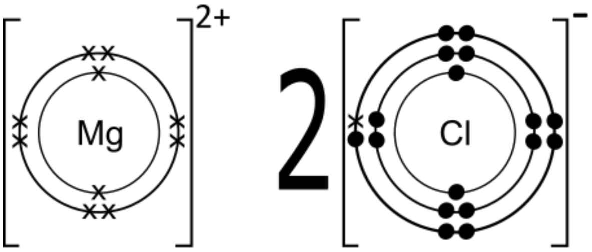
Element	Atomic number	Electronic configuration	To obtain a full shell
Magnesium	12	2,8,2	loses 2 electrons
Chlorine	17	2,8,7	gains 1 electron

So, Mg needs to lose 2 electrons while 2 chlorine atoms each gain 1 electron to have a full outer electron shell. The diagram shows what happens when they react:



magnesium chloride

Magnesium forms an ion with a 2+ charge and chlorine forms two chloride ions with a 1- charge each. In the case where there are two similar ions, one can draw just one ion and then write the multiplier in front of the respective ion as shown below.



magnesium chloride

Note that the number of electrons lost or gained by an atom is equal to the valency of that same atom.

3.6.2.1.4 Sodium oxide

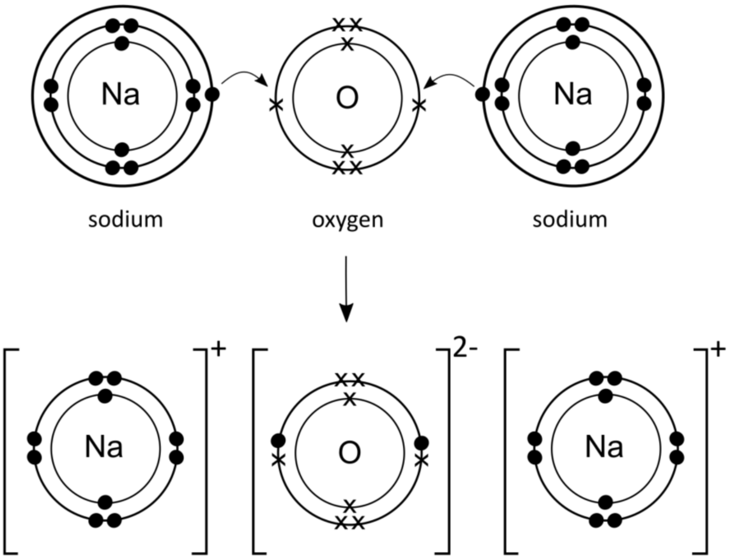
When a piece of sodium is cut with a knife it immediately reacts with oxygen in the air to form sodium oxide.

sodium + oxygen → sodium oxide

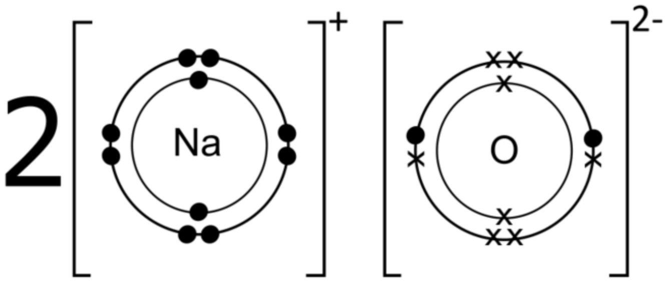
$$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$$

Element	Atomic number	Electronic configuration	To obtain a full shell
Sodium	11	2,8,1	loses 1 electron
Oxygen	8	2,6	gains 2 electrons

Two sodium atoms each lose an electron and oxygen gains these 2 electrons for all to have a full outer electron shell. Consequently, a sodium ion Na^+ and an oxide ion O^{2-} are formed. The diagram below shows what happens when they react:



sodium oxide



sodium oxide

Draw dot cross diagrams showing all electron shells of: LiF, Na_2S, and AlF_3.

3.6.2.2 Properties of ionic compounds

Ionic compounds are formed from two oppositely charged ions, e.g., Na^+ and Cl^-. These oppositely charged ions attract to form large three-dimensional structures called salts.

- These ions are held by strong electrostatic forces, and so a large amount of energy is needed to separate them which is why they have high melting and boiling points.
- Ionic compounds are usually crystalline solids at RTP. Ionic compounds have a regular, repeating pattern called a crystal lattice structure. This arrangement results from the strong electrostatic attractions between the positively and negatively charged ions.
- When molten or in aqueous solution, ionic compounds conduct electricity since they are made of charged particles which are free to move. Ionic solids do not conduct electricity since the ions are not free to move such as when they are in the solid state.
- Ionic compounds are usually soluble in water but insoluble in organic solvents (e.g., ethanol, benzene, tetrachloromethane) although exceptions exist.

3.6.3 Writing chemical formulae of ionic substances

By the end of this section, you should be able to:

- Work out the formulae of ionic compounds from the charge on the ions. *(Metal ions limited to groups 1 and 2, aluminium, zinc, lead(II), silver, copper(II) and iron(II and III). Non-metal ions limited to groups 6 and 7. Polyatomic ions limited to carbonate, hydrogencarbonate, nitrate, sulfate, hydroxide, and ammonium.)* (3.2h)
- Work out the formulae of ionic compounds from the charge on the ions. *(Limited to copper(I), nitrite, sulfite, and phosphate.)* (3.3h)

When writing a chemical formula, you need to remember the charges formed by the ions and their valency. Here are the ions of the first 18 elements excluding groups 4 and 5.

	Group 1	Group 2	Group 3	Group 6	Group 7	Group 0
	Li^+ Na^+	Mg^{2+}	Al^{3+}	O^{2-} S^{2-}	F^- Cl^-	none
	lose 1e	lose 2e	lose 3e	gain 2e	gain 1e	
Charge	+1 ion	+2 ion	+3 ion	-2 ion	-1 ion	no ions formed
Valency	1	2	3	2	1	none

Note that:

- Hydrogen and metals lose electrons to form positive ions. The ions have the same names as the atoms.
- Non-metals form negative ions with the names ending in **-ide**.
- Normally, group 0 elements do not form ions. Their atoms already have stable outer shells, so they do not need to gain or lose electrons.

The following is a list of steps for writing a chemical formula for ionically bonded substances.

1. Look at the name of the chemical compound and write the symbols of the ions that make it up. Metals are written before non-metals.

   For example, aluminium chloride:

   aluminium ion, Al^{3+} chloride ion, Cl^-
2. Note the valencies of the ions making up the ionic compound.

   Al - valency 3 Cl - valency 1
3. The valencies need to be balanced if they are not balanced already.
   - Al - valency 3: one ion is needed
   - Cl - valency 1: three ions are needed to balance the aluminium ion
4. Therefore, for every aluminium ion, there has to be **three** chloride ions.
5. Resulting chemical formula: **$AlCl_3$**

There are other methods used to write chemical formulae such as the following:

1. Write down the name of the ionic compound.
2. Write down the symbols of its ions.
3. The compound must have an overall charge of zero, so balance the ions until the positive and negative charges add up to zero.
4. Write down the formula without charges.

Here are some examples for writing the following chemical formulae:

Sodium fluoride	**Magnesium chloride**	**Potassium oxide**
Na^+ F^-	Mg^{2+} Cl^-	K^+ O^{2-}
	Cl^-	K^+
NaF	**$MgCl_2$**	**K_2O**

1. Write down the chemical formula of the following compounds.

potassium chloride calcium sulfide magnesium fluoride lithium oxide

aluminium bromide sodium sulfide barium oxide potassium iodide

rubidium oxide barium chloride magnesium sulfide aluminium oxide

2. Write down the name of the following compounds.

$CaBr_2$ BaO LiI MgF_2

Na_2O Al_2S_3 K_2S $SrCl_2$

Most of the **transition metal ions have multiple valencies** as shown in the following table. The valency is written as a roman numeral in brackets in the chemical name without a space between the name and the bracket.

Compounds	Ion	Valency	Example
Iron compounds	**Fe^{2+}** iron(II) compounds	2	iron(II) chloride $FeCl_2$
	Fe^{3+} iron(III) compounds	3	iron(III) chloride $FeCl_3$
Copper compounds	**Cu^+** copper(I) compounds	1	copper(I) oxide Cu_2O
	Cu^{2+} copper(II) compounds	2	copper(II) oxide CuO

Note that:

- Zinc and silver do not have multiple valencies. Zinc forms **Zn^{2+}** ions and silver forms **Ag^+** ions.
- Lead is found in group 4 and commonly forms **Pb^{2+}** ions e.g. lead(II) chloride, $PbCl_2$.

Note that roman numerals are only needed when writing the names of substances containing transition metals or lead (that is metals that have variable valency). Roman numerals are not written when the metal has a fixed valency e.g., magnesium chloride. They are also not written in the formula of the compound.

Polyatomic ions

All ions that you met so far were formed from single atoms. Ions can also exist in groups of atoms and are known as polyatomic ions.

A **polyatomic ion** is a group of covalently chemically combined atoms
that is charged and exists in several compounds.

Polyatomic ions do not exist on their own but are always accompanied with oppositely charged ions. The following table contains some common polyatomic ions and their charges.

+1	-1	-2	-3
NH_4^+ Ammonium	**OH^-** Hydroxide	**CO_3^{2-}** Carbonate	**PO_4^{3-}** Phosphate
	NO_3^- Nitrate	**SO_4^{2-}** Sulfate	
	NO_2^- Nitrite	**SO_3^{2-}** Sulfite	
	HCO_3^- Hydrogencarbonate		

Let us have a look at some examples of how to write a chemical formula using polyatomic ions.

1. **copper(II) carbonate**
   Copper ions (Cu^{2+}) have a charge of 2+, that is a valency of 2.
   Carbonate ions (CO_3^{2-}) have a charge of 2-, that is a valency of 2.
   Chemical formula is **$CuCO_3$**

2. **aluminium hydroxide**
   Ions present: Al^{3+}, OH^-
   3 OH^- are required to balance out the charges of Al^{3+}.
   The 3 OH^- ions are grouped in brackets when writing the formula.
   Chemical formula is **$Al(OH)_3$**.

   OH^-
   Al^{3+} OH^-
   OH^-

   It would be incorrect to write $AlOH_3$ because the hydroxide ion exits as a single entity where the oxygen and hydrogen atoms have a single covalent bond between them and at the same time, they have a single negative charge.

3. **ammonium sulfate**
   Ions present: NH_4^+, SO_4^{2-}
   2 NH_4^+ ions are needed to balance out the charges of the SO_4^{2-}.
   The 2 NH_4^+ ions are grouped in brackets when writing formula.
   Chemical formula is **$(NH_4)_2SO_4$**.

   NH_4^+
   SO_4^{2-}
   NH_4^+

From the above examples note that:

- For compounds that contain transition metals such as iron and copper that have variable valencies, large Roman numerals enclosed in brackets is written after the name of the metal. This will indicate the valency the metal ion is in.
- In examples 2 and 3 we must consider the polyatomic ion as one entity and not as separate atoms. When multiple polyatomic ions of the same type exist in a compound, we use brackets to enclose the polyatomic ion and a subscript outside the bracket shows how many polyatomic ions are needed. For example, $Ca(OH)_2$, $Al_2(SO_4)_3$.

Summary

Charge	**+1**	**+2**	**+3**	**-3**	**-2**	**-1**
Valency	**1**	**2**	**3**	**3**	**2**	**1**
Ions	Li^+ lithium	Mg^{2+} magnesium	Al^{3+} aluminium	PO_4^{3-} phosphate	O^{2-} oxide	F^- fluoride
	Na^+ sodium	Ca^{2+} calcium	Fe^{3+} iron(III)		S^{2-} sulfide	Cl^- chloride
	K^+ potassium	Sr^{2+} strontium			CO_3^{2-} carbonate	Br^- bromide
	Rb^+ rubidium	Ba^{2+} barium			SO_4^{2-} sulfate	I^- iodide
	H^+ hydrogen	Cu^{2+} copper(II)			SO_3^{2-} sulfite	OH^- hydroxide
	NH_4^+ ammonium	Fe^{2+} iron(II)				NO_3^- nitrate
	Cu^+ copper(I)	Pb^{2+} lead(II)				NO_2^- nitrite
	Ag^+ silver	Zn^{2+} zinc				HCO_3^- hydrogen carbonate

Note that:

- The ending **"-*ide*"** indicates the presence of a non-metal ion in a binary compound such as chloride, Cl^- or oxide, O^{2-}.
- The ending **"-*ate*"** indicates a polyatomic ion containing oxygen, e.g. sodium hydrogencarbonate ($NaHCO_3$), magnesium sulfate ($MgSO_4$) and sodium nitrate ($NaNO_3$).
- The ending **"-*ite*"** indicates a polyatomic ion that also has oxygen in it but when there are two polyatomic ions of the same non-metal that have different number of oxygen atoms, the one that has the lowest amount gets its name ending in "-ite". The one that has the higher amount of oxygen gets its name ending in "-ate". For example,

  magnesium sulfite: $MgSO_3$ and magnesium sulfate: $MgSO_4$

  sodium nitrite: $NaNO_2$ and sodium nitrate: $NaNO_3$

1. Give the formulae of the following polyatomic ions (including their charge):
   a) Carbonate
   b) Phosphate
   c) Hydrogencarbonate
   d) Sulfate
   e) Ammonium
   f) Hydroxide
   g) Nitrite
   h) Nitrate

2. Give the names and chemical formulae of the compounds that copper and iron form with oxygen.

3. Write chemical formulae for the following ionic compounds:

sodium chloride	aluminium bromide	barium iodide	potassium oxide
iron(II) carbonate	copper(II) chloride	lead(II) sulfide	manganese(IV) oxide
silver carbonate	zinc hydroxide	lithium sulfite	iron(III) nitrate
calcium nitrite	aluminium sulfate	copper(I) oxide	ammonium phosphate
copper(I) sulfide	iron(II) hydroxide	aluminium nitrate	calcium hydrogen carbonate
iron(III) sulfate	lead(II) nitrate	silver sulfate	ammonium carbonate
aluminium oxide	rubidium oxide	barium hydroxide	potassium sulfite
zinc fluoride	iron(II) hydroxide	strontium nitrate	sodium phosphate

4. Name the following ionic compounds:

NaI	$ZnBr_2$	MgO	$BaCl_2$
$CuCl_2$	NH_4Cl	$Fe(NO_3)_2$	$Ca(OH)_2$
$AgNO_3$	$Mg(HCO_3)_2$	$PbSO_4$	$CuCl$
Fe_2O_3	$Cu(NO_3)_2$	CaF_2	$PbCO_3$
$(NH_4)_2SO_3$	K_3PO_4	$LiNO_2$	$MnCl_2$
$Mg(NO_3)_2$	Ag_2S	$NaHCO_3$	$Zn(NO_3)_2$
$CuSO_4$	Li_2SO_3	$FeCl_2$	NH_4OH

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3.6.4 Balancing chemical equations - revisited

Balancing chemical equations has already been tackled in section 2.6.3 where it was limited to equations involving covalently bonded substances. This section will deal with balancing chemical equations for different reaction.

Let’s consider the reaction between sodium metal and chlorine gas to produce the salt sodium chloride. The following table shows the steps needed to balance this equation.

1. Write a word equation.	sodium + chlorine → sodium chloride
2. Write down the correct formulae. From now onwards, chemical formulae MUST NEVER be CHANGED.	$Na + Cl_2 \rightarrow NaCl$
3. Count the number of atoms for each element on either side.	$Na + Cl_2 \rightarrow NaCl$ Na: 1 Na: 1 Cl: 2 Cl: 1
4. Note that Cl is unbalanced. Add coefficient on right hand side to balance Cl.	$Na + Cl_2 \rightarrow 2NaCl$ Na: 1 Na: 2 Cl: 2 Cl: 2
5. Note that Na is now unbalanced. Add coefficient on left hand side to balance Na.	$2Na + Cl_2 \rightarrow 2NaCl$ Na: 2 Na: 2 Cl: 2 Cl: 2

Let’s consider another reaction between aluminium and chlorine to produce aluminium chloride. The following table shows the steps needed to balance this equation.

1. Write a word equation.	aluminium + chlorine → aluminium chloride
2. Write down the correct formulae. From now onwards, chemical formulae MUST NEVER CHANGE.	$Al + Cl_2 \rightarrow AlCl_3$
3. Count the number of atoms for each element on either side.	$Al + Cl_2 \rightarrow AlCl_3$ Al: 1 Al: 1 Cl: 2 Cl: 3
4. Note that Cl is unbalanced. The least common multiple of 2 and 3 is 6. Add coefficients on left hand side and right-hand side to balance Cl.	$Al + 3Cl_2 \rightarrow 2AlCl_3$ Al: 1 Al: 2 Cl: 6 Cl: 6
5. Note that Al is now unbalanced. Add coefficient on left hand side to balance Al.	$2Al + 3Cl_2 \rightarrow 2AlCl_3$ Al: 2 Al: 2 Cl: 6 Cl: 6

Balance the following equation involving polyatomic ions.

1. Unbalanced equation.	$NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$
2. Count the number of atoms for each element on either side of the equation. If the polyatomic ion is unchanged on each side of the equation count it as a whole entity.	$NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$ Na: 1 Na: 2 O: 1 O : 1 H: 3 H: 2 SO_4: 2 SO_4: 2
3. Look for an atom/ion that appears once on each side of the equation and check whether it is balanced. Na is unbalanced. Add a coefficient on the left-hand side to balance Na.	$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$ Na: 2 Na: 2 O: 2 O : 1 H: 4 H: 2 SO_4: 2 SO_4: 2
4. Note that O and H are now unbalanced. Add coefficient on the right-hand side to balance O and H.	$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$ Na: 2 Na: 2 O: 2 O : 2 H: 4 H: 4 SO_4: 2 SO_4: 2

1. Balance the following chemical equations:

a. $Fe + S \rightarrow FeS$

b. $C + O_2 \rightarrow CO_2$

c. $C + O_2 \rightarrow CO$

d. $HCl + Ca(OH)_2 \rightarrow CaCl_2 + H_2O$

e. $H_2SO_4 + KOH \rightarrow K_2SO_4 + H_2O$

f. $Zn + HCl \rightarrow ZnCl_2 + H_2$

g. $KI + Cl_2 \rightarrow KCl + I_2$

h. $C_2H_5OH + O_2 \rightarrow CO_2 + H_2O$

i. $Pb(NO_3)_2 \rightarrow PbO + NO_2 + O_2$

j. $NH_3 + O_2 \rightarrow NO + H_2O$

2. Complete the following word equations. Write a balanced chemical equation. The first one has been done for you. Note the diatomic molecules!

a. lithium + oxygen → ______________________________

$4\,Li + O_2 \rightarrow 2Li_2O$

b. aluminium + chlorine → ______________________________

____

c. magnesium + oxygen → ______________________________

____

3. Balance the following chemical equations:

a. $H_2 + Br_2 \rightarrow HBr$

b. $Ca + O_2 \rightarrow CaO$

c. $Na + Cl_2 \rightarrow NaCl$

d. $H_2 + F_2 \rightarrow HF$

e. $CaCO_3 + HCl \rightarrow CaCl_2 + CO_2 + H_2O$

f. $H_2SO_4 + KOH \rightarrow K_2SO_4 + H_2O$

g. $Mg + N_2 \rightarrow Mg_2N_3$

h. $CaO + HCl\ (aq) \rightarrow CaCl_2 + H_2O$

i. $SO_2 + O_2 \rightarrow SO_3$

j. $Na_2CO_3 + HNO_3 \rightarrow NaNO_3 + H_2O + CO_2$

k. $Zn + AgNO_3 \rightarrow Zn(NO_3)_2 + Ag$

l. $NaNO_3 \rightarrow NaNO_2 + O_2$

m. $Fe + HCl \rightarrow FeCl_2 + H_2$

n. $CH_4 + O_2 \rightarrow CO_2 + H_2O$

4. Write down the word equations in terms of chemical symbols. Balance the equations.

- sodium hydroxide + hydrochloric acid → sodium chloride + water

  HCl

- zinc + lead(II) nitrate → lead + zinc nitrate

- iron + sulfuric acid → iron(II) sulfate + hydrogen

  H_2SO_4

- ammonium sulfate + potassium hydroxide → potassium nitrate + water + ammonia

- magnesium + sulfuric acid → magnesium sulfate + hydrogen

  H_2SO_4

- barium chloride + copper(II) sulfate → copper(II) chloride + barium sulfate

3.7 Solvent action of water, solution, and solubility

The **solubility of a solute** is the maximum quantity of solute that can dissolve in a certain quantity of solvent at a specified temperature.

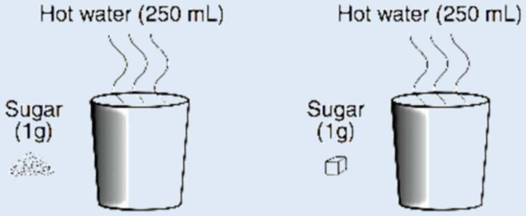
The main factors that influence solubility are:

- **The nature of the solute and solvent -** While only 1 gram of lead(II) chloride can be dissolved in 100 grams of water at room temperature, 200 grams of zinc chloride can be dissolved in the same amount of water. The great difference in the solubility of these two substances is the result of differences in their nature.
- **Temperature -** Generally, an increase in the temperature of the solution increases the solubility of a solid solute although some exceptions exist. For all gases, solubility decreases as the temperature of the solution rises.
- **Pressure -** For solid and liquid solutes, changes in pressure have practically no effect on solubility. For gaseous solutes, an increase in pressure increases solubility and a decrease in pressure, decreases solubility. (When the cap on a bottle of soft drink is removed, pressure is released, and the gaseous solute (CO_2), bubbles out of solution. This escape of gas from a solution is called **effervescence.**)

Some of the factors determining how fast a substance dissolves in a solvent are:

- **Size of the particles -** When a solute dissolves, the action takes place only at the surface of each particle. When the total surface area of the solute particles is increased, the solute dissolves more rapidly. Breaking a solute into smaller pieces increases its surface area and hence its rate of dissolving.
- **Stirring -** With liquid and solid solutes, stirring brings fresh portions of the solvent in contact with the solute, thereby increasing the rate of solution.
- **Amount of solute already dissolved -** When there is little solute already in solution, dissolving takes place relatively rapidly. As the solution approaches the point where no solute can be dissolved, dissolving takes place more slowly.
- **Temperature -** For liquids and solid solutes, increasing the temperature not only increases the amount of solute that will dissolve but also increases the rate at which the solute will dissolve. For gases, the reverse is true that is an increase in temperature decreases both solubility and rate of solution.

The following experiment was carried to find out how quickly sugar dissolves in hot water.



a. Name the variable that is being changed in this experiment (independent variable).
b. Name the variable that can be measured to find out how quickly sugar dissolved in water.
c. Name the controlled variables and explain why they were kept constant.
d. Predict a result for this experiment.
e. Design experiment to investigate another factor affecting solubility of sugar in water.

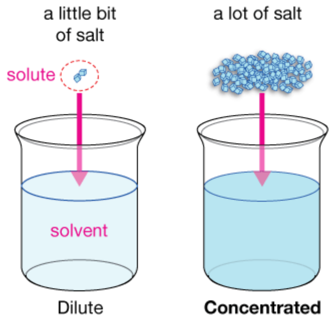
3.7.1 Dilute and concentrated solutions

By the end of this section, you should be able to:

- Distinguish between dilute, concentrated, and saturated solutions. (3.2i)

A **saturated solution** is one which cannot dissolve any more solute at a given temperature. We are often concerned with how much solute is dissolved in a given amount of solution. The concentration of a solution is described with two related and relative terms – **dilute** and **concentrated**.

- A **dilute** solution is one in which there is a relatively small amount of solute dissolved in the solution.
- A **concentrated** solution contains a relatively large amount of solute dissolved in the solution.



The above terms do not provide any quantitative information (actual numbers), but they are often useful in comparing solutions in a more general sense. These terms do not tell us whether the solution is saturated or unsaturated, or whether the solution is "strong" or "weak". These last two terms are defined in LO4 when we discuss acids and bases. Do not confuse these terms.

3.7.2 Solubility rules

By the end of this section, you should be able to:

- Predict solubility of salts in water using the solubility rules. (3.2j)

The following table shows the solubility of substances in water at room temperature taken at 25 °C. It should serve as a helpful guide to decide which substances are soluble or insoluble when constructing balanced chemical equations.

Soluble	Insoluble
• All nitrates. • All hydrogencarbonates. • All group 1 metal salts. • All ammonium salts. • Halides except silver and lead halides. • Sulfates except barium, calcium, and lead sulfates.	• Carbonates except group 1 metal and ammonium carbonate. • Metal oxides except group 1 and 2 metal oxides that react with water. • Hydroxides except group 1 metal and ammonium hydroxides.

Source: SEC06 2025 syllabus

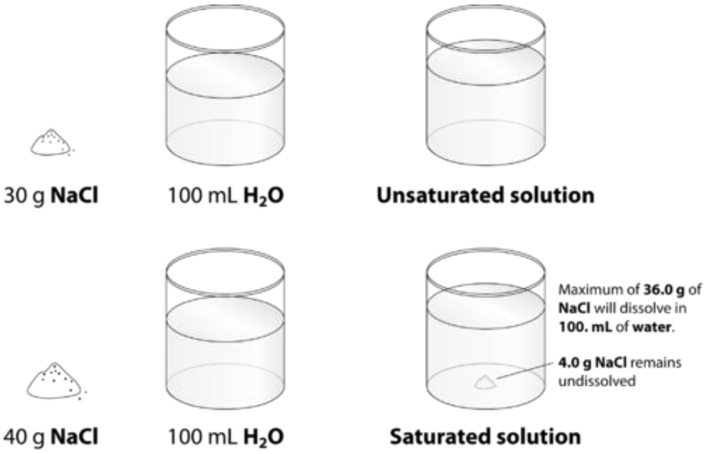
One should note that although this table might convey the idea that solubility is clean cut between being soluble or not, this cannot be further from the truth. Some substances like calcium hydroxide and calcium sulfate are **sparingly soluble** in water. This means that they are mostly insoluble but are able to dissolve slightly in water. In the table they are placed under the insoluble section. For a substance to be considered as soluble it has to be mostly soluble while for a substance to be considered as insoluble it has to be mostly insoluble!

3.7.3 Solubility of solids and gases in water

By the end of this section, you should be able to:

- Distinguish between dilute, concentrated, and saturated solutions. (3.2i)
- Interpret solubility curves of salts/gases in water. (3.3j)

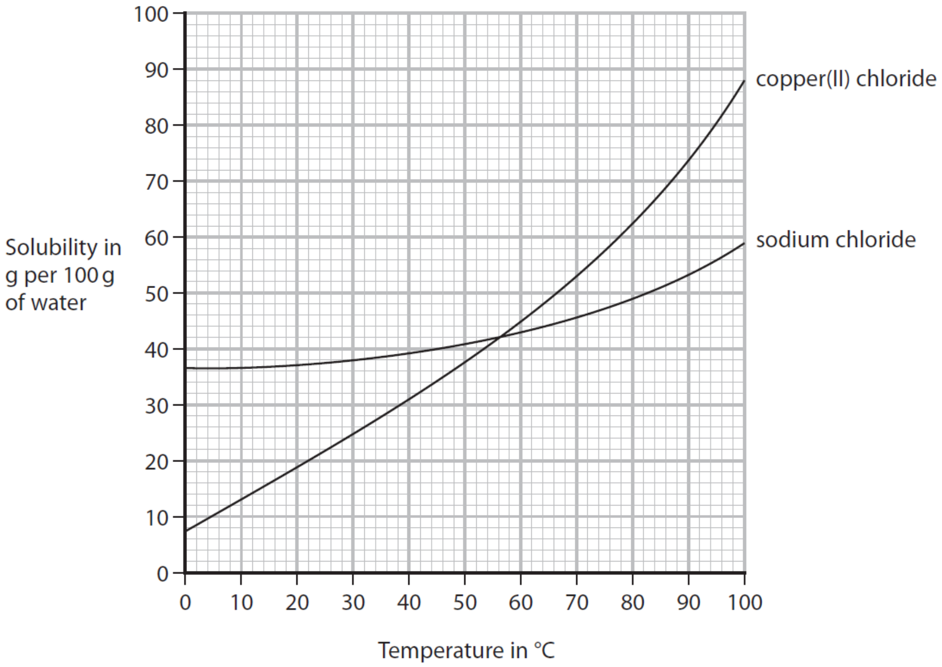
A saturated solution is one that cannot take in any more solute, and the extra solute remains undissolved. If the saturated solution is heated, more solute dissolves in it.



Source: https://www.coursehero.com/study-guides/cheminter/saturated-and-unsaturated-solutions/

A **saturated solution** is a solution in which no more solute will dissolve at a given temperature.

The solubility of a substance in a solvent is usually shown graphically by plotting the amount of solute in grams that dissolves in 100 g of solvent against temperature. The graph is known as a **solubility curve** and shows how the solubility of a substance varies with temperature. The graph below shows the solubility of copper(II) chloride and sodium chloride at different temperatures.



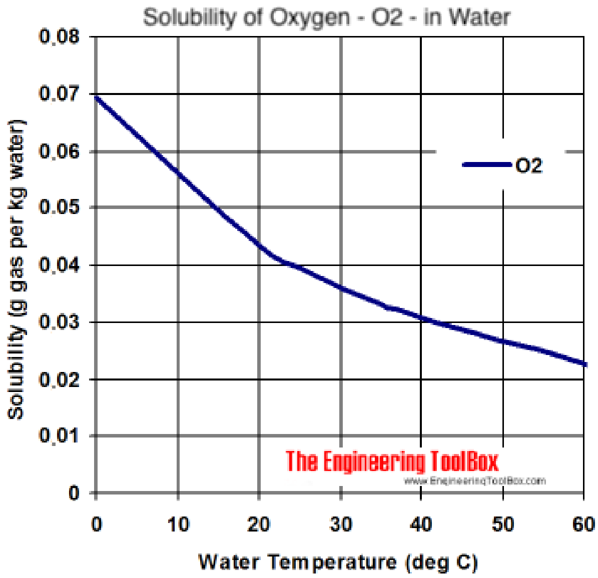
Source: Edexcel

Note that:

- Solubility is a physical property of a substance. It is defined as the mass of solute in grams dissolved per 100 g of solvent.
- The solubility of copper(II) chloride increases rapidly, whereas the solubility of sodium chloride increase steadily on increasing temperature.
- Each point on the curve represents a saturated solution at a given temperature. Any point below the curve represents an unsaturated solution for that solute. Example at 70 °C, 46 g of sodium chloride dissolve in 100 g of water and 53 g of copper(II) chloride dissolve in 100 g of water. These are saturated solutions.
- When a solution is cooled the crystals become less soluble in water. This implies that the solute will come out of the solution and will crystallise. For example, when a solution of copper(II) chloride is cooled from 70 °C to 21 °C then 33 g (53 g - 20 g) should crystallise from the solution.
- Generally, the **solubility of most solids increases as temperature rises**. At higher temperatures more energy is available for breaking of bonds in the crystals. This enables the solvent action of water to be more effective.

With reference to how salts crystallise out of solution when they cool, in real life situations the quantity of salt that crystallises is less than that expected as the solution then forms a **super saturated** solution. This solution is used to grow crystals in it.

On the other hand, the **solubility of gases decreases with increasing temperature** as shown in the graph below.



Increasing the temperature causes the kinetic energy of the gaseous molecules to increase. The higher kinetic energy causes the molecules to move about faster breaking intermolecular bonds and escaping from solution.

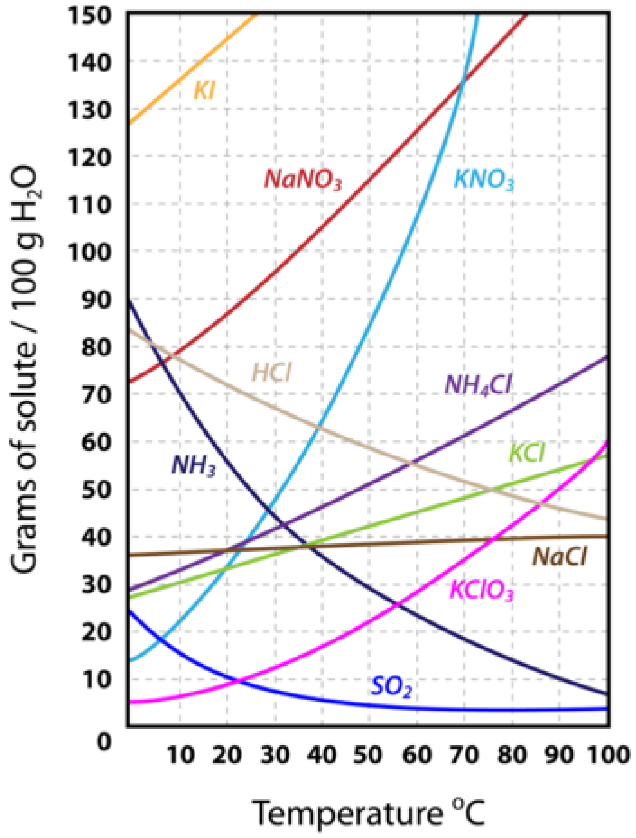
Increasing the temperature of sea water due to global warming implies that less oxygen will be available for marine ecosystems, and this can have a devastating effect.

Read the article “Why are our oceans getting warmer?” published online by National Geographic.

Different salts have different solubilities in water. The solubility of several substances is plotted on the same axis for comparison.

As a general rule, one notes that the solubility of solids increases with increasing temperature. On the other hand, the solubility of substances in the gaseous phase such as ammonia and sulfur dioxide shown in the diagram below, decreases with increasing temperature.

Note that the solubility shown by the graph represents the point at which 100 g of water cannot take in any more solute at a particular temperature and so all points on the graph represent saturated solutions!



Solubility curves for several compounds. From the CK-12 Foundation -- Christopher Auyeung

By using the solubility curve above, find:

a. The solubility of sodium nitrate at 35 °C.

b. The solubility of ammonia at 65 °C.

c. The mass of potassium nitrate that dissolves in 200 g of water at 50 °C.

d. The mass of hydrogen chloride that dissolves in 50 g of water at 45 °C.

e. The temperature at which sodium nitrate and potassium nitrate have the same solubility.

f. The mass of ammonium chloride that crystallizes from a 100 cm^3 saturated solution which is cooled from 70 °C to 30 °C.

g. The mass of ammonia that escapes from a 50cm^3 saturated solution which is heated from 10 °C to 90 °C.

h. Will a saturated solution be obtained if 30 g of sodium chloride is added to 100 cm^3 of water at 20 °C? Explain.

3.8 Hard and soft water

By the end of this section, you should be able to:

- Distinguish between hard and soft water using simple chemical tests. *(Limited to: Lathering of soap.)* (3.1k)
- Explain the difference between hard and soft water. (3.2k)
- Investigate the differences between hard and soft water. *(E.g. Using soap solution and boiling water.)* (3.3k)
- Describe the risks and benefits of hard water including issues of health and economics. *(E.g. the need of calcium by the body, clogging of hot water pipes and limescale on electric heating elements.)* (3.1l)
- Describe where hardness, both temporary and permanent, and limescale come from. *(With reference to groundwater.)* (3.2l)
- Explain, using chemical reactions, where hardness, both temporary and permanent, and limescale come from. *(With reference to groundwater.)* (3.3l)
- State why water softening is important in hard water areas. (3.1m)
- Describe different methods for removing water hardness. *(Using ion exchange resin, boiling water, distillation, and addition of washing soda.)* (3.2m)
- Explain, using chemical equations where appropriate, the effectiveness of different methods for removing water hardness. *(Using ion exchange resin, boiling water, distillation, and addition of washing soda.)* (3.3m)

In section 3.2.1 we said that ground water is the only natural resource of potable water in Malta. Ground water is collected when rainwater seeps through the ground collecting into underground reservoirs called water tables. In areas where sedimentary rocks are present, the rocks containing limestone (calcium carbonate) will react with the naturally slightly acidic rainwater resulting in dissolved calcium hydrogencarbonate. Such water is called **hard water.** Ground water also contains any other soluble salts that may be present in the bedrock as explained in the next section.

3.8.1 Water hardness

Water is said to be **hard water** when it **contains dissolved calcium and/or magnesium ions** (irrespective of any other ions present in water). Similarly, if water does not have calcium and/or magnesium ions dissolved in it, it is said to be **soft water**.

Rainwater collects various substances as it falls through the atmosphere. Clean rainwater combines with carbon dioxide to form carbonic acid. This makes rainwater slightly acidic with a pH in the range between 5 – 6.

$$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$

Although CO_2 makes rainwater slightly acidic, rainwater is not considered to be acid rain.

Many rocks in different parts of the world contain limestone (calcium carbonate) or dolomite (magnesium carbonate). Carbonic acid, reacts with these rocks, very slowly dissolving them as it passes through and over them.

The dissolved substances are **calcium** and **magnesium hydrogencarbonates**.

$$CaCO_3(s) + H_2CO_3(aq) \rightleftharpoons Ca(HCO_3)_2(aq)$$

$$MgCO_3(s) + H_2CO_3(aq) \rightleftharpoons Mg(HCO_3)_2(aq)$$

Rocks may also contain gypsum (calcium sulfate, $CaSO_4.2H_2O$). These crystals can be found in clay slopes all over Malta and Gozo. **Calcium sulfate** is sparingly soluble in water. **Magnesium sulfate**, also called Epsom salt, is also sparingly soluble in water and is found in the groundwater of many countries.


Calcium sulfate crystals found in a local clay slope.

All of the salts mentioned, i.e. calcium hydrogen carbonate, magnesium hydrogen carbonate, calcium sulfate and magnesium sulfate may all be present in ground water.

There are two kinds of hardness in water:

Temporary Hardness	**Permanent hardness**
Caused by the presence of dissolved calcium (or magnesium) hydrogen carbonate, $Ca(HCO_3)_2$ (aq) and $Mg(HCO_3)_2$ (aq).	Caused by the presence of dissolved calcium (or magnesium) ions, such as $CaSO_4$(aq) and $MgSO_4$(aq), except hydrogencarbonates.
Calcium (or magnesium) hydrogen carbonate is formed by the reaction of acidified rainwater with limestone or dolomite.	Water percolating through rock containing gypsum dissolves calcium (or magnesium) sulfate.
Can be removed by boiling because calcium (or magnesium) hydrogen carbonate thermally decomposes.	Cannot be removed by boiling but can be removed by other methods.

3.8.2 Softening of hard water

The basic principle involved in softening hard water is the removal of the calcium or magnesium ions from water. This can be done in one of the following ways.

3.8.2.1 Boiling

Only **temporary hardness** can be removed by boiling.

When heated, calcium and magnesium hydrogen carbonate thermally decompose producing calcium and magnesium carbonate respectively which are insoluble, thereby removing the calcium and magnesium ions from solution, rendering the water soft.

The layer of insoluble substance formed on containers used to boil water is commonly known as **scale** as shown in the adjacent picture.

If the scale is formed due to calcium carbonate, the term used is **limescale**.

$$Ca(HCO_3)_2(aq) \rightarrow CaCO_3(s) + H_2O(l) + CO_2(g)$$

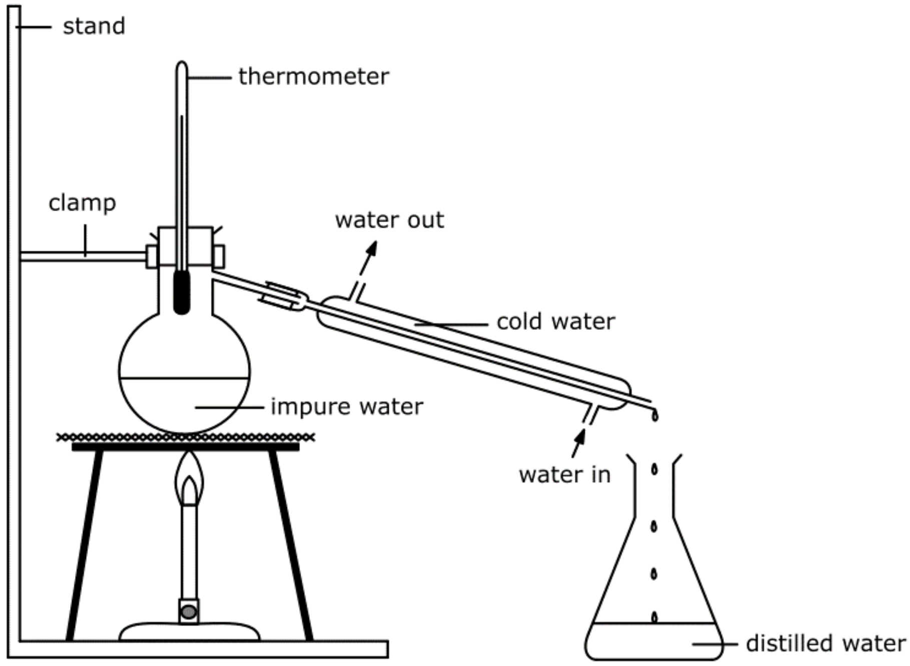
$$Mg(HCO_3)_2(aq) \rightarrow MgCO_3(s) + H_2O(l) + CO_2(g)$$

Permanent hardness cannot be removed by boiling.

3.8.2.2 Distillation

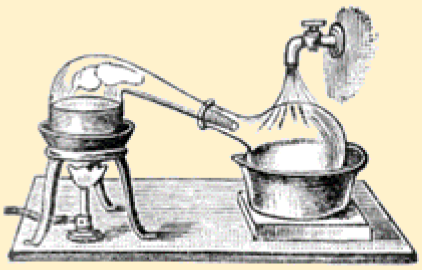
Distillation is a separating technique used to separate the solvent (in this case water) from any dissolved solutes. It is used to remove all soluble ions to produce **pure water**. Both types of hardness can be removed by distillation.

In the laboratory, distillation is used to obtain the solvent from a mixture of a **soluble solid** and the **solvent**. The mixture is heated in a round bottomed flask and the solvent evaporates. Note that a round bottomed flask is used whenever a flask is going to be heated to prevent breakage. Anti bumping granules are placed in the round bottom flask to prevent vigorous boiling. The thermometer is used to monitor the temperature of the vapour. At standard pressure, pure water distils at 100 °C. The condenser provides a cool environment by having tap water circulating through it. The solvent vapour is cooled in the Liebig condenser and is collected as the **distillate**.



Distillation

Distillation is one of the oldest known separation techniques. Its history can be traced back to ancient civilizations. The earliest evidence of distillation dates to around 3000 BCE in ancient Mesopotamia, where clay distillation pots were used to produce perfumes and fragrances through the distillation of plant materials.



The ancient Egyptians practiced distillation as early as 1500 BCE, primarily for extracting essential oils and perfumes from plants and flowers. In China, distillation is thought to have been used around the same time, with evidence of alcohol distillation dating back to the 9th century BCE. The knowledge of distillation techniques gradually spread to various regions, including the Middle East and Europe, during the Middle Ages. Arab scholars and alchemists made significant contributions to the development and understanding of distillation processes.

3.8.2.3 Washing soda

Both temporary and permanent hardness are removed by adding washing soda ($Na_2CO_3.10H_2O$) to the hard water. Permanent hardness is caused by calcium or magnesium sulfates. These substances cannot be decomposed by heating so boiling cannot remove permanent hardness.

The addition of sodium carbonate would remove such hardness because the carbonate ions from the washing soda react with the calcium/magnesium ions to form a precipitate.

For permanent hardness:

$$CaSO_4(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + Na_2SO_4(aq)$$

For temporary hardness:

$$Ca(HCO_3)_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + 2NaHCO_3(aq)$$

In both instances the dissolved calcium ions (which is the cause for hardness) are removed as a precipitate (i.e., an insoluble calcium salt). This is a precipitation reaction.

Removing water hardness

3.8.2.4 Ion-exchange

This method can be used to remove both types of hardness. Water is passed through a container filled with a suitable resin in the form of small granules. This resin contains sodium ions, and when hard water is passed through the resin, the calcium (or magnesium) ions are exchanged with the sodium ion.

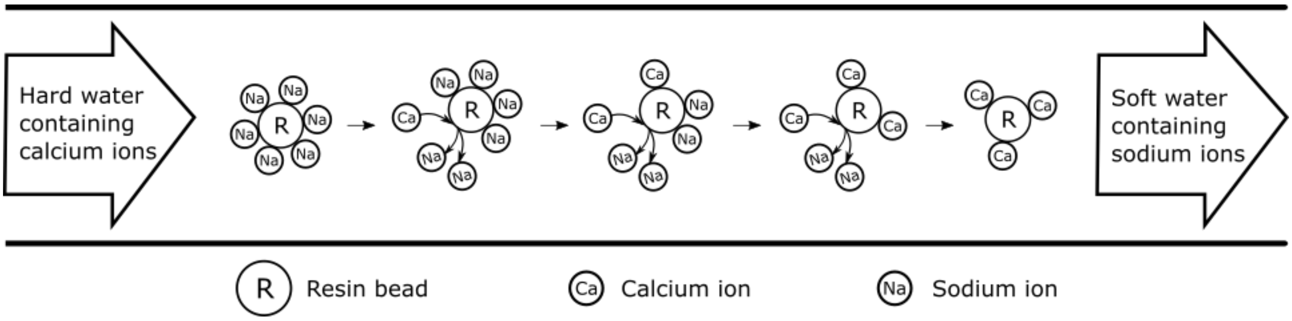
For temporary hardness:

$$Ca(HCO_3)_2(aq) + Na_2\text{-resin}(s) \rightarrow Ca\text{-resin}(s) + 2NaHCO_3(aq)$$

For permanent hardness:

$$CaSO_4(aq) + Na_2\text{-resin}(s) \rightarrow Ca\text{-resin}(s) + Na_2SO_4(aq)$$

When all the Na^+ ions have been removed from the resin, the resin can be chemically recharged by placing it in a concentrated solution of NaCl.



1. Name the ions present in water that makes it hard.
2. Explain the differences between temporary and permanent hardness.
3. Hardness in water may be removed by several means. Discuss.

3.8.3 Testing for water hardness

One method to show the difference between hard and soft water is to add soap to it.

Soft water readily forms a lather with soap.

Hard water does not form a lather readily but forms **scum**.

Soaps consist of the soluble sodium (or potassium) salts of a few organic acids, e.g., stearic acid (octadecanoic acid) and palmitic acid (hexadecanoic acid). Ordinary soap is a mixture of sodium stearate, sodium palmitate and sodium oleate; "soft soaps" have potassium salts instead of sodium salts. Since the formulae of these salts are so complicated, e.g., $C_{17}H_{35}COONa$ (sodium stearate), we use St for the stearate (palmitate or oleate) polyatomic ion such that the formula is written as: **NaSt**

Soaps readily dissolve in water such that the acid anions and sodium (or potassium) cations separate and spread out:

$$\underset{\text{Solid soap}}{NaSt(s)} \rightarrow \underset{\text{Soap solution}}{Na^+(aq) + St^-(aq)}$$

If soap is used on **soft water**, a good lather is formed almost immediately because the stearate part of the soap dissolves in water to form a soapy solution.

If on the other hand, the **water is hard**, the lather does not form immediately. In fact, a larger amount of soap has to be used if any lather is to form. When soap is added to hard water, the water becomes cloudy and a layer of solid material (a precipitate called **scum**) forms on the surface of the water. This precipitate forms because of the reaction of the dissolved calcium/magnesium ions in the hard water with soap.

$$\underset{\text{soap}}{2NaSt(aq)} + Ca(HCO_3)_2(aq) \rightarrow \underset{\text{scum}}{Ca(St)_2(s)} + 2NaHCO_3(aq)$$

This reaction is an example of a precipitation reaction because a solid material is obtained by mixing two soluble solutions.

An experiment is carried to estimate the amount of water hardness in different samples of water. Soap is added a little at a time (e.g., 1 cm^3) to a known volume of water. The water is shaken after each addition to see if a permanent lather forms which lasts for about 1 minute. The amount of soap required to just produce a lather can be used to estimate the hardness in water.

Testing for water hardness with soap solution.

Go-Lab: Cleaning capacity of soap with hard and soft water.

Maryanne tested 3 samples of water with soap solution. She recorded how much soap was needed to get a permanent lather before and after boiling the samples of water.

	Soap solution added (cm^3)	
Water sample	Before boiling	After boiling
Sample A	8	7
Sample B	1	1
Sample C	7	2

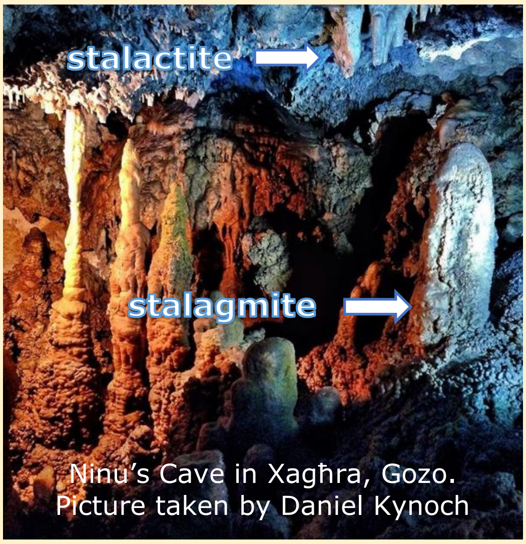
a) Determine which sample could be distilled water. Explain your answer.
b) Which sample contains temporary hard water? Explain your answer. Also support your answer with a balanced chemical equation.
c) Describe a method that can be used to remove permanent hardness.
d) Explain the chemical difference between permanent and temporary hard water.
e) Explain why boiling is not a good idea to remove temporary hardness from water on a large scale.
f) Describe how can you remove limescale from hot water pipes or kettles.

Stalagmites and Stalactites

Caves are frequently found in limestone areas. As water trickles down through the ground to the roof of a cave, a chemical reaction produces a dilute solution of calcium hydrogen carbonate as shown below:

$CaCO_3(s) + H_2O(l) + CO_2(g) \rightleftharpoons Ca(HCO_3)_2(aq)$

As drops of water containing calcium hydrogen carbonate hang from the cave roof, the water evaporates slowly, and the calcium hydrogen carbonate decomposes to give calcium carbonate which is deposited on the roof. Over hundreds of years, these deposits form stalactites.



Ninu's Cave in Xagħra, Gozo.
Picture taken by Daniel Kynoch

Sometimes the drops of water fall to the ground before they evaporate. They then evaporate later and again form calcium carbonate deposits on the ground just under the deposits on the roof. These form stalagmites.

This reaction is a reversible reaction as shown below:

$Ca(HCO_3)_2(aq) \rightleftharpoons CaCO_3(s) + H_2O(l) + CO_2(g)$

3.8.4 Advantages and disadvantages of hard water

The following are **advantages** of hard water:

- The calcium salts in the water help form healthy teeth and bones.
- Calcium carbonate is important for shell and egg formation in many animals.

In the past when lead pipes were used, as soft water flowed through lead pipes, some lead which is poisonous would dissolve and be consumed by humans. Hard water forms a layer of lead(II) carbonate and sulfate on the pipes which prevents this from happening.
More information on lead toxicity: https://mayocl.in/2OAQsuA

The following are **disadvantages** of hard water:

- Soap wastage increases because all the dissolved calcium and magnesium salts have to be precipitated before soap can start to function (by producing lather).
- The "scum" produced is difficult to remove from fabrics.
- The dissolved calcium and magnesium salts interfere with the dyeing processes.
- When hard water is boiled or evaporated, an off-white solid deposit is usually left behind (have a look at your kettle at home if you use tap water for making tea or coffee!).

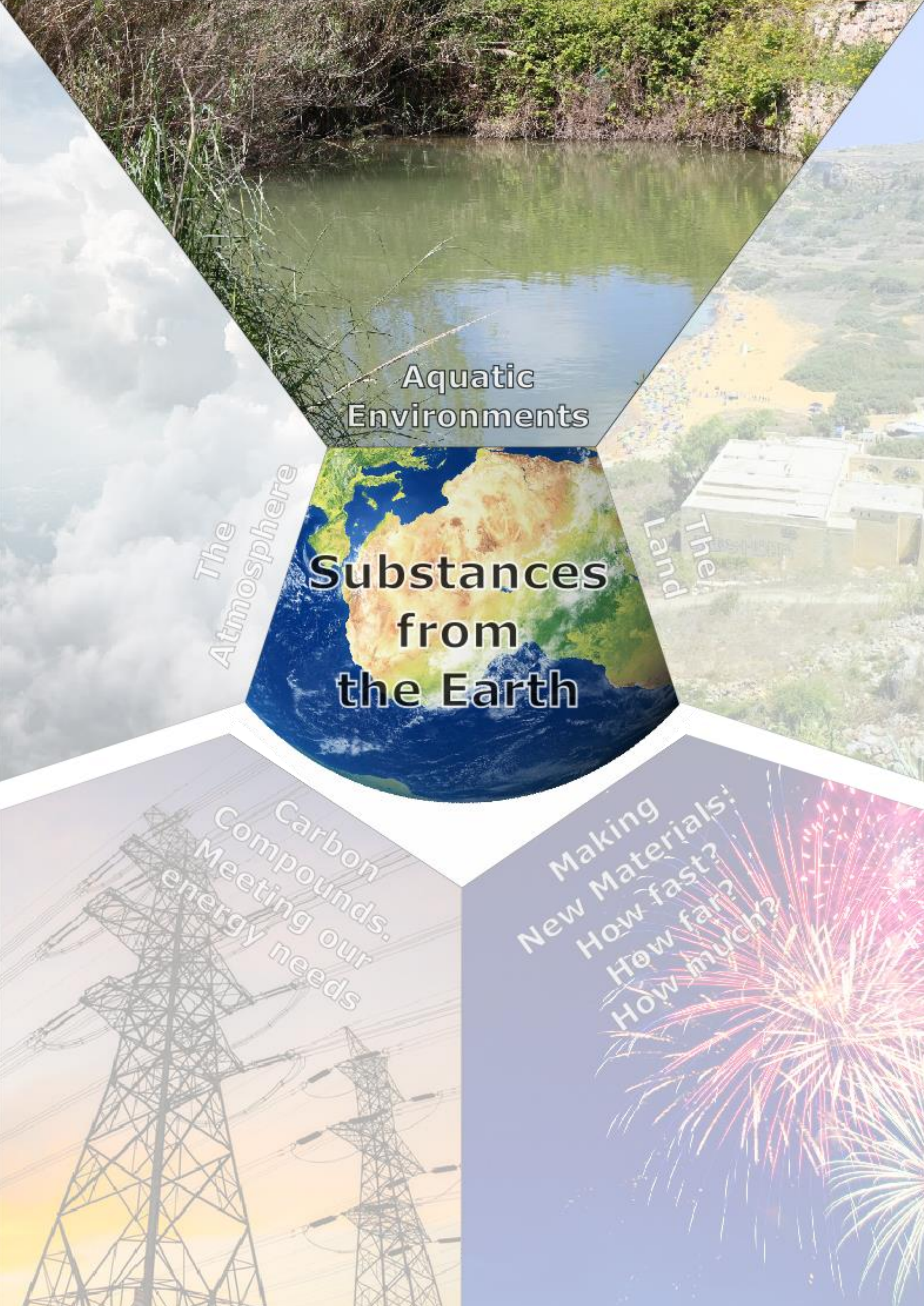
  This deposit consists of calcium carbonate originating from the thermal decomposition of calcium hydrogencarbonate that was present in the hard water.

  $$Ca(HCO_3)_2(aq) \rightarrow CaCO_3(s) + H_2O(l) + CO_2(g)$$

  This phenomenon causes "*furring*" in kettles and boilers as well as blockages in water pipes. It can also occur in household appliances that heat water such as washing machines, dishwashers, etc., blockages in hot water pipes (e.g., in steam engines, laundries, central heating systems, etc.). The formation of limescale reduces the efficiency of the heating element as more energy is needed to heat the water.
- In the lab the calcium and magnesium salts may interfere with any reaction that takes place in water.

Because of these and other disadvantages it would be advantageous if hard water could be softened to remove the soluble calcium and magnesium ions.

Detergents such as washing soap liquid, and laundry detergents are not made of sodium stearate, but are made from reacting hydrocarbons with concentrated sulfuric acid and converting them into the sodium salt. These synthetic detergents can be used with hard water and no soap is wasted. They are more soluble in water and have a stronger cleansing action than soap. Whereas soap has been used since 2500 BCE, detergents started being produced at the beginning of the 20th century because during the 1st World War there was a shortage of oils and fats to make soap. Synthetic detergents were produced as an alternative to soap.



Substances from the Earth

Aquatic Environments

The Atmosphere

The Land

Carbon Compounds. Meeting our energy needs

Making New Materials: How fast? How far? How much?

4 Acids, bases, and salts (LO 4)

Many **acids** exist naturally in plants and food; a sour taste usually is indicative that a substance contains an acid. The word acid comes from the Latin word "acidus" meaning sour.

Ascorbic acid (Vitamin C) is found in fruits. 	**Carbonic acid** is found in fizzy drinks. 
Citric acid is found in oranges and lemons. 	**Ethanoic acid** is found in vinegar which is used for pickling. 

- **Hydrochloric acid** is produced by the cells in the lining of our stomachs.
- **Lactic acid** is found in sour milk.
- **Nitric acid** is used for making fertilisers and explosives.
- **Phosphoric acid** is used in anti-rust paint and for making fertilisers.
- **Sulfuric acid** is used in car batteries and for making fertilisers.

Alkalis or bases are generally found in cleaning reagents. Dishwasher tablets, washing powder and cleaning liquids all contain bases.

- Oven cleaner or drain unblocker contains **sodium hydroxide** or caustic soda.
- Baking soda contains **sodium hydrogen carbonate**.
- Bases are found in indigestion tablets which may contain **magnesium hydroxide** or **calcium carbonate**.

One way of producing **salts** is by the reaction of acids and bases. Potassium chloride, magnesium sulfate and copper(II) sulfate are examples of salts.

- Salts are ionic compounds with typical high melting and boiling points.
- Most are soluble in water (refer to solubility rules).
- Solutions of soluble salts are known as electrolytes that is they conduct elelctricity but at the same time decompose.

copper(II) sulfate crystals

4.1 Indicators

By the end of this section, you should be able to:

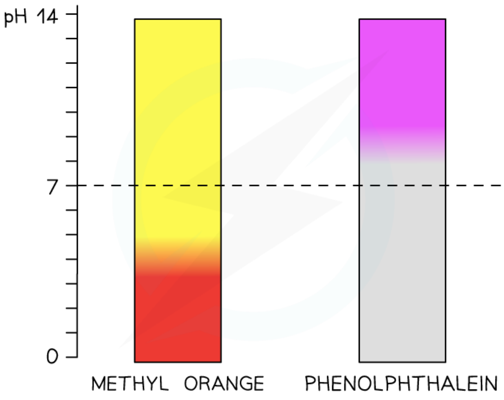
- Use indicators and the pH scale to distinguish between acidic, alkaline, and neutral solutions. *(E.g. Using litmus, universal indicator, phenolphthalein, and methyl orange indicators.)* (4.1a)
- Classify a substance as acid, base, or alkali. *(E.g., Using litmus, universal indicator, phenolphthalein and methyl orange indicators, and pH scale.)* (4.2a)

Indicators are substances which are used to detect the presence of acidic, alkaline, or neutral substances. Indicators are usually obtained from lichens and other plant material.

A colour change is obtained when alkalis and acids interact with indicators. The following table shows the colour change given by three common indicators in different solutions:

Colour change given by acids and alkalis with indicators.

Indicator	Acid	Neutral	Alkaline
Methyl Orange	pink	orange	yellow
Litmus	red	purple	blue
Phenolphthalein	colourless	colourless	pink



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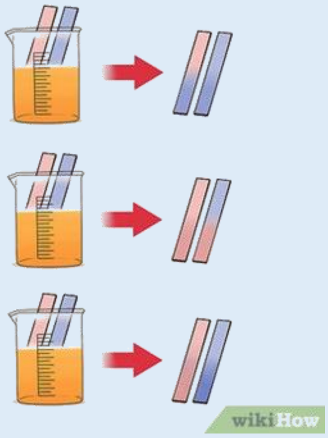
Anna tested different solutions using blue and red litmus paper. She obtained the following results.

- **Solution A** turns red litmus paper blue, but blue litmus paper is unchanged.
- **Solution B** turns blue litmus paper red, but the red litmus paper is unchanged.
- No colour change is produced when blue and red litmus papers are dipped in **solution C**.

a. Determine the nature of solution A, B and C.

b. Explain why it is incorrect to conclude that when a blue litmus paper does not change colour, the solution is alkaline.

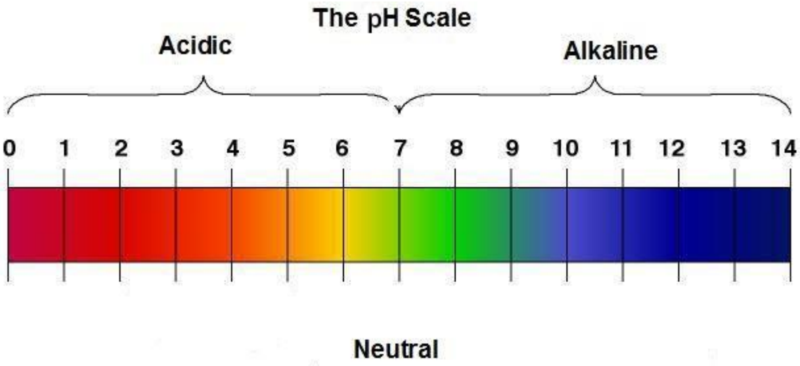
c. Hence describe how the litmus test should be conducted correctly.



Universal indicator does not give one colour change, but a series of colour changes as shown below. This demonstrates that not all acids and alkalis are of the same strength.

Universal indicator is used to indicate the **strength of the acid** or **alkali** not just whether the solution is acidic or alkaline like litmus, methyl orange and phenolphthalein.

The strength of acids and alkalis is measured using the **pH scale.**



The pH scale varies from 0-14.

- **All acids have a pH less than 7**. A pH value of 0 indicates a very strong acid and turns universal indicator red. Weak acids like ethanoic acid in vinegar (pH 4) changes it to orange. Whilst a pH of 6 indicates a weak acid and universal indicator turns yellow.
- **A solution that has a pH value of 7 is neutral**, i.e., neither acidic nor alkaline. Pure water (pH 7) makes universal indicator green.
- **All alkalis have a pH greater than 7**. Weak alkalis like soaps (pH 8-9) make universal indicator paper appear dark green, whilst stronger alkalis turn the paper blue (pH 10) or purple (pH 14).

Rough pH values can be obtained using universal indicator. More accurate pH readings are given with a digital pH meter. Note that different brands of universal indicator may show different colours for a particular pH. So, one must use the reference pH chart that is supplied with the indicator when it is bought.

The Beer Chemist and the pH Scale

The pH scale was introduced by the Danish chemist S.P.L. Sørensen (1868–1939) whilst working in a brewery. He wanted to measure the concentration of H^+ ions in solution. The amount of H^+ ions in solution depend on the strength and concentration of solution. They affect the colour of indicators. If a base is slowly added, the acid will be neutralised.

PhET Chemistry: pH scale

4.2 Acids

By the end of this section, you should be able to:
- Classify a substance as acid, base, or alkali. (4.2a part).
- Explain the difference between strong and weak acids (4.3a part)

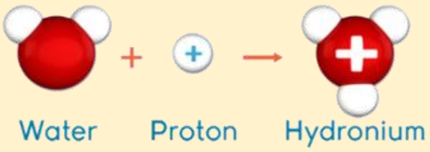
Common laboratory acids include:
- Hydrochloric acid – $HCl(aq)$
- Sulfuric acid – $H_2SO_4(aq)$
- Nitric acid – $HNO_3(aq)$

Definitions of acids have varied over the years as scientists discovered more about them.

At this level, an acid is best defined as:

An **acid** is a substance which releases hydrogen ions (H^+) when dissolved in water.

The hydrogen ion (H^+) does not exist on its own in aqueous solutions. Since a **H^+ ion** is a **proton**, it bonds to water molecules to form the hydronium ion (**H_3O^+**). This ion is also known as the hydroxonium or the oxonium ion.



At SEC level, we still refer to the hydrogen ion and represent it in equations as H^+.

4.2.1 Strong and weak acids

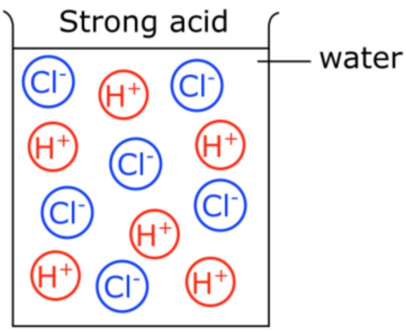
The **strength of an acid** depends on the **concentration of H^+ ions formed in solution**.

Strong acids produce a high concentration of H^+ ions because all their molecules split into ions. We say that the acid has ionised completely in water.

Examples of substances that produce strong acids include hydrochloric acid, $HCl(aq)$; nitric acid, $HNO_3(l)$ and sulfuric acid, $H_2SO_4(l)$.

$$HCl(g) \xrightarrow{\text{water}} H^+(aq) + Cl^-(aq)$$

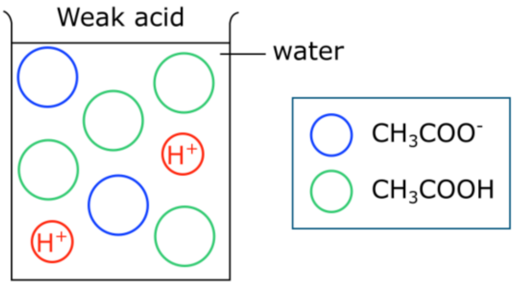
$$H_2SO_4(l) \xrightarrow{\text{water}} 2H^+(aq) + SO_4^{2-}(aq)$$



Note that when an acid is given in the liquid state this means that it is highly concentrated, while when the acid is given in the aqueous state this means that the acid is dilute.

Weak acids produce a low H^+ ion concentration because their molecules partially ionise in water, that is they do not ionise completely when in solution, e.g., ethanoic acid.

$$CH_3COOH(l) \overset{water}{\rightleftharpoons} CH_3COO^-(aq) + H^+(aq)$$



Two other weak acids include carbonic acid, H_2CO_3 and sulfurous acid, H_2SO_3.

$$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$

$$SO_2(g) + H_2O(l) \rightleftharpoons H_2SO_3(aq)$$

The following table lists some common weak and strong acids.

Weak and strong acids

Acid	Formula	Strength	Occurrence
Nitric acid	HNO_3	Strong	Found in the lab
Sulfuric acid	H_2SO_4	Strong	Found in the lab
Hydrochloric acid	HCl	Strong	Produced by the stomach
Ethanoic acid	CH_3COOH	Weak	Found in vinegar
Citric acid	$C_6H_8O_7$	Weak	Found in the juice of citrus fruits (e.g., lemons and oranges)
Carbonic acid	H_2CO_3	Weak	Found in fizzy drinks

Recognising Acids

In 1903, the Swedish chemist Svante Arrhenius (1859-1927), won the Nobel Prize for his work on ionisation. He introduced the idea of compounds ionising, or splitting, into their constituent ions in solution, for example, HCl becoming H^+ and Cl^-. He explained how the strength of an acid in water depends on the concentration of hydrogen ions in it.

4.2.2 Properties of acids

The following is a list of properties shown by acids. They:

- are sour.
- are soluble in water.
- are often corrosive.
- contain hydrogen ions $H^+(aq)$ in solution.
- change the colour of indicators.
- are neutralised by bases, alkalis, and carbonates.
- react with some metals to release hydrogen.
- react with carbonates to release carbon dioxide.

4.3 Bases and alkalis

By the end of this section, you should be able to:

- Classify a substance as acid, base, or alkali. (4.2a part).
- Explain the difference between strong and weak alkalis (4.3a part)
- Represent the reaction of an alkali with an ammonium salt using chemical equations. (4.2d)

Bases and alkalis are commonly used in our daily routines.

- Ammonia is found in cleaning liquids and used in the manufacture of fertilisers.
- Calcium hydroxide (slaked lime) is used to treat soil which is too acidic.
- Calcium oxide (lime) is used in the manufacture of cement, mortar, and concrete.
- Magnesium hydroxide is found in antacid indigestion tablets.
- Sodium hydroxide is found in oven/drain cleaners and used in the manufacture of soap.

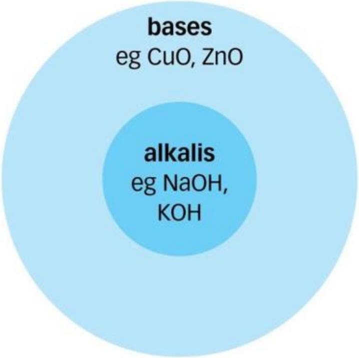
Bases are metallic oxides, hydroxides, and carbonates. Bases are **insoluble in water**.

A **base** is a substance that accepts hydrogen ions.

Soluble bases are called alkalis.

Common laboratory alkalis include.

- Sodium hydroxide – $NaOH(aq)$
- Potassium hydroxide – $KOH\ (aq)$
- Calcium hydroxide – $Ca(OH)_2(aq)$
- Ammonium hydroxide – $NH_4OH(aq)$, also referred to as ammonia solution $NH_3\ (aq)$.



The table below shows examples of insoluble metal oxides and hydroxides which are known as bases. The soluble bases are known as alkalis.

Insoluble bases	Soluble bases (alkalis)
Copper(II) oxide, CuO	Sodium hydroxide, $NaOH$
Iron(III) oxide, Fe_2O_3	Potassium hydroxide, KOH
Copper(II) hydroxide, $Cu(OH)_2$	Calcium hydroxide, $Ca(OH)_2$
Magnesium oxide, MgO	Ammonium hydroxide, NH_4OH

4.3.1 Strong and weak alkalis

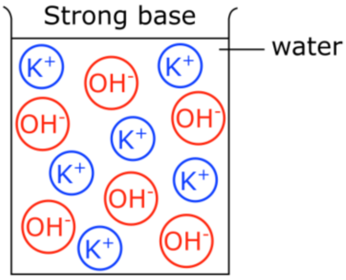
Just as there are weak and strong acids, there are weak and strong alkalis.

The strength of an alkali depends on the amount of OH^- ions in solution.

The more an alkali ionises in water the stronger it is.

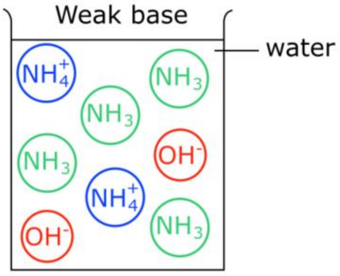
A **strong alkali** is one that ionises completely in water to form an alkaline solution.

$$KOH(s) \xrightarrow{water} K^+(aq) + OH^-(aq)$$



A **weak alkali** is one that does not ionise completely in solution.

$$NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$



On the pH scale, strong alkalis have a pH value over 11 whilst weak alkalis have a pH value between 8 and 11.

4.3.2 Properties of alkalis

An alkali is a solution of a base in water. Therefore, all soluble bases form alkaline solutions.

The following is a list of properties shown by alkalis:

- Alkalis contain hydroxide ions $OH^-(aq)$ in solution.
- Alkalis have a 'soapy' feel. This is because they react with natural oils on our skin to form soap. It is this soap that gives them their 'slippery' feel.
- Alkalis interact with indicators to bring about colour changes.
- Alkalis (also bases) neutralise acids to produce a **salt** and **water**. This type of reaction is called a **neutralisation reaction**.
- Alkalis react with ammonium salts to produce ammonia gas as in the following reaction.

alkali + ammonium salt → salt + water + ammonia

$$NaOH(aq) + NH_4Cl(s) \rightarrow NaCl(aq) + H_2O(l) + NH_3(g)$$

Test for ammonia gas: It turns damp red litmus paper blue.

4.4 Metallic and non-metallic oxides

By the end of this section, you should be able to:

- Identify carbon dioxide, sulfur dioxide and nitrogen dioxide as examples of acidic oxides. (2.2p)
- Explain how some gases react with water to produce acidic solutions. *(E.g. acidic oxides such as carbon dioxide, nitrogen dioxide and sulfur dioxide.)* (2.3p)
- Identify water and carbon monoxide as examples of neutral oxides. (2.2q)
- Identify basic oxides by their reaction with acids and the metal's position in the Periodic Table. (4.2b)
- Identify amphoteric oxides by their reaction with acids and alkalis as well as the metal's position in the Periodic Table. *(Chemical equations for their reactions with alkalis are not required.)* (4.3b)

4.4.1 Acidic and neutral oxides

Most non-metal oxides are **acidic oxides**. When acidic oxides mix with water, they react together to form acidic solutions. Examples of acidic oxides include carbon dioxide (CO_2), sulfur dioxide (SO_2) and nitrogen dioxide (NO_2). Carbon dioxide gives clean rain a slight acidic pH of around 5 – 6. Sulfur dioxide and nitrogen dioxide make rain more acidic which is why we speak of acid rain which has a pH value around 3 – 4 depending on the amount of pollution in the air.

The following reactions show how the acidic oxides mentioned, form acidic solutions. Acidic oxides form acidic solutions as can be noted by the aqueous hydrogen ions present.

Carbon dioxide reacts with water to form carbonic acid (H_2CO_3).

$$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$

Sulfur dioxide reacts with water to form sulfurous acid (H_2SO_3).

$$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$$

Nitrogen dioxide reacts with water to form two oxides, nitrous (HNO_2) and nitric (HNO_3) acid.

$$2NO_2(g) + H_2O(l) \rightarrow HNO_2(aq) + HNO_3(aq)$$

Water, carbon monoxide (CO), nitrogen monoxide (NO), and dinitrogen oxide (N_2O) are non-metal oxides that have a neutral pH, so they are said to be **neutral oxides**. They do not react with either acids or bases/alkalis.

Summary:

	Non-metallic oxides	
	Acidic oxides	**Neutral oxides**
Examples	carbon dioxide, CO_2 sulfur dioxide, SO_2 nitrogen dioxide, NO_2	water, H_2O carbon monoxide, CO nitrogen monoxide, NO dinitrogen oxide, N_2O
pH	acidic	neutral
Reactions	React with bases/ alkalis	Do not react with either acids or bases/alkalis

4.4.2 Basic and amphoteric oxides

Most metallic oxides are **basic oxides**, e.g. magnesium oxide (MgO) and copper(II) oxide (CuO). Oxides do not dissolve in water to form solutions.

There are some basic oxides, like sodium oxide (Na_2O), potassium oxide (K_2O) and calcium oxide (CaO) that react with water to form alkalis as shown by the following chemical reactions:

$$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$$

$$K_2O(s) + H_2O(l) \rightarrow 2KOH(aq)$$

$$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$$

Sodium and potassium hydroxide solutions are strong alkalis with a pH of 14.

Calcium hydroxide is partially soluble in water, and it is known as lime water. It has a pH of 12.

Some metallic oxides are known to be **amphoteric oxides**. These include aluminium oxide (Al_2O_3), lead(II) oxide (PbO) and zinc oxide (ZnO). These metal oxides **react with both acids and alkalis**. Amphoteric oxides are formed by metals that are close to the non-metal section of the Periodic Table.

Summary:

	Metallic oxides	
	Basic oxides	**Amphoteric oxides**
Examples	calcium oxide, CaO magnesium oxide, MgO copper(II) oxide, CuO	aluminum oxide, Al_2O_3 lead(II) oxide, PbO zinc oxide, ZnO
Reactions	react with acids	react with both acids and alkalis

The following equations show aluminium oxide (an amphoteric oxide) reacting with an acid and an alkali:

$$Al_2O_3(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2O(l)$$

$$Al_2O_3(s) + 2NaOH(aq) + 3H_2O(l) \rightarrow 2NaAl(OH)_4(aq) \text{ *}$$

* You do not need to know this equation for your SEC or annual exam.

Amphoteric oxides.

Classify the following oxides as acidic, basic, neutral, or amphoteric.

carbon dioxide
carbon monoxide
nitrogen monoxide
lead(II) oxide
copper(II) oxide
aluminium oxide
nitrogen dioxide
sulfur dioxide
water

Go-Lab: Properties of acids and bases

4.5 Acid reactions

By the end of this section, you should be able to:

- Represent reactions of non-oxidising acids with bases/alkalis, carbonates/ hydrogencarbonates, and fairly reactive metals, using chemical equations. (4.2c)

4.5.1 Reactions of acids with bases

Non-oxidising acids include dilute hydrochloric and sulfuric acid. Nitric acid (dilute or concentrated) and concentrated sulfuric acid act as oxidising agents and are not used as exemplars. The concept of oxidising agents will be dealt with in Year 10.

a. Dilute acids react with alkalis to form a salt and water. This is also called a **neutralisation** reaction because the acid and alkali neutralise each other.

alkali + acid → salt + water

$$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$$

b. Dilute acids react with metal oxides / hydroxides to form a salt and water. This is a **neutralisation** reaction as well. Metal oxides / hydroxides that react with acids only, are called **basic oxides / hydroxides**.

basic oxide + acid → salt + water

$$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$$

basic hydroxide + acid → salt + water

$$Cu(OH)_2(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + 2H_2O(l)$$

c. Dilute acids react with metal carbonates and hydrogencarbonates to form a salt, carbon dioxide and water.

carbonate + acid → salt + carbon dioxide + water

$$PbCO_3(s) + 2HCl(aq) \rightarrow PbCl_2(aq) + CO_2(g) + H_2O(l)$$

hydrogencarbonate + acid → salt + carbon dioxide + water

$$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

Cake baking

Baking powder is normally made of three different parts:

an acid, **a base, and a filler**.

All three need to be **dry** powders that can be mixed.

For example, baking soda (a base), cream of tartar (an acid) and corn starch (the filler) are three common ingredients.

When you add water to baking powder, the dry acid and base go into solution and start reacting to produce **carbon dioxide** bubbles.

1. **Single-acting** baking powder produces all its bubbles when it gets wet.
2. **Double-acting** baking powder produces bubbles again when it gets hot.

4.5.2 Reactions of acids with metals

Dilute acids react with reactive metals to release hydrogen gas and form a salt.

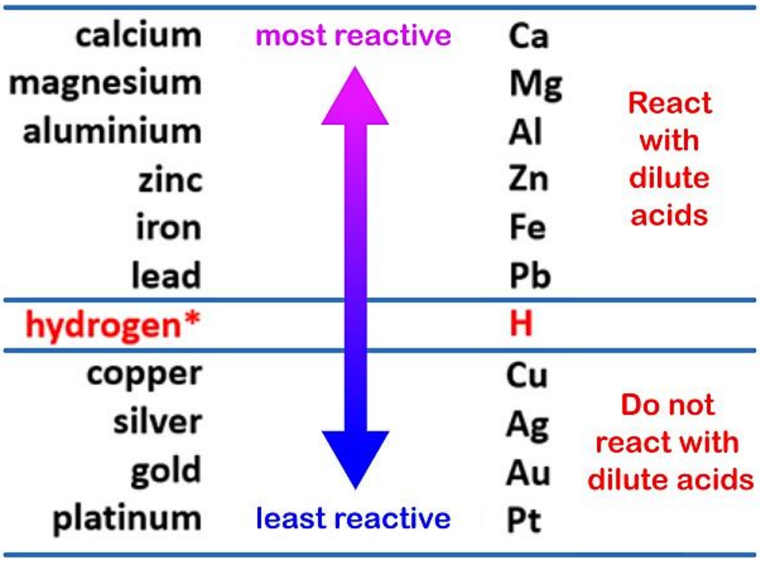
metal + acid → salt + hydrogen

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

Calcium, magnesium, aluminium, iron, and zinc react readily with dilute acids. The oxide layer which covers these metals needs to be removed for them to react.

One must note that:

- Highly reactive metals such as potassium and sodium must **never** be added to acids as they undergo explosive reactions which pose a health hazard and so are not included in the reactivity series below.
- Copper, silver, and platinum do not react with dilute acids. This does not apply to nitric acid.



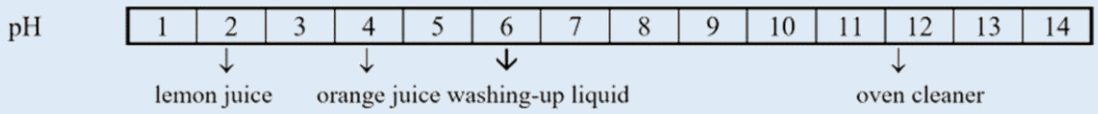
THE REACTIVITY SERIES OF METALS:

- **Reactivity** of metals is described as ***how easily a metal atom loses electrons*** (or how quickly it donates electrons), normally from their outermost shell. ***Metals high in the series are more reactive*** because they are able to lose their valence electrons more easily than metals which are further down the series to form positive ions.
- Also, only metals ***more reactive than hydrogen*** can displace it from dilute acids.

* Hydrogen is not a metal and is included in this list as a reference point.

Note that the reactivity series shown above is **not examinable** in Year 9. It will be discussed further in Year 10.

1. The pH of chemicals found around the home may be tested using pH paper. Some typical results are shown below.



   a. Answer the following questions using the above information.
      i. What is the pH of oven cleaner?
      ii. Which is the most acidic solution?
      iii. What would be the pH of a neutral solution?
   b. When a wasp stings someone it injects a liquid in the person's skin. An old-fashioned remedy for wasp stings is to rub vinegar on them.
      i. Suggest the pH of the liquid which wasps inject into the skin.
      ii. What name is given to the reaction which happens between this liquid and the vinegar?
      iii. Name the colourless, tasteless liquid that is produced in this reaction.

2. This question is about acids.
   a. Define an acid.
   b. What is the difference between a strong and a weak acid?
   c. Give an example of a strong and a weak acid.

3. What is the difference between a base and an alkali?
4. Describe a simple experiment to show that sodium hydroxide and zinc oxide are basic.
5. Acids undergo different reactions. Complete the reactions below by writing a full balanced chemical equation including state symbols:
   a. $HCl\ (aq)\ +\ KOH\ (aq)\ \rightarrow$
   b. $H_2SO_4\ (aq)\ +\ CuO\ (s)\ \rightarrow$
   c. $HNO_3\ (aq)\ +\ CaCO_3\ (s)\ \rightarrow$
   d. $HCl\ (aq)\ +\ NaHCO_3\ (s)\ \rightarrow$
   e. $H_2SO_4\ (aq)\ +\ Zn\ (s)\ \rightarrow$

6. Hydrogen is produced when an acid is added to a metal.
   a. Describe what is observed when magnesium is added to dilute hydrochloric acid.
   b. Write a balanced chemical equation for the reaction.
   c. Describe the test used to identify hydrogen gas.
   d. Explain why copper does not react with dilute hydrochloric acid.

7. Carbon dioxide is produced when sulfuric acid is added to potassium hydrogen carbonate.
   a. Write a balanced chemical equation for the reaction.
   b. Describe the test used to identify the presence of carbon dioxide.
   c. Write a balanced chemical equation for the test of carbon dioxide.

8. Ammonia is produced when an alkali reacts with an ammonium salt.
   a. Write a balanced chemical equation for the reaction of ammonium sulfate and sodium hydroxide.
   b. Describe a chemical test that is used to identify the presence of ammonia.

4.6 Applications of acid-base concepts in real-life situations

By the end of this section, you should be able to:

- Apply acid-base concepts to the real world. *(E.g. In terms of solutions to environmental issues such as acid rain, neutralisation of acid soils and excess stomach acidity.)* (4.2f)
- Investigate acid-base concepts in real-life applications. (4.3f)

4.6.1 Acid rain

Read the following article entitled "Acid Rain" at Acid Rain - Young Reporters for the Environment (yremalta.org).

> "Acid rain causes several environmental problems. Acidic gases are produced when fossil fuels, like coal and oil are burned in power stations, factories and in our own homes.
>
> These gases are blown in the sky and mixed with droplets in the atmosphere creating weak solutions of nitric and sulfuric acids. When precipitation occurs, these solutions fall as acid rain.
>
> Acid rain poisons fish in lakes, it damages the leaves of trees, stopping them photosynthesising. Ancient statues or buildings that contain calcium carbonate and are renowned for their architectural beauty, all show signs of corrosion.
>
> We can prevent acid rain by conserving energy. The less electricity we use in our homes, the fewer chemicals, power plants emit. Vehicles are also major fossil fuel users so we can walk, ride a bike, or use public transportation.
>
> Reducing pollution and controlling emissions can help to fight acid rain for a better environment!"

1. Which gases contribute towards acid rain?
2. Name one gas that acidifies rainwater slightly, but it is not responsible for acid rain.
3. This article mentions human activity as a source of acid rain. Name one natural source of acidic gases that causes acid rain as well.
4. Fossil fuels are made of a mixture of hydrocarbons which contain hydrogen and carbon only. Describe how sulfur dioxide and oxides of nitrogen are produced when fossil fuels are burned.
5. Give a balanced chemical equation for the interaction of acid rain on limestone.
6. The article suggests several solutions to reduce acid rain. State what you, as a student who does not own a car, can do to contribute towards reducing acid rain.

4.6.2 Neutralisation of acid soils

Read the following article entitled "Soil Acidity" at Soil Acidity | Fact Sheets | soilquality.org.au. Here is a summary of the main points of the article.

Soil pH is a measure of the concentration of hydrogen ions in the soil solution. The lower the pH of soil, the greater the acidity. A soil with a pH of 4 has 10 times more acid than a soil with a pH of 5 and 100 times more acid than a soil with a pH of 6.

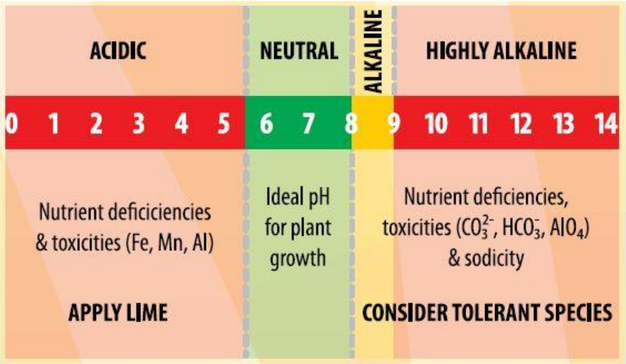
pH should be maintained at above 5.5 in the topsoil and 4.8 in the subsurface.

Acidic soils cause significant losses in production and where the choice of crops is restricted to acid tolerant species and varieties, profitable market opportunities may be reduced. In pastures grown on acidic soils, production will be reduced, and some legume species may fail to persist because acidic soil, particularly in the subsurface, will also restrict root access to water and nutrients.

Soil acidification is a natural process accelerated by agriculture. Soil acidifies because the concentration of hydrogen ions in the soil increases. The main cause of soil acidification is inefficient use of nitrogen, followed by the export of alkalinity in produce. Ammonium based fertilisers are major contributors to soil acidification. Ammonium nitrogen is readily converted to nitrate and hydrogen ions in the soil. If nitrate is not taken-up by plants, it can leach away from the root zone leaving behind hydrogen ions thereby increasing soil acidity. Most plant material is slightly alkaline and removal by grazing, or harvest leaves residual hydrogen ions in the soil. Over time, as this process is repeated, the soil becomes acidic. Major contributors are hay, especially lucerne hay and legume crops. Alkalinity removed in animal products is low, however, concentration of dung in stock camps adds to the total alkalinity exported in animal production.

A well-maintained soil pH will maintain the value of the soil resource, maximize crop, and pasture choice and avoid production losses due to low pH.

Agricultural lime may need to be applied to maintain pH, or to recover pH to an appropriate level. Lime sand, from coastal dunes, crushed limestone and dolomitic limestone are the main sources of agricultural lime. Carbonate from calcium carbonate and magnesium carbonate is the component in all these sources that neutralises acid in soil.



If the topsoil pH is above 5.5 and the subsurface pH above 4.8, only maintenance levels of liming will be required to counter on-going acidification caused by productive agriculture.

Forensic Science

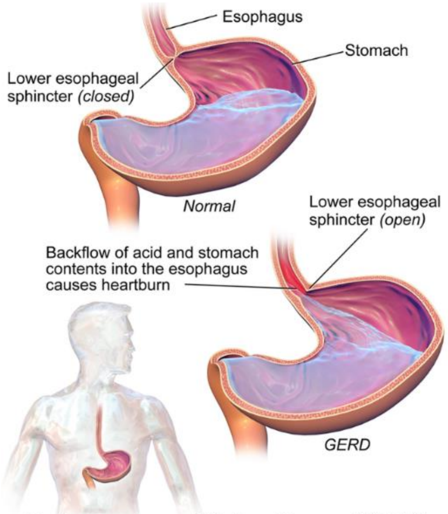
Chemistry is often used by forensic scientists to solve a crime. The pH of soil samples from a suspect's shoes or perhaps from the tread of a car tyre can be measured to compare it with solid at the scene of the crime.

Samples of the two soils are mixed thoroughly with distilled water and then filtered. Drops of universal indicator solution or a pH probe will be used to compare the pH of the soil samples.

4.6.3 Excess stomach acidity

Read the following adapted passage entitled "Heartburn" accessed at http://bit.ly/2knkJSQ on 20/8/2019.

> "Heartburn is caused by a build-up of excessive amounts of stomach acid, particularly HCl. This acid is used to digest the food we eat, but it can often back up into the oesophagus causing that burning sensation many of us are familiar with. The symptoms of heartburn can be treated with a mild base, which acts to neutralize the excess HCl. An antacid that can be bought over the counter such as Rennie®, work in similar ways. … Following this acid-base reaction, carbonic acid quickly degrades into CO_2 and H_2O."



Gastroesophageal Reflux Disease (GERD)

Following is a table of the pH of some substances we use daily.

Substance	pH
Baking soda	11
Orange juice	3
Tonic water	8.2
Bleach	13

You are suffering from acid indigestion.

a. Why is it not a good idea to drink bleach to remove the excess acid?

b. What is the reaction between an acid and a base called?

c. Would you rather drink water, tonic water, or orange juice? Explain.

d. Baking soda is sodium hydrogen carbonate while milk of magnesia is magnesium hydroxide. By writing balanced chemical equations including state symbols for their reaction with stomach acid (hydrochloric acid), suggest which one is best to ingest (eat) giving reasons for your answer.

Farming Chemistry

Ammonium nitrate is a common fertiliser. Calcium oxide is added to soil to reduce its acidity. Whilst it would be easier to mix the two chemicals together and distribute them in the soil together, farmers first apply the fertiliser and then lime the soil some days later.

This is done to prevent ammonia gas forming and escaping into the air. This is not desirable since the crops will not be able to absorb the nitrogen for growth and because ammonia is harmful if inhaled.

4.7 Salts

By the end of this section, you should be able to:

- Prepare a pure dry sample of an insoluble salt from named starting substances. (4.2g)
- Prepare a pure dry sample of a soluble/insoluble salt from different starting substances. *(Limited to metal with acid, carbonate with acid, base with acid, alkali with acid, and precipitation reactions.)* (4.3g)
- Represent the precipitation of an insoluble salt using chemical equations. (4.2e)

Salts are chemical substances that are made of positively charges ions called **cations** and negatively charged ions called **anions**. Salts may be soluble or insoluble in water.

The solubility of a salt determines which method should be used to produce it.

- **Insoluble salts are prepared by precipitation**. The precipitate can then be filtered, washed with distilled water, and dried.
- There are two methods to prepare **soluble salts**. It depends on the type of salt being produced. Soluble salts are prepared in solution. The solution is then evaporated to remove water and salt crystals form.

The following table lists salts which are soluble/insoluble in water.

Soluble	Insoluble
• All nitrates. • All hydrogencarbonates. • All group 1 metal salts. • All ammonium salts. • Halides except silver and lead halides. • Sulfates except barium, calcium, and lead sulfates.	• Carbonates except group 1 metal and ammonium carbonate. • Metal oxides except group 1 and 2 metal oxides that react with water. • Hydroxides except group 1 metal and ammonium hydroxides.

Source: SEC 06 2025 syllabus

Salts used as Pesticides

Certain salts are used as pesticides to kill insects, pests, weeds, and fungi. These include copper(II) sulfate, iron(II) sulfate, mercury(II) chloride, sodium arsenate and sodium chlorate.

4.7.1 Making salts

Salts can be prepared by using different techniques as shown in the next sub sections.

4.7.1.1 The action of an acid on excess insoluble solid

This method is used to prepare soluble calcium, magnesium, aluminium, zinc, iron(II) and copper(II) salts.

Salts such as magnesium sulfate can be prepared using the following reactions.

a) Reaction of an acid and a metal

$$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$$

b) Reaction of an acid on the insoluble **metal oxide**, **hydroxide**, or **carbonate**,

$$MgO(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2O(l)$$

$$Mg(OH)_2(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + 2H_2O(l)$$

$$MgCO_3(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + CO_2(g) + H_2O(l)$$

The solid reagent can be the metal, metal oxide, metal hydroxide or metal carbonate.

The following experimental procedure could be followed.

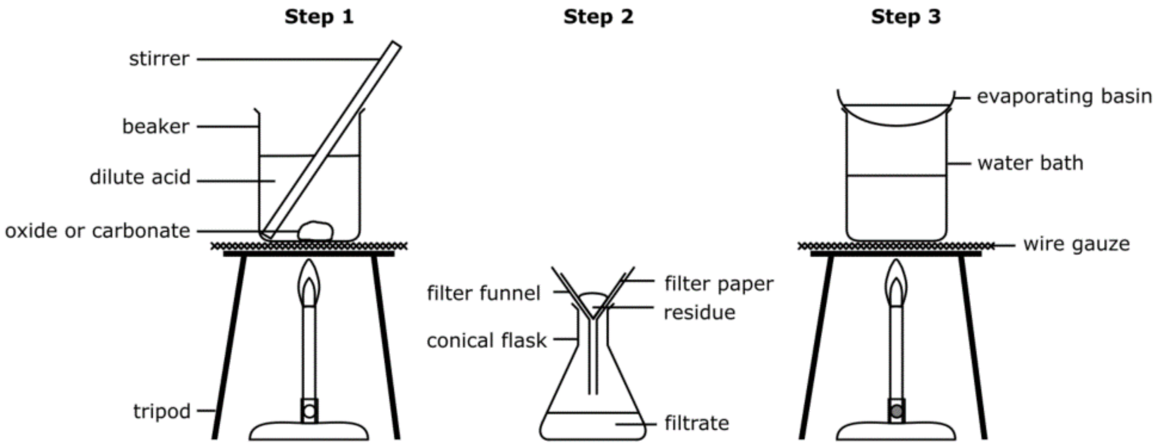
Step 1: 50 ml of dilute acid are poured in a beaker.

Step 2: The solid reagent is added to the acid.

Step 3: Excess of solid reagent is added to the acid to ensure that all the acid has reacted. The unreacted solid remains at the bottom of the beaker.

Step 4: The unreacted solid reagent is filtered from the mixture.

Step5: The solution of the soluble salt is poured in an evaporating dish. The filtrate is heated to evaporate water and obtain a concentrated solution of the salt. A cold glass rod is dipped into the hot concentrated solution. If salt crystals form on the end of the rod, the solution is ready to crystallise and so it is left to stand until the remaining solvent evaporates slowly. This process is known as **crystallisation**.



Preparation of copper(II) chloride by reaction of dilute acid with excess insoluble solid.

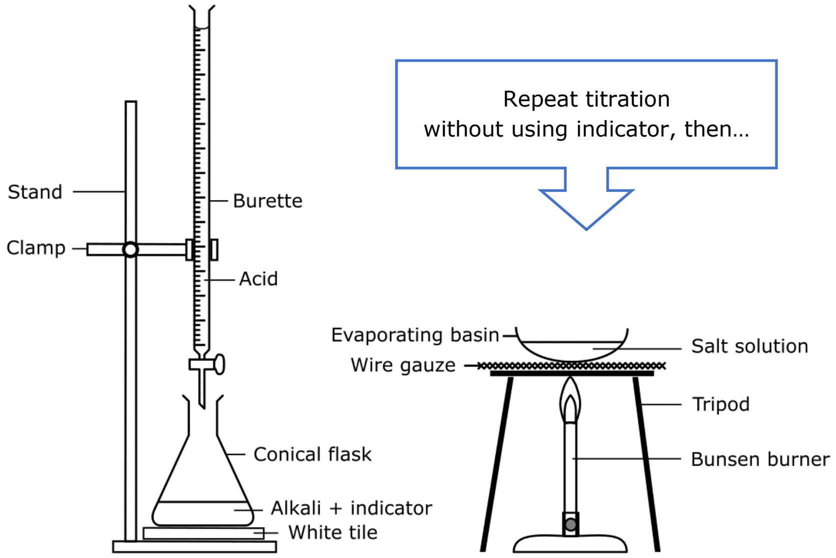
4.7.1.2 Titration

To obtain sodium, potassium and ammonium salts, an alkali or soluble carbonate can be added to a dilute acid. Example sodium sulfate can be produced as follows:

$$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

$$Na_2CO_3(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2O(l) + CO_2(g)$$

Where two aqueous solutions are mixed together, it is difficult to determine when the reaction has occurred completely. To obtain a pure salt one must ensure that the alkali/carbonate added to the acid must be just enough for complete neutralisation to happen. Failing to do so, would result in the salt being contaminated with acid or alkali/carbonate. To avoid this inconvenience, a method called titration is employed.



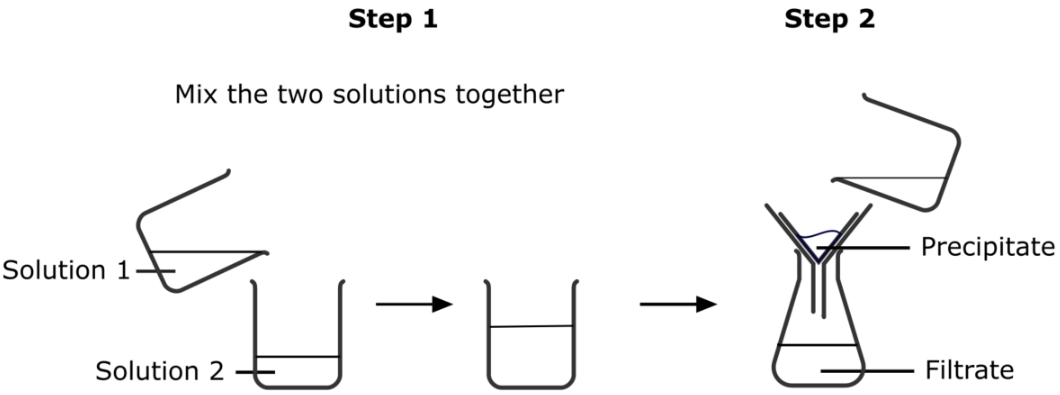
Titration is a process where a known volume of alkali/carbonate solution is transferred by means of a volumetric pipette to a conical flask. A few drops of indicator are placed in the flask. Acid is then added by means of a burette to the alkali/carbonate until the indicator changes colour showing that the end point has been reached. At this point one knows how much of the acid solution must be mixed with the alkali/carbonate solution. To obtain the pure salt, the newly determined quantities of acid and alkali/carbonate are mixed without the indicator. Water is evaporated from the resulting solution to obtain the pure salt. Refer to section 3.5.3 for separation techniques to obtain soluble salts.

The following steps describe how titration should be performed.

1. 25 cm^3 of 0.1 mol/dm^3 of the alkaline solution (e.g., sodium hydroxide) is measured using a volumetric pipette and is placed in a conical flask.
2. Two drops of indicator (e.g., phenolphthalein) are added. The indicator changes colour. (in phenolphthalein’s case it turns pink in alkali). Place the conical flask on a white tile.
3. The acid is placed in a burette, which is clamped vertically above the flask.
4. The acid is added a little at a time and the flask is swirled in a controlled way, to allow the acid and the alkali to mix.
5. When the indicator suddenly turns colourless, there is no need to add more acid.
6. The volume of acid, which has reacted, can be read from the burette.
7. The reaction is carried out again, **but this time there is no need to add the indicator**. 25 cm^3 of alkali are put in the flask and the correct volume of acid is added.
8. The solution is poured in an evaporating dish. Water is evaporated until the crystals of the dry salt are obtained.

4.7.1.3 Precipitation

This method is suitable for producing **insoluble salts**. Solutions of two soluble salts are mixed and an insoluble salt is formed.



For example, to prepare lead(II) sulfate; solutions of lead(II) nitrate and sodium sulfate may be used.

$$\mathbf{Pb}(NO_3)_2(aq) + Na_2\mathbf{SO_4}(aq) \rightarrow \mathbf{PbSO_4}(s) + 2NaNO_3(aq)$$

The $PbSO_4$ precipitate is filtered, washed with distilled water to remove impurities and dried.

Other examples of reactions that produce insoluble salts are:

$$\mathbf{Pb}(NO_3)_2(aq) + 2Na\mathbf{Cl}(aq) \rightarrow \mathbf{PbCl_2}(s) + 2NaNO_3(aq)$$

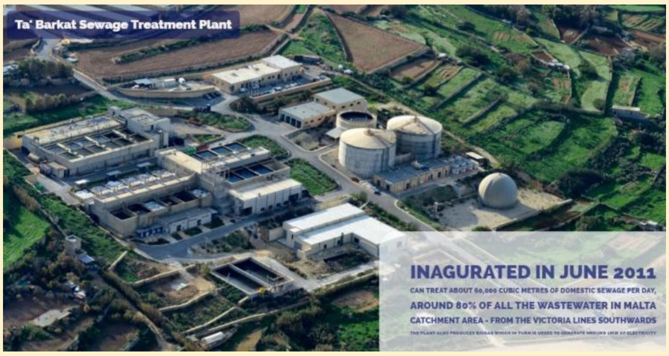
$$\mathbf{Ag}NO_3(aq) + Na\mathbf{Cl}(aq) \rightarrow \mathbf{AgCl}(s) + NaNO_3(aq)$$

$$\mathbf{Ba}Cl_2(aq) + Na_2\mathbf{SO_4}(aq) \rightarrow \mathbf{BaSO_4}(s) + 2NaCl(aq)$$

Crash Course Chemistry: Precipitation Reactions

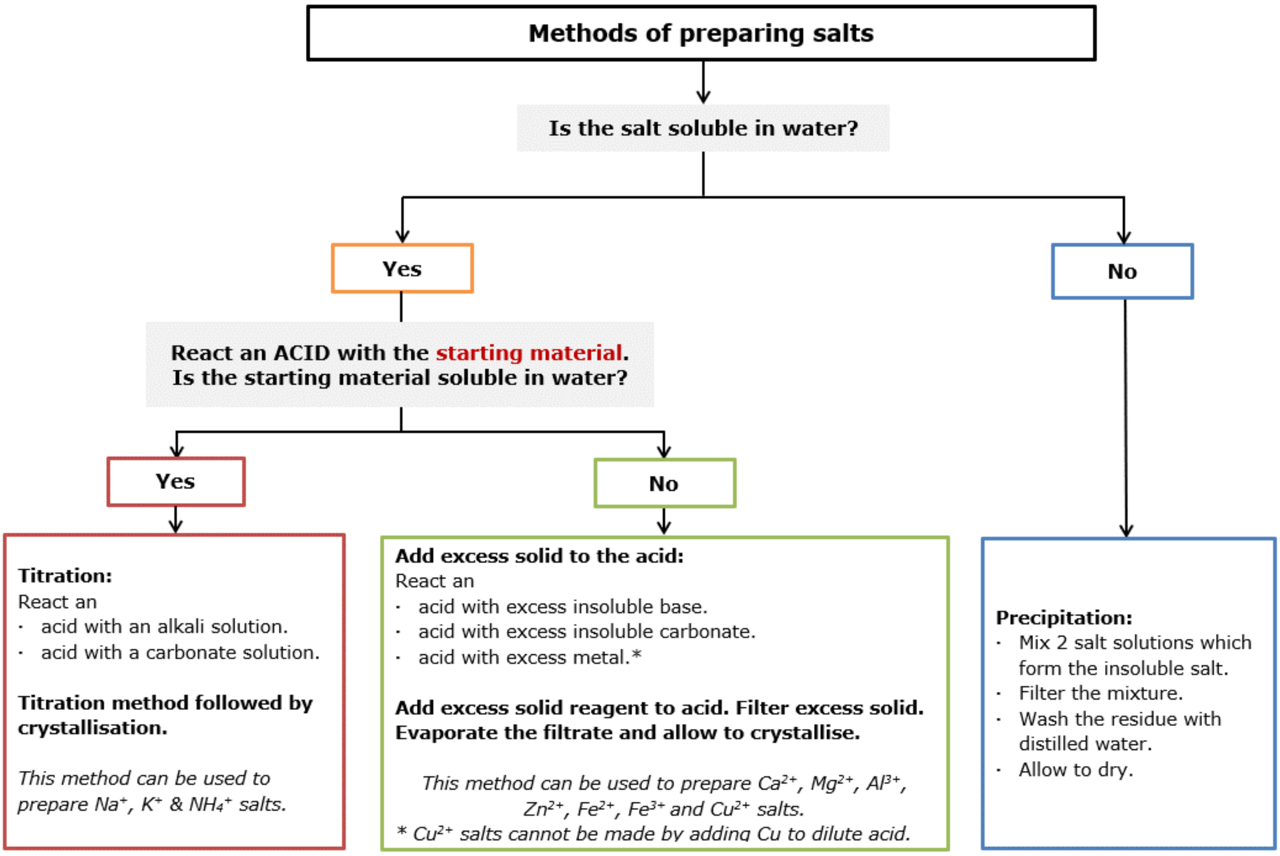
Preparation of lead(II) iodide by precipitation.

Precipitation for Clean Water



Chemical precipitation is used in wastewater treatment to remove metallic cations, and anions like fluoride, cyanide, phosphates, and other organic molecules. Since many hydroxide and sulfide compounds of heavy metals are insoluble in water, adding chemicals like sodium hydroxide (NaOH) or sodium sulfide (Na_2S) result in precipitation. Barium chloride is also used to precipitate detergents and oily emulsions.

Flowchart to determine methods of preparation of salts.



Methods of preparing salts

Is the salt soluble in water?

Yes

React an ACID with the starting material.
Is the starting material soluble in water?

Yes

Titration:
React an
- acid with an alkali solution.
- acid with a carbonate solution.

Titration method followed by crystallisation.

This method can be used to prepare Na^+, K^+ & NH_4^+ salts.

No

Add excess solid to the acid:
React an
- acid with excess insoluble base.
- acid with excess insoluble carbonate.
- acid with excess metal.*

Add excess solid reagent to acid. Filter excess solid. Evaporate the filtrate and allow to crystallise.

This method can be used to prepare Ca^{2+}, Mg^{2+}, Al^{3+}, Zn^{2+}, Fe^{2+}, Fe^{3+} and Cu^{2+} salts.
** Cu^{2+} salts cannot be made by adding Cu to dilute acid.*

No

Precipitation:
- Mix 2 salt solutions which form the insoluble salt.
- Filter the mixture.
- Wash the residue with distilled water.
- Allow to dry.

4.7.2 Some useful salts

$NaHCO_3$	Baking soda is used due to its leavening action.
$Na_2CO_3.10H_2O$	Washing soda as a water softener.
$CaSO_4.\frac{1}{2}H_2O$	Plaster of Paris for plaster casts and boards.

Salts for Medical Purposes

If you've ever broken a bone, you're very familiar with hydrated calcium sulfate ($CaSO_4.2H_2O$) found in the plaster of Paris which is used to make your cast. Patients who suffer from anaemia, take iron(II) sulfate heptahydrate ($FeSO_4.7H_2O$) in the form of iron supplement pills. People who need a low intake of sodium salt take potassium chloride as a replacement.

Epsom salt (magnesium sulfate heptahydrate - $MgSO_4.7H_2O$) and Glauber salt (sodium sulfate decahydrate – $Na_2SO_4.10H_2O$) are used as laxatives. Patients who need an X-ray of their digestive system drink barium sulfate solution ($BaSO_4$) known as barium meal. This helps the internal soft organs like the intestine appear on X-ray film. Potassium permanganate can kill bacteria and is therefore used as a disinfectant.

1. Zinc sulfate can be made by neutralising zinc oxide with an acid.
   a. Name the acid used to make zinc sulfate.
   b. Write a balanced chemical equation for this reaction.
   c. Briefly describe how a pure sample of zinc sulfate is produced.
   d. Draw labelled diagrams of each stage for this preparation.
   e. Give two precautions for this preparation and explain why they are taken.
   f. Suggest another suitable reagent that can be used instead of zinc oxide to prepare zinc sulfate.

2. Hydrochloric acid is a strong acid.
   a. Explain the term “strong acid”.
   b. What effect does hydrochloric acid have on blue litmus paper?
   c. Hydrochloric acid reacts with potassium hydroxide solution.
      i. Name the type of reaction taking place.
      ii. Write a balanced chemical equation for this reaction.
      iii. Both hydrochloric acid and potassium hydroxide are colourless solutions. Briefly describe the procedure used to determine how much acid is required to react with the alkali.
      iv. A salt is produced in this reaction. Describe how a sample of dry the salt can be produced.

3. Calcium oxide reacts with hydrochloric acid to form calcium chloride and water.
   a. Write a balanced chemical equation for the reaction.
   b. Calcium oxide is a base. Explain the term “base”.
   c. Farmers often add calcium oxide to the soil. Explain why they do this.

4. Excess stomach acid can be cured by taking antacid tablets or a suspension. Tablets can contain magnesium carbonate and the suspension contains magnesium hydroxide.
   a. Write a balanced chemical equation for the reaction of magnesium carbonate and hydrochloric acid. Include state symbols.
   b. Write a balanced chemical equation for the reaction of magnesium hydroxide and hydrochloric acid.

5. A solution of barium chloride is added to a solution of iron(II) sulfate to produce barium sulfate as one of the products.
   a. Write a balanced chemical equation for the reaction. Include state symbols.
   b. Name the type of reaction taking place.
   c. Give a use of barium sulfate.
   d. Describe how to prepare a sample of pure dry barium sulfate.
   e. Give a precaution for this experiment and explain why it is taken.
   f. Draw labelled diagrams of each stage of the experiment.
   g. Suggest another two chemicals that can be used to prepare barium sulfate and give a balanced chemical equation for this reaction. Include state symbols.

5 Periodic Tables of the Elements

PERIODIC TABLE OF THE ELEMENTS

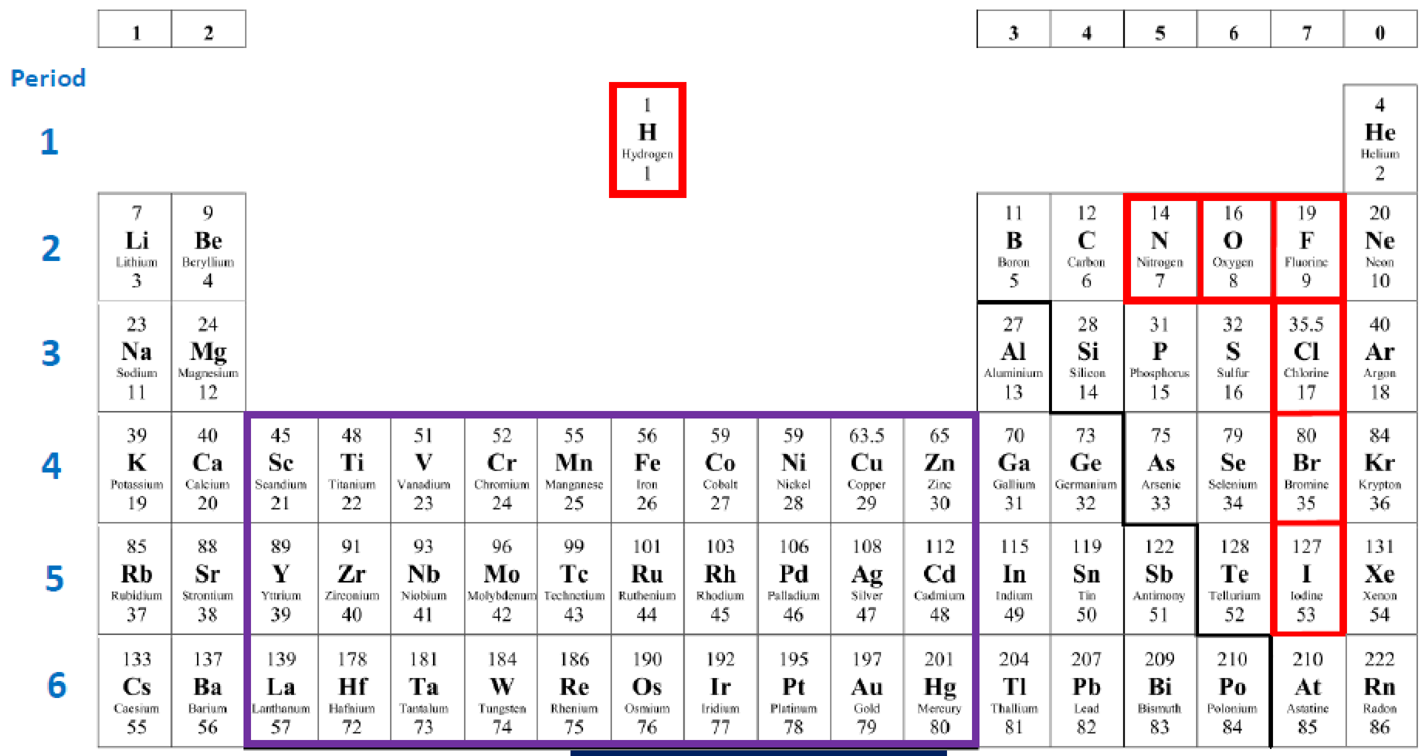
1	2											3	4	5	6	7	0
							1 **H** Hydrogen 1										4 **He** Helium 2
7 **Li** Lithium 3	9 **Be** Beryllium 4											11 **B** Boron 5	12 **C** Carbon 6	14 **N** Nitrogen 7	16 **O** Oxygen 8	19 **F** Fluorine 9	20 **Ne** Neon 10
23 **Na** Sodium 11	24 **Mg** Magnesium 12											27 **Al** Aluminium 13	28 **Si** Silicon 14	31 **P** Phosphorus 15	32 **S** Sulfur 16	35.5 **Cl** Chlorine 17	40 **Ar** Argon 18
39 **K** Potassium 19	40 **Ca** Calcium 20	45 **Sc** Scandium 21	48 **Ti** Titanium 22	51 **V** Vanadium 23	52 **Cr** Chromium 24	55 **Mn** Manganese 25	56 **Fe** Iron 26	59 **Co** Cobalt 27	59 **Ni** Nickel 28	63.5 **Cu** Copper 29	65 **Zn** Zinc 30	70 **Ga** Gallium 31	73 **Ge** Germanium 32	75 **As** Arsenic 33	79 **Se** Selenium 34	80 **Br** Bromine 35	84 **Kr** Krypton 36
85 **Rb** Rubidium 37	88 **Sr** Strontium 38	89 **Y** Yttrium 39	91 **Zr** Zirconium 40	93 **Nb** Niobium 41	96 **Mo** Molybdenum 42	99 **Tc** Technetium 43	101 **Ru** Ruthenium 44	103 **Rh** Rhodium 45	106 **Pd** Palladium 46	108 **Ag** Silver 47	112 **Cd** Cadmium 48	115 **In** Indium 49	119 **Sn** Tin 50	122 **Sb** Antimony 51	128 **Te** Tellurium 52	127 **I** Iodine 53	131 **Xe** Xenon 54
133 **Cs** Caesium 55	137 **Ba** Barium 56	139 **La** Lanthanum 57	178 **Hf** Hafnium 72	181 **Ta** Tantalum 73	184 **W** Tungsten 74	186 **Re** Rhenium 75	190 **Os** Osmium 76	192 **Ir** Iridium 77	195 **Pt** Platinum 78	197 **Au** Gold 79	201 **Hg** Mercury 80	204 **Tl** Thallium 81	207 **Pb** Lead 82	209 **Bi** Bismuth 83	210 **Po** Polonium 84	210 **At** Astatine 85	222 **Rn** Radon 86

Key:

a
X
y
b

relative atomic mass
SYMBOL
Name
atomic number

PERIODIC TABLE OF THE ELEMENTS



Period \ Group Number	1	2											3	4	5	6	7	0
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Key:

a	*relative atomic mass*
X	***SYMBOL***
Yyyy	***Name***
b	*atomic number*

🟥 **Diatomic Elements**

🟪 **Transition Metals**

Metals / Non-metals Border

Group Names:

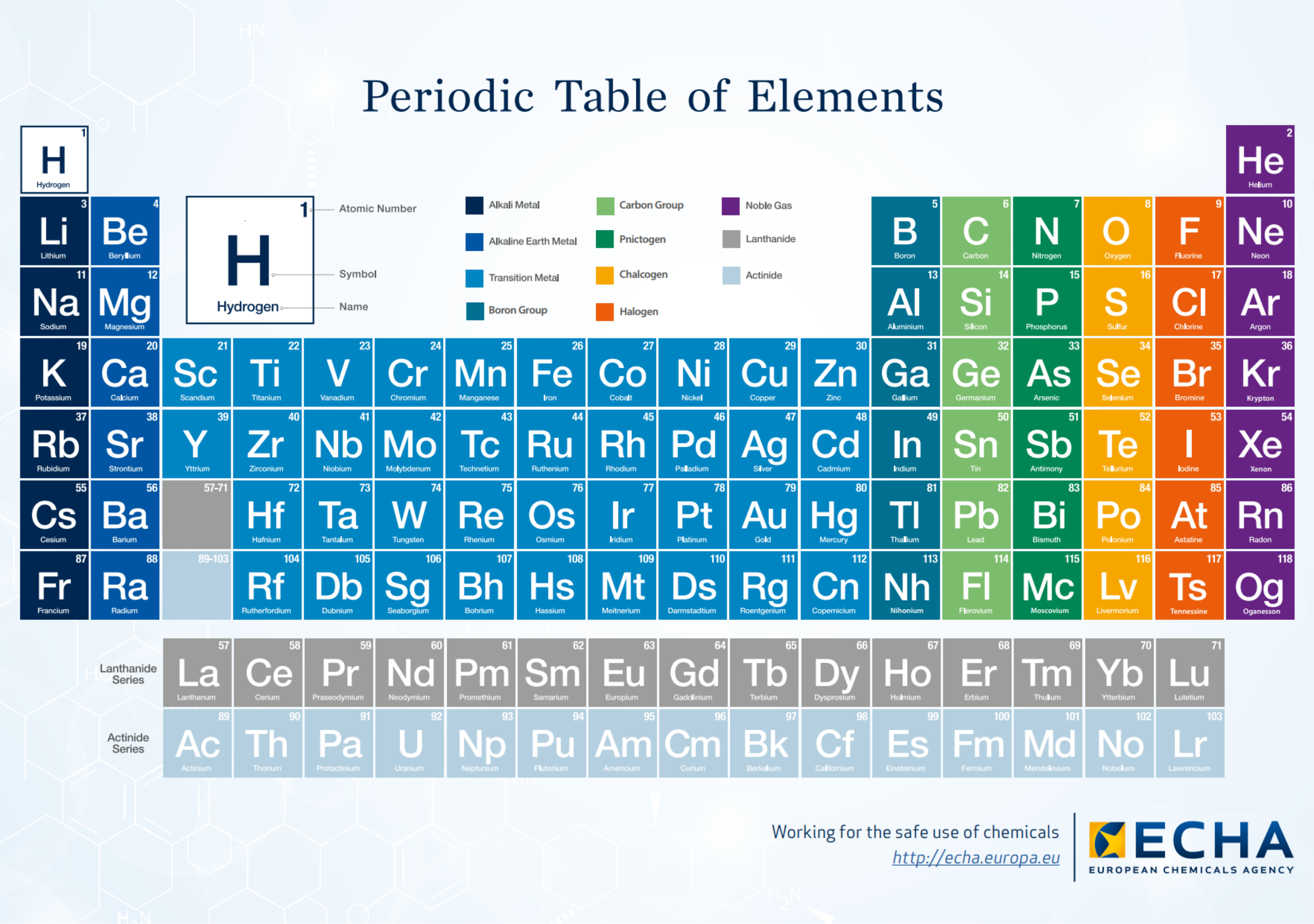
1 – Alkali Metals
2 – Alkali Earth Metals
7 – Halogens
0 / 8 – Noble (Inert) gases

Remember:

- **Group Number =** number of ELECTRONS in OUTER SHELL (Valence Electrons)
- **Period (Row) Number =** number of ELECTRON SHELLS in atom.

Remember:

- **Atomic Number = number of PROTONS**
- **Relative Atomic Mass = PROTONS + NEUTRONS**
- **In a NEUTRAL atom, number of PROTONS = number of ELECTRONS**



Periodic Table of Elements

1 Atomic Number
H Symbol
Hydrogen Name

Alkali Metal
Alkaline Earth Metal
Transition Metal
Boron Group
Carbon Group
Pnictogen
Chalcogen
Halogen
Noble Gas
Lanthanide
Actinide

1 H Hydrogen																	2 He Helium
3 Li Lithium	4 Be Beryllium											5 B Boron	6 C Carbon	7 N Nitrogen	8 O Oxygen	9 F Fluorine	10 Ne Neon
11 Na Sodium	12 Mg Magnesium											13 Al Aluminium	14 Si Silicon	15 P Phosphorus	16 S Sulfur	17 Cl Chlorine	18 Ar Argon
19 K Potassium	20 Ca Calcium	21 Sc Scandium	22 Ti Titanium	23 V Vanadium	24 Cr Chromium	25 Mn Manganese	26 Fe Iron	27 Co Cobalt	28 Ni Nickel	29 Cu Copper	30 Zn Zinc	31 Ga Gallium	32 Ge Germanium	33 As Arsenic	34 Se Selenium	35 Br Bromine	36 Kr Krypton
37 Rb Rubidium	38 Sr Strontium	39 Y Yttrium	40 Zr Zirconium	41 Nb Niobium	42 Mo Molybdenum	43 Tc Technetium	44 Ru Ruthenium	45 Rh Rhodium	46 Pd Palladium	47 Ag Silver	48 Cd Cadmium	49 In Indium	50 Sn Tin	51 Sb Antimony	52 Te Tellurium	53 I Iodine	54 Xe Xenon
55 Cs Cesium	56 Ba Barium	57-71	72 Hf Hafnium	73 Ta Tantalum	74 W Tungsten	75 Re Rhenium	76 Os Osmium	77 Ir Iridium	78 Pt Platinum	79 Au Gold	80 Hg Mercury	81 Tl Thallium	82 Pb Lead	83 Bi Bismuth	84 Po Polonium	85 At Astatine	86 Rn Radon
87 Fr Francium	88 Ra Radium	89-103	104 Rf Rutherfordium	105 Db Dubnium	106 Sg Seaborgium	107 Bh Bohrium	108 Hs Hassium	109 Mt Meitnerium	110 Ds Darmstadtium	111 Rg Roentgenium	112 Cn Copernicium	113 Nh Nihonium	114 Fl Flerovium	115 Mc Moscovium	116 Lv Livermorium	117 Ts Tennessine	118 Og Oganesson
	Lanthanide Series	57 La Lanthanum	58 Ce Cerium	59 Pr Praseodymium	60 Nd Neodymium	61 Pm Promethium	62 Sm Samarium	63 Eu Europium	64 Gd Gadolinium	65 Tb Terbium	66 Dy Dysprosium	67 Ho Holmium	68 Er Erbium	69 Tm Thulium	70 Yb Ytterbium	71 Lu Lutetium	
	Actinide Series	89 Ac Actinium	90 Th Thorium	91 Pa Protactinium	92 U Uranium	93 Np Neptunium	94 Pu Plutonium	95 Am Americium	96 Cm Curium	97 Bk Berkelium	98 Cf Californium	99 Es Einsteinium	100 Fm Fermium	101 Md Mendelevium	102 No Nobelium	103 Lr Lawrencium	

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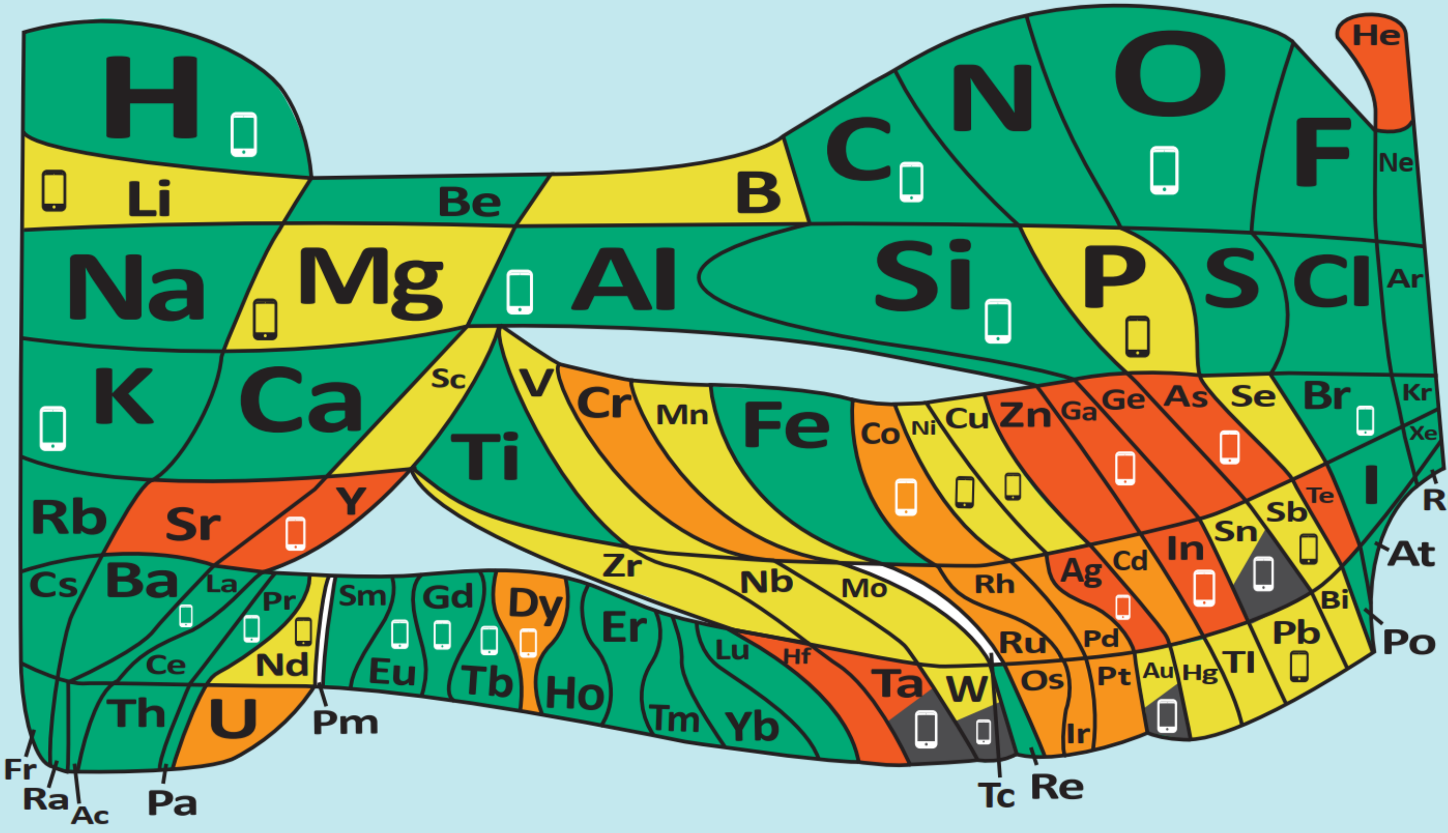
ECHA
EUROPEAN CHEMICALS AGENCY

United Nations
Educational, Scientific and
Cultural Organization

2019 IYPT
International Year
of the Periodic Table
of Chemical Elements

The 90 natural elements that make up everything

How much is there? Is that enough?



🟥 **Serious threat in the next 100 years**

🟧 **Rising threat from increased use**

🟨 **Limited availability, future risk to supply**

🟩 **Plentiful Supply**

☐ **Synthetic**

⬛ **From conflict minerals**

▯ **Elements used in a smart phone**

Inspired by WF Sheehan's 'A Periodic Table with Emphasis' published in Chemistry, 1976, 49, 17-18'

Read more and play the video game **http://bit.ly/euchems-pt**

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