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SCIENCE TEACHERS' ASSOCIATION OF WESTERN AUSTRALIA

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CHEMISTRY

Exploring Chemistry Year 11- Experiments, Investigations & Problems

Second Edition

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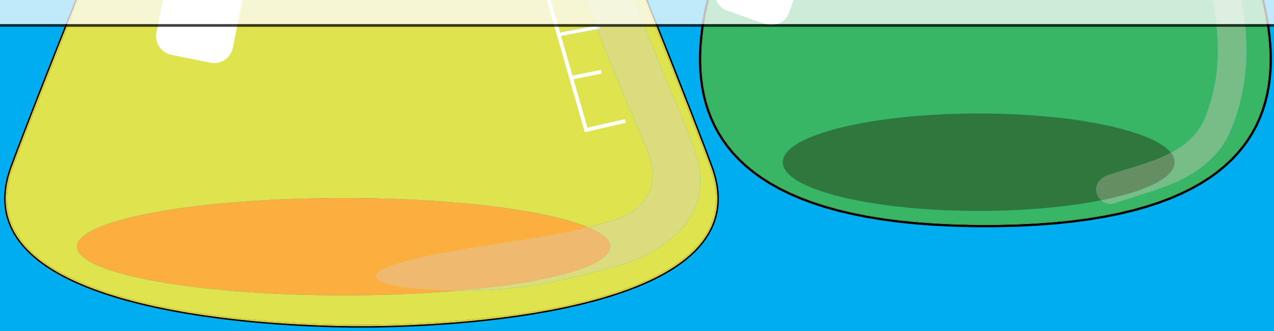

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CONTENTS

SAFETY RULES

Many experiments in chemistry use potentially dangerous chemicals and procedures. Providing you follow appropriate laboratory and safety rules, the risks associated with laboratory work can be minimised. The following is taken from the Australian Standards for laboratory practice (AS 2243.1-1990).
Always follow these rules.

Behaviour

1. Never adopt a casual or reckless attitude in the laboratory and always be conscious of the potential hazards.
2. Never run in the laboratory or along corridors.

Clothing

3. Ensure that personal clothing is suitable for laboratory conditions, e.g. non-slip, closed-in footwear. Do not wear open-toed shoes in the laboratory.
4. Always wear eye protection when in the laboratory.
5. Use protective clothing and devices appropriate to the operation being carried out, giving due thought to the work being carried out near you.

Due care

6. Always exercise care when opening and closing doors and entering or leaving the laboratory.
7. Take additional care when carrying any potentially hazardous substance.
8. Clean up spills immediately and report accidents or breakages to the teacher.

Emergency

9. Keep all fire escape routes completely clear at all times.

Food and poisons

10. Do not handle or consume food in the laboratory.
11. Regard all substances as hazardous unless there is definite information to the contrary.
12. Always use a fume hood when working with highly toxic, volatile or odorous substances.
13. Wash skin areas that come in contact with chemicals, irrespective of the concentration.
14. Dispose of specialised wastes (e.g. broken glassware and organic substances) in containers reserved for the particular type of waste.

SAFE LABORATORY PRACTICES

Following safe laboratory practices and being aware of potential hazards can avoid most accidents in the laboratory.

Acids and alkalis

- If spilt on the skin, these should be thoroughly washed off under running water for 20 minutes. Affected clothing should be removed immediately.
- If acids need to be diluted, always add the acid to the water. Never add water to acids.

Burns

- Take care not to handle hot glassware.
- Extinguish Bunsen burners when not in use.
- If a Bunsen burner is to be left for a short time, adjust the flame to an easily visible yellow colour.

Chemical spills

- If chemicals are spilt on the skin, remove affected clothing and wash the affected area thoroughly with running water.
- Do not carry bottles by the neck.

Cuts

- Take care when setting up glassware to avoid putting the glass under stress. Never force glass tubing through a cork or rubber stopper. Lubricate the tubing with water or inert grease and gently work the tubing into the stopper.
- When boring a cork or stopper, lubricate the borer with 1:1 ethanol/glycerol and place the stopper on a piece of wood, not on your hand.

Eye injuries

- Safety glasses must be worn. If you wear prescription glasses, you should wear over-glasses, safety goggles or a visor.
- Acids and alkalis are particularly dangerous. If any material enters the eye, the eye should be washed immediately and very thoroughly with running water. Note the location of the laboratory eyewash.
- Eye injuries occurring in the laboratory (including splashes) must be referred to medical staff.

Fire

- Never panic.
- Safety of people comes first. Know the escape routes and keep them clear.
- Remember the location of the fire extinguisher and fire blanket. Do not use a naked flame in the presence of flammable organic solvents such as ethanol, propan-2-one and petroleum spirits.
- If a person's clothing catches fire, put the victim on the floor and extinguish the fire with a laboratory coat or fire blanket.
- If a fire gets out of control, turn off gas and electrical appliances if possible, evacuate, shut the doors of the laboratory, and raise the alarm.

Heating test tubes

- This can be dangerous if the top of the test tube is pointed towards others and if the test tube is heated too rapidly. Point the top of the test tube away from others and heat gently by holding the test tube above the flame.

Ingestion

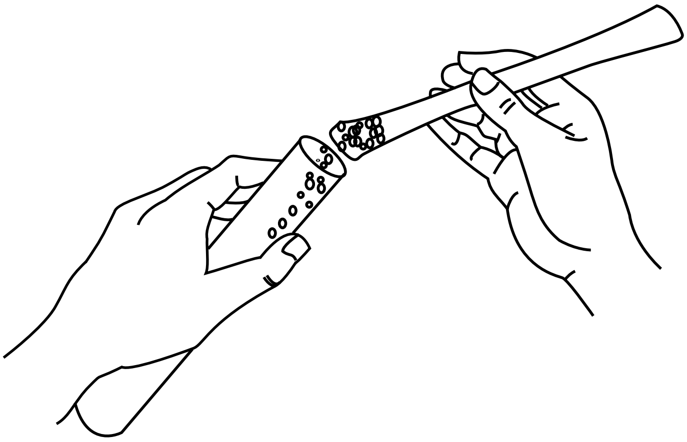
- Never taste chemicals.
- Do not pipette by mouth. Use manual pipette pumps.

Poisonous gases or vapours

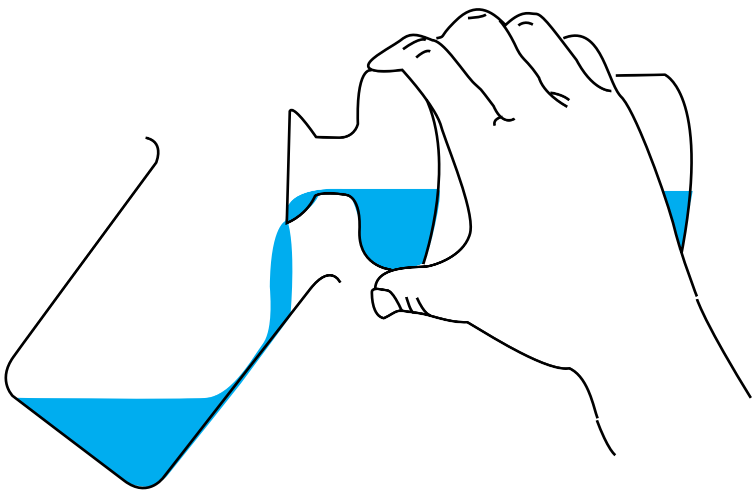
- Generally avoid inhaling any gases or vapours.
- If testing the odour of gases, gently waft the gas towards your nose and cautiously sniff.
- Use the fume hood to generate or use poisonous gases or vapours.

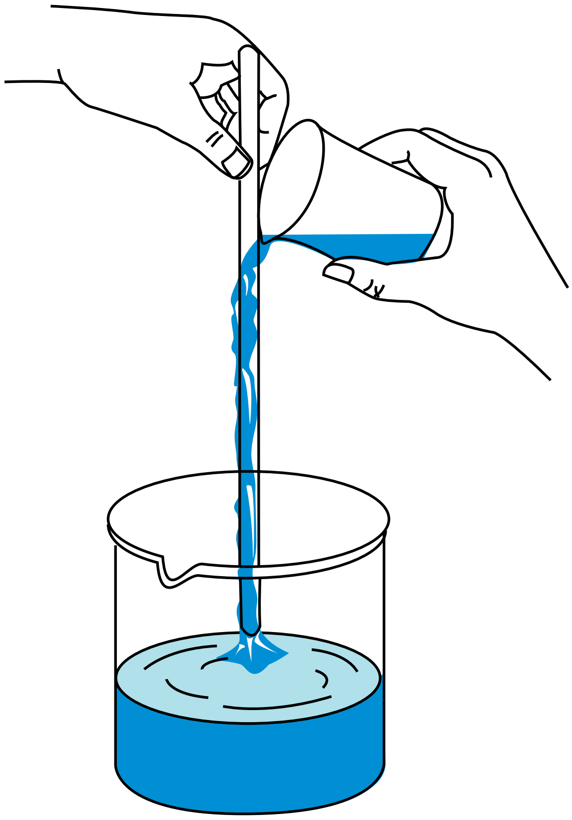
Transferring solids

A spatula is used to transfer small quantities of solid material from a stock bottle to a test tube, beaker or weighing bottle. The spatula must be thoroughly cleaned between uses by washing with water and thoroughly drying. Solid material must never be returned to a stock bottle as this may cause contamination. Discard any excess into the rubbish bin.



Transferring liquids by pouring

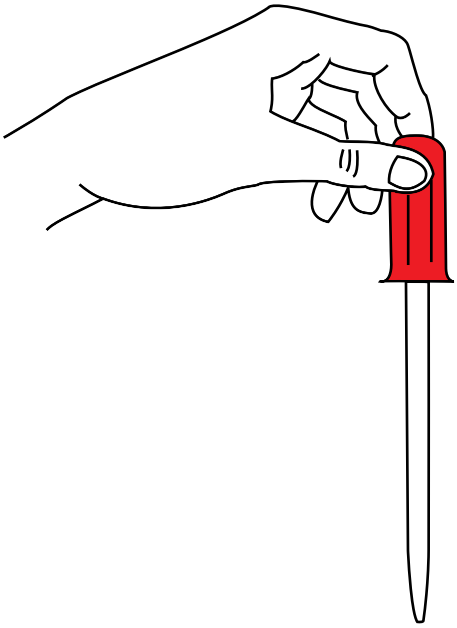




When pouring from a reagent bottle, the receiving vessel (beaker, flask, test tube, etc.) should be tilted so that the liquid can run down the side of the vessel without splashing.

When pouring from a beaker, the liquid stream can be directed using a stirring rod. The stirring rod is held against the beaker lip and directs the liquid stream into the receiving vessel.

Transferring liquids with a teat pipette



An eye dropper also known as teat pipette or Pasteur pipette, can be used if a small volume of liquid is required. Sometimes reagent bottles are provided with their own teat pipette. Be careful not to exchange these teat pipettes between bottles or to allow the teat pipettes to touch the bench or the insides of test tubes or beakers.

If the reagent solution is not provided with a teat pipette, do not use your own dropper to make the transfer, as this will contaminate the liquid. Wash a test tube or beaker, rinse with distilled water, rinse with a small amount of reagent and then transfer some of the reagent into the test tube or beaker. Use a clean teat pipette to obtain the required quantity of liquid from the test tube or beaker.

Do not invert the teat pipette when transferring liquids. Hold the teat pipette upright as atmospheric pressure is sufficient to support the liquid in the teat pipette. If the teat pipette is inverted, liquid may run into the rubber bulb at the end of the teat pipette, which tends to damage the rubber and contaminate liquids transferred on future occasions.

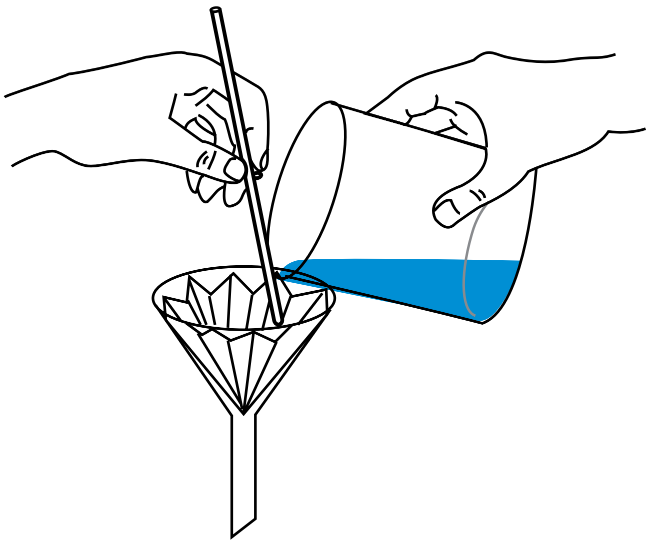
Filtration

Filtration is a technique used to separate a solid from a liquid. Depending on the situation, the solid may be an impurity that is discarded, or it may be the required product. In gravity filtration, separation of the mixture is achieved by adding it to a filter paper supported in a filter funnel. The filter funnel is itself supported by either a funnel holder (metal ring) or a filter funnel stand (wooden stand). The filter paper is folded in one of two ways.

Method 1: The filter paper is folded into halves and then quarters. One of the segments is opened and the filter paper is placed into the funnel.

Method 2: A fluted filter paper is obtained by folding the filter paper so that there are sixteen segments. The filter paper is then opened. A maximum filtering area is obtained with minimum contact between the filter paper and the funnel. (This type of folded filter paper is shown in the image).

Filtering to separate a solid:
Place the filter paper in the filter funnel and moisten it with a small amount of water. The liquid to be filtered is transferred from the beaker using a glass rod to direct the liquid into the filter paper. Care must be taken to ensure that the liquid does not come above the edge of the filter paper, and that the glass rod does not come in contact with the filter paper.



If the solid is required, extra liquid may need to be added to the beaker to wash out any residual material. Depending on the solubility of the solid, the filtrate or distilled water from a wash bottle may be used.

Weighing

The accuracy with which you want to weigh a particular substance will depend on the purpose of the activity. In some experiments you will only want to carry out a rough weighing within 0.1 g, but in other experiments you will want to do an accurate weighing, within 0.001 g or better. The accuracy of your measurement depends on the equipment available within your school laboratory and the best you might be able to do is 0.01 g.

Using a top-loading electronic balance.

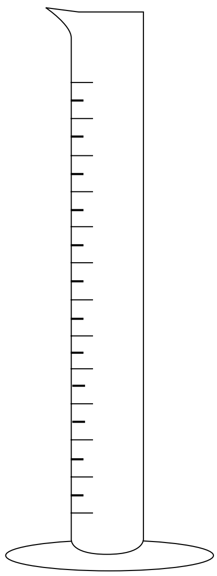
1. Place the container on the balance pan and either tare the balance so the readout is 0.0 g or 0.00 g or 0.000 g (depending on the precision of the balance being used), or record the balance reading.
2. Carefully add the required amount of substance.

Measuring liquid volumes

A range of glassware is used for measuring liquid volumes. This includes graduated cylinders, pipettes, burettes and volumetric flasks.

A graduated cylinder is used to deliver variable approximate volumes of liquid with moderate accuracy. Graduated cylinders, or measuring cylinders as they are often called, are available in various sizes:

10 mL (with 0.1 mL graduations)
25 mL (0.5 mL)
50 mL (1 mL)
100 mL (1 mL)
250 mL (2 mL)
500 mL (5 mL)
1 L (10 mL)



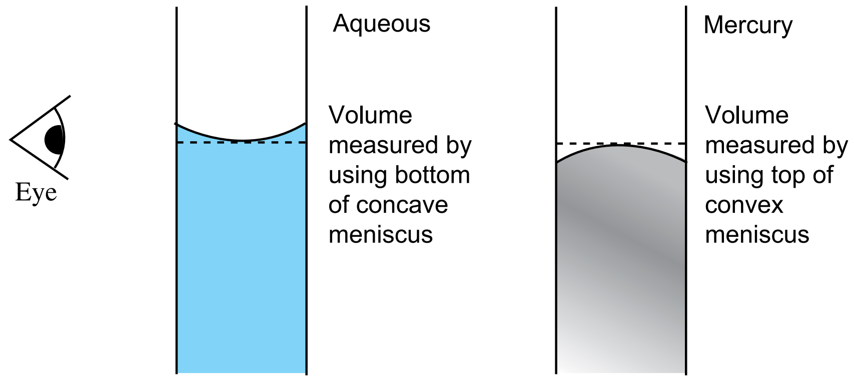
Cleanliness of glassware

In all situations it is essential that glassware is clean and free from grease. To test if glassware is clean, pour in some distilled water and swirl it gently. The glassware is clean if, when the water is poured out, an unbroken film remains on the glass surface, which drains uniformly under gravity. If the glass is dirty or greasy, the water will form droplets that stick to the sides of the glass and will cause errors in volume measurement.

If glassware is dirty, it should be cleaned with dilute detergent solution. This is then drained, the glassware rinsed several times with tap water, and finally with distilled water.

The liquid meniscus

The measurement of a liquid volume involves comparing the liquid level in a glass container with an accurately etched or graduated line on the container. The surface of a liquid is curved and in water occurs as a concave meniscus. The bottom, curved line of the meniscus is used when measuring the volume. In some intensely coloured aqueous solutions, such as potassium permanganate solution, it may be necessary to use the top of the meniscus, although this procedure is less accurate as the top of the meniscus is more difficult to identify.

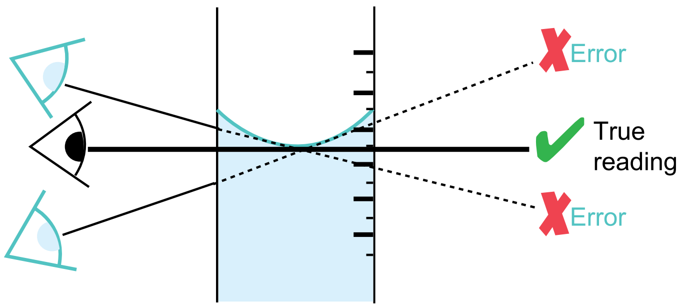


If the liquid is mercury, as in a barometer, the meniscus is convex and the top curved line of the meniscus is used. When using a graduated cylinder or burette, the bottom of the meniscus may not coincide with a particular graduated line. In these cases, where the bottom of the meniscus lies between two graduations, the volume reading is estimated from the position of the meniscus relative to the two adjacent graduations.

A useful rule of thumb is that the uncertainty in a measurement is half the smallest scale division.

Parallax error

Reading analogue scales such as those on metre rules or measuring cylinders, is more challenging than reading digital scales. A major difficulty in reading analogue scales is 'parallax error'.



A Guide to writing practical reports

Introduction

In general your practical report should reflect as accurately as possible all measurements and observations you made. It should also clearly illustrate your knowledge and understanding of the subject covered by the practical work. Practical reports should include the following headings:

Aim, purpose or hypothesis

This should be a short sentence or two, describing briefly what you are trying to find out from the experimental procedure you are conducting.

Equipment list and procedure

These should only be included for investigations where you need to design the experiment yourself. There is no point copying an equipment list and instructions when writing a report for a skill building experiment.

Results and/or observations

Results are numerical quantities which you measure during the experiment.

- You should record all measured quantities in your report.
- Record these quantities in a suitable table, particularly if there are a large number of measurements.
- You must include units for each quantity that you measure and record.
- Each measurement must be written to the appropriate number of significant figures. You may also be required to estimate the uncertainty involved in each measurement.

Observations are qualitative descriptions of what you actually see, hear, smell and detect by touch during an experiment.
What you infer, calculate, or conclude from the observations or from your knowledge about the experiment must not be included as observations.

Example 1

If you place a piece of zinc metal into hydrochloric acid one observation you might make is that colourless bubbles are produced on the surface of the zinc.
This is an **observation.**
From your knowledge of this type of reaction you may infer that the bubbles contain hydrogen gas.
This is an **inference** not an observation.

Example 2

If you drop an object one observation you may make is that you hear it hit the ground.
This is an **observation.**
You may infer that some of the kinetic energy of the object is converted to sound energy.
This is an **inference** not an observation.

Processing of results and questions

Written answers or calculations

- Include answers to all questions in your report.
- You will often use calculations to generate new quantities from numerical results. Full details of these calculations must be included in your report. If large numbers of similar calculations are required, then details of only one example of each type should be included.
- Always quote numerical results to an appropriate number of significant figures.
- One-word written answers are not acceptable. You must always include some justification or explanation in your answer.

Significant figures represent an approximate system of indicating the degree of accuracy of results. The precision with which an instrument is manufactured, the measurement scale provided and the skill of the experimenter contribute to the level of uncertainty of a measurement. For more detail read the 'Uncertainty and error' section on page 12.

Graphs

- These should be drawn on graph paper and to an appropriate scale.
- These should not be smaller than half an A4 page.
- Label with a heading/title.
- Label the axes with a suitable scale, the quantity being graphed and the units used.
- You should draw the line or curve of best fit (by eye) since the quantities plotted are measured values and so must contain uncertainties.

Conclusion

The conclusion for an experiment or investigation should relate to the aim, purpose or hypothesis of the experiment or investigation. It should answer the question asked or implied in the aim. The conclusion might be stated as the answer to one of the questions. If this is so, then there is no need to write a separate conclusion, otherwise a conclusion must be included.

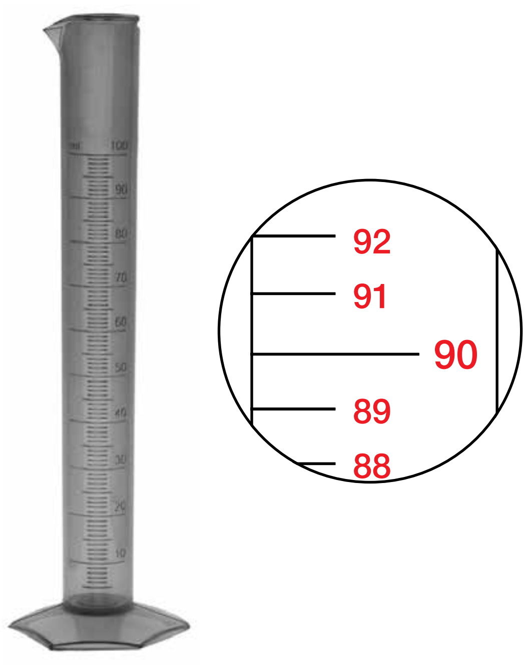
If the aim of the experiment is to experimentally verify a quantity, which has a known, generally accepted value, some comparison between the measured and established value should be made.

Uncertainty in measurement

Measurement is an important part of laboratory work in Chemistry. When any measurement is made, such as determining mass, volume or temperature, it is subject to a degree of uncertainty. What this means is that any quantitative data you record actually occurs within a range of values that could represent the true value.

Different pieces of apparatus have different levels of uncertainty associated with them and so it is important to carefully consider the piece of equipment you select for a task. Think of the variety of glassware available for measuring volume. There are beakers, measuring cylinders, volumetric pipettes, burettes and volumetric flasks of various grades, shapes and sizes. Selection of the most appropriate piece of glassware depends on the accuracy that is required for the task you are carrying out and your budget. The volume of water in a beaker being used as a water bath to cool a test-tube does not need to be measured to the nearest 0.1 mL whereas the volume of water added to reactions when investigating rate of reaction is critical.

Uncertainty in analogue equipment/instruments



If you were measuring 90 mL using a 100 mL measuring cylinder the uncertainty of that measurement would be +/- 0.5 mL as the smallest scale division is 1.0 mL. This means that the true value of the volume measured is in the range 89.5- 90.5 mL. The measured volume could be above or below the 90 mL you are trying to achieve.

A beaker would be unsuitable for measuring a volume of 100.0 mL accurately. The most accurate measuring device for a volume of 100.0 mL in a school lab would usually be a 100.0 mL volumetric flask. There is only one line on the flask so you cannot calculate the uncertainty using the half the smallest scale division rule; however, in this case there is generally a manufacturer’s uncertainty written on the apparatus. Different grades of glassware can be purchased. The best quality grades have the lowest uncertainty. A Grade A volumetric flask has an uncertainty of +/-0.08 mL.

Uncertainty in digital equipment

When reading a digital scale, the uncertainty is plus or minus the smallest scale division. For example if you use an electronic balance that reads to one decimal place its uncertainty would be +/- 0.1 g. A balance that reads to two decimal places has an uncertainty of +/- 0.01 g.

Experimental error

All measurements have a degree of uncertainty resulting in experimental error. The experimental error in a result is the difference between the experimental value and the published or theoretical value. There are two types of experimental error: random error and systematic error. Both should be considered when evaluating any quantitative investigation.

Random error

Random errors arise from measurements that have an equal chance of being above or below the actual value. Values taken from any analogue or digital scale have random error that can be stated in the uncertainty of the reading. Random errors can never be completely eliminated from measurements. Random errors affect the reliability or precision of the results. Precision refers to how close the values are to each other. A reliable method is one that produces consistent results with similar data readings.

If a number of experimental trials are carried out it is possible to reduce the effect of random errors on the final result. If an anomalous result (outlier) is obtained, repeating trials and working out an average minimizes the impact of the outlier if there is no justification for discarding it from the data set.

Another source of random error that must be considered is that humans may not be as accurate as the device being used for the measurement. For example human reaction time must be considered if using a stopwatch. The stopwatch may read to hundredths of a second; however human reaction time is far greater than this at around 0.1 s. In this case the uncertainty of the measuring device has minimal effect on the results compared with the human impact. Reaction time might impact on experimental data in a random manner.

An experiment with low random error will be more reliable or precise; however, it might not necessarily produce accurate results. The precision of a measurement refers to the consistency among a number of measurements made in the same way.

Systematic error

Systematic errors are a result of flaws in the experimental method or apparatus that lead to a result that is always either above or below the true value. Examples of things that result in systematic errors include the type of equipment used, calibration of the equipment and human judgement.

Systematic errors cannot be reduced by increasing the number of experimental trials carried out. They must be overcome by altering the method in some way, generally by using different equipment. Enthalpy changes such as the heat of neutralisation can be estimated experimentally by measuring temperature change when acid and alkali are mixed together in a beaker. The increase in temperature and other data can be used to estimate the enthalpy of reaction. The enthalpy change obtained is lower than the true value due to heat loss from the reaction vessel. The heat loss is the most significant source of systematic error in this experiment. Altering the method to minimize heat loss by adding insulation to the beaker will achieve a more accurate result. Human colour perception could impose systematic or random errors in an experiment. You should use your common sense when evaluating a particular experiment to judge whether the error is random or systematic. Removing the need for human judgment in the experimental design by choosing a piece of equipment that can measure colour intensity (a colorimeter) or using a digital pH meter that gives continuous monitoring of pH rather than a chemical indicator that changes colour can reduce systematic error.

SECTION 1
Experiments and Investigations

Unit 1

Chemical fundamentals: structure, properties and reactions

Design, conduct, evaluate and communicate investigations into the properties of elements,compounds and mixtures and the energy changes involved in chemical reactions.

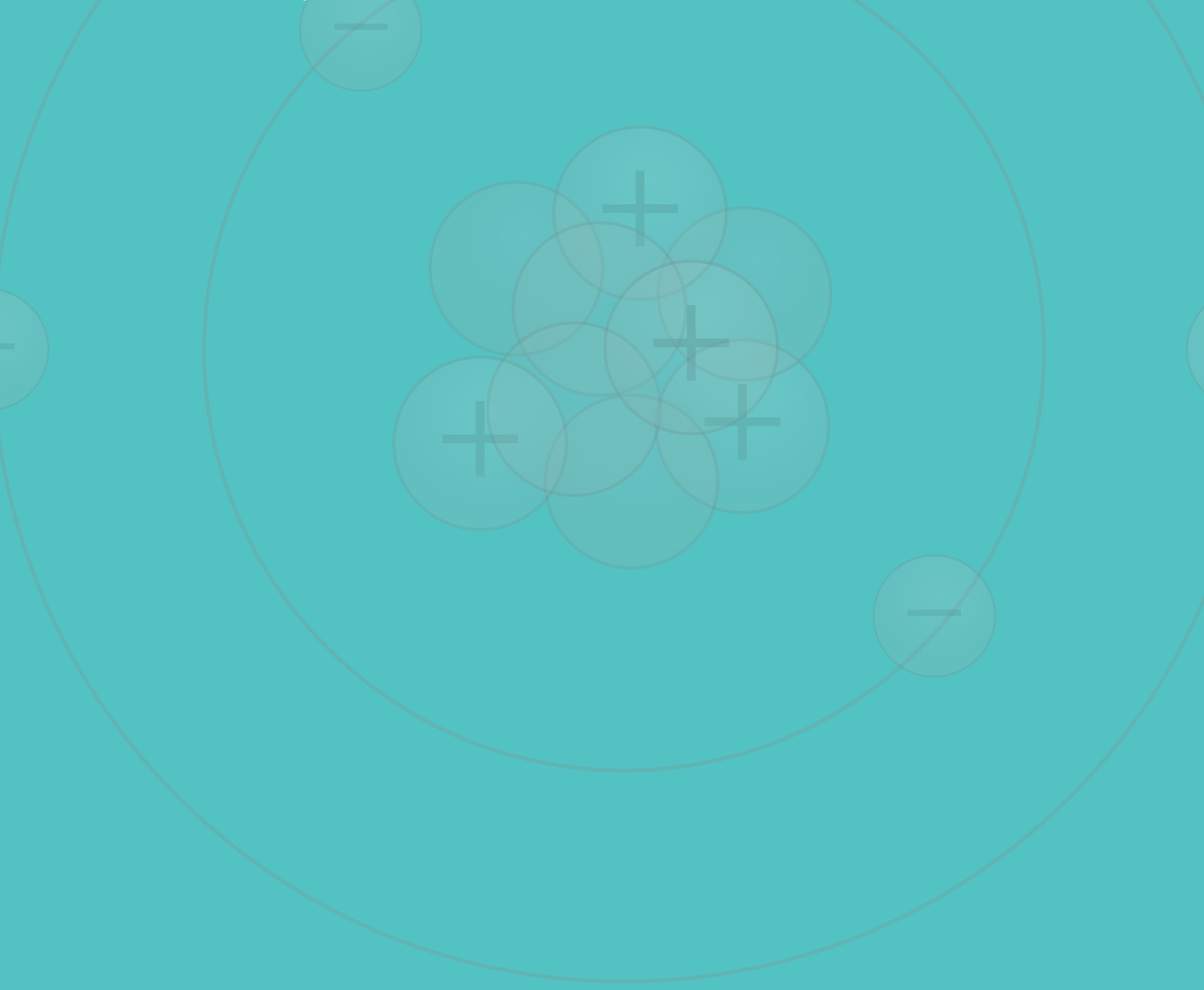
Safely conduct investigations, collecting valid and reliable data, using equipment appropriate to accurately measure temperature change and mass, flame tests, separation techniques and heat of reaction.

Record and represent data in meaningful and useful ways, including using appropriate graphic representations and correct units and symbols. Process data to identify trends, patterns and relationships. Identify sources of random and systematic error and estimate their effect on measurement results. Make and justify conclusions using evidence from investigations.

Properties and structures of atoms

Design, conduct, evaluate and communicate investigations of the science of atomic models to understand and explain the structure and properties of elements.

The results from scientific investigations have built our understanding of the atom. The Greek philosopher, Democritus, was the first to describe matter as indivisible tiny particles - “atomos”. Since then scientist including, Dalton, Thomson, Millikan, Rutherford, Bohr, Chadwick and Schrödinger progressively developed models of atomic structure making reliable predictions about the mass, charge and location of the sub-atomic particles.



Experiment 1: Flame tests

The characteristic colours of fireworks are due to the presence of particular metal salts. The energy to raise the electrons in the metal ions comes from the exothermic chemical reaction between oxygen in the air and gunpowder. Almost immediately, excited atoms emit photons with characteristic wavelengths as the electrons fall from higher energy levels to lower energy levels. Copper gives a green display, sodium gives yellow, calcium gives red, and potassium gives violet. In industry we use this property to identify elements, and, using a modification of this 'flame test' in an atomic absorption spectrometer (AAS) we can even measure concentrations of specific elements.
You may have noticed the different flame colours that you see when a glossy magazine is burnt or that cooking salt spilled near a gas flame causes the flame to go an orange/yellow colour. These are also examples of where metal ions gain energy and then release it in the form of light.

Aim

The purpose of this experiment is to examine the colour produced by different salts when they are flame tested.

Part A – Microscale version of flame test

Equipment

- Nichrome wire loop
- 4 test tubes
- texta to label test tubes
- test tube rack
- 1 mol L^{-1} **HCℓ** solution (1 dropper bottle of clean fresh HCℓ per group)
- distilled water in bottle
- Bunsen burner
- matches
- small samples (pea sized) of salts such as:
  potassium nitrate [KNO_3]
  magnesium nitrate [$Mg(NO_3)_2$]
  copper(II) sulfate [$CuSO_4$]
  iron(II) sulfate [$FeSO_4$]
  strontium nitrate [$Sr(NO_3)_2$]
  barium nitrate [$Ba(NO_3)_2$]
  calcium nitrate [$Ca(NO_3)_2$]
  lithium nitrate [$LiNO_3$]
  sodium chloride [NaCℓ]
  sodium nitrate [$NaNO_3$]

Procedure

1. Add 5 mL 1.0 mol L^{-1} HCℓ to two test tubes. Label them "1" and "2".
2. Add 5 mL of distilled water to a further two test tubes. Label them "3" and "4".
3. Light a Bunsen burner, and adjust the flame to a hot blue colour.
4. Dip the nichrome wire loop into the first test tube of HCℓ.
5. Tap off the excess HCℓ, and bring the loop into the hottest part of the flame above the blue cone.
6. Heat until the flame ceases to glow.
7. Rinse off the nichrome loop with distilled water, then repeat with tube 2 (HCℓ), 3 (distilled water) and 4 (distilled water).

8. Take the now clean nichrome wire loop, add a drop of distilled water to it, and pick up a small amount of potassium nitrate salt.
9. Bring the loop to the hottest part of the Bunsen flame, and observe and record the colour of the light produced. Leave the loop in the flame until most of the colour is gone.
10. Rinse the nichrome loop in distilled water, then repeat the cleaning procedure in steps 3-7. You may replace the distilled water in Tubes 3 and 4 if you need to.
11. Repeat steps 8-10 for the remaining salts.

Results and observations

Salt	Colour observed
KNO_3	
$Mg(NO_3)_2$	
$CuSO_4$	
$FeSO_4$	
$Sr(NO_3)_2$	
$Ba(NO_3)_2$	
$Ca(NO_3)_2$	
$LiNO_3$	
$NaC\ell$	
$NaNO_3$	

Questions

1. What component of the salt is responsible for the colours observed? What is your evidence for this conclusion?

2. Why must the nichrome wire loop be cleaned before each test?

3. Explain how the coloured light observed is produced.

4. Give a practical use of this type of test.

5. Often different colours are used to describe the same flame colour, e.g. lilac, mauve or purple for potassium.
   (a) Why might this be a problem?

   (b) Is this a random or systematic error? Justify your choice.

Experiment 2: Spectroscopy

Chemists use spectroscopy to identify elements in distant gas clouds or stars. Spectroscopy works on the principle that substances absorb energy and emit energy as electromagnetic radiation. The radiation is absorbed or emitted in bundles (called photons) whose wavelengths, which correspond to colours, are specific to the atoms that make up the substance. For example, sodium atoms emit yellow light very strongly.

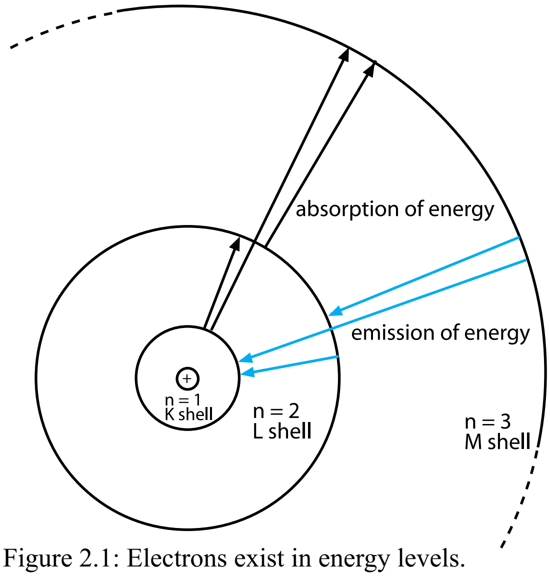


Figure 2.1: Electrons exist in energy levels.

For many years now, astronomers and other scientists have used photons of electromagnetic radiation (including visible light) from our Sun and other stars to identify which elements they contain, and even the proportions in which they occur.

The state of an element where its electrons are in their lowest possible energy levels is called the 'ground state'. When an element is heated the electrons in their ground state absorb energy and move to higher energy levels, called 'excited states'. When the electrons return to lower energy levels, energy is emitted as light. The light can be seen as distinct colours and the colours emitted are characteristic of the element. This is because the colour of the light emitted is determined by the difference in energy between the excited and ground states. As each element has a unique arrangement of electrons surrounding its atomic nucleus, it will also have a unique spectrum. We see this in Figure 2.1 and 2.2. A popular metaphor is that the spectrum of an element is its 'fingerprint'.

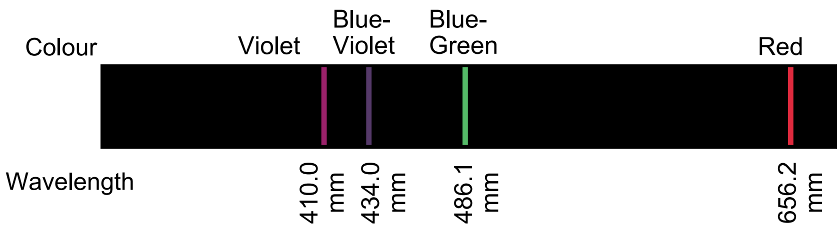


Figure 2.2: Emission spectrum for hydrogen in the visible region of the electromagnetic spectrum.

Equipment

- hand held spectroscope
- gas emission (discharge) tubes (for example hydrogen, helium, neon)
- power supply

SAFETY NOTE:
- **Do not point the spectroscope directly at the sun.**
- **Do not touch the tubes when they are powered by the high voltage power supply**

Procedure and observations

Use the spectroscope to observe the spectra of natural daylight by pointing it at a region of blue sky. Do not point the spectroscope directly at the sun. Describe what you observe.

2. Darken the room as much as possible and observe the light emitted from various light sources provided in the lab. Sketch the emission spectra you observe in the table below.

Light source (gas in the tube)	Colour observed (approximate wavelength - nm)  400 450 500 550 600 650 700

Processing of results and questions

1. Name the colours that can be seen in the visible spectrum using the spectroscope.

2. Explain why the spectrum you observe for natural daylight is referred to as a **continuous spectrum**.

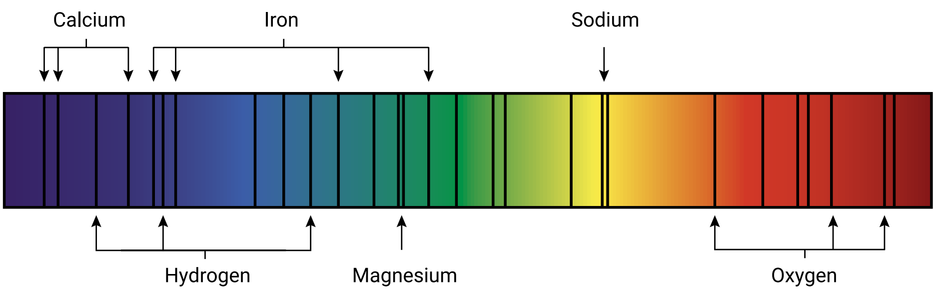
3. Explain why the room must be darkened for viewing the light emitted from the gas discharge tubes.

4. Compare and contrast the emission spectra from the gas discharge tubes compared with the continuous spectrum of daylight.

5. Describe how atomic emission spectroscopy could be used to identify elements.

6. Atomic absorption spectroscopy is a technique related to atomic emission spectroscopy. In atomic absorption spectroscopy specific wavelengths of light are absorbed by the sample. This is shown in Figure 2.3 and you might have observed the gaps in the spectrum of sunlight with black lines indicating wavelengths absorbed.

The Sun's spectral lines



(a) Did you observe any black absorption lines for the daylight (sunlight) spectrum?. If you did, use figure 2.3 to help list the elements identified by your spectrum.

(b) Explain why absorption spectra contain dark lines, and how this technique can also be used to identify elements.

Properties and structures of materials

Design, conduct, evaluate and communicate investigations into materials and models of bonding to understand and explain the structure and properties of compounds.

Matter can be manipulated to create new materials, composites and devices leding to commercialisation. The structure and properties of materials are fundamental to their selection in the design and manufacture of goods. The use and disposal of materials can propose risks to health, safety and the environment. Regulations including labelling and disposal of new and existing materials including nano materials continue to be introduced.

Experiment 3: Recognising mixtures

Scientists study the effect of chemicals on our health, however sometimes these chemicals can be part of a mixture which means scientists need to recognise and understand mixtures. Mixtures can be heterogeneous or homogeneous. A heterogeneous mixture has an irregular composition. Concrete and fruitcakes are examples of heterogeneous mixtures. Homogeneous mixtures on the other hand have the same composition throughout. Solutions are examples of homogeneous mixtures. How would you classify blood?

During this laboratory activity, you will develop an understanding of different types of mixtures including solutions.

Aim

To observe and identify heterogeneous and homogeneous mixtures. To attempt to make a supersaturated solution.

Equipment

- Bunsen burner
- tripod and gauze mat
- matches
- watch glass
- thermometer
- beakers (2 × 100 mL) measuring cylinder (25 mL)
- potassium nitrate solid (6 g)
- 100% orange juice
- 1% saline solution
- salad dressing (oil, water, salt)
- solder (lead free)
- fruit scone (or fruit cake)
- charcoal
- boiling beads (or boiling chips)
- hand lens or magnifying glass

SAFETY NOTE:
- **Do not point the spectroscope directly at the sun.**
- **Do not touch the tubes when they are powered by the high voltage power supply**

Part A: Homogeneous or heterogeneous

Procedure

Examine the 6 mixtures and record your observations in the table drawn below.

	Mixture	Observations	Classification
1	100% orange juice		
2	1% saline solution		
3	Solder		
4	Fruit scone or cake		
5	Air		
6	Salad dressing		

Part B Saturated and Supersaturated Solutions

1. Use a solubility graph to determine the mass of salt that you would require to make a 15 mL saturated solution of potassium nitrate at 25 °C.

2. Using the measuring cylinder, pour 15 mL of distilled water into one of the 100 mL beakers. Measure and record the temperature of the water. Cool it in an ice bath so it is 20 °C.
   Predict (based on your recorded temperature and value calculated in procedure #1) the type of solution that you would form by dissolving the salt in the 15 mL of water: unsaturated, saturated or supersaturated. Write your prediction in the results table.

3. Weigh the calculated mass of potassium nitrate in procedure #1 and attempt to dissolve it in 15 mL of distilled water at 20 °C
   Observe and describe the type of solution that you have created in procedure #3: unsaturated, saturated or supersaturated.

   Explain your observations (e.g. explain why your predicted solution type is different from, or the same as, your observed solution type).

4. Now, gently heat your mixture, with continuous stirring, to dissolve any remaining solid. Stop heating the\ mixture once you are convinced all solid has dissolved or when boiling is about to occur. Describe and record your observations including the temperature of the water.

5. Remove any stirring implement but leave the thermometer in the water and allow it to cool. Be very careful not to touch or bump the beaker or thermometer.

6. Record the temperature at which you observe the formation of any solid

Observations and results

Mass of potassium nitrate required	
Prediction of type of solution	
Observation - type of solution	
Explanation	
Observation on heating	
Temperature at which solid forms	

Processing of results and questions

1. Complete the following table:

Characteristic		
Composition		
Number of Phases-possible		
Particles of mixture visible with the naked eye.		
Examples		

2. Describe how you would make 100 mL of the following types of sodium chloride solution given that the solubility of sodium chloride in water is 35.9 g/100 mL at 25 °C.

   a) unsaturated

   ____

   ____

   b) saturated

   ____

   ____

   c) supersaturated.

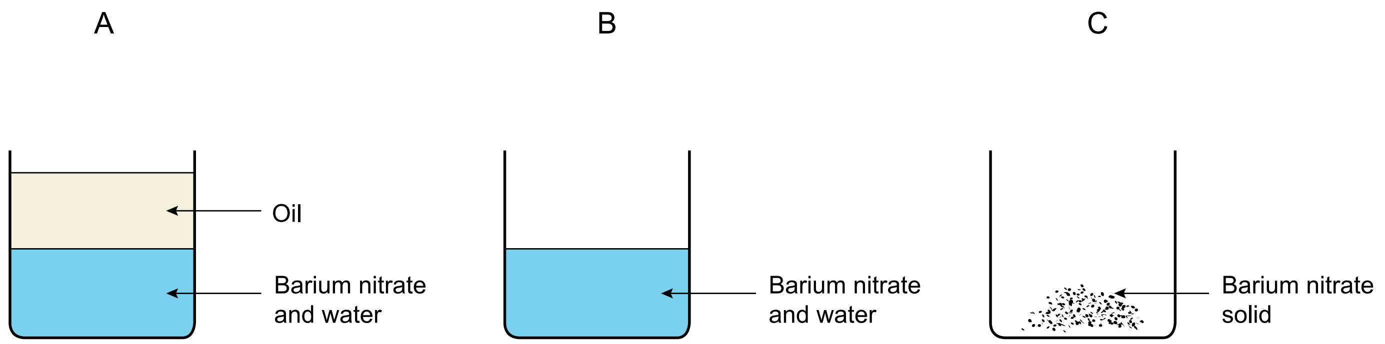
   ____

   ____

3. Compare temperature records in Part B. Did you manage to make a super saturated solution? Explain your answer.

4. You are given a solution of sodium nitrate and must determine if it is saturated or unsaturated. Explain how you could determine this and any materials and equipment you would require.

5. A student is presented with three beakers containing:
A. Oil, barium nitrate and water.
B. Barium nitrate and water
C. Barium nitrate solid



Use the following labels to correctly identify the three beakers and explain your answer:
Homogenous mixture, heterogeneous mixture and pure substance.

Beaker	Label	Explanation
A		
B		
C		

Experiment 4: Filtration and crystallisation

Separation techniques are used every day in kitchens and industry. Fruit juices like orange juice contain pulp which many people find unpleasant. Therefore, fruit juices are often processed by filtration to remove pulp. In the case of solids mixed together, separation is often achieved by filtration and crystallisation. To be successful, this requires that the components of the mixture have different solubilities in a particular solvent. In this experiment you will examine the solubilities of some common chemicals in order to complete separation of two mixtures.

Aim

To separate mixtures of charcoal/sodium chloride and copper(II)chloride/sodium chloride.

Equipment

- electronic balance
- filter funnel
- filter funnel stand
- matches
- spatula
- Bunsen burner, tripod and gauze mat
- glass rod
- beakers (two 100 mL)
- graduated cylinder (25 mL)
- boiling chip
- ethanol (25 mL)
- distilled water
- wash bottle
- hot hand holders
- watch glass
- filter paper (Whatman No. 1-two 12.5 cm sheets)
- charcoal
- sodium chloride solid (4 g)
- copper(II) chloride solid (4 g)
- sodium chloride/charcoal mixture (4 g)

Procedure

Part A: Solubility

1. Test the solubility of a small sample (half a teaspoon) of ethanol, charcoal, sodium chloride and copper(II) chloride in separate 15 mL lots of distilled water. Record your observations and classify each sample as soluble or insoluble in the 'water (solvent)' column of the table below.

2. Record your observations and classify each sample as soluble or insoluble in the 'ethanol (solvent)' column of your table.

Observations and results

Examine the 6 mixtures and record your observations in the table drawn below.

Solute	**Solubility**	
	Water (solvent)	Ethanol (solvent)
Charcoal		
Sodium chloride		
Copper(II) chloride		
Ethanol		
Water		

Part B: Separation of a sodium chloride and charcoal mixture

1. Place approximately 4 g of the salt/charcoal mixture in a 100 mL beaker and add about 15 mL of distilled water. Stir the mixture for about two minutes to allow the salt to dissolve.
2. Set up a filter funnel stand with funnel and filter paper. Filter the mixture and collect the filtrate in a 100 mL beaker.
3. Wash the solid with a further 5 mL of distilled water but do not add this to the filtrate. Record the appearance of the solid.
4. Add a boiling chip to the filtrate. Heat the solution with a Bunsen burner and boil gently to reduce the volume.
5. When crystals of sodium chloride appear, turn off the Bunsen burner and allow the solution to cool.
6. Record the appearance of the sodium chloride crystals.

Observations and results

	Observations
Appearance of solid in filter paper	
Appearance of Sodium chloride crystals	

Part C: Separation of a sodium chloride and copper(II)chloride mixture

Based on your results from Part A, predict which solvent will best separate this mixture.

Prediction:

1. Weigh out approximately 2 g of sodium chloride and record its mass accurately. Add this to a 100 mL beaker.
2. Weigh out approximately 2 g of copper(II) chloride and record its mass. Add this to the same 100 mL beaker.
3. Add about 15 mL of ethanol to the beaker and stir the mixture for about two minutes.
4. Weigh a piece of filter paper to use to filter the undissolved solid.
5. Wash this solid with about 5 mL of ethanol. Note the appearance of the solid.
6. Allow the solid in the filter paper to dry (overnight) and reweigh. Calculate the mass of solid in the filter paper.
7. Place about 5 mL of the filtrate on a watch glass and allow the ethanol to evaporate. Record the appearance of the crystals that form.

Observations and results

	Observations
Weight of filter paper	
Appearance of solid in filter paper	
Weight of filter paper with solid	
Weight of sodium chloride solid	
Appearance of crystals in the watch glass	

Processing of results and questions

1. In Part B, what property enabled you to completely separate the sodium chloride and the charcoal mixture?

2. In part C, if you had used water instead of ethanol as the solvent, what would have happened? Explain your answer fully, making reference to your results from part A.

3. a) Calculate the percentage of sodium chloride recovered from the mass of the solid in the weighed filter paper.

   b) Describe a systematic error in this experimental design that would result in less than 100% recovery of the sodium chloride from the original mixture.

5. Where do the crystals first form in the solution?

6. What factors affect the size and shape of the crystals formed?

Experiment 5: Distillation demonstration

Distillation can be used to separate two or more liquids with different boiling points, used in the refining of crude oil. It can also be used to separate a solvent from a solution, as in the distillation of seawater to obtain pure water. In this experiment you will obtain pure water by the distillation of a copper(II) sulfate solution.

Aim

To perform the process of distillation.

Equipment

- balance
- retort stand, boss head and clamp
- Bunsen burner
- test tubes (two small; two large, one with stopper and delivery tube)
- boiling chip (one)
- matches
- graduated cylinder (25 mL)
- copper(II) sulfate solution [$CuSO_4$] 0.1 mol L^{-1} (15 mL)

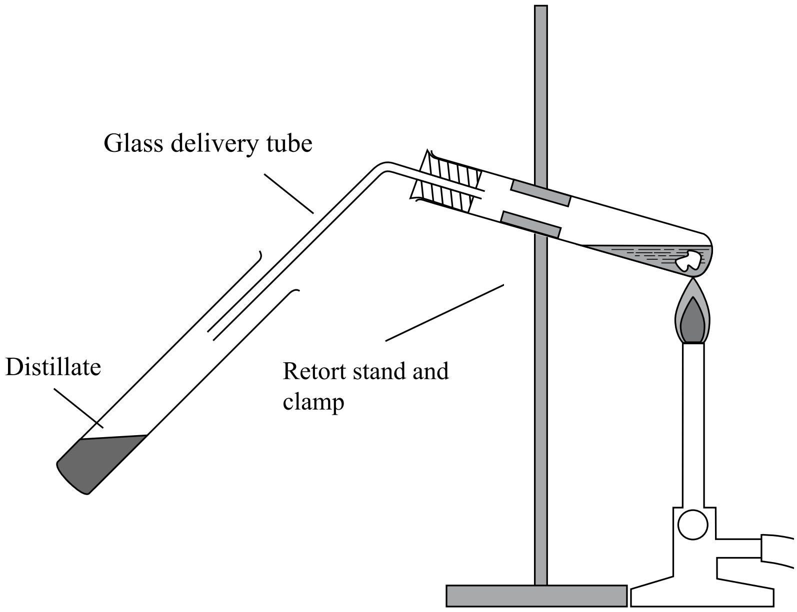


Figure 5.1

Procedure

1. Place about 15 mL of 0.1 mol L^{-1} $CuSO_4$ into a large test tube fitted with a delivery tube as shown in Figure 3.1. Clamp the test tube to a retort stand. Record the colour of the solution.
3. Add a boiling chip and boil the solution using a Bunsen burner. Collect the distillate in a small test tube. Ensure that none of the solution boils over directly into the small test tube.
4. When you have collected a small amount of distillate, view it carefully and write observations.

Observations and results

	Observation
Colour of copper(II)sulfate solution	
Colour of distillate	

Processing of results

1. Draw a flow diagram to show the heat transfer and phase changes in the distillation process.

2. Why is the colour of the distillate different from the copper(II) sulfate solution? Explain.

3. A student completes this experiment but uses sodium chloride instead of copper(II) sulfate. Would there be a difference in the colour of the solution and the distillate?

4. How could the student who used sodium chloride prove with a chemical test there is a difference between the original solution and the distillate?

5. What is the requirement of two liquids if you wish to separate them by distillation?

6. A supply of ethanol is suspected of contamination by methanol. Would it be possible to separate the two by distillation?

Experiment 6: Separation using paper chromatography

In the separation of substances using chromatography, the sample to be separated, liquid or gas, is passed over an inert substance. In this experiment paper chromatography is used to separate the pigments in leaves. The leaf components that are least soluble in water and those that adhere most strongly to the paper will move more slowly. The components that are more soluble in water or adhere less strongly to the paper, travel more quickly up the paper.

Aim

To separate the chlorophyll pigments in leaves.

Equipment

- test tube (2 large)
- capillary tube
- beaker (100 mL)
- sodium chloride solution 1% solution 20 mL
- methylated spirits
- assorted leaves at least 2 types (spinach and geranium and plants with purple foliage are recommended)
- mortar and pestle
- strips of chromatography or filter paper to fit the large test tubes

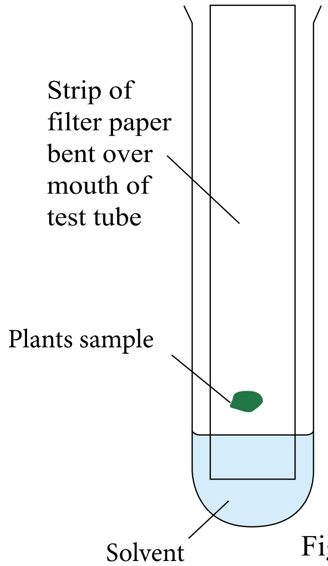


Figure 6.1

Procedure

1. Put several spinach leaves in a pestle with several drops of saline and grind them until they are thoroughly liquefied. (Alternatively spinach leaves could be placed in a blender with saline and blended thoroughly for 5-10 minutes, or until all of the leaves are liquefied.)
2. Using a capillary tube take a small amount of the liquid from the ground spinach (avoid solid pieces). Place carefully on a piece of chromatography paper about 1 cm from the bottom.
3. Put about 1-2 mL of methylated spirits into the clean test-tube.
4. Place the chromatography paper into the test-tube so that the sample being tested is not sitting in the methylated spirits. The diagram above shows the basic set up.
5. Leave the sample for 5 minutes for the components to separate.
6. Repeat steps 1-5 with a second sample of leaves

Observations and results

Draw a diagram of the results of the separation, or paste the dry chromatagrams in the spaces below.

Processing of results

1. Are the components of the extracted plant sample composed of one pure pigment or different pigments?

2. Compare the properties of the plant pigments that enabled them to be separated using paper chromatography.

Questions

1. Research the different pigments that are in leaves. How many different pigments exist and what are their colours?

Pigment	Colour

2. From your research in question 1, try to identify the separated components you have found in this experiment. Base your judgement on the colour of the pigment and the distance they have moved compared with information found in your research.

Experiment 7: Bonding and conductivity

When elements join together to form compounds, the atoms can be held together by ionic or covalent bonds. In ionic substances the atoms achieve a noble gas electron configuration by the transfer of electrons from one atom to another. The positive and negative ions are held together by electrostatic forces to form an ionic lattice. When ionic compounds like aluminium oxide, $A\ell_2O_3$, are molten or dissolved in water the ions can move through the liquid and are therefore able to conduct an electric current.

In covalent bonds the atoms achieve a noble gas electron configuration by sharing electrons between atoms. When molten, covalent substances do not conduct an electric current as there are no mobile charge carriers (such as electrons or ions) that can move through the liquid. When dissolved in water some covalent substances react to produce ions and hence a conducting solution.

In this experiment the electrical conductivities of some substances in the molten state and in aqueous solution will be investigated. From this you should be able to draw some conclusions about the nature of the bonding in these substances.

Part A Demonstration: Electrical conductivity of molten ionic and covalent molecular substances

Equipment PART A

- D.C. power supply (0-12 V)
- Bunsen burner, tripod, pipe clay triangle
- hot plate
- crucible with lid (at least two required)
  ammeter or globe in a socket (6 V, 0.5 A)
  electrical leads (four, two with alligator clips)
- paper towel
- sandpaper
- tongs
- beaker (100 mL)
- wire electrode system (This can be made by pushing two 10 cm lengths of 16-gauge stainless steel wire through a cork so that the wires are about 1 cm apart as shown in figure 7.1.)
- perspex screen to shield against crucible splash
- mat to protect benches from NaOH spills
- sodium hydroxide [NaOH] (5 g) *
- candle wax (5 g) **

SAFETY NOTES:
***   corrosive**
**** Flammable**

- **This demonstration must be conducted in a fume hood**
- **The electrolysis of the molten salt should be done in a fume hood.**
- **The electricity should be switched off as soon as the demonstration is complete.**
- **Be particularly careful handling the molten substances**
- **Care must be taken when heating the candle wax. If it catches fire, use tongs to place the lid on the crucible. This will extinguish the fire.**

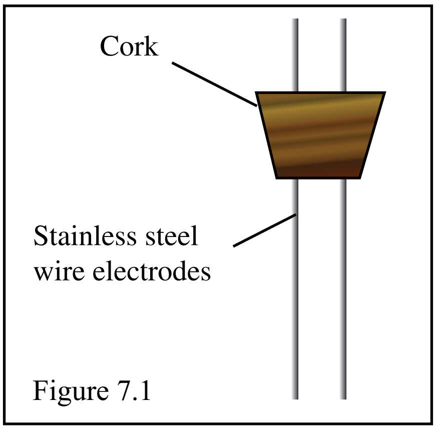


Figure 7.1

Procedure PART A

1. Connect a 6 V D.C. supply, ammeter and/or globe and wire electrode system in series as shown in Figure 7.2.

2. Place some sodium hydroxide pellets in a crucible to a depth of about 0.5 cm. Place the crucible in a pipe clay triangle supported by a tripod.

3. Gently heat the crucible until the sodium hydroxide melts. Do not heat the crucible excessively. Apply only enough heat to keep the sodium hydroxide just molten.

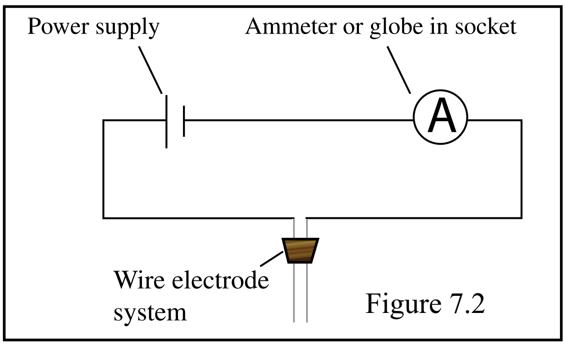


Figure 7.2

4. Place the electrode system into the melt and note the conductivity in terms of the brightness of the globe or the ammeter reading. It is necessary to continue gentle heating as there will be some solidification around the electrodes.

5. Remove the electrodes and clean them thoroughly with some sandpaper.

6. Place some candle wax in a crucible to a depth of 0.5 cm. Place the crucible on a hotplate, melt the candle wax and keep molten while you test its conductivity as per step 4.

7. Remove and clean the electrodes.

8. Record results in table.

9. Clean up when cool.

Processing of results

Part B: Electrical conductivity of liquid and aqueous ionic and covalent molecular substances

Equipment PART B

- baby oil (mineral)
- ethanol (50 ml) **
- distilled water (50 mL)
- 50 mL 0.1 mol L^{-1} solutions of:
  - sucrose [$C_{12}H_{22}O_{11}$]
  - sodium chloride [$NaC\ell$]
  - sodium hydroxide [NaOH] *
  - hydrochloric acid [$HC\ell$] *
- plate electrode system (Made from two stainless steel plates 4 cm × 7 cm. Separate plates by about 0.3 cm, using two plastic spacers, tape together so that it fits inside a 100 mL beaker as shown in figure 7.3)

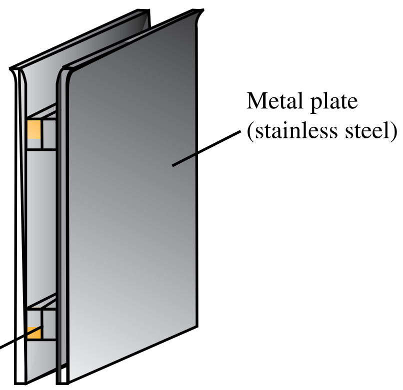


Figure 7.3

DISPOSAL.
Collect all used baby oil into a waste beaker and return to laboratory technicians

Procedure PART B

1. Connect a 6 V D.C. supply, ammeter and/or globe, and plate electrode system in series (similar to the circuit in Figure 7.2).

2. Place 50 mL of baby oil into a 100 mL beaker. Place the plate electrode system into the liquid and note the conductivity in terms of the brightness of the globe or the ammeter reading. Record your results.

3. Remove the electrodes and dry them with a piece of paper towel.

4. Repeat the procedure using, in turn, ethanol, distilled water, 0.1 mol L^{-1} solutions of sucrose, $NaC\ell$, NaOH, and $HC\ell$. Tabulate your results.

Observations and results

Substance	Conductivity
Baby oil	
Ethanol	
Distilled water	
0.1 M sucrose	
0.1M HCℓ	
0.1M NaOH	
0.1M HCℓ	

Processing of results

1. Which of the pure substances (not the solutions) tested in parts A and B conducted an electric current? What can you say about the nature of the bonding present in these substances? List the species that were the charge carriers in each case.

2. List the pure substances tested in the experiment that did not conduct an electric current. What can you say about the type of bonding present in these substances? What species are present in each case?

3. Which of the solutions (sucrose, NaCℓ, NaOH and HCℓ) conducted an electric current? List the species present in each case and state why the solutions conducted or not.

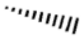
4. In the solid state, both NaCℓ and NaOH are ionic lattices. What happens to these substances when they dissolve in water?

Experiment 8: Molecular models

You can purchase petrol with different octane ratings such as 91, 95 and 98. The octane rating of petrol is its anti-knock rating or detonation resistance. 'Detonation resistance' is about burning more slowly; overly fast burning (detonation) creates a very fast expansion of gas, pushing unevenly on the piston and leading to lower efficiency and possible engine damage. In general, higher octane ratings give slower, smoother combustion. Modern high performance motors require the higher octane fuels. Octane rating is determined by comparing the fuel to a mixture of iso-octane (2,2,4-trimethylpentane, a **constitutional isomer*** of octane) and normal heptane (the straight chain isomer of heptane). Iso-octane is given an octane rating of 100 and heptane is given an octane rating of zero. 91-octane petrol, for example, has the same anti-knock rating as a mixture of 91% (by volume) iso-octane and 9% (by volume) heptane. This does not mean, however, that the petrol contains these hydrocarbons in these proportions.

*****_Constitutional isomers have the same molecular formula but adifferent arrangement of atoms. They were previously called structural isomers (IUPAC name change 2021)._

To draw three dimensional diagrams of models you should consider using the following method to represent bonds in molecules:

Bond Orientation	Symbol
In the plane of the page	—
Into the page	
Out of the page	

Number of carbon atoms in the chain	Prefix
1	meth-
2	eth-
3	prop-
4	but-
5	pent-
6	hex-
7	hept-
8	oct-

Aim

To build molecular models of various hydrocarbons and their isomers.

Equipment

Molecular model kit or polystyrene balls and toothpicks or plasticine and toothpicks

Procedure

A. alkanes

Methane [CH_4]

1. Use the molecular model kit provided to construct a model of the molecule.

2. Carefully observe the model of the molecule. Does it look the same from all sides? Describe the shape.

3. Draw a diagram of the model of your methane molecule.

Ethane [C_2H_6]

4. Remove one hydrogen atom from your model of the methane molecule and replace it with another CH_3 group. This group should be identical to the one you hold. Are the bond angles still the same as for methane?

5. Rotate the ends of the molecule about the C-C bond. Record your observations.

6. Draw a diagram of your ethane model.

Dichloroethane [$C_2H_4C\ell_2$]

7. Remove two hydrogen atoms from your model (from the same carbon atom) of the ethane molecule and replace them with two chlorine atoms.

8. Construct another dichloroethane molecule different from the one you have already made. These two molecules are structural isomers. They have the same molecular formula but different structural formulas, that is, the order of attachment of atoms is different.

9. Draw diagrams or write formula of your two dichloroethane molecules and label them: 1,1-dichloroethane and 1,2-dichloroethane, as appropriate.

Hexane [C_6H_{14}]

10. Construct a model of straight chain hexane.

11. Write the formula for straight chain hexane. ____________________

12. Of the four additional isomers with formula C_6H_{14}, two have a five-carbon chain and two have a four-carbon chain. Write the name of each isomer, using IUPAC nomenclature.

Cyclohexane [C_6H_{12}]

13. Arrange the model of the straight chain hexane so that the two terminal carbon atoms are adjacent. Remove one hydrogen atom from each terminal carbon and join the two carbon atoms to form a cyclic structure. Is the molecule planar? Describe the shape.

____

____

14. Draw a diagram of your cyclohexane molecule.

B. Alkenes

Ethene [C_2H_4]

1. Construct a model of the ethene molecule, remembering it has a double bond.

2. Try to rotate the two ends of the molecule. What prevents rotation about the carbon-carbon double bond?

____

3. Draw a diagram of your ethene molecule.

Dichloroethene [$C_2H_2C\ell_2$]

4. Remove two hydrogen atoms from your model of the ethene molecule and replace them with chlorine atoms
5. Make another model of the ethene molecule and by replacing two hydrogen atoms with two chlorine atoms, construct a different isomer of $C_2H_2C\ell_2$.
6. Construct a third isomer of $C_2H_2C\ell_2$.
7. Draw diagrams of the three dichloroethene isomers and write their names.

Hexene [C_6H_{12}] and Cyclohexene [C_6H_{10}]

8. Draw diagrams of three straight chain isomers of hexene. Draw their structures and write their names.
9. Construct a model of one of these isomers of hexene.
10. Convert this to a model of cyclohexene in the same way in which you constructed a model of cyclohexane.
11. Draw a diagram of your cyclohexene molecule.

Questions

1. Using examples from this activity distinguish between conformational isomerism and *cis-trans* isomerism (now known as *E-Z* or constitutional isomerism).
2. Which groups of aliphatic (straight chain) hydrocarbons display *E-Z* isomerism?
3. Explain why some groups of aliphatic (straight chain) hydrocarbons do not display *E-Z* isomerism.
4. Research differences in the physical properties between two structural isomers and two *E-Z* isomers. Write an account of your findings.

Experiment 9: Reactivity of hydrocarbons

A chemical plant involved in the extraction of non-polar substances from vegetable matter requires an unreactive solvent for the extraction process. As the plant chemist you need to determine the reactivity of various hydrocarbons that could be used.

Three classes of hydrocarbons are alkanes (saturated hydrocarbons containing only single C-C bonds), alkenes (unsaturated hydrocarbons containing a C=C double bond) and aromatics (derivatives of benzene). You already know that hydrocarbons react with halogens and are susceptible to oxidation as evidenced by their combustion.

Aim

1. Investigate the reactivities of the classes of hydrocarbons: alkanes, alkenes and aromatics, under similar conditions, with respect to:
   (a) oxidation with acidified potassium permanganate solution and
   (b) halogenation with bromine water.
2. Determine the class or classes of hydrocarbon that would be suitable to be used as a chemical plant solvent.

Equipment

- test tubes and stoppers (seven)
- 1 mL droppers (five)
- sulfuric acid [H_2SO_4] 2 mol L^{-1} (2 mL)
- bromine water [Br_2] (5 mL)
- potassium permanganate solution [$KMnO_4$] 0.01 mol L^{-1} (4 mL)
- hydrocarbons you will be using are:
  alkane: cyclohexane [C_6H_{12}] (3 mL)
  alkene: cyclohexene [C_6H_{10}] (3 mL)
  aromatic: toluene [$CH_3C_6H_5$] (3 mL)

SAFETY NOTE:
The use of bromine water and toluene is restricted to use in Senior High Schools only. Bromine is a Schedule 7 Dangerous Poison. Toluene is a Category 3 Carcinogen. This entire experiment must be conducted in a fumehood and Part B must be done as a teacher demonstration only.
- **Bromine water is acutely toxic by inhalation, do not breathe the vapour.**
- **If bromine water comes in contact with your skin, immediately wash the affected area with copious quantities of water**
- **Cyclohexane, cyclohexene and toluene are highly flammable and acutely toxic and must be handled with extreme care**
- **Do not let these liquids come in contact with your skin and avoid inhaling their vapours.**

Procedure

A: Reaction of hydrocarbons with acidified permanganate solution

1. Make the acidified permanganate solution. In a test tube place 4 mL of 0.01 mol L^{-1} $KMnO_4$ and 2 mL of 2 mol L^{-1} H_2SO_4.
2. Into three separate, labelled test tubes place 1 mL of cyclohexane, cyclohexene and toluene respectively.
3. Add 1 mL of the acidified $KMnO_4$ solution, prepared in Step 1 above, to each of the test tubes containing the hydrocarbons. Shake each test tube gently and record any change that takes place in the aqueous layer (top layer) over about 5 minutes. Record your observations in the table below.

Observations and results

Hydrocarbon	Observations of any reaction with acidified $KMnO_4$ solution

B: Reaction of hydrocarbons with bromine water

1. Into three separate, labelled test tubes place 1 mL of cyclohexane, cyclohexene and toluene respectively.
2. Add 1 mL of bromine water to each of the test tubes containing the hydrocarbons. Shake each test tube gently and record any colour change

Observations and results

Hydrocarbon	Observations of any reaction with bromine water

Processing of results

1. Using any reactions with the acidified permanganate solution (the oxidising agent), describe the relative reactivity of the three classes of hydrocarbons.

2. Using any reactions with the bromine water describe the relative reactivity of the three classes of hydrocarbons.

3. Write equations for any observed reactions of bromine water with the hydrocarbons.

Questions

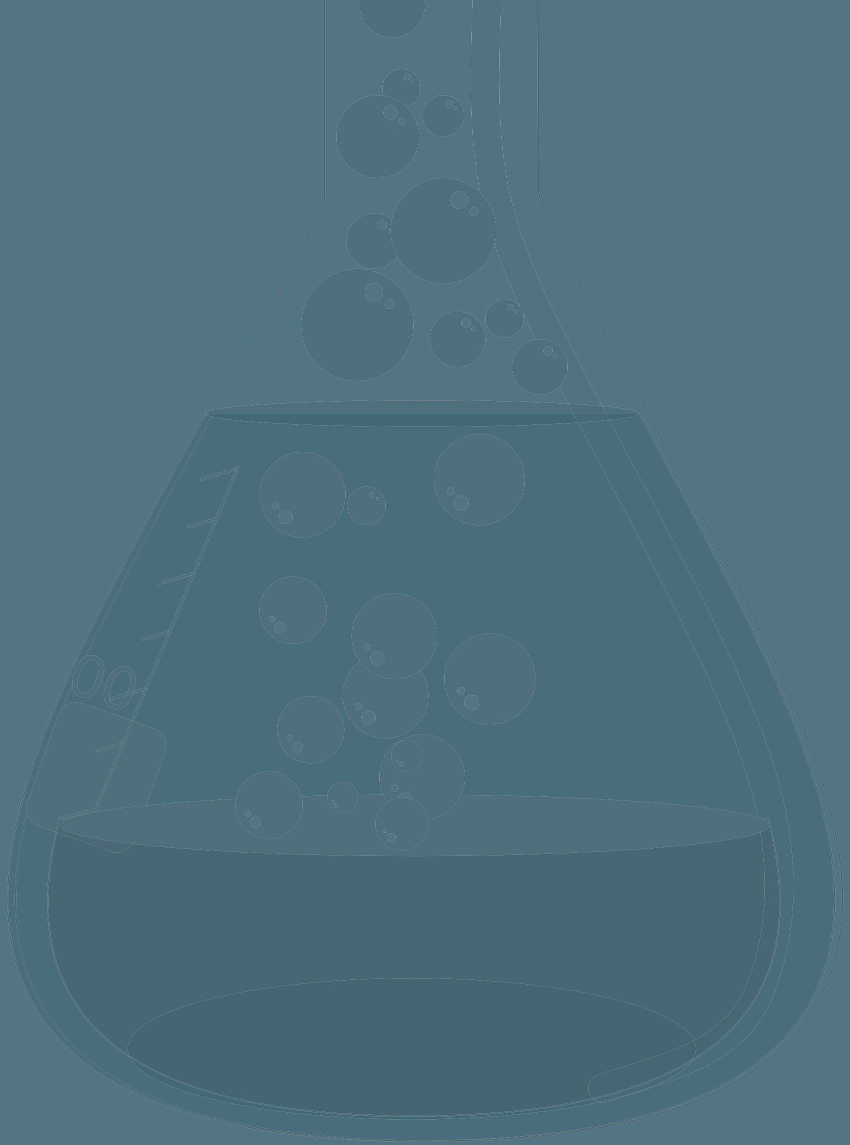
1. Which class or classes of compounds tested would be suitable as a solvent in the chemical plant described? Explain your choice.

Chemical reactions: reactants, products and energy change

Design, conduct, evaluate and communicate investigations into a range of reaction types including energy changes involved in chemical reactions.

Chemical reactions and phase changes involve enthalpy changes, observable as changes in the temperature of the surroundings and/or the emission of light. Reactions releasing energy are described as exothermic and reactions that absorb energy from the surroundings are called endothermic reactions. They can be explained in terms of the Law of Conservation of Energy and the breaking and forming of bonds. Chemical reactions can be represented by chemical equations and heat energy released or absorbed by the system to, or from the surroundings, can be represented in thermochemical equations.

The foods we eat and fuels we use have differences in energy output. Fossile fuels and biofuels also differ in their carbon emissions. How widely we use these fuels and their future use is not only determined by their chemistry, but is influenced by social, economic, cultural and political values.



Experiment 10: Types of chemical reactions

Many of the tests performed for crime scene investigations involve chemical reactions. The results help to provide answers to investigators' questions and links the criminals and the crimes. Whilst you cannot see the reaction happening at the molecular level, usually the results are indicated by something obvious like a colour change or the production of a gas.

During a chemical reaction, chemical substances, or reactants, interact yielding a product that is different from the reactants. During this reaction process, bonds are broken and new ones formed. In chemistry there are many different types of chemical reactions. Chemists use chemical equations as shorthand representations for reactions.

Aim

The purpose of this experiment is to observe several important types of chemical reaction and to write equations for these reactions.

Equipment

- safety glasses
- narrow (small) test tube
- balance
- spatula
- Bunsen burner
- matches
- metal tongs
- wax taper dropper
- beaker (100 mL)
- steel wool
- zinc strips (two) copper strip (one)
- copper(II) carbonate $[CuCO_3]$ (2 g)
- calcium carbonate $[CaCO_3]$ marble chips (2 g)
- magnesium ribbon [Mg] (two 3 cm strips)
- Universal Indicator ™ (5 drops) + chart
- test tubes (one large fitted with rubber stopper and plastic gas delivery tubing, two small)

5 mL solutions of:
- sodium iodide [NaI] 0.1 mol L^{-1}
- sodium hydroxide [NaOH] 2 mol L^{-1} 10 mL
- hydrochloric acid $[HC\ell]$ 2 mol L^{-1}
- limewater $[Ca(OH)_2]$ saturated

30 mL, 0.1 mol L^{-1} solutions of:
- copper(II) sulfate solution $(CuSO_4]$
- sodium hydroxide solution [NaOH]
- hydrochloric acid $[HC\ell]$
- lead (II) nitrate solution $[Pb(NO_3)_2]$

Procedure, observations and results

Decomposition of a metal carbonate by heating

Place a spatula of copper(II) carbonate into a large test tube and fit the test tube with a stopper and delivery tube. Heat the test tube with a Bunsen burner and pass any gas evolved through 5 mL of limewater in another test tube.

SAFETY NOTE:
- **Carefully and gently heat the solid.**
- **Remove the delivery tube from the limewater before removing the heat to avoid sucking the water back into the reaction tube.**

Reactants		Reaction	Products	
Name	Description	Observations	Name	Description
copper(II) carbonate				
Limewater test				

Decomposition of a carbonate with an acid

Place a spatula of calcium carbonate (marble chips) into a large test tube, add 2 mol L^{-1} hydrochloric acid to a depth of about 2-3 cm and fit the test tube with the stopper and delivery tube as before. Again, note the effect of any gas evolved on the limewater solution.

Reactants		Reaction	Products	
Name	Description	Observations	Name	Description
Limewater test				

Oxidation of a metal

For safety reasons this experiment will be done as a teacher demonstration or video demonstration.

SAFETY NOTE:
- **Do not look directly at the burning magnesium.**

Reactants		Reaction	Products	
Name	Description	Observations	Name	Description

Reaction of a reactive metal with a dilute acid

Place another 3 cm strip of magnesium in a small test tube and add 2 mol L^{-1} hydrochloric acid to a depth of about 3 cm. Note the reaction and collect any gas evolved by inverting another small test tube and holding it directly above the reaction tube. Test the gas evolved by placing a lighted taper in the inverted test tube.

Reactants		Reaction	Products	
Name	Description	Observations	Name	Description
‘Pop’ test				

Chemical reactions: reactants, products and energy change

Precipitation reactions

1. Place 2-3 mL of 0.1 mol L^{-1} $Pb(NO_3)_2$ solution into a test tube and add about the same volume of 0.1 molL^{-1} NaI solution.
2. Place 2-3 mL of 0.1 mol L^{-1} $CuSO_4$ solution into a test tube and add about the same volume of 1 mol L^{-1} NaOH solution.

Reactants		Reaction	Products	
Name	Description	Observations	Name	Description

Metal displacement reactions

1. Place about 25 mL of 0.1 mol L^{-1} of $CuSO_4$ solution into a 100 mL beaker then place a freshly cleaned zinc strip.

Reactants		Reaction	Products	
Name	Description	Observations	Name	Description

2. Place about 25 mL of 0.1 mol L^{-1} of $Pb(NO_3)_2$ solution into a 100 mL beaker then place a freshly cleaned zinc strip.

Reactants		Reaction	Products	
Name	Description	Observations	Name	Description

3. Place about 25 mL of 0.1 mol L^{-1} of $AgNO_3$ solution into a 100 mL beaker then place a freshly cleaned copper strip.

Reactants		Reaction	Products	
Name	Description	Observations	Name	Description

Neutralisation reaction (reaction of an acid and a base)

Place 5 mL of 0.1 mol L^{-1} $HC\ell$ into a small test tube and add 2 drops of Universal Indicator ™. Using a dropper, gradually add 0.1 mol L^{-1} NaOH solution until about 6 mL of the base has been added. Note any colour changes that take place.

Reactants		Reaction	Products	
Name	Description	Observations	Name	Description

Processing of results

1. Write general word equations and balanced chemical equations for each of the reactions observed.

Decomposition of a metal carbonate by heating:

Decomposition of a metal carbonate with an acid:

Oxidation of a metal:

Reaction of a reactive metal with a dilute acid:

Precipitation reactions:

Metal displacement reactions:

Neutralisation reaction) reaction of an acid and a base)

2. Write a balanced chemical equation for the following reactions:
   (a) heating magnesium carbonate;

   (c) burning of iron in pure oxygen;

   (d) adding dilute sulfuric acid solution to calcium;

   (f) the reaction between lead metal and copper(II) sulfate solution to produce lead(II) sulfate and copper metal;

   (g) the reaction between sulfuric acid solution and potassium hydroxide solution to form potassium sulfate and water

Experiment 11: Exothermic and endothermic processes

Chemical reactions are accompanied by the release or absorption of energy. Reactions that release energy are described as exothermic. Exothermic reactions lose heat to the surroundings, which become warm. Reactions that absorb energy are described as endothermic. Endothermic reactions gain heat and the surroundings become cold. The energy released in chemical reactions was previously stored as chemical potential energy in the reactants. This stored chemical potential energy is called the heat content or enthalpy of a substance.

Aim

The purpose of this experiment is to investigate different chemical processes and, based on the temperature changes you observe, classifying them as endothermic or exothermic.

Equipment

SAFETY NOTE:
* **Sodium hydroxide and barium hydroxide are corrosive. If contact with the skin occurs wash with large quantities of water for 20 minutes.**

- thermometer (-10 to 110 °C)
- 100 mL beaker
- steel wool
- plastic teaspoon or spatula
- distilled water
- saturated copper(II) sulfate solution [$CuSO_4$] (25 mL)
- The following solids (about a teaspoonful of each):
- sodium hydroxide [NaOH]*
- ammonium chloride [$NH_4C\ell$]
- sodium acetate [NaCH3COO]
- sodium chloride [$NaC\ell$]
- barium hydroxide [$Ba(OH)_2$]*
- ammonium thiocyanate [NH_4SCN]

Procedure Part A: Solution processes

1. Place about 30 mL of water in a 100 mL beaker and measure its temperature. Add one teaspoonful of sodium hydroxide pellets, stir gently, and record any change in the temperature.

2. Repeat the procedure using, in turn, a teaspoonful of ammonium chloride, sodium acetate and sodium chloride. Tabulate your results.

Results

Solute	Initial temperature °C	Final temperature °C	Endothermic or exothermic

Procedure Part B: reaction between ammonium thiocyanate and barium hydroxide

1. Place about half a teaspoonful each of solid ammonium thiocyanate and solid barium hydroxide into separate test tubes and measure and record their temperatures

2. Transfer the ammonium thiocyanate solid to the test tube containing the solid barium hydroxide. Stopper the test tube and shake gently until a solution is formed.

3. Remove the stopper and smell cautiously. Try to identify one of the products. Measure and record the temperature of the solution.

Results

Solute	Initial temperature °C	Final temperature °C	Endothermic or exothermic

Procedure Part C: reaction between iron and copper(II) sulfate solution

1. Place about 25 mL of saturated copper(II) sulfate into a 100 mL beaker and record the temperature.
2. Place a cylinder of steel wool about 1 cm thick into the solution and hold it under the surface with the thermometer. Record any evidence of reaction and change in temperature.

Results

Solute	Initial temperature °C	Final temperature °C	Endothermic or exothermic

Processing of results

Write equations for the processes or reactions in parts A, B and C and classify each reaction as either exothermic or endothermic by placing heat on the appropriate side of the equation.

Experiment 12: Foods and fuels

Most of the energy our body needs comes from fats and carbohydrates.
Carbohydrates are broken down in the intestines to glucose. The glucose is transported in the blood to cells where it is oxidised to produce CO_2, H_2O and energy:

$C_6H_{12}O_6(s) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(\ell)$ $\Delta H = -2\ 816$ kJ

The breakdown of fats also produces CO_2 and H_2O.
Any excess sugar in the body is stored as fats. About 100 kJ per kilogram of body weight per day is required to keep the body functioning at a minimum level.

Compound	Fuel Value (kJ g^{-1})
Fats	38
Carbohydrates	17
Proteins	17

Most chemical reactions used for the production of heat are combustion reactions. The energy released when 1 g of material is burnt is called its 'fuel value'. Since all heats of combustion are exothermic, fuel values are reported as positive.

Not all chemical reactions are combustion reactions, yet every chemical reaction has a particular heat of reaction. This is usually expressed as the molar heat of reaction. The molar heat of oxidation of glucose, shown earlier, is 2 816 kJ. How are these values determined?

Example

Consider the neutralisation reaction between hydrochloric acid and sodium hydroxide solution:

$HC\ell(aq) + NaOH(aq) \rightarrow H_2O(\ell) + NaC\ell(aq)$ or $H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$

The molar heat of reaction for this reaction is the heat released when one mole of hydrochloric acid is neutralised with one mole of sodium hydroxide solution.

Aim

The purpose of this investigation is to experimentally determine the molar heat of this neutralisation reaction to measure the temperature increase that occurs when solutions of hydrochloric acid and sodium hydroxide are mixed.

Equipment

- polystyrene cup
- graduated cylinder (100 mL)
- thermometer (-10 to 110 °C)
- hydrochloric acid [$HC\ell$] 1 mol L^{-1} (50 mL)
- sodium hydroxide solution [NaOH] 1 mol L^{-1} (50 mL)

Procedure

1. Place 50 mL of 1 mol L^{-1} hydrochloric acid solution in the polystyrene cup and measure its temperature.

2. Measure 50 mL of 1 mol L^{-1} sodium hydroxide solution into a measuring cylinder and record its temperature.

3. Add the sodium hydroxide solution to the hydrochloric acid, stir briefly with the thermometer, and note the maximum temperature reached.

Observations and results

Reactant	Initial temperature °C	Final temperature °C
Average		

Processing of results and questions

1. Using the average temperature of the two reactants as the initial temperature, calculate the change in temperature of the solution.

2. Calculate the amount of heat energy released, as follows. Assume the heat absorbed by the cup and small losses to the surroundings are negligible.

Heat energy released	=	mass of solution	×	specific heat of solution	×	temperature change
	=	0.100 kg	×	(4.18×10^3 J K^{-1} kg^{-1})	×	(ΔT K)
	=	_______J				

3. In question 2 you calculated the heat released when 50 mL of 1 mol L^{-1} $HC\ell$ is neutralised with 50 mL of 1 mol L^{-1} NaOH.
   (a) Calculate the molar heat of reaction.

   (b) Is this reaction exothermic or endothermic?

   (c) Write the molar value of ΔH for the neutralisation reaction using the appropriate sign.

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Rates of chemical reactions

Design, conduct, evaluate and communicate investigations into chemical reactions, including the prediction and identification of products and the measurement of the rate of reaction.

Understanding the chemistry of reactions enables scientists to manipulate the rate of reactions. Reactions can be sped up to save time and money in production processes, while it is important for others such as food spoilage reactions to be slowed down. Catalysts play an important role in many industrial processes. They are used to increase or decrease the rates of reactions to make them more economical. Motor vehicles have catalytic converters to reduce the amount of pollution from carbon monoxide, unburnt fuel and nitrogen oxides that are emitted.

Investigation 13: Measuring the rate of reaction

Some reactions are fast, while other reactions are slow. The combustion reaction of an explosion is fast, while the spoiling of food is slow. The corrosion of a metal is even slower and can take years.

How slow or fast a reaction proceeds, is a measure of its rate. Rate of a chemical reaction can be measured by how fast a reactant is consumed or by how fast a product is formed. Rate has the units per second (s^{-1}).

Task

To plan and conduct, an investigation to measure the rate of reaction of calcium carbonate (marble chips, $CaCO_3$) with hydrochloric acid solution ($HC\ell$).

Equipment

- marble chips [$CaCO_3$]
- 250 mL conical flask
- cotton wool
- delivery tube
- stopwatch
- electronic balance (0.01g)
- 25 mL and 50 mL measuring cylinders or gas syringe
- 1 L beaker or trough
- one holed rubber stopper to fit conical flask
- 50 mL of 2 mol L^{-1} hydrochloric acid [$HC\ell$]*
- graph paper

SAFETY NOTE:
- **Check your plan with your teacher before you commence.**

* **Hydrochloric acid is corrosive**

Pre-planning for the investigation

Check your pre-planning ideas with your teacher before writing your investigation plan.

1. Write an equation for the reaction.

____

2. List variables that you could measure to determine the rate of this reaction.

____

____

____

3. Select one variable to measure and draft an aim or hypothesis for your investigation. Explain your selection.

____

____

____

____

____

4. Sketch your equipment setup showing how you will carry out your investigation and measure the rate of the reaction.

Plan and conduct the investigation

Write your report on a separate piece of paper.

Your investigation plan should include

- Aim or hypothesis
- List of the chemicals and equipment (provided below)
- Safety notes
- Procedure
    - o Labelled diagram of the setup
    - o List of steps to follow
- Results and observations:
    - o Collect and record observations and results of the reaction in a table as you proceed.
    - o If time and resources allow, replicate the data collection.
- Processing of results: Draw a graph of your results, identify any trends, and explain your data. (*hint: rate has the units s^{-1})
- Conclusion: Write a conclusion based on your aim or hypothesis
- Evaluation:
    - o Evaluate the effectiveness of your procedure and describe any modifications you would make to improve the accuracy of the results and your organisation of the investigation.
    - o Discuss whether the data are sufficient to support your conclusion.

Experiment 14: Concentration and rate of reaction

A study of the factors that influence rates of reaction is important in developing an understanding of how environmental, biological and industrial processes proceed. Different chemical reactions occur at different rates. Some reactions, such as many neutralisation and precipitation reactions, occur very rapidly. Other reactions, such as the air oxidation of vitamin C and the rusting of iron, are relatively slow. In this experiment you will investigate how one factor, concentration, can affect the rate of reaction.
The reaction between sodium thiosulfate and hydrochloric acid can be represented by the following balanced ionic equation:

$S_2O_3^{2-}(aq) + 2\,H^+(aq) \rightarrow S(s) + SO_2(aq) + H_2O(\ell)$

The time taken for the precipitate of sulfur to obscure a cross marked on a piece of paper placed under the reaction flask is used as a measure of the reaction rate. The less time taken for the cross to 'disappear', the faster is the rate of the reaction. In this activity you will investigate the rate of this reaction as a function of the sodium thiosulfate concentration.

Equipment

- sodium thiosulfate solution [$Na_2S_2O_3$] 0.25 mol L^{-1} (150 mL)
- hydrochloric acid [$HC\ell$] 2 mol L^{-1} (30 mL)
- distilled water
- graduated cylinders (10 mL, 25 mL and 100 mL)
- conical flasks (two 100 mL)
- stopwatch
- white paper with black texta cross

SAFETY NOTE:
- **Avoid inhaling sulfur dioxide evolved while waiting for the dot to disappear.**
- **If asthmatic, carry out the experiment in a fume cupboard.**

Procedure

1. Place 45 mL of 0.25 mol L^{-1} $Na_2S_2O_3$ in a 100 mL conical flask. Put the flask over a cross marked on a piece of white paper as shown in Figure 14.1 below.

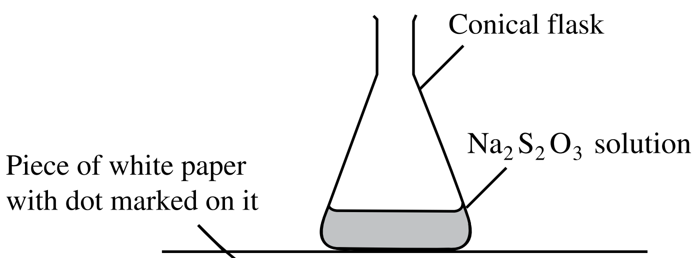


Figure 14.1

2. Add 5 mL of 2 mol L^{-1} $HC\ell$ and briefly agitate to ensure mixing of the reactants. Start a stopwatch at the moment of addition.

3. Note and record in the table on the following page the time taken for the cross to 'disappear' when it is viewed through the solution from directly overhead. The formation of solid sulfur causes the cross to be obscured.

4. Repeat the experiment using various sodium thiosulfate concentrations, made up as indicated in the results table.

Results

Vol. of 0.25 mol L^{-1} $Na_2S_2O_3$ (mL)	Vol. of H_2O added (mL)	Total volume after adding and mixing 5 mL HCℓ	Concentration of $Na_2S_2O_3$ (mL) on mixing (mol L^{-1})	Time for cross to disappear (s)	$\frac{1}{time}$ (s^{-1})
45	0	50	0.225		
35	10	50			
25	20	50			
15	30	50			
5	40	50			

Processing of results and questions

1. Calculate $\frac{1}{time}$ for each experiment and enter the results into the table.

2. Plot a graph of $\frac{1}{time}$ (a measure of the reaction rate) against sodium thiosulfate concentration on the grid below.

3. What effect does the concentration of sodium thiosulfate have on the reaction rate?

4. If the concentration of sodium thiosulfate is doubled, what happens to the rate of the reaction?

5. Identify two random errors in this experiment. How could these errors be minimised?

6. Identify two systemic errors in this experiment. How could these errors be minimised?

7. State two ways that you could improve this experiment's reliability.

8. State two ways that you could improve this experiment's validity.

9. Explain, in terms of collision theory, why increasing the concentration of the thiosulfate ion would have this effect on the reaction rate.

10. What is the advantage of representing the reaction rate as $\frac{1}{\textit{time for cross to disappear}}$ instead of time to disappear?

Experiment 15: Temperature and the rate of reaction.

Temperature can have a significant impact on the rate of reactions. We use refrigeration to reduce rates of spoiling of food. We heat other reactions to increase their rate.
The reaction observed is the same as in the previous experiment, between sodium thiosulfate and hydrochloric acid.

$S_2O_3^{2-}(aq) + 2\,H^+(aq) \rightarrow S(s) + SO_2\,(aq) + H_2O(\ell)$

The time taken for the sulfur precipitate to obscure a cross marked on a piece of paper placed under the reaction flask is used as a measure of the reaction rate. The less time taken for the cross to 'disappear', the faster is the rate of reaction.

Aim

To investigate the effect of temperature on the rate of reaction.

Equipment

- hydrochloric acid [HCℓ] 2 mol L^{-1} (30 mL)
- distilled water conical flask (100 mL)
- test tube
- thermometer (-10 to 110 °C)
- stopwatch
- Bunsen burner, tripod and gauze mat
- paper towel
- beaker (500 mL) to be used as a water bath
- sodium thiosulfate solution [$Na_2S_2O_3$] 0.25 mol L^{-1} (150 mL)
- graduated cylinders (10 mL, 25 mL and 100 mL)

SAFETY NOTE:
- **Avoid inhaling sulfur dioxide evolved while waiting for the cross to disappear. If asthmatic, carry out the experiment in a fume cupboard.**

Procedure

1. Place 15 mL of 0.25 mol L^{-1} $Na_2S_2O_3$ and 30 mL of water into a 100 mL conical flask.
2. Place 5 mL of 2 mol L^{-1} HCℓ into a test tube.
3. Heat the solutions separately in a water bath until they are slightly above 20 °C. Remove the flask from the water bath and dry the outside of the flask with paper towel.
4. Put the flask over a cross marked on a piece of white paper and add the acid to the thiosulfate solution. Start a stopwatch at the moment of addition.
5. Note the time taken for the cross to 'disappear' and record the average temperature during the reaction.
6. Repeat the procedure for temperatures of about 30, 40, 50 and 60 °C. In each case you should measure the average temperature for the reaction.
7. Record the temperatures at the start and end of each experiment, and the time taken for the cross to disappear.

Results

Effect of temperature on reaction time

Temperature at the start of the reaction (T_1)	Temperature at the end of the reaction (T_2)	Average temperature $\frac{T_1+T_2}{2}$	Time for the cross to disappear	1/time

Processing of results and questions

1. Complete the results table by calculating the average temperature and 1/time for each sample.
2. Plot a graph of 1/time against temperature.

3. What effect does increasing the temperature have on the reaction rate?

4. Suggest two factors that contribute to the change in reaction rate with temperature.

5. Identify two random errors in this experiment. How could these errors be minimised?

6. Use collision theory to explain the effect of temperature on reaction rate. Include in your answer a sketch of the Maxwell-Boltzmann distribution of particle speed, and a reaction coordinate vs energy diagram.

Experiment 16: Catalysts and rate of reaction

Catalysts are substances that increase the rate of chemical reactions but are not consumed in these reactions. Catalysts therefore take part in chemical reactions but are regenerated at the end of the reactions.

Catalysts are important in many industrial processes such as the manufacture of ammonia and sulfuric acid and in biochemical processes catalysed by biological catalysts called enzymes.

Aim

Investigate the effect of catalysts on the rate of the three different reactions:

A. The reaction between solutions of acidified permanganate ion and oxalic acid

In this reaction the purple permanganate ion reacts with oxalic acid to form carbon dioxide and the very pale pink manganese(II) ion.
The equation is:

$2\ MnO_4^-\ (aq) + 5\ H_2C_2O_4\ (aq) + 6\ H^+(aq) \rightarrow 2\ Mn^{2+}(aq) + 10\ CO_2\ (g) + 8\ H_2O(\ell)$
purple pale pink/colourless products

B. The decomposition of hydrogen peroxide

Hydrogen peroxide is unstable and slowly decomposes over time according to the equation:

$2\ H_2O_2(aq) \rightarrow O_2(g) + 2\ H_2O(\ell)$

C. The reaction between a solution of tartrate ion and hydrogen peroxide

In this reaction hydrogen peroxide oxidises the tartrate ion to carbon dioxide according to the equation:

$C_4H_4O_6^{\ 2-}(aq) + 5\ H_2O\ (aq) + 2\ H^+(aq) \rightarrow 4\ CO_2\ (g) + 8\ H_2O(\ell)$

Equipment

- matches
- Bunsen burner
- sulfuric acid [H_2SO_4] 2 mol L^{-1} (60 mL)
- potassium permanganate solution [$KMnO_4$] 0.02 mol L^{-1} (10 mL)
- saturated solution of oxalic acid [$H_2C_2O_4$] (10 mL)
- saturated solution of manganese(II) sulfate [$MnSO_4$] (1 mL)
- fresh hydrogen peroxide solution [H_2O_2] 6% (50 mL)
- potassium sodium tartrate [$KNaC_4H_4O_6$] (about 1 g)
- manganese(IV) oxide [MnO_2] (a few grains)
- cobalt(II) chloride-6-water solution [$CoC\ell_2.6H_2O$] 10% solution (1 mL)
- $NaC\ell$ solution (1mol L^{-1} in dropper bottle)
- distilled water
- beakers (two 250 mL and two 100 mL) thermometer (-10 to 110 °C)
- plastic teaspoon

Procedure, observations and results

A. **reaction between solutions of acidified permanganate ion with oxalic acid**

1. In each of two 250 mL beakers place 25 mL of 2 mol L^{-1} H_2SO_4 , 125 mL of distilled water, 3 mL of 0.02 mol L^{-1} $KMnO_4$ and 3 mL of saturated oxalic acid solution. Describe any evidence of reaction.

2. To one of the beakers add about 1 mL of saturated manganese(II) sulfate solution and record your observations.

B. **Decomposition of hydrogen peroxide**
1. Place about 20 mL of approximately 6% hydrogen peroxide solution into each of two 100 mL beakers. Is there any evidence of reaction? ____________
2. Into one of the beakers add a few grains of manganese(IV) oxide (MnO_2) and record your observations.

____

____

C. **Reaction between a solution of tartrate ion and hydrogen peroxide**
1. Dissolve 1 g of potassium sodium tartrate in 10 mL of water and add 10 mL of 6% hydrogen peroxide. Add 5 drops of 1 mol L^{-1} $NaC\ell$ to the mixture.
2. Heat the solution to 65 °C and remove the Bunsen burner. Is there any evidence of reaction? ____________
3. Add about 1 mL of 10% cobalt(II) chloride solution and carefully observe any changes.

____

____

Processing results

1. Identify the species acting as a catalyst in each reaction.
   A. ____
   B. ____
   C. ____
2. In reaction C,
   (a) what evidence is there that the catalyst was not consumed in the reaction?

   ____

   ____

   (b) Is there any evidence that the catalyst took an active part in the reaction?

   ____

   ____

Investigation 17: Decomposition of hydrogen peroxide

Some contact lenses are sterilised with hydrogen peroxide. A metal disc acts as a catalyst to speed up the reaction in which the hydrogen peroxide sterilises the contact lenses. The catalyst also ensures that all the hydrogen peroxide has decomposed to form oxygen gas and water before the contact lens wearer inserts the clean lens into their eye. Several metal compounds also speed up the rate at which hydrogen peroxide decomposes to form oxygen gas and water. The equation for the reaction is:

$$2\ H_2O_2(aq) \rightarrow O_2(g) + 2\ H_2O(\ell)$$

The rate of this reaction can be measured in terms of the rate at which oxygen gas is released and collected using the equipment shown in Figure 17.1

Task

To investigate factors that affect the rate at which hydrogen peroxide reacts to form oxygen gas and water.

Equipment

- test tube rack
- plastic syringe (30-50 mL)
- dropper
- stopwatch
- electronic balance
- stoppered side arm test tube with the side arm connected with a tightly fitting 30-40 cm length 5 mm diameter plastic hose to a 30 mL plastic syringe (see Figure 17.1)
- hydrogen peroxide [H_2O_2] 10 volume (100 mL)
- potassium iodide [KI] (about 1 g)
- 10 mL graduated cylinder or graduated pipette and pipette filler

groups investigating the amount of catalyst will also need:
- potassium iodide [KI] (another 2 g)

groups investigating the type of catalyst will also need:
- manganese(IV) oxide [MnO_2] (about 1 g)
- liver (about 1 g)

groups investigating the effect of temperature will also need:
- test tubes (eight) beakers (four 250 mL) ice (100 g)
- kettle to produce hot water thermometer (-10 to 110 °C)

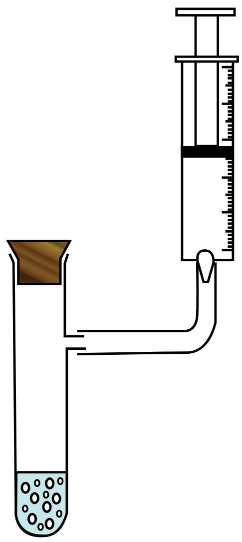


Figure 17.1: Decomposition of hydrogen peroxide

Preliminary trial - measuring the rate of decomposition of H_2O_2

Conduct the following trial to become familiar with the reaction and the equipment you will use. The trial illustrates how the collection of oxygen gas can be used to measure the rate of the reaction.

1. Weigh out about 0.3 g of KI onto a filter paper.
2. Set up the equipment shown in the diagram with the syringe plunger pushed in.
3. Transfer the weighed KI to the test tube.
4. Measure 5 mL of 10 vol H_2O_2 into a graduated cylinder or graduated pipette.
5. Quickly transfer the H_2O_2 into the test tube, stopper firmly, start the stop clock and swirl the contents in the test tube to ensure mixing and dissolving of the KI.
6. Observe and record the rate at which oxygen gas is collected in the syringe.

Planning the investigation

1. List the variables (factors) that could affect the rate of decomposition of hydrogen peroxide.

2. Plan an investigation to determine how one of these factors affects the rate of decomposition of hydrogen peroxide to form oxygen gas and water. Write an hypothesis for your investigation.

3. Fill in the table below.

	Variable/s	Unit/s	How the variable will be measured
Independent variable			
Dependent variable			
Factors kept constant (controlled variables)			

4. Briefly outline your preliminary plan to test the hypothesis. List the chemicals and equipment you need and identify the safety requirements. Remember that you may need to modify your plan after your preliminary trials.

SAFETY NOTE:

- **Check your plan with your teacher before you comence.**
- **Handle the hydrogen peroxide solution carefully as it can cause burns.**
- **Manganese(IV) oxide (MnO_2) is a more effective catalyst than KI for this reaction; use a smaller mass to prevent too rapid a production of oxygen.**

Conducting the investigation - preliminary trials

1. Carry out preliminary trials to refine your technique and to determine the ranges for the variables involved.
2. Describe what you learned from the preliminary trials and any modifications you make to your initial plan. You may have determined factors such as the range over which you will collect data, the number of trials to conduct, how to accurately use the equipment, how to work cooperatively together and so on.

Conducting the investigation - collecting data

Carry out the investigation and record your results in a table

Processing of results

1. Graph your data (plot the independent variable on the horizontal axis).
2. Discuss your data and relate your findings to your hypothesis.

Evaluating the investigation

Evaluate the effectiveness of your procedure and describe any modifications you would make to improve it. You may discuss factors that would improve the accuracy of your results such as sample size and selection, measurement errors and the control of variables. As well, you may address more general organisational factors such as the allocation of tasks among group members and the nature of the apparatus and how it was set up.

Intermolecular forces and gases

Investigate the properties of materials such as electrical conductivity, solubility, melting and boiling points, to understand that the properties are determined by the type of bonding and intermolecular forces. Understanding intermolecular forces enables the separation, testing and identification of many substances, including gaseous materials. Chromatographic techniques, for example, have a wide range of analytical and forensic applications. They are used to determine the components of a wide range of mixtures in various settings.

Chromatographic techniques, include thin layer chromatography (TLC), gas chromatography (GC), and high performance liquid chromatography (HPLC). The decision to use a particular chromatographic technique depends on a number of factors, including the properties of the substances being separated, the amount of substance available for analysis and the sensitivity of the equipment.

Investigate the behaviour of gases, and use the Kinetic Theory to predict the effects of changing temperature, volume and pressure in gaseous systems.

Experiment 18: Intermolecular forces

When substances evaporate, they absorb energy from the surroundings and produce a cooling effect (temperature decrease). This is shown by the general equation below:

$$\text{liquid} + \text{energy} \rightarrow \text{gas}$$

Evaporation is the change of state from liquid to gas that takes place at any temperature below the boiling point. (Figure 18.1)

The magnitude of the temperature decrease is related to the strength of the intermolecular forces present between the liquid molecules.

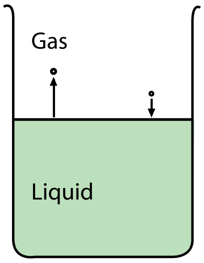


Figure 18.1: Evaporation occurs at the surface of a liquid.

Aim

To compare the cooling effects of pentane, methanol, ethanol, propan-1-ol and butan-1-ol and draw conclusions about the relative strengths of their intermolecular forces.

Equipment

- temperature sensors (two or more depending on the number of sensors that the interface can accommodate)
- computer or graphic calculator loaded with temperature program, computer or graphic calculator interface
- If you don't have a computer and sensing equipment, this experiment can be done using four thermometers.

- 10 mL of:
  pentane*
  methanol*
  ethanol*
  propan-1-ol*
  butan-1-ol*
- filter paper squares (five, approximately 2 cm × 2 cm)
- small elastic bands (fibre, one for each temperature sensor)
- test tubes (five small)
- test tube rack

SAFETY NOTE

***Flammable**

- **Alcohols and hydrocarbons are highly flammable liquids and should be kept away from naked flames.**
- **Butan-1-ol and Propan-1-ol are corrosive to skin. Wear gloves.**
- **Methanol is acutely toxic by ingestion, inhalation and in contact with skin. Wear safety glasses and conduct experiment in a fumehood.**
- **Severe Health Hazard & Environmental Hazard**

Procedure

1. List the five liquids in the 'liquid' column of the results table.
2. Ensure that the temperature sensors are connected to the interface and that the interface is connected to a computer or graphic calculator. Check that the temperature program has been loaded.
3. Check that the temperature sensors have been calibrated. If this needs to be done, follow the software instructions.
4. Wrap a small piece of filter paper (2 cm × 2 cm) around each of the temperature sensors and secure with a small rubber band.
5. Place each of the liquids into a different test tube each to a depth of approximately 3 cm.
6. Place one sensor in the hexane for approximately 45 seconds to ensure the filter paper is fully saturated. Record the initial temperature of the hexane.
7. Remove the sensor from the test tube and place on the bench so that the filter paper end projects over the edge of the bench. It can be fastened with tape if needed. Allow the temperature to reach a minimum and then record in the results table.
8. Calculate the change in temperature that occurred when the hexane evaporated.
9. Repeat steps 6-8 for the remaining four liquids.

Results

Liquid	Initial Temperature (°C)	Final Temperature (°C)	Change in temperature
pentane			
methanol			
ethanol			
propan-1-ol			
butan-1-ol			

Processing of results

1. Rank the liquids in terms of their cooling effect, from greatest to least.

2. Explain how the cooling effect (temperature change) relates to the rate of evaporation of the liquid.

Questions

1. Which of the liquids had the strongest intermolecular forces? Which had the weakest? Use data gathered to support your answer.

2. Two of the liquids, pentane and butan-1-ol have very similar molar masses, 72.15 g mol L^{-1} and 74.12 g mol L^{-1} respectively, but they have very different cooling effects. Explain this difference in ΔT based on their intermolecular forces.

3. All of the alcohols contain hydrogen bonding between molecules yet have vastly different cooling effects. Using the results of this experiment explain the difference between the strength of the intermolecular forces between molecules of methanol and molecules of butan-1-ol.

Experiment 19: Bonding and solubility

Processing our resources often involves solvent chemistry. Bauxite, for example, is first mixed with sodium hydroxide solution. The soluble aluminium salts are dissolved, while most impurities are insoluble in the basic (caustic) solution and are separated as insoluble solids.
Ethanol, a renewable resource, is an additive to our octane petroleum fuels. This is possible because ethanol is soluble in petrol.
The solubility of a substance depends on the solvent being used to dissolve it. For example, sodium chloride is quite soluble in water but is insoluble in paraffin oil. Conversely, petrol is insoluble in water but is quite soluble in paraffin oil.
In this experiment water and paraffin oil, two very different solvents, are used.

Aim

To compare the effectiveness of water and paraffin oil as solvents for solutes with different types of bonding.

Equipment

SAFETY NOTE:
- **Oil must not be poured down the sink.**

- test tube rack
- test tubes (nine)
- beaker (100 mL)
- pipette to transfer oil
- distilled water in wash bottle
- paraffin oil (baby oil) (10 mL)
- popsticks or spatulas to transfer crystals (5×)
- a few crystals of the ionic substances:
  sodium chloride [$NaC\ell$]
  copper(II) sulfate-5-water [$CuSO_4.5H_2O$]
  calcium carbonate [$CaCO_3$]
- a few crystals/mL of the following polar covalent molecular substances:
  sucrose [$C_{12}H_{22}O_{11}$]
  urea [NH_2CONH_2]
  ethanol [CH_3CH_2OH] (1 mL)
- a few crystals/mL of the following non-polar covalent molecular substances:
  iodine [I_2]
  naphthalene [$C_{10}H_8$]
  sunflower oil (1 mL)

Procedure

1. Record the solubility of the nine solutes in water and paraffin oil. Record your results as insoluble (i), slightly soluble (ss) or soluble (s) in the results table.
2. Place 1 mL of paraffin oil into each of nine test tubes.
3. Test the solubility of sodium chloride by adding a few crystals (about a matchheads worth) to one of the test tubes of paraffin oil. Shake the test tube gently and record the results.
4. Repeat this procedure for each of the solutes in turn. For liquid solutes add about 2-3 drops to the paraffin oil.
5. Discard the contents of the test tubes as directed by your teacher and clean the test tubes thoroughly.
6. Place 1-2 mL of distilled water into the cleaned test tubes and test the solubility of each of the solutes in water. Record your results.
7. Discard the contents of the test tubes as directed by your teacher and clean as before.

Results

Substance	Dissolves in water?	Dissolves in paraffin oil?

Processing of results and questions

1. What generalisations can you make about the solubility of the solutes tested in:
   (a) water? ______________________________

   ____

   (b) paraffin oil? ______________________________

   ____

2. If two pieces of clothing were stained with chromium(III) chloride and lubricating grease respectively, what types of solvent would be needed to remove the stains?

   ____

   ____

   ____

   ____

Investigation 20: Chromatography – a bonding competition

Chromatography is a laboratory technique used to separate mixtures. A mixture dissolved in a solvent called the *mobile phase* is passed through a *stationary phase*, like paper or silica gel. Paper chromatography was used in Experiment 6. Chromatography works because the mobile phase, stationary phase and mixture/compound have different polarities. These differences in polarity mean differences in the attraction of mixture components to the stationary phase and in their solubility in the solvent. This results in their separation as they travel through the stationary phase at different rates.

In paper chromatography, the paper is made of cellulose, a polar substance. The more polar compounds within a mixture being separated will bond more readily with the polar cellulose paper. As a result the components become separated. The more polar the component the less it travels up the paper while the less polar components travel further up the paper.

Separation of compounds using chromatography is therefore based on a competition. The compounds in the sample and the mobile phase compete for bonding places on the stationary phase.

Task

Identify the most effective solvent, from a selection of solvents, for the separation of the coloured dyes from black ink using paper chromatography.

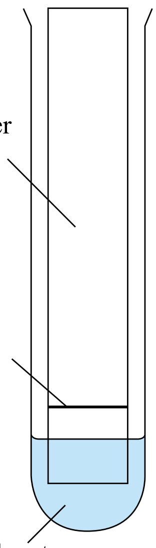


Figure 20.1

Equipment

- test tubes (five large)
- strips of chromatography or filter paper to fit into the large test tubes (five)
- household solvents - 5 mL samples:
  sodium chloride solution [NaCℓ] 1% solution
  cloudy ammonia solution
  methylated spirits
  tap water
  black ink: use ink from a black felt tipped pen or ballpoint pen

Planning the investigation

1. Plan an investigation using paper chromatography to determine the most effective solvent, (mobile phase), to use to separate the coloured dyes from black ink. In your plan include a table of the variables considered in the investigation.

	Variable/s	Unit/s	How the variable will be measured/manipulated
Independent variable			
Dependent variable			
Factors kept constant (controlled variables)			

2. Briefly outline your plan. A list of equipment and solvents and a sample setup has been provided but, with your teacher's permission, alternatives may be used. List any additional chemicals and equipment you need. Identify safety requirements.
   Check the proposed procedure with your teacher.

SAFETY NOTE:
- Check your plan with your teacher before you commence.
- At the end of the investigation dispose of the solvents appropriately – a method of solvent disposal must be included in your procedure.

Conducting the investigation

1. Conduct some preliminary trials. You may need to clarify how the independent and dependent variables will be measured.

2. Describe what you learned from these initial trials and list modifications that you made to your original plan.

3. Conduct the investigation, collecting and recording the data in a table as you proceed. If time allows replicate the data collection and average the results if appropriate.

Processing the results

Analyse your data and relate your conclusions to the original purpose of the investigation. Include in your discussion:
- The different coloured components in black ink and their relative polarities.
- Competition of the coloured components for binding positions on the polar cellulose paper, i.e. the stationary phase.
- Comparison of the solubilities of the different coloured components with each of the solvents.

Conclusion

Which is the most effective solvent for the separation of the coloured dyes from black ink using paper chromatography? Explain your choice.

Evaluating the investigation

1. Evaluate the effectiveness of your procedure and describe any modifications you would make to improve it. You may discuss factors that would improve the accuracy of your results such as sample size and selection, measurement errors and the control of variables. You may also address more general organisational factors such as the allocation of tasks among group members and the nature of the apparatus and how it was set up.

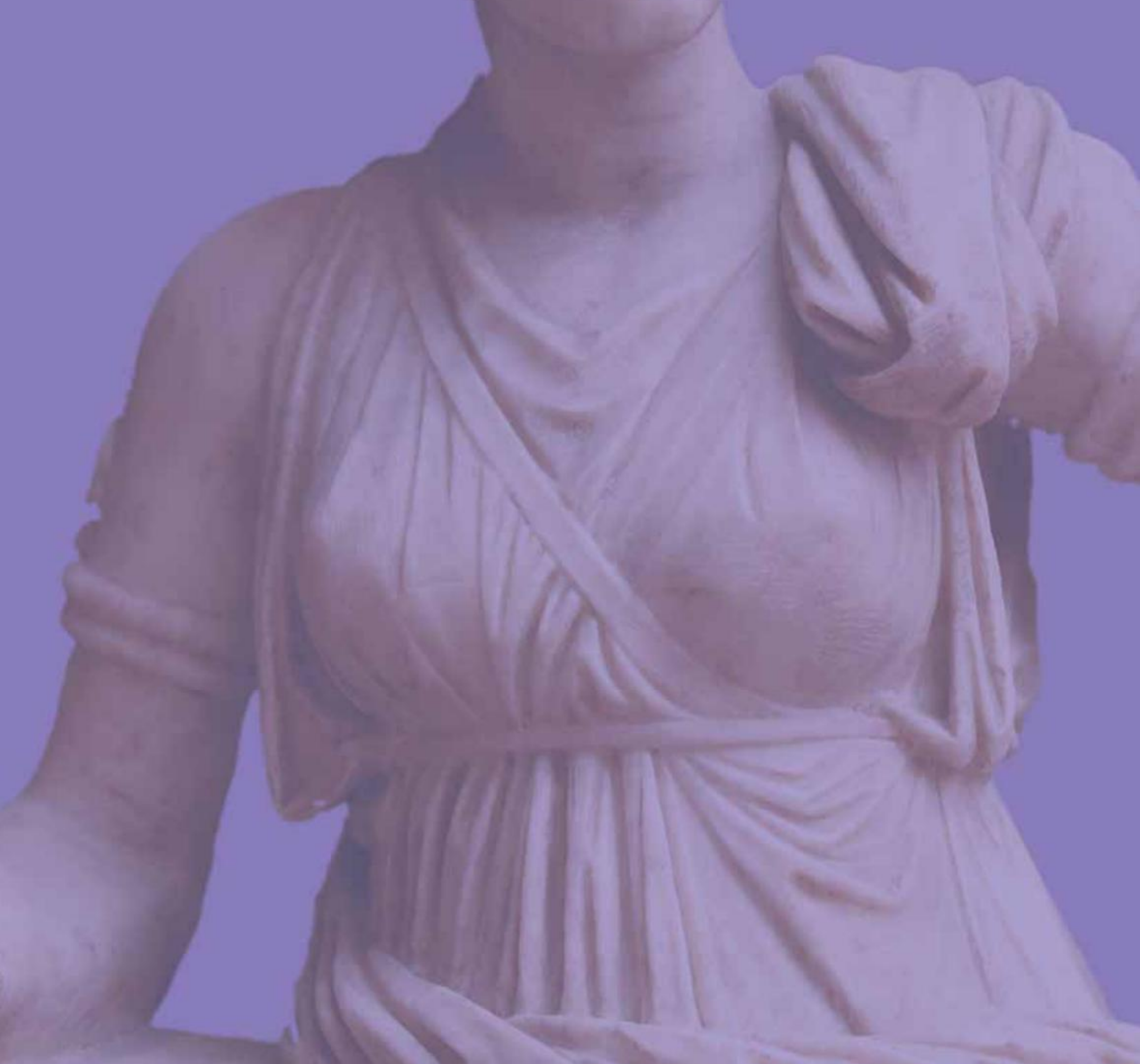
2. Discuss your confidence in the findings of the investigation.

3. Recommend procedures that can be carried out and investigations that need to be done to further improve the effectiveness of the solvent you have identified as the best to use for the separation of black ink.

Aqueous solutions and acidity

Design, conduct, evaluate and communicate investigations into the properties and behaviour of water, aqueous solutions and acids and bases.

Understanding the characteristic properties of water including the properties of aqueous solutions underpins an understanding of the physical, chemical and biological processes on Earth. They explain the solubility of substances in water, and the acidity of solutions, which can then be applied to the supply of drinking water and the identification of pollutants that cause corrosion of metal and stone structures.



Experiment 21: Solubility rules

In chemistry a precipitation reaction occurs when two salt solutions are added together to form an insoluble salt, known as the precipitate. An example of this is shown in the equation below: where silver nitrate solution is added to potassium chloride solution and a precipitate of silver chloride is formed.

$$AgNO_3\,(aq) + KC\ell\,(aq) \rightarrow AgC\ell\,(s) + KNO_3\,(aq)$$

$$Ag+\,(aq) \quad + C\ell^-\,(aq) \quad \rightarrow AgC\ell\,(s)$$

Transition metals are known to form precipitates with different colours depending on their elemental identity and their ion charge (oxidation state) within the precipitate. Chemists need to know the solubilities of ions to identify the presence of certain cations and anions present in a solution. It also allows them to remove specific ions from water and prepare pigments with specific colours.

Aim

In this experiment the solubilities of nitrates, chlorides, sulfates and carbonates will be examined by observing the results of mixing solutions of the sodium salts with solutions containing different cations.

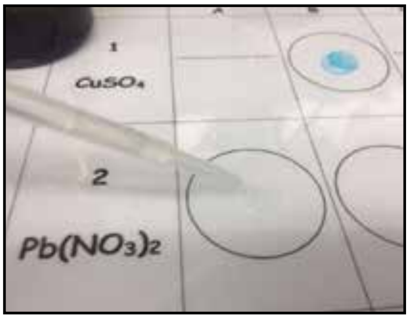
Equipment

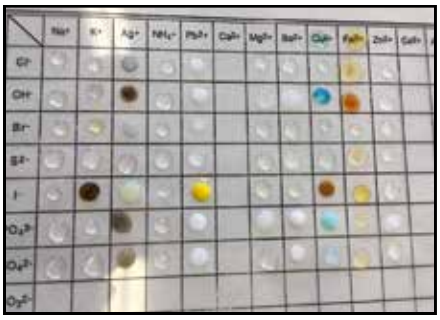
This experiment is designed to be done on a microscale rather than in test tubes to limit the amount of chemical waste.

Dropper bottles of 0.1 mol L^{-1} solutions of:

- sodium nitrate [$NaNO_3$]
- sodium chloride [$NaC\ell$]
- potassium nitrate [KNO_3]
- silver nitrate [$AgNO_3$]
- lead(II) nitrate [$Pb(NO_3)_2$]
- magnesium sulfate [$MgSO_4$]
- iron(II) sulfate [$FeSO_4$]
- copper(II) sulfate [$CuSO_4$]
- sodium sulfate [Na_2SO_4]
- sodium carbonate [Na_2CO_3]sodium iodide [NaI]
- sodium iodide [NaI]
- zinc sulfate [$ZnSO_4$]
- chromium(III) sulfate [$Cr_2(SO_4)_3$]
- aluminium sulfate [$A\ell_2(SO_4)_3$]
- barium chloride [$BaC\ell_2$]
- calcium chloride [$CaC\ell_2$]
- iron(III) chloride [$FeC\ell_3$]

Spotting tile (see photos for examples)





Procedure

1. Add one drop of $NaNO_3$ solution to 12 squares on the spotting tile so that all the cations (positive ions) can be tested.
2. To the first square add one drop of the KNO_3 solution and record your observations in the results table.
3. To the second square add one drop of the $AgNO_3$ solution and record your observations in the results table.
4. Continue to add one drop of each of the solutions listed so that each cation can be tested.
5. Once complete, wash the tile and start again, this time by adding one drop of $NaC\ell$ solution to 12 squares on the spotting tile.
6. Test each cation solution with the $NaC\ell$ solution by adding one drop of each one to a different square. Record your results in the table.
7. Once complete, wash the tile.
8. Complete steps 5-7 for Na_2SO_4 Na_2CO_3 NaI solutions so they are tested with each cation.

Results

Solution to be tested	Cation present	$NaNO_3$ anion added NO_3^-	NaCℓ anion added $C\ell^-$	Na_2SO_4 anion added SO_4^{2-}	Na_2CO_3 anion added CO_3^{2-}	NaI anion added I^-
KNO_3	K^+	no visible reaction				
$AgNO_3$	Ag^+	no visible reaction	white ppt.			
$Pb(NO_3)_2$	Pb^{2+}					
$MgSO_4$	Mg^{2+}					
$FeSO_4$	Fe^{2+}					

Processing of results

1. Examine your observations in column 3 of the results table. What generalisation can be made about the solubilities of nitrates?

2. Examine columns 4, 5, 6 and 7 of the table in the same way as you did for column 3. What generalizations can be made about the solubilities of the chlorides, sulfates, carbonates and iodides?

Questions

1. Write ionic equations for all cases in which a precipitate was formed.

2. Lead(II) ions must be removed from wastewater before it is discharged as it is toxic if allowed to build up. Based on your results from this experiment, discuss how you might remove the lead(II) ions from the water.

Experiment 22: pH of materials

Our environment contains a variety of solutions. The most obvious being oceans, lakes and rivers but also including soil moisture and ground water. Some of these may become contaminated by the deliberate or inadvertent addition of substances that are in common usage in industry, agriculture or in the home. One of the effects of this contamination is to change the pH of these solutions.

The pH of solutions can be measured using indicators and electronic pH meters. Indicators are materials that change colour in different pH solutions. Universal Indicator ™ is a commonly used indicator. It is a mixture of several indicators each exhibiting different colours in acids and bases. Its colour results from the combination of the colours of all the indicators present. Because it exhibits a range of colours over a large range of pH values, Universal Indicator ™ can be used to estimate the pH of solutions.

Aim

To test the pH, using Universal Indicator ™, of a variety of materials found around the home that are used and released into the environment.

Equipment

- test tubes (several)
- Universal Indicator ™ solution (5 mL)
- Universal Indicator ™ colour chart

5 mL solutions of:

- washing powder
- household ammonia
- baking soda
- sugar salt
- orange juice
- lemon juice
- garden or builder’s lime
- toothpaste
- lemonade
- vinegar
- milk
- aspirin
- antacid preparation
- soluble garden fertiliser
- drain or oven cleaner
- mouth wash

SAFETY NOTE:
- **Garden or builder’s lime and drain or oven cleaner should not be allowed to come into contact with your skin.**
- **Handle these substances with extreme care.**
- **Wear gloves and safety glasses.**

Procedure

Prediction:
In the results table write A (acid), B (base) or N (neutral) next to each test sample listed.

1. Preparation of samples.
   (i) Solutions can be tested directly.
   (ii) Solutions that have solids mixed with them should be decanted or filtered.
   (iii) Solids should be mixed with water, stirred thoroughly and then the solution should be treated as in (i) or (ii) as appropriate.
2. Place about 2 mL of each solution into separate test tubes.
3. Add two drops of Universal Indicator ™ and, using the colour chart provided with the indicator, note the pH of the solutions, and record in your results table.

Results

Material	Prediction	pH	Material	Prediction	pH

Material	Prediction	pH	Material	Prediction	pH

Processing of results

1. What percentage of your predictions were correct? ____________
2. In the box on the right, draw a scaled diagram from pH 0 to pH 14 that lists the materials tested against their pH, from most acidic to neutral to most basic.

Questions

1. (a) Name the component in each of the acidic solutions responsible for the solutions being acidic. You may have to research some of these.

   ____

   (b) Choose one of the acidic substances and write a balanced ionic equation to show why the solution is acidic.

   ____

2. (a) Name the component in each of the basic materials responsible for the solutions being basic. You may have to research some of these.

   ____

   ____

   (b) Choose one of the basic substances and write a balanced ionic equation to show why the solution is basic.

   ____

Further investigation

Investigate the pH of rainwater and/or groundwater in your area.

- Predict the pH you expect the water to have.
- Explain your predictions.
- Collect water samples.
- Measure and record the pH readings of the rainwater and/or groundwater.
- Explain any differences between predicted pH and measured pH values.

Experiment 23: Conductivity of acids and bases

Acids and bases are electrolytes, that is, they conduct an electric current when in aqueous solution. For acids and bases electrolyte strength is also referred to as acid or base strength. Strong acids and bases exist essentially as ions in aqueous solution. Weak acids and bases are those in which only a small proportion of the molecules or ions react with water to form hydronium ions, H_3O^+ (hydrogen ion attached to a water molecule) or hydroxide ions, OH^- in aqueous solution.

In this experiment you will investigate the electrical conductivity of acid and base solutions of different strengths and concentrations.

Aim

To identify the effect of electrolyte strength and concentration on conductivity.

Equipment

- power supply (0-12 V)
- electrical leads (four)
- beakers (two 100 mL)
- plate electrode system as shown in Figure 23.1
- switch
- ammeter
- 50 mL samples of the following solutions:
  - hydrochloric acid [$HC\ell$] 1 mol L^{-1}
  - hydrochloric acid [$HC\ell$] 0.1 mol L^{-1}
  - hydrochloric acid [$HC\ell$] 0.01 mol L^{-1}
  - hydrochloric acid [$HC\ell$] 0.001 mol L^{-1}
  - nitric acid [HNO_3] 1 mol L^{-1}
  - acetic acid [CH_3COOH] 1 mol L^{-1}
  - sodium hydroxide [$NaOH$] 1 mol L^{-1}
  - ammonia [NH_3] 1 mol L^{-1}

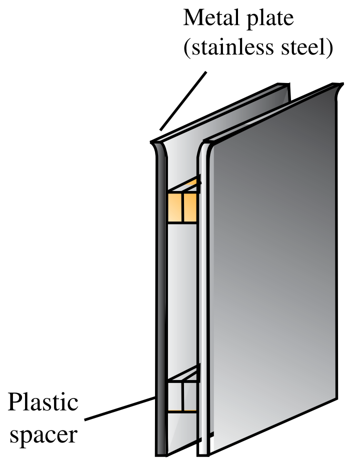


Figure 23.1. Plate Electrode System

Procedure

Please note this experiment should be set up as 8 stations. Students can set up one station then move around the room and take readings at each station.

1. Set up the circuit and the electrodes as shown in Figures 23.2 and 23.3. Set the power supply to 6 V DC.
2. Place 50 mL of 1 mol L^{-1} $HC\ell$ (or one of the other designated solutions) in a 100 mL beaker.
3. Put the electrodes into the solution and briefly close the switch.
4. Record the reading on the ammeter.
5. Move to the next station and repeat steps 3-4 for the other samples: 0.1 mol L^{-1} $HC\ell$, 0.01 mol L^{-1} $HC\ell$, 0.001 mol L^{-1} $HC\ell$ and 1.0 mol L^{-1} solutions of HNO_3, $NaOH$, CH_3COOH and ammonia solution, (NH_3 (aq)).

Figure 23.2

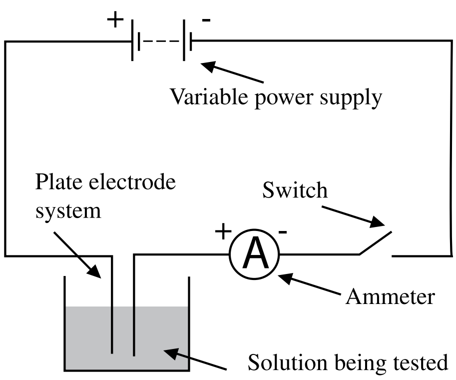


Figure 23.3: Circuit diagram

SAFETY NOTE:
- **Connect the red terminal of the power supply to the red terminal of the ammeter that uses the least sensitive scale. If the reading is less than the full scale reading of the next more sensitive scale switch the = red terminal to use that more sensitive scale.**

Observations and results

Substance	Conductivity(A)
hydrochloric acid 1 mol L^{-1}	
hydrochloric acid 0.1 mol L^{-1}	
hydrochloric acid 0.01 mol L^{-1}	
hydrochloric acid 0.001 mol L^{-1}	
nitric acid 1 mol L^{-1}	
acetic acid 1 mol L^{-1}	
sodium hydroxide 1 mol L^{-1}	
ammonia 1 mol L^{-1}	

Processing of results

1. If you had to classify these substances into 3 groups, what rule would you use? Compare your rule with others used by the class.

2. Look at your results for the HCℓ solutions. What conclusion can you draw concerning the conductivity of a solution and the concentration of ions in it?

3. Were the conductivities of 1.0 mol L^{-1} HCℓ and 1.0 mol L^{-1} CH_3COOH different? If so, explain why they were different.

4. Were the conductivities of 1.0 mol L^{-1} NaOH and 1.0 mol L^{-1} NH_3 different? Explain why they were different.

5. Acids and bases are referred to as being ‘strong’ or ‘weak’ as distinct from being ‘concentrated’ or dilute’. Explain how these two variables, electrolyte strength and solution concentration, affect the conductivity of aqueous solutions.

6. Classify the following solutions correctly from the choices in the table headings.

Substance	acid or base	strong or weak	concentrated or dilute
8 mol L^{-1} $HC\ell$			
0.005 mol L^{-1} Na_2CO_3			
glacial CH_3COOH			
6 mol L^{-1} NaOH			
0.0001 mol L^{-1} HNO_3			

Experiment 24: Acid reactions

Understanding the properties of acids and bases is important as they have wide ranging applications in both the context of chemistry and in our daily lives. Commonly encountered acids include sulfuric acid contained in car batteries, hydrochloric acid in the stomach and acetic acid, also known as ethanoic acid, found in vinegar. Commonly used bases include sodium hydrogen carbonate, also known as sodium bicarbonate, found in baking soda, sodium hydroxide found in caustic soda and ammonia used in cleaning products. Acids react to neutralise bases and react with many metals and metal compounds.

Aim

In this experiment you will investigate the reactions of acids with bases, metals, metal oxides, hydroxides, carbonates and hydrogen carbonates.

SAFETY NOTE:
Acids and bases are corrosive to a large variety of materials including your skin, clothing and books. They cause skin burns and eye damage. Avoid contact with your skin, but if skin or eye contact does occur, immediately irrigate the affected area with copious quantities of water for a minimum of 15 minutes.
You must wear safety glasses for duration of the experiments.

Equipment

- Bunsen burner
- teat pipettes (2)
- test tubes (7)
- single hole stopper to fit test tube
- glass or plastic delivery tubing
- matches
- taper
- hydrochloric acid, $HC\ell$, 2 mol L^{-1} (50 mL)
- vinegar (10 mL)
- saturated limewater solution, $Ca(OH)_2$ (50 mL)
- two small pieces each of aluminium, $A\ell$; copper, Cu; iron, Fe; magnesium, Mg and zinc, Zn

about 1-2 g each of the metal oxides and hydroxide:

- copper(II) oxide [CuO]
- iron(III) oxide [Fe_2O_3]
- aluminum oxide [Al_2O_3]
- calcium hydroxide [$Ca(OH)_2$]

about 1-2 g each of the metal carbonates and hydrogen carbonate:

- calcium carbonate [$CaCO_3$]
- sodium carbonate [Na_2CO_3]
- sodium hydrogen carbonate [$NaHCO_3$]

- 10 mL of 2 mol L^{-1} sodium hydroxide solution [$NaOH$]
- blue litmus paper
- red litmus paper
- litmus solution in dropper bottle

Procedure

1. Carry out the following reactions carefully and be sure to follow all safety precautions.
2. Solutions of hydrochloric acid and sodium hydroxide are clear and colourless. Containers used to store these solutions must be clearly labelled.
   You can check whether a solution is an acid or a base by using an indicator such as litmus indicator paper. To establish the test, do the following and record your observations.
   (a) Place a drop of 2 mol L^{-1} $HC\ell$ onto pieces of red and blue litmus paper.
   (b) Repeat using 2 mol L^{-1} NaOH.

A Neutralisation of a strong acid with a strong base

1. Place 10 drops of 2 mol L^{-1} HCℓ into a test tube and add three drops of litmus solution or small pieces of red and blue litmus paper. Note the final colour of the litmus.
2. Add 2 mol L^{-1} NaOH dropwise until there is a permanent colour change. Record all observations.

Observations and results

Reactants (name and description)	Observations of any reaction	Products (name and description)

B Reaction of acids with metal oxides and hydroxides

1. Place about a level spatula, of CuO in a test tube. Add around 2 cm of 2 mol L^{-1} HCℓ, and gently heat the mixture, without boiling and allow to stand. Record your observations.
2. Repeat step 1, replacing CuO with Fe_2O_3, Al_2O_3 and then with $Ca(OH)_2$.

Observations and results

Metal oxide and hydroxide reactants (name and description)	Observations of reaction with HCℓ	Products (name and description)

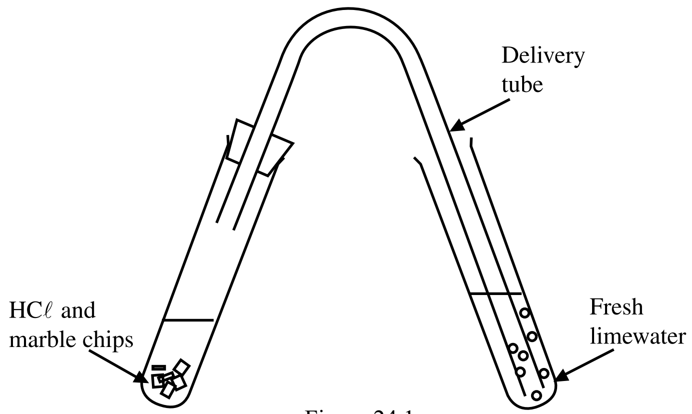


Figure 24.1

C Reaction of acids with metal carbonates and hydrogen carbonates

1. Place about a level spatula of $CaCO_3$, marble chips, in a test tube.
   Add around 2 cm of 2 mol L^{-1} $HC\ell$. Set up the test tube and delivery tubing as shown in figure 24.1, with around 2 cm of limewater in the collection test tube.
2. Record your observations.
3. Repeat steps 1 and 2, replacing $CaCO_3$ with Na_2CO_3 and then with $NaHCO_3$.
4. Repeat steps 1 and 2 with $CaCO_3$ using vinegar in place of the $HC\ell$ solution.

Observations and results

Metal carbonate and hydrogen carbonate reactants (name and description)	Observations of reaction with $HC\ell$	Products (name and description)

Limewater test		
Reactants	Observations of reaction of product gas with limewater	Products (name and description)

D Reaction of acids with metals.

Dispose of the solid metals by tipping the contents of the test tubes containing unreacted metals into the container provided by your teacher. DO NOT dispose into the sink.

SAFETY NOTE:
Acids will react with a large variety of materials including your skin, clothing and books. Take care not to spill the acid, particularly onto your skin. If spills do occur wash the affected area with large quantities of tap water for 20 minutes. You must wear safety glasses at all times during this experiment.

1. Carry out the following reactions carefully. All members of a group should make observations of each reaction, then write them into their own results table.

2. Using the steel wool or emery paper thoroughly clean the surface of the metals.

3. Place the small pieces of Mg, Zn, Aℓ, Fe and Cu into separate test tubes. Add 2-3 mL of 2 mol L^{-1} HCℓ to each. (Note: not all the metals will react)

4. If a gas is evolved, collect it by inverting another test tube over the mouth of the reaction tube. Keep the collection tube inverted and hold a lighted taper near the open end as shown in figure 24.1.

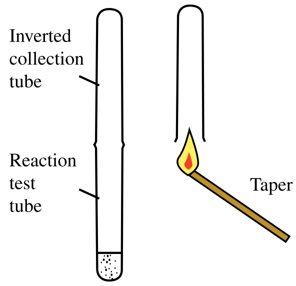


figure 24.1

Observations and results

Reactant metals (name and description)	Observations of reaction with HCℓ	Products (name and description)

Pop test		
Reactants	Observations of reaction of product gas	Products (name and description)

Processing of results and questions

1. Compare and contrast the reaction observations of vinegar and hydrochloric acid with calcium carbonate.

2. Write an ionic equation for each of the reactions you observed, including any reaction that occurs with the gas collected.

3. (a) Which one of the metals appears to be the most reactive with acids?

(b) Explain your decision based on the observations you made.

(c) Which one of the metals appears to be the least reactive with acids?

(d) Explain your decision based on the observations you made.

4. List examples of where acids have a detrimental effect on metal and stone structures or containers, identify the type of acid reaction involved and suggest how this effect could be minimised.

SECTION 2
Chemical understanding and problem solving

Unit 1

Chemical fundamentals: structure, properties and reactions

Through the investigation of appropriate contexts, students explore how evidence from multiple disciplines and individuals have contributed to developing understanding of atomic structure and chemical bonding. They explore how scientific knowledge is used to offer reliable explanations and predictions, and the ways in which it interacts with social, economic and ethical factors.

Students use science inquiry skills to develop their understanding of patterns in the properties and composition of materials. They investigate the structure of materials by describing physical and chemical properties at the macroscopic scale, and use models of structure and primary bonding at the atomic and sub-atomic scale to explain these properties. They are introduced to the mole concept as a means of quantifying matter in chemical reactions.

Commonly encountered quantities and units in chemistry

Quantity	SI Unit	SI Symbol	Common unit	Common symbol	Comments
Amount of substance	mole	mol			
Concentration	mole per cubic metre	mol m^{-3}	mole per litre	mol L^{-1}	Common unit is more convenient
Potential difference	volt	V			
Energy	joule	J	kilojoule	kJ	
Mass	kilogram	kg	gram	g	Common unit is more convenient
Molar mass	kilogram per mole	kg mol^{-1}	gram per mole	g mol^{-1}	Common unit is more convenient
Molar volume	cubic metre per mole	m^3 mol^{-1}	litre per mole	L mol^{-1}	Common unit is more convenient
Pressure	pascal	Pa	kilopascal	kPa	Pressures, in examination questions are stated in kilopascals (kPa)
Relative atomic mass (atomic weight)	These are all dimensionless quantities				
Relative molecular mass (molecular weight)					
Relative formla mass (formula weight)					
Temperature	kelvin	K	degree celsius °C		Both sets of units are useful K = °C + 273.15
Time	second	s			
Volume	cubic metre	m^3	cubic decimetre litre millimetre	dm^3 L mL	Common units are convenient 1L = $1dm^3$ = 1000 mL = 1000 cm^3 = 1×10^{-3} m^3

Set 1: Scientific notation and unit conversions

Every measurement consists of two essential parts:

- a number, and
- a unit

Number

Byron measured 1.8 m tall with a mass of 67 kg.

Unit

Base units relevant to the following problem solving sets include:

Quantity	SI base unit	Symbol
Length	metre	m
Mass	kilogram	kg
Time	second	s
Absolute temperature	kelvin	K
Amount of substance	mole	mol

Derived units are those defined by various operations with units, such as

- multiplication,
- division,
- conversion, and
- raising to any power.

Examples of derived units are:

- dm^3 or L for volume
- $g\ cm^{-3}$ or $g\ mL^{-1}$ for density
- $mol\ L^{-1}$ for concentration
- $g\ mol^{-1}$ for molar mass

Metric prefixes
Decimal multiples and decimal fractions of SI units are represented by standard prefixes, and each prefix has the standard symbol as shown in the table.

Power of 10	Prefix	Symbol
10^{12}	tera	T
10^{9}	giga	G
10^{6}	mega	M
10^{3}	kilo	k
10^{2}	hecto	h
10^{1}	deca	da
10^{-1}	deci	d
10^{-2}	centi	c
10^{-3}	milli	m
10^{-6}	micro	μ
10^{-9}	nano	n
10^{-12}	pico	p

Converting between units:
To convert a mass of 9.213 kg to a mass in g we can replace the kg unit with $\times 10^3$ g.

Since: 1 kg = 1×10^3 g
then: 9.213 kg = 9.213×10^3 g
= 9 213 g

Exponential notation

This notation, also known as scientific notation, is used for convenience when writing very large or very small numbers.
To express a number using exponential notation, it is written as a number between 1 and 10 multiplied by the appropriate power of 10.
For example 0.056 expressed in exponential notation is written as 5.6×10^{-2} while 167 000 would be written as 1.67×10^5.

Arithmetic using exponential notation is governed by the following rules:

Addition and subtraction Before adding or subtracting, numbers must be expressed to the same powers of 10.	**Multiplication** When multiplying powers of 10 add their indices algebraically.	**Division** When dividing powers of 10 by each other, subtract the index of the denominator from that of the numerator.
Example: $(2.04 \times 10^5) + (4.7 \times 10^4)$ $= (2.04 \times 10^5) + (0.47 \times 10^5)$ $= (2.04 + 0.47) \times 10^5$ $= 2.51 \times 10^5$	Example: $(5 \times 10^5) \times (4 \times 10^2)$ $= (5 \times 4) \times 10^{5+2}$ $= 20 \times 10^7$ $= 2 \times 10^8$	Examples: $\frac{1.8 \times 10^8}{6 \times 10^5} = \frac{1.8 \times 10^{8-5}}{6}$ $= 0.3 \times 10^3$ $= 3 \times 10^2$

Set 1: Exercises

1. Express the following numbers using exponential notation:

(a) 329 ____________

(b) 1006 ____________

(c) 0.5731 ____________

(d) 67 240 000 000 ____________

(e) 0.04 ____________

(f) 0.000 000 078 ____________

2. Convert the following measurements into the required units:

(a) 2.643 kg to g ____________

(b) 0.012 kg to g ____________

(c) 2.50 tonne to g ____________

(d) 439 mg to g ____________

(e) 0.2846 g to mg ____________

(f) 6.72×10^8 mg to kg ____________

3. Complete the following:

(a) 10.2 m = ____________ cm

(b) 1.26 cm = ____________ m

(c) 1.46 m = ____________ mm

(d) 143 267 mm = ____________ m

(e) 109 nm = ____________ m

(f) 141 mm^2 = ____________ m^2

(g) 8.3 mm^3 = ____________ m^3

(h) 1220 cm = ____________ m

(i) 15 mm = ____________ cm

(j) 10 dm = ____________ m

(k) 1.9 µm = ____________ cm

(l) 15 cm^2 = ____________ mm^2

(m) 4.9 cm^3 = ____________ m^3

(n) 1.67 mL = ____________ L

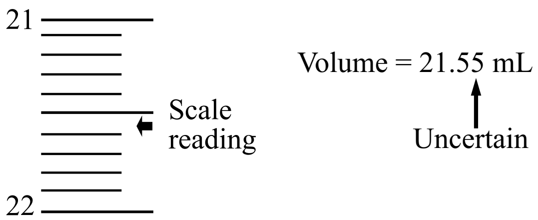
Set 2: Significant figures

Uncertainty and significant figures in measurement

When handling experimental data it is important to consider the appropriate number of significant figures to use and not be tempted to record values with large numbers of figures that may be produced by your calculator when processing data. Significant figures in a number are those digits that are known with certainty plus the first digit that is uncertain. All measurements have some level of uncertainty. These arise from the quality of the equipment used, the type of scale and the skill of the experimenter. Some apparatus will have a manufacturers' uncertainty assigned; however, there are cases where you will make a judgement.

Burettes are used in titration experiments to accurately measure the volume of a reactant added to another. A typical 50 mL burette has an analogue scale with a smallest scale division of 0.1 mL. A manufacturer's uncertainty of +/- 0.05 mL (representing half the smallest scale division) is often assigned. Depending on your skill level you may judge that you can read the scale on the burette to the nearest 0.01 mL and so reduce the range of uncertainty for each measurement to + /- 0.02 mL.

The burette reading in the diagram of 21.55 mL is estimated to the nearest 0.05 mL. The true value lies in a range from 21.50-21.60 mL.



Significant figures in a number are those digits that are known with certainty plus the first digit that is uncertain. In the reading of 21.55 mL the 2, 1 and 5 are known with certainty and the second decimal place 5 is uncertain. This value has 4 significant figures.

	Number	Number of significant figures
1.	713	3
2.	7.03×10^3	3
3.	11.05	4
4.	0.027	2
5.	9.9643×10^{-7}	5

Whole numbers ending with one or more zeros such as 430 and 500 are ambiguous with respect to the number of significant figures because it is unclear whether the terminating zeros are significant or merely serve to locate the decimal point.

The use of scientific notation avoids ambiguity as shown by the following examples:

1. If 430 has 3 significant figures it is written as 4.30×10^2

2. If 430 has 2 significant figures it is written as 4.3×10^2

3. If 500 has 1 significant figure it is written as 5×10^2

Set 2: Exercises

1. How many significant figures are given in each of the following

(a) 454 ________ (c) 100.10 ________ (e) 6000 ________ (g) 14.0 ________

(b) 0.3750 ________ (d) 6.07×10^{-11} ________ (f) 0.0003 ________ (h) 1.103 ________

2. Round off the following measurements to three significant figures:

(a) 7.248 ________ (b) 0.017428 ________ (c) 6.275×10^{-3} ________ (d) 1.1252×10^{8} ________

3. Evaluate the following and express the answer to the correct number of significant figures:

(a) $2.65 + 2.5 + 3.61$

____

(b) $727 + 18.42 - 85.496$

____

(c) $(2.464 \times 10^{-3}) - (8.7643 \times 10^{-4})$

____

(d) $8734.2 + (2.81 \times 10^{4}) - (3.432 \times 10^{3})$

4. Evaluate the following and quote your answer in exponential notation to the correct number of significant figures.

(a) 7.325×9.27 ________

(b) $(7.81 \times 10^{4}) \times 0.031 \times 22.4$ ________

(c) $\dfrac{8.426}{2.98}$ ________

(d) $\dfrac{3.564 \times 10^{6}}{7.2538 \times 10^{-2}}$ ________

(e) $\dfrac{(5.68 \times 10^{2}) + (2.10 \times 10^{-1})}{2.1 \times 10^{-1}}$ ________

(f) $\dfrac{(8.23 \times 10^{2}) - (7.94 \times 10^{2})}{8.262 \times 10^{-2}}$ ________

5. Calculate the volume of a rectangular block with the dimensions 3.0 cm by 2.0 cm by 1.55 cm.

____

____

____

____

Set 3: Random and systematic errors

All measurements have a degree of uncertainty resulting in experimental error. The experimental error in a result is the difference between the experimental value and the literature or theoretical value. There are two types of experimental error: random error and systematic error. Both should be considered when evaluating any quantitative investigation.

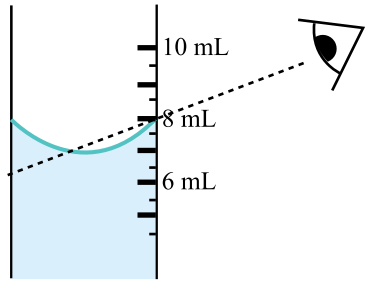
Random errors come from measurements that have an equal chance of being above or below the actual value.

Systematic errors are a result of flaws in the experimental method or apparatus that lead to a result that is always either above or below the true value.

Errors are discussed in detail on pages 12 and 13.

Set 3: Exercises

1. A student measured 8.0 mL of water in a measuring cylinder as shown.



   a) What would her reading be?

   b) Describe the error/s in the student's technique.

   c) Does this error in technique result in random or systematic errors in the volume measurements?

   d) What is the impact of this error in technique on the actual volumes measured?

   e) What is the uncertainty (measurable random error) associated with reading from this measuring cylinder?

2. Digital pH meters are calibrated with solutions of known pH.

   a) What kind of error occurs if a solution thought to be pH 7.0 but which is actually pH 7.2 is used in the calibration of the meter?

   b) Explain your answer.

3. Describe a systematic error that could occur when:

   a) measuring the volume of a liquid using a measuring cylinder.

   b) measuring the mass of reactants on a top pan balance.

4. Describe how random errors result when:

   a) Estimating pH using a colour chart following the addition of Universal Indicator ™ to samples of acids and bases.

   b) Measuring temperature using a thermometer.

   c) Judging flame colour when metal salts are heated in a Bunsen burner flame.

   d) Calculating the R_f values (distance the sample migrates/ total distance moved by the solvent.) of samples separated by paper chromatography.

5. Describe how it is possible to obtain a set of experimental results that are very similar in value but that are not accurate.

6. Describe how the random uncertainty in measuring the mass of reactants can be minimized.

7. A student is investigating the rate of a reaction using a digital stopwatch. The results were recorded to one hundredth of a second. Explain why the results are not accurate to one hundredth of a second.

8. Repeated chemical analysis of a sample of hematite-rich iron ore indicated the percentage by mass of iron present was 56.35%, 56.35%, 56.37%, 56.34% respectively. More accurate analysis then revealed the actual amount of iron present was 57.702%.

   a) Comment on the precision of the original method and the accuracy of results.

   b) Comment on the impact of random and systematic error on the results.

   c) How could the accuracy of the results be improved?

9. The energy contained in food can be estimated by burning food and using the energy released to heat a known mass of water. This experiment can be done with cracker biscuits. The sample is ignited using a Bunsen burner and then held below a test tube of water. The temperature rise in the water can be used to estimate the energy in the food.

    a) Identify variables that should be controlled in this experiment.

    b) Explain why the energy content discovered by this technique is always lower that the actual value.

10. Neutralisation reactions are exothermic. The enthalpy change in such reactions can be calculated by finding the heat released into solution. An experiment was carried out adding 100 mL of 1 mol L^{-1} sodium hydroxide added to 100 mL of 1 mol L^{-1} hydrochloric acid in a polystyrene coffee cup. The temperature change in the solution is measured using a thermometer and then used along with the mass of solution and specific heat capacity of water to calculate the enthalpy change (ΔH).

    The experimental value calculated in this experiment was only 50 % of the theoretical value.

    a) Identify the major source of systematic error in this experiment.

    b) Identify sources of random error.

    c) Suggest improvements that would increase the reliability (validity) of the results.

    d) Suggest improvements that would improve the accuracy of the result.

11. Two experiments were carried out to determine the formula of hydrated copper(II) sulfate. In the first experiment (Experiment 1) copper sulfate was added to a crucible and heated for two minutes. Recognising a systematic error in experiment 1 the student carried out a second experiment (Experiment 2) heating an identical sample repeatedly until a constant mass was obtained.
The measurements made are shown in the table below:

		Experiment 1	Experiment 2
	Time of heating (minutes)	Mass (g)	Mass (g)
crucible		12.20	12.20
hydrated copper sulfate and crucible		13.20	13.20
hydrated copper sulfate		1.00	1.00
anhydrous copper sulfate and crucible	2	12.95	12.94
	3		12.89
	4		12.84
	5		12.84

a) State a source of measurable random error in this experiment.

b) Explain why this measurement is considered to have random uncertainty.

c) Assuming there is no uncertainty in taring (zeroing) the balance, explain why the overall uncertainty of measurements used to obtain the mass of water in Experiment 1 is +/- 0.02 g

d) Describe the systematic error present in the first experimental design.

e) Describe the impact of the error on the number of moles of water (x) in the hydrated formula: $CuSO_4 \cdot xH_2O$.

f) Explain how the method in Experiment 2 overcame this systematic error.

g) The student carrying out the second method weighed the sample 4 times. Do the additional measurements increase the uncertainty of the final value they obtain? Explain your answer.

12. The energy released when different fuels undergo combustion can be estimated by burning a known mass of the fuel and using the heat released to increase the temperature of water. The temperature change along with other data (shown below) can be used to calculate the molar heat of combustion (ΔH).
The results obtained for four different alcohols are shown in the table.

Fuel	Formula of alcohol	Average mass of fuel burned (g)	Average change in temperature of 200 mL of water (K)	Literature value for the molar heat of combustion (kJ mol^{-1})	Experimental value for the molar heat of combustion (kJ mol^{-1})
methanol	CH_3OH	2.40	42	726	385
ethanol	CH_3CH_2OH	2.22	45	1367	
propan-1-ol	$CH_3CH_2CH_2OH$	2.64	50	2021	

The heat released by each alcohol can be calculated as follows:

$q = mc\Delta T$
q= the heat released in joules
m= mass of water heated in g.
The density of water is 1 g mL^{-1} therefore the 200.0 mL of water heated is converted to 200.0 g for the calculation.
c is the specific heat capacity of water = 4.18 J K^{-1} g^{-1}

Example calculation for methanol
$q = 200.0 \times 4.18 \times 42 = 35{,}112$ J

The molar heat of combustion for each fuel can be calculated using the formula $\Delta H = q/n$

Example calculation for methanol:
$n(CH_3OH) = m/M = 0.0911$
$\Delta H = 35{,}112/.0911 = 385422$ J mol^{-1} = 390 kJ mol^{-1}
(2 significant figures)

a) Write balanced chemical equations for the complete combustion of each alcohol to water and carbon dioxide.

b) Calculate the heat released for the other fuels.

c) Calculate the experimental value for the molar heat of combustion for the other fuels.

d) Describe the trend in the literature values of molar heat of combustion shown in question 12.

e) The experimental values are all lower than the literature values. What major source of systematic error accounts for this difference?

f) Calculate the percentage difference between the experimental values obtained in this experiment and the literature values for each of the fuels.

g) Unlike the other two fuels it was observed that the propan-1-ol burned with a sooty flame and left a deposit of black powder on the underside of the flask. Use this evidence to account for the result for propan-1-ol in part f.

h) Suggest improvements to the experimental design that could improve the accuracy of the results obtained.

Properties and structures of atoms

- Represent elements by symbols.
- Describe a model of an atom as a central nucleus (protons and neutrons), surrounded by electrons in distinct energy levels, held together by electrostatic forces of attraction between the positive nucleus and negative electrons.
- Represent the location of electrons within atoms using electron configurations.
- Explain chemical bonds by the arrangement of electrons in the atom and in particular by the stability of the valence electron shell.
- Understand the structure of the periodic table is based on the atomic number and the properties of the elements.
- Describe trends across periods and down main groups, including atomic radii, valencies, 1st ionisation energy and electronegativity as exemplified by groups 1, 2, 13–18 and period 3 elements of the periodic table.
- Describe the analytical techniques of flame tests and atomic absorption spectroscopy (AAS) used to identify elements and rely on electron transfer between atomic energy levels shown by line spectra.
- Define elements, ions and isotopes of atoms and represent them in the form (IUPAC) or X-A for isotopes.
- isotopes of an element have the same electron configuration and possess similar chemical properties but have different physical properties.
- Calculate relative atomic masses of the elements.
- Analyse and interpret mass spectra to determine the isotopic composition of elements and to determine relative atomic mass.

Set 4: Elements and symbols

Many element names have ancient origins as they have a long history. Elements identified in the modern era are often named by the scientists who have isolated or created them for the first time. Any proposed names must be ratified by the International Union of Pure and Applied Chemistry (IUPAC) before they can appear on the Periodic Table.

Element symbols are a short hand form of the element name. All symbols consist of up to 3 letters, where the first letter is always a capital and the remaining letters are lowercase.

Some symbols represent the Latin name of an element rather than the name in current use. Sodium, for example, has the symbol Na, which is derived from natrium, its Latin name.

Set 4: Exercises

1. Write the symbols of the following elements

(a) fluorine
(b) calcium
(c) manganese
(d) tungsten
(e) silver
(f) uranium
(g) platinum
(h) iodine
(i) neon
(j) bromine
(k) strontium
(l) astatine
(m) chromium
(n) copper
(o) barium

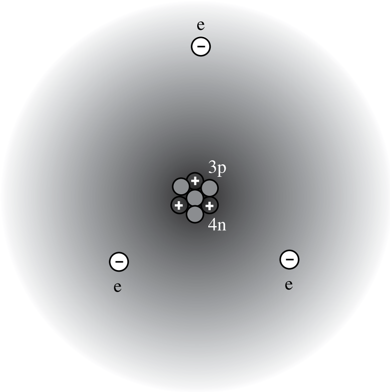
2. Write the names for the following symbols

(a) Na
(b) Rb
(c) Cℓ
(d) Au
(e) Li
(f) Pb
(g) Ni
(h) Aℓ
(i) Fe
(j) Co
(k) P
(l) Kr
(m) Zn
(n) K
(o) Sn

Set 5: Atoms and isotopes

Set 5: Exercises

1. a) Label the parts of the atom shown below. Note: This diagram is not to scale, the nucleus is hundreds of times smaller than shown.



b) Complete the following using the atom shown.

Number of protons __________

Number of neutrons __________

Number of electrons __________

c) Use a Periodic Table to identify the element. __________

2. Complete the table below.:

Particle	Charge	Mass relative to a proton
proton		
neutron		
electron		

3. If an electron were removed from an atom:
   a) How would it affect the overall mass of the atom? __________

   b) How would it affect the overall charge of the atom? __________

4. Complete the sentence:
Isotopes of an element are atoms of the same element and hence have the same number of __________ and __________, but have different numbers of __________.

5. Isotopes are regularly represented by the following symbol : ${}^{A}_{Z}E$ eg. ${}^{63}_{29}Cu$

For the element copper, identify:

E (element symbol): __________

A (mass number): __________

Z (atomic number): __________

6. Complete the following table:

Symbol	Element	Mass No (A)	Atomic No (Z)	No of neutrons
${}^{14}_{6}C$		14	6	
	chlorine		17	18
${}^{56}_{26}Fe$		56	26	
		31	15	
	silver	108		61
	carbon	12		
	sodium	23	11	
${}^{64}_{29}Cu$		64	29	
	calcium	40		20
	carbon			7

7. Which element in the table in question 6 is represented by more than one isotope?

____

8. Use the format: ${}^{A}_{Z}E$ to represent the following isotopes of hydrogen.

(a) hydrogen-1

____

(b) hydrogen-2

____

(c) hydrogen-3

____

Set 6: Atomic structure and the periodic table

An electron configuration describes the arrangement of electrons surrounding the nucleus of an atom. Electrons occupy regions of space around the nucleus which we refer to as 'shells'. Electrons in different shells have different energies. The energy of the electrons depends upon their distance from the nucleus and the shell they occupy. The electron configuration indicates how many electrons are in each energy level or shell.

Electrons, in their ground state, occupy the lowest energy levels or shells possible.

Examples

1. Oxygen: atomic number (Z) = 8, there will also be eight electrons in a neutral atom of oxygen, two in the first energy level and the remaining six in the second energy level. The electron configuration of oxygen would be written as: 2, 6
2. Oxide ion: O^{2-}, Z = 8, there will be eight electrons plus two extra electrons as it has a negative two charge. Therefore there will be two electrons in the first energy level and the remaining eight will be in the second energy level. The electron configuration of the oxide ion would be written as: 2, 8

Set 6: Exercises

1. Complete a table similar to the following to show details of the first 20 elements. As an example the element nitrogen has been done for you.

Z	Name	Symbol	Metal/ non-metal	Electron configuration	Valence electron behaviour
1					
2					
3					
4					
5					
6					
7	Nitrogen	N	Non-metal	2, 5	Shares or gains 3 e^-
8					
9					
10					
11					
12					
13					
14					
15					
16					
17					
18					
19					
20					

2. (a) Which elements appear to always lose electrons?

(b) Describe the position of these elements on the Periodic Table.

3. (a) Which elements appear to always gain electrons?

(b) Describe the position of these elements on the Periodic Table.

4. (a) Which elements do not gain or lose electrons?

(b) Describe the position of these elements on the Periodic Table.

5. Is there a pattern between the number of valence electrons and whether or not electrons are gained or lost?

6. Complete the following table:

	Symbol	Name	Atomic number	Number of protons	Number of electrons	Number of neutrons
Example	$^{19}_{9}F$	Fluorine	9	9	9	10
A	$^{16}_{8}X$					
B	$^{17}_{8}X$					
C	$^{16}_{8}X$					
D	$^{17}_{9}X$					

7. In the table in question 6, which species are isotopes of each other? Explain your answer.

8. There is a saying that protons give an atom its identity and electrons give an atom its personality. Explain what this means.

9. Write the electron configurations for the following species:
(a) S^{2-}

(b) $A\ell^{3+}$

(c) K

(d) C

10. Which elements are represented by the following electron configurations?
(a) 2, 8, 7

(b) 2, 8, 8, 2

(c) 2, 8, 3

11. Which of the following electron configurations represent elements in which electrons are not in the ground state? Explain.
(a) 2, 7, 8, 1

(b) 1, 8, 8

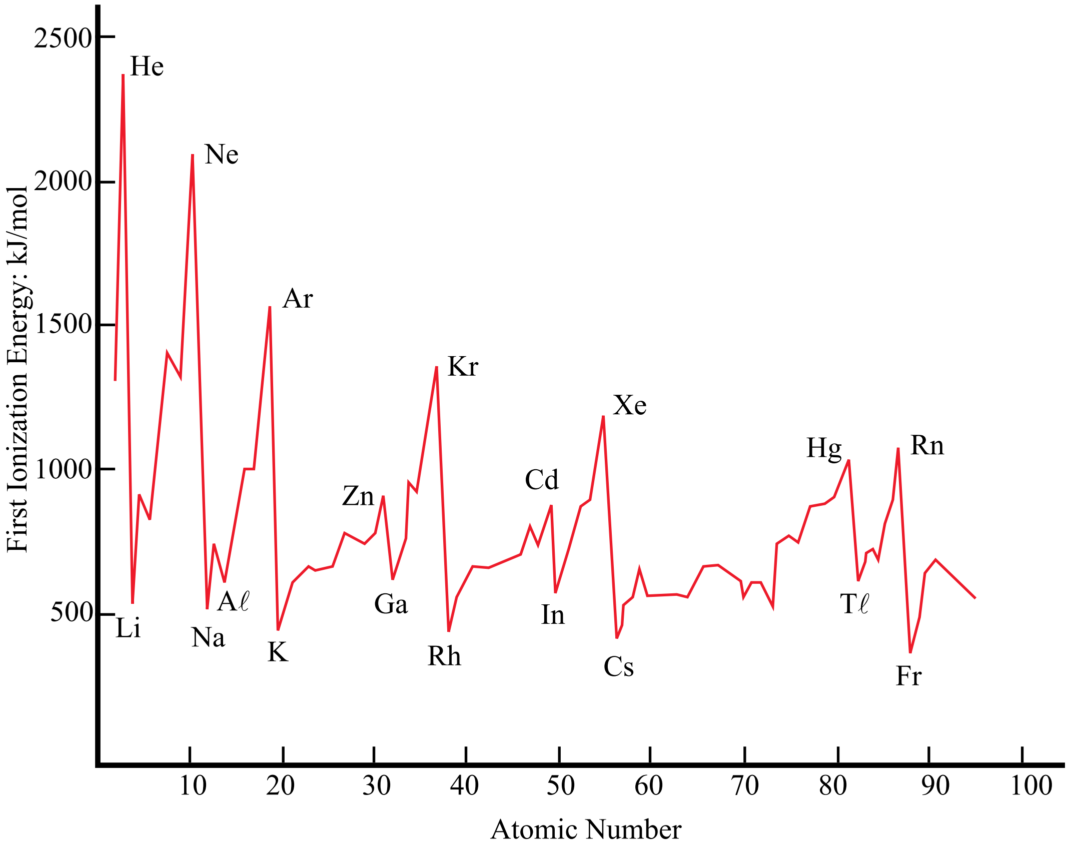
(c) 2, 8, 7, 1

Set 7: Ionisation energy

Ionisation energy is the amount of energy required to remove one mole of electrons from one mole of atoms in the gaseous state. The process can be represented by:

$$X + \text{energy} \rightarrow X^{+} + e^{-}$$

The first ionisation energy is the energy required to remove the outer most electron from an atom forming a 1+ ion. The second ionisation energy is the energy required to remove the next electron, forming a 2+ ion. If the outer most electron is very loosely bound as in Group 1 and 2 elements the amount of energy required will be lower than for elements where the outer most electrons are more strongly bound ie. Group 17 and 18. The graph shows the first ionisation energy for elements in the Periodic Table.



- The ionisation energies are highest for Group 18 elements and lowest for Group 1 elements.
- As you move across a Period the ionisation energies steadily increase.
- As you move down a Group the ionisation energies steadily decrease.

Set 7: Exercises

1. State what is meant by ionisation energy using a sodium atom to illustrate your answer.

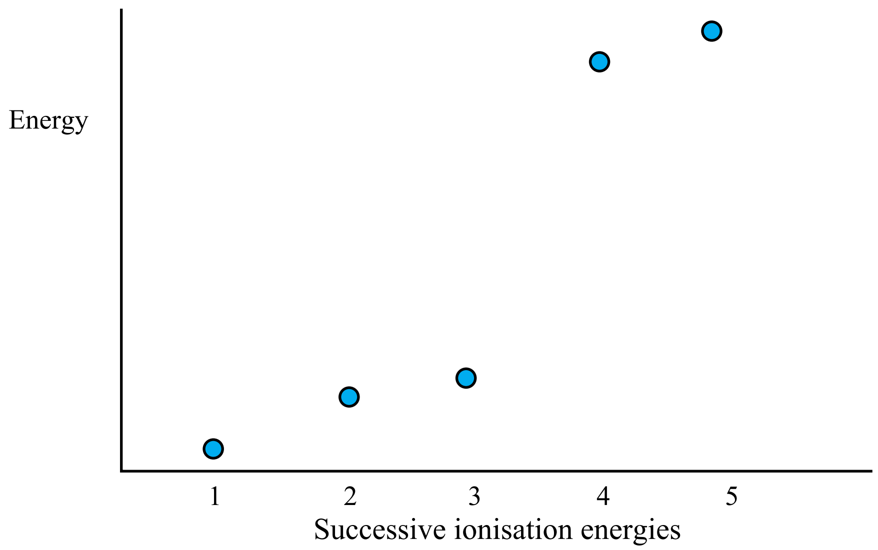
2. Why would you expect an ionised atom to be in the gaseous state?

3. "The first ionisation energy of atoms shows periodicity" State a reason for this periodicity.

4. Is there a relationship between ionisation energy and metal/non-metal properties of atoms? Explain your answer.

5. Explain why magnesium will not easily form 3+ ions.

6. Use the following graph showing consecutive ionisation energies for an unknown element to answer the questions.



(a) How many valence electrons does this element have?

(b) Justify your answer to (a).

(c) Will all the elements in the same Group as this element have similar shaped graphs? Explain.

Use the table below to answer questions 7, 8 and 9.

Element	1st ionisation energy (kJ mol^{-1})	2nd ionisation energy (kJ mol^{-1})	3rd ionisation energy (kJ mol^{-1})	4th ionisation energy (kJ mol^{-1})
A	418	3052	4420	5877
B	2080	3952	6122	9371
C	737	1450	7732	10542
D	899	1757	14849	21006
E	800	2427	3659	25025
F	590	1145	4912	6491

7. How many valence electrons does element F have? Explain your answer.

8. Which of these elements could be positioned before F in the Periodic Table? Explain your answer.

9. Which of these elements could be positioned after F in the Periodic Table? Explain your answer.

10. How can knowing the number of valence electrons help in determining the type of bonding that will occur in an element?

Set 8: Periodic trends

The Periodic Table that we use today was developed over many years and ordered all known elements into an arrangement that made it possible to make predictions about properties both physical and chemical.

Going from left to right the horizontal rows are called Periods. The vertical columns are called Groups.

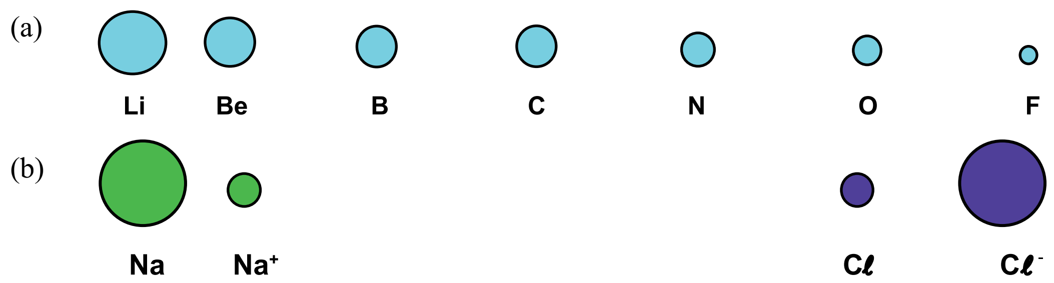
Periodic trends and patterns:
When going from left to right across a Period in the Periodic Table:

- atomic radii decrease
- electronegativity increases
- number of valence electrons increases
- ionisation energy increases
- electrical and thermal conductivity decrease
- elements change from metal to metalloid to non-metal
- bonding in elements changes from metallic to covalent network to covalent molecular

Set 8: Exercises

1. List 5 general trends or patterns observed down a Group on the Periodic Table.

2. Write two conclusions that can be made from the following diagram showing radii of atoms and ions.



3. Explain the changes in the electronegativity of elements as you move:

   (a) to the right across a Period of the table and

   (b) up a Group on the table.

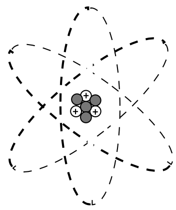
4. A new element is discovered and found to have 2 valence electrons. Write a paragraph describing the physical and chemical properties of this new element. In which Group of the Periodic Table would you place it?

5. The type of bonding changes as you move across the Periodic Table from metallic to covalent network to covalent molecular. Relate the changes in bonding to the number of valence electrons present.

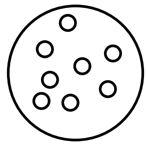
Set 9: Properties and structures of atoms

1. The Atom Timeline: A student searches for images of the structure of the atom and finds many different images, including those shown on the right. Answer the following questions about the structure of the atom.

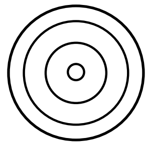
A



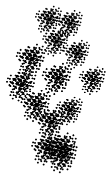
B



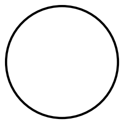
C



D



E



   (a) Why are there different representations of the atom?

   (b) What do the different images convey about the understanding of the structure of the atom?

   (c) Why do you think the structure shown in Set 5 is more accurate than structure A?

   (d) Sequence the structures to make a timeline of our understanding of atomic structure. Include the structure in Set 5.

   (e) Briefly explain what evidence was used by scientists to determine the different structural representations shown.

2. Find out how emission spectroscopy is used to help identify plastics for recycling.

3. Under what circumstances would atomic absorption spectroscopy (AAS) be used?

4. With the aid of a diagram explain how an AAS works.

5. Describe how the following isotopes are used in medical research:

(a) Technetium-99

b) Molybdenum-99

(c) Cobalt-60.

6. Each of the following scientists have made significant contributions to our understandings of the atomic model. For each scientist, list the year of their discovery and describe their contribution, including a diagram where appropriate.

Dalton
Thomson

Rutherford

Bohr

Chadwick

Set 10: Relative atomic mass and mass spectroscopy

Relative atomic mass

Scientists have known for a long time that atoms of different elements have different masses. Originally the lightest element was given a value of 1 and all other elements were measured relative to that. As the mass was a relative value (a ratio), it had no units. Today, the masses of atoms have been more accurately determined and we know that an atom of hydrogen-1 (H-1) has a mass of 1.67353×10^{-27} kg. These tiny masses are awkward to work with and so scientists still often use relative masses.

The unified atomic mass unit (u), or dalton (Da), is defined as 1/12 the mass of an atom of C-12 in its ground state and has a mass of $1.660\ 538\ 782(83) \times 10^{-27}$ kg. Using this value, an atom of H-1 will have a mass of 1.007 94(7) u. (http://goldbook.iupac.org/U06554.html)

Mass Spectrometry

Atomic mass spectrometry can be used to qualitatively and quantitatively identify elements present in a compound or mixture. These techniques can be used to measure concentrations as low as a few parts per billion. The basic principle that underlies the use of mass spectroscopy is that when charged particles move through a magnetic field they change direction. When the charge on the particles is the same, particles with the lower mass will experience a greater degree of deflection than heavier particles. When the mass is the same, the particles with a higher charge will experience a greater degree of deflection.

The first step in the process, vaporisation or atomisation, involves the separation of the substance into atoms in the gaseous phase. The atoms are then ionised, usually into +1 ions.

There are a number of ways of ionising a sample. The two methods most commonly used for mass spectroscopy are:

- inductively coupled plasma (ICP)
- electric spark

The ions produced will always be positively charged and usually single positive (+1).

Electric sparks were originally used to produce the ions. These have been replaced by ICP. When ICP is used, the atoms are usually converted to singly charged ions.

After the sample has been ionised it is passed through a mass analyser. The stream of ions is accelerated through a electric field. The charged particles are deflected by the magnetic field. The deflected particles then hit a detector which measures the number of particles and the amount of deflection. From this a mass spectrum is produced.

Mass spectra for chlorine showing the abundance of Cℓ-35 and Cℓ-37.

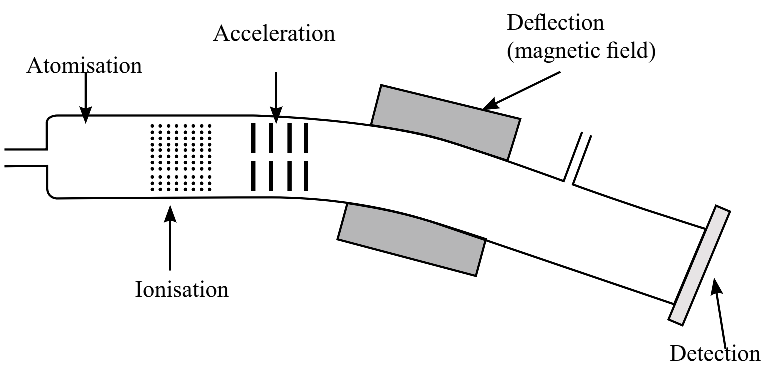
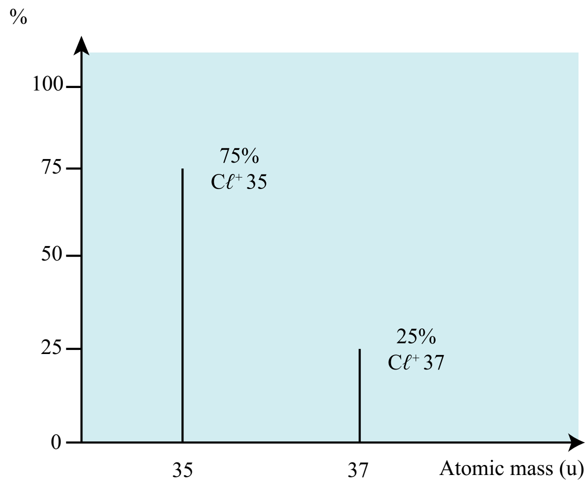


Figure 10.1: Mass spectrometer



The separation is based on mass-to-charge ratio, so as long as the charge on the ions is +1, then the mass to charge ratio will be the same as the mass of the ion. For the exercise below, you may assume that the charge on the ions produced is always +1, although you need to be aware that ions with other charges can be produced.
The process of acceleration and deflection must occur in a vacuum as there should not be any gas particles that the ions could bounce off.

In summary, the process follows the steps atomisation, ionisation, acceleration, deflection and detection. The heavier the ion, the less it will be deflected within the magnetic field. While ions of different charge can be produced in a mass spectrometer, the most common charge is +1.

Calculating relative atomic mass from isotopic composition
Atoms of the same element with a different mass number or number of neutrons are called isotopes. The relative atomic mass and molar mass of elements are derived from the average weight of the various isotopes of an element. Knowing the relative isotopic masses and amounts of each isotope, the relative atomic mass can be calculated.

Atomic mass spectroscopy is an analytical technique that can be used to determine the relative abundances of the isotopes. The technique is described in detail elsewhere.

Boron consists of two main isotopes, Boron-11 and Boron-10. In a natural sample of boron, 80.1% will be Boron-11, while the remainder is Boron-10. Calculate the relative atomic mass of boron.

In order to calculate the relative atomic mass, each isotopic mass (which is close enough to the mass number for us to use that value) is multiplied by the percentage and divided by 100.

$$A_r(B) = \frac{10(19.9) + 11\,(80.1)}{100} = 10.8$$

Given the relative atomic mass and the isotopes, it is also possible to calculate the relative abundance of them. The simplest examples consist of two isotopes. To determine the relative abundance, simultaneous equations are used. The sum of the percentage abundances will always equal 100.

Boron consists of two isotopes, boron-10 and boron-11. The relative atomic mass of boron is 10.80. What is the percentage abundance of each isotope?

Let x be the percentage of boron-10 and y be the percentage of boron-11.

$$x + y = 100 \qquad \text{equation 1}$$

$$10.80 = \frac{10(x) + 11\,(y)}{100} \qquad \text{equation 2}$$

Equation 1 can be rearranged: $y = 100 - x$ equation 3

The value for y can then be substituted into equation 2

$$10.80 = \frac{10(x) + 11\,(100\text{-}x)}{100}$$

The equation can then be arranged to determine the value for x

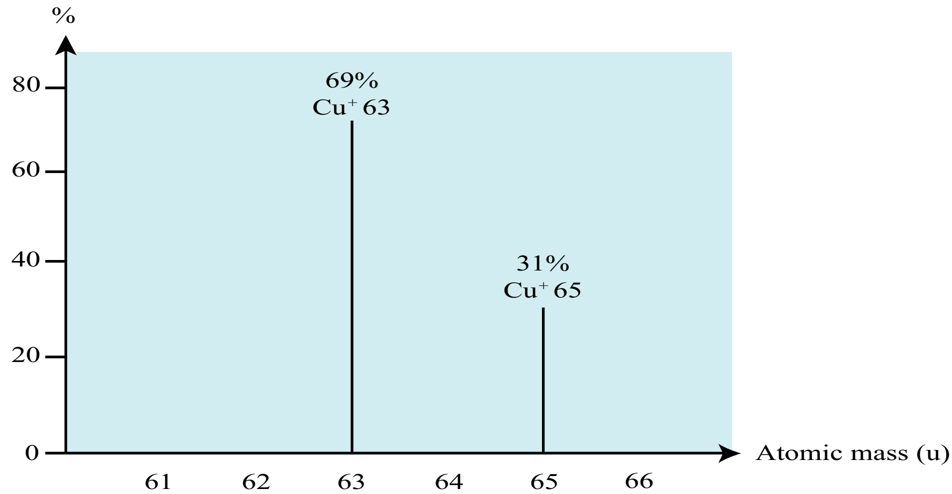
$$1080 = 10x + 1100 - 11x \qquad x = 20.0\ \%$$

Substitute the value of x into equation 3

$$y = 100 - 20 = 80.0\ \%$$

Atomic mass spectroscopy spectra

A mass spectrum produced will plot the number of ions (relative amount) against the mass-to-charge ratio or mass where the ions are singly charged.



In the example above, a sample of copper was analysed. There are two isotopes of copper present. The peaks at 63 and 65 represent +1 ions. Based on this information, the relative atomic mass of copper can be calculated.

From the spectrum above, the percentage abundance of Cu-63 is 69 and the percentage abundance of Cu-65 is 31. Given this data, we can calculate the relative atomic mass of Cu:

$$Ar(Cu) = (69 \times 63) + (31 \times 65)/100 = 63.62$$

Set 10: Exercises

1. Chlorine exists naturally as two isotopes, Cℓ-35, which accounts for 75.78% of all chlorine and Cℓ-37, which accounts for 24.22%. Calculate the relative atomic mass of chlorine.

2. There three isotopes of silicon and their relative abundances are: Si-28 (92.2%), Si-29 (4.68%) and Si-30. Calculate the relative atomic mass of silicon.

3. Four isotopes of lead and their relative abundances are Pb-204 (1.40%), Pb-206 (24.1%), Pb-207 (22.1%) and Pb-208 (52.4%). Calculate the relative atomic mass of lead.

4. A particular atom of barium has a mass of 138 while the relative atomic mass of barium is 137.33. Explain the difference between these two values.

5. Given the three isotopes of magnesium and their relative abundances of Mg-24 (79.0%), Mg-25(10.0%) and Mg-26, calculate the relative atomic mass of magnesium.

6. Copper exists as two isotopes, Cu-63 and Cu-65. Given that the relative atomic mass of copper is 63.55, determine the percentage of each isotope present.

7. Two samples of bromine, collected from different continents were found to have relative atomic masses of 79.9 and 80.9 respectively. What proportion of the stable isotopes of bromine, Br-79 and Br-81, are present in each sample?

8. Refer to the Figure 10.1 in Set 10 to help answer the questions.
   (a) Briefly describe each of the following processes that occur in a mass spectrometer
      (i) ionisation

      (ii) acceleration

      (iii) deflection

      (iv) detection

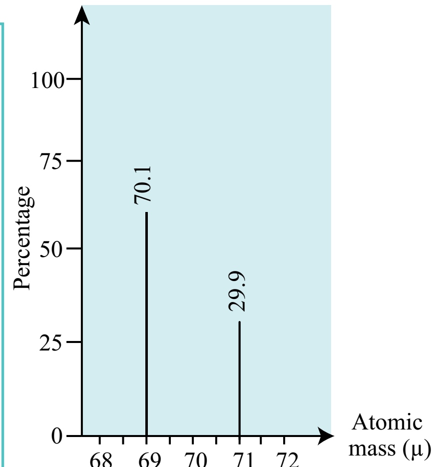
   (b) The amount of deflection of a charged particle depends on the size of the charge and the mass of the particle.
      (i) How does the size of the positive charge impact on the degree of deflection?

      (ii) How does the mass of the particle impact on the degree of deflection?

      (iii) Which of the following particles will experience the greatest deflection in each case?
         I Cu^{+}-63 or Cu^{2+}-63

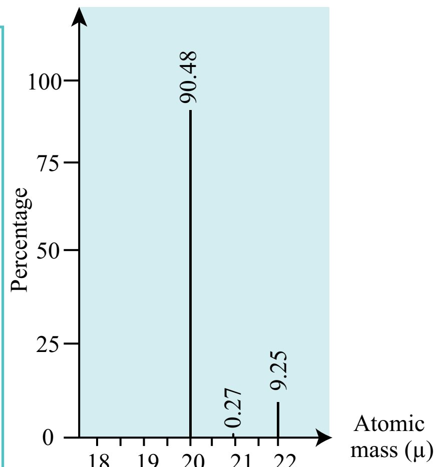
         II B^{+}-10 or B^{+}-11

9. Gallium has two isotopes. The following mass spectra was produced. Calculate the relative atomic mass of Gallium. (Ga – 69: 70.1 %; Ga – 71: 29.9 %)



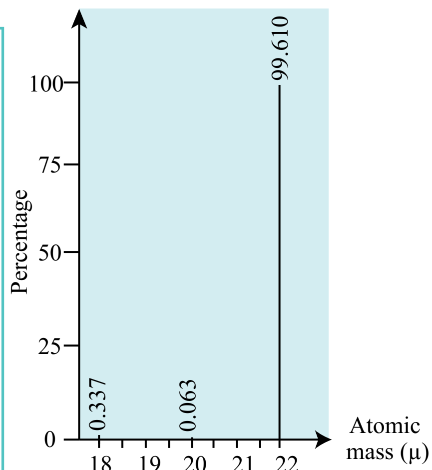
Spectrum for question 9: Gallium

10. The following mass spectra was produced from a sample of neon. Calculate the relative atomic mass of neon.
20 – 90.48 % 21 – 0.27 % 22 – 9.25 %



Spectrum for question 10: Neon

11. The following mass spectra was produced from a sample of argon. Calculate the relative atomic mass of argon.
36 – 0.337 % 38 – 0.063 % 40 – 99.610 %



Spectrum for question 11: Argon

Set 11: Molar mass

The mole concept

In order to perform calculations based on chemical reactions, we need to be able to express quantities of chemicals in terms of their mass. The mass of one atom, molecule or formula unit is too small to be practical; therefore a relative scale was devised.

Masses of atoms, molecules or formula units are compared to the mass of an atom of carbon-12. This gives us ‘relative atomic mass’.

1 unified atomic mass unit = $\frac{1}{12}$ of the mass of 1 atom of ${}^{12}_{6}C$

In chemistry it is common usage to omit the unit when reporting relative atomic mass. Whilst this is useful, chemists found it more useful to determine the number of atoms in 12.00 g of carbon-12 and this number was designated one mole of atoms. It was defined by IUPAC in 2019 to be exactly $6.02214076 \times 10^{23}$ particles. The mole concept is universally accepted and enables us to calculate molar masses of elements and compounds and perform stoichiometric calculations.

Molar mass

For ease of understanding we will refer to molar mass (M), which can be the mass of a mole of:

- atoms, eg Fe, Ne
- molecules, eg CO_2, NH_3
- formula units, eg $CuSO_4$, NaOH

The molar mass of a substance is numerically equal to the relative atomic, molecular or formula mass, but has the units g mol^{-1}.

The molar masses of elements are shown on the Periodic Table on the inside back cover of this book.

Examples

Calculate the molar mass of the following stating your answer to 3 significant figures:

1. Magnesium (Mg): M(Mg) = 24.3 g mol^{-1}

2. Ammonia (NH_3)

$$\begin{aligned} M(NH_3) &= M(N) + 3\times M(H) \\ &= (14.01) + (3 \times 1.008) \\ &= 17.034 \\ &= 1.70 \times 10^1 \text{ g mol}^{-1} \end{aligned}$$

3. Alumina ($A\ell_2O_3$)

$$\begin{aligned} M(A\ell_2O_3) &= 2\times M(A\ell) + 3\times M(O) \\ &= (2 \times 26.98) + (3 \times 16.00) \\ &= 101.96 \\ &= 1.02 \times 10^2 \text{ g mol}^{-1} \end{aligned}$$

4. Copper(II) sulfate-5-water ($CuSO_4{\cdot}5H_2O$)

$$\begin{aligned} M(CuSO_4{\cdot}5H_2O) &= M(Cu) + M(S) + 4 \times M(O) + 10 \times M(H) + 5 \times M(O) \\ &= (63.55) + (32.01) + (4 \times 16.00) + (10 \times 1.008) + (5 \times 16.00) \\ &= 249.64 \\ &= 2.50 \times 10^2 \text{ g mol}^{-1} \end{aligned}$$

Isotopes

Most elements have several naturally occurring isotopes. This is why the molar masses of elements are not whole numbers.

Set 11: Exercises

1. Calculate the molar masses of

(a) potassium hydroxide, KOH

(b) copper(II) chloride, $CuC\ell_2$

(c) calcium hydroxide, $Ca(OH)_2$

(d) ammonium oxalate, $(NH_4)_2C_2O_4$

(e) sodium carbonate-10-water, $Na_2CO_3{\cdot}10H_2O$

(f) tungsten carbide, WC

(g) methane, CH_4

(h) sucrose, $C_{12}H_{22}O_{11}$

(i) silver chloride

(j) zinc iodide

(k) iron(III) sulfate

(l) sulfur dioxide

(m) sulfuric acid

2. Carbon dating is possible due to the composition of the C-14 isotope in living organisms.

(a) How is a C-14 atom similar to and different from a C-12 atom?

(b) Are there any other isotopes of carbon?

(c) Which isotope of carbon is the most abundant and how you can verify this?

(d) How does carbon dating work?

3. Describe the composition and uses of the three isotopes of hydrogen: protium, deuterium and tritium.

Set 12: Moles, particles and mass

Atoms, molecules and ions are sub-microscopic with incredibly small masses. Yet we measure their mass and handle atoms in the laboratory all the time. So how do we do it? Small particles are dealt with in large numbers. We don't measure sugar and sand out as single grains. We use sensible units like teaspoons of sugar and truckloads of sand. Atoms are dealt with in a similar way where the sensible unit is called the mole. A mole of anything contains exactly $6.022\ 140\ 76 \times 10^{23}$ particles.

This very, very large number, $6.022\ 140\ 76 \times 10^{23}$ is called Avogadro's number. Avogadro's number of particles is called a mole. Therefore one mole of any substance will contain $6.022\ 140\ 76 \times 10^{23}$ atoms, molecules or formula units. The mole is a unit of measure used in chemistry.

For example:

- A mole of iron (Fe) consists of $6.022\ 140\ 76 \times 10^{23}$ atoms of iron.
- A mole of carbon dioxide molecules (CO_2) consists of $6.022\ 140\ 76 \times 10^{23}$ molecules of carbon dioxide.
- A mole of copper(II) sulfate ($CuSO_4$) consists of $6.022\ 140\ 76 \times 10^{23}$ formula units of copper(II) sulfate.

Moles and mass

The relationship between the number of moles (n) and mass of a substance (m), in grams, is:

$$n = \frac{m}{M}$$

where M= molar mass of a substance in g mol^{-1}.

Examples

1. How many moles of calcium atoms are there in 30.0 g of calcium?

$$n = \frac{m}{M}$$

$$= \frac{30.0}{40.08} \qquad M(Ca) = 40.08 \text{ g mol}^{-1}$$

$$n = 7.49 \times 10^{-1} \text{ mol of Ca atoms}$$

2. How many moles of benzene (C_6H_6) molecules have a mass of 390.0 g?

$$n = \frac{m}{M}$$

$$= \frac{390.0}{78.107} \qquad M(C_6H_6) = 78.107 \text{ mol}^{-1}$$

$$n = 4.993 \text{ mol of } C_6H_6 \text{ molecules}$$

3. Calculate the mass of 5.0 moles of barium sulfate ($BaSO_4$) formula units.

$$n = \frac{m}{M}$$

$$m = nM$$

$$= (5.0)(233.36) \qquad M(BaSO_4) = 233.36 \text{ g mol}^{-1}$$

$$m = 1.2 \times 10^3 \text{ g}$$

Remember to check how many significant figures to report in your answer. The general rule is to use the least number of significant figures of a measurement provided in the question, unless otherwise stated by the question.

Set 12: Exercises

1.Calculate the number of moles contained in each of the following:

(a) 72.0 g of magnesium (Mg)

(b) 4.00×10^2 g of calcium carbonate ($CaCO_3$)

(c) 104 g of ethyne (C_2H_2)

2. Calculate the mass of each of the following:

(a) 4.75 moles of lithium atoms

(b) 0.25 moles of sodium hydroxide formula units

(c) 9.00 moles of carbon monoxide molecules

3. Calculate the number of moles of molecules or formula units contained in each of the following:

(a) 28.0 g of nitrogen

(b) 232 g of butane (C_4H_{10})

(c) 3.90 g of sodium peroxide (Na_2O_2)

4. Calculate the number of moles of hydrogen peroxide (H_2O_2) molecules in 119 g of hydrogen peroxide.

5. Calculate the number of moles of sucrose ($C_{12}H_{22}O_{11}$) molecules in a can of cool drink that contains 39.8 g of sucrose.

6. Calculate the number of atoms in the following:
(a) 1.00 mole of gold (Au)

(b) 0.25 moles of sodium (Na)

(c) 2.33 moles of chlorine gas ($C\ell_2$)

(d) 0.0894 moles of phosphorus (P_4)

7. Calculate the total number of ions in each of the following:
(a) 1.11 moles of sodium chloride ($NaC\ell$)

(b) 2.98 moles of aluminium oxide ($A\ell_2O_3$)

(c) 1.11×10^{-5} moles of magnesium fluoride (MgF_2)

(d) 0.222 moles of calcium sulfate ($CaSO_4$)

8. Calculate the total number of atoms in each of the following

(a) 0.189 moles of barium phosphate ($Ba_3(PO_4)_2$)

(b) 25.0 g of bromine (Br_2)

(c) 0.0678 g of nitric acid (HNO_3)

(d) 12.5 g of acetic acid (CH_3COOH)

Set 13: Interpretation of formulae

The formula of a substance indicates the relative number of atoms of each element in the substance. For example:
One molecule of glucose, $C_6H_{12}O_6$, contains 6 atoms of carbon, 12 atoms of hydrogen and 6 atoms of oxygen.
Also one mole of glucose molecules, $C_6H_{12}O_6$ contains 6 mol of carbon atoms, 12 mol of hydrogen atoms and 6 mol of oxygen atoms.

Examples

1. How many moles of oxygen atoms are contained in 5.00 moles of calcium phosphate ($Ca_3(PO_4)_2$) formula units?
   n (oxygen atoms) $= 8 \times n(Ca_3(PO_4)_2)$ $= (8)\,(5.00)$ $= 40.0$ mol

2. What mass of oxygen atoms is contained in 132 g of carbon dioxide (CO_2)?

   (a) Find the number of moles of CO_2 using $n = \frac{m}{M}$ $\quad M(CO_2) = 44.01 \text{ g mol}^{-1}$

   $$n\,(CO_2) = \frac{m}{M} = \frac{132}{44.01} = 3.00 \text{ mol}$$

   (b) Find the number of moles of oxygen from the number of moles of CO_2

   $$n(O) = 2 \times n(CO_2) = (2)\,(3.00) = 6.00 \text{ mol}$$

   (c) Find the mass of O using $n = \frac{m}{M}$ $\quad M(O) = (16.0 \text{ g mol}^{-1})$

   $$n(O) = \frac{m}{M}$$

   $$m(O) = nM = (6.00)\,(16.0) = 96.0\text{g}$$

Set 13: Exercises

1. Calculate the number of moles of:

   (a) hydrogen atoms in 2.50 moles of calcium hydroxide ($Ca(OH)_2$) formula units

   (b) nitrogen atoms in 0.0500 moles of ammonium nitrate (NH_4NO_3) formula units

2. Calculate the number of moles of:
   (a) phosphorus trichloride ($PC\ell_3$) molecules which contain 21.0 moles of chlorine atoms

   (b) potassium permanganate ($KMnO_4$) formula units which contain 2.00 moles of oxygen atoms

3. Calculate the mass of:
   (a) oxygen in 795 g of copper(II) oxide (CuO)

(b) potassium in 1.04 g of potassium sulfate (K_2SO_4)

(c) calcium in 38.4 g of calcium oxalate (CaC_2O_4)

4. What mass of:

(a) sulfur dioxide (SO_2) contains 193 g of sulfur?

(b) ammonium sulfate ($(NH_4)_2SO_4$) contains 0.0960 g of hydrogen?

(c) octane (C_8H_{18}) contains 36.0 g of carbon?

5. Given 2.50 moles of ammonium phosphate ($(NH_4)_3PO_4$) formula units, calculate the number of moles of:

(a) ammonium (NH_4^+) ions

(b) phosphate (PO_4^{3-}) ions

(c) nitrogen atoms

(d) hydrogen atoms

(e) phosphorus atoms

(f) oxygen atoms

6. Calculate the mass of iron in 6.40×10^2 g of iron(III) oxide.

7. What mass of ethanol (C_2H_5OH) contains 144 g of carbon?

8. What mass of oxygen is contained in 2.50 moles of oxalic acid ($H_2C_2O_4{\cdot}2H_2O$)?

9. How many moles of nitrogen atoms are contained in 1.80×10^4 g of urea ($CO(NH_2)_2$)?

10. Calculate the mass of soap, sodium stearate ($NaC_{17}H_{35}COO$), that contains 1.25 g of carbon.

Set 14: Percentage composition

Today we are very interested in knowing the exact composition of many of the products which surround us. When you buy foods, there are consumer panel listing the percentage of the different ingredients present. Medications also have the amount of active ingredient listed. There is increasing consumer demand that producers should supply specific information to consumers. Chemists analyse food and medical goods to determine the exact percentage compositions of the ingredients.
The percentage composition of a chemical compound specifies the percentage by mass of each of the different elements present in the compound.

Examples

$$\% \text{ mass} = \frac{\text{mass of part}}{\text{mass of whole}} \times 100\ \%$$

1. Calculate the percentage composition of ammonium sulfate, $(NH_4)_2SO_4$ to 3 significant figures.

$\% N = \frac{28.02}{132.144} \times 100 = 21.2\ \%$

$\% H = \frac{8.064}{132.144} \times 100 = 6.10\ \%$

$\% S = \frac{32.064}{132.1} \times 100 = 24.3\ \%$

$\% O = \frac{64.0}{132.144} \times 100 = 48.4\ \%$

$\underline{100\ \%}$

$M(N) = 14.01 g\ mol^{-1}$
$M(H) = 1.008\ g\ mol^{-1}$
$M(S) = 32.06\ g\ mol^{-1}$
$M(O) = 16.0\ g\ mol^{-1}$

$M((NH_4)_2SO_4) = 132.144\ g\ mol^{-1}$

Note: The percentage should add up to 100. The difference in the calculated value is due to rounding.

2. Calculate the percentage of water of crystallisation in iron(II) sulfate-7-water, $FeSO_4{\cdot}7H_2O$ to 3 significant figures.

$\%H_2O = \frac{\text{mass}(H_2O)}{M(FeSO_4{\cdot}7H_2O)}$

$\%H_2O = \frac{7(18.016) \times 100}{278.22}$ $\quad M(FeSO_4{\cdot}7H_2O) = 278.022\ g\ mol^{-1}$

$\%H_2O = 45.4\ \%$ $\quad M(H_2O) = 18.016\ g\ mol^{-1}$

3. A sample of soft solder consisted of 4.77 g of tin and 9.54 g of lead. Calculate the percentage composition of the solder to 3 significant figures.

$\% Sn = \frac{m\ (Sn)}{m\ (\text{solder})} \times 100$

$= \frac{4.77}{(4.77 + 9.54)} \times 100$

$= 33.3\ \%$

$\% Pb = 100 - 33.3 = 66.7\ \%$

Set 14: Exercises

1. Calculate the percentage by mass of each element in:

   (a) sodium hydroxide ($NaOH$)

   (b) acetic acid (CH_3COOH)

   (c) copper(II) sulfate-5-water ($CuSO_4 \cdot 5H_2O$)

   (d) potassium phosphate (K_3PO_4)

2. Calculate the percentage by mass of:

   (a) chlorine in calcium chloride

   (b) sulfur in chromium(III) sulfide

(c) oxygen in potassium permanganate

(d) nitrogen in ammonium nitrate

3. Calculate the percentage by mass of water in:

(a) sodium carbonate-10-water

(b) nickel(II) sulfate-6-water

(c) Barium chloride-2-water

4. An alloy is prepared by melting together 25.44 g of bismuth, 15.36 g of lead and 7.20 g of tin.

(a) Calculate the percentage composition of the alloy.

(b) Calculate the mass of each metal required to make 150.0 g of the alloy.

5. An alloy used in aircraft construction consists of aluminium, copper and magnesium. An 11.34 g sample of the alloy was treated with alkali to dissolve the aluminium, leaving a residue of mass 2.73 g. This residue was treated with dilute hydrochloric acid solution to dissolve the magnesium. The remaining residue had a mass of 0.900 g. Calculate the percentage composition of the alloy.

6. A 3.030 g sample of zinc was heated in oxygen to produce zinc oxide. The mass of zinc oxide produced was 3.771 g. Calculate the percentage composition of the zinc oxide.

7. A 16.00 g sample of copper(II) oxide was reduced by reaction with hydrogen gas to yield 12.77 g of pure copper. Calculate the percentage composition of the original copper(II) oxide.

8. Copper is found in such minerals as chalcopyrite ($CuFeS_2$) and malachite ($CuCO_3 \cdot Cu(OH)_2$).

(a) Calculate the percentage by mass of copper in each.

(b) What mass of chalcopyrite must be smelted to produce 1.00×10^2 kg of copper?

9. Darling Range bauxite contains high levels of the mineral gibbsite ($A\ell_2O_3{\cdot}3H_2O$) with kaolinite clay ($A\ell_2O_3{\cdot}2SiO_2{\cdot}2H_2O$). Calculate the percentage of aluminium and the percentage of water in each of these substances.

10. Ilmenite from Capel contains approximately 53.5% titanium dioxide (TiO_2).

(a) Calculate the percentage of titanium in titanium dioxide.

(b) What mass of ilmenite must be refined to produce 1.00 tonne of pure titanium metal?

Properties and structures of materials

- Define elements, ions and compounds and write chemical formulae (for example, O, O_2 , O^{2-} , SO_4^{2-} , H_2O).
- Describe nanomaterials as substances that contain particles in the size range 1–100 nm and have specific properties relating to the size of these particles which may differ from those of the bulk material.
- Apply an understanding of differences in the physical properties of substances in a mixture, including particle size, solubility, density, and boiling point, to separate them.
- Use the type of bonding within ionic, metallic and covalent substances to explain their physical properties, including melting and boiling points, conductivity of both electricity and heat and chardness.
- Explain chemical bonds are electrostatic attractions that arise because of the sharing or transfer of electrons between participating atoms and that the valency is a measure of the bonding capacity of an atom.
- Describe the allotropes of carbon, including graphite, diamond and fullerenes, with significantly different structures and physical properties
- Explain the properties of covalent molecular substances, including low melting point, in terms of their structure and the weak intermolecular forces between molecules and their non-conductivity in the solid and liquid/molten states can be explained by the absence of mobile charged particles in their molecular structure.
- Write and name molecular formulae using the number and type of atoms present in the molecules.
- Calculate the percentage composition of compounds using the relative atomic masses of the elements in the compound and the formula of the compound.
- Use IUPAC nomenclature to name and draw molecular structural formulae (condensed or showing bonds) of hydrocarbons, alkanes and alkenes from C 1 - C 8
- Identify and write characteristic reactions of alkanes, alkenes and benzene such as combustion, addition reactions for alkenes and substitution reactions for alkanes and benzene.

Set 15: Compounds and formulae

Chemical symbols and formulae are a shorthand way of communicating information about the structure and composition of elements and compounds. For a molecule the formula indicates what elements and how many of each type are present. So a molecule of carbon dioxide which is represented by the formula CO_2 consists of one carbon atom and two oxygen atoms. For metals, covalent network structures and ionic compounds the formulae indicate the ratio of elements present. So calcium chloride which is represented by the formula $CaC\ell_2$ consists of a lattice of positive calcium ions and negative chloride ions. For every calcium ion there are two chloride ions present, but the total number of ions is not known.

Writing the formulae of molecular compounds

Common names are often used for molecular compounds. Some common names and their formulae that you should know are shown in the table (below).

Common Name	Formula
Water	H_2O
Hydrogen peroxide	H_2O_2
Ammonia	NH_3
Methane	CH_4
Hydrogen chloride	$HC\ell$
Sulfuric acid	H_2SO_4
Phosphoric acid	H_3PO_4
Nitric acid	HNO_3
Acetic acid	CH_3COOH

For molecular compounds without common names the name of the compound usually includes information about the number of each element present. For example sulfur dioxide is written as SO_2. Prefixes, listed in the following table, are used to show the number of atoms of each element.

Prefix	Number of atoms
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8

Rules for naming molecular compounds:	Example: CO_2
1. The name of the element closer to the bottom left of the Periodic Table is written first.	Carbon is further to the left than oxygen so is named first
2. The name of the second element is changed to end with the suffix *-ide*	Oxygen is the second element so will be written as *oxide*
3. Prefixes are then used to identify the number of atoms of each element in the compound. (If there is only one atom of the first named element no prefix is required.)	Carbon – no prefix needed Oxygen – 2 atoms so use the prefix di - *Carbon dioxide*

Writing the formulae of ionic compounds
Writing the formula for an ionic compound requires knowledge of the charge on the ions present in the compound. As the overall charge in a compound must be zero, the charges on the ions present must add up to zero. The name only contains the name of the ions. Write the positive ion first.

Examples

1. Writing the formula of sodium fluoride

Ions involved:	Na^{+}	F^{-}	
Use subscript numbers to balance charge	Na_1^{+}	F_1^{-}	
Charges on the ions add to zero	+1	-1	= 0

NaF

2. Writing the formula of calcium fluoride

Ions involved:	Ca^{2+}	F^{-}	
Use subscript numbers to balance charge	Ca_1^{2+}	F_2^{-}	
Charges on the ions add to zero	+2	-2	= 0

CaF_2

Set 15: Exercises

1. Name the following molecular substances

(a) CO

(d) N_2S

(b) SO_2

(e) P_2Br_4

(c) $PC\ell_5$

(f) SF_6

2. Write the formula of each of the following molecular compounds.

(a) nitrogen monoxide

(d) sulfur trioxide

(b) nitrogen dioxide

(e) water

(c) dinitrogen tetroxide

(f) pentaphosphorus decaoxide

(g) hydrogen chloride

(h) hydrogen iodide

(i) phosphorus tribromide

(j) ammonia

3. Write the formula for each of the following ionic compounds.
   (a) lithium chloride
   (b) silver iodide
   (c) potassium nitrate
   (d) caesium ethanoate
   (e) barium bromide
   (f) copper(II) sulfate
   (g) manganese(IV) oxide
   (h) nickel nitrate
   (i) aluminium oxide
   (j) chromium(III) sulfate
   (k) lead(IV) phosphate
   (l) ammonium dichromate

4. Name the following compounds
   (a) CO
   (b) N_2O
   (c) H_2S
   (d) H_2O_2
   (e) H_3PO_4
   (f) Ca_3N_2
   (g) HNO_3
   (h) $CoHPO_4$
   (i) $CuC\ell$
   (j) $FeSO_4$

Set 16: Bonding and properties

Using the general properties of pure substances such as conductivity as solid, conductivity as liquid, melting point, hardness and brittleness, it is possible to see that most of these substances fall into one of four categories. We can link these properties to the types of bonding present within the substances. The four categories of substances are metallic, ionic, covalent network and covalent molecular. Each of these categories has a characteristic set of physical properties.

Set 16: Exercises

1. From the physical properties listed in the table, predict the bonding class: metallic, ionic, covalent network or covalent molecular.

	Colour	Hardness	Electrical conductivity as solid	Electrical conductivity as liquid	Melting point (°C)	Bonding class
(a)	yellow	brittle	no	no	113	
(b)	yellow	soft	yes	yes	1064	
(c)	yellow	brittle	no	yes	402	
(d)	white	soft	no	no	37	
(e)	white	brittle	no	yes	801	
(f)	white	brittle	no	no	146	
(g)	brown	hard	yes	yes	1085	

2. In terms of the metallic bonding model state reasons for each of the following properties:

(a) high melting point

(b) electrical conductivity

(c) thermal conductivity

(d) malleability

3. In terms of the covalent bonding model state reasons for each of the following properties:

(a) high melting point

(b) electrical conductivity

(c) hardness

(d) brittleness

4. Diamonds, graphite and fullerene are all allotropes of carbon.

(a) Define the term allotrope and give an example of one other element that exists as allotropes.

(b) For each of the three allotropes, describe	Diamond	Graphite	Fullerene
i. Colour			
ii. Melting point			
iii. Electrical conductivity			
iv. Hardness			

(c) For each of the three allotropes, explain, in terms of their bonding and using diagrams, the following properties

Diamond
i. Melting point

ii. Electrical conductivity

iii. Hardness

Graphite
i. Melting point

ii. Electrical conductivity

iii. Hardness

Fullerene

i. Melting point

ii. Electrical conductivity

iii. Hardness

5. In terms of the ionic bonding model state reasons for each of the following properties:

(a) high boiling point

(b) electrical conductivity when solid

(c) electrical conductivity when liquid/molten

(d) brittleness

(e) hardness

Set 17: Properties and structures research

1. What is the composition of solder? Compare the individual properties of the two metal components with the properties of the alloy.

2. How and why does the amount of carbon in steel change the properties of iron?

3 . Why are carbon steel alloys more brittle and less malleable than the pure metal itself?

4. Perspex is said to remember it has been deformed for a while. Explain. Are there any metals or alloys with this property?

5. How can we use freezing point to decide if something is pure?

6. Nano-particles are very small – about 1/2000th the diameter of a human hair. Matter at the nanoscale can be manipulated to create new materials, composites and devices. The different characteristics of nanomaterials can be used to provide commercially available products. As products are designed on the basis of properties, which are different from the bulk material, their use can be associated with potential risks to health, safety and the environment and this has led to regulations being developed to address new and existing nanoform materials

   (a) Identify and explain three potential benefits of using nano-particles.

(b) Some people have concerns about the safety of nano-particles. Identify and explain three problems that might be associated with their use.

7. Zinc oxide is commonly used as sunscreen in Australia. Zinc oxide sunscreen exists in both the nanoscale and non-nanoscale forms.

(a) For each form, list the advantages and disadvantages.

(b) What concerns have been raised about the use of zinc oxide sunscreens?

(c) What does the current research suggest about how safe it is to use nanoscale zinc oxide sunscreens?

Set 18: Electron dot diagrams

Electron dot diagrams are a useful tool to help you understand the bonding arrangements of atoms. From these diagrams it is possible to determine the shape of molecules and also their polarity. They are quite easy to draw for simple molecules as long as you follow some simple guidelines. Try to satisfy a full outer level of electrons (the octet rule) in all atoms except hydrogen (2 electrons) by:

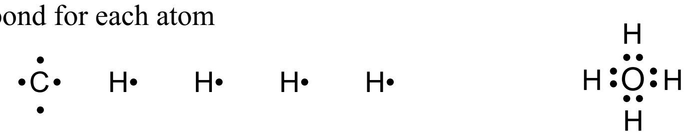
- Determining the total number of valence electrons from all the atoms involved;
- Using a pair of electrons to form a single bond between each pair of bonded atoms;
  If there are not enough electrons, consider a multiple covalent bond.
- Placing the remaining electrons to satisfy the octet rule. Be sure the electrons are drawn in pairs.

Examples

1. **Methane CH_4**

Carbon has four outer shell electrons and each of the hydrogen atoms has one electron, giving a total of 8 valence electrons.

- Trying a single covalent bond for each atom



A single bond between each carbon and hydrogen will satisfy all the elements.

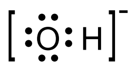
2. **Carbon dioxide CO_2**

Each oxygen has six outer shell electrons and carbon has four outer shell electrons, giving a total of 16 valence electrons.

- Try a single covalent bond for each atom
  Single bonds will not satisfy the carbon, there are not enough electrons.
- If there are not enough electrons, consider a multiple covalent bond. A double bond between each oxygen and carbon will satisfy all the elements.
- Place the remaining electrons to satisfy the octet rule. Be sure the electrons are drawn in pairs.

3. **Hydroxide ion OH^-** (IUPAC now uses HO^-)

Oxygen has six outer shell electrons and hydrogen has one outer shell electron. As a single negative ion, an extra electron has been gained, giving a total of 8 valence electrons.



A single bond between each oxygen and hydrogen will satisfy the elements, with the extra electron completing an octet for oxygen and is shown as a -1 charge on the ion.

4. **Ionic substances**

Square brackets are used when representing ions in electron dot diagrams of ions and ionic substances.
E.g. in NaCℓ, sodium has one valence electron and when forming an ionic bond becomes a positive ion by losing its valence electron to a chlorine atom.
Chlorine has seven valence electrons and accepts an electron when forming ionic bonds becoming a negative ion.

sodium atom chlorine atom sodium chloride

1. Draw electron dot diagrams for the following atoms and ions:

(a) Na	(d) S^{2-}
(b) Br	(e) H^{+}
(c) P	(f) N^{3-}

2. Draw electron dot diagrams for the following covalent molecular compounds (The central atom is in **bold** type) :

(a) $H_2\mathbf{O}$	(d) $H_2\mathbf{S}$
(b) $\mathbf{C}H_3C\ell$	(e) $\mathbf{C}HC\ell_3$
(c) $\mathbf{P}H_3$	(f) $H\mathbf{C}N$

3. Draw electron dot diagrams for the following ionic compounds:

(a) NaOH	(d) K_2S
(b) $CaC\ell_2$	e) $Mg(OH)_2$
(c) Fe_2O_3	(f) $AgNO_3$

Set 19: Naming and drawing hydrocarbons

The name of any simple organic compound is based on its structure. It ensures that, given the structure, a name can be constructed and if given a name, a structure can be drawn. The rules that govern this naming convention were adopted and first published by the Union of Pure and Applied Chemistry (IUPAC) in 1958. This set of rules has become known as the IUPAC Nomenclature of Organic Compounds.

Aliphatic hydrocarbons

For straight chain (aliphatic) hydrocarbons the rules can be summarised as follows.

1. The prefix of the name is used to identify the number of carbon atoms in a continuous chain (the chain length).

Number of carbon atoms	Prefix	Prefix for alkyl group
1	meth-	methyl
2	eth-	ethyl
3	prop-	propyl
4	but-	butyl
5	pent-	pentyl
6	hex-	hexyl
7	hept-	heptyl
8	oct-	octyl

2. The suffix of the name is used to identify the type of aliphatic hydrocarbon, that is, alkane, alkene or alkyne (alkynes are not in the current WA syllabus).

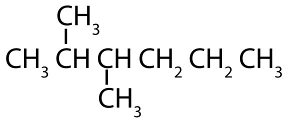
Type of hydrocarbon	Functional group	Suffix
alkane	$C-C$	-ane
alkene	$C=C$	-ene
alkyne	$C\equiv C$	-yne

3. Groups other than alkyl groups can result from addition and substitution reactions. The names of halogen groups are as follows.

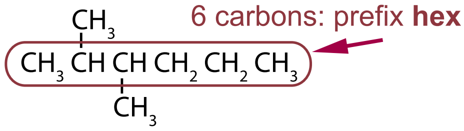
Type of substituted hydrocarbon	Functional group	Name of group
fluorocarbon	-F	fluoro-
chlorocarbon	-Cℓ	chloro-
bromocarbon	-Br	bromo-
iodocarbon	-I	iodo-

The steps for naming hydrocarbons

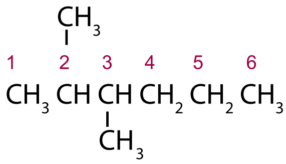
The following structure will be used to illustrate the rules for naming hydrocarbons in each step:



Step 1: Identify the longest continuous carbon chain. This determines the prefix for the name.



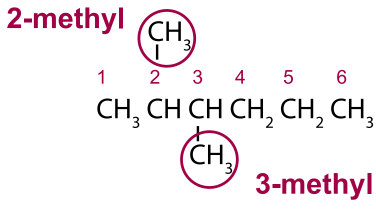
Step 2: Assign a number to each carbon atom in the longest continuous chain starting at the end of the chain that results in the lowest possible number for the position of functional groups.



Step 3: Identify the presence and location of any double bond.This determines the suffix of the name: -ane (only single bonds), -ene (double bond)

suffix: **- ane**
(only single bonds present)

Step 4: Identify all groups attached to the longest carbon chain.

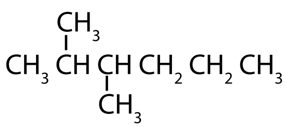


Step 5: Add the names of the groups in front of the chain name, indicating the position of the group using the number of the carbon to which it is attached. If a group occurs more than once, use the prefix di-, tri-, tetra-, etc to indicate the number present and indicate the position of each as described above. A number for the location of groups is often used even though it may be redundant. The locating number is placed as close to the part of the name referring to the group, as in butan-1-ene.

2,3-dimethyl

Step 6: Substituent groups are written alphabetically on the basis of the group name. Numbers are separated from each other by commas and numbers are separated from names by hyphens.

2,3-dimethylhexane



The steps for naming hydrocarbons

The compound 1-fluoro-2-methylbutane will be used to illustrate the rules for drawing structural formulas of hydrocarbons:

Step 1: Identify and draw a carbon skeleton of the parent chain.

1-fluoro-2-methyl**butane**

C - C - C - C

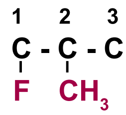
Step 2: Number the carbons in the chain.

1 2 3 4

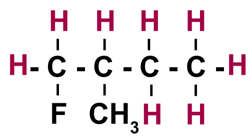
C - C - C - C

Step 3: Add the funtional groups to the appropriate carbons on the chain.

1-fluoro-2-methylbutane

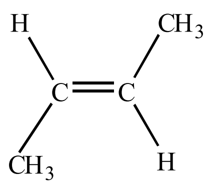


Step 4: Add hydrogen atoms to give each carbon four (4) bonds.

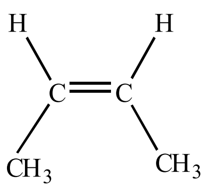


E-Z (cis-trans) isomers

In alkenes, rotation about the double bond is not possible. This results in the existance of *E-Z* isomers, this type of isomer was formerly called *cis-trans* isomers. The atoms and their bonding is the same in *E-Z* isomers, but the arrangement of atoms in space can be different. The two different isomers are distinguished by placing the prefix ***E-*** or ***Z-*** in front of the name, as follows:



E-but-2-ene
(*trans*-but-2-ene)



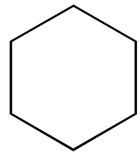
Z-but-2-ene
(*cis*-but-2-ene)

Alicyclic compounds

These are ring structures formed when the ends of a carbon chain join.
The rules for naming alicyclic (carbons forming rings) hydrocarbons are mostly the same as for those used to name aliphatic (straight chains) hydrocarbons. There are some differences though, which include the following.

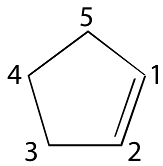
1. The prefix cyclo- is placed in front of the name indicating the number of carbon atoms.

cyclohexane (C_6H_{12})



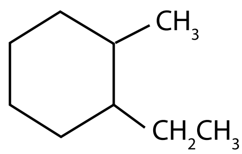
2. The carbon atoms are numbered starting at the functional group and for double bonds the numbering is such that the carbon atoms either side of the multiple bond have the numbers 1 and 2.

cyclopentene (C_5H_8)



3. If more than one type of group is attached the numbering starts at the group that is first in alphabetical order.

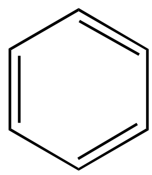
1-ethyl-2-methylcyclohexane (C_9H_{18})



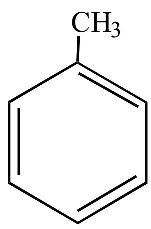
Aromatic compounds based on benzene

The naming of these compounds is very similar to the alicyclic compounds and because the compounds we will consider are based on the benzene structure, the names are based on the name benzene. This group is called 'aromatic'. For example:

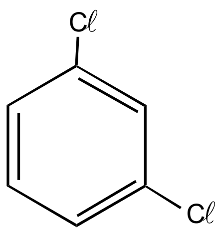
(a) benzene (C_6H_6)



(b) methylbenzene, also known as toluene (C_7H_8)



(c) 1,3-dichlorobenzene ($C_6H_4C\ell_2$)



Set 19: Exercises

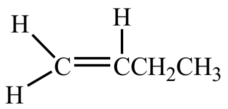
Petroleum is a complex mixture of hydrocarbons and can include some of the substances in **questions 1 and 2**.

1. Write systematic names for the following compounds.

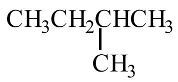
(a)

$CH_3CH_2CH_2CH_2CH_3$

(b)



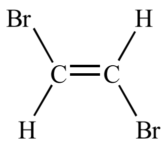
(c)



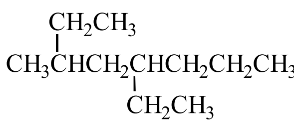
(d)

$CHC\ell_2CH_2C\ell$

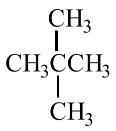
(e)



(f)



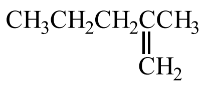
(g)



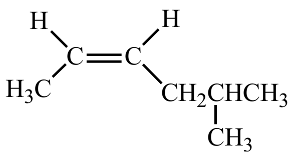
(h)

$CF_3CH_2CH_2CH_2CH_3$

(i)



(j)



2. Draw structural formulae for the following compounds.

(a) 2,2,4-trimethylpentane

(b) dichlorodifluoromethane

(c) 3-ethyl-2-methylpent-2-ene

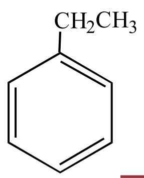
(d)	4,4-diethyloctane
(e)	5,5-dichloro-4-methylhex-1-ene
(f)	*E*-hept-3-ene
(g)	1,1-dichloro-*cis*-but-2-ene
(h)	5-ethylhept-1-ene

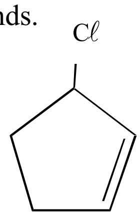
3. Petroleum contains many substances that have the same formula but different structures. Draw the structural isomers and write systematic names for

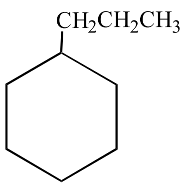
(a)	all the isomers of: (i) pentane	
	(ii) pentene	
(b)	four isomers of C_4H_9Br	

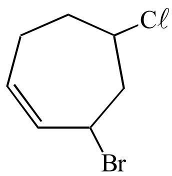
Alicyclic and aromatic compounds are often used as starting materials for important agricultural chemicals including pesticides and herbicides. Examples of substituted alicyclic and aromatic compounds are included in **questions 4 and 5**.

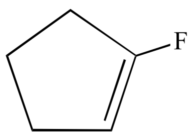
4. Write systematic names for the following compounds.

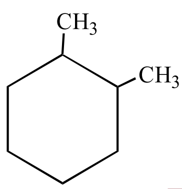
(a) 

(b) 

(c) 

(d) 

(e) 

(f) 

5. Draw structural formulae for the following compounds.

(a) fluorocycloheptane

(b) 4-methylcyclopentene

(c) butylbenzene

(d) 1,2-difluorobenzene

(e) 1,3-dibromobenzene

(f) 1-ethyl-4-methylbenzene

6. (a) Petroleum contains a number of different compounds with the formula C_4H_8
Draw all the isomers of C_4H_8

(b) Benzene, with some of its hydrogen atoms substituted with halogens, can be a starting material for the synthesis of pesticides.
Draw all the isomers of dichlorobenzene.

Set 20: Reactions of hydrocarbons

The reactions you will deal with in this section are combustion, substitution and addition reactions.

Examples

Complete and partial combustion of hydrocarbons as shown by the combustion of propane.

Complete combustion

When a fuel burns in an excess of air, it receives enough oxygen for complete combustion

hydrocarbon + oxygen → carbon dioxide + water

$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g) \quad \Delta H° = –2220\ kJ\ mol^{–1}$

Incomplete combustion

When oxygen supply is limited water is still produced, but carbon monoxide or solid carbon (soot) are produced instead of carbon dioxide.

$C_3H_8(g) + 3½\ O_2(g) \rightarrow 3\ CO(g) + 4\ H_2O(g) \quad \Delta H = –1579.4\ kJ\ mol^{–1}$

$C_3H_8(g) + 2\ O_2(g) \rightarrow 3\ C(s) + 4\ H_2O(g) \quad \Delta H = –1038.5\ kJ\ mol^{–1}$

Substitution in alkanes as shown in the chlorination of propane

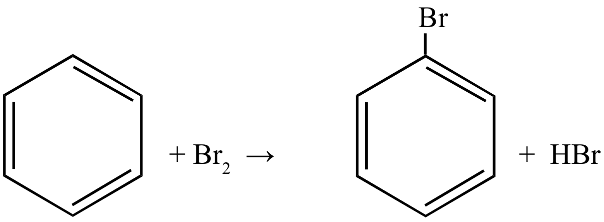
$C_3H_8 + C\ell_2 \rightarrow C_3H_7C\ell + HC\ell$

Addition in alkenes as shown in the bromination of pent-2-ene

$CH_3CH_2CH = CHCH_3 + Br_2 \rightarrow CH_3CH_2CHBrCHBrCH_3$

In addition reactions involving unsymmetrical reagents such as water or hydrogen halides, **Markovnikov's rule** can be applied to identify the most likely product. *Most hydrogen atoms add to the carbon atom that already has the greater number of hydrogen atoms.*

Substitution in benzene as shown in the bromination of benzene



Set 20: Exercises

1. Hydrocarbons are mostly used as fuels as they all undergo combustion to produce heat. Write an equation showing each of the following substances ignited in excess oxygen.
   (a) propane.

   (b) ethene.

   (c) benzene.

   (d) ethylbenzene.

2. Hydrocarbons are important starting materials for the production of many useful substances. Write equations for and name any organic products formed in the reactions between

(a) ethane and chlorine in the presence of ultraviolet radiation

(b) propene and bromine

(c) but-2-ene and hydrogen chloride

(d) pent-2-ene and hydrogen in the presence of a platinum catalyst

(e) cyclopentene and hydrogen fluoride

(f) benzene and bromine

(g) propene in excess chlorine

3. As a research chemist you are employed to investigate the conditions that give the best yield of halocarbons starting with simple hydrocarbons. You are required to make the following substances. If there is more than one type of reaction that could produce the product, use the reaction that occurs most readily.

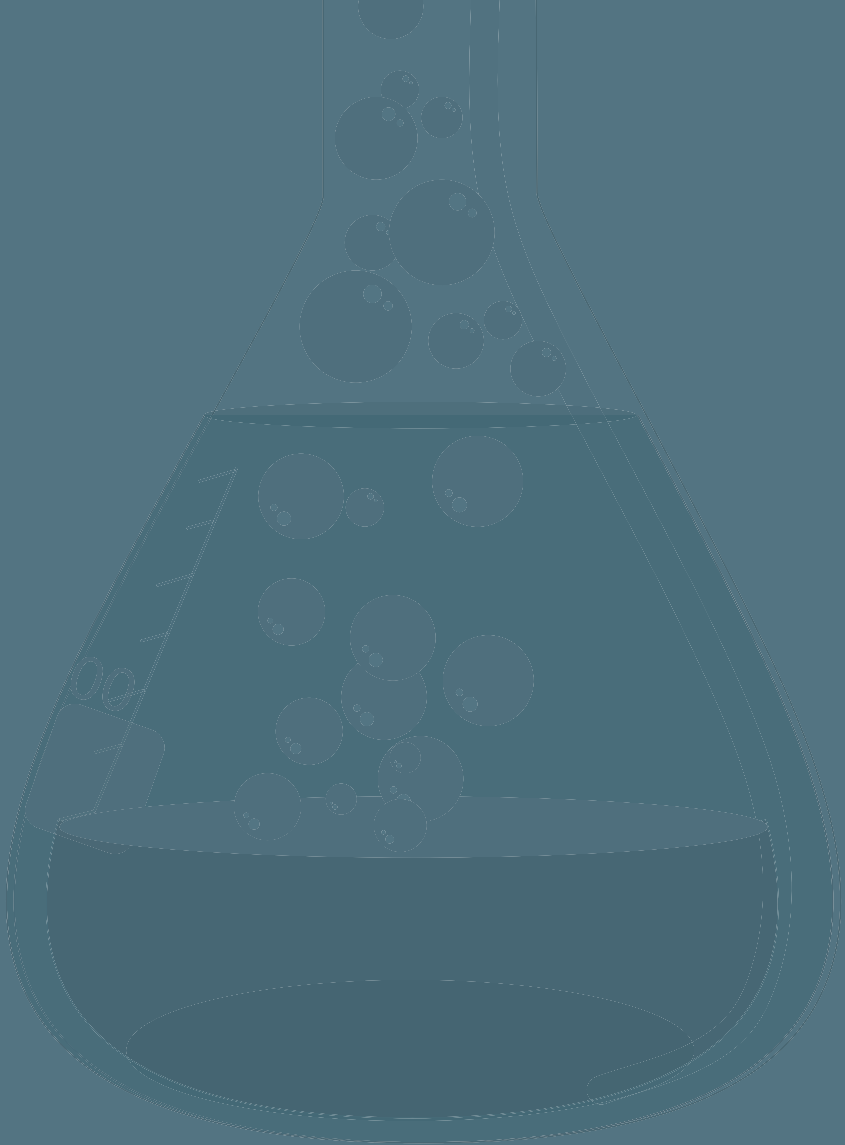
   (i) In each case draw the structure and name the starting hydrocarbon and name any other reagents required.
   (ii) Write equations for each step of the process.

(a)	chlorofluoromethane
(b)	chloroethane
(c)	1,2-dichloroethane
(d)	2-chloropropane
(e)	iodobenzene
(f)	chlorobenzene

Chemical reactions: reactants, products and energy change

- Write and balance chemical equations.
- Identify and describe energy changes for endothermic and exothermic reactions and explain in terms of the Law of Conservation of Energy and the breaking of existing bonds and forming of new bonds.
- Define the mole as the amount of matter equal to Avogadro’s number of particles.
- Use the mole concept to calculate the masses of reactants and products in a chemical reaction.

The foods we eat and fuels we use have differences in energy output. Fossil fuels and biofuels also differ in their carbon emissions. How widely we use these fuels and their future use is not only determined by their chemistry, but is influenced by social, economic, cultural and political values.



Set 21: Reactions, equations and observations

An important skill in chemistry is the ability to predict the products when given some reactants. This requires knowledge of the general reaction types. From the products, an equation can be written and the observations expected can be described. Observations include the dissolving or formation of solids (with colours), evolution of gases (including colours and odour) and changes in solution colour. They may also include the formation or absorption of heat and even the production of light and sound.

Examples

Solid sodium carbonate is added to dilute hydrochloric acid solution.

Step 1 Identify the type of reaction.
This is an acid + carbonate reaction and has the general equation
Acid + carbonate → salt + water + carbon dioxide

Step 2 From the general equation, predict the products and write the ionic equation.
$Na_2CO_3(s) + 2\ H^+(aq) \rightarrow 2\ Na^+(aq) + H_2O(\ell) + CO_2(g)$

Step 3 Describe the observations expected
White solid dissolves; colourless, odourless gas evolved; colourless solution formed.

Set 21: Exercises

For the following reactants, write balanced equations and describe the expected observations.

1. Marble chips (calcium carbonate) are added to excess dilute nitric acid.

2. Magnesium metal is added to excess dilute hydrochloric acid.

3. Baking soda (sodium hydrogencarbonate) is added to excess vinegar (acetic acid).

4. Sodium hydroxide solution is added to sulfuric acid solution.

5. Solid sodium hydroxide is added to excess dilute hydrochloric acid.

6. Excess dilute sulfuric acid is added to solid cobalt carbonate.

7. Barium chloride solution is added to excess dilute sulfuric acid.

8. Lead(II) nitrate solution is added to excess sodium iodide solution.

9. Gold(III) chloride solution is added to excess solid copper.

10. A small piece of sodium metal is added to water.

Set 22: Stoichiometry

A driver has been stopped by the police as part of a random breath test. There are a number of analytic techniques that can be used to determine the amount of alcohol in the driver's blood. Many of these techniques rely on reacting the alcohol with other chemicals and measuring the amount of products formed. In order to determine the amount of alcohol an understanding of the relationship between the reactants (including alcohol) and the products in the reaction is required. This relationship, expressed with the aid of a chemical equation, is called stoichiometry.

For example the equation representing the burning of ethanol is:

$$C_2H_5OH\ (\ell) + 3\ O_2\ (g) \rightarrow 2\ CO_2\ (g) + 3\ H_2O\ (g)$$

This indicates that one mole of C_2H_5OH reacts with three moles of O_2 to form two moles of CO_2 and three moles of H_2O.

Using reaction stoichiometry, the relationship between the number of moles, masses and gaseous volumes of reactants and products can be calculated. We will start with the relationship between masses and moles.

Examples

1. When a mixture of hydrogen and oxygen is sparked an explosive reaction occurs in which water is formed. If 0.480 moles of hydrogen gas is mixed with excess oxygen and sparked, calculate:
   (a) the number of moles of water formed
   (b) the mass of water formed
   (c) the number of moles of oxygen consumed
   (d) the mass of oxygen consumed.

   (a)(i) Write a balanced equation:

   $$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$$

   (ii) Identify the unknown and relate it to the quantity which is known:

   $n(H_2O) = n(H_2)$

   $= 0.480$ mol

   (b) Convert the moles to mass (m=nM):

   $n(H_2O) = 0.480$

   $m(H_2O) = 0.480\ (18.0)$ $\quad$ $M(H_2O) = 18.0\ g\ mol^{-1}$

   $= 8.60$ g

   (c) Relate the unknown to the known (using the equation):

   $n(O_2) = ½ \times n(H_2)$

   $= (½)(0.480)$

   $= 0.240$ mol

   (d) Convert moles to mass (m = nM):

   $n\ (O_2) = 0.240$

   $m\ (O_2) = (0.240)\ (32.0)$ $\quad$ $M(O_2) = 32.0\ g\ mol^{-1}$

   $= 7.68$ g

2. 17.8 g of chlorine gas reacts with hydrogen to form hydrogen chloride gas.
Calculate:
(a) the number of moles of hydrogen chloride formed and;
(b) the mass of hydrogen chloride formed.

(a) (i) Write a balanced equation:

$$H_2(g) + C\ell_2(g) \rightarrow 2\ HC\ell(g)$$

(ii) Identify the unknown and relate it to the quantity which is known:

$n(HC\ell) = 2 \times n(C\ell_2)$

$= 2 \times \frac{17.8}{70.9}$ $\qquad M(C\ell_2) = 70.9\ g\ mol^{-1}$

$n(HC\ell) = 0.502\ mol$

(b) Convert the moles to mass

$n(HC\ell) = 0.502$

$m(HC\ell) = (0.502)(36.458)$ $\qquad M(HC\ell) = 36.458\ g\ mol^{-1}$

$= 18.3\ g$

Set 22: Exercises

1. The equation for the combustion of butane is:

$$2\ C_4H_{10}(g) + 13\ O_2(g) \rightarrow 8\ CO_2\ (g) + 10\ H_2O(g)$$

Calculate the number of moles of:

(a) CO_2 produced in the combustion of 1.00 mole of C_4H_{10}

(b) H_2O produced in the combustion of 3.00 moles of C_4H_{10}

(c) O_2 consumed in the combustion of 0.600 moles of C_4H_{10}

2. 0.0300 moles of magnesium oxide is completely reacted with nitric acid to form a solution of magnesium nitrate: $MgO + 2\ HNO_3 \rightarrow Mg(NO_3)_2 + H_2O$

or $MgO(s) + 2\ H^+(aq) \rightarrow Mg^{2+}(aq) + H_2O(\ell)$

Calculate the number of moles and mass of:

(a) nitric acid required

(b) magnesium nitrate formed

3. When silver nitrate solution is added to a solution of calcium chloride, a white precipitate of silver chloride is produced

$$CaC\ell_2\,(aq) + 2\ AgNO_3(aq) \rightarrow Ca(NO_3)_2\,(aq) + 2\ AgC\ell(s)$$

If 0.200 moles of silver chloride is formed, calculate the number of moles and masses of :

(a) silver nitrate required

(b) calcium chloride required

(c) calcium nitrate formed in solution

4. A sample of copper(II) oxide was reacted with sulfuric acid and the solution evaporated to dryness to yield 3.14 g of $CuSO_4{\cdot}5H_2O$:

$$CuO(s) + H_2SO_4\,(aq) + 4\ H_2O\ (\ell) \rightarrow CuSO_4{\cdot}5H_2O(s)$$

(a) Calculate the mass of sulfuric acid required.

(b) Calculate the moles of copper(II) oxide dissolved.

5. Write a balanced equation for the decomposition of potassium chlorate ($KC\ell O_3$) into potassium chloride and oxygen gas by heating. How many moles of oxygen would be formed from the decomposition of 0.800 mol of $KC\ell O_3$?

6. Write a balanced equation for the decomposition of mercury(II) oxide into mercury and oxygen gas (O_2). What mass of oxygen would be formed from the decomposition of 0.240 mol of mercury(II) oxide?

7. Excess nitric acid was added to 3.00 g of calcium carbonate:

$$CaCO_3(s) + 2\ HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(\ell) + CO_2(g)$$

or $$CaCO_3(s) + 2\ H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(\ell) + CO_2(g)$$

Calculate the masses of:

(a) nitric acid consumed

(b) carbon dioxide produced

(c) calcium nitrate formed in solution

8. In the processing of uranium, one of the steps involves converting UO_2 to UF_6.

$$UO_2(s) + 4\ HF(g) + F_2(g) \rightarrow UF_6(g) + 2\ H_2O(\ell)$$

For 7.50 kg of UO_2, calculate:

(a) the mass of hydrogen fluoride required

(b) the mass of fluorine required

(c) the mass of UF_6 produced

Set 23: Energy changes

Observable changes in chemical reactions and physical changes can be described and explained at an atomic and molecular level. The energy changes in chemical reactions come from breaking chemical bonds of reactants and forming new chemical bonds in products. The breaking of bonds generally requires energy from the surroundings and is an endothermic process. The formation of new bonds generally releases energy to the surroundings and is an exothermic process. The difference in energy between the breaking and forming of bonds is called the heat of reaction or enthalpy, delta H (ΔH).

$$\Delta H = E_P \text{ (products)} - E_R \text{ (reactants)}$$

Energy profile diagrams can be used to represent energy changes in chemical systems (physical changes and chemical changes). Endothermic and exothermic reactions are summarised in the table below, which also includes their energy profile diagrams.

<table>
<tr>
<td>Endothermic reactions
• absorb energy from the surroundings
• the surroundings become cold
• energy is a reactant
• enthalpy, ΔH is positive (+ ΔH)</td>
<td>Exothermic reactions
• release energy to the surroundings
• the surroundings become hot
• energy is a product
• enthalpy, ΔH is positive (+ ΔH)</td>
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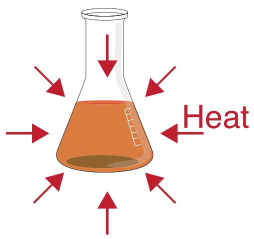
<td>
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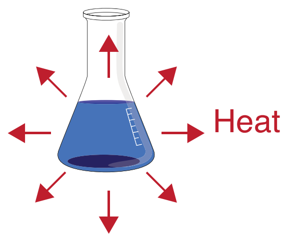
<td>
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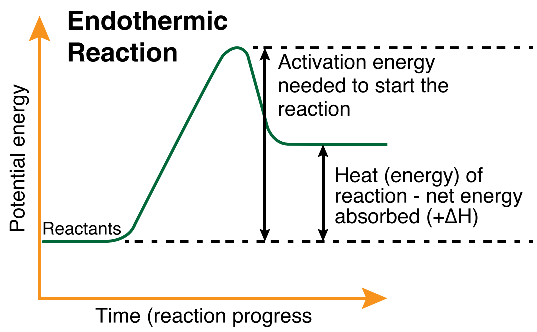
<td>
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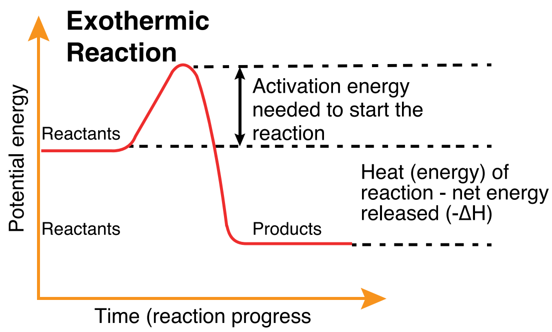
<td>
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Example

Glucose is oxidised to produce CO_2, H_2O and energy as shown in the equation:

$C_6H_{12}O_6(s) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(\ell)$ $\Delta H = -2\ 816$ kJ

(a) How much energy is released when 1.00 mol of glucose is oxidised?

$\Delta H = -2\ 816$ kJ per mol of glucose $\therefore$ 2 816 kJ is released.

(b) How much energy is released when 360.0 g of glucose is oxidised?

$$\begin{aligned} M(\text{glucose}) &= 6\,M(C) + 12\,M(H) + 6\,M(O) \\ &= 6(12.01) + 12(1.008) + 6(16.00) \\ &= 72.06 + 12.096 + 96.00 \\ &= 180.156 \text{ g mol}^{-1} \end{aligned}$$

$$n(\text{glucose}) = \frac{m}{M(\text{glucose})} = \frac{360.0}{180.156} = 1.998 \text{ mol}$$

Energy released

$= 2\ 816 \times 1.998$

$= 5\ 627$ kJ

Set 23: Exercises

1. Identify the following reactions as endothermic or exothermic (tick appropriate box)

Reaction	Endothermic	Exothermic
(a) combustion		
(b) respiration		
(c) melting		
(d) boiling		
(e) freezing/solidification		
(f) acid-base neutralisation		
(g) dissociation of sodium hydroxide in water		

2. Photosynthesis can be represented by the following reaction

$6\ CO_2(g) + 6\ H_2O(\ell) \rightarrow C_6H_{12}O_6(aq) + 6\ O_2(g) \quad \Delta H = +\ 2803\ kJ$

(a) How much energy is required to produce 1.00 mole of glucose?

(b) How much energy is required to convert 1.00 mole of carbon dioxide into glucose?

(c) How much energy is required to convert 45.0 g of water into glucose and oxygen?

(d) What mass of oxygen is produced when 1540.0 kJ of energy is consumed?

3. The pop test is used to confirm the presence of hydrogen. It is a highly exothermic reaction and can be represented by the following equation:

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(\ell) \quad \Delta H = -484\ kJ$$

(a) How much energy is produced when 1.00 mole of hydrogen is reacted with oxygen?

(b) How much energy is produced when 5.00 moles of water is formed?

(c) What mass of oxygen is required to produce 255 kJ of energy?

(d) Write the equation for the decomposition of water into hydrogen and oxygen and include the enthalpy change in the equation.

4. Methanol is used as a reactant in fuel cells. It can be produced using the following reaction:
$2\ H_2(g) + CO(g) \rightarrow CH_3OH(g) + 90\ kJ$
(a) Rewrite the equation showing the heat of reaction as a ΔH.

(b) How much energy will be produced when 100.0 g of hydrogen is consumed?

(c) What mass of carbon monoxide is required to produce 25.0 kJ of energy?

(d) What mass of methanol will be produced when 555 kJ of energy is produced?

5. The formation of photochemical smog occurs when chemicals such as ozone and nitrogen monoxide react in the presence of sunlight.
$O_3(g) + NO(g) \rightarrow O_2(g) + NO_2(g)$
(a) Ozone and oxygen are allotropes. What does this mean?

(b) This reaction is exothermic. Rewrite the equation to include 'heat' in the reaction.

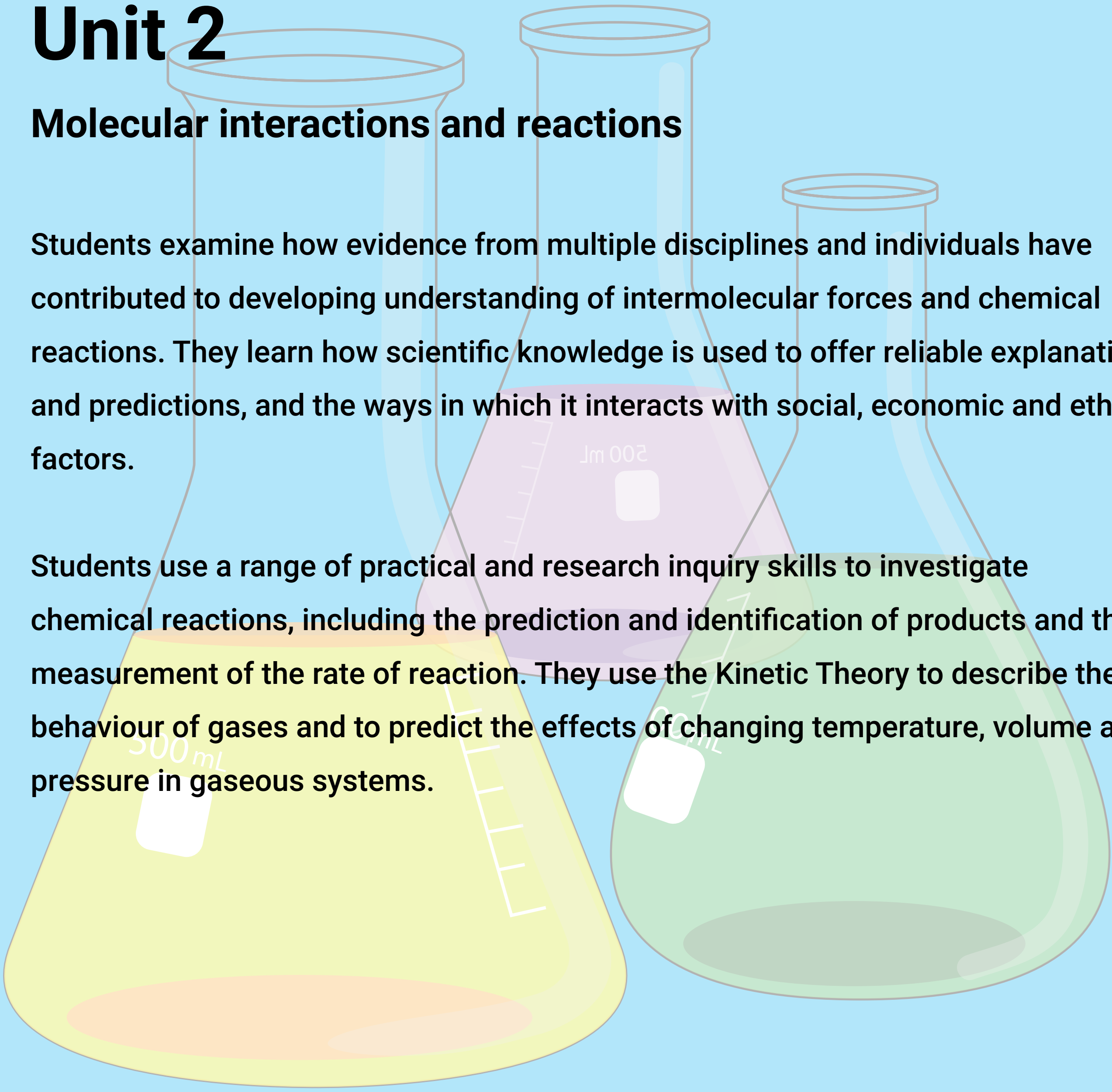
(c) Given that the activation energy is 210 kJ and the heat of reaction is -200 kJ, draw a reaction profile diagram for this reaction.

6. Separately list common examples of endothermic and exothermic reactions or processes in and around the home e.g. combustion, hot packs, change of phase. Select the reaction that has the most significance to you. Describe and explain the processes involved. Include an energy profile diagram and equations.

7. Resusable heat packs usually contain sodium acetate. Describe how these heat packs work and why they are recharged by heating in boiling water. In your answer include reaction profile diagrams and equations where appropriate.

Unit 2

Molecular interactions and reactions



Students examine how evidence from multiple disciplines and individuals have contributed to developing understanding of intermolecular forces and chemical reactions. They learn how scientific knowledge is used to offer reliable explanations and predictions, and the ways in which it interacts with social, economic and ethical factors.

Students use a range of practical and research inquiry skills to investigate chemical reactions, including the prediction and identification of products and the measurement of the rate of reaction. They use the Kinetic Theory to describe the behaviour of gases and to predict the effects of changing temperature, volume and pressure in gaseous systems.

Rates of chemical reactions

- Explain how the rate of a chemical reaction can be measured by the rate of formation of products or the depletion of reactants.

- Use the collision theory to explain and predict the effects of concentration, temperature, pressure, the presence of catalysts and surface area on the rate of chemical reactions.

- Define the activation energy as the minimum energy required for a chemical reaction to occur and is related to the strength and number of the existing chemical bonds; the magnitude of the activation energy influences the rate of a chemical reaction.

- Draw energy profile diagrams to represent the enthalpy changes and activation energy associated with a chemical reaction, including showing the transition state and catalysed and uncatalysed pathways.

- Describe the action of a catalysts, including enzymes and metal nanoparticles, on the rate of certain reactions by providing an alternative reaction pathway with a reduced activation energy, increasing the proportion of collisions that lead to a chemical change.

Set 24 Rates of reaction

The rates of chemical reactions can be altered in a number of ways, for example, by changing temperature, pressure (for systems involving gases), concentration, degree of sub-division or adding a catalyst.
The collision theory can be used to explain the observed impact of these changes on the rate of a reaction. The collision theory states that, in order for a successful reaction to occur, particles must collide with

(a) sufficient energy (the activation energy) and
(b) the appropriate orientation.

Figures 24.1 and 24.2 show energy profile diagrams for an exothermic and an endothermic reaction while figure 24.3 shows the distribution of molecular energies in a system.

Figure 24.1
Exothermic reaction

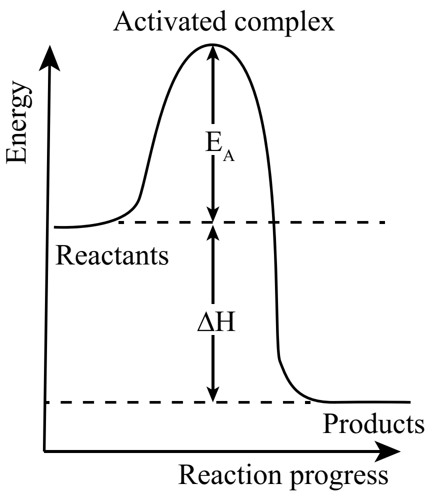


Figure 24.2
Endothermic reaction

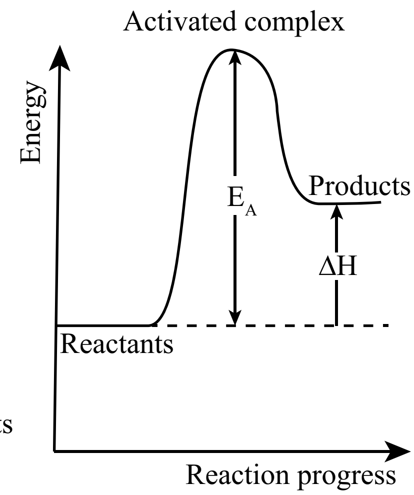
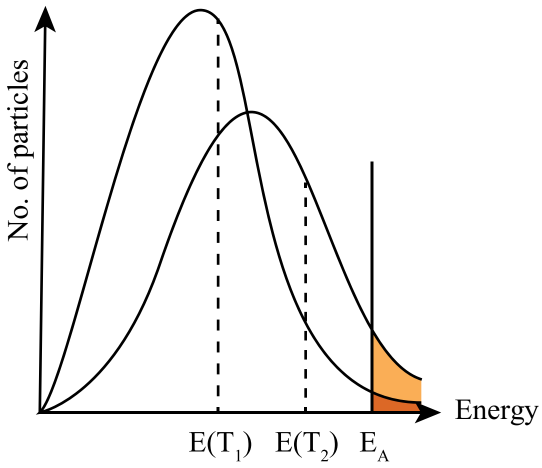


Figure 24.3
Molecular energy distribution



Exothermic reaction:
- Energy released to surroundings
- Surroundings warmer
- ΔH is negative (-ΔH)

Endothermic reaction:
- Energy absorbed from surroundings
- Surroundings cooler
- ΔH is positive (+ΔH)

E(T) = average energy at temperature T
$T_1 < T_2$
The proportion of molecules with energy equal to or greater than the activation energy, E_A (shaded areas) increases with an increase in temperature

Example

Explain why powdered calcium carbonate will react more quickly than the same mass of chips of calcium carbonate when added to dilute hydrochloric acid.

A powered mass of calcium carbonate will have a much greater surface area than chips. A greater surface area increases the frequency of collision of particles and increases the frequency of successful collisions, making the rate of reaction greater.

Set 24: Exercises

1. The following energy values were measured during a reaction.
   E (reactants) = 110 kJ
   E (products) = 130 kJ
   E_A = 30 kJ
   ΔH = +20 kJ

   (a) Draw and label an energy profile diagram for the reaction.

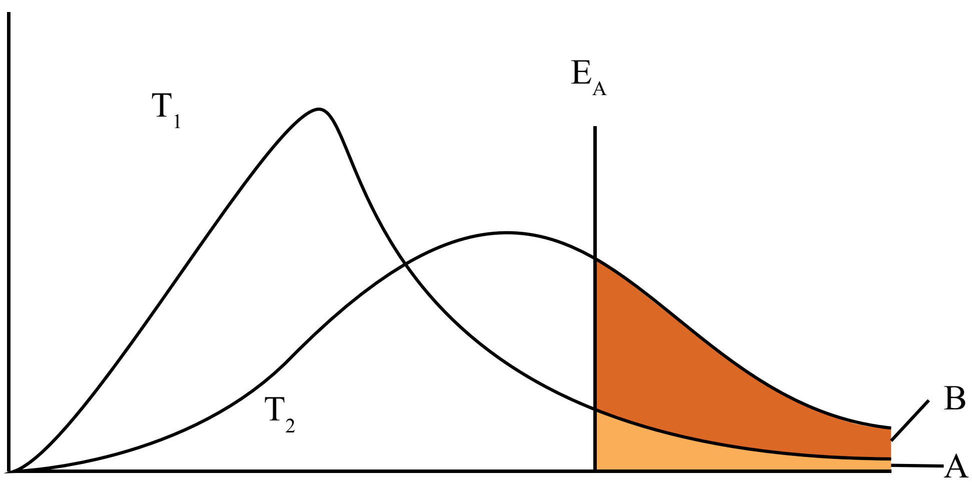
   (b) Is the reaction described endothermic or exothermic? Explain

   (c) This reaction is a reversible reaction. For the reverse reaction, state the
   (i) ΔH

   (ii) E_A

   (d) On your diagram for (a), draw in a different colour and label, the reaction pathway for the reaction if a catalyst was used.

2. The energy distribution diagram shows the distribution of kinetic energy for reacting particles at two different temperatures T_1 and T_2. Use this diagram to answer the following questions.



(a) Label the axes.

(b) Which is the higher temperature, T_1 or T_2? Explain

(c) Explain what the line, E_A represents.

(d) What does area A represent?

(e) What does area B represent?

3. Industrial processes, in order to be economic, require that products are formed as rapidly as possible at a reasonable cost. The production of ammonia through the Haber process involves an exothermic reaction between nitrogen gas and hydrogen gas to produce ammonia gas.

(a) Write a balanced equation for the reaction, including the heat of reaction.

(b) Explain how increasing the temperature will increase the rate of this reaction.

(c) Explain how increasing the pressure of the system will increase the rate of reaction.

(d) Explain how the use of a catalyst will increase the rate of reaction.

(e) Draw an energy profile diagram for the reaction.

4. A student was experimenting with ways of increasing the reaction rate between marble chips (calcium carbonate) and hydrochloric acid solutions.
(a) Write a balanced ionic equation for the reaction between calcium carbonate and hydrochloric acid.

(b) List three ways the student could increase the rate of reaction and explain how each change causes the increase.

(c) What impact would using acetic acid rather than hydrochloric acid have on the rate of the reaction?

5. When carbon monoxide reacts with oxygen to produce carbon dioxide, 566 kJ of energy is produced per mole of oxygen gas consumed.
(a) Write a balanced equation for the reaction, including the heat of reaction.

(b) Draw an energy profile diagram for this reaction given that the activation energy for the reaction is 250 kJ.

(c) Annotate the reaction profile diagram above to show the impact of adding a catalyst to this reaction.

(d) List two other ways of increasing the rate of this reaction and explain how they cause the increase.

6. Flammable liquids do not burn. It is the mixture of the substance's vapour and oxygen that actually ignites. Suggest why it is the vapour that burns and not the liquid itself.

7. List five examples where rates of reaction have been altered in processes occurring in and around the home. Apply the collision theory to describe and explain each of the examples you list.

8. Enzymes are catalysts found in biological systems. Describe how enzymes act as catalysts.

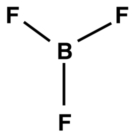
Intermolecular forces and gases

- Explain observable properties, including vapour pressure, melting point, boiling point and solubility, in terms of the nature and strength of intermolecular forces within a covalent molecular substance.
- Use the valence shell electron pair repulsion (VSEPR) theory and Lewis structure diagrams to explain, predict and draw the shapes of molecules.
- Use an understanding of molecular shape, symmetry, and differences in the electronegativity of atoms involved in the bond formation to explain the polarity of molecules.
- Use the shape and polarity of molecules to explain and predict the nature and strength of intermolecular forces, including London dispersion forces, dipole-dipole forces and hydrogen bonding.
- Use data from chromatography techniques, including thin layer chromatography (TLC), gas chromatography (GC), and high-performance liquid chromatography (HPLC), to determine the composition and purity of substances, and variation in strength of the interactions between atoms, molecules or ions in the mobile and stationary phases to explain the separate of components of mixtures.
- Use the Kinetic Theory to explain the behaviour of an ideal gas, including the qualitative relationships between pressure, temperature and volume.
- Use the mole concept to calculate the mass of substances and volume of gases (at standard temperature and pressure) involved in a chemical reaction.

Set 25: Molecular shape

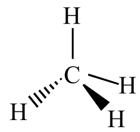
Polarity of a molecule is determined by the electron distribution in the bonds of the molecule and by the shape of a molecule. Shape is determined by the Valence Shell Electron Pair Repulsion Theory (VSEPR Theory). This states that electrons repel each other and will position themselves as far from each other as possible. The shapes you should be familiar with are shown in the diagram below.

Trigonal Planar



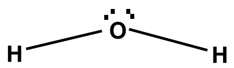
Bond angles for BF_3 are 120°

Tetrahedral



Bond angles for CH_4 are 109.5°

V-shaped

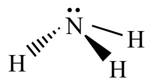


Bond angles for H_2O are 104.5°

Linear

Bond angles for CO_2 are 180°

Trigonal Pyramidal



Bond angles for NH_3 are 107°

There are two requirements for a molecule to be polar:

- The molecule should have one or more polar bonds and
- the molecule should not be symmetrical.

Bonds are polar if the elements bonded have differing electronegativities (electron attracting power).
In the molecule of ammonia (NH_3) in the diagram above, nitrogen and hydrogen have differing electronegativities so polar bonds are present and the molecule shape is asymmetrical. This makes the ammonia molecule polar. Methane (CH_4), on the other hand, has four polar bonds but the shape is symmetrical so the resulting molecule is non-polar.

Examples

When determining the shape and polarity of a molecule start by drawing an electron dot diagram.

1. **Ammonia (NH_3)**: Nitrogen has 5 valence electrons. Hydrogen has one valence electron.
The electron dot diagram is:

```
    ..
H : N : H
    ..
    H
```

Shape: As well as 3 bonding pairs of electrons there is a lone pair of electrons in the molecule and these also contribute to the shape of ammonia. The lone pair requires its share of space so the shape is **trigonal pyramidal** rather than trigonal planar.

Polarity: In the molecule of ammonia (NH_3), nitrogen and hydrogen have differing electronegativities so polar bonds are present. The molecule shape is trigonal pyramidal (asymmetrical). This makes the ammonia molecule **polar**.

2. **Methane (CH_4)**: Carbon has four valence electrons. Hydrogen has one valence electron. The carbon and hydrogen share electrons as shown in the electron dot diagram:

```
    H
    ..
H : C : H
    ..
    H
```

The molecule is symmetrical as the atoms space out evenly creating a **tetrahedral** shape. It has four polar bonds of equal strength and the shape is symmetrical so the resulting methane molecule is **non-polar.**

Set 25: Exercises

1. Explain why the electronegativity of iodine is much less than for fluorine.

2. Cℓ----------------------Cℓ
   Cℓ----------------------H
   Cℓ----------------------Na

   (a) On the diagram above show by drawing a cross, where the shared pair of electrons would be most likely to be found between the two elements that are bonded. Explain why they would be in that position.
   (b) What is the significance of the position of the electrons to the bond formed?

3. Which of the following species would have polar bonds?
   (a) I_2 (b) CO_2 (c) PH_3 (d) $CH_3C\ell$ (e) O_2

4. For the species in question 3, draw electron dot diagrams and determine the shape and the polarity of the molecules.

5. Draw diagrams to represent each of the five shapes of molecules discussed in this set. Which shapes are always going to be polar? Explain why.

6. Use examples to explain how it is possible to have a non-polar molecule with polar bonds.

7. For the following species, draw the electron dot diagram and determine its shape.

(a) SO_2

shape:

(b) NH_3

shape:

(c) H_2O

shape:

(d) SO_3

shape:

(e) NO_3^-

shape:

(f) SO_4^{2-}

shape:

(g) CO_3^{2-}

shape:

(h) H_2O_2

shape:

(i) C_2H_2

shape:

(j) $PC\ell_3$

shape:

Set 26: Intermolecular forces

Observable properties of covalent molecular substances can be explained by the type and strength of intermolecular forces present between the molecules.

London dispersion forces are present between all molecules and are caused by temporary dipoles created by transitory imbalances in the electron cloud around a molecule. These temporary dipoles then induce dipoles in neighbouring molecules. The strength of London dispersion forces are primarily determined by the number of electrons in the molecule, that is, the larger the number of electrons, the greater the strength of the London dispersion forces. The surface area of the molecules will also impact on the strength of the London dispersion forces. The greater the surface area, the greater the London dispersion forces. This explains why isomers, with the same number electrons, have different boiling points.

Dipole-dipole forces of attraction will occur between polar molecules in addition to the London dispersion forces. When comparing polar and non-polar molecules with similar numbers of electrons, differences in properties such as boiling point, can be explained by the presence of the additional dipole-dipole forces.

Hydrogen bonding is an extreme form of dipole-dipole force. Where a very polar covalent bond between hydrogen and a very electronegative element such as fluorine, oxygen or nitrogen, exists in a molecule, the electron density around the hydrogen will be very low. As a result, the hydrogen atom will have a relatively high positive charge. The low electron density also allows the hydrogen atom to approach an unbonded pair of electrons on a neighbouring molecule's oxygen, nitrogen or fluorine (which are all relatively negatively charged) more closely. This combination of close proximity and relatively large difference in charge causes the dipole-dipole force of attraction to be much stronger than usual. This explains the higher than expected melting and boiling points in substances such as water, ammonia and hydrogen fluoride.

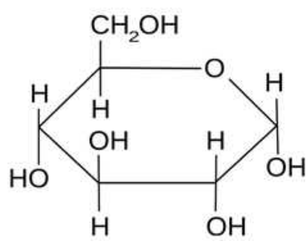


Figure 26.1: Glucose molecule

Set 26: Exercises

1. For the following substances, identify all the type(s) of intermolecular forces present between the molecules.

(a) octane (C_8H_{18})

(b) water (H_2O)

(c) ethanol (C_2H_5OH)

(d) pure nitric acid (HNO_3)

(e) glucose ($C_6H_{12}O_6$) (see Figure 26.1)

(f) iodine (I_2)

(g) hydrogen chloride ($HC\ell$)

(h) methane (CH_4)

2. Label 1-4 the following substances in order of increasing boiling point and justify your choices.

(a) HF HI $HC\ell$ HBr

(b) H_2O NH_3 O_2 $CH_3C\ell$

3. Given the following substances and their boiling points, label 1-4 the substances in order of increasing vapour pressure at room temperature. Justify your choices.

Substance	Boiling point (°C)
H_2O	100
HNO_3	83
Octane	125
Mercury	357

4. Explain how drawing an electron dot diagram can help you to predict the solubility of a substance.

5. Draw a dot diagram and predict the solubility of each of the following substances in water. Explain your decisions.

(a) methane (CH_4)

(c) ammonia (NH_3) (e)

iodine (I_2)

(b) ethanol (C_2H_5OH)

(d) hexane (C_6H_{14})

Set 27: Chromatography

All forms of chromatography are used to separate mixtures into their components. In each case there is a stationary phase and a mobile phase. The components of the mixture show a degree of preference for either the mobile or stationary phase. The stationary phase can be a solid (or a solid covered by a liquid) and the mobile phase can be a liquid or a gas.

Set 27: Exercises

Thin layer chromatography

1. As explained above, in all types of chromatography there is a mobile phase and a stationary phase. In Thin Layer Chromatography (TLC)
   a) what is the mobile phase usually made of?

   b) how is the stationary phase constructed?

2. A simple TLC experiment is to be set up using a plate and a mixture of dyes to be tested. The final result is shown in Figure 27.1.
   a) Write a "method" for this experiment.

b) Why was a pencil rather than a ball point pen used to mark the starting position?

   c) Why is the beaker covered with a lid?

   d) Calculate the Rf factors for the red and the blue dye.

   e) List two variables you controlled in this experiment

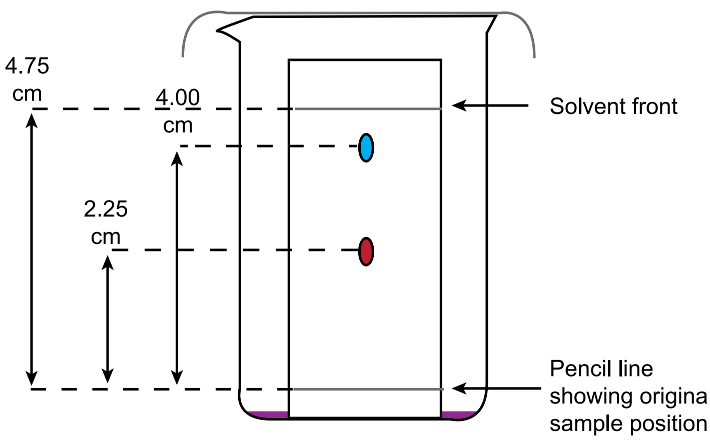


Figure 27.1: TLC example

Gas chromatography (GC)

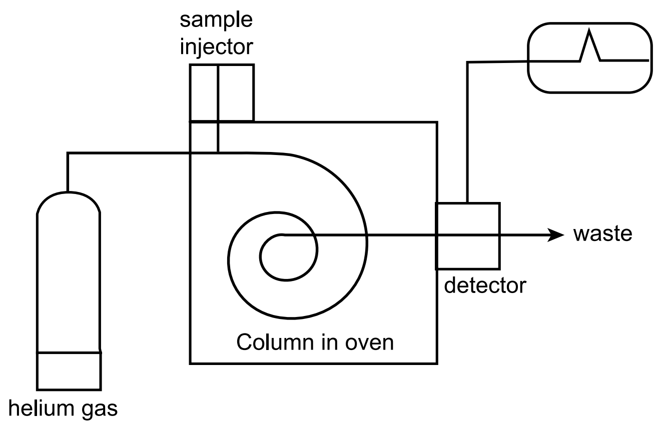


Figure 27.2: Gas chromatography

3. In gas chromatography (GC) the mobile phase is a gas (often helium) and the stationary phase is a liquid dispersed over a solid. The sample is either preferentially attracted to the gas or to the liquid on the solid material (column packing material). The time a sample takes to travel through the column is called its Retention Time – the more it is attracted to the liquid the longer its retention time.

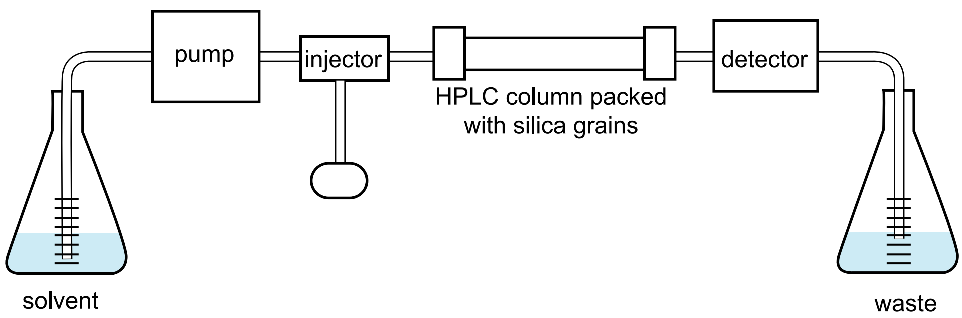
   a) Name the stationary phase and the mobile phase in gas chromatography.

   b) Explain why a sample might have a very short retention time

   c) If the temperature of the oven were raised why would the retention time decrease?

High-performance liquid chromatography (HPLC)

HPLC is a form of chromatography where, rather than allowing liquids to elute, or drip under gravity they are forced by a high pressure pump. Pressures of up to 40,000 kPa are often used. This means that the stationary phase can be very fine grained which gives an increased surface area and more interactions with the sample in the mobile phase. Hence it tends to produce better quality results.



4. As with GC the sample components are identified by their retention time. The retention time depends on the pump pressure, the size and chemical properties of the stationary phase, the solvent and its temperature.

   a) Explain clearly the advantages of using a high pressure pump to force the solvent through the stationary phase.

b) Consider the diagram above of the HPLC system. How will the retention time for a polar molecule differ from a less poplar molecule?

5. As in GC the retention time is calculated from when the sample is injected until it appears at the detector. List two factors that will determine the retention time.

6. List three examples of mixtures that are commonly separated using high performance liquid chromatography (HPLC) and explain why this is the most appropriate method of separation for each mixture.

7. List three examples of mixtures that are common separated using high performance liquid chromatography (HPLC) and explain why this is the most appropriate method of separation for each mixture.

Set 28: Solutions

1. A student recorded the solubility of sugar at different temperatures.

Temperature °C	Solubility (g/100 g water)
0	179
20	204
40	238
60	287
80	362
100	487

(a) Draw a solubility curve for sugar with temperature on the horizontal axis.

(b) Describe the pattern shown by the graph.

(c) Determine the solubility of sugar at:

(i) 30 °C

(ii) 70 °C

(d) Use the information in the student's table to identify an unsaturated, saturated and a super saturated solution of sugar at 20 °C.

(e) Classify the following sugar solutions as unsaturated, saturated or super-saturated.

(i) 200 g of sugar dissolved in 100 g of water at 40 °C

(ii) 200 g of sugar dissolved in 50 g of water at 60 °C

(iii) 50 g of sugar dissolved in 20 g of water at 100 °C

2. A 2.000 kg (approximately 2.0 L) sample of recycled water on a property was examined for total dissolved solids (TDS). The purpose was to determine if, after treatment, its TDS concentration was at an acceptable level for use in irrigation. The sample was evaporated to dryness and the total dissolved solids remaining weighed 3.45 g.
Parts per million can be calculated as follows:

$$\text{ppm} = \frac{\text{mg of solute}}{\text{kg of solution}}$$

Water up to 2,500 ppm can be used for irrigation. Use the information in Figure 28.1 to classify the salinity of the sample and state if it is suitable for irrigation.

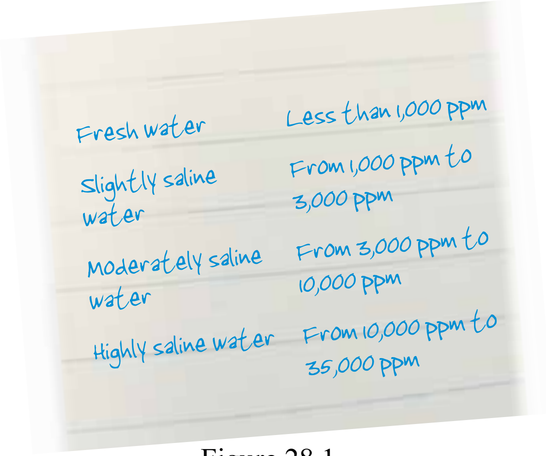


Figure 28.1

3. Seawater contains about 35,000 ppm of salt. How many grams of salt would be obtained if 1.000 kg (approximately 1 L) of seawater were evaporated to dryness?

4. Describe the chemical tests that would allow you to distinguish between solid samples of the following barium salts: $BaCO_3$, $Ba(NO_3)_2$, $BaC\ell_2$ and $BaSO_4$.

5. Use the solubility table (Appendix) to identify and describe any precipitate that forms when the following solutions are mixed.
(a) $NaC\ell$ and $AgNO_3$

(b) $Pb(NO_3)_2$ and KI

(c) K_2SO_4 and $Ba(OH)_2$

(d) $CuSO_4$ and NaOH

(e) $(NH_4)_3PO_4$ and $FeC\ell_2$

6. Write ionic equations for any precipitation reactions occurring in question 5.

7. Classify the following as strong, weak or non-electrolytes:

tap water

sea water

sugar solution

copper sulfate solution

hydrochloric acid solution

Set 29: Solution concentrations

Many reactions occur in aqueous solutions. Stoichiometry in solutions uses concentration in calculations. Concentration is defined as the amount of solute per unit volume of solution. The amount of solute is usually expressed in moles. The relationship can be expressed as:

$$\mathbf{n = cV}$$

where **n** is the number of moles of solute (mol),
c is the concentration of solute in moles per litre (mol L^{-1}), and
V is the volume of the solution in litres (L).

Note: Solution concentrations can be expressed using 'c' or '[ ]' for example, c(NaCℓ) or [NaCℓ] both refer to the concentration of NaCℓ in solution, expressed in mol L^{-1}.

Examples

1. Calculate the number of moles of sodium chloride in 15.0 mL of a 2.50 mol L^{-1} NaCℓ solution.

$$n = cV$$
$$= (2.50)\,(15.0 \times 10^{-3})$$
$$n = 3.75 \times 10^{-2} \text{ mol of NaC}\ell$$

2. Calculate the concentration in mol L^{-1} of an iron(II) sulfate-7-water solution which contains 5.56 g of the salt dissolved in distilled water and made up to 750.0 mL of solution.

(a) Calculate the number of moles of solute:

$$n = \frac{m}{M}$$
$$= \frac{5.56}{278.0} \qquad M(FeSO_4{\cdot}7H_2O) = 278.0 \text{ g mol}^{-1}$$
$$= 2.00 \times 10^{-2} \text{ mol}$$

(b) Calculate the concentration of $FeSO_4{\cdot}7H_2O$ solution.

$$c = \frac{n}{V}$$
$$= \frac{2.00 \times 10^{-2}}{750 \times 10^{-3}}$$
$$c = 2.67 \times 10^{-2} \text{ moI L}^{-1}$$

Ionic substances and some covalent substances, such as strong acids, produce almost 100% ions when dissolved in H_2O. For example

$$NaC\ell(s) \rightarrow Na^+(aq) + C\ell^-(aq)$$

$$HC\ell(g) \rightarrow H^+(aq) + C\ell^-(aq) \quad \text{or} \quad HC\ell(g) + H_2O(\ell) \rightarrow H_3O^+(aq) + C\ell^-(aq)$$

Knowing the concentration of the solute and its formula, the concentration of ions formed in solution can be calculated.

3. In 25.0 mL of a 0.200 mol L^{-1} solution of aluminium sulfate, calculate:
   (a) the $[A\ell^{3+}]$ (c) $n(A\ell^{3+})$
   (b) the $[SO_4^{2-}]$ (d) $n(SO_4^{2-})$

   (i) Write a balanced ionic equation:
   $A\ell_2(SO_4)_3(s) \rightarrow 2A\ell^{3+}(aq) + 3SO_4^{2-}(aq)$

   (ii) From the balanced ionic equation relate the unknowns to the known:

(a) $c(A\ell^{3+}) = 2 \times c(A\ell_2(SO_4)_3)$
$= (2)\ (0.200)$
$c(A\ell^{3+}) = 0.400 \text{ mol L}^{-1}$

(b) $c(SO_4^{2-}) = 3 \times c(A\ell_2(SO_4)_3)$
$= (3)\ (0.200)$
$c(SO_4^{2-}) = 0.600 \text{ mol L}^{-1}$

(c)(i) Calculate the number of moles of $A\ell_2(SO_4)_3$
$n(A\ell_2(SO_4)_3) = c(A\ell_2(SO_4)_3)\ V$
$= (0.200)\ (25.0 \times 10^{-3})$
$= 5.00 \times 10^{-3} \text{ mol}$

(ii) Relate the unknown numbers of moles to the known number of moles of $A\ell_2(SO_4)_3$:
$n(A\ell^{3+}) = 2 \times n\ (A\ell_2(SO_4)_3)$
$= (2)\ (5.00 \times 10^{-3})$
$n(A\ell^{3+}) = 1.00 \times 10^{-2} \text{ mol}$

(d) $n(SO_4^{2-}) = 3 \text{ x } n\ (A\ell_2(SO_4)_3)$
$= (3)\ (5.00 \times 10^{-3})$
$n(SO_4^{2-}) = 1.50 \times 10^{-2} \text{ mol}$

Set 29: Exercises

1. Calculate the concentrations of the following solutions:
   (a) 0.223 moles of copper(II) sulfate in 125.0 mL of solution

   (b) 1.17 moles of sodium chloride in 2.05 L of solution

   (c) 0.0335 moles of silver nitrate in 250.0 mL of solution

2. Calculate the number of moles of:
   (a) potassium nitrate in 105 mL of 2.55 mol L^{-1} potassium nitrate solution

   (b) sodium carbonate in 2.50 L of 0.112 mol L^{-1} of sodium carbonate solution

   (c) potassium permanganate in 660 mL of 0.230 mol L^{-1} of potassium permanganate solution

3. Calculate the mass of solute that must be used in order to prepare each of the solutions listed below:
   (a) 630.0 mL of 1.26 mol L^{-1} $KC\ell$ solution from $KC\ell$

(b) 250.0 mL of 0.265 mol L^{-1} Na_2CO_3 solution from $Na_2CO_3{\cdot}10H_2O$

(c) 1.05 L of 0.420 mol L^{-1} $H_2C_2O_4$ solution from $H_2C_2O_4{\cdot}2H_2O$

4. Calculate the number of moles of:
   (a) chloride ions in 25.0 mL of 0.200 mol L^{-1} barium chloride solution

   (b) sulfate ions in 550 mL of 2.56 mol L^{-1} sodium sulfate solution

   (c) nitrate ions in 2.20 L of 2.02×10^{-3} mol L^{-1} lead(II) nitrate solution

5. In a solution prepared by dissolving 10.0 g of potassium carbonate in 220.0 mL of distilled water, determine:
   (a) the concentration, in mol L^{-1}, of K^+ (aq)

   (b) the concentration, in mol L^{-1}, of CO_3^{2-} (aq)

6. Calculate the volume of 4.0 mol L^{-1} nitric acid solution required to prepare 250.0 mL of 0.250 mol L^{-1} solution.

7. Calculate the concentration in mol L^{-1} of ammonium ions in a solution prepared by mixing 360.0 mL of 0.250 mol L^{-1} ammonium sulfate solution with 675.0 mL of 1.20 mol L^{-1} ammonium nitrate solution. Assume solution volumes are additive.

8. A laboratory assistant has a solution of 0.120 mol L^{-1} $KMnO_4$. What volume of this solution must be diluted to produce 500.0 mL of a 0.025 mol L^{-1} solution?

9. What volume of water must be added to 150.0 mL of 1.10 mol L^{-1} sulfuric acid solution to prepare a 0.210 mol L^{-1} solution?

10. 25.6 g of anhydrous sodium carbonate is dissolved in 200.0 mL of water.
(a) Calculate the concentration in mol L^{-1} of this solution. Assume no volume change.

(b) A 20.0 mL sample of this solution was placed in a 150.0 mL flask and 80.0 mL of water added. Determine the concentration of sodium ions in this diluted solution.

Set 30: Mixtures

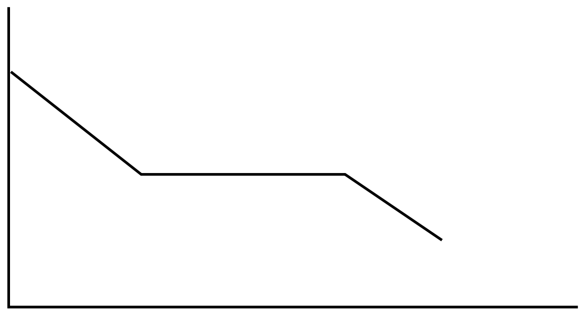
Physical separation techniques are used to purify components or to determine the composition of a mixture. Separation techniques rely on the individual substances having different physical properties. These properties include solubility, level of attraction to a stationary phase, particle size, density, melting and boiling point.

- Filtration relies on different particle size
- Distillation relies on different boiling points.
- Chromatography relies on the solubility of substances in a particular solvent and their levels of attraction to a stationary phase.

Set 30: Exercises

1. Describe using examples the differences between a homogeneous and a heterogeneous mixture.

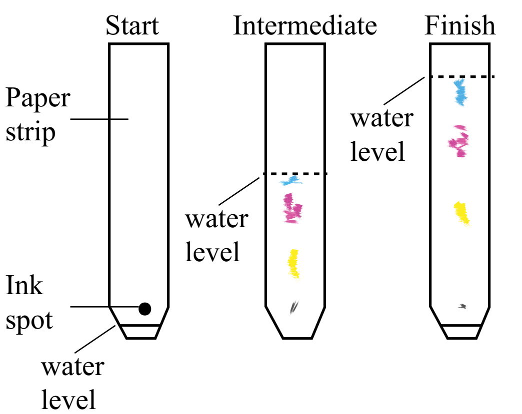
2. The graph (right) shows the cooling curve of ethanol.
   (a) Place the following labels on the graph: time, temperature, and the time during which freezing occurs.
   (b) Research the freezing temperature of "pure" ethanol and add this to the graph. Describe the meaning of pure using ethanol as an example.



   (c) Ethanol produced by distillation is not pure. State how this mixture should be described and name the other substance present.

3. The diagram (right) shows the chromatographic separation of the components of a sample of black ink.
   (a) Describe how paper chromatography can be used to separate the different dyes in ink.
   (b) Use the final chromatogram to describe the ink mixture.

Chromatographic separation of black ink



   (c) Which of the dyes show the least adhesion to the paper and the most solubility in water. Explain.

(d) Each separated dye can be assigned an "Rf value". Describe what is meant by an "Rf value" and how it can be determined.

(e) State which colour dye has the largest Rf value.

4. Describe how you could separate and collect the first substance (in bold) from each mixture:

(a) **sucrose** and sand

(b) **sand** and sodium chloride

(c) **water** and copper sulfate

(d) **pink** and blue dyes in purple text ink

(e) **octane** from crude oil

5. Explain the difference between simple distillation used to obtain water from salty water and fractional distillation to separate the components of crude oil.

Set 31: Kinetic theory

1. Aerosol cans, including deodorant sprays, have labels warning against exposing used cans to excessive heat. Record the warning label of an aerosol can that you have at home. Explain the reasons behind this warning.

2. A car or bicycle tyre becomes hotter during use. Assuming the volume of the tyre remains constant how does the pressure of air inside the tyre change? Explain in terms of the Kinetic Theory of matter.

3. A helium balloon, when released, rises up into the air and disappears out of sight.
   (a) Why does the helium balloon rise?
   (b) Describe the pressure changes both in and around a helium balloon as it rises. Use diagrams to help your explanation.
   (c) Why will the balloon eventually burst?

4. Using the Kinetic Theory explain why the boiling point of a liquid is a sharp definite temperature. Why does water boil at a temperature lower than 100 °C at 2 000 m above sea level?

5. Explain why liquids exert a vapour pressure. Why is the boiling point of a solution higher than the boiling point of its solvent?

6. Some cooks will tell you that adding a tablespoon of salt to a pot of water will make the boiling point of the liquid greater than 100 °C, and so your pasta is cooked faster than it would be in unsalted water. Investigate the validity of this statement, and suggest other methods that could be used to decrease the cooking time of pasta.

7. As a liquid is heated its vapour pressure increases. When the vapour pressure reaches atmospheric pressure boiling occurs. Knowing this, compare and contrast the boiling point of water in Perth (at sea level) to that at Kalgoorlie (about 500 m). What impact will this have on cooking in boiling water?

8. **Healthy water:**
   (a) Why is dissolved oxygen in rivers and lakes important?

   (b) How does the solubility of oxygen gas in water change when the water temperature is increased?

   (c) What is thermal pollution?

(d) Discuss the impact of thermal pollution on the concentration of dissolved gases in natural waterways.

9. **SCUBA Diving:**
SCUBA diving, a popular underwater pastime, depends on a supply of compressed air to allow breathing under water. Compressed air is air kept under a pressure greater than atmospheric pressure.
(a) Why is the air in SCUBA tanks kept under pressure?

(b) List the gases added to the breathing tanks of a SCUBA diver and explain why they are used.

(c) Nitrogen is not included in the air in scuba tanks for deep and prolonged dives because at higher pressures a large quantity of nitrogen dissolves in the blood. The increased solubility of nitrogen in the blood at high pressures is responsible for the two conditions: *rapture of the deep* (nitrogen narcosis) and the *bends*. Explain these conditions.

Set 32: Gas volumes

From our understanding of Kinetic Theory we know that gas volumes are independent of the identity of the gas. Gas volume is dependent upon the number of particles, its pressure and temperature. This means that, for an ideal gas, one mole of any gas will occupy the same volume under the same pressure and temperature conditions. This enables us to develop a relationship between volume and number of moles of a gas.

Molar volume of gases

The molar volume of a gas is the volume occupied by 1 mole of the gas. The molar volume for an ideal gas, and for some real gases, is 22.71 L at S.T.P. (100 kPa, 0 °C)

The relationship between the number of moles (n) of a gas and its volume (V) in litres at S.T.P. is

$$n = \frac{V(\text{in litres at S.T.P.})}{22.71}$$

Examples

1. Calculate the volume occupied by 2.75 moles of carbon monoxide at S.T.P.

$$n = \frac{V}{22.71}$$

$$V = n \times 22.71$$

$$= (2.75)\,(22.71)$$

$$V = 62.5 \text{ L}$$

2. What is the volume of 9.60 g of oxygen at S.T.P.?

(a) Calculate the number of moles.

$$n(O_2) = \frac{9.60}{32.0} \qquad M(O_2) = 32.0 \text{ g mol}^{-1}$$

$$= 0.300 \text{ mol}$$

(b) Find the volume at S.T.P

$$n = \frac{V}{22.71}$$

$$V = n \times 22.71$$

$$= (0.300)\,(22.71)$$

$$V = 6.81 \text{ L at S.T.P.}$$

Set 32: Exercises

1. Calculate the volume occupied by:

(a) 6.50 moles of carbon dioxide at S.T.P.

(b) 0.850 moles of hydrogen at S.T.P.

2. Calculate the number of moles of:

(a) methane (CH_4) in 4.50 L of the gas at S.T.P.

(b) oxygen in 25.0 mL of the gas at S.T.P.

3. Calculate the volume of
   (a) 100 g of propane (C_3H_8) at S.T.P.

   (b) carbon dioxide gas at S.T.P. produced when a 1.00 kg block of dry ice, which is solid CO_2, sublimes.

4. Calculate the volume occupied by the following masses of gases at S.T.P.
   (a) 1.34 g of methane (CH_4)

   (b) 2.00 g of ethane (C_2H_6)

   (c) 58.6 g of ammonia (NH_3)

   (d) 0.566 g of sulfur dioxide (SO_2)

5. 4.18 g of a gas occupies 1.00 L at S.T.P. Calculate the molar mass of the gas.

6. 3.49 g of a gas occupies 1.00 L at S.T.P. Calculate the molar mass of the gas.

7. What mass of oxygen gas will occupy 1.00 L at S.T.P.?

8. What is the mass of the following?
   (a) 2.55 L of sulfur trioxide (SO_3)

   (b) 89.5 L of nitrogen (N_2)

   (c) 0.00253 L of argon

   (d) 41.2 L of chlorine ($C\ell_2$)

Set 33: Stoichiometry and gas volumes

Determine the volume of hydrogen (measured at S.T.P.) that is produced when 45.6 g of zinc metal is added to excess hydrochloric acid.

(i) Write an equation for the reaction

$$Zn(s) + 2\ HC\ell(aq) \rightarrow ZnC\ell_2(aq) + H_2(g)$$

$$Zn(s) + 2\ H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$$

(ii) Calculate the number of moles of zinc

$$n(Zn) = \frac{m}{M}$$

$$= \frac{45.6}{65.38}$$

$$= 0.697 \text{ mol}$$

(iii) From the equation determine the number of moles of hydrogen

$$n(H_2) = n(Zn)$$

$$n(H_2) = 0.697 \text{ mol}$$

(iv) Calculate the volume of hydrogen

$$V = n \times 22.71$$

$$= 0.697 \times 22.71$$

$$= 15.8 \text{ L}$$

Set 33: Exercises

1. Calculate the volumes of gas at S.T.P. produced in each of the following reactions:

(a) When 0.250 moles of $CaCO_3$ is treated with excess hydrochloric acid:

$$CaCO_3(s) + 2\ HC\ell(aq) \rightarrow CaC\ell_2(aq) + H_2O(\ell) + CO_2(g)$$

or $$CaCO_3(s) + 2\ H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(\ell) + CO_2(g)$$

(b) When 0.150 moles of potassium permanganate reacts with excess concentrated hydrochloric acid according to the equation

$$2\ MnO_4^-(aq) + 10\ C\ell^-(aq) + 16\ H^+(aq) \rightarrow 2\ Mn^{2+}(aq) + 5\ O_2(g) + 8\ H_2O(\ell)$$

(c) When 0.0300 moles of hydrogen peroxide disproportionates (reacts with itself) according to the equation

$$2\ H_2O_2(aq) \rightarrow O_2(g) + 2\ H_2O(\ell)$$

2. If 4.60 g of sodium reacts with water, calculate the volume of hydrogen produced at S.T.P.:

$$2\ Na(s) + 2\ H_2O(\ell) \rightarrow 2\ NaOH(aq) + H_2(g)$$

or $$2\ Na(s) + 2\ H_2O(\ell) \rightarrow 2\ Na^+(aq) + 2\ OH^-(aq) + H_2(g)$$

3. Calculate the volume of ammonia produced at S.T.P. when 22.8 g of ammonium sulfate is heated with excess moist calcium hydroxide:

$$(NH_4)_2SO_4(s) + Ca(OH)_2(s) \rightarrow CaSO_4(s) + 2\ H_2O(\ell) + 2\ NH_3(g)$$

4. When 25.9 g of impure limestone was treated with excess hydrochloric acid, 5.6 L of carbon dioxide was produced at S.T.P.:

$$CaCO_3(s) + 2\ HC\ell(aq) \rightarrow CaC\ell_2(aq) + H_2O(\ell) + CO_2(g)$$

or $$CaCO_3(s) + 2\ H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(\ell) + CO_2(g)$$

Calculate the percentage purity of the limestone.

5. A 15.3 g piece of steel, containing only iron and carbon, was treated with excess hot acid to form 6.03 L of hydrogen at S.T.P.:

$$Fe(s) + 2\ H^+(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$$

Calculate the percentage of iron in the steel.

6. The reactions involved in the manufacture of nitric acid can be represented as follows:

$$N_2 + 3\ H_2 \rightarrow 2\ NH_3$$
$$4\ NH_3 + 5\ O_2 \rightarrow 4\ NO + 6\ H_2O$$
$$2\ NO + O_2 \rightarrow 2\ NO_2$$
$$4\ NO_2 + 2\ H_2O + O_2 \rightarrow 4\ HNO_3$$

Calculate:

(a) the mass of nitrogen, and

(b) the volume of nitrogen at S.T.P. required to manufacture 25.0 kg of nitric acid.

7. Methane (CH_4) burns in oxygen to form carbon dioxide and water. Write a balanced equation for the process. If 2.50 kg of water is produced at S.T.P., calculate the volumes of:

(a) carbon dioxide produced

(b) methane consumed at the same temperature and pressure.

8. Sodium cyanide, which is used in the extraction of gold, can be made from sodium metal, carbon and ammonia:

$$2\ Na(s) + 2\ C(s) + 2\ NH_3(g) \rightarrow 2\ NaCN(s) + 3\ H_2(g)$$

Calculate the mass of sodium cyanide that would be formed by reacting 0.500 tonne of sodium metal with 762 kL of ammonia gas at S.T.P. in the presence of excess carbon.

9. It has been suggested that methane gas be used as an alternative fuel to petrol (assume C_8H_{18}) in motor vehicles, as it produces less of the greenhouse gas, CO_2 (*g*), for the same energy output.

(a) Write separate equations for the combustion of each fuel, methane (CH_4) and petrol (C_8H_{18}).

(b) The combustion of 206 g of methane produces as much energy as 239 g of petrol. Determine the volume of CO_2 (*g*) produced at STP, for each mass of fuel.

(c) Determine the percentage change in the volume of CO_2 (*g*) emission when methane is used as an alternative fuel to petrol.

(d) Using your results from part (c), does methane produce less CO_2 greenhouse gas emission as claimed?

10. A 0.7941 g sample of cast iron was heated in oxygen to convert the carbon it contained to carbon dioxide. The carbon dioxide produced was absorbed in potassium hydroxide solution to form potassium hydrogencarbonate.

$$C(s) + O_2(g) \rightarrow CO_2\ (g)$$
$$CO_2(g) + KOH(aq) \rightarrow KHCO_3(aq)$$

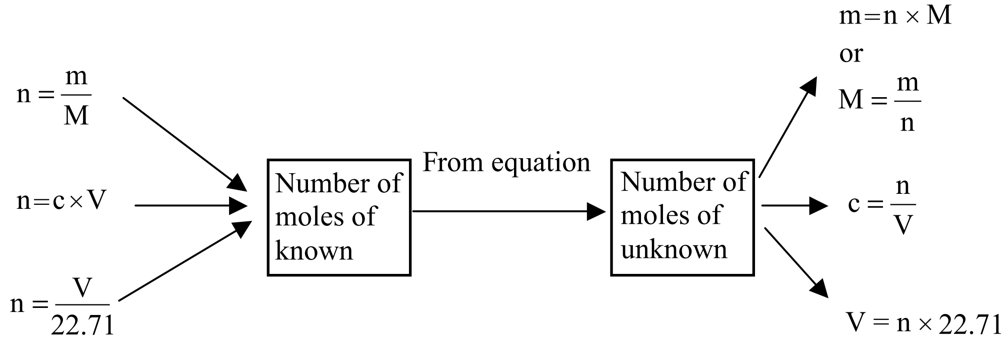
The potassium hydroxide solution increased in mass by 0.0732 g.
Calculate:
(a) The number of moles of carbon dioxide gas that dissolved in the potassium

(b) The volume of carbon dioxide produced at STP.

(c) The percentage of carbon in the cast iron.

Set 34: Reacting masses and gaseous and solution volumes

Using the stoichiometric relationship between the numbers of moles of reactants and products in a chemical reaction, a range of calculations can be carried out involving the masses, gaseous volumes and solution volumes of reactants and products. These calculations can be represented schematically as follows:



Set 34: Exercises

1. 25.0 mL of 0.0227 mol L^{-1} silver nitrate solution is added to excess sodium chloride solution.
   (a) Write an equation for the reaction.

   (b) Calculate the mass of the precipitate formed.

2. 47.3 L of butane (C_4H_{10}) measured at S.T.P. is burned in excess oxygen.
   (a) Write an equation for the reaction.

   (b) What volume of carbon dioxide (measured at S.T.P.) is produced?

   (c) What mass of water is produced?

3. 1.34 g of solid anhydrous sodium carbonate is added to excess 0.125 mol L^{-1} hydrochloric acid solution.
   (a) Write an equation for the reaction.

   (b) What volume of CO_2 (measured at S.T.P.) is produced?

   (c) What volume of hydrochloric acid would be required to react completely with the sodium carbonate?

4. Magnesium reacts with dilute hydrochloric acid according to the equation

   $$Mg(s) + 2\,H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$$

   If 0.720 g of magnesium is treated with 0.950 mol L^{-1} hydrochloric acid, calculate:
   (a) the volume of hydrochloric acid needed to react with all the magnesium

   (b) the volume of hydrogen gas formed at S.T.P.

(c) the mass of magnesium chloride which would be produced in solution

5. Hydrogen gas is produced by adding sulfuric acid to aluminium according to the equation

$$2\ A\ell(s) + 6\ H^+(aq) \rightarrow 2\ A\ell^{3+}(aq) + 3\ H_2(g)$$

If 19.6 L of hydrogen is produced at S.T.P., calculate:

(a) the volume of 6.00 mol L^{-1} sulfuric acid required

(b) the mass of aluminium consumed

6. (a) What volume of 0.260 mol L^{-1} sodium iodide is needed to precipitate all the lead ions in 25.0 mL of 0.212 mol L^{-1} lead(II) nitrate?

(b) Calculate the mass of lead(II) iodide precipitated.

7. Carbon dioxide can be produced by reacting calcium carbonate with hydrochloric acid. If 10.1 g of calcium carbonate is treated with 1.07 mol L^{-1} hydrochloric acid, calculate:

(a) the volume of $HC\ell$ needed to completely react with the $CaCO_3$

(b) the volume of CO_2 produced at S.T.P.

(c) the concentration of calcium chloride in the final solution, assuming no change in the volume of solution

8. A calcite crystal ($CaCO_3$) with a mass of 5.60 g is placed in a beaker with 20.0 mL of dilute hydrochloric acid. The solution is warmed gently until the reaction is complete. The remaining calcite is removed and dried and found to have a mass of 5.09 g. Calculate the concentration of the hydrochloric acid.

9. An excess of hydrochloric acid was added to a sample of silver carbonate. After the reaction the silver chloride residue was washed and dried and found to have a mass of 5.74 g:

$$Ag_2CO_3(s) + 2\ HC\ell(aq) \rightarrow 2\ AgC\ell(s) + H_2O(\ell) + CO_2(g)$$

Calculate

(a) the mass of the silver carbonate sample

(b) the number of moles of the acid consumed

(c) the volume of carbon dioxide gas, measured at S.T.P., given off during the reaction

Aqueous solutions and acidity

- Use molecular shape and hydrogen bonding between molecules to explain the unique physical properties of water, including melting point, boiling point, density in solid and liquid phases and surface tension.

- Classify solutions as saturated, unsaturated or supersaturated.

- Define the concentration of a solution as the quantity of solute dissolved in a quantity of solution and can be represented in a variety of ways, including by the number of moles of the solute per litre of solution ($mol\ L^{-1}$) and the mass of the solute per litre of solution ($g\ L^{-1}$) or parts per million (ppm)

- Identify by observing the colour of the solution and observing various chemical reactions, including precipitation and acid-base reactions the presence of specific ions in solutions.

- Explain the solubility of substances in water, including ionic and polar and non-polar molecular substances, by applying an understanding of the intermolecular forces, including ion-dipole interactions between species in the substances and water molecules, and that solubility is also affected by changes in temperature.

- Use the Arrhenius model to explain the behaviour of strong and weak acids and bases in aqueous solutions.

- Classify aqueous solutions as acidic, basic or neutral using indicator colour and the pH scale.

- Define pH as a measure of the acidity of solutions and state if it is dependent on the concentration of hydrogen ions in the solution.

- Predict products and observations for the reactions of acids and bases, including reactions of acids with bases, metals and carbonates and the reactions of bases with acids and ammonium salts.

- Write ionic equations to represent the reacting species and products in acid and base reactions.

- Use the mole concept to calculate the mass of solute, solution concentrations and volumes involved in a chemical reaction.

Set 35: Ionic equations

Ionic equations are written for reactions that take place in solution. They represent only the species that are involved in the reaction. The following rules apply to the writing of ionic equations.

1. Strong electrolytes are written in ionic form.
2. Weak electrolytes are written in molecular form.
3. Non-electrolytes are written in molecular form.
4. Insoluble substances are written as the formulae of the substances.
5. Gases are written in molecular form (except for noble gases).
6. Ionic equations should not include species that do not change (spectator ions).
7. Equations must be balanced in terms of atoms and electrical charge.

Knowledge of solubility rules is invaluable when writing ionic equations.
See the table solubility rules on page 207.

Examples

1. Write a balanced ionic equation for the dissolution of solid $NaC\ell$ in water. $NaC\ell$ is a solid, which, when dissolved, dissociates into ions as follows:

   $NaC\ell(s) \rightarrow Na^+(aq) + C\ell^-(aq)$

2. Write a balanced ionic equation showing what reaction occurs when dilute hydrochloric acid is added to solid calcium carbonate. Hydrochloric acid contains H^+ and $C\ell^-$ ions since it is a strong electrolyte. The products of adding an acid to a metal carbonate are a salt, water and carbon dioxide.
   (a) Write a balanced formula equation (including state symbols is expected):

   $CaCO_3(s) + 2\ HC\ell(aq) \rightarrow CaC\ell_2(aq) + H_2O(\ell) + CO_2(g)$

   (b) Use the rules for writing ionic equations to write relevant species as ions:

   $CaCO_3(s) + 2\ H^+(aq) + 2\ C\ell^-(aq) \rightarrow Ca^{2+}(aq) + 2\ C\ell^-(aq) + H_2O(\ell) + CO_2(g)$

   Note: Calcium chloride is a soluble salt.

   (c) Omit spectator ions and write the final balanced ionic equation:

   $CaCO_3(s) + 2\ H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(\ell) + CO_2(g)$

3. Write a balanced ionic equation showing the products formed when a solution of sodium iodide is added to a solution of lead(II) nitrate.
   (a) Write a balanced formula equation:

   $2\ NaI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2\ NaNO_3(aq)$

   Note: Lead(II) iodide is an insoluble salt whereas sodium nitrate is a soluble salt.

   (b) Write relevant species as ions:

   $2\ Na^+(aq) + 2\ I^-(aq) + Pb^{2+}(aq) + 2\ NO_3^-(aq) \rightarrow PbI_2(s) + 2\ Na^+(aq) + 2\ NO_3^-(aq)$

   (c) Omit spectator ions and write the balanced ionic equation:

   $2\ I^-(aq) + Pb^{2+}(aq) \rightarrow PbI_2(s)$

Set 35: Exercises

Write balanced ionic equations for the following reactions:

1. The dissolving of potassium chloride.

2. The dissolving of barium nitrate.

3. The dissolving of sodium hydroxide.

4. Dilute nitric acid added to a solution of potassium hydroxide.

5. Zinc oxide powder added to sulfuric acid solution.

6. Dilute nitric acid being added to solid calcium carbonate.

7. A piece of clean lead added to a solution of silver nitrate.

8. The addition of a silver nitrate solution to a solution containing magnesium chloride.

9. Carbon dioxide gas being bubbled through limewater solution.

10. Phosphoric acid added to solid silver carbonate.

11. Hydrogen sulfide gas bubbled through silver nitrate solution.

|12. Mixing of nitric acid and sodium carbonate solutions.

13. Sodium iodide solution added to lead(II) acetate solution.

14. The addition of excess sodium hydroxide solution to a suspension of aluminium hydroxide in water to produce a sodium tetrahydroxyaluminate ($Na[A\ell(OH)_4]$) solution.

Set 36: The pH scale

Concentration of hydrogen ions is written like this $[H^+]$ using square brackets. The pH scale is a convenient way of expressing concentrations of hydrogen ions, $[H^+]$ in aqueous solutions. The definition of pH is expressed by the equation:

$\mathrm{pH} = -\log_{10} [H^+]$ where pH is a number without units.

The pH scale has a usual range of 0 to 14 as illustrated in the diagram below:

pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
$[H^+]$	$1=10^0$	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}	10^{-9}	10^{-10}	10^{-11}	10^{-12}	10^{-13}	10^{-14}
Acidic pH < 7								Neutral pH = 7	Basic pH > 7						

Examples

1. Hydrochloric acid is one of the most common industrial acids with many uses. It is produced in Western Australia in large quantities. Low concentration solutions of the acid can be standardised and used in analysis of basic solutions.

   Calculate the $[H^+]$ and pH of a 0.0200 mol L^{-1} $HC\ell$ solution.

   (a) Write an ionisation equation:

   $$HC\ell(aq) \rightarrow H^+(aq) + C\ell^-(aq)$$

   (b) Calculate the hydrogen ion concentration:

   As $HC\ell$ is a strong acid, it is fully ionised.
   So the $[H^+] = 0.0200$ mol L^{-1}

   (c) Calculate the pH of the solution:

   $$\mathrm{pH} = -\log_{10} [H^+] = \log_{10} (0.0200) = 1.700$$

2. Orange juice has a pH of 4.50. What is the $[H^+]$ in orange juice?

   Calculate the hydrogen ion concentration:

   $$\mathrm{pH} = -\log_{10} [H^+]$$
   $$-4.50 = \log_{10} [H^+]$$
   $$[H^+] = 10^{-4.50} = 3.16 \times 10^{-5} \text{ mol L}^{-1}$$

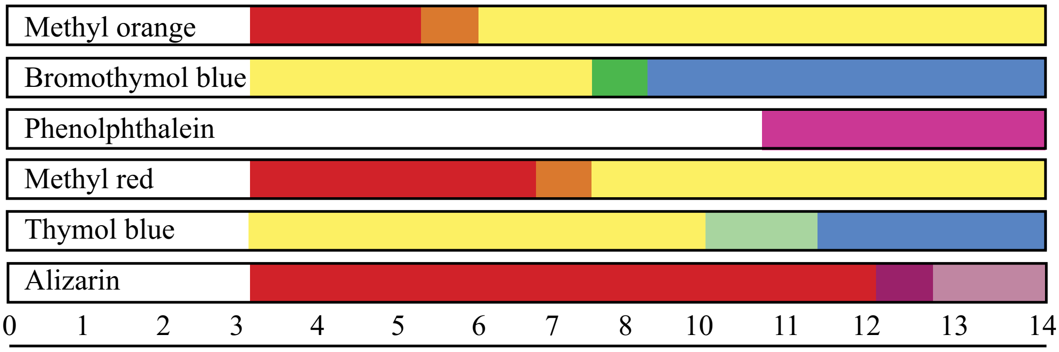
3. A dilution problem.
   Joan, a laboratory technician, requires 250.0 mL of 0.210 mol L^{-1} hydrochloric acid solution, but finds she has only 2.00 mol L^{-1} hydrochloric acid solution. What volume of 2.00 mol L^{-1} acid should Joan use to prepare 250 mL of the new solution?
   As this is a dilution problem we can use the dilution equation:

   $$c_1V_1 = c_2V_2$$
   $$0.250 \times 0.210 = 2.00 \times V_2$$
   $$V_2 = 26.3 \text{ mL}$$

   Joan should use 26.3 mL of 2.00 mol L^{-1} acid and make it up to 250 mL with distilled water.

Set 36: Exercises

In all of the following questions the temperature is assumed to be 25 °C
Questions 1 and 2 use the following table that shows the colour ranges of some acid-base indicators.



1. A solution is yellow in methyl orange, blue in bromothymol blue and colourless in phenolphthalein. Underline the pH range of the solution?
   (a) 4.0 to 6.0
   (b) 6.0 to 7.0
   (c) 7.0 to 9.0
   (d) 8.0 to 10.0

2. Mary used indicators to determine whether three colourless solutions were acidic or basic. The indicators used are shown in the table above. Samples of each solution were tested with the indicators. The colours of the resulting solutions are shown in the table.

Indicator added	Colour of Solution A	Colour of Solution B	Colour of Solution C
Methyl orange	yellow	yellow	yellow
Methyl red	yellow	yellow	yellow
Thymol blue	blue	blue	yellow
Alizarin	purple	red	red

Mary concluded that each of the three solutions tested was basic. Is she right? ☐ Yes ☐ No

3. Calculate the $[H^+]$ and pH of the following solutions.
   (a) A 0.100 mol L^{-1} HCℓ solution.

   (b) A 0.00500 mol L^{-1} HNO_3 solution.

   (c) A 2.00 mol L^{-1} HCℓ solution

Set 37: Concentration calculations of acids and bases

Moles per litre and grams per litre

Concentration is defined as the amount of solute per unit volume of solution. The units most commonly used for acid and base concentration is the mole per litre (mol L^{-1}).

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

where **c** is the concentration of solute in moles per litre (mol L^{-1}),
n is the number of moles of solute (mol), and
V is the volume of the solution in litres (L)

Note: Solution concentrations can be expressed using 'c' or '[ ]' that is, c(NaOH) or [NaOH] both refer to the concentration of NaOH in solution, expressed in mol L^{-1}.

Examples

1. Determine the concentration in mol L^{-1} of sulfuric acid when 1.50 moles of acid is made up to 600.0 mL with distilled water.

$$c = \frac{n}{V} = \frac{1.50}{0.600} = 2.50 \text{ mol L}^{-1}$$

2. Calculate the concentration in mol L^{-1} of a potassium hydroxide solution which contains 5.56 g of the solute dissolved in 750.0 mL of solution.

(a) Calculate the number of moles of solute:

$$n = \frac{m}{M} = \frac{5.56}{56.11} = 0.0991 \text{ mol} \qquad M(KOH) = 56.11 \text{ g mol}^{-1}$$

(b) Calculate the concentration of the KOH solution:

$$c = \frac{n}{V} = \frac{0.0991}{0.750} = 0.132 \text{ mol L}^{-1}$$

3. Find the number of moles of sodium hydroxide in 15.0 mL of a 2.50 mol L^{-1} NaOH solution

$$n = cV = (2.50)(0.0150) = 3.75 \times 10^{-2} \text{ mol of NaOH}$$

Concentration of ions

Strong acids and strong bases produce almost 100% ions when dissolved in H_2O. Knowing the concentration of an acid or base solution and its formula, the concentration of ions present in solution can be calculated.
For example in 25.0 mL of a 0.200 mol L^{-1} solution of barium hydroxide ($Ba(OH)_2$) the concentrations of the individual ions are:

(a) $c(Ba^{2+}) = c(Ba(OH)_2) = 0.200 \text{ mol L}^{-1}$

(b) $c(OH^-) = 2 \times c(Ba(OH)_2) = 2 \times 0.200 = 0.400 \text{ mol L}^{-1}$

Dilutions

When a solution is diluted, the volume increases and the concentration decreases. The total amount of solute, that is, the number of dissolved moles, remains the same. This can be expressed as:

$$n_1 \text{ (before dilution)} = n_2 \text{ (after dilution)}$$

$$c_1 V_1 = c_2 V_2$$

Example

To prepare 250.0 mL of 0.200 mol L^{-1} nitric acid, what volume of concentrated 15.6 mol L^{-1} nitric acid would be required?

(a) Calculate the known number of moles:

$$n(\text{dilute } HNO_3) = c(HNO_3)\,(V) = (0.200)\,(250.0 \times 10^{-3}) = 5.00 \times 10^{-2} \text{ mol}$$

(b) Relate the unknown number of moles of concentrated HNO_3 to the known number of moles of dilute HNO_3:

$$n(\text{concentrated } HNO_3) = n(\text{dilute } HNO_3)$$

$$(15.6)(V) = 5.00 \times 10^{-2}$$

$$V = \frac{5.00 \times 10^{-2}}{15.6}$$

$$V = 3.21 \times 10^{-3} \text{ L}$$

Set 37: Exercises

1. 25.6 g of anhydrous sodium carbonate is dissolved in 200.0 mL of water.

   (a) Calculate the concentration in grams per litre of this solution. Assume no volume change.

   (b) Calculate the concentration in mol L^{-1}.

   (c) A 20.0 mL sample of this solution was placed in a 150.0 mL flask and 80.0 mL of distilled water was added. Determine the concentration of sodium ions in this diluted solution.

2. Calculate the concentration of
   (a) chloride ions in 25.0 mL, of 0.200 mol L^{-1} hydrochloric acid solution

   (b) hydroxide ions in 2.20 L of 2.02×10^{-3} mol L^{-1} barium hydroxide solution

3. 10.0 g of calcium nitrate is dissolved in 220.0 mL of distilled water. Calculate the concentration in mol L^{-1} of:
   (a) Ca^{2+} ions

   (b) NO_3^- ions

4. Calculate the volume of 10.0 mol L^{-1} nitric acid solution required to prepare 500.0 mL of 0.500 mol L^{-1} solution.

.

5. Calculate the concentration in mol L^{-1} of hydroxide ions in a solution prepared by mixing 360.0 mL of 0.250 mol L^{-1} sodium hydroxide solution with 675.0 mL of 1.20 mol L^{-1} potassium hydroxide solution. Assume solution volumes are additive

6. What volume of water must be added to 150.0 mL of 1.10 mol L^{-1} sulfuric acid solution to prepare a 0.210 mol L^{-1} solution?

7. Battery acid is made by diluting (98.0% by mass) sulfuric acid with water. If 1.00 kg of the 98.0% acid was mixed with water to make up 3.00 L of battery acid, calculate the concentration of the resulting acid in g L^{-1}.

8. Slaked lime (calcium hydroxide) can be used to reduce soil acidity and is sparingly soluble in water. In a solution containing 10.0 mg of calcium hydroxide in 1.00 L of solution, calculate the concentration in mol L^{-1} of:

(a) calcium ions;

(b) hydroxide ions;

10. Where vegetables are grown hydroponically, fertilisers are applied in solution. Ammonium nitrate, potassium sulfate and calcium dihydrogenphosphate are typical compounds used.
(a) Write equations to illustrate how each of these compounds result from acid-base reactions.


Question 10: Hydroponics (photo: Joel Malcolm)

(b) Calculate the mass of solute required to prepare each of the solutions listed:
(i) 10.0 L of 1.50 mol L^{-1} ammonium nitrate

(ii) 2.50 L of 2.80 mol L^{-1} potassium sulfate

(iii) 500.0 mL of 0.100 mol L^{-1} calcium dihydrogenphosphate

11. Concentrated sulfuric acid is to be diluted before it is used to dissolve a sample of iron prior to analysis. Calculate the volume of 18.0 mol L^{-1} sulfuric acid required to prepare 800.0 mL of a 2.50 mol L^{-1} solution.

12. A mixture of hydrochloric acid and nitric acid is used to digest meat allowing analysis for lead, copper, cadmium and zinc. Calculate the concentration of hydrogen ions, in mol L^{-1}, in such a solution prepared by mixing 3.60×10^2 mL of a 14.3 mol L^{-1} nitric acid solution with 675 mL of 12.1 mol L^{-1} hydrochloric acid solution. Assume the solution volumes are additive.

Set 38: Acid and base reaction stoichiometry

Example (mass to volume calculation)

A student carried out a laboratory experiment hoping to produce enough hydrogen gas to fill a 2.00 L balloon at S.T.P. The student reacted 3.00 g of magnesium in excess 2.0 mol L^{-1} hydrochloric acid solution. Will enough hydrogen be produced?

(a) Write an equation for the reaction

$$Mg(s) + 2\ HC\ell(aq) \rightarrow MgC\ell_2(aq) + H_2(g)$$

$$Mg(s) + 2\ H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$$

(b) Calculate the number of moles of magnesium

$$n(Mg) = \frac{m}{M} = \frac{3.00}{24.3} = 0.123\ mol$$

(c) Calculate the number of moles of hydrogen gas produced

$$n(H_2) = n(Mg) = 0.123\ mol$$

(d) Calculate the volume of hydrogen gas produced (1 mole of a gas occupies 22.71 L at STP)

$$V(H_2) = n(H_2) \times 22.71 = (0.123) \times 22.71 = 2.79\ L$$

The reaction will produce enough gas to fill the 2.00 L balloon.

Example (volume to moles and mass calculation)

Brick cleaning acid reacts with limestone (calcium carbonate) producing carbon dioxide gas. How many moles and what mass of limestone is needed to produce 2.00 L of carbon dioxide gas at STP?

(a) Write an equation for the reaction

$$CaCO_3(s) + 2\ HC\ell(aq) \rightarrow CaC\ell_2(aq) + CO_2(g)$$

$$CaCO_3(s) + 2\ H^+(aq) \rightarrow Ca^{2+}(aq) + CO_2(g)$$

(b) Calculate the number of moles of carbon dioxide gas (1 mole of a gas occupies 22.71 L at S.T.P.)

$$n(CO_2) = \frac{V(CO_2)}{22.71} = \frac{2}{22.71} = 0.0881\ mol$$

(c) Calculate the number of moles of calcium carbonate needed

$$n(CaCO_3) = n(CO_2) = 0.0881\ mol$$

(d) Calculate the mass of calcium carbonate needed

$$m(CaCO_3) = n(CaCO_3) \times M(CaCO_3) = 0.0881 \times 100.09 = 8.81\ g$$

Set 38: Exercises

1. A 20.0 mL sample of stomach juices obtained from a patient suffering from hyperacidity was neutralised using 10.4 mg of solid aluminium hydroxide. Assuming that the volume of stomach juice present in the patient's stomach at the time was 250.0 mL and the only acid present was hydrochloric acid, calculate the mass of acid present in the stomach.

2. Sweat is a solution containing mainly sodium, potassium and chloride ions. During vigorous exercise, particularly in hot and humid conditions this solution often runs off the skin. In doing so it dissolves significant amounts of hydrogen ions from the skin. The concentration of hydrogen ions in sweat is typically 1.00×10^{-5} g L^{-1}. Calculate the mass of solid calcium hydroxide required to neutralise 10.0 mL of sweat.

3. A 2.89 g sample of sandstone, containing only calcium carbonate and silicon dioxide, is analysed by reacting it with hydrochloric acid. 10.7 mL of 2.50 mol L^{-1} hydrochloric acid solution is required for complete reaction. Calculate the percentage by mass of calcium carbonate in the sandstone sample.

4. Hydrochloric acid is used to decrease the pH of swimming pool water. To determine the amount of acid to be added, the concentration of the acid needs to be known. To determine the concentration of some hydrochloric acid, 5.60 g of pure calcium carbonate in the form of calcite was placed into 20.0 mL of the acid. When the reaction was completed the remaining calcite is removed, washed in distilled water and dried. It was found to have a mass of 5.09 g. Calculate the concentration of the hydrochloric acid.

Question 4: Hydrochloric acid

5. To determine the purity of a sample of cave limestone (mostly calcium carbonate) an analyst measured the mass of a small sample of limestone to be 2.59 g. He placed enough hydrochloric acid into a beaker to react with the limestone sample. The beaker and acid had a mass of 110.61 g. He then placed the limestone sample into the acid and allowed it to completely react. After the reaction was completed the mass of the beaker and its contents was found to be 112.22 g. Calculate the percentage purity of the limestone.

6. Sodium carbonate can be used to neutralise acid spills in the environment. It can be used as a solid or as a solution that can gradually seep into the soil.

(a) Calculate the concentration, in mol L^{-1}, of a sodium carbonate solution produced from 2.50×10^2 g of sodium carbonate-10-water dissolved into 2.00 L of water to produce 2.10 L of solution.

(b) If the acid spilt was hydrochloric acid, calculate the mass of sodium chloride produced in the soil if all of the 2.10 L of the sodium carbonate solution reacted with acid in the soil.

7. Slaked lime ($Ca(OH)_2$) is sometimes used to reduce soil acidity. To neutralise the top 200 mm of soil in a paddock required 1.00×10^2 g of slaked lime per square metre. Calculate the number of moles of hydrogen ions per square metre in the top 2.00×10^2 mm of soil.

8. 9.00×10^4 L of rain water with an average hydrogen ion concentration of 1.00×10^{-4} g L^{-1} was collected and stored in a galvanised iron tank. After several months the pH of the water was found to be 7.00 ($[H^+] = 1.00 \times 10^{-7}$ mol L^{-1}). Assuming the zinc was the only material in contact with the water, calculate the

(a) mass of zinc that would have dissolved

(b) the concentration of zinc ion, in mol L^{-1}, in the rain water.

ANSWERS

Set 1: Scientific notation and unit conversions

1. (a) 3.29×10^2 (b) 1.006×10^3 (c) 5.731×10^{-1} (d) 6.724×10^{10} (e) 4×10^{-2}
   (f) 7.8×10^{-8}
2. (a) 2.643×10^3 g (b) 1.74×10^{-2} g (c) 2.50×10^6 g (d) 4.39×10^{-1} g (e) 2.846×10^2 mg
   (f) 6.72×10^2 kg
3. (a) 1.02×10^3 cm (b) 1.26×10^{-2} m (c) 1.46×10^3 mm (d) 1.43267×10^2 m (e) 1.09×10^{-7} m
   (f) 1.41×10^{-4} m^2 (g) 8.3×10^{-9} m^3 (h) 1.22×10^1 m (i) 1.5 cm (j) 1 m
   (k) 1.9×10^{-4} cm (l) 1.5×10^3 mm^2 (m) 4.9×10^{-6} m^3 (n) 1.67×10^{-3} L

Set 2: Significant figures

1. (a) 3 (b) 4 (c) 5 (d) 3 (e) 1,2,3 or 4
   (f) 1 (g) 3 (h) 4
2. (a) 7.25 (b) 0.0174 (c) 6.28×10^{-3} (d) 1.13×10^8
3. (a) 8.8 (b) 6.60 × 10+ (c) 1.588 × 10+ (d) 3.34×10^4
4. (a) 6.79×10^1 (b) 5.4×10^4 (c) 2.83 (d) 4.913×10^7 (e) 2.7×10^3
   (f) 3.51×10^2
5. (a) 9.3 cm^3

Set 3: Random and systematic errors

1. (a) 7.0 mL (b) parallax error. (c) systematic error (d) the volume is overestimated (e) ± 0.25 mL
2. (a) systematic (b) the pH readings from the meter will all be above the true value.
3. (a) parallax error if the meniscus of the liquid is always viewed either from above or below the level of the liquid.
   (b) not taring (zeroing) the balance before weighing.
4. (a) differences in human perception of colour.
   (b) reading an analogue scale so the uncertainty is half the smallest scale division.
   (c) this relies on human perception of colour which varies with individuals.
   (d) reading with a ruler so the uncertainty is half the smallest scale division.
5. Experiments that have systematic errors may have results that are not accurate. If the systematic error is consistent throughout all trials the results obtained from repeating trials can be close in value and reproducible but may not be accurate.
6. By using a digital top pan balance reading to as many decimal places as possible, e.g. using a balance that reads to 2 decimal places rather than one place will decrease the uncertainty from ± 0.1 to ± 0.01 g.
7. The reading is limited by the reaction time of the student operating the stopwatch, which is much longer than one hundredth of a second and on average is around 0.1 s.
8. (a) the results are precise but not accurate.
   (b) the random error is relatively low as the range of values in the data set is just 0.03%. A systematic error has resulted - average of data being over 1% below the more accurate value.
   (c) the accuracy of analysis is improved by altering the method of analysis not by repetition of the method.
9. (a) the mass of cracker, volume of water heated, distance burning cracker held below test tube, stirring of the test tube.
   (b) energy loss to the surroundings (air, test tube) in the form of heat and light.
10. (a) heat loss to the surroundings
   (b) measurement of the volumes of the solutions, the reading of the thermometer
   (c) repeat the experiment several times and calculate an average.
   (d) alter the method to minimise heat loss by improved insulation around cup, using a lid, and using a stirrer.
11. (a) the measurement of the masses
   (b) the true value for the mass recorded could be above or below the recorded value.
   (c) two measurements each with an uncertainty of +/- 0.01 g are used to obtain the mass of water (13.20-12.95)
   (d) the method does not heat the sample to constant mass and so not all the water has been removed.
   (e) it will be an underestimate.
   (f) the sample was heated to a constant mass which ensured that all of the water was removed from the hydrated sample.
   (g) no, only one value is used in the calculations so the uncertainty of each measurement is irrelevant.

12. (a) $2\ CH_3OH + 3\ O_2 \rightarrow 2\ CO_2 + 4\ H_2O$ $C_2H_5OH + 3\ O_2 \rightarrow 2\ CO_2 + 3\ H_2O_2$ $C_3H_7OH + 9\ O_2 \rightarrow 6\ CO_2 + 8\ H_2O$
(b) ethanol 37,620 J, propan-1-ol 41,800 J
(c) ethanol 781 kJ mol^{-1} propan-1-ol 952 kJ mol^{-1}
(d) As the molecular mass of the alcohol increases the molar heat of combustion increases.
(e) Heat loss to the surroundings including the air and the container of the water.
(f) Methanol 53 %, ethanol 57 %, propanol 47%.
(g) The sooty deposit indicates that the propan-1-ol did not combust fully into carbon dioxide and water. The soot indicates that carbon was produced.
(h) burning the fuels in oxygen enriched air to encourage complete combustion.

Set 4: Elements and symbols

1. (a) fluorine F (b) calcium Ca (c) manganese Mn (d) tungsten W (e) silver Ag
(f) uranium U (g) platinum Pt (h) iodine I (i) neon Ne (j) bromine Br
(k) strontium Sr (l) astatine At (m) chromium Cr (n) copper Cu (o) barium Ba

2. (a) Na sodium (b) Rb rubidium (c) Cℓ chlorine (d) Au gold (e) Li lithium
(f) Pb lead (g) Ni nickel (h) Aℓ aluminium (i) Fe iron (j) Co cobalt
(k) P phosphorus (I) Kr krypton (m) Zn zinc (n) K potassium (o) Sn tin

Set 5: Atoms and isotopes

1. (a) label nucleus and electron cloud (b) 3 protons, 4 neutrons and 3 electron (c) Li 2
2

Particle	Charge	Mass
Proton	+1	1
Neutron	0	1
Electron	-1	1/1836

3. (a) Negligible difference (b) Gives a +1 charge.
4. protons, electrons, neutrons.
5. Copper: E = Cu ; A = 63: Z = 29 6
6.

Symbol	Element	A	Z	Neutrons
$^{14}_{6}C$	Carbon	14	6	8
$^{35}_{17}C\ell$	Chlorine	35	17	18
$^{56}_{26}Fe$	Iron	56	26	30
$^{31}_{15}Ga$	Gallium	31	15	16
$^{108}_{47}Ag$	Silver	108	47	61
$^{12}_{6}C$	Carbon	12	6	6
$^{23}_{11}Na$	Sodium	23	11	12
$^{64}_{29}Cu$	Copper	64	29	35
$^{40}_{20}Ca$	Calcium	40	20	20
$^{13}_{6}C$	Carbon	13	6	7

7. Carbon
8. (a) Hydrogen-1 $^{1}_{1}H$ (b) Hydrogen-2 $^{2}_{1}H$ (c) Hydrogen-3 $^{3}_{1}H$

Set 6: Atomic structure and the Periodic able

1.

Z	Name	Symbol	Metal/ Non-metal	Electron configuration	Valence electron behaviour
1	Hydrogen	**H**	Metal	1	Lose 1, gain 1, share 1
2	Helium	**He**	N/A	2	N/A
3	Lithium	**Li**	metal	2,1	Lose 1
4	Beryllium	**Be**	metal	2,2	Lose 2
5	Boron	**B**	Metal	2,3	Share 3
6	Carbon	**C**	Non-metal	2,4	Share 4
7	Nitrogen	**N**	Non-metal	2,5	Lose 3, share 3
8	Oxygen	**O**	Non-metal	2,6	Gain 2, share 2
9	Fluorine	**F**	Non-metal	2,7	Gain 1, share 1
10	Neon	**Ne**	N/A	2,8	N/A
11	Sodium	**Na**	metal	2,8,1	Lose 1
12	Magnesium	**Mg**	metal	2,8,2	Lose 2
13	Aluminium	**Aℓ**	Metal	2,8,3	Lose 3
14	Silicon	**Si**	Non-metal	2,8,4	Share 4
15	Phosphorus	**P**	Non-metal	2,8,5	Lose 3, share 3
16	Sulfur	**S**	Non-metal	2,8,6	Gain 2, share 2
17	Chlorine	**Cℓ**	Non-metal	2,8,7	Gain 1, share 1
18	Argon	**Ar**	N/A	2,8,8	N/A
19	Potassium	**K**	metal	2,8,8,1	Lose 1
20	Calcium	**Ca**	metal	2,8,8,2	Lose 2

2. (a) Metals (b) Left hand side of the table

3. (a) Non-metals (b) Right hand side of the table

4. (a Noble gases (also boron and some Group 14 elements (e.g. carbon, silicon))
   (b) Far right hand side of the Table (also middle of the table)

5. The number of electrons likely to be gained or lost by an element can be predicted based on the Group the element belongs to on the Periodic Table. Elements in Groups 1 - 3 will lose electrons (the number lost will be the same as the Group number). Elements in the middle of the main Groups tend to share electrons (this only applies to the higher elements in the group) and elements in Groups 15 - 17 tend to gain electrons. Elements in Group 18 do not gain or lose electrons.

6.

Particle	Symbol	Atomic number	Number of protons	Number of electrons	Number of neutrons
Example	$^{19}_{9}\mathbf{F}$	9	9	9	10
A	$^{16}_{8}\mathbf{O}$	8	8	8	8
B	$^{17}_{8}\mathbf{O}$	8	8	8	9
C	$^{16}_{8}\mathbf{O}^{2-}$	8	8	10	8
D	$^{17}_{9}\mathbf{F}$	9	9	9	8

7. A, B and C are isotopes and D and the example are isotopes. They have the same atomic number and different mass numbers.

8. The number of protons of an atom defines which element (identity) it is while the chemical behaviour of an atom can be explained in terms of its electron configuration (personality)

9. (a) 2,8,8 (b) 2,8 (c) 2,8,8,1 (d) 2,4

10. (a) chlorine (b) calcium (c) aluminium

11. (a) not ground state (b) not ground state (c) not ground state

ANSWERS

Set 7: Ionisation energy

1. Using sodium as an example, the ionization energy is the energy required to remove a mole of electrons from the third shell of a mole of sodium atoms in the gaseous state.
2. High energy is needed so the sample of element changes to gaseous phase before ionisation occurs.
3. Periodicity is shown by the first ionisation energy. For example the first ionisation energy of successive elements increase across a Period up to Group 18. When you move to the next element in the sequence you start filling a new Period, the first ionization energy drops and then increases again as you move to successive elements across the Period.
4. Atoms with low first ionisation energies tend to lose electrons, while those with higher first ionisation energies tend to gain electrons.
5. The third electron is a whole shell closer to the nucleus resulting in a much higher ionisation energy than the first or second electron.
6. (a) 3 (b) A large jump in ionisation energy represents a change between electron shells.
   (c) Yes, same number of valence electrons.
7. 2, a large increase in energy occurs between the second and third ionisation energies.
8. A, the large increase in ionisation energies after first ionisation energy which indicates one valence electron.
9. E, the large increase in ionisation energies occurs after the third ionisation energy.
10. Atoms with a small number of valence electrons will tend to form metallic bonds, while atoms with a larger number of valence electrons will tend to share electrons in covalent bonds.

Set 8: Periodic Trends

1. The metallic character increases.
   They have the same number of electrons.
   The atomic radius increases.
   The electronegativity decreases.
   The first ionisation energy decreases.
2. (a) Across the Period (left to right) atomic radii get smaller.
   (b) As sodium loses the electron in the third energy level, so size of the ion is smaller than the atom. The chlorine gains an electron and so completes its third energy level and the size of its ion as a consequence is larger.
3. (a) The electronegativity increases as number of protons in nucleus is increasing and electrons are attracted more closely.
   (b) Electronegativity increases as you move up a Group - same number of electrons but closer to nucleus.
4. It could be placed in Group 2, a metal forming strong ionic bonds with non-metals.
5. 1-3 electrons metallic; 3-4 electrons covalent network; 5-7 electrons covalent molecular.

Set 9: Properties and structures of atoms

1. (a) The different images represent the sum of knowledge to a point of time in history - they included the solid sphere, the "plum pudding" model, Bohr model and the Quantum Mechanical model.
   (b) They show increasing complexity of the structure of the atom
   (c) It represents a dynamic model
   (d) E,B,A,C,D, Set 5
   (e) Dalton, Thomson, Rutherford, Bohr, Chadwick and Planck
2. All organic materials absorb and reflect light in the near-infrared (NIR) region of the spectrum, and the particular pattern of reflectance, how much at each wavelength is that material's spectrum. The NIR spectrum gives as a chemical 'fingerprint' of that material and this fingerprint can be used to positively identify it.
3. To identify atoms in a sample.
4. Your diagram should include an atomiser, a radiation source, wavelength selector, detector, amplifier and signal processor
5. (a) Tc - radioactive tracer
   (b) Mo - used to manufacture Tc 99
   (c) Co - as a tracer and to sterilize medical equipment
6. Dalton - atomic theory, 1803. Thompson - electron, 1897. Ruth-erford - nucleus and electrons, 1944. Bohr- shells, 1913. Chadwick - neutron, 1932.

Set 10: Relative atomic mass and mass spectroscopy

1. 3.55×10^1 2. 2.81×10^1 3. 2.07×10^2 4. Another isotope of barium with smaller atomic mass

5. 2.43 101 6. 72.5% of Cu-63 and 27.5% of Cu-65 7. **Sample A**: 60.0% of Br-79 and 40.0% of Br-81
Sample B: 5.0% of Br-79 and 95.0% of Br-81

8. (b) (i) The greater the charge the larger the deflection (ii) The larger the mass the less the deflection
(iii) I the higher charge Cu-63 is 2+ ions the greater the deflection;
II the lower mass B-10 single + the greater the deflection

9. 69.60 10. 20.19 11. 39.99

Set 11: Molar mass

1. (a) 56.11 g mol^{-1} (b) 134.5 g mol^{-1} (c) 133.3 g mol^{-1} (d) 74.09 g mol^{-1} (e) 124.1 g mol^{-1} (f) 286.1 g mol^{-1}
(g) 195.9 g mol^{-1} (h) 16.04 g mol^{-1} (i) 342.3 g mol^{-1} (j) 143.3 g mol^{-1} (k) 319.2 g mol^{-1} (l) 142.0 g mol^{-1}
(m) 399.9 g mol^{-1} (n) 64.06 g mol^{-1} (o) 98.08 g mol^{-1}

2. (a) C-14 has two more neutrons in the nucleus than C-12 (b) C-13
(c) C-12 is the most abundant because the RAM of Carbon is 12.01, which is closest to the RAM of C-12 than for the other isotopes C-13 and C-14. This can be confirmed by measuring the isotopic distribution in a mass spectrometer.

3. (a) Normal hydrogen: 1 proton, 1 electron. Deuterium:1 proton, 1 electron, 1 neutron. Tritium: 1 proton, 1 electron, 2 neutrons.
(b) Deuterium and tritium are used in nuclear research for experimental fusion reactors. Deuterated water (heavy water) is used as a moderator in some reactor designs. Tritium is used extensively in biological research as a radioactive tracer.Set

12: Moles, particles and mass

1. (a) n = 2.96 mol (b) n = 4.00 mol (c) n = 3.99 mol

2. (a) m = 33.0 g (b) m = 1.0 10^1 g (c) m = 252 g

3. (a) n = 1.00 mol (b) n = 3.99 mol (c) n = 0.0500 mol

4. n = 3.50 mol 5. n = 0.116 mol

6. (a) $N = 6.02 \times 10^{23}$ (b) $N = 1.5 \times 10^{23}$ (c) $N = 2.81 \times 10^{24}$ (d) $N = 2.15 \times 10^{23}$

7. (a) $N = 1.34 \times 10^{24}$ (b) $N = 8.97 \times 10^{24}$ (c) $N = 2.00 \times 10^{19}$ (d) $N = 2.67 \times 10^{23}$

8. (a) $N = 1.48 \times 10^{24}$ (b) $N = 1.88 \times 10^{23}$ (c) $N = 3.24 \times 10^{21}$ (d) $N =1.00 \times 10^{24}$

Set 13: Interpretation of formulae

1 (a) $N = 2.5 \times 6.02214076 \times 10^{23} \times 2 = 3.01 \times 10^{24}$ hydrogen atoms
(b) $N = 0.0500 \times 6.02214076$ x $10^{23} \times 2 = 6.022 \times 10^{22}$ hydrogen atoms

2 (a) $n(PC\ell_3) = 21.0 \div 3 = 7.00$ mol (b) $n(KMnO_4) = 2.00 \div 4 = 0.500$ mol

3 (a) $n = m/M = 795 \div (63.55+16.00) = 9.99$ mol CuO $\therefore$ there are 19.98 mol of O $\therefore$ m = 19.98 x 16.00 = 319.8 g oxygen
(b) $n = m/M = 1.04 \div ((40.08)+(12.01\times2)+(4\times16.00)) = 5.968 \times 10^{-3}$ mol K_2SO_4 $\therefore$ $2 \times 5.968 \times 10^{-3}$ mol of K = 1.19×10^{-2} mol $\therefore$ $m = 39.10 \times 1.19 \times 10^{-2} = 0.467$ g
(c) $n = m/M = 38.4 \div ((39.10\times2)+32.06)+(4\times16.00)) = 0.2203$ mol = mol Ca $m=n \times M = 8.83$ g

4 (a) $n = m/M = 193 \div 32.06 = 6.0199$ mol S $\therefore$ 6.0199 mol SO2 $m = n \times M = 386$ g SO_2
(b) $n = m/M = 0.0960 \div 1.008 = 0.095238$ mol of H $\therefore$ 0.01190 mol of $(NH_4)_2SO_4$ $m = n \times M = 6.0199$ mol S $\therefore$ 6.0199 mol SO_2 $m = n \times M = 386$ g SO_2
(c) $n = m/M = 36.0 \div 12.01 = 2.9975$ mol C $\therefore$ $2.9975 \div 8$ mol C_8H18 = 0.37468 mol $m = n \times M = 42.79$ g

5 (a) $2.50 \times 3 = 7.50$ mol NH_4^+ (b) 2.50 mol PO_4^{3-} (c) $3 \times 2.50 = 7.50$ mol N
(d) $6 \times 2.50 = 15.0$ mol H (e) 2.50 mol P (f) $3 \times 2.50 = 10.0$ mol O

6 Fe_2O_3 $n = m / M = 640 / ((55.85\times2)+(16.00\times3)) = 4.0075$ mol $\therefore$ 8.0150 mol Fe $m=n\times M = 448$ g

7 C_2H_5OH $n = m / M = 144 / 12.01 = 11.990$ mol C $\therefore$ 5.995 mol C_2H_5OH $m=n\times M = 270$ g C_2H_5OH

8 2.5 mol oxalic acid = 6×2.5 mol O $m=n\times M = 240$ g O

9 $(CO(NH_2)_2)$ $n = m / M = 1.80 \times 104 / 60.062 = 299.6$ mol $\therefore$ 599 mol N

10 $NaC_{17}H_{35}COO$ $n = m / M = 1.25 / 12.01 = 0.10407$ mol C so 6.122×10^{-3} mol of soap $m=n\times M = 32.16$ g soap

ANSWERS

Set 14: Percentage composition

1. (a) NaOH: Na 57.5%,O 40.0%, H 2.50%
   (b) CH_3COOH: C 40.0%, O 53.3%, H 6.70%
   (c) $CuSO_4.5H_2O$: Cu 25.4%,S 12.9%, O 57.7%, H 4.00%
   (d) K_3PO_4: K 55.3%, P 14.6%, O 30.1%
2. (a) 63.9% C (b) 48.0% S (c) 40.5% O (d) 3.50% N
3. (a) 62.9% (b) 41.1% (c) 14.7%
4. (a) 53.00% Bi, 32.00% Pb, 15.0% Sn (b) Bi 79.5 g, Pb 47.8 g, Sn 22.7 g
5. 7.94% Cu, 16.1% Mg, 75.9% A 6. 80.35% Zn, 19.65% O 7. 79.9% Cu, 20.1% O
8. (a) chalcopyrite 34.6% Cu malachite 57.48% Cu (b) 2.89×10^5 g or 2.89×10^2 kg
9. Gibbsite: 34.59% A , 34.63% H_2O Kaolinite: 20.90% A , 13.95% H_2O
10. (a) 59.9% Ti (b) 3.12 tonnes

Set 15: Compounds and formulae

1. (a) carbon monoxide (b) sulfur dioxide (c) phosphorus pentachloride (d) dinitrogen monosulfide
   (e) diphosphorus tetrabromide (f) Sulfur hexafluoride
2. (a) NO (b) NO_2 (c) N_2O_4 (d) SO_3 (e) H_2O (f) P_5O_{10} (g) $HC\ell$ (h) HI (i) PBr_3 (j) NH_3
3. (a) $LiC\ell$ (b) Agl (c) KNO_3 (d) $CsCH_3COO$ (e) $BaBr_2$ (f) $CuSO_4$ (g) MnO_2 (h) $Ni(NO_3)_2$
(i) $A\ell_2O_3$ (j) $Cr_2(SO_4)_3$ (k) $Pb_3(PO_4)_4$ (l) $(NH_4)_2Cr_2O_7$
4. (a) carbon monoxide (b) dinitrogen monoxide (c) hydrogen sulfide (d) hydrogen peroxide (e) phosphoric acid
   (f) calcium nitride (g) nitric acid (hydrogen nitrate) (h) cobalt(II) hydrogenphosphate (i) copper(I) chloride
   (j) iron(II) sulfate

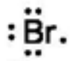
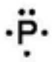
Set 16: Bonding and properties

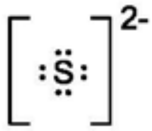
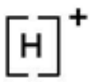
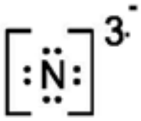
1. (a) Covalent molar (b) Metallic (c) Ionic (d) Covalent molar (e) Ionic (f)Covalent molar (g) Metallic
2. (a) Strong electrostatic attraction between positive metal ions and delocalised electrons requires a large amount of energy to overcome the attraction and melt the solid.
   (b) Delocalised electrons able to move through the lattice.
   (c) Delocalised electrons are able to transfer heat energy from its source to cooler parts of the lattice.
   (d) As the bonding is non-directional (delocalised electrons) layers of ions can slide over each other when a force applied.
3. Silicon dioxide is a covalent network structure. All of the bonds between the atoms are covalent bonds.
   (a) Strong 3D covalent primary bond
   (b) There are no free charge carriers, so it does not conduct electricity.
   (c) It is hard because it is held together by strong, 3D primary bonds.
   (d) Brittle because the bonds arc directional and when broken do not reform.
4. (a) Allotropes are different bonding arrangements of the same element e.g. phosphorus.
   (b)

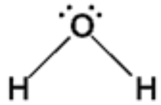
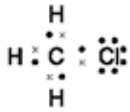
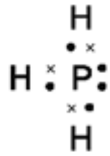
Allotrope	Colour	Melting point °C	Electrical conductivity	Hardness
Diamond	Colourless	3550 (sublimes)	Non-conductor	Hard
Graphite	Black	3550 (sublimes)	Conductor	Soft
Fullerene(C_{60})	Brown	600 (sublimes)	Non-conductor	Soft

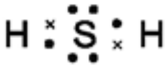
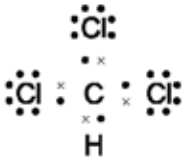
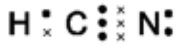
   (c) Diamond has a JD primary bonded structure so has a high melting point, has no free charge carriers so does not conduct, is hard due the primary bonding between its atoms.
   Graphite has a high melting point due to the extensive primary bonding between its atoms, conducts because of free electrons between the sheets and is soft because the sheets will slide over each other.
   Fullerene is molecular and has weak bonds between the molecules. It is therefore soft, melts easily but does not conduct as it has no free charge carriers.
5. (a) Strong non-directional ionic bonds require high temperature to break the bonds.
   (b) There arc no charge carriers so it docs not conduct when solid.
   (c) Ions are free to move when the substance is molten and act as charge carriers.
   (d) Brittle because if the ions are displaced they will be adjacent to an ion of the same charge and will repel.
   (e) Sodium chloride is hard because its ions arc bound together using strong primary ionic bonds.

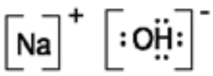
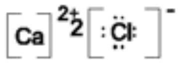
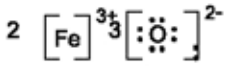
Set 18: Electron dot diagrams

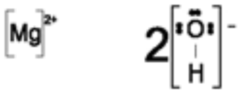
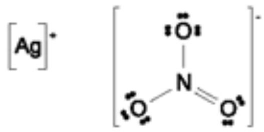
1. (a)  (b)  (c) 

   (d)  (e)  (f) 

2. (a)  (b)  (c) 

   (d)  (e)  (f) 

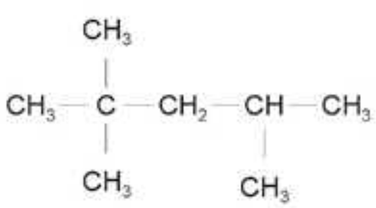
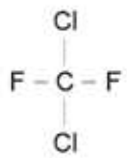
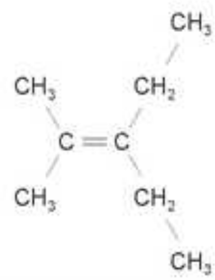
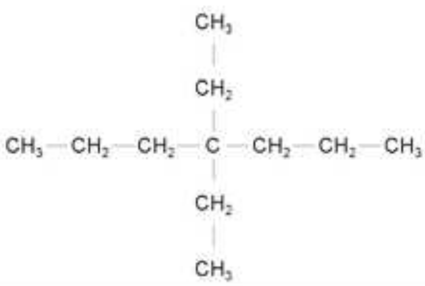
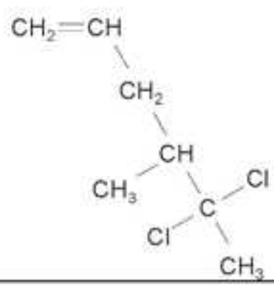
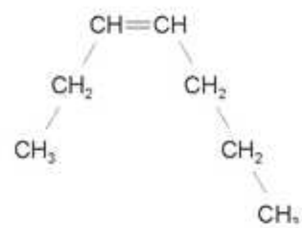
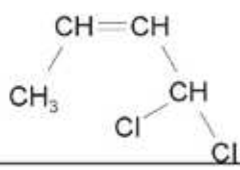
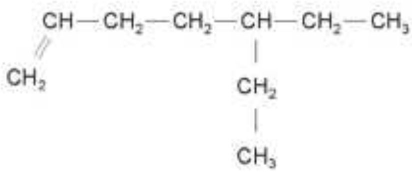
3. (a)  (b)  (c) 

   (d)  (e)  (f) 

Set 19: Naming and drawing hydrocarbons

1. (a) pentane (b) but-1-ene (c) 2-methylbutane (d) 1,1,2-trichloroethane (e) *E*-1,2-dibromoethene (f) 5-ethyl-3-methyloctane (g) 2,2-dimethylpropane (h) 2-methylpent-l-ene (i) 1,1,1-trifluoropentane (j) 5-methyl-*Z*-hex-2-ene

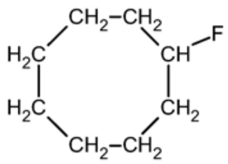
2.

a		b		c	
d		e		f	
g		h			

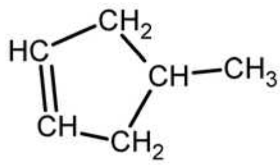
3. (a) (i) Pentane, 2-methylbutane, 2,2-dimethylpropane (ii) cis-pent-2-ene and trans-pent-2-ene
   (b) 1-bromobutane 2-bromobutane 1-bromo-2-methylpropane 2-bromo-2-methylpropane

4. (a) ethylbenzene (b) 3-chlorocyclopentene (c) propylcyclohexane (d) 3-bromo-5-chlorocycloheptene
   (e) 1-fluorocyclopentene (f) 1,2-dimethylhexane

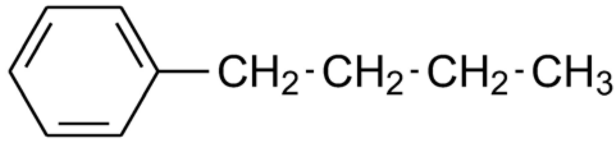
5. (a)



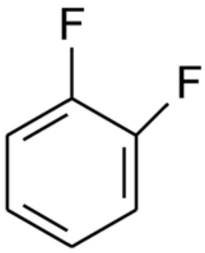
(b)



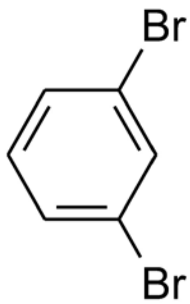
(c)



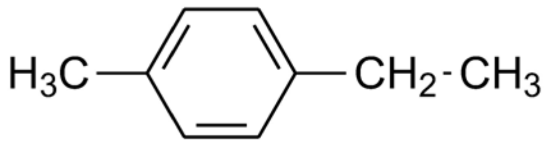
(d)



(e)

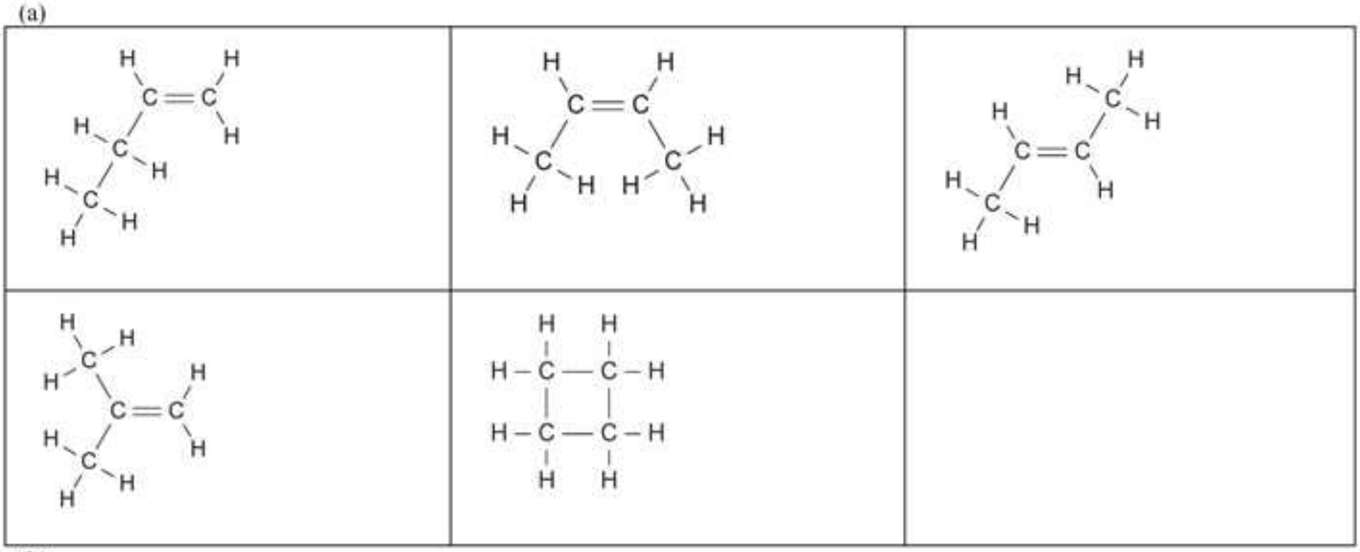


(f)



6.

(a)



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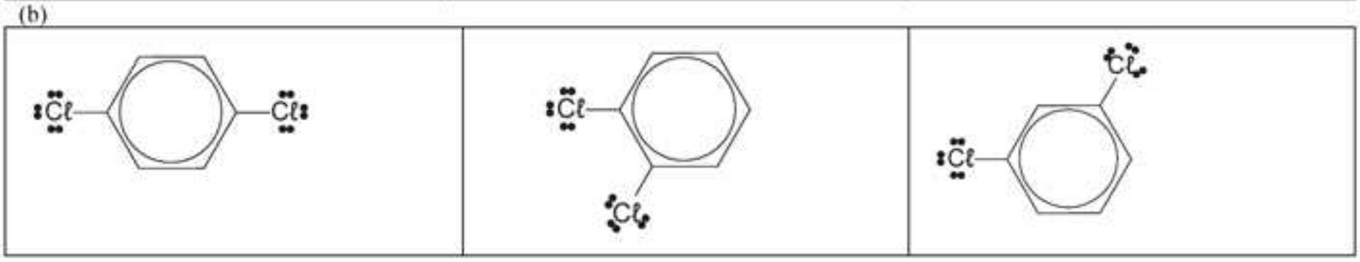
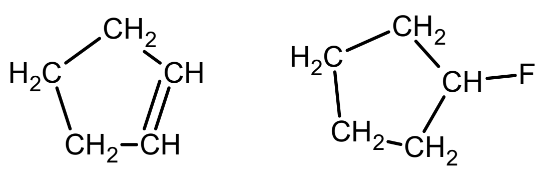


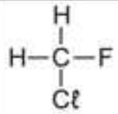
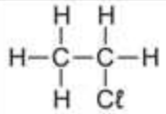
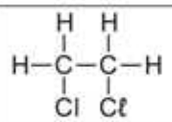
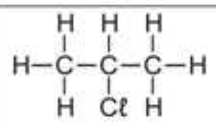
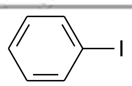
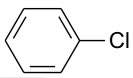
Set 20: Reactions of hydrocarbons

1. (a) $C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$ (b) $2\ C_2H_4 + 5\ O_2 \rightarrow 4\ CO_2 + 2\ H_2O$
   (c) $2\ C_6H_6 + 15\ O_2 \rightarrow 12\ CO_2 + 6\ H_2O$ (d) $2\ C_6H_5CH_2CH_3 + 21\ O_2 \rightarrow 16\ CO_2 + 10\ H_2O$

2. (a) $CH_3CH_3 + C\ell_2 \xrightarrow{U.V.} CH_3CH_2C + HC\ell$ (b) $CH_3CH{=}CH_2 + Br_2 \rightarrow CH_3CHBrCH_2Br$
   (c) $CH_3CH{=}CHCH_3 + HC\ell \rightarrow CH_3CHC\ CH_2CH_3$
   (d) $CH_3CH_2CH{=}CHCH_3 + H_2 \xrightarrow{Pt} CH_3CH_2CH_2CH_2CH_3$
   (e)



   (f) benzene + $Br_2 \rightarrow$ bromobenzene + HBr
   (g) $CH_3CH{=}CH_2 + C\ell_2 \rightarrow CH_3CHC\ell CH_2C\ell$

	Structure	Starting Hydrocarbon	Other Reagents	Equation
3(a)a		methane	fluorine chlorine	$CH_4 + C\ell_2 \xrightarrow{UV} CH_3C\ell + HC\ell$ $CH_3C\ell + F_2 \xrightarrow{UV} CH_2FC\ell + HF$
3(b)		ethene	hydrogen chloride	$C_2H_2 + HC\ell \rightarrow CH_3CH_2C\ell$
3(c)		ethene	chlorine	$C_2H_2 + C\ell_2 \rightarrow CH_2C\ell CH_2C\ell$
3(d)		propene	hydrogen chloride	$CH_3CHCH_2 + HC\ell \rightarrow CH_3CHC\ell CH_3$
3(e)		benzene	iodine	$C_6H_6 + I_2 \xrightarrow{UV} C_6H_5I + HI$
3(f)		benzene	chlorine	$C_6H_6 + C\ell_2 \xrightarrow{UV} C_6H_5C\ell + HC\ell$

21: Equations and observations

Set1. $CaCO_3(s) + 2\ H^+(aq) \rightarrow Ca_2+(aq) + CO_2(g) + H_2O(\ell)$
White solid dissolves; colourless, odourless gas evolved; colourless solution formed.

2. $Mg(s) + 2\ H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$
Silver solid dissolves; colourless, odourless gas evolved; colourless solution formed.

3. $NaHCO_3(s) + 4\ CH_3COOH(aq) \rightarrow Na^+ (aq) + CH_3COO^- (aq) + CO_2(g) + H_2O(\ell)$
White solid dissolves; colourless, odourless gas evolved; colourless solution formed.

4. $H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$ No visible reaction; heat evolved.

5. $H^+(aq) + NaOH(s) \rightarrow Na^+(aq) + H_2O(\ell)$ White solid dissolves; heat evolved.

6. $CoCO_3(s) + 2\ H^+(aq) \rightarrow CO_2(aq) + CO_2(g) + 4\ H_2O(\ell)$
Pink solid dissolves; colourless, odourless gas evolved, pink solution formed.

7. $Ba^{2+}(aq) + SO^{2-}(aq) \rightarrow BaSO\ (s)$ White precipitate formed.

8. $Pb^{2+}(aq) + 2\ I^-(aq) \rightarrow PbI_2(s)$ Yellow precipitate formed.

9. $2\ Au^{3+}(aq) + 3\ Cu(s) \rightarrow 2\ Au(s) + 3\ Cu2+(aq)$
Yellow/gold precipitate formed on surface of copper, yellow solution turns blue.

10. $2\ Na(s) + 2\ H_2O(\ell) \rightarrow 2\ Na+(aq) + 2\ OH^- (aq) + H_2(g)$
Solid reacts vigorously, may ignite; colourless, odourless gas evolved, colourless solution formed.

11. $K_2CO_3(s) + 2\ H^+(aq) \rightarrow 2\ K^+(aq) + CO_2(g) + H_2O(\ell)$
White solid dissolves; colourless, odourless gas evolved; colourless solution formed.

12. $H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$ No visible reaction; heat evolved.

13. $Ag+(aq) + Br-(aq) \rightarrow AgBr(s)$ two colourless solutions mixed and a cream ppt forms.

14. $Ni(s) + Cu^{2+}aq) \rightarrow Ni^{2+}(aq) + Cu(s)$
Silver metal added to blue solution. Brown/black precipitate forms solution turns green.

15. $Fe^{2+}(aq) + 2\ OH^-(aq) \rightarrow Fe(OH)_2(s)$
A green solution and a colourless solution are mixed and a green ppt forms.

ANSWERS

Set 22: Stoichiometry

1. (a) 4.00 mol (b) 15.0 mol (c) 3.90 mol

2. (a) $n(HNO_3)$ = 0.0600 mol $m(HNO_3)$ = 3.78 g (b) $n(Mg(NO_3)_2)$ = 0.0300 mol $m(Mg(NO_3)_2)$ = 4.45 g

3. (a) 0.200 mol m= 33.9 g (b) 0.100 mol m = 11.1 g (c) 0.100 mol m = 16.4 g

4. (a) m H_2SO_4 required = 1.23 g (b) Mol CuO = 0.0126 mol

5. 2 $KC\ell O_3$ 2 KC + 3 O_2 1.20 mol

6. 2 HgO 2 Hg + O_2 3.84 g

7. (a) m HNO_3 = 3.78 g (b) m CO_2 = 1.32 g (c) m $Ca(NO_3)_2$ = 4.92 g

8. (a) m HF = 0.222 g (b) m F_2 = 0.106 g (c) m UF6 = 0.978 g

9. (a) m = CaO = 258 kg (b) m CO_2 = 202 kg

Set 23 Energy changes

1. Exothermic a, b, e, f, g Endothermic c, d

2. (a) 2.80 103 kJ (b) 467 kJ (c) 1.70 103 kJ (d) 105 g

3. (a) 242 kJ (b) 1210 kJ (c) 16.9 g (d) 2 $H_2O(\ell)$+ 484 kJ 2 $H_2(g)$ + $O_2(g)$

4. (a) 2 $H_2(g)$ + CO(g) $CH_3OH(g)$ H = - 90 kJ (b) 2232 kJ (c) 7.78 g (d) 198 g

5. (a) They are molecules representing different structural combinations of an element
   (b) $O_3(g)$ + NO (g) $O_2(g)$ + $NO_2(g)$ + heat

6. Endothermic: boiling water, melting ice, cooking an egg and baking bread. Exothermic: rusting of iron, souring of milk, making ice blocks and burning a candle. Boiling water: When water boils energy is absorbed by the liquid, water molecules so they have enough energy to separate from the attraction for one another in the liquid and sufficient energy (vapour pressure) to overcome the pressure of the atmosphere to be released as gas particles.

7. A re-usable sodium acetate heat pack contains a supersaturated solution of sodium acetate. The supersaturated solution is unstable, crystallizes easily and is exothermic, so the pack gets hot. To restore the pack ready to be used again, simply redissolve the crystals into the solution, by heating the pack in hot water. CH_3COONa (aq) → CH_3COONa (s) + Heat

Set 24: Rates of reactions

1 (b) endothermic as the ΔH is positive (c) (i) ΔH = -20kJ (ii) Ea = 10kJ (d) The Δ is the same but the EA is lower

2. (a) Vertical axis: Number of particles, Horizontal axis: Energy (b) $T_2 > T_1$ as it has a higher average kinetic energy
   (c) EA the activation energy for the reaction.
   (d) Area A proportion of particles with energy equal to or greater than activation energy at temperature T_1.
   (e) Area B increase in proportion of particles with energy equal to or greater than activation energy when T_1 is increased to T_2.

3. (a) Heat + $N_2(g)$ + 3 $H_2(g)$ 2$NH_3(g)$
   (b) Increasing temperature has two effects. It increases the KE of the particles increasing frequency of collisions and successful collision, also a greater proportion of particles can achieve the required Ea so both effects increase the rate.
   (c) Increasing the pressure increases the concentration so collisions can occur more frequently.
   (d) A catalyst provides an alternate pathway requiring a lower Ea, a greater proportion of particles achieve the new Ea and rate of reaction is greater.

4. (a) $CaCO_3(s)$ + 2 $H^+(aq)$ $Ca2^+$ (aq) + $H_2O(\ell)$ + $CO_2(g)$
   (b) Increase the surface area of the reaction by crushing the marble chips into smaller particles. Increase the concentration of the acid for more collisions. Heat the mixture to provide more particles with energy equal to or greater than the EA.
   (c) Decrease the rate of reaction. Acetic acid is a weaker acid so [H^+] available to collide and react is less.

5. (a) $2\ CO(g) + O_2(g)$ $2\ CO_2(g)$ + Heat
   (d) Increase the pressure, which increases the concentration causing more collision and a taster reaction. Increase the temperature to increase the number of particles with energy equal to or greater than the EA.

6. Vapour particles collide with oxygen to react.

7. Souring of milk left out of the fridge - the higher temperature means that more milk particles have enough energy for the souring reaction to take place. Dissolving sugar or coffee in hot water - the higher temperature means that more particles have enough energy for the dissolving process to take place at a faster rate. Stirring when cooking - increases the surface area causing more collisions per second and a greater chance of a successful collision.
   Cutting potatoes for baking or boiling - increases the surface area and enables more collisions per second and a greater chance of a successful collision.

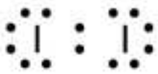
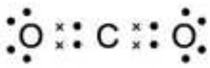
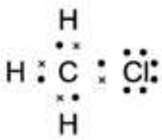
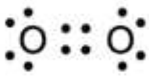
8. Enzymes work by helping with the orientation of complex particles creating more reaction collisions per second with suitable orientation for reactions to be successful.

Set 25 Molecular shape

1. The electronegativity of iodine is much less compared with fluorine as the outer electreons of iodine are much further from the nucleus and therefore held with less attractive force compared with electrons of fluorine.

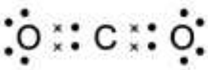
2. (a) Cℓ e- Cℓ — Electron in the middle as each atom attracts with same force
   Cℓ e- H — Electron closer to chlorine as it has a stronger attraction
   Cℓ e- Na^+ — Electron transferred/pulled away from sodium by chlorine
3. (b) (c) (d)

4.

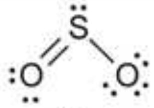
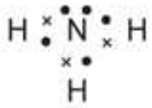
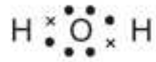
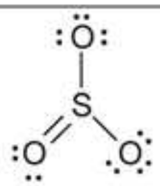
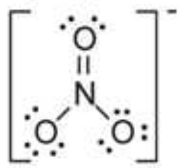
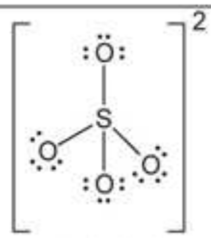
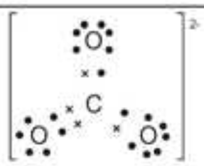
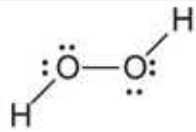
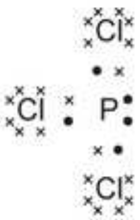
	electron dot diagram	shape	polarity
a		linear	non-polar
b		linear	non -polar
c		pyramidal	polar
d		tetrahedral	Polar
e		linear	non-polar

5. The bent and pyramidal shapes will always be polar as they are asymmetrical due to the lone pair of electrons.

6. The bonds within carbon dioxide for example are polar as carbon and oxygen have different electronegativities. The overall molecule is non-polar as the sum of the two dipoles cancel leaving the molecule non-polar.



7. Polarity affects the type of intermolecular force present. Non-polar molecules only have dispersion forces. Polar molecules can have dipole forces (or hydrogen bonding) as well as dispersion forces.

a	 bent	b	 trigonal pyramid	c	 bent
d	 Trigonal planar	e	 Trigonal planar	f	 tetrahedral
g	 Trigonal planar	h		i	 linear
j	 Trigonal pyramid				

Set 26: Intermolecular forces I

1. (a) dispersion forces (b) dispersion forces, hydrogen bonds
(c) hydrogen bonds and dispersion forces (d) hydrogen bonds and dispersion forces
(e) hydrogen bonds and dispersion forces (f) dispersion forces
(g) dipole-dipole forces and dispersion forces (h)dispersion forces

2. (b) Boiling points $O_2 < CH_3C < NH_3 < H_2O$

3. Vapour pressure at room temperature: HNO_3 > octane > water > mercury

4. Drawing an electron dot diagram shows if a molecule is polar and will engage in dipole-dipole bonding with water and so be soluble.

5. (a) insoluble (b) soluble (miscible) (c) insoluble
(d) soluble (e) insoluble

Set 27: Chromatography

1. a) A liquid solvent b) Usually a piece of glass with alumina or silica gel coating

2 a) Select a beaker just taller than the TLC plate. Place a half cm of solvent in the beaker. Place a pencil line one cm from the bottom of the plate and place a drop of the sample mixture on the line. Place the end of the plate in the solvent and cover the beaker.
b) A pencil line is made of graphite, which will not contaminate the results. The dyes in ball point pen might elute.
c) Placing a lid on the beaker creates a closed system. It allows the solvent to saturate the atmosphere in the beaker and establish dynamic equilibrium. It stops the solvent evaporating.

d) Red : 2.25 ÷ 4.75 = 0.47 Blue : 4.00 ÷ 4.75 = 0.84

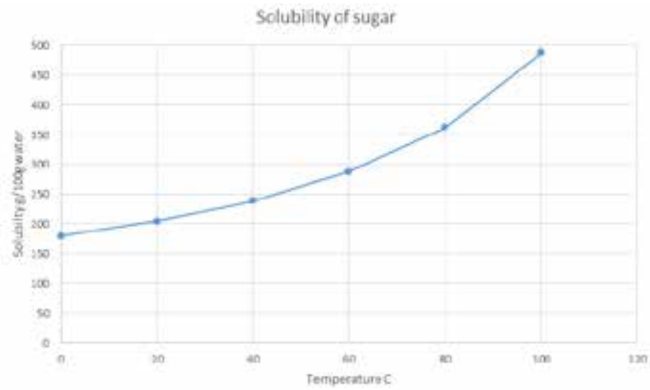
e) temperature, saturated atmosphere (do not remove the lid)

3. a) Helium and Diatomaceous earth with an adsorbed high boiling point liquid.
b) It might have no attraction for the stationary phase, it might have a very low boiling point.
c) The sample would be a gas and evaporate out of the liquid stationary phase quickly and then be carried on by the carrier gas.

4 a) The stationary phase can be very fine-grained giving it a large surface area and allowing good contact between the sample and the stationary phase. This means the sample components are better separated.
b) Silica is a polar substance and the polar molecules will be attracted and have a longer retention time than less polar molecules.

5. Pressure of the solvent, temperature, surface area of the stationary phase

Set 28: Solutions



1. (a) See graph
(b) Sugar solubility increases rapidly with temperature
(c) 30oC = 220 g/100g water, 70o = 320g/100g water
(d) At 20°C unsaturated < 204g/100g water
Saturated solution = 204g/100g water,
Supersaturated > 204g/100g water
(e) unsaturated, supersaturated, unsaturated

2. 1725 ppm. Suitable for use

3. 35g

4. Step 1: Add a little distilled water to each the insoluble one is barium sulfate.
Step 2: Add acid to each the carbonate will fizz.
Step 3: Add silver nitrate solution to the remaining 2. The chloride will produce a white precipitate. Remaining sample is barium nitrate.

5. (a) silver chloride white precipitate
(b) lead (II) iodide yellow precipitate
(c) barium sulfate white precipitate
(d) copper hydroxide - pale blue solid
(e) iron (III) phosphate-green precipitate forms

6. (a) $Ag^{+}(aq) + C\ell^{-}(aq) \rightarrow AgC\ell\,(s)$
(b) $Pb^{2+}(aq) + 2\,I^{-}(aq) \rightarrow PbI_2(s)$
(c) $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4\,(s)$
(d) $Cu^{2+}(aq) + 2\,OH^{-}(aq)\ Cu(OH)_2(s)$
(e) $Fe^{2+}(aq) + PO_4^{3-}(aq) \rightarrow Fe_3(PO_4)_2(s)$

7. Strong: tap water, sea water, copper sulfate, hydrochloric acid Non-electrolyte: sugar

Set 29: Solution concentrations

1 (a) 1.78 mol L^{-1} (b) 0.571 mol L^{-1} (c) 1.78 mol L^{-1}

2. (a) 0.268 mol (b) 0.280 mol (c) 0.152 mol

3. (a) 59.2 g (b) 19.0 g (c) 55.6 g

4. (a) 0.0100 mol (b) 1.41 mol (c) 0.00889 mol

5. (a) 0.658 mol L^{-1} of K^{+} ion (b) 0.329 mol L^{-1} of K^{+} ion

6. 16 mL

7. 0.956 mol L^{-1}

8. 104 mL

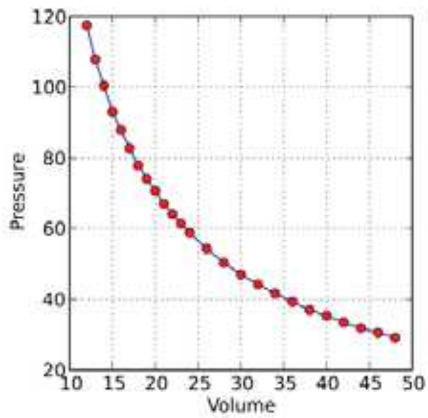
9. 786 mL total final volume. So we need to add 786-150 = 636 mL

10. (a) 1.21 mol L^{-1} (b) 0.484 mol L^{-1}

Set 30: Mixtures

1. Homogeneous mixture uniform composition and properties, heterogeneous mixture non-uniform composition and properties. Homogeneous mixtures: aqueous solution of sodium chloride, air, brass. Heterogeneous mixtures: concrete, orange juice with pulp, sandy water.
2. (a) x axis - time, y axis - temperature/°C. Freezing when the temperature is constant line is horizontal.
   (b) The freezing point of pure ethanol is -114°C labelled on y axis. Pure ethanol contains only one substance - ethanol molecules.
   (c) Homogeneous, water.
3. (a) The different dyes in the ink have different levels of attraction to the paper and different solubilities in water. As a result the dye travel different distances up the paper as the solvent moves up.
   (b) The ink mixture contains 3 components: a yellow, purple and blue dye.
   (c) The blue dye as it has travelled the furthest along with the solvent front.
   (d) An Rf value is a retardation factor or retention value. It is obtained by measuring the distance each dye has moved from where it started and dividing this by the total distance moved by the solvent.
   (e) Blue dye
4. (a) Add water to dissolve sucrose, filter the insoluble sand, evaporate the water from filtrate, crystallise sucrose.
   (b) Add water to dissolve sodium chloride, filter to collect the sand. Rinse with water and dry.
   (c) Distillation.
   (d) Paper chromatography.
   (e) Fractional distillation.
5. In simple distillation boiling points of components are very different, one component with the lower boiling point is removed. Fractional distillation involves the separation of multiple components into fractions with particular boiling points.

Set 31: Kinetic theory



1. Boyle's Law: At constant temperature, the volume of a given quantity of gas is inversely proportional to its pressure: $P_1V_1 = P_2V_2$.
2. Increasing temperature increases the average kinetic energy of the particles leading to more collisions with the walls of the container resulting in a higher pressure.
3. Air particles bombard the inner wall of tyre with greater frequency and force resulting in higher pressure.
4. (a) Helium gas is less dense than air so the balloon will float/rise in air.
   (b) as the balloon rises the air pressure around it decreases, the balloon expands and the pressure inside lowers
   (c) as the balloon rises the external pressure continues to drop and the balloon expands so much that it bursts.
5. Boiling point occurs when the vapour pressure of the liquid being heated reaches atmospheric pressure (pressure above liquid). Atmospheric pressure decreases with altitude so the liquid can boil at a lower temperature.
6. The pressure created by vapour molecules of a substance above its liquid (or solid) form in a closed container. Solute particles interfere with the escape of solvent particles so they require more energy to escape.

Set 32: Gas volumes

1.	(a) 148 L	(b) 19.3 L	2.	(a) 0.198 mol	(b) 1.10×10^{-3} mol		
3.	(a) 51.5 L	(b) 516 L	4.	(a) 1.90 L	(b) 1.51 L	(c) 78.1 L	(d) 0.200 L
5.	94.9 g mol^{-1}		6.	79.3 g mol^{-1}			
7.	1.41 g occupies 1L		8.	(a) 8.98 g	(b) 1.10×10^2 g	(c) 4.45 10^{-3} g	(d) 129 g

Set 33: Stoichiometry and gas volumes

1. (a) 5.68 L (b) 8.52 L (c) 0.341 L

2. 2.27 L 3. 7.81 L 4. 95.0 % pure 5. 97.0 % pure

6. (a) 5557 g (b) 4.50×10^3 L

7. $CH_4 + 2\ O_2\ CO_2 + 2\ H_2O$
(a) V CO_2 = 1576 L (b) V CH_4 = 1576 L

8. m NaCN = 1.06 tonnes

9. (a) $CH_4 + 2\ O_2\ CO_2 + 2\ H_2O$ $2\ C_8H_{18} + 25\ O_2\ 16\ CO_2 + 18\ H_2O$
(b) CH_4: V CO_2 = 291.7 L C_8H_{18}: V CO_2 = 380.3 L
(c) % change: (380.3-291.7) 380.3 = 23.3 % drop in CO_2 output
(d) yes

10. (a) n CO_2 = 0.001664 mol (b) V = 0.03778 L (c) 2.52 %

Set 34: Reacting masses and gaseous and solution volumes

1. (a) $AgNO_3(aq) + NaC\ell\ (aq) \rightarrow AgC\ell\ (s) + NaNO_3(aq)$ (b) m $AgC\ell$ formed = 0.0814 g

2. (a) $2\ C_4H_{10}(g) + 13\ O_2(g) \rightarrow 8\ CO_2(g) + 10\ H_2O(g)$ (b) 189 L (c) 188 g

3. (a) $Na_2CO3(s) + 2\ HC\ell\ (aq) \rightarrow 2\ NaC\ell\ (aq) + CO_2(g) + H_2O(\ell)$ (b) 0.287 L (c) 0.202 L

4. (a) 0.0623 L (b) 0.672 L (c) 2.82 g

5. (a) 0.144 L (b) 15.5 g

6. (a) 0.0408 L (b) 2.44 g

7. (a) 0.189 L (b) 2.29 L (c) 0.534 mol L^{-1}

8. $CaCO_3(s) + 2\ HC\ell\ (aq) \rightarrow CaC\ell_2(aq) + CO_2(g) + 2\ H_2O(\ell)$ 0.509 mol L^{-1}

9. (a) 5.52 g (b) 0.0400 mol (c) 0.454 L 10. (a) 14.2 g (b) 28.4 L

Set 35: Ionic equations

1. $KC\ell\ (s)\ K^+(aq) + C\ell^-(aq)$

2. $Ba(NO_3)_2(s)\ Ba^{2+}(aq) + 2\ NO_3^-(aq)$

3. $NaOH(s)\ Na^+(aq) + OH^-(aq)$

4. $H^+(aq) + OH^-(aq)\ H_2O(\ell)$

5. $ZnO(s) + 2\ H^+(aq)\ Zn^{2+}(aq) + H_2O(\ell)$

6. $CaCO_3(s) + 2\ H^+(aq)\ Ca^{2+}(aq) + CO_2(g) + H_2O(\ell)$

7. $Pb(s) + 2\ Ag^+(aq)\ Pb^{2+}(aq) + 2\ Ag(s)$

8. $Ag^+(aq) + C\ell^-(aq)\ AgC\ell\ (s)$

9. $CO_2(g) + Ca^{2+}(aq) + 2\ OH^-(aq)\ CaCO_3(s) + H_2O(\ell)$

10. $2\ H_3PO_4(aq) + 3\ Ag_2CO_3(s)\ 2\ Ag_3PO_4(s) + 3\ CO_2(g) + 3\ H_2O(\ell)$

11. $H_2S(g) + 2\ Ag^+(aq)\ Ag_2S(s) + 2\ H^+(aq)$

12. $2\ H^+(aq) + CO_3^{2-}(aq)\ H_2O(\ell) + CO_2(g)$

13. $Pb_2+(aq) + 2\ I^-(aq)\ PbI_2(s)$

14. $A\ \ell\ (OH)_3(s) + OH^-(aq)\ [A\ell(OH)_4]^-(aq)$

ANSWERS

Set 36: The pH scale

1. (c)

2.

Indicator added	Colour of solution A	Colour of solution B	Colour of solution C
Conclusion	∴ >12	∴ between 10 - 11	∴ between 6 - 8
	Basic	Basic	Basic, neutral or acidic

3. (a) $[H^+] = 0.100$ mol L^{-1} pH = 1.00 (b) $[H^+] = 0.00500$ mol L^{-1} pH = 2.30 (c) $[H^+] = 2.00$ mol L^{-1} pH = 0.301

4. (a) Lemon juice $[H^+] = 1.00 \times 10^{-3}$ mol L^{-1} (b) Dish washing solution $[H^+] = 1.00 \times 10^{-11}$ mol L^{-1}
(c) Pool acid $[H^+] = 10.0$ mol L^{-1} (d) Orange juice $[H^+] = 2.75 \times 10^{-5}$ mol L^{-1}
(e) Swimming pool water $[H^+] = 2.51 \times 10^{-8}$ mol L^{-1}

5. Concentration changed by a factor of 1000 6. $[H^+] = 1.80$ mol L^{-1} 7. 603 mL

8. 4.23 g citric acid 9. pH = 2.30 10. pH = 5.83 11. 5.31 mL of the hydrochloric acid solution

Set 37: Solutions of acids and bases

1. (a) 128 g L^{-1} (b) 1.21 mol L^{-1} (c) 0.484 mol L^{-1}

2. (a) 0.200 mol L^{-1} (b) 4.04×10^{-3} mol L^{-1}

3. (a) 0.277 mol L^{-1} (b) 0.554 mol L^{-1} 4. 0.025 L 5. 0.870 mol L^{-1} 6. 0.636 L to be added

7. 327 g L^{-1} 8. 1.09 g 9. (a) 0.000135 mol L^{-1} (b) 0.000270 mol L^{-1}

10. (a) $HNO_3(aq) + NH_3(aq) \rightarrow NH_4NO_3(aq) \rightarrow NH_4NO_3(s)$
$H_2SO_4(aq) + 2\ KOH(aq) \rightarrow K_2SO_4(aq) + H_2O(\ell)$
$Ca(OH)_2\ (aq) + 2\ H_3PO_4(aq) \rightarrow Ca(H_2PO_4)_2\ (aq) + H_2O(\ell)$
(b) $m(NH_4NO_3) = 1200$ g
$m\ (K_2SO_4) = 1220$ g
$m\ (Ca(H_2PO_4)_2) = 11.8$ g

11. 0.111 L 12. 12.9 mol L^{-1}

13. (a) add 1.077×10^6 L (b) too much rainwater (c) Add a low cost base

14. add 1.50 L of distilled water

Set 38: Acid and base reaction stoichiometry

1. 0.182 g 2. 3.70 10^{-6} g 3. 46.3 % $CaCO_3$ 4. 0.510 mol L^{-1} 5. 86.1 %

6. (a) 0.4160 mol L^{-1} (b) 102 g 7. 2.70 mol 8. (a) 291 g (b) 4.96 10^{-5} mol L^{-1}