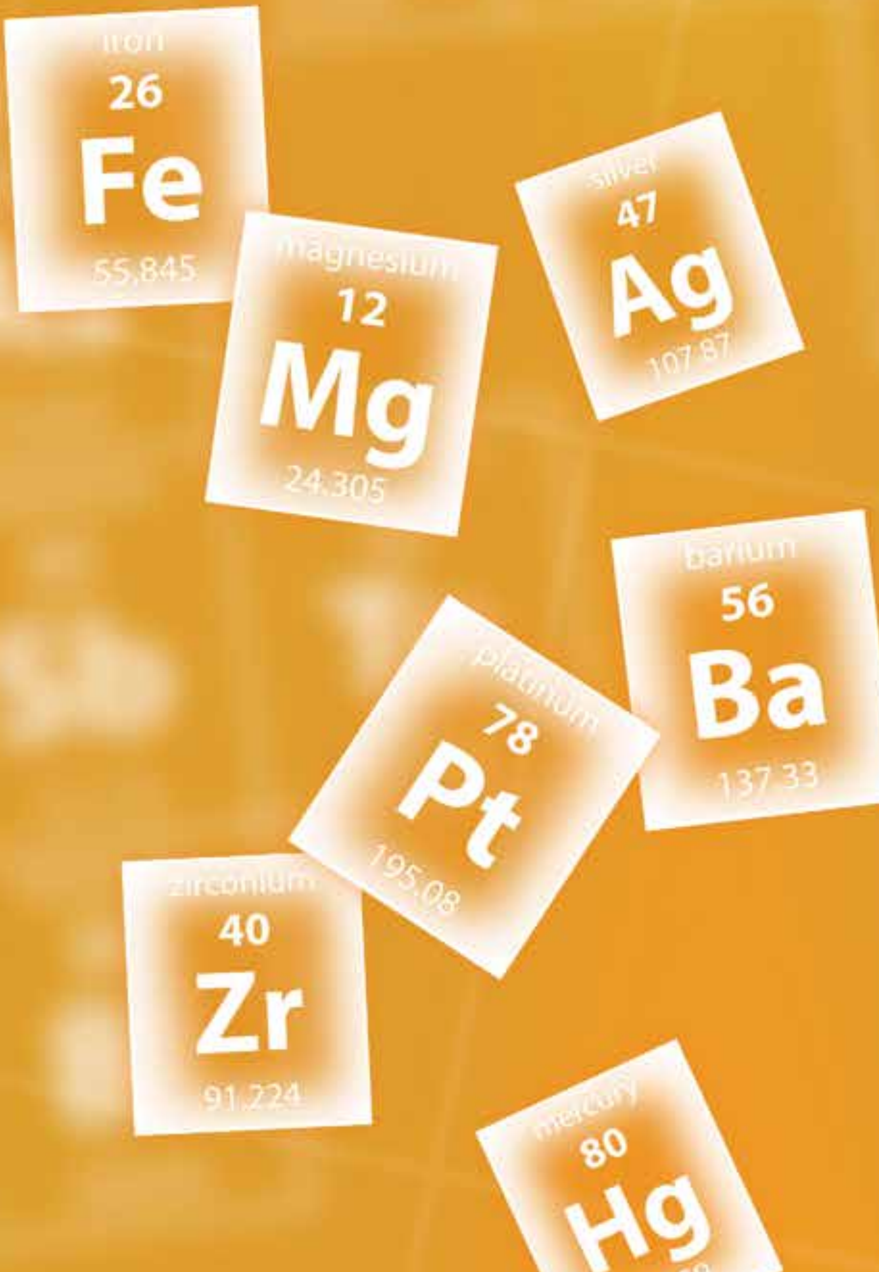
EXPLORING

CHEMISTRY

YEAR 11 - EXPERIMENTS, INVESTIGATIONS & PROBLEMS



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SCIENCE TEACHERS' ASSOCIATION OF WESTERN AUSTRALIA

YEAR 11 ATAR

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YEAR 11 ATAR

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How to use this book

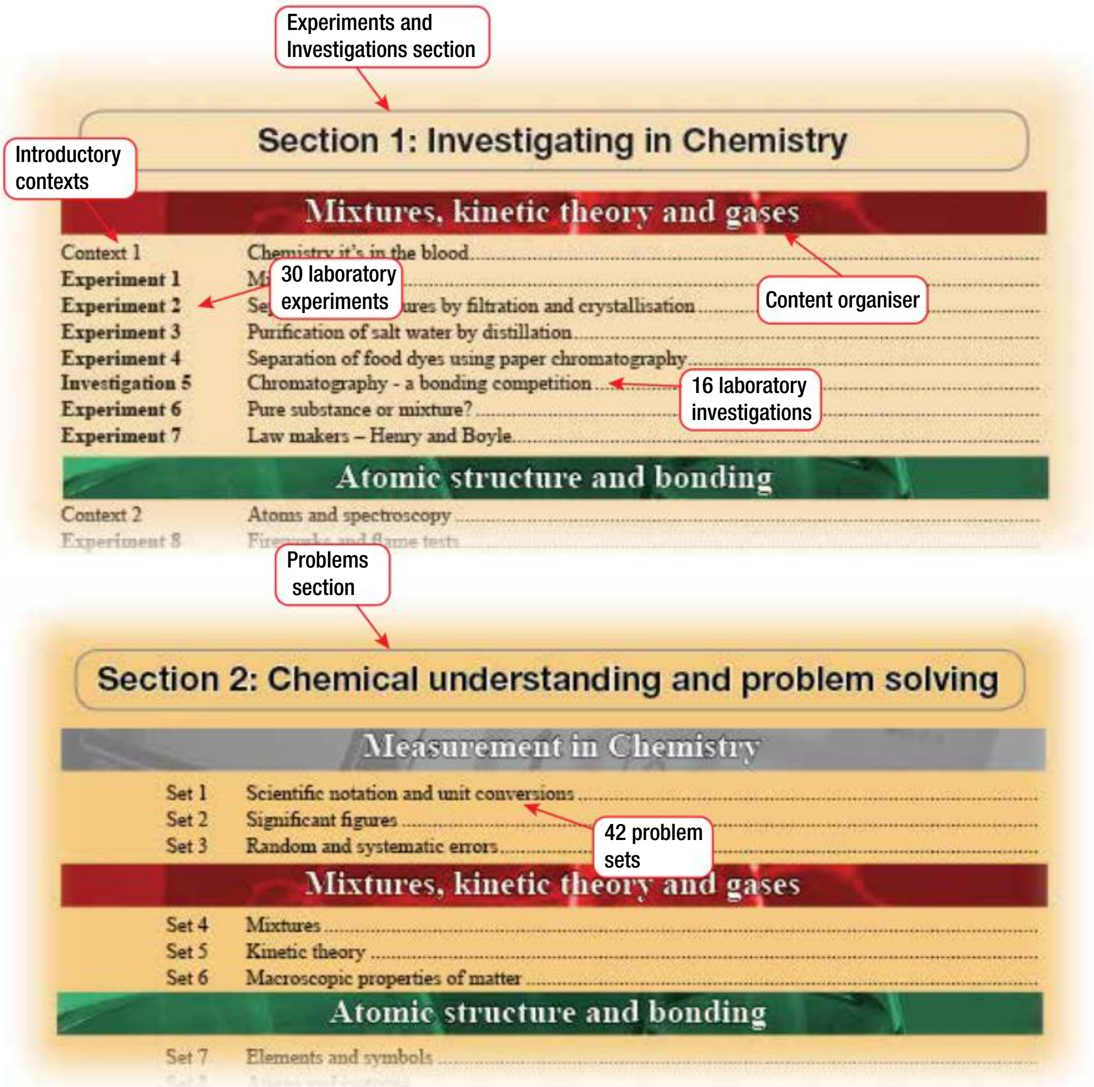
Exploring Chemistry Year 11: Experiments, Investigations and Problems, is divided into two sections, Section 1: Investigating in chemistry and Section 2: Problem-solving and quantities in chemistry.

Investigating in chemistry
Section 1 of the resource contains 30 experiments and 16 investigations helping with exploration and communication of chemical understandings through investigative processes. They represent opportunities to do experiments and to plan and conduct investigations, to analyse the data collected and to evaluate investigation plans, processes and findings.

Problem-solving and quantities in chemistry
Section 2 of the resource contains 42 problem sets requiring the application and understanding of problem-solving techniques to problems in a chemical context.

Content organisers
Exploring Chemistry Year 11: Experiments, Investigations and Problems is organised around six main areas of content: Mixtures, kinetic theory and gases; Atomic structure and bonding; Chemical reactions and stoichiometry; Solutions and acidity; Energy changes and rates of reaction; and Organic chemistry. Each content area is highlighted in a different colour to help navigate the book.

Introductory contexts
There are one or more context passages that introduce each content area. Their purpose is to help highlight the applications of chemistry.



Laboratory safety

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 not in use.
- If the 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, acetone 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.

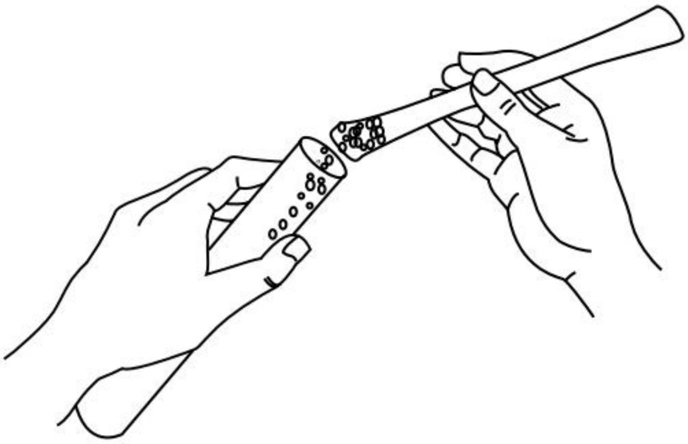
Poisonous gases or vapours

- 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.
- Report mercury spills immediately to the teacher.

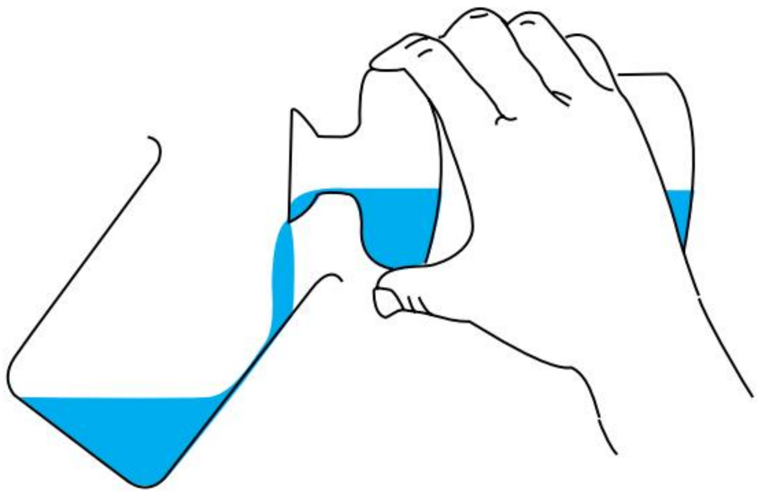
Laboratory techniques

Transferring solids

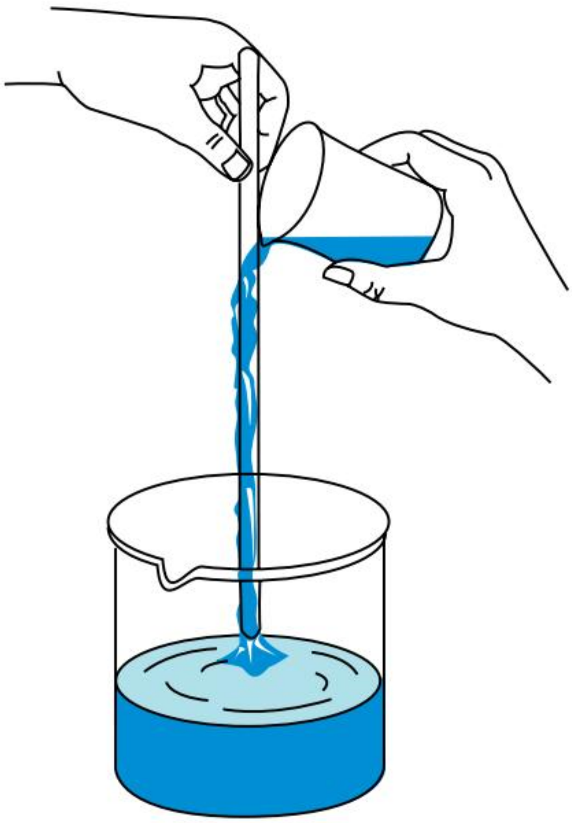
A spatula is used to transfer small quantities of solid material from a stock bottle to a test tube, beaker or weighing bottle. The spatula must be thoroughly cleaned between uses by washing with water and thoroughly drying. Solid material must never be returned to a stock bottle as this may cause contamination.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



Transferring liquids by pouring

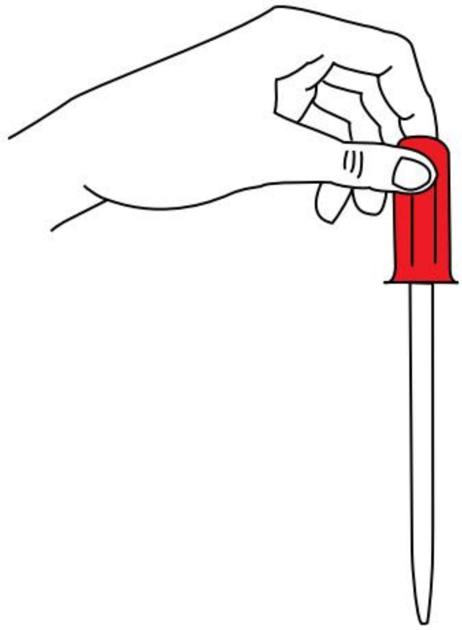


When pouring from a reagent bottle, the receiving vessel (beaker, flask, test tube, etc.) should be tilted so that the liquid can run down the side of the vessel without splashing.



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 dropper



A dropper or Pasteur pipette can be used if a small volume of liquid is required. Sometimes reagent bottles are provided with their own dropper. Be careful not to exchange these droppers between bottles or to allow the droppers to touch the bench or the insides of test tubes or beakers.

If the reagent solution is not provided with a dropper, 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 dropper to obtain the required quantity of liquid from the test tube or beaker.

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

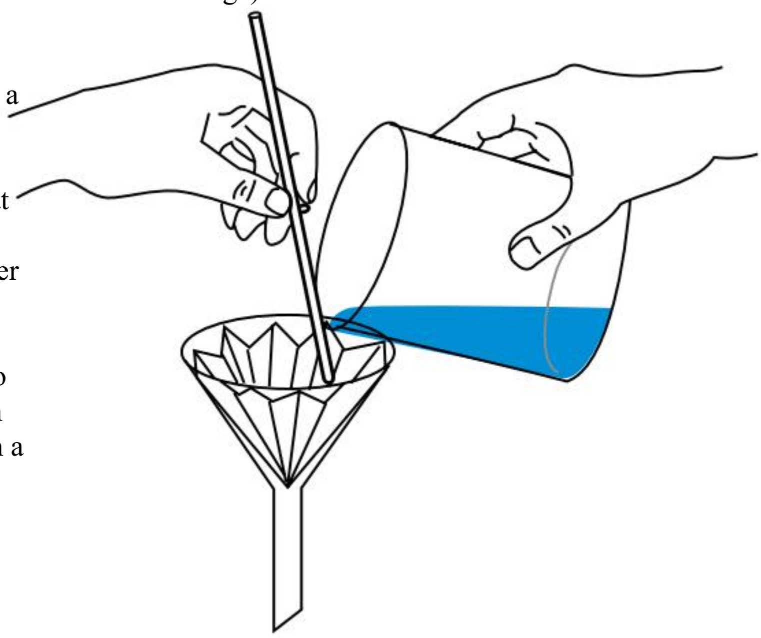
Filtration

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.

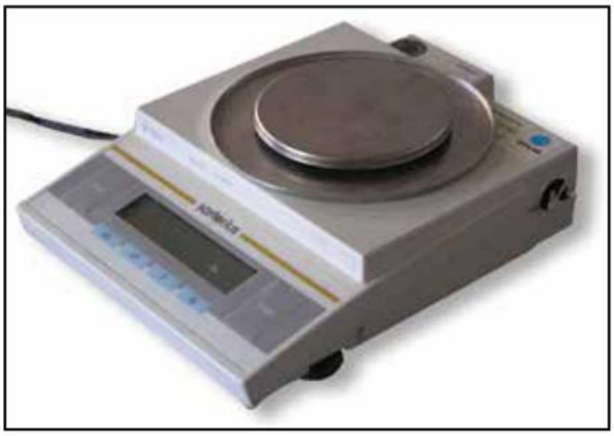
Laboratory techniques

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 may 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 accuracy of the balance being used), or record the balance reading.
2. Carefully add the required amount of substance.

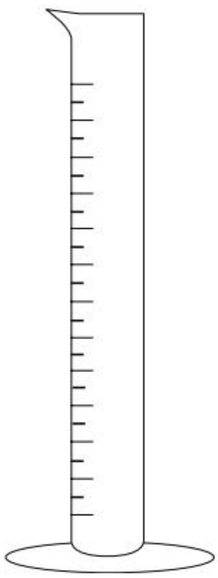


Measuring liquid volumes

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, come 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)



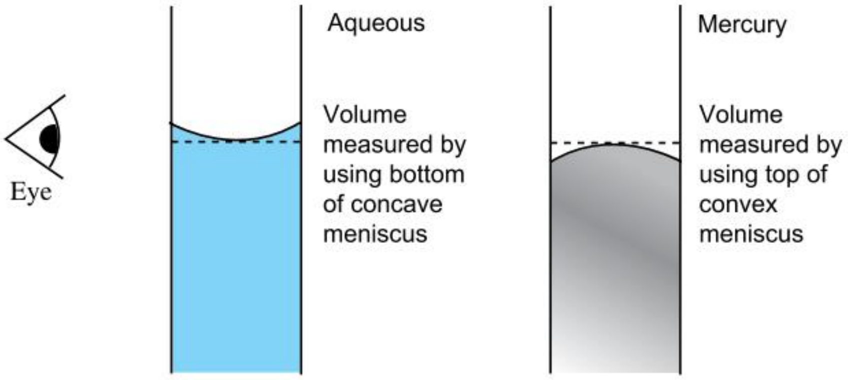
Cleanliness of glassware

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 may cause errors in volume measurement.

If glassware is dirty, it should be cleaned with dilute detergent solution. This is then drained, 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, 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.

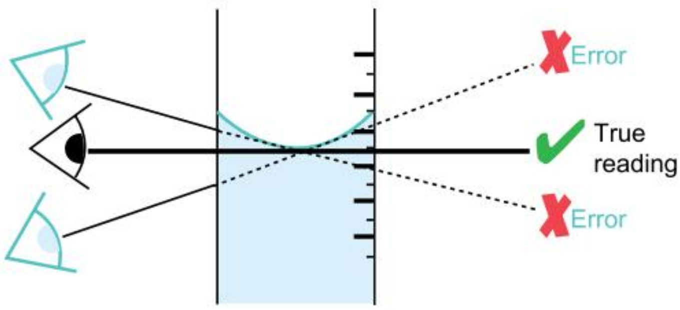


If the liquid is mercury, as in a barometer, the meniscus is convex and the top curved line of the meniscus is used. When using a graduated cylinder or burette, the bottom of the meniscus may not coincide with a particular graduated line. In these cases, where the bottom of the meniscus lies between 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 IX.

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.

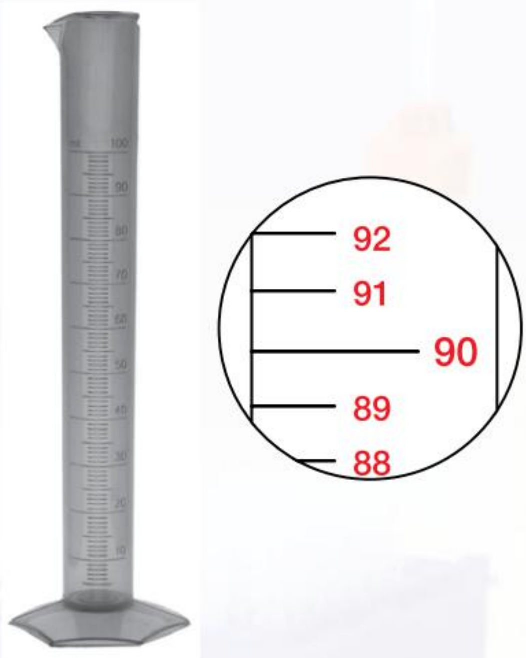
Uncertainty and error

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 come 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: Investigating in Chemistry

Exploring Chemistry Year 11: Experiments, Investigations and Problems **provides opportunities for students to continue developing their science inquiry skills.**

Section 1: Investigating in Chemistry involves identifying and posing questions; planning, conducting and reflecting on experiments and investigations; processing, analysing and interpreting data; and communicating findings.

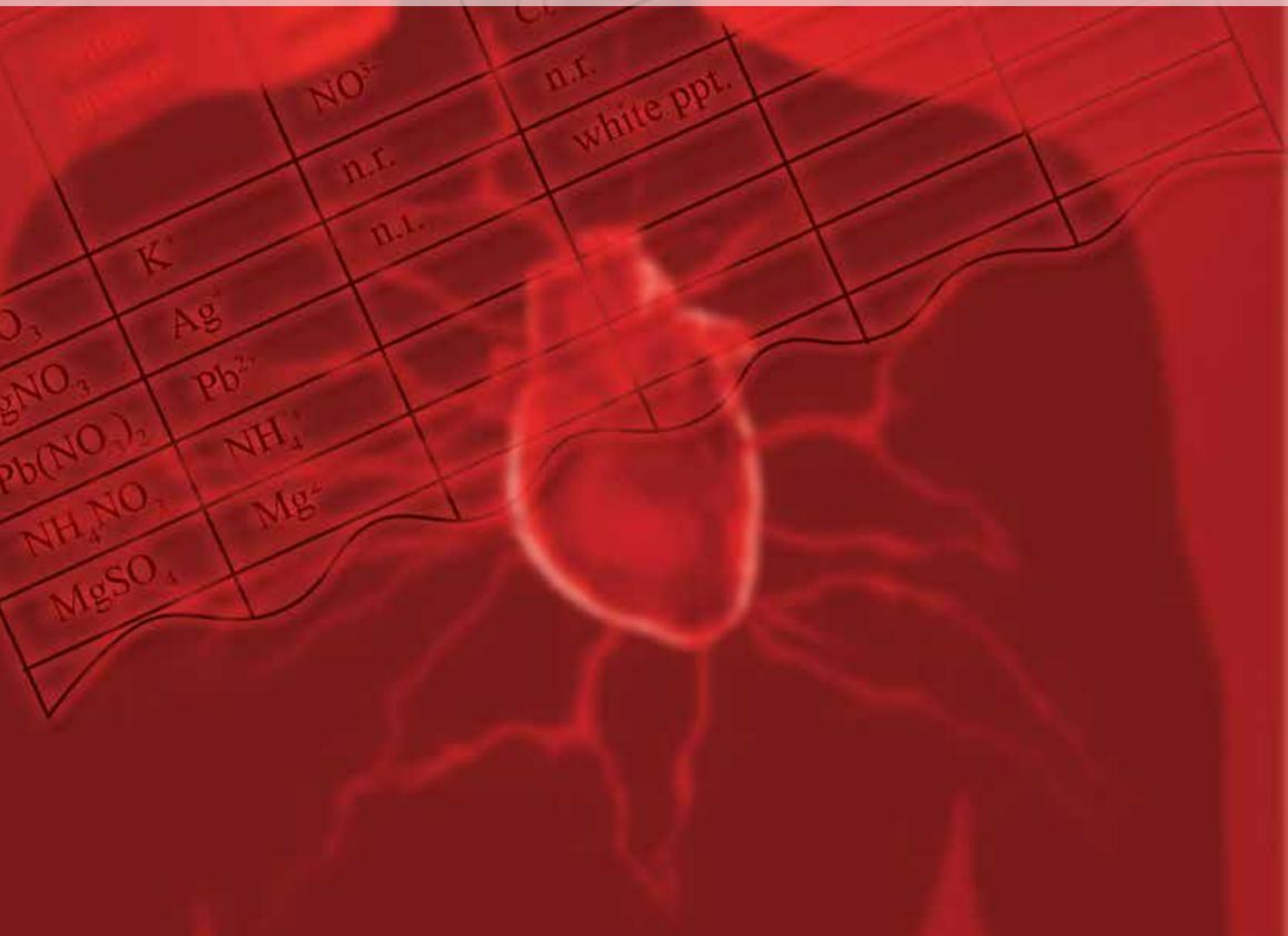
The chemistry experiments and investigations in this book are activities in which chemical concepts are explored and ideas, predictions or hypotheses are tested and conclusions are drawn in response to a question or problem. The collection and analysis of data is an important component of the experiments and investigations. This can involve collecting or extracting information and reorganising data in the form of tables, graphs, flow charts, diagrams, text, keys, spreadsheets and databases.

26
Fe
12
Mg
47
Ag
56
Ba
78
Pt
40
Zr
80
Hg

Mixtures, kinetic theory and gases

The experiments and investigations in the mixtures, kinetic theory and gasses section of *Exploring Chemistry Year 11: Experiments, Investigations and Problems*, provides opportunities for students to explore:

- **pure substances - materials with distinct measurable properties, including melting and boiling points, reactivity and hardness**
- **mixtures - materials with properties dependent on the identity and relative amounts of the substances that make up the mixture**
- **how differences in the physical properties of substances in a mixture, including particle size, solubility and boiling point, can be used to separate them**
- **chromatography techniques and how the data collected can be used to determine the composition and purity of substances**
- **the kinetic theory and how it can be used to explain the behaviour of gaseous systems**
- **the behaviour of an ideal gas, including the qualitative relationships between pressure, temperature and volume**



Chemistry! It's in the blood

The chemistry of our bodies is intriguing. A myriad of different chemicals constantly circulate throughout your body in your blood. These chemicals enable processes such as digestion, cell repair, growth and respiration to take place. For example, where would we be without GnRH, gonadotrophin-releasing hormone? The chemical GnRH is released by the hypothalamus in the brain telling the body to begin puberty.

Some simple yet very special body chemistry involves the blood. Blood is the foremost ingredient of the circulatory system. It functions primarily as a carrier, transporting nutrients and waste chemicals around the body. These chemicals are dissolved in the plasma, which is the liquid part of the mixture in blood. Like all mixtures, blood can be separated into its component materials by physical processes.

The solid components of blood are not visible with the naked eye, but they can be clearly seen using a microscope. As with other mixtures involving suspended matter, the solid components of blood can be separated by filtration. Despite this, separation of the solid components in blood is most commonly achieved by centrifuging (see photo). This involves spinning the blood quickly and letting the different densities of the components do the work.

Centrifuged blood reveals three visible layers - plasma, the buffy coat and the red blood cells. Plasma is the yellow-coloured liquid portion that forms the top half.

The buffy coat is just below the plasma. It is a thin cream-coloured layer of platelets and white blood cells. The red blood cells form the heavy bottom portion. As well as these components, plasma also contains important dissolved electrolytes.

Centrifuge

'Pure' water, 'Pure' apple juice, 'Pure' orange juice! Really?

Only a few substances around the home are actually pure. Have a look at a bottle of 'pure spring water'. The analysis will often state that the water contains chloride, magnesium, bicarbonate and so on. So bottled water is more a mixture than a pure substance. To be pure, a material should include only one type of substance, which could be either an element or compound. Mixtures, like bottled water, contain a combination of elements and/or compounds that are simply added together, but not chemically combined.

Pure apple juice and pure orange juice, like bottled water, are mixtures and not really pure. In general, long life apple juice is a homogeneous mixture with the same clear colour throughout and the same amount of apple juice in all parts of the bottle. Long life orange juice on the other hand is a heterogeneous mixture. The orange juice usually contains some solid pieces of orange, which tend to fall to the bottom of the bottle. As a result its composition is not uniform; the top of the bottle will have less orange than the bottom of the bottle.

Adapted with the permission of WestOne Services (Chemistry Unit 2A -SCIENCE857)

Experiment 1: Mixtures

Notes

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 homogeneous mixtures. How would you classify blood?

During this laboratory activity, you will develop an understanding of different types of mixtures including solutions and examine the different solubilities of some solutes in two solvents: water and ethanol.

Equipment

Bunsen burner
tripod and gauze mat
matches
watch glass
glass rod
thermometer
beakers (2 × 100 mL)
measuring cylinder (25 mL)
sodium chloride solid
100% orange juice
100% apple juice
100% prune juice
solder
fruit scone (or fruit cake)
charcoal
sodium chloride
copper(II) chloride
ethanol
distilled water
boiling beads (or boiling chips)
hand lens or magnifying glass

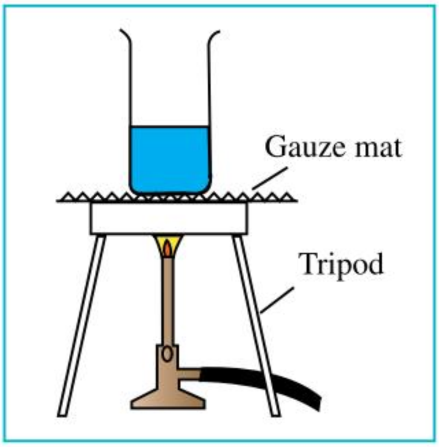


Figure 1.1

SAFETY NOTE:
- **Do not consume fruit juices or fruit scone/cake from this experiment.**

Procedure

Part A: homogeneous or heterogeneous?

1. Examine the 5 mixtures and record your observations in a table similar to the one drawn below.

	Mixture	Observations	Classification
1	100% orange juice		
2	100% apple juice		
3	100% prune juice		
4	solder		
5	fruit scone (or fruit cake)		

2. Classify the mixtures as either homogeneous or heterogeneous.

Notes

Procedure **Part B: solutions**

1. Sodium chloride, has a solubility of 35.9 g/100 mL at 25 °C.
   Determine the mass of salt that you would need to make a saturated solution if only 15 mL of solution was required.

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.

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.*

3. Weigh the calculated mass of sodium chloride in procedure #1 and attempt to dissolve it in 15 mL of distilled water at room temperature.

Observe *- and describe the type of solution that you have created in procedure #2: 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. If necessary, 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.

Experiment 1: Mixtures

Notes

Procedure **Part C: 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 a table similar to that shown.

Solute	Solubility	
	Water (solvent)	Ethanol (solvent)
charcoal		
sodium chloride		
copper(II) chloride		
ethanol		
water		

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

Processing of results and questions

1. Describe how you would make 100 mL of the following types of sodium chloride solution given that the solubility of $NaC\ell$ in water is 35.9 g/100 mL at 25 °C.
   a) unsaturated
   b) saturated
   c) supersaturated

2. Compare temperature records in Part B. Did you manage to make a supersaturated solution? Explain.

3. Ethanol is the alcohol in alcoholic drinks like wine and beer. Ethanol is said to be miscible in water. What does this mean? Is miscibility different from solubility?

Experiment 2: Separation of mixtures by filtration and crystallization

Notes

The separation of a mixture of two solids can often be achieved by filtration and crystallisation. To be successful, this requires that the components of the mixture have different solubilities in a particular solvent. The purpose of this experiment is to separate sodium chloride/charcoal and sodium chloride/copper(II) chloride mixtures.

Equipment

filter funnel stand
Bunsen burner, tripod and gauze mat
beakers (two 100 mL)
boiling chip
distilled water
filter funnel
glass rod
graduated cylinder (25 mL)
ethanol (25 mL)
watch glass
filter paper (Whatman No. 1-three 12.5 cm sheets)
sodium chloride/charcoal mixture (4 g)
sodium chloride/copper(II) chloride mixture (8 g)

Procedure

A Separation of a sodium chloride and charcoal mixture

1. Place 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 with filter paper on a filter funnel stand. Filter the mixture and collect the filtrate in a 100 mL beaker.
3. Wash the solid with a further 5 mL of water but do not add this to the filtrate. Note and 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.

B Partial separation of a sodium chloride and copper(II) chloride mixture

1. Place 4 g of the sodium chloride/copper(II) chloride mixture in a 100 mL beaker and dissolve it in about 15 mL of distilled water.
2. Warm the solution with a Bunsen burner and boil gently until crystals begin to appear in the solution. Turn off and remove the Bunsen burner and cool the solution. Be aware the Bunsen burner barrel will be hot.
3. Filter the solid formed and wash it with about 5 mL of ethanol. Record the appearance of the solid and the filtrate.

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

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 accurately. 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.

Experiment 2: Separation of mixtures by filtration and crystallization

Notes

5. Filter off the undissolved solid using the weighed filter paper. 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 completely (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.

Processing of results and questions

1. What property enabled you to completely separate the sodium chloride/charcoal mixture?
2. In part B you were able to separate some pure sodium chloride but the filtrate contained both sodium chloride and copper(II) chloride. What can you say about the solubility of these two compounds in water?
3. In part C you were able to more completely separate the sodium chloride and the copper (II) chloride using ethanol as a solvent.
    a. Explain why this was possible using ethanol but not when water was used as a solvent in part B.
    b. Calculate the percentage of sodium chloride recovered from the mass of solid in the weighed filter paper
    c. Describe a systematic error in this experimental design that would result in less than 100% recovery of the sodium chloride from the initial mixture.

Experiment 3: Separation of salt water by distillation

Distillation can be used to separate two or more liquids with different boiling points, as 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 saltwater solution.

Equipment

retort stand, boss head and clamp
Bunsen burner
test tubes (two small; two large, one with stopper and delivery tube)
boiling chip (one)
graduated cylinder (25 mL)
sodium chloride solution [$NaC\ell$] 0.1 mol L^{-1} (15 mL)
silver nitrate solution [$AgNO_3$] 0.1 mol L^{-1} (5 mL)
dropper

Procedure

1. Place about 2 mL of 0.1 mol L^{-1} $NaC\ell$ into a test tube, add five drops of 0.1 mol L^{-1} $AgNO_3$ solution and record your observations.
2. Place about 10 mL of 0.1 mol L^{-1} $NaC\ell$ into a large test tube fitted with a delivery tube as shown in Figure 3.1. Clamp the test tube to a retort stand.
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, add to it five drops of 0.1 mol L^{-1} $AgNO_3$, and record your observations.

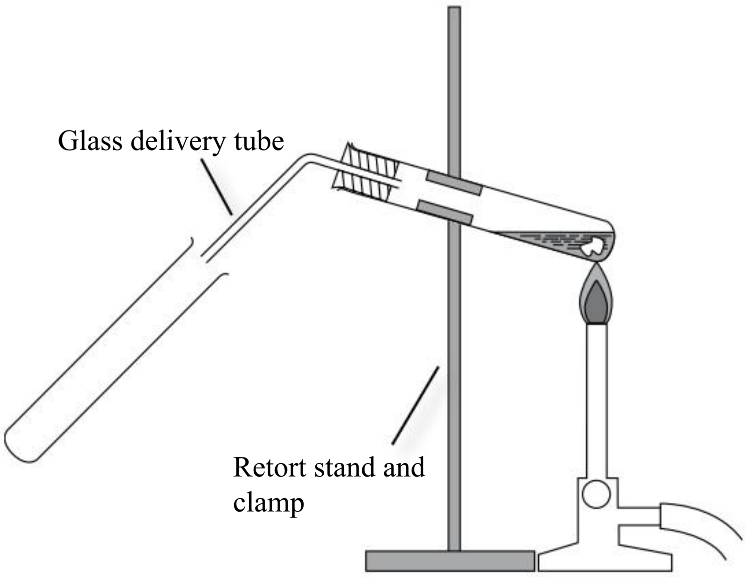


Figure 3.1

Notes

Processing of results and questions

1. Name the precipitate formed when silver nitrate solution was added to the sodium chloride solution. Write an equation for the reaction.
2. What did testing the distillate with silver nitrate solution indicate?

Experiment 4: Separation of food dyes using paper chromatography

Notes

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 food dyes. In this technique the components that are least soluble in water and those that adhere most strongly to the paper will move more slowly than those that are more soluble in water or adhere less strongly to the paper.

Equipment

test tubes (five large)
capillary tube
beaker (100 mL)
sodium chloride solution [$NaC\ell$] 1% solution (20 mL)
assorted food colourings - especially black, green, blue, yellow, red (1-2 drops) or
assorted Smarties® or M&M's®, including black and brown (one of each colour)
strips of chromatography or filter paper to fit into the large test tubes (five)
distilled water

Procedure

1. If Smarties® or M&M's® are being used, place each of the Smarties® or M&M's® in turn into a small beaker, add about five drops of water and swirl the contents until the colour dissolves. Avoid dissolving any underlying sugar.
2. With a capillary tube containing one of the food colourings or Smarties® or M&M's® dyes draw a fine line of dye across a strip of filter paper about 2 cm from one end.
3. Similarly prepare four more filter paper strips with four other food colourings or dyes.
4. Prepare five large test tubes containing 1% $NaC\ell$ solution to a depth of about 1 cm.
5. Place the paper strips into the test tubes but ensure that the level of the liquid is below the line on the paper and that the paper strip does not cling to the sides of the test tube. Bend the top of the paper strip over the mouth of the test tube to keep it in position.
6. Allow the chromatograms to develop.
7. When the liquid nears the top of the paper strip remove the strip from the test tube and record your observations.

Processing of results and questions

1. Identify those food colourings or dyes that consisted of a single component, and those that contained more than one component.
2. What property of the food colourings enables them to be separated in this way?

Investigation 5: 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. It 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.

Notes

The task

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

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
- kerosene

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)			

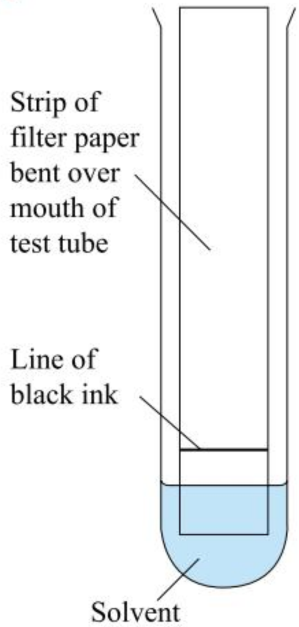


Figure 5.1

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 the chemicals and equipment you need and identify safety requirements.

3 Check the proposed procedure with your teacher.

Investigation 5: Chromatography - a bonding competition

Notes

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.

Experiment 6: Pure substance or mixture?

Notes

Our blood is 83% water; our muscles are about 75% water; our brain 74%, while our bones are about 23% water. Although most people don't think of it as a nutrient, water is the most vital nutrient in our daily diet. You can survive for weeks without food but you can only survive a few days without water.

Good drinking water is a homogeneous mixture, a solution, containing important nutrients. Sometimes it is important, particularly in scientific research experiments, to use pure water. Pure water obtained by distillation is called distilled water, and water obtained by the removal of ions is called de-ionised water. So how do we tell the difference between pure water and a water solution? In general how do we tell the difference between a pure substance and a mixture?

During this laboratory experiment, you will further develop your understanding of the differences between pure substances and mixtures by examining freezing points.
You can do this experiment using one of two methods:
Method A: using a glass thermometer
Method B: using a digital temperature sensor

Equipment

Method A: using a glass thermometer

test tubes (two large)
graduated cylinders (10 mL and 25 mL)
beakers (100 mL and 1 L)
thermometer (-10 to 110 °C)
stopwatch
glass rod
distilled water
glycerol
crushed ice (about 500 g)
household salt (200 g)

Procedure

Method A: using a glass thermometer

1. Fill a 1L beaker with crushed ice and add about 50 g of household salt so that it is well mixed with the ice.
2. Measure 10 mL of distilled water into a large test tube and place a thermometer in the water.
3. Place the test tube into the ice/salt mixture and quickly pack the ice around the test tube. Note the temperature and immediately start the stopwatch. The first temperature reading is recorded at time zero.
4. Continue to take temperature readings at 30 second intervals, while gently stirring with the thermometer. Take readings until all the water has solidified and the temperature starts to fall again. Take care not to snap the thermometer.
5. Into a 100 mL beaker make up a solution containing 16 mL of water and 4 mL of glycerol. Thoroughly mix the solution. Measure out 10 mL of this solution into a large test tube and place a thermometer in it.
6. Place the test tube into the ice/salt mixture and follow the same procedure as used previously with the water.

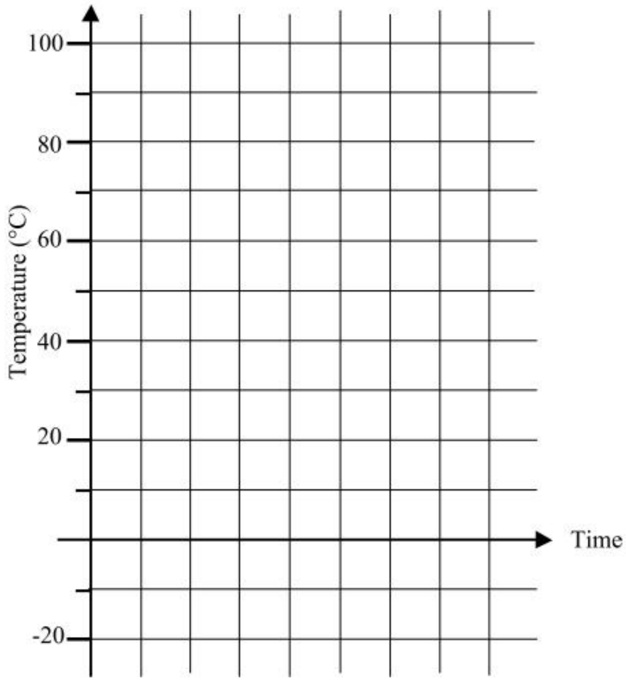
Experiment 6: Pure substance or mixture?

Notes

Processing of results and questions

Method A: using a thermometer

1. Plot both cooling curves on a piece of graph paper and identify the freezing points of water and your glycerol/water solution.



2. Describe any differences between the cooling curve of distilled water (pure substance) compared to a water solution (mixture).
3. Why was salt added to the ice?

Equipment

Method B: using a temperature sensor

In addition to the equipment required for Procedure A, the following will be needed:
computer or graphic calculator loaded with temperature program
suitable computer/calculator interface
two temperature sensors capable of measuring to -10 °C
laboratory tray
permanent marker

Procedure

Method B: using a temperature sensor

1. Ensure that both 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.

SAFETY NOTE:
- **To protect the electrical equipment from water, perform the experiment in a laboratory tray.**

2. Prepare an ice-bath at a temperature of approximately -10 °C. To do this, place a 1 L beaker in a laboratory tray, half fill the beaker with crushed ice and add about 50 g (2 tablespoons) of salt. Use a stirring rod to mix in the salt. Check the temperature with a thermometer or a temperature sensor. If necessary, add more salt to lower the temperature.

3. If the temperature sensors have not been calibrated, it will be necessary to calibrate them before starting. Calibrate the temperature sensors at temperatures of approximately -10 °C and 20 °C by following the calibration instructions in the software program. Use the ice-bath to determine the lower fixed point and place some water in a 100 mL beaker to determine the upper fixed point.

4. Identify which temperature sensor is responsible for each set of data on the screen. This can be done by warming one of the sensors with your hand and observing the response on the screen.

5. Using a 100 mL beaker make up a solution containing 16 mL of distilled water and 4 mL of glycerol. Thoroughly mix the solution. Measure out 10 mL of this solution into a large test tube.

6. Into the other large test tube place 10 mL of distilled water. Label the test tubes with the permanent marker.

7. Place the sensors in the test tubes, making sure you know which sensor is in each tube.

8. Record the temperatures in the test tubes until the first points appear on the screen (15 to 25 s). Then place both test tubes containing the sensors into the ice-bath.

9. Gently stir the sensors in the test tubes until the water and the glycerol/water solution start to freeze. Do not attempt to stir when frozen because you might break the sensors.

10. Allow the experiment to run for about 15 minutes.

11. Print out the graph if possible.

Processing of results and questions

Method B: using a temperature sensor

1. Either plot both cooling curves on a piece of graph paper or obtain a printout of the graph. From the graph identify the freezing points of the water and the glycerol/water solution.

2. Describe any differences between the cooling curve of distilled water (pure substance) compared to a water solution (mixture).

3. Why was salt added to the ice?

Notes

Experiment 7: Law maker - Robert Boyle

Notes

Pressure is not measured in kilograms. For this experiment by using a constant diameter syringe we can use kg to indicate pressure.

Irish chemist Robert Boyle in 1662, investigated the relationship between gas pressure and gas volume.

In this experiment you will investigate Boyle's Law. The forces between gas molecules are relatively weak, so the molecules have very little effect on one another and the gas molecules are so far apart that the volume of the gas molecules is negligible compared to the distance between them. For this reason the relationship observed between the pressure and volume of a sample of air will be general for all gases. A fixed amount of air trapped in a syringe is used and the pressure applied to the gas is increased, by pushing in the syringe. The volume of air will be recorded for each value of pressure. You will graph the values of pressure and volume to determine a mathematical relationship between pressure and volume.

This experiment may be performed either by using a graduated syringe with a sealed end and a set of masses (Method A) or by using a pressure sensor (Method B) that is interfaced to a computer or a graphic calculator.

Equipment **Method A: using a graduated syringe**

graduated syringe with sealed end - about 50 mL
set of masses (six 1 kg or twelve 500 g masses)
stand and clamp
balance

Procedure **Method A: using a graduated syringe**

1. Make a table of your results as shown below.

Pressure due to masses (kg)	Volume (mL)			Average volume (mL)	Total pressure (kg)	$\frac{1}{V}$ (mL^{-1})	PV
	Trial 1	Trial 2	Trial 3				
0							
1							
2							

2. Weigh the plunger on an accurate balance and record its mass.

3. Adjust the plunger so that it is just below the highest calibration on the syringe and ensure that the syringe is air-tight. Clamp the syringe in an upright position and read the volume of air trapped in the syringe as precisely as possible.

4. Carefully place a kilogram mass on top of the plunger and again read the volume of air in the syringe. Remove the mass and place it back on twice more so as to record three volume measurements.

5. Repeat procedure # 3 using a mass of 2 kg, and continue increasing the masses until a mass of 6 kg has been added.

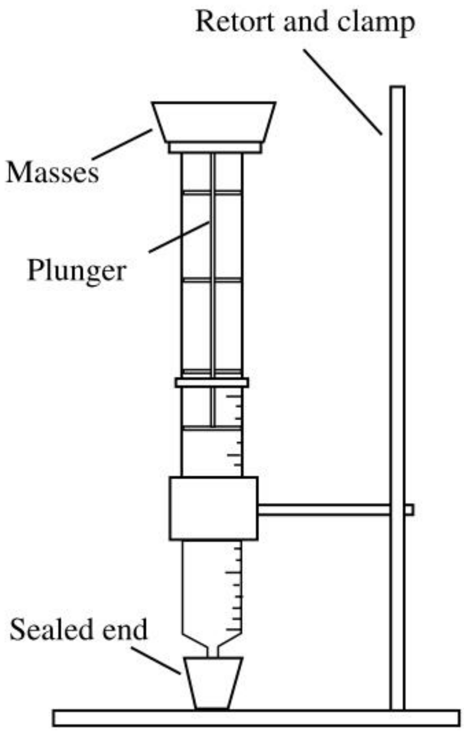


Figure 7.1

SAFETY NOTE:
- **The syringe may fail under pressure.**
- **Wear safety glasses.**
- **Use a method to support the weights so that they cannot topple from the top of the syringe.**

Experiment 7: Law maker - Robert Boyle

Notes

Processing of results and questions

Method A: using a graduated syringe

To calculate the pressure acting on the gas in the syringe it is necessary to take account of atmospheric pressure, the mass of the plunger and the masses placed on top of the plunger. You will use the sum of the masses of the plunger, the added masses and the air column above the syringe as a measure of the pressure exerted on the gas. This is a reasonable procedure because all these masses experience the same gravitational acceleration and act on the same cross-sectional area of the syringe.

1. To calculate the mass of air acting on the syringe you need to multiply the mass per unit area of an air column above the earth (1.03 kg cm^{-2}) by the cross sectional area of the syringe.

   mass of air acting on syringe
   =(mass/unit area of air column) × (cross sectional area of syringe in cm^2)
   = (1.03 kg cm^{-2}) × (πr^2)

2. Calculate the total pressure (in kilograms) in each case by calculating the total mass acting on the syringe. Make sure all your masses are expressed in kilograms. Record these values in the table.

   pressure = mass of air column + mass of plunger + added masses (in kg)

3. Using the data from the various trials, determine the average gas volume for each pressure and record these values in the table.

4. Draw a graph of these results, plotting the pressure (in kilograms) on the vertical axis and the average volume (in millilitres) on the horizontal axis. Draw a smooth curve of best fit through the points.

5. Use the average volumes to determine values of $\frac{1}{V}$. Record these values, to the appropriate number of significant figures, in the table.

6. Draw another graph, this time plotting the pressure (in kilograms) on the vertical axis and values of $\frac{1}{V}$ on the horizontal axis. Rule a line of best fit through the points.

7. Calculate the values of the products of total pressure and average volume and record these in the table.

8. From your results what can you conclude about the relationship between the pressure and volume of a gas?

9. Why do you think that the values of the PV products were not exactly the same?

10. Under what circumstances would the gas volume approach zero?

Equipment **Method B: using a pressure sensor**

computer or graphic calculator loaded with suitable software
computer or graphic calculator interface
pressure sensor
graduated syringe with sealed end

Procedure **Method B: using a pressure sensor**

1. Ensure that the pressure sensor is connected to the interface and that the interface is connected to the computer or graphic calculator.
2. Check that the program has been loaded and has been set up with pressure on the vertical axis and volume on the horizontal axis.
3. Check that the pressure sensor has been calibrated. If this needs to be done, follow the software instructions. Assume atmospheric pressure is 1.00 atmosphere. Now take pressure-volume readings following the procedures in steps 4 and 5.
4. With the air valve open, adjust the syringe to its largest volume. Close the air valve and enter on the screen the volume of gas in the syringe. Also enter the air pressure reading of 1.00 atmosphere.
5. Take at least four more readings. Do this by pushing the syringe in to decrease the volume. Hold the syringe steady and record the pressure and volume readings. Repeat this until you have at least five pressure-volume readings.
6. Examine the graph of pressure versus volume. Decide whether the relationship is direct or inverse. Then, using the curve of best fit procedure, draw the graph.
7. Obtain printouts of the graph and the data table.

Processing of results and questions **Method B: using a pressure sensor**

1. Use the values from the data table to calculate PV (pressure x volume) for each pair of readings and comment on the results.
2. Using a computer spreadsheet or a graphic calculator, plot pressure (P) versus the reciprocal of volume ($\frac{1}{V}$) to confirm the relationship between pressure and volume. To do this you will to create a new data column ($\frac{1}{V}$). Enter the formula ($\frac{1}{V}$) and the appropriate units (mL^{-1}). Change the axis of the graph and print the graph.
3. What conclusion about the relationship between the pressure and volume of a gas can you make from the results?
4. Under what conditions would it be possible to obtain a gas volume that approaches zero?

Notes

If: $PV = k$

then: $P = \frac{k}{V}$

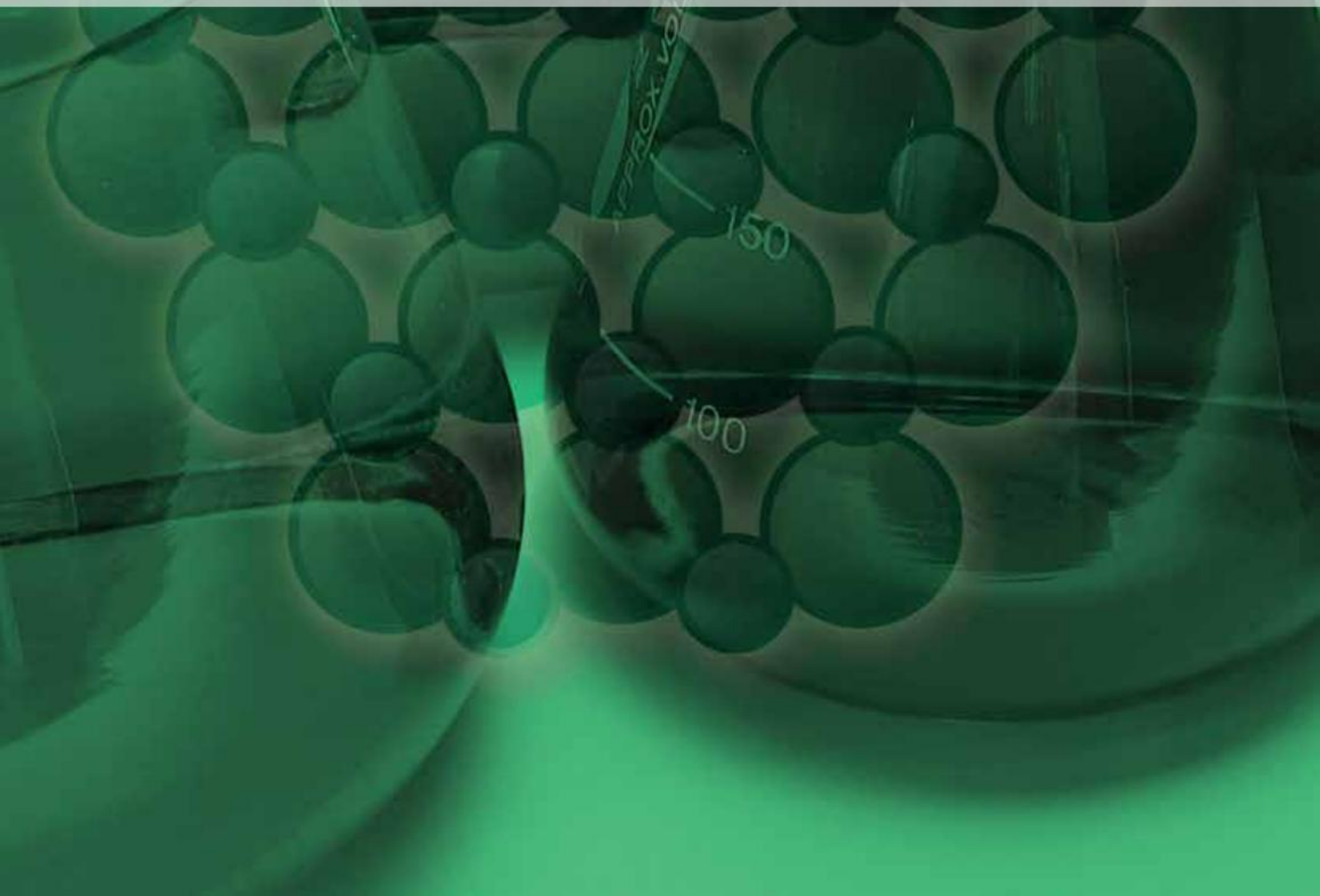
By plotting P vs $\frac{1}{V}$

we can determine k, the slope of the graph.

Atomic Structure and Bonding

The experiments and investigations in the atomic structure and bonding section of *Exploring Chemistry Year 11: Experiments, Investigations and Problems*, provides opportunities for students to explore:

- **how the atomic model and models of bonding explain the structure and properties of elements and compounds**
- **how atoms can be modelled as a nucleus, surrounded by electrons in distinct energy levels, held together by electrostatic forces of attraction between the nucleus and electrons; the location of electrons within atoms can be represented using electron configurations**
- **analytical techniques such as flame tests and atomic absorption spectroscopy (AAS) and how they can be used to identify elements**
- **how observable properties, including vapour pressure, melting point, boiling point and solubility, can be explained by considering the nature and strength of intermolecular forces between covalent molecular substances**
- **the type of bonding within ionic, metallic and covalent substances and how it explains their physical properties, including melting and boiling points, hardness, conductivity of both electricity and heat**



Atoms and spectroscopy

Throughout history, humans have pondered the question: From what and how did our Universe begin?

Our best current theory suggests that it all began some 15 billion years ago from a massive expansion of very dense matter – the **Big Bang theory.**

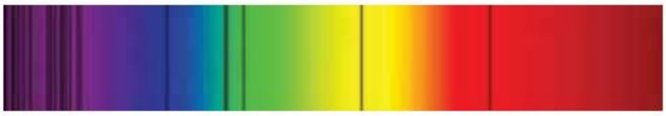
Astronomers have discovered that our Universe is expanding, and that it contains billions of galaxies filled with stars. They have inferred the presence of, black holes, gas clouds, neutron stars and planets. Astronomers seek to understand the evolution of our Universe, and what it is made of.

A revolutionary new radio telescope, the Square Kilometer Array (SKA) will utilize radio astronomy to add to our understanding of the evolution of the Universe.

The **emission spectrum** of hydrogen matches some of the black lines in the absorption spectrum of the Sun. In this case, the lines tell us that hydrogen atoms exist in the outer layers of the Sun, and they absorb some of the light as it passes through the outer layers. Such absorption lines have also been used to identify other elements in the outermost layer of the Sun and other stars. Sodium emits strongly in the yellow part of the spectrum. The corresponding dark absorption line in the yellow part of the solar spectrum tells us that sodium atoms exist in the Sun's outer layers.

Experiments involving emission spectra of electromagnetic radiation from gases provide evidence that the electrons in an atom have discrete energy values.

Hot bodies like the Sun produce a continuous spectrum of colour. Closer examination of the Sun's spectrum (below) shows dark lines. Such an absorption spectrum is created by atoms absorbing energy. These dark lines, which appear throughout the solar spectrum, are called Fraunhofer lines because they were discovered by Joseph von Fraunhofer in 1814. Absorption lines appear at specific frequencies and are the result of the absorption of energy by elements in the outer layers of the Sun.



Visible spectrum of the Sun showing Fraunhofer absorption lines

Monitoring Pollution

A device known as an atomic absorption spectrophotometer has been found to be highly effective in monitoring pollution, particularly soil and water contamination. The device is able to measure very, very small concentrations of a wide selection of elements, particularly metals. It measures the absorption spectra of elements.

Atomic absorption spectroscopy, AAS, was developed in Australia by a team of CSIRO scientists led by Alan Walsh in the 1950s. It is now used throughout the world for accurate analysis in the parts per million (ppm) range.

AAS works by pumping an aqueous solution of a sample into a flame. The flame has light of a specific wavelength (colour) shining through it. Some of the light is absorbed by the sample. By measuring the amount of light that is absorbed, the concentration of the element that absorbs it can be determined.

Experiment 8: Fireworks and flame tests

Fireworks

Skyshows explode around the country to celebrate Australia day.

The characteristic colours of **fireworks** are due to particular metal elements. In this case, the energy to excite the atoms 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 '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. The purpose of this experiment is to examine the colour produced by different salts when they are flame tested.

Notes

Equipment

watch glass
matches
ethanol (10-20 mL)
tooth picks

small samples (pea sized) of salts such as:

copper(II) sulfate [$CuSO_4$]
potassium nitrate [KNO_3]
sodium chloride [$NaC\ell$]
barium nitrate [$Ba(NO_3)_2$]
iron(II) sulfate [$FeSO_4$]
strontium nitrate [$Sr(NO_3)_2$]
magnesium nitrate [$Mg(NO_3)_2$]
lithium nitrate [$LiNO_3$]
calcium nitrate [$Ca(NO_3)_2$]

Procedure

1. Place a small amount of the salt in a clean watch glass.
2. Add 1-2 mL of ethanol to the sample and stir with a tooth pick.
3. Place the watch glass on a heat proof mat.
4. Carefully light the ethanol.
5. When the flame burns low observe the colour produced.
6. Repeat for each salt using a clean watch glass each time.

Processing of results and questions

1. What component of the salts is responsible for the colours observed?
2. What assumption is made about the ethanol in the flame test?
3. Why must a clean watch glass be used for each test?
4. Explain how the coloured light observed is produced.
5. Give a practical use of this type of test.
6. 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.

SAFETY NOTE:
- **Ethanol is flammable and it is essential that amounts used are as indicated. Keep the stock solution away from any naked flame.**
- **Do this experiment on a heatproof mat.**

Experiment 9: Emission spectra

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. 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.

When an element is heated the electrons in their ground state absorb energy and move to higher energy levels. When they 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. We see this in Figure 9.1 and 9.2

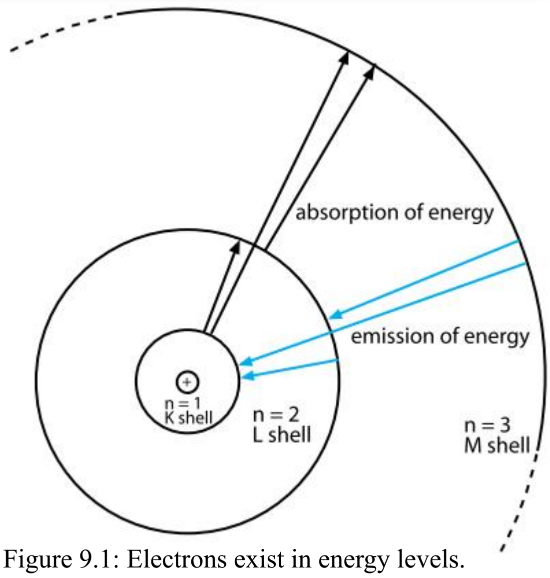


Figure 9.1: Electrons exist in energy levels.

The energy differences between energy levels determine the wavelengths of the electromagnetic radiation that can be absorbed or emitted; and the energy associated with each level is unique to the element involved. Thus, the emission spectrum produced is unique to atoms of a particular element. A popular metaphor is that the spectrum of an element is its 'fingerprint'.

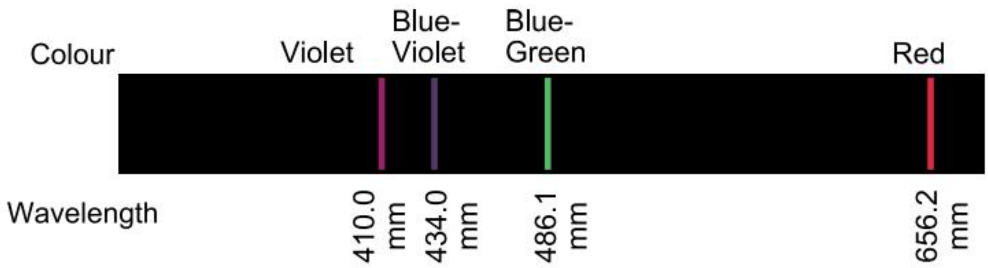


Figure 9.2: Emission spectrum for hydrogen in the visible region of the electromagnetic spectrum.

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**

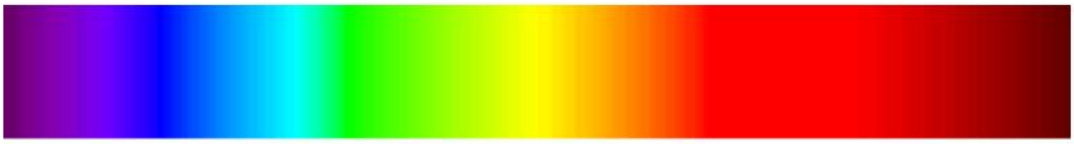
Notes

Experiment 9: Emission spectra

Procedure

1. Use the spectroscope to observe the spectra of natural daylight. 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 a table like the one below.

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

Processing of results and questions

Notes

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

2. Explain why what you observe with natural daylight is referred to as a continuous spectrum.

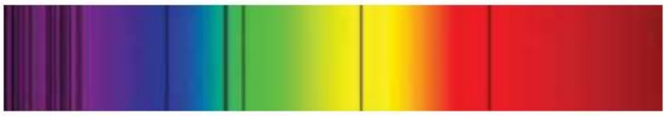
3. Explain why the room has to be darkened for viewing the light emitted from the gas discharge tubes.

4. How do the emission spectra from the gas discharge tubes compare with the spectrum of visible light?

5. Compare and contrast the emission spectra observed from different gas discharge tubes.

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

7. Atomic absorption spectroscopy is a technique related to atomic emission spectroscopy but in this case the specific wavelengths of light absorbed by the sample are detected. You observed the gaps in the spectrum of light with black lines indicating wavelengths absorbed.



Visible spectrum of the Sun showing Fraunhofer absorption lines

Explain how this technique can also be used to identify elements.

Experiment 10: Metal crystals

Notes

Equipment

Per group:
stereo microscope (with microscope lamp if not self-illuminated) – dark stage plate gives best results
Petri dish, plastic
dropper bottle of 0.1 mol L^{-1} copper (II) sulfate
dropper bottle of 0.1 mol L^{-1} tin (II) chloride
sandpaper

Additional materials for Part A - Teacher demonstration:
copper foil, wire or strip, zinc strip
dropper bottle of 0.1 mol L^{-1} silver nitrate solution

Procedure

Part A - Teacher demonstration: Copper

1. Clean the copper foil with sandpaper and lay it in the Petri dish.
2. Add a drop of silver nitrate solution along the edge of the metal. Do not flood the slide.
3. Place the Petri dish under the microscope and carefully focus on the edge of the metal where it is touching the solution. Use highest power and overhead illumination for best results. Remember there is no cover slip on the liquid so be careful not to touch the microscope lens on the surface of the liquid.
4. Observe for a few minutes and describe the shape of the crystals that form.

SAFETY NOTE:
- **Silver nitrate causes burns and stains skin and clothing.**
- **Copper(II) sulfate and tin(II) chloride are irritants to skin and eyes.**

Clean up
Place Petri dishes with copper and silver nitrate into a waste container filled with water.

Procedure

Part B: Zinc

1. Clean the zinc strip with sandpaper and lay it in the Petri dish.
2. Add a drop of copper (II) sulfate solution along the edge of the metal. Do not flood the slide.
3. Place the Petri dish under the microscope and carefully focus on the edge of the metal where it is touching the solution. Use highest power and overhead illumination for best results. Remember there is no cover slip on the liquid so be careful not to touch the microscope lens on the surface of the liquid.
4. Observe for a few minutes and describe the shape of the crystals that form.
5. Repeat steps 1 to 4 using tin (II) chloride solution and compare the crystal shapes that form over time.

Investigation 11: Desktop materials

The task

Notes

Your science teacher wants to put a new surface on the desks in the school laboratory. She is undecided whether to simply sand and repaint them or whether to stick a laminated surface on to the desks. Your class discusses the necessary requirements of the desk surface and decide that testing the paint and laminate to see which will be most durable would be a good idea. The teacher gives you samples of the laminated surface and a piece of wood painted in the recommended paint. What sorts of tests would you recommend should be done to the samples to decide which is the most suitable surface? Ensure you state how each test is conducted and any safety precautions that should be taken.

Planning the investigation

1. Plan an investigation to determine the most suitable surface for the desks in your laboratory.
2. Make a list of the types of physical activity the surface must withstand.
3. Make a list of the sorts of solvents that might be spilled on the surface that could potentially damage it.
4. Write out your proposed procedure describing each test, list the chemicals and equipment you need and identify any safety requirements to complete each test.
5. Check the proposed procedure with your teacher.

Conducting the investigation

1. Conduct the investigation, collecting and recording the observations of each test in a table as you proceed.
2. If time allows, replicate the data collection.

Processing the data

1. Which material will you choose and why?
2. Why is it important to test materials that consumers use?

Evaluating the investigation

1. 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.
2. Discuss whether the data are sufficient to support your conclusion.

Investigation 12: Physical properties and structure

Wax candle

Notes

What makes substances different and determines their properties?
Copper is a metal, silicon dioxide is a covalent network substance, sodium chloride is an ionic substance and candle wax is a covalent molecular substance. Their properties are characteristic of the class of material that they represent. Materials are selected for particular applications based on their properties.

We use a metal such as copper for electrical wiring. Copper is a good electrical conductor and is ductile (can be drawn out as a wire). Silicon dioxide (sand), on the other hand, is a non-conductor of electricity is brittle and can be formed into glass. Like silicon dioxide, common table salt (sodium chloride), will not conduct electricity when in solid form. Candle wax is soft, has a low melting point and is an insulator of electricity. Because of these properties it has many applications that the other three materials do not.

The properties of materials are dependent on the bonding between their particles. Metals have metallic bonds; the free valence electrons allow conductivity and ductility of the solid. Covalent network substances contain strong covalent bonds between atoms making them hard, brittle and non-conducting solids.

Ionic substances consist of a lattice of positive and negative ions. The ionic bond is strong and the presence of ions means that when molten or dissolved in a solvent they are free to move and conduct an electric current.

Covalent molecular substances like wax consist of molecules weakly bonded to other molecules. The weak intermolecular forces mean that they are soft and easy to melt and the molecules have no charge so they cannot conduct in the solid or liquid state.

The purpose of this experiment is to use the relationships between the physical properties and the structure of ionic, metallic, covalent network and covalent molecular substances to classify samples of four types of solids.

Equipment

Samples of the four types of substances: metals, covalent network, ionic and covalent molecular substances. (Material samples that could be used include copper, lead, tin, sand, carbon block, ceramic tile, glass, sodium chloride, potassium nitrate, copper sulfate crystals, candle wax, paraffin wax, soap, naphthalene
power supply (0-12 V)
ammeter or globe in a socket (6 V, 0.5 A)
electrical leads (four, two with alligator clips)
4 × 100 mL beaker
water
spatula

The task

Develop a dichotomous key that can be used to classify:

- metallic substances
- covalent network substances
- soluble ionic substances
- covalent molecular substances.

Planning the investigation

1 Plan a series of observations and tests that can be used to distinguish the different types of material: metals, covalent network substances, soluble ionic substances and covalent molecular substances.

2 Briefly outline your plan. List the materials and equipment you need and identify safety requirements.

3 Check the proposed procedure with your teacher.

SAFETY NOTE:
- **DO NOT conduct conductivity tests of molten solids**
- **Check your plan with your teacher before you commence.**

Conducting the investigation

1 Conduct some preliminary trials. You should clarify that your tests work and do distinguish the substances. You should make sure that the tests are carried out in an appropriate order to minimise repetitions and to simplify classification.

2 Construct a dichotomous key from your refined trials.

3 Conduct a test of your dichotomous key. If time allows have someone else trial your classification key.

Processing the results

Discuss the success of your dichotomous key.

Evaluating the investigation

1 Evaluate the effectiveness of the procedures/tests used to distinguish materials in your key.

2 Describe how you would modify your classification key to include insoluble ionic substances. Include descriptions of any extra tests that would need to be conducted.

Notes

Experiment 13: Bonding and conductivity

Notes

Australia is the world's largest refiner of bauxite and the largest producer of gem and industrial diamonds. Petroleum products also provide significant royalty income to the WA Government.

Bauxite contains aluminium oxide $A\ell_2O_3$, the main source of aluminium ($A\ell$) metal. Aluminium oxide $A\ell_2O_3$, is an ionic compound and must undergo considerable chemical refining to eventually isolate the aluminium metal. On the other hand, diamond needs only to be separated from its source rock. Diamond is a three dimensional network of covalently bonded atoms of carbon. Petroleum exists as a covalent molecular substance. Atoms in petroleum molecules are held together by strong covalent bonds with weak forces between molecules.

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 (see Set 13). 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 available charge carriers 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.

Equipment

Part A

D.C. power supply (0-12 V)
Bunsen burner and tripod
pipe clay triangle
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)
filter paper
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 13.1.)
sodium hydroxide [NaOH] (5 g)
silver nitrate [$AgNO_3$] (5 g)
candle wax (5 g)
naphthalene [$C_{10}H_8$] (5 g)
sulfur [S_8] (5 g)

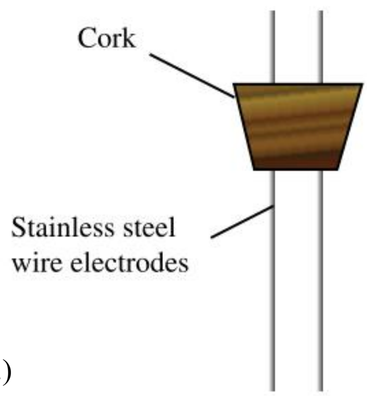


Figure 13.1

Equipment **Part B**

kerosene (50 mL)
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 (This can be made from two stainless steel plates 4 cm x 7 cm. These plates can be separated by about 0.3 cm, using two plastic spacers, and the system taped together so that it fits inside a 100 mL beaker as shown in figure 13.2)

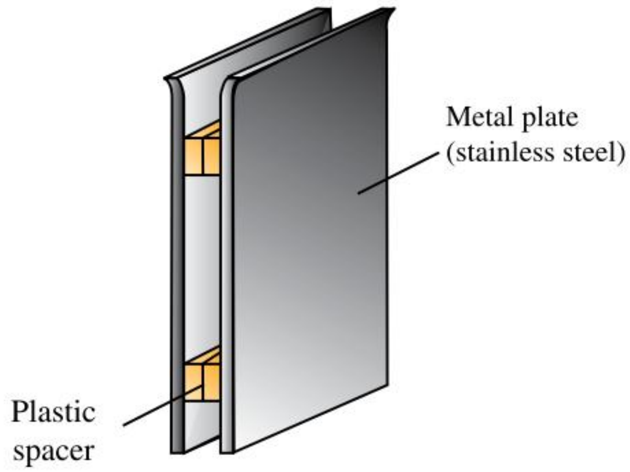


Figure 13.2

SAFETY NOTE:

- **Sodium hydroxide is very corrosive and must not be allowed to come in contact with your skin.**
- **Handle the NaOH pellets with a spatula or plastic spoon.**
- **Handle the $AgNO_3$ carefully as it produces bad stains.**
- **Be particularly careful handling the molten substances.**

Notes

Procedure **Part A: Electrical conductivity of ionic and covalent molecular substances**

(For safety reasons your teacher will conduct this part of the experiment as a demonstration.)

1. Connect a 6 V D.C. supply, ammeter and/or globe and wire electrode system in series as shown in figure 13.3.
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.

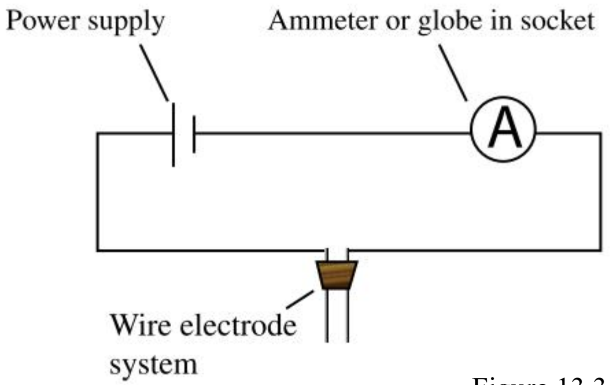


Figure 13.3

Experiment 13: Bonding and conductivity

Notes

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.
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. Repeat the procedure using silver nitrate instead of sodium hydroxide.
7. Repeat the procedure using, in turn, candle wax, sulfur and naphthalene. Record your results in a suitable table.

SAFETY NOTE:

- **Care must be taken when heating the candle wax, sulfur and naphthalene.**
- **If they catch fire, use tongs to place the lid on the crucible. This will extinguish the fire.**
- **After the crucibles have cooled wash out those containing the NaOH and $AgNO_3$ and scrape out the others.**
- **Heat sulfur in a fume cupboard, taking care not to inhale SO_2 fumes. SO_2 fumes can induce asthma symptoms if inhaled.**

Procedure

Part B: electrical conductivity of some liquids and aqueous solutions

1. Connect a 6 V D.C. supply, ammeter and/or globe, and plate electrode system in series (similar to the circuit in Figure 13.3).
2. Place 50 mL of kerosene 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 filter paper.
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.

Processing of results and questions

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, $NaC\ell$, NaOH and $HC\ell$) conducted an electric current? List the species present in each case and state why the solutions are conducting or not.
4. In the solid state, both $NaC\ell$ and NaOH are ionic lattices. What happens to these substances when they dissolve in water?
5. Pure $HC\ell$ is a covalent molecule. What happens to each $HC\ell$ molecule when it is dissolved in water?

Experiment 14: Bonding and solubility

Notes

Processing our resources often involves solvent chemistry. Bauxite for example is first mixed with sodium hydroxide. 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 kerosene. Conversely, petrol is insoluble in water but is quite soluble in kerosene.

In this experiment water and kerosene, two very different solvents, are used. The purpose of the experiment is to compare the effectiveness of water and kerosene as solvents for solutes with different types of bonding.

Equipment

test tube rack
test tubes (nine)
beaker (100 mL)
distilled water
kerosene (20 mL)
a few crystals of the ionic substances:
- sodium chloride [$NaC\ell$]
- copper(II) sulfate-5-water [$CuSO_4{\cdot}5H_2O$]
- 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 [$C_{10}H_8$]
- motor oil (~1 mL)

Procedure

1. Prepare a results table in which to record the solubility of the nine solutes listed above in water and kerosene. Record your results as insoluble (i), slightly soluble (ss) or soluble (s).
2. Place 1-2 mL of kerosene into each of nine test tubes.
3. Test the solubility of sodium chloride in kerosene by adding a few crystals to one of the test tubes. Shake the test tube gently and record the result in your table.
4. Repeat this procedure for each of the solutes in turn. Where the solute is a liquid add about 2-3 drops to the kerosene.
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.

Processing of results and questions

1. What generalisations can you make about the solubility in water and kerosene of the solutes tested?

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?

Experiment 15: The strength of intermolecular forces

***Intermolecular attraction*:**

- *dispersion forces,*
- *dipole-dipole forces and*
- *hydrogen bonding.*

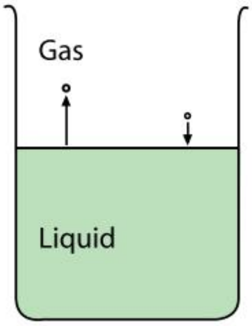


Figure 15.1: Evaporation occurs at the surface of a liquid

Notes

Fossil fuels were formed from the remains of dead plants. The fossilised remains were buried and, over hundreds of millions of years, exposed to extreme pressure and heat in the Earth's crust.

Fossil fuels range from volatile gases, like methane to liquids, like petroleum to solids like bituminous materials and almost pure carbon, like coal. At room temperature the variation of fossil fuels from solid to liquid to gas reflects the change in melting and boiling points of the hydrocarbons. The small molecules exist as gases the mid-sized molecules existing as liquids, while the large hydrocarbon molecules are solids at room temperature.

Differences in melting and boiling points of the various fossil fuel hydrocarbons are due to differences in the strength of the bonds between the molecules. The larger the molecule the stronger the intermolecular forces between molecules.

The strength of intermolecular forces also influences the rate of evaporation of liquids. Evaporation is the change of state from liquid to gas that takes place at any temperature below the boiling point (see Figure 15.1). When substances evaporate they absorb energy from the surroundings and produce a cooling effect: liquid + energy → gas

In this experiment you will compare the cooling effects of 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:
- methanol
- ethanol
- propan-1-ol
- butan-1-ol

filter paper squares (four, approximately 2 cm × 2 cm) Small elastic bands (one for each temperature sensor)
test tubes (four small)
test tube rack

Procedure

1. Construct a table similar to the one on the next page for the results and list the four liquids in the 'liquid' column in the 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.

Liquid	Initial temperature	Final temperature	Change in temperature

Notes

4. Identify which temperature sensor is responsible for each set of data produced on the screen. This can be done by warming one of the sensors with your hand and observing the response on the screen.
5. Wrap a small piece of square filter paper (2 cm × 2 cm) around each of the temperature sensors. The paper should be level with the end of the sensors and secured with a small rubber band that is as far away from the end of the sensor as possible.
6. Place methanol, ethanol, propan-1-ol and butan-1-ol in each of the four test tubes to a depth of approximately 3 cm.
7. Place one sensor in the methanol and the other in the ethanol for approximately 45 seconds to ensure that the filter paper is saturated. (The number of sensors used depends on the interface.) Then record the initial temperature of the solvents.
8. While continuing to collect data, remove the sensors from the test tubes and place them on the bench so that the filter paper end of the sensor projects over the edge of the bench. Continue to collect data and to graph the temperature until the temperature appears to reach a minimum.
9. Calculate the change in temperature that occurred when the methanol and ethanol evaporated. Enter your results in a table.
10. If possible, store this data so that data for the remaining liquids can be added to them.
11. Repeat steps 7-9 for the other liquids.
12. Obtain a printout of your results.

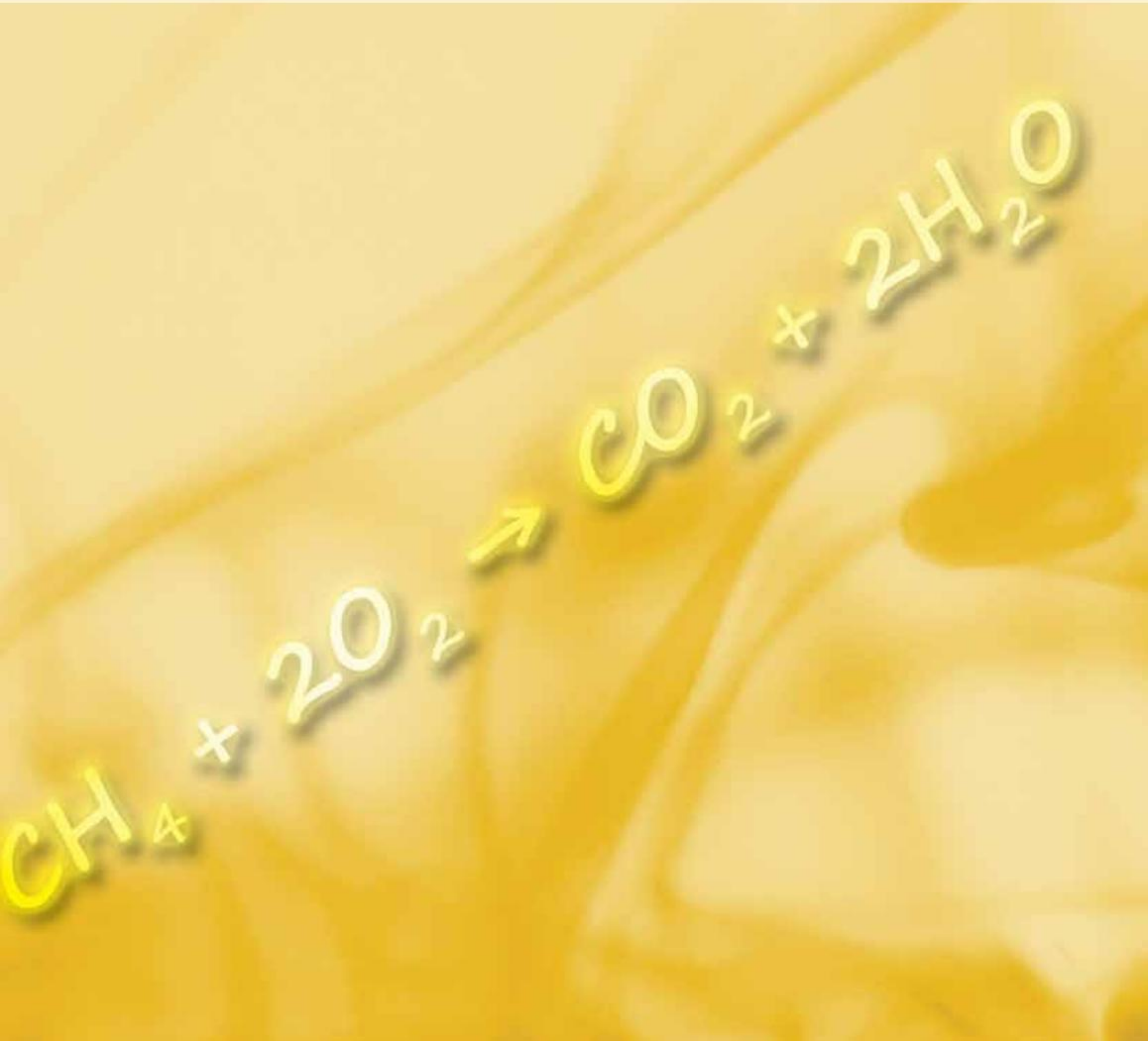
Processing of results and questions

1. Rank the alcohols, from greatest to least, in terms of their cooling effect.
2. Explain how the cooling effect relates to the rate of evaporation of the liquid.
3. Explain how the cooling effect and rate of evaporation relate to the strength of intermolecular forces in the liquids. Support your answer with examples from this experiment.

Chemical reactions and stoichiometry

The experiments and investigations in the chemical reactions and stoichiometry section of *Exploring Chemistry Year 11: Experiments, Investigations and Problems*, provides opportunities for students to explore:

- **different types of chemical reactions, including precipitation reactions and how they can be represented by chemical equations**
- **how the presence of specific ions in solutions can be identified by observing the colour of the solution**
- **chemical reactions quantitatively using the mole concept as it relates mass, moles and molar mass and, with the Law of Conservation of Mass**



Forensic science

Forensic science relies on laboratory testing to help solve crime. Many of the tests performed involve the identification of specific substances and provide evidence that links criminals to a crime. Laboratory research is pivotal to any investigation.

The results of forensic tests involving chemical reactions are often indicated by something obvious like a colour change. 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. Energy in various forms, like heat or light, may be released or absorbed.

In crime scene investigations, luminol is commonly used to detect the presence or absence of blood traces. This colour change is very striking, giving a unique blue glow effect, called luminescence. In order to give off the glow, the luminol is activated with an oxidant, usually acidified hydrogen peroxide solution. Even if blood were present, but has been cleaned or removed by the suspect, the colour change will still occur.

If a catalyst such as iron (Fe) is present, a series of reactions occur. Blood has iron in the haemoglobin molecule, so by testing evidence that is suspected to contain blood with luminol the bright blue glow should appear. Reactions during the luminol test produce excited electrons. As the electrons settle back down to the ground state, excess energy in the form of photons of visible strong blue light are liberated, but only for around 30 seconds. This is why the test is usually conducted in a dark room and a camera is at hand to photograph the image.

A Forensic Scientist from the Western Australian Police Service (WAPS) explains: “The luminol test for blood traces is common. It is conducted by police forensic investigators in the field and laboratory - routinely yet sparingly - and only when there is strong evidence to support possible foul play.”

The luminol chemiluminescence test is an incredibly simple and accurate indicator of the presence of blood but it can have a negative impact on the retrieval of DNA results. Our WAPS forensics expert suggests that “for the purposes of finding an offender, it is pointless to be able to identify blood, but not the person involved by DNA ID.” Consequently WA police and laboratory staff do only about six tests a year.

A sink used to clean up blood, following a gruesome WA homicide. The sink appeared clean to the naked eye. A luminol test reveals the presence of blood.

Photos courtesy of Western Australian Police Service

Experiment 16: Empirical formula of magnesium oxide

Notes

Many analytical techniques provide information about the relative amounts of elements present in a compound. Such techniques are most frequently used when new compounds are prepared and the formula of the compound is to be determined. From the data obtained in these types of experiments the empirical formula of compounds can be determined.

The empirical formula of a compound is the simplest whole number ratio of atoms of each element in the compound. In this experiment you will determine the empirical formula of magnesium oxide.

Equipment

balance such as an electronic top loading balance
crucible and lid
Bunsen burner
matches
magnesium ribbon [Mg] (0.2 g, about 20 cm)
steel wool
crucible tongs
pipe clay triangle
tripod

Procedure

1. Draw up a table for your results as shown below:

Materials	Mass +/- uncertainty (g)
crucible and lid	= g
crucible, lid & magnesium	= g
magnesium	= g
crucible, lid & magnesium oxide	= g
magnesium oxide	= g

2. Obtain a clean, dry crucible and lid and heat them for 5 minutes over a Bunsen burner flame. Allow them to cool.
3. Thoroughly clean the surfaces of a 20 cm strip of magnesium ribbon with steel wool.
4. Coil the magnesium ribbon around a pencil so that it will fit into the crucible.
5. Weigh the crucible and lid on a balance and record the mass in your results table.
6. Place the magnesium ribbon into the crucible, replace the lid and reweigh.

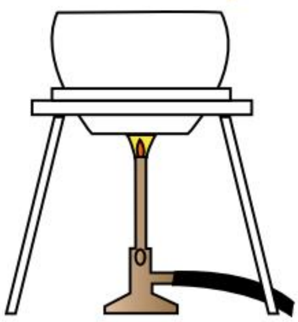


Figure 16.1

SAFETY NOTE:
- **Care must be taken when burning magnesium ribbon.**
- **Do not touch the hot crucible with your hands. Always use crucible tongs.**

Notes

7. Carefully heat the crucible and its contents with the lid off as shown in Figure 16.1 until the magnesium begins to glow. Immediately replace the lid and heat the crucible strongly.
8. Continue to heat the crucible for about 10 minutes, occasionally lifting the lid with tongs to provide oxygen for the reaction. Try to avoid the loss of any magnesium oxide smoke when the lid is lifted.
9. When all the magnesium has reacted, remove the lid and heat strongly for a further 5 minutes. Take note of your observations as the reaction proceeds.
10. Replace the crucible lid and allow the crucible and contents to cool.
11. Reweigh the crucible with its contents and lid.
12. The contents of the crucible can be safely disposed of in a waste bin.

Processing of results and questions

1. Record your qualitative observations of the magnesium before and after treatment with the steel wool. Also record your observations as the reaction occurs.
2. Record your quantitative primary data in your table. Indicate the uncertainty of your measured masses by considering the number of decimal places that can be read by the balance you used.
3. Process the results as follows. Calculate the mass of oxygen that has reacted with the magnesium and the number of moles of Mg and O and to 1 decimal place. Share these results with the class.
4. Calculate the ratio of n(Mg):n(O) and write the empirical formula of magnesium oxide.
5. Explain the need for heating the crucible and lid at the start of the experiment.
6. Explain why it was necessary to rub the magnesium with steel wool prior to weighing.
7. Write a balanced equation for the reaction between oxygen and magnesium.
8. Describe one source of random error. This is an error that may give a result above or below the actual value.

One possible source of systematic error is that if the sample is overheated it may form magnesium nitride.

9. Estimate the effect this error would have on the results you obtained for mass after heating and how it would affect the empirical formula.
10. Describe another source of systematic error.
11. Estimate how this systematic error would have affected the results you obtained for mass after heating and how it would affect the empirical formula.
12. Suggest a way that one of the systematic sources of error could be reduced.

Optional:
A spreadsheet could be set up to analyse and share class results.

Experiment 17: Water of crystallisation of barium chloride

Notes

Gravimetric techniques (analysis by weight) can be used to determine the water of crystallisation of hydrated compounds. Many ionic compounds that crystallise out of an aqueous solution have water molecules fixed in regular positions throughout the crystal lattice. These compounds are said to be hydrated. The number of water molecules per formula unit of the compound is shown as part of the formula. For example $CuSO_4{\cdot}5H_2O$, an important fungicide and $Na_2CO_3{\cdot}10H_2O$, washing soda, which is used as a water softening agent, have respectively five and ten water molecules of crystallisation per formula unit of compound. Heating the hydrated compound can drive off these water molecules of crystallisation. If the remaining compound has no water molecules of crystallisation; it is said to be anhydrous.

The purpose of this experiment is to determine the number of water molecules of crystallisation of hydrated barium chloride, $BaC\ell_2{\cdot}xH_2O$. To do this, a known mass of the hydrated compound will be heated so that the mass of water driven off can be determined. From this, the empirical formula of the hydrate will be calculated.

Equipment

balance
crucible and lid
Bunsen burner
hydrated barium chloride [$BaC\ell_2{\cdot}xH_2O$] (2 – 3 g)
pipe clay triangle
tripod
desiccator
crucible tongs

Procedure

SAFETY NOTE:
- **Do not touch the hot crucible with your hands. Always use crucible tongs.**
- **Barium chloride is poisonous.**
- **If barium chloride comes into contact with your skin wash with copious amounts of tap water for 20 minutes.**

1. Obtain a clean, dry crucible and lid and heat them strongly for 5 minutes over a Bunsen burner flame. Allow them to cool in a desiccator.
2. Weigh the crucible and lid on an accurate balance. Add 2-3 g barium chloride crystals and immediately reweigh.
3. Place the lid on the crucible, leaving a small gap, and heat for 5 minutes with a low flame. Then heat strongly for 20 minutes. Allow the crucible to cool in the desiccator and reweigh when cool.
4. Repeat the strong heating for a further 5 minutes, cool in the dessicator and again reweigh. Repeat until constant weight has been obtained, that is, to within 1 or 2 milligrams. This procedure is called "heating to constant mass".

Processing of results and questions

1. Determine the mass of anhydrous $BaC\ell_2$ and the mass of water driven off in the heating process.
2. Calculate the number of moles of anhydrous $BaC\ell_2$ and the number of moles of water lost during the heating process.
3. Calculate the ratio $n(BaC\ell_2) : n(H_2O)$. What is the empirical formula of the compound?
4. Write the formula of hydrated barium chloride in the usual way.

Investigation 18: Determination of the formula of Epsom salts

Epsom salts is the common name of hydrated magnesium sulfate. Epsom salts can be purchased at the supermarket and can be used as bath salts to soothe tired muscles and can help to heal skin problems. They can also be used to add magnesium and sulfur to soils deficient in these elements.

The formula can be written as: $MgSO_4 \cdot xH_2O$ where x is the number of water molecules per formula unit.

Notes

Planning the investigation

1. Plan an investigation to find the percentage of water present in a 2-4 g sample of Epsom salts and determine the value of x in the formula $MgSO_4 \cdot xH_2O$

2. Write out your proposed procedure including a list of the equipment you will need and identify any safety issues.

SAFETY NOTE:
- **Check your procedure with your teacher before commencing.**
- **The solid product of this reaction can be discarded in the regular rubbish bin after the experiment.**

Conducting the investigation

Conduct the investigation, making observations, and collecting and recording data in a table as you proceed.

Processing the data

1. Calculate the percentage of water present in Epsom salts.
2. Determine the value of x for $MgSO_4 \cdot xH_2O$ and write the formula of the hydrated salt.

Evaluating the investigation

1. Compare your results with others in your class. If class data is not available repeat the experiment and assess the similarity of the data from each trial.

2. The value of x obtained experimentally is rounded to a whole number. Explain why this is done.

3. Evaluate the procedure you used and suggest possible sources of error.

4. Suggest improvements to your method that could minimise the errors identified.

5. Research why it is called ‘Epsom Salts’.

Investigation 19: Conservation of mass

Notes

An understanding of the Law of Conservation of Mass is fundamental to chemistry. During a chemical reaction atoms are neither created nor destroyed, but are rearranged. As a result of this, the total mass of the reactants consumed will always equal the total mass of the products formed.

The task

Design an experiment that verifies the Law of Conservation of Mass using the reaction between sodium hydrogencarbonate and hydrochloric acid.

Equipment

sodium hydrogencarbonate powder
hydrochloric acid solution (2 mol L^{-1})
other equipment as specified by your investigation design

Planning the investigation

1. Plan an investigation to demonstrate the Law of Conservation of Mass.
2. Write out a proposed procedure for the investigation including any equipment and chemicals you require. Identify any safety requirements.
3. Check the proposed procedure with your teacher

Conducting the investigation

1. Conduct the investigation.
2. If time allows replicate the experiment.

Processing the data

Does your investigation confirm the Law of Conservation of Mass? Justify your answer by referring to the data you have collected.

Evaluating the investigation

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. You should consider the following:

What other reactions could be used to verify the Law of Conservation of Mass?
What problems, if any, were there in using a reaction that produces a gas?

Investigation 20: Determination of molar mass

Police have provided you and other laboratory teams with samples of a white solid found in the shoe tread of a number of crime suspects. They have also provided a container of sodium hydrogencarbonate found at the crime scene. It had been spilled on the floor and the police believe that it had been walked through by the perpetrator of the crime. Police are looking to match the white solid shoe sample to the sodium hydrogencarbonate found at the scene. Do you have a match?

The task

Examine the white solid shoe sample evidence and prepare a report for the police. Based on the evidence you examined conclude your report with a statement that either supports or refutes the police officers' suspicions about the suspect.

Equipment

evidence sample of unknown white solid (approximately 8 - 10 g)
sodium hydrogencarbonate (2 g)
hydrochloric acid [$HC\ell$] 3 mol L^{-1} (60 mL)
graduated cylinder (100 mL)
beaker (250 mL)
balance

Notes

Planning the investigation

1. In planning this investigation it is necessary to determine
   a) if our white solid is a hydrogencarbonate.
   b) the molar mass of the sample if it is a hydrogencarbonate.

The first step of the procedure would be to test a trace amount of the evidence sample to see if it is a hydrogencarbonate.

The second part in the procedure, if the hydrogencarbonate test proves positive, would be to apply an understanding of the conservation of mass (Investigation 19) and measure the mass of carbon dioxide produced when the hydrogencarbonate sample provided from the shoe is reacted with hydrochloric acid. Use this mass to determine the molar mass of the hydrogencarbonate in the reaction.

Finally compare the molar mass of the sample with that of sodium hydrogencarbonate.

2. Following is a procedure for conducting the investigation. The equipment and chemicals required is listed above. Identify any safety requirements.

Conducting the investigation

1. Test some of the sodium hydrogencarbonate with a few drops of hydrochloric acid. Test a trace amount of the evidence sample with hydrochloric acid. Tabulate your results and observations. Compare the two sets of observations to see if the evidence sample behaved the same way as the sodium hydrogencarbonate.

Investigation 20: Determination of molar mass

Notes

2. If your sample has a positive result in #1, accurately weigh a sample of the unknown hydrogencarbonate and record the mass in an appropriate table.
3. Pour approximately 60 mL of hydrochloric acid into a beaker and record the total mass of the beaker and acid.
4. Add the hydrogencarbonate and wait for it to react completely. You may need to swirl the beaker gently to make sure that all of the carbon dioxide has escaped from the solution.
5. Weigh the beaker and solution and record the mass.
6. If time allows replicate the experiment.

Processing of results and questions

1. Write a balanced equation for the reaction of sodium hydrogencarbonate with hydrochloric acid. Explain the observations in procedure #1 and provide an explanation as to why you decided that your evidence sample was a hydrogencarbonate or not.
2. Calculate the difference between the mass of the beaker, acid solution and hydrogencarbonate and the mass of the beaker and solution after the reaction. This reduction in mass is due to the mass of carbon dioxide lost.
3. From the mass of carbon dioxide produced, calculate the number of moles of carbon dioxide.
4. Assume the following equation is the reaction of the hydrogencarbonate evidence sample.
$$XHCO_3(s) + HC\ell(aq) \rightarrow XC\ell(aq) + H_2O(\ell) + CO_2(g)$$
Calculate the number of moles of $XHCO_3$ consumed in the reaction.
5. From the mass of $XHCO_3$ and the number of moles, calculate the molar mass of the hydrogencarbonate.
6. What is the most likely identity of your sample of hydrogencarbonate? (Compare the molar mass of your sample with that of sodium hydrogencarbonate).
7. What assumption has been made about the solubility of carbon dioxide in water? Find out the solubility of carbon dioxide at the temperature at which you conducted the experiment. Is the assumption a reasonable one? Justify your answer.
8. List possible sources of systematic and random errors in this experiment and state how you would go about minimising them.

Report for the police

Prepare a report for the police on the evidence provided. Conclude your report with a statement that either supports or refutes the police officers' suspicions.

Evaluating the investigation

1. 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.
2. Discuss whether the data are sufficient to support your conclusion.
   - How conclusive do you think that your results are?
   - What further testing needs to be done, if any, to confirm your conclusion?

Experiment 21: Types of chemical reactions

Notes

Many of the tests performed for crime scene investigations involve chemical reactions. The results help to provide answers to investigators questions and links between 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. The purpose of this experiment is to observe several important types of chemical reactions and to write equations for these reactions.

Equipment

test tubes (one large fitted with rubber stopper and plastic gas delivery tubing, two small)
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

5 mL solutions of:
- sodium bromide [NaBr] 0.1 mol L^{-1}
- sodium hydroxide [NaOH] 2 mol L^{-1}

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

30 mL, 0.1 mol L^{-1} solutions of:
- silver nitrate solution [$AgNO_3$]
- copper(II) sulfate solution ($CuSO_4$]
- sodium hydroxide solution [NaOH]
- hydrochloric acid [$HC\ell$]
- lead(II) nitrate solution [$Pb(NO_3)_2$]

Procedure

1. Prepare a table including space for reactants, observations and products for the following reactions.

Decomposition of a carbonate by heating

2. 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.**

Decomposition of a carbonate with an acid

3. 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 limewater solution.

Experiment 21: Types of chemical reactions

Notes

Oxidation of a metal

4. Place a protective mat on the laboratory bench. Cut a 3 cm strip of magnesium ribbon. Hold the ribbon with tongs and heat it in a Bunsen burner flame above the mat.

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

Reaction of a reactive metal with a dilute acid.

5. 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.

Precipitation reactions

6. Place 2-3 mL of 0.1 mol L^{-1} $AgNO_3$ solution into a test tube and add about the same volume of 0.1 mol L^{-1} NaBr solution.
7. Place 2-3 mL of 0.1 mol L^{-1} $CuSO_4$ solution into a test tube and add about the same volume of 2 mol L^{-1} NaOH solution.

Metal displacement reactions

8. Place about 25 mL of 0.1 mol L^{-1} $CuSO_4$ solution into a 100 mL beaker and place into the beaker a freshly cleaned zinc strip.
9. Place about 25 mL of 0.1 mol L^{-1} $Pb(NO_3)_2$ solution into a 100 mL beaker and place into the beaker a freshly cleaned zinc strip.
10. Place about 25 mL of 0.1 mol L^{-1} $AgNO_3$ solution into a 100 mL beaker and place into the beaker a freshly cleaned copper strip.

Neutralisation reaction (reaction of an acid and a base)

11. Place about 20 mL of 0.1 mol L^{-1} $HC\ell$ into a beaker and add 5 drops of Universal Indicator ™. Using a dropper, gradually add 0.1 mol L^{-1} NaOH solution until about 25 mL of the base has been added. Note any colour changes that take place.

Processing of results and questions

1. Write general word equations and balanced chemical equations for each of the reactions observed.
2. Write word equations and balanced chemical equations for the following reactions:
   (a) heating magnesium carbonate;
   (b) adding dilute sulfuric acid to copper(II) carbonate;
   (c) burning of iron in pure oxygen;
   (d) adding dilute sulfuric acid to calcium;
   (e) mixing solutions of barium nitrate and sodium sulfate to precipitate barium sulfate;
   (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 and potassium hydroxide solution to form potassium sulfate and water.

Solutions and acidity

The experiments and investigations in the solutions and acidity section of *Exploring Chemistry Year 11: Experiments, Investigations and Problems*, provides opportunities for students to explore:

- **the presence of specific ions in solutions by observing the colour of the solution and observing precipitation and acid-base reactions**
- **the solubility of substances in water, including ionic and polar and non-polar molecular substances**
- **the Arrhenius model to explain the behaviour of strong and weak acids and bases in aqueous solutions**
- **indicator colour and the pH scale to classify aqueous solutions as acidic, basic or neutral**
- **patterns of the reactions of acids and bases, including reactions of acids with bases, metals and carbonates**

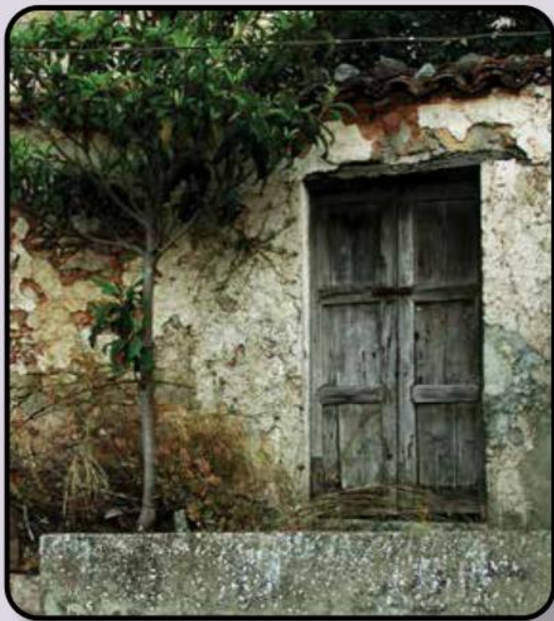
Photo supplied courtesy of Pfizer (Perth) Pty Limited

Chemistry - our body and buildings

Vitamin E – The most abundant fat-soluble antioxidant in the body is found in almonds, in many oils, wheat germ, safflower, corn and soybean, and also in mangoes, nuts, broccoli and other foods.

Vitamin C – The most abundant water-soluble antioxidant in the body can be found in large amounts in many fruits and vegetables but is also found in cereals, fish, beef, and poultry.

Acid rain
The effect of acids and bases on the natural environment and on the many environments created as a result of human activity is enormous. Some effects are rapid and dramatic and some are slower and long term. Acid rain has been shown to have adverse impacts on forests, freshwaters and soils, killing insect and aquatic lifeforms as well as causing damage to buildings and having impacts on human health.



What has taste got to do with it?
Have you ever wondered why lemons and most unripe fruit taste sour, why many antacid powders and other medicines taste bitter? Citrus fruits like lemon contain citric acid and ascorbic acid, vitamin C (which is in many fruits). Acids taste sour. The bitter taste of many materials is due to the base or bases that they are made of. Bases are bitter.

Christmas Island and Nauru Rocks!
Phosphoric Acid (H_3PO_4) can be made by treating rock containing phosphorus (called rock phosphate) with sulfuric acid. Rock phosphate has been obtained from Christmas Island and Nauru. The main use of phosphoric acid has been in the production of superphosphate fertiliser.

What has pH to do with our hair and skin?
Chemical technology is responsible for a huge range of personal hygiene and beauty products: soaps, deodorants, shampoos, conditioners, perfume, toothpaste, moisturisers and make-up. Many of these products claim to improve the appearance of our hair, skin, and teeth or to treat a medical condition.

Most modern shampoos, conditioners and shower gels are formulated to be slightly acidic, having a pH around 4.0-5.0. Why?

The hair and skin are both covered by a very thin fluid layer comprised of oil, salt and water. This layer, called the mantle, is acidic with a pH around 4.5 to 5.0. This acid mantle helps to maintain the moisture balance in our hair and skin. Exposure to the environment and washing disturbs the pH balance of the mantle. Using pH-balanced products can restore the pH of the hair and skin.

Experiment 22: Solubility rules

Plants and animals access to nutrients is dependent, in many instances, on the nutrient's solubility. Most of the food that we consume provides us with water and energy-yielding nutrients such as carbohydrates (sugars and starches), lipids (fats), and proteins. Besides these major nutrients, our bodies require smaller amounts of a variety of other molecules and ions, collectively known as vitamins and minerals.

Vitamins are organic molecules containing the elements C, H, N and O. They are needed in trace amounts to help catalyse many of the biochemical reactions in the body. Minerals, on the other hand, are typically inorganic elements consumed in the form of a salt containing the mineral element and other ions. Minerals, like vitamins, perform a wide variety of functions in the body. Magnesium ions Mg^{2+} and zinc ions Zn^{2+} for example help the workings of enzymes; sodium ions Na^+ and potassium ions K^+, calcium ions Ca^{2+}, and chloride ions $C\ell^-$, help to maintain electrical and water balance in the body, transmit nerve impulses, and stimulate muscle contraction. Others such as phosphate ions, form compounds responsible for bone growth and structure.

Notes

In order to use the nutrients that we get from our food, they must be soluble. Digestion processes enable the nutrient minerals to be eventually dissolved into the blood. The blood carries these nutrients to where they can be reassembled and used by the body.

Fertilisers provide nutrients for plants. Ionic substances dissolve in water to varying extents. The different solubilities of these substances have several applications. For example, chemists use knowledge of solubilities to identify the presence of certain ions, both cations (+ ions) and anions (- ions), and also to prepare other compounds.

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

test tubes (several)
test tube rack
Solutions of the following (all 0.1 mol L^{-1}, about 20 mL):
- sodium nitrate [$NaNO_3$]
- sodium chloride [$NaC\ell$]
- sodium sulfate [Na_2SO_4]
- sodium carbonate [Na_2CO_3]

Solutions of the following in dropper bottles (all 0.1 mol L^{-1}):
- 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 [$CuSO_4$]
- 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 [$FeC\ell_3$]

beakers (two 100 mL)
droppers (two)

Experiment 22: Solubility rules

Notes

n.v.r = no visible reaction

Procedure

1. Draw up a results table as shown below.

Solution to be tested	Cation present	$NaNO_3$ anion added NO_3^-	$NaC\ell$ anion added $C\ell^-$	Na_2SO_4 anion added SO_4^{2-}	Na_2CO_3 anion added CO_3^{2-}
KNO_3	K^+	n.v.r.	n.v.r.		
$AgNO_3$	Ag^+	n.v.r.	white ppt.		
$Pb(NO_3)_2$	Pb^{2+}				
$MgSO_4$	Mg^{2+}				
$FeSO_4$	Fe^{2+}				

2. Place about 1 mL of each of the solutions to be tested (KNO_3, $AgNO_3$, $Pb(NO_3)_2$ etc) into separate clean test tubes.
3. Place about 20 mL of 0.1 mol L^{-1} $NaNO_3$ into a clean 100 mL beaker. Use a dropper to add about 1 mL of the $NaNO_3$ solution to each of the solutions in the test tubes. Where a precipitate occurs indicate its colour in the table. If no reaction is observed indicate with the letters 'n.v.r.' (no visible reaction).
4. When all of the solutions have been tested with the $NaNO_3$ solution carefully clean the test tubes, and repeat the procedure with the 0.1 mol L^{-1} $NaC\ell$ solution (column 4), 0.1 mol L^{-1} Na_2SO_4 solution (column 5) and 0.1 mol L^{-1} Na_2CO_3 solution (column 6).

Processing of results and questions

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

2. Examine columns 4, 5 and 6 of the table in the same way as you did for column 3 in question 1.

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

4. Research and classify the cations tested either as an essential body nutrient or poison.

5. What ions are provided in a common lawn fertiliser? List the salts present on the label of a common lawn fertiliser.

6. How could you remove lead (II) ions from wastewater before it is discharged?

Investigation 23: Identify the unknowns I

The situation

A freak storm has forced rainwater to flood the chemical storage room. Labels have smudged and some have completely fallen off a number of bottles.

The task

Design and carry out an investigation that will identify the set of unlabelled solutions of different cations with the same anion (nitrate ion) and a second set of solutions of different anions with the same cation (sodium ion).

Identification of cations

Set 1 (unknown cations):	0.1 mol L^{-1} nitrate solutions of iron(II), lead, silver, copper(II) and magnesium, randomly labelled A, B, C, D and E.
Set A (known test solutions):	0.1 mol L^{-1} solutions of sodium hydroxide, ammonia, sodium chloride and potassium iodide.

Identification of anions

Set 2 (unknown anions):	0.1 mol L^{-1} solutions of sodium carbonate, sodium hydroxide, sodium chloride, sodium sulfate or sodium nitrate randomly labelled as unknown solution F, G, H, I and J.
Set B (known test solutions):	0.1 mol L^{-1} solutions of lead nitrate, silver nitrate, iron(II) sulfate (fresh), copper(II) sulfate and sulfuric acid.

Planning the investigation

1. Plan an investigation to correctly identify the unknown cations and anions.
2. Write out your proposed plan, list the equipment required and identify the safety requirements.

SAFETY NOTE:
- **Check your plan with your teacher before you start.**
- **Sulfuric acid is corrosive! Handle with care.**

Conducting the investigation

1. Conduct the investigation, collecting and recording the data as you proceed.

Processing the data

1. From your observations identify each of the unknowns in the set. Include the reasons for your identifications and write chemical equations for all reactions that were observed.
2. Write an ionic equation showing how each precipitate forms.

Evaluating the investigation

1. Evaluate the effectiveness of your procedure and describe any modifications you would make to improve it.
2. Discuss your confidence in the findings of the investigation.

Notes

Use precipitation reactions to determine which cation you have in each bottle. You should make a record of each experimental test you do and explain how/why your testing determines the identity of the cation.

Investigation 24: Identify the unknowns II

Notes

Some substances are more soluble in water than others. The different solubilities of these substances have several applications. For example, chemists use knowledge of solubilities to identify the presence of certain ions, to prepare compounds in horticulture and in re-hydration. It also helps in various forensic tests and separation methods.

The task

Apply the solubility rules to identify a set of unknown solids or solutions.

Your teacher will allocate you one or more of the following sets of solids or solutions. Each member of the set provided will be labeled only as A, B, C ... etc.

Set 1: 0.1 mol L^{-1} solutions of $BaC\ell_2$, $CuSO_4$, H_2SO_4, $NaC\ell$ randomly labelled as A, B, C, D.

Set 2: 0.1 mol L^{-1} solutions of $AgNO_3$, $ZnSO_4$, $BaC\ell_2$, $NaC\ell$ randomly labelled as G, H, I, J.

Set 3: 0.1 mol L^{-1} solutions of H_2SO_4, $BaC\ell_2$, Na_2CO_3, $Pb(NO_3)_2$, KOH and 0.2 mol L^{-1} $HC\ell$ randomly labeled as M, N, O, P, Q, R.

Set 4: Solid samples of $BaC\ell_2$, $Zn(NO_3)_2$, Na_2CO_3, $CaCO_3$ randomly labelled as W, X, Y, Z. (Distilled water, dilute hydrochloric acid and limewater solution are also provided to test the unknown solid samples.)

It will be necessary for you to plan a procedure to identify the unknowns provided, carry out the procedure and evaluate your results.

Equipment

spatula
test tube racks
test tubes (several)
distilled water
beakers (two 100 mL)

Unknown samples in Set 1, 2, 3 or 4 as listed above with each chemical labelled only as A, B, C etc.
set 1-3 solutions in dropper bottles
set 4 solid samples in small sample bottles; if Set 4 unknowns are used, the following solutions in labelled dropper bottles as well as delivery tubes are also provided:
hydrochloric acid [$HC\ell$] 1 mol L^{-1}
saturated limewater solution [$Ca(OH)_2$]

Planning the investigation

1. Plan an investigation to identify the unknown solids or solutions provided. For Sets 1-3 you may only use the solutions provided. For Set 4 the solid samples can be tested with the hydrochloric acid, limewater and distilled water provided. For each test, use a minimum amount of sample - for solids about a pea-sized sample; for solutions about 2-3 mL.

2. Write out your proposed plan, list the equipment required and identify the safety requirements.

3. Check the proposed plan with your teacher.

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

Conducting the investigation

Conduct the investigation, collecting and recording the data in a table as you proceed.

Processing the data

From your observations identify each of the unknowns in the set. Include the reasons for your identifications and write chemical equations for all reactions that were observed.

Evaluating the investigation

1. Evaluate the effectiveness of your procedure and describe any modifications you would make to improve it.

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

Notes

Experiment 25: Electrolytes and conductivity

Notes

Electrolyte materials can be found in all sorts of places in and around the home. Your medicine cabinet for example, may contain vitamins and minerals. Have you any sports drinks in the refrigerator? Do you have a swimming pool? Do you have plant fertilisers in the garden shed? All of these materials contain electrolytes.

Electrolytes are chemical compounds like salts that, when they dissolve, the resulting solution contains free ions that are capable of conducting electricity. Blood plasma contains important dissolved electrolytes. The dissolved electrolytes in the blood contribute to the general functioning of organs such as the kidney and heart, regulating the body's fluid and pH balance. Electrolytes also carry electrical currents that help nerve and muscle function. Important electrolytes include ions of magnesium, calcium, potassium, sodium, chloride and phosphate.

In this experiment you will investigate a number of different solutions that could be found around the home.

Equipment

power supply (0-12 V)
plate electrode system as shown in Figure 25.1
switch
ammeter (or 2.5 V globe)
electrical leads (four)
beakers (two 100 mL)
distilled water (50 mL)
50 mL of each of the following:
tap water
sugar solution
1 mol L^{-1} acetic acid [CH_3COOH] in vinegar, found in the kitchen
1 mol L^{-1} ammonia solution found in the laundry
1 mol L^{-1} hydrochloric acid [$HC\ell$] found in the pool shed
1 mol L^{-1} potassium nitrate [KNO_3] in some fertilisers found in the shed
swimming pool water if you have some available

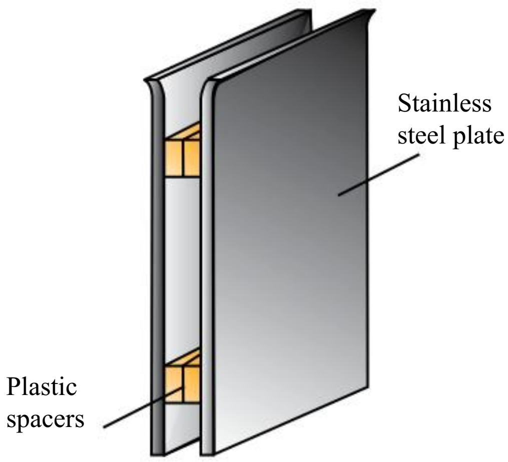


Figure 25.1: Plate electrode

Figure 25.2: Conductivity test set-up for solutions

Procedure

1. Set up the circuit and the electrodes as shown in Figure 25.2. Set the power supply to 6V.

2. Place 20 mL of 1 mol L^{-1} $HC\ell$ in a 100 mL beaker, put the electrodes into the solution and briefly close the switch. If a glow is not detected or the ammeter reading is very low on the 6 V power setting increase the voltage to 8 V (up to a maximum of 12 V). Tabulate your results.

3. Record the reading on the ammeter or record the brightness of the globe.

4. Wash the beaker and electrodes with distilled water and then, in turn, test the conductivity of 20 mL samples of: distilled water, tap water, sugar solution, acetic acid 1 mol L^{-1} , ammonia solution 1 mol L^{-1} , potassium nitrate 1 mol L^{-1} and swimming pool water.

Processing of results and questions

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

2. What general rule can you find from the class results?

3. Explain your results using diagrams where appropriate and by describing the type of species present in each of the different solutions.

4. Read the label of a garden or lawn fertiliser. Draw up a table of the active ingredients contained in the fertiliser, giving the name and formula and have a column that classifies the component as a non-electrolyte, weak electrolyte or strong electrolyte.

Notes

Experiment 26: pH of materials around the home

Notes

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.

In this experiment you will 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
samples of:

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

Procedure

Notes

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.**

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 your results in a suitable table.

Processing of results and questions

1. List all those solutions that were neutral.

2. (a) List all those solutions that were acidic.
   (b) What is the name of the component in each of these that is responsible for the solutions being acidic? You may have to research some of these.
   (c) Choose one of these substances and write a balanced ionic equation to show why the solution is acidic.

3. (a) List all those solutions that were basic.
   (b) What is the name of the component in each of these that is responsible for the solutions being basic? You may have to research some of these.
   (c) Choose one of these substances and write a balanced ionic equation to show why the solution is basic.

4. Investigate the pH of rainwater and groundwater in your area.
   - Predict the pH you expect the water to have.
   - Explain your predictions.
   - Measure and record the pH readings of the rainwater and groundwater.
   - Explain any differences between predicted pH and measured pH values.

Experiment 27: The right pH shampoo for your hair

An understanding of the chemistry of hair care, including the role of pH, helps in the development of better hair care products. pH is important when it comes to shampoo. When deciding on which product to buy, would you choose a low or high pH brand? Do such shampoos actually strengthen hair?

A solution of pH 7 is said to be neutral. The hydrogen ion, H^+, and hydroxide ion, OH^- , concentrations are equal. An acidic solution has more H^+ than OH^- giving a pH less than 7, while a basic solution has more OH^- than H^+ and a pH greater than 7.

Notes

Equipment

test tubes (several)
Universal Indicator ™ solution (5 mL) or you can use a pH probe
Universal Indicator ™ colour chart
samples of different hair shampoos
eyedroppers or Pasteur pipettes (enough for one per shampoo sample)
dimple tray or 5 test tubes
solutions of hydrochloric acid [$HC\ell$] 0.01 mol^{-1} (pH = 2) and 0.0001 mol^{-1} (pH = 4)
tap water (pH = 7)
solutions of sodium hydroxide [NaOH] 0.0001 mol^{-1} (pH = 10.0) and 0.01 mol^{-1} (pH =12.0)
5 wooden splints (pop sticks)
5 Petri dishes
25 strands of the same type of hair; you could use your own, from your hair brush or through a local hairdresser
sticky tape.
wash bottle of distilled water
microscope
hair drier (optional)

Procedure

Part A: measuring the pH of different shampoos

1. Using either a dimple tray or test tubes collect a few drops of each of the different shampoos provided using an eyedropper or pipette. Make sure you have labelled your samples so you don't mix up your results. Record any label information about the pH of the shampoo.
2. Add a drop or two of Universal Indicator ™ to each of your shampoo samples. Record the colour and pH of each of the samples in a table similar to that below.

Shampoo name	Label information	Universal Indicator ™ colour	pH

Procedure

Part B: testing the effect of pH on hair

1. Collect and clean 5 Petri dishes and rinse with distilled water.

2. Label the Petri dishes and the wooden splints pH 2.0, pH 4.0, pH 7.0, pH 10.0, and pH 12.0. Add 10 mL of the appropriate pH solution to each of the Petri dishes.The setup is shown in Figure 27.1.

Figure 27.1: Equipment for part B

3. Bundle the hair into 5 sets of 5 strands and tape one end of each set to separate wooden splints.

4. Lay each wooden splint across the top of the Petri dish with matching label and allow the free ends of the hair to suspend and soak in the test solution for 10 minutes.

5. Remove the hair samples, rinse them with distilled water and suspend so the hair dries in the air as shown in Figure 27.2. A hair drier could be used but is not essential.

Figure 27.2: Drying

Notes

Experiment 27: The right pH shampoo for your hair

Notes

6. Carry out the following observations and tests of the properties of each of the hair bundles. Record your observations in a table similar to that shown. Rank the texture from 1 (roughest) to 5 (smoothest) and resilience (strength) 1 (easiest to break) to 5 (hardest to break).
   (a) Observe the texture - roughness, smoothness.
   (b) Hold the ends of the hair bundle and pull gently to observe the hair's resilience - its ability to stretch and contract without breaking.
   (c) Using the microscope examine the hair, sketch what you see and record any additional observations about its texture etc.

pH Treatments	Texture	Resilience	Sketch of Hair
pH 2.0			
pH 4.0			
pH 7.0			
pH 10.0			
pH 12.0			

Processing of results and questions

1. Did the shampoos differ greatly in pH? Record the range.
2. Compare label information on the shampoo bottles with your observations of measured pH values.
3. Is your preferred hair shampoo low or high pH?
4. What pH treatment provided the hair sample with the best resilient properties?
5. Which pH treatment gave the hair sample the least resilient characteristics?
6. Using your results identify the best shampoo pH range for hair.
7. Will the results from this experiment influence your future choice of hair shampoo? Explain.
8. What is the purpose of using conditioners after washing your hair?

Investigation 28: Determining soil pH using Universal Indicator ™

Some plants are very sensitive to soil pH. Potatoes and strawberries, for example, grow best in more acidic soils while cabbage and cauliflower prefer alkaline conditions. The majority of food crops tend to grow best in a neutral or slightly acidic soil. Generally, soils in dry climates tend to be alkaline and those in moist climates tend to be acidic.

The acidity or alkalinity of the soil is measured by pH. You can buy pH test kits at most nurseries, or hardware stores. The better kits also have booklets describing how to interpret your result.

Sample 1

Sample 2

Sample 3

The task

You are required to find the pH of three soil samples.

Sample 1 obtained from a newly developed beach side suburb near extensive limestone outcrops.

Sample 2 obtained from the local nursery. It is the soil used by the nursery to plant seedlings into pots.

Sample 3 (optional) obtained from a paddock on a farm that was used to graze sheep. To achieve the best pasture the farmer used large amounts of ammonium nitrate as a fertiliser on the paddock every year.

Equipment

Write a list of equipment you need to complete this experiment.

Planning the investigation

1. Write the steps you will use to conduct this experiment. You should include a description of the method you will use to prepare the sample and the precautions you will take to ensure that the results are accurate.

2. Draw up a suitable table to record your results.

Conducting the investigation

Conduct the investigation collecting and recording the observations and results in a table.

Processing the data

Write a report summarising your results.
For each of the soils write a possible explanation for the observed pH.

Evaluating the investigation

Describe any improvements you could make to your procedure that could give you a more accurate result.

Notes

Experiment 29: Electrical conductivity of acids and bases

Notes

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.

Equipment

power supply (0 to 12 V)
plate electrode system as shown in Figure 29.1
switch
ammeter
electrical leads (four)
beakers (two 100 mL)
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}

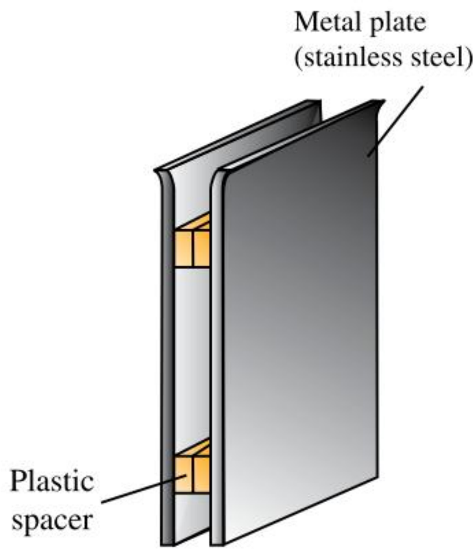


Figure 29.1. Plate Electrode System

Procedure

Set up the circuit and the electrodes as shown in Figures 29.2 and 29.3. Set the power supply to 6 V DC.

Figure 29.2

CAUTION!
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.

2. Place 50 mL of 1 mol L^{-1} $HC\ell$ in a 100 mL beaker, put the electrodes into the solution and briefly close the switch.

3. Record the reading on the ammeter.

4. Wash out the beaker and electrodes with distilled water and then, in turn, test the conductivity of :
   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)). Record your results in a suitable table.

Notes

Processing of results and questions

1. Explain why these solutions conduct an electric current. Use an equation in each of your answers.

2. Look at your results for the $HC\ell$ 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} $HC\ell$ 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? If so, 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.

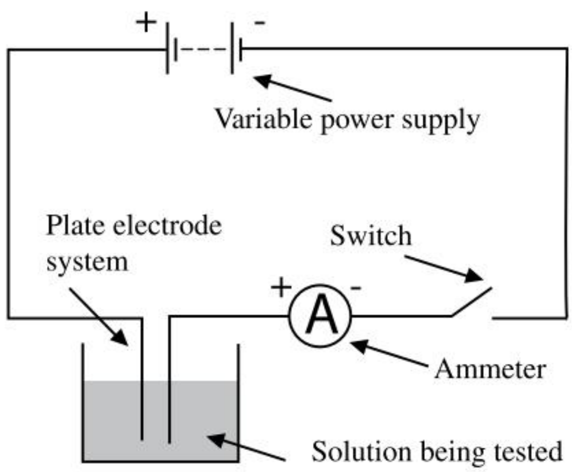


Figure 29.3: Circuit diagram

Investigation 30: Acid or base? Strong or weak?

Notes

IMPORTANT

Before you attempt this investigation you should complete Experiment 29 Electrical conductivity of acids and bases

The chemistry of acids and bases can be used to explain why:

- good shampoo has a low pH;
- lemons and most unripe fruit taste sour;
- antacid powders work;
- some medicines taste bitter;
- acid rain is so damaging.

You can identify an acid from a base using an indicator. The electrical conductivities of aqueous solutions give an indication of the concentration of ions in the solutions. It is now time to apply your understanding of acids and bases to carry out the following investigation.

The task

You will be given 0.100 mol L^{-1} solutions of:

hydrochloric acid [$HC\ell$]
sodium hydroxide solution [NaOH]

Samples of the following household items:

vinegar	washing powder	body wash
shampoo	soda water	soft drink
fruit juice	antacid powder	sodium chloride
milk		

Using knowledge and skills gained by conducting experiments involving electrical conductivity and the use of indicators, you will be required to:

(i) identify the pH of each solution and
(ii) determine its electrolyte (acid/base) strength.

Research each household item to determine the primary source of its pH and electrolyte strength.

Planning the investigation

1. Plan a laboratory investigation to identify the solution pH and to classify it as acidic, basic or neutral.
2. Plan a second investigation to determine the strength and electrolyte classification of each solution.
3. Plan your alternative methods or qualitative test/s that you will use to confirm your pH and strength classifications of the solutions provided.
4. Write out each procedure, and list chemicals, equipment and safety procedures needed to complete the tasks.
5. Have your teacher check your plans before proceeding.

Conducting the investigation

Conduct each part of the investigation collecting and recording the observations and results in a table as you proceed.

Processing the data

Write a report summarising your results. Your report must include your conclusion, and an explanation of the reasons you reached your conclusions as a result of your observations.

Evaluating the investigation

Write an evaluation of your experimental designs, focusing on any improvements that could be made so that the accuracy of the results could be improved.

Notes

Experiment 31: Acid reactions with some metal compounds

Concrete

Notes

Understanding the properties of acids and bases is vital to the understanding of the effect of water that has a low pH on metal compounds such as metal oxides, hydroxides and carbonates found in rocks and soils as well as in many modern structures made of concrete and mortar.

Commonly available acids include spirits of salts (hydrochloric acid), battery acid (sulfuric acid) and vinegar (acetic acid), and commonly available bases include soda ash (sodium carbonate), baking soda (sodium hydrogencarbonate), caustic soda (sodium hydroxide) and ammonia.

In this experiment you will investigate the effect of acids (using hydrochloric acid and vinegar) on some oxides, hydroxides, carbonates and hydrogen carbonates commonly found in rocks, soils and building and household materials.

Equipment

Bunsen burner
dropper
test tubes (7)
matches
hydrochloric acid, $HC\ell$, 2 mol L^{-1} (40 mL)
vinegar (10 mL)
saturated limewater solution, $Ca(OH)_2$ (40 mL)
about 1 g each of the metal oxides and hydroxides:
- copper(II) oxide [CuO]
- iron(III) oxide [Fe_2O_3]
- lead(II) oxide [PbO]
- calcium hydroxide [$Ca(OH)_2$]

about 1 g each of the metal carbonates and hydrogencarbonates:
- calcium carbonate [$CaCO_3$]
- sodium carbonate [Na_2CO_3]
- sodium hydrogencarbonate [$NaHCO_3$]

about 10 mL of sodium hydroxide solution (caustic soda) [$NaOH$] 2 mol L^{-1}
blue litmus paper
red litmus paper
litmus solution

Procedure

Notes

1. Draw up a table that can be used to summarise the observations for the experiments you will perform. Your table should include the following headings:
   (a) the reactants involved in the reaction you are performing;
   (b) your observations of the reaction: these observations should be thorough and clearly written;
   (c) products identified from your observations and an explanation of how your observations allowed you to identify those products.

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

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

3. Solutions of hydrochloric acid and sodium hydroxide are colourless and, like all solutions, they are clear. Make sure that containers used to store these solutions are 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} $HC\ell$ 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 observations.

B Reaction of acids with metal oxides and hydroxides

1. Place a small quantity (about the size of two match heads) of CuO in a test tube. Add 4 to 5 mL of 2 mol L^{-1} $HC\ell$, gently heat the mixture but do not boil it, then allow it to stand.
2. Record your observations.
3. Repeat step 1 and 2 with all the other oxides and hydroxides.

Experiment 31: Acid reactions with some metal compounds

Notes

C Reaction of acids with metal carbonates and hydrogencarbonates

1. Place a small quantity (about the size of two match heads) of $CaCO_3$ in a test tube. Add 2 to 3 mL of 2 mol L^{-1} $HC\ell$. Set up the test tube as shown in figure 31.1.

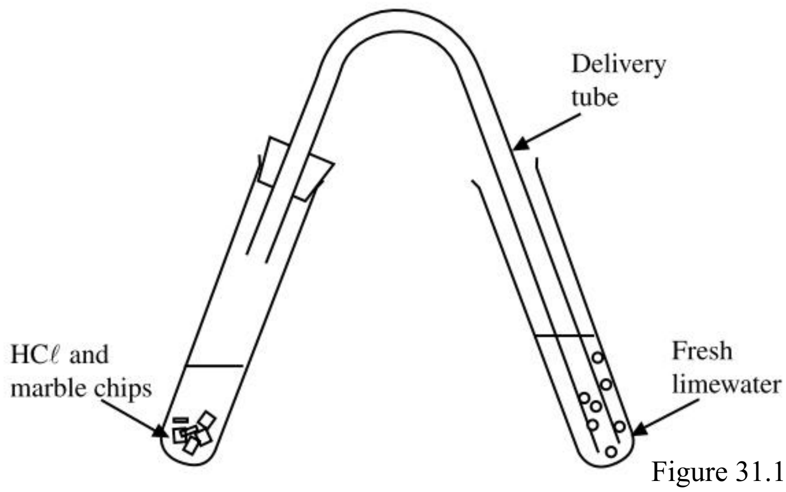


Figure 31.1

2. Record your observations.
3. Repeat step 1 and 2 with Na_2CO_3 and $NaHCO_3$.
4. Repeat steps 1 and 2 with $CaCO_3$ using vinegar in place of the $HC\ell$ solution.

Processing the results and questions

1. Write an ionic equation for each of the reactions you observed.

2. Compare and contrast the reactions of sodium hydroxide and ammonia with hydrochloric acid.

3. Why do you think it was necessary to warm some of the solutions?

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

5. Acid rain is a problem in industrial areas that burn sulfur-containing coal, diesel and fuel oil. Research the pH of acid rain and the effect of acid rain on limestone, marble and sandstone buildings and on aluminium oxide in soil. Use one or more of the reactions observed in this experiment to either support or refute your research findings regarding the impact of acid rain on buildings.

Experiment 32: Reactions of dilute acids with metals

To investigate the effect of acids on metals we will be using hydrochloric acid. Other commonly available acids such as battery acid (sulfuric acid) or ethanoic acid (vinegar) could also be used. This experiment will help to develop your understanding of acids and why metal structures are susceptible to acid rain.

Notes

Equipment

Bunsen burner
dropper
test tubes (six)
taper and matches
30 mL 2 mol L^{-1} hydrochloric acid, $HC\ell$
two small pieces each of aluminium, $A\ell$; copper, Cu; iron, Fe; magnesium, Mg and zinc, Zn
steel wool or emery paper

Procedure

1. Draw up a table that can be used to summarise the observations for the experiments you will perform. Your table should include the following headings:
   (a) The reactants involved in the reaction you are performing.
   (b) Your observations of the reaction. Observations should be thorough and clearly written.
   (c) Products identified from your observations and an explanation of how your observations allowed you to identify those products.

2. 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.

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

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 water for 20 minutes. You must wear safety glasses at all times during this experiment.

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

5. Place the small pieces of Mg, Zn, $A\ell$, Fe and Cu into separate test tubes. Add 2-3 mL of 2 mol L^{-1} $HC\ell$ to each. If there is no sign of reaction warm the mixture carefully but do not boil it. (Note: not all the metals will react - do not confuse air bubbles that appear in heated water with gas evolved by a reaction)

Experiment 32: Reactions of dilute acids with metals

Notes

6. 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 32.1.

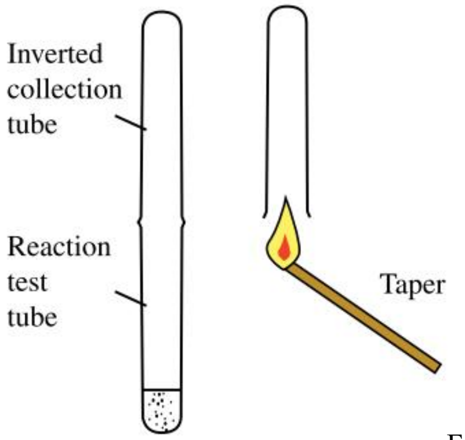


Figure 32.1

Processing of results and questions

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

2. (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.

3. Why do you think it was advisable to warm some of the solutions?

4. Research an example of where acids have a detrimental effect on metal structures or containers and suggest how this effect could be eliminated or reduced.

Experiment 33: Metal and non-metal oxides

Atmospheric pollution can have natural sources, like volcanic eruptions, but the term is usually used to refer to the gases produced from human activities, such as burning fossil fuels (coal, natural gas, petrol, diesel and oil), deforestation and agriculture. Humankind over the last 200 years, in particular, has significantly altered the composition of the atmosphere through pollution. Although air is still made up mostly of oxygen (21%) and nitrogen (78%), the levels of many of the trace gases (carbon dioxide, methane and nitrous oxides for example) have increased through pollution. Human processes have also released some completely new gases like CFCs into the atmosphere. Some of these trace gases, when at higher concentrations, are harmful to both humans and the environment.

Some air pollutants, particularly some non-metal oxides, return to Earth in the form of acid rain. This low pH rain damages crops and forests, corrodes statues and buildings, and can kill fish and other animal and plant life in lakes and streams.

In this experiment you will investigate the properties and reactions of metal and non-metal oxides. Many of these oxides are poisonous and pose a hazard to your health. It is imperative that you follow the safety procedures closely and carry out the procedures in an appropriately ventilated area or preferably in a fumehood.

Notes

Equipment

deflagrating spoon (×2 or thoroughly clean the spoon after each use)
gas jar (×4 or thoroughly clean the jar after each use) and lid
a sheet of white A4 paper (no ink)
distilled water
crucible
Bunsen burner
magnesium ribbon (about 2 cm)
solid sulfur (about 1 g)
crucible tongs
blue litmus paper (four small pieces)
red litmus paper (one small piece)
Universal Indicator ™ solution
watch glass or Petri dish
matches

Deflagrating spoon

Procedure

Your teacher may prefer to carry out this experiment as a demonstration in a fume cupboard.

SAFETY NOTE:

Burning sulfur

- **Do not touch the hot end of the deflagrating spoon with your hands and do not place the hot spoon on an unprotected bench.**
- **Sulfur dioxide is a poisonous gas that must be handled with care.**
- **Do not breathe the SO_2: it has a choking odour.**
- **This experiment should be carried out in a fumehood.**
- **If asthmatics are present it must be carried out in a fume cupboard.**

Experiment 33: Metal and non-metal oxides

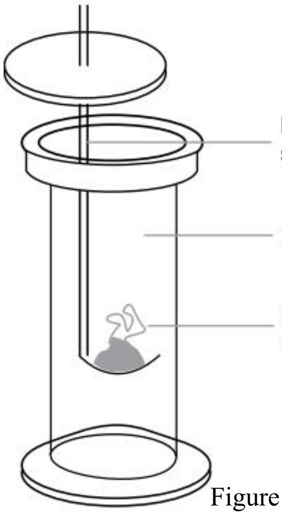


Figure 33.1
Burning sulfur

Sulfur dioxide

1. Set up the equipment to burn the sulfur and to collect the non-metal oxide, sulfur dioxide as shown in figure 33.1. Draw up a table to record your results and observations.
2. Carry out the experiment in a fumehood on a protective mat.
3. Pour a few milliliters (5 mL) of distilled water into the bottom of the gas jar. Add a drop or two of Universal Indicator ™ solution or a small piece of red and blue litmus paper.
4. Test your deflagrating spoon to make sure that when placed into the gas jar it seals without the spoon touching the water.
5. Place a small amount of sulfur on the deflagrating spoon.
6. Light a Bunsen burner and hold the deflagrating spoon of sulfur above the flame. Allow the sulfur to start burning. Once alight place the spoon into the gas jar so that the jar is sealed.
7. Once the sulfur has stopped burning carefully swirl the water around in the gas jar so that it dissolves the sulfur dioxide gas produced.
8. Record all observations: colour change of indicator, colour of gas and carefully detect its odour.
9. Set aside in the fumehood until all the gas is removed.

Notes

Carbon dioxide

1. Draw up a table to record your results and observations.
2. Loosely crumple half a piece of white A4 paper, so that it will drop to the bottom of a gas jar without becoming jammed.
3. Hold the paper with metal tongs, light the paper with a match and quickly drop the burning paper into the gas jar and cover it with the lid.
4. Once the paper has stopped burning and the jar has cooled pour about 5 mL of water into the jar. Replace the lid and swirl the water to dissolve the carbon dioxide gas.
5. Add a few drops of Universal Indicator ™ or a piece of blue litmus paper.
6. Record all observations and colour changes.
7. This can also be done with a burning candle in place of the paper.

Magnesium oxide

SAFETY NOTE:
Care must be taken when burning magnesium ribbon. The bright light produced can harm the retina in your eye, so do not look directly at the burning metal.

1. Hold the magnesium ribbon with the metal tongs in a Bunsen burner flame. Burn the magnesium and collect the metal oxide, magnesium oxide, in a crucible. Draw up a table to record your results and observations.
2. Add a drop or two of Universal Indicator ™ solution to the white magnesium oxide powder.
3. Record all observations: colour change of indicator, colour of flame and solid and gas produced.

Processing of results and questions

Notes

1. Write an equation for the combustion of each of the solids (S, C and Mg).

2. Classify the oxides produced as acid or basic. Describe any trend observed in the acidic or basic properties of metal and non-metal oxides.

3. From your observations what can you say about the solubility of each of the oxides in water? Write equations to show the reactions that occur when each of the oxides dissolves in water.

4. Nitrogen dioxide is a poisonous brown gas (see figure 33.2). It can be formed in the laboratory by the reaction of concentrated nitric acid and copper metal. (Your teacher may demonstrate this reaction in a fume cupboard for you.)
   a) If the nitrogen dioxide was dissolved in water and Universal Indicator ™ added what colour change would you expect?
   b) Explain, with the support of equations, what caused the Universal Indicator ™ colour change.

Figure 33.2
Nitrogen dioxide gas

5. Metals are added to fireworks to give the bright light and colours you see when they explode. Most of these metals are much more difficult and dangerous to burn than magnesium.
   a) Will the oxides of these metals form acidic or basic oxides?
   b) If the metal oxides were dissolved in water and Universal Indicator ™ added what colour change would you expect?
   c) Explain, with the support of equations, what caused the Universal Indicator ™ colour change.

Investigation 34: Acid rain

Notes

IMPORTANT
Before you attempt this investigation you should complete one or more of Experiments 31, 32 and 33 .

Acid rain
The main contributors to acid rain are sulfur dioxide; oxides of nitrogen; and, to a lesser extent, carbon dioxide. These compounds are produced from the burning of fossil fuels. They are gases and can be carried many kilometres in the atmosphere before dissolving in raindrops, forming acids, and falling as rain or snow. Acid rain has been shown to have adverse effects on forests, waterways and soils. It kills insect and aquatic lifeforms, damages buildings, and impacts on human health.

The tasks

Plan and conduct an investigation that models the chemistry of the formation of acid rain; its effect on a metal carbonate of your choice, and the rate at which your acid rain reacts with this metal carbonate.

Background research

1. Research the properties of metal oxides and hydroxides, and non-metal oxides and hydroxides.
2. Research the formation of acid rain, identifying the chemicals and reactions involved, and the typical pH of acid rain.
3. Research the chemistry of the reactions between acid rain and metal carbonates.

Equipment

Write a list of equipment and chemicals that you would require to form a sample of acid rain, and test its effects on a metal carbonate.

Planning the investigation

1. Plan how to form your own acid rain in the laboratory.
2. Plan a laboratory investigation to test the effects of your acid rain on a metal carbonate and to determine the rate of reaction (dissolution) between the metal carbonate and acid rain.
3. Write out each procedure, list the required chemicals and equipment, and outline the safety procedures needed to complete the tasks.
4. Have your teacher check your plans before proceeding.

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

Conducting the investigation

Conduct the investigation, collecting and recording measurements, making and recording observations and entering the results in a table as you proceed.

Processing the data

1. Write balanced equations for any reactions observed.
2. Use your results to determine the rate at which the metal carbonate dissolves when under the influence of acid rain.

Evaluating the investigation

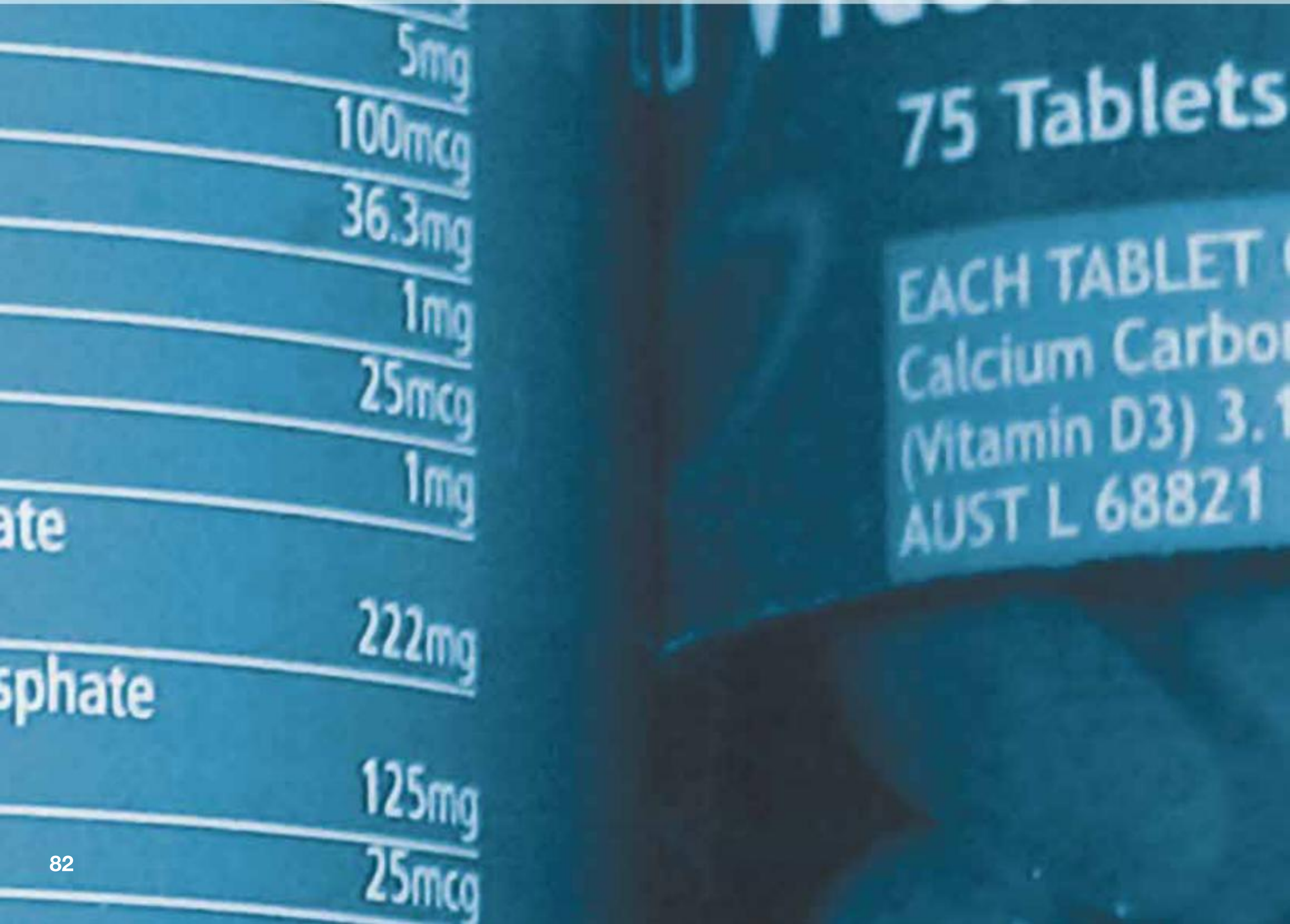
1. Discuss and explain differences between the actual dissolution rates of, say, a marble statue ($CaCO_3$), and your experimentally determined dissolution rate of the metal carbonate.
2. Write an evaluation of your experimental design, focusing on any improvements that you could make so that:
   - the accuracy of the results could be improved;
   - it would better model the real world.

Notes

Energy changes and rates of reaction

The experiments and investigations in the energy changes and rates of reaction section of *Exploring Chemistry Year 11: Experiments, Investigations and Problems*, provides opportunities for students to explore:

- **enthalpy changes in chemical reactions and phase changes through observable changes in the temperature of the surroundings and/or the emission of light**
- **endothermic and exothermic reactions**
- **how food and fuels (fossil fuels including coal, oil, petroleum and natural gas and biofuels including biogas, biodiesel and bioethanol) can be compared in terms of their energy output, suitability for purpose, and the nature of products of combustion**
- **the rate of chemical reactions by measuring the rate of formation of products or the depletion of reactants**
- **the effects of concentration, temperature, pressure, the presence of catalysts and surface area on the rate of chemical reactions**



Reactions and energy

Exploding Silos
Silos are large concrete cylinders commonly used for bulk storage of grain, coal, cement, wood products, flour, sugar and certain metals. The dust of most of these materials can explode. The air inside the silo becomes mixed with fine particles, such as grain dust in a concentration that makes it highly flammable. A spark can trigger an explosion of the fine dust cloud.

I said put the milk in the fridge!
Ever been told to put the milk back in the refrigerator after using it otherwise it will go sour?
We are surrounded by chemical reactions. Some are fast, like the gas burning in a BBQ, and some are slow, like the gradual corrosion of the hotplate and metal surrounds of the BBQ. Sometimes we want to make reactions happen more quickly, like when we cook chips, but sometimes we want to slow down the chemical processes. Milk going sour is a chemical reaction that we want to slow down. All chemical reactions are made to go faster by adding heat. So to slow down the souring of milk, do as you are told and put it back in the fridge.

Catalytic converters
Catalytic converters in cars normally contain the metals platinum and rhodium as surface catalysts. The reactions that they speed up are those that remove unwanted and toxic fumes, such as nitrogen monoxide (NO), carbon monoxide (CO) and unburnt octane (C_8H_{18}), from the car exhaust.
Adapted with the permission of WestOne Services (Chemistry Unit 2A -SCIENCE857)

Respiration
Our bodies rely on the breakdown of carbohydrates to glucose ($C_6H_{12}O_6$), which then undergoes the exothermic respiration reaction below to provide us with energy.
$C_6H_{12}O_6(aq) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(\ell)$ $\Delta H = -2\ 880\ kJ\ mol^{-1}$

Our bodies manage this energy by using the chemical adenosine diphosphate (ADP). The energy released from this respiration reaction will allow the ADP to be converted into adenosine triphosphate (ATP), as shown below.
$ADP + PO_4^{3-} \leftrightarrows ATP$ $\Delta H = +30.6\ kJ\ mol^{-1}$

The ATP can than convert back to ADP later, releasing this energy to the part of the body where it is required. This is a complex process, but it shows how an exothermic reaction can lead to an endothermic reaction, allowing for the transfer of stored chemical energy within the body.
Adapted with the permission of WestOne Services (Chemistry Unit 2A -SCIENCE857)

Experiment 35: Exothermic and endothermic processes

Notes

Chemical reactions are accompanied by the release or absorption of energy. Reactions in which energy is released are described as **exothermic,** they lose heat and the surroundings become warm. Those in which energy is absorbed are **endothermic,** they 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.

In this experiment you will be observing and classifying chemical processes as endothermic or exothermic based on the temperature changes you observe.

Equipment

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]
- sodium acetate [$NaCH_3COO$]
- barium hydroxide [$Ba(OH)_2$]
- ammonium chloride [$NH_4C\ell$]
- sodium chloride [$NaC\ell$]
- ammonium thiocyanate [NH_4SCN]

SAFETY NOTE:

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

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.
3. Tabulate your results under the headings: solute, initial temperature, final temperature, and classify the reaction as exothermic or endothermic.

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 in a test tube. Stopper the test tube and shake gently until a solution is formed.
2. Remove the stopper and smell cautiously. Try to identify one of the products. Measure and record the temperature of the solution.

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.

Processing of results and questions

Write equations for the processes or reactions in parts A, B and C and classify each reaction as either exothermic or endothermic.

Experiment 36: Foods and fuels

Notes

This experiment is outside the requirements of Year 11 chemistry.

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 1g 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?

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.

In this activity you will experimentally determine the molar heat of this neutralisation reaction by measuring 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)

Experiment 36: Foods and fuels

Notes

Procedure

1. Place 50 mL of 1 mol L^{-1} hydrochloric acid 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.

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.

$$\text{Heat energy released} = \text{mass of solution} \times \text{specific heat of solution} \times \text{temperature change}$$

$$= 0.100\text{ kg} \times (4.18 \text{ x } 10^3 \text{ J K}^{-1}\text{ kg}^{-1}) \times (\Delta T \text{ K})$$

$$= \_\_\_\_\_\_\_ \text{J}$$

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.

Investigation 37: Designing a cold pack

Notes

Endothermic and exothermic reactions impact our energy needs, both biological and commercial and they have applications in the health industry. Cold packs are applied to sporting and other injuries to decrease the blood supply to the injured area by cooling it. They reduce bruising and swelling that may result from an injury. Some commercially available cold packs depend on an endothermic solution process. A typical cold pack consists of a sturdy, well-sealed plastic bag that contains a sachet of water and crystals of a solute. When required, the water sachet is punched or squeezed to break it open. The sealed plastic bag is then shaken thoroughly to dissolve the solute in the water. The manufacturers claim that the temperature of the resulting solution can be lowered to about 1°C for around 20 minutes.

The task

1. Determine the most suitable solute for use in a cold pack.
2. Determine the optimum ratio of solute to solvent that will produce the maximum cooling effect.

Equipment

thermometer (-10 to 100 °C)
beaker (100 mL)
test tubes (several)
test tube rack
balance
graduated cylinder (10 mL or 25 mL)
plastic teaspoon or spatula
distilled water
samples of the following solids:
- urea (about 1-2 g)
- sodium acetate [$NaCH_3COO$] (about 1-2 g)
- sodium chloride [$NaC\ell$] (about 1-2 g)
- calcium chloride [$CaC\ell_2$] (about 1-2 g)
- selected solid for part B (about 20 g)

Planning the investigation

Task 1: which solute is most suitable for use in a cold pack?

1. Briefly outline your plan to determine the most suitable solute for the cold pack. Your investigation should be carried out on a small scale using, for example, 1-2 g of solute and 5 mL of water. List the chemicals and equipment you need and identify the safety requirements.
2. Check your proposed plan with your teacher.

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

Conducting the investigation

Conduct the investigation, collecting and recording your results in a table as you proceed.

Investigation 37: Designing a cold pack

Notes

Processing the data

1. Write equations to illustrate the dissolving of the solutes and indicate whether each reaction was endothermic or exothermic.
2. Analyse your data and discuss your findings.

Planning the investigation

Task 2: what is the optimum ratio of solute to solvent that will produce the maximum cooling effect?

1. Write a hypothesis for the investigation.
2. State the independent and dependent variables.
3. Copy and complete in the following table.

	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 determine the ratio of solute to solvent that will produce the maximum cooling effect. Your investigation should be carried out on a small scale using, for example, about 5 mL of water in each trial. List the chemicals and equipment you need and identify the safety requirements. Remember that you may need to modify your plan after some preliminary trials.
5. Check your proposed plan with your teacher.

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

Conducting the investigation - preliminary trials

1. Conduct some preliminary trials to determine the quantities of solute and solvent you will use.
2. Describe what you learned from these initial trials and modifications that you made to your original plan.

Conducting the investigation - collecting data

1. Conduct the investigation, collecting and recording the data in a table as you proceed.
2. If time and materials allow, replicate the data collection and average the results.

Processing the data

1. Draw a graph to illustrate your data (plot the independent variable on the horizontal axis).
2. Discuss your data and relate them to your hypothesis.

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. 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.

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

3. How might cost factors affect the final design specification for the cold pack?

Notes

Investigation 38: Measuring the rate of reaction

Notes

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 from weeks to 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.

The task

Investigate the rate of reaction of calcium carbonate with hydrochloric acid.

Planning the investigation

1. Plan an investigation to measure the rate of reaction of calcium carbonate (marble chips, $CaCO_3$) with hydrochloric acid ($HC\ell$).
2. Write an equation for the reaction. Make a list of the variables that you could measure in order to determine the rate of this reaction. Select one variable to measure as the reaction proceeds. Design an experiment that will enable you to measure the rate of reaction.
3. Write out your proposed procedure, include a labelled diagram of the set up that you will use, a list of the chemicals and equipment that you may require and identify any safety requirements for your investigation.

Equipment

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

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

Conducting the investigation

1. Conduct the investigation, collecting and recording your observations of the reaction in a table as you proceed.
2. If time and resources allow, replicate the data collection.

Processing the data

1. You should graph your results.
2. Identify any trends and explain your data.

Evaluating the investigation

1. 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.
2. Discuss whether the data are sufficient to support your conclusion.

Experiment 39: Reactant concentration and rate of reaction

Notes

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 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

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. Copy the table below into your laboratory notebook and record your results directly into it.

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

Experiment 39: Reactant concentration and rate of reaction

Notes

2. 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 39.1 below.

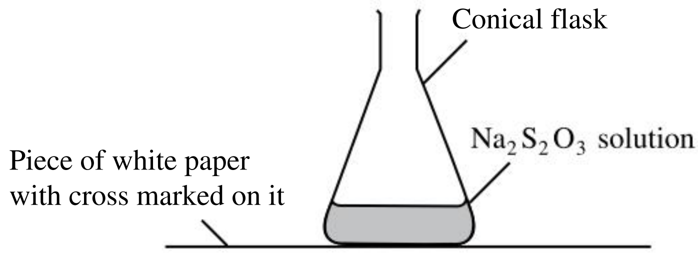


Figure 39.1

3. 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.
4. Note and record 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.
5. Repeat the experiment using various sodium thiosulfate concentrations, made up as indicated in the table on page 91.

Processing of results and questions

1. Calculate 1/time for each experiment and enter the results into the table.
2. Plot a graph of 1/time (a measure of the reaction rate) against sodium thiosulfate concentration.
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?

Experiment 40: Temperature and 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 purpose of this experiment is to investigate the effect of temperature on reaction 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.

Equipment

hydrochloric acid [$HC\ell$] 2 mol L^{-1} (30 mL)
conical flask (100 mL)
thermometer (-10 to 110 °C)
Bunsen burner, tripod and gauze mat
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)
distilled water
test tube
stopwatch
paper towel

SAFETY NOTE:
- **Avoid inhaling sulfur dioxide evolved while waiting for the cross to disappear. If asthmatic, carry out 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. Measure 5 mL of 2 mol L^{-1} $HC\ell$ 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 will need to measure the average temperature for the reaction.
7. Construct a table showing the temperature, times taken for the cross to disappear, and 1/time.

Processing of results and questions

1. Plot a graph of 1/time against temperature.
2. What effect does increasing the temperature have on the reaction rate?
3. Suggest two factors that contribute to the change in reaction rate with temperature.
4. Identify two random errors in this experiment. How could these errors be minimised?

Notes

Experiment 41: Catalysts and rate of reaction

Notes

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.

In this experiment the effect of catalysts on the rate of three different reactions will be observed. The three reactions are:

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 almost colourless manganese(II) ion.
The equation is:

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

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:

$$C_4H_4O_6^{2-}(aq) + 5\ H_2O_2(aq) + 2\ H^+(aq) \rightarrow 4\ CO_2(g) + 8\ H_2O(\ell)$$

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 [$CoC\ell_2\cdot 6H_2O$] 10% solution (1 mL)
distilled water
beakers (two 250 mL and two 100 mL)
thermometer (10 to 110 °C)
plastic teaspoon

Procedure

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

SAFETY NOTE:

- **Oxalic acid is poisonous and should not be allowed to come in contact with your skin.**
- **Handle the $H_2C_2O_4$ solution carefully.**

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. Is there any evidence of reaction?
2. To one of the beakers add about 1 mL of saturated manganese(II) sulfate solution and record your observations.

Procedure

B. decomposition of hydrogen peroxide

SAFETY NOTE:

- **Hydrogen peroxide can cause burns if it comes in contact with your skin.**
- **Handle the H_2O_2 solution carefully.**
- **Add only a few grains of manganese(IV) oxide to the H_2O_2 otherwise the reaction will be too vigorous.**

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.

Procedure

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.
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 of results and questions

1. Which species was acting as a catalyst in each reaction?
2. In reaction **C**, what evidence is there that the catalyst was not consumed in the reaction? Is there any evidence that the catalyst took an active part in the reaction?

Notes

Investigation 42: The decomposition of hydrogen peroxide

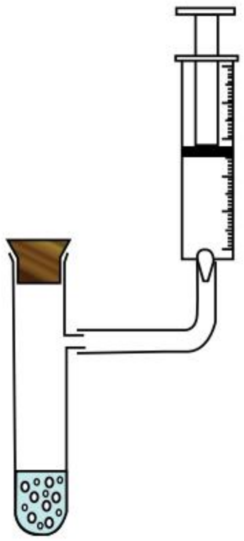


Figure 42.1: Decomposition of hydrogen peroxide

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 42.1

The task

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

Notes

Equipment

test tube rack, plastic syringe (30-50 mL), dropper, stop watch, 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
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)

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

Notes

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. Copy and fill in the following table.

	Variable/s	Unit/s	How the variable/s 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 commence.**
- **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 avoid 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.

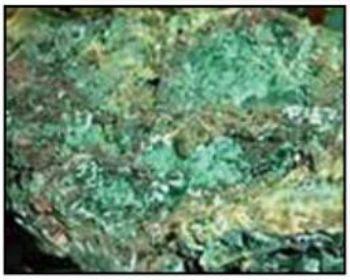
Processing the data

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

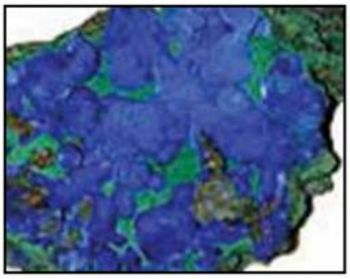
Evaluating the investigation

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.

Investigation 43: Metallurgist needed - copper leaching plant


Malachite


Polished malachite


Azurite (on malachite)


Heap preparation in Chile; note the pool liner and the general size of the heap


Reticulation system on top of a heap

An extractive metallurgist determines the most efficient and environmentally friendly way to extract a mined metal from the rock and ore in which it is trapped. This investigation places you in the role of an extractive metallurgist. You are required to examine the process of leaching of copper from an oxide ore. It consists of two separate tasks.

The task

1. Using the copper leaching method determine the percentage of copper in the ore sample. Obtain results by manually comparing ore copper concentration with a set of standard copper solutions.

2. Plan and conduct an investigation that explores how varying one factor in the process will improve the speed of production of copper. Use the ore sample provided for task 1.

Background to the copper heap leach process

The modern world uses around 10,000,000 tonnes of copper each year. There are two major types of copper minerals, sulfides and oxides. Once these ores are mined it is necessary to extract the metal from the rock in the most efficient and environmentally friendly way possible. The expert who undertakes this difficult task is called an extractive metallurgist (sometimes a metallurgical engineer or just a metallurgist) and these scientists are vital to any successful operation.

Sulfides cannot easily be processed by solution chemistry and have to be treated by a high temperature process (pyrometallurgy) producing sulfur dioxide (SO_2), which has to be completely captured for environmental preservation. The copper oxide minerals are much more readily processed by solution chemistry (hydrometallurgy) as they react readily with acids in the same way as all metal oxides.

$$\text{acid} + \text{oxide} \rightarrow \text{water} + \text{salt}$$

$$2\,H^+(aq) + CuO(s) \rightarrow H_2O(\ell) + Cu^{2+}(aq)$$

The major copper oxide minerals present are actually copper hydroxide carbonates called malachite ($Cu_2(OH)_2CO_3$) and azurite ($Cu_3(OH)_2(CO_3)_2$). Malachite is used in jewellery and is a bright green colour whilst azurite is named for its deep blue colour. These minerals react with acids as you would expect for hydroxides and carbonates forming water and carbon dioxide in addition to a metal salt which is soluble in acid.

$$Cu_2(OH)_2CO_3 + 2\,H_2SO_4 \rightarrow 2\,CuSO_4 + 2\,H_2O + CO_2$$

Copper ores are not composed entirely of one mineral and the copper mineral is usually present in relatively small concentrations. A typical ore contains around 1-2% copper, so every tonne of ore only contains 10-20 kg copper. The remaining mass is composed of both reactive (e.g. dolomite, $MgCa(CO_3)_2$) and unreactive minerals (e.g. quartz, SiO_2). These extra minerals are known as gangue.

Industrially, the copper ore is mined and stacked into big heaps on pads which are carefully prepared using extra thick pool liners to prevent acid leakage. The heaps are then irrigated with acid, using a reticulation system similar to that used in gardens, to dissolve the copper. The solution percolates through the heap and is collected in drains and fed to a pond where the solution is treated to recover the copper as metal. After the copper has been chemically removed, the remaining solids, called tailings, are stored so they do not harm the environment and rehabilitated so the landscape is the same as it was before mining.

Equipment

a clean 600 mL plastic drink bottle with cap
75 g crushed ore*
Note: make up synthetic ore one week prior to lab.
sulfuric acid [H_2SO_4] 1% (300 mL)
copper(II) sulfate-5-water [$CuSO_4{\cdot}5H_2O$] (5 g)
6 x 50 mL screw-cap sample bottles
10 mL pipette
pipette fillers
50 mL volumetric flask and stopper
several 100 mL volumetric flasks and stoppers
distilled water
magnifying glass or low powered microscope

* **A synthetic ore** can be made by mixing 10 g of copper hydroxide carbonate (or 10g copper sulfate + 5 g calcium carbonate) with 80 g of fine sand and 10 g of cement powder. Add water, mix thoroughly, spread out and leave to set. Once set, the ore can then be crushed and ground as required to give a range of particle sizes which can be used in leaching experiments.

Conducting the investigation

Task 1

The copper ore leach process – part 1

Measure out approximately 75 g of crushed ore (or ore substitute). Carefully pour 300 mL of 1% H_2SO_4 solution into a clean 600 mL plastic bottle. Now add the crushed ore using a funnel. Cap the bottle tightly and shake the mixture for 45-60 s, occasionally opening the cap slightly to release the gas, until all of the powder has been exposed to the acid solution. Set the container aside to settle.

If possible look at the larger ore particles under a low power microscope or magnifying glass to see how the copper compounds are distributed within the ore.

Preparing a set of standard solutions

Collect and clean a set of six 50 mL screwcap sample bottles and a suitable rack, one 50 mL and several 100 mL volumetric flasks and a 10 mL and 25 mL pipette.

1. Measure out 3.93 g of copper(II) sulfate pentahydrate and add to a 100 mL volumetric flask. Add about 80 mL distilled water, cap the flask and shake well assuring that all the salt is dissolved. Once dissolved, carefully add more distilled water until the bottom of the meniscus just touches the mark etched on the neck of the flask. Pour about 40 mL of this solution into a sample bottle and label it 1.0 g L^{-1} Cu.
2. Carefully pipette 10.0 mL of this solution into a clean 100 mL volumetric flask. Add distilled water until about 1 cm below the mark etched on the flask neck, cap the flask and shake well. Add more distilled water until the bottom of the meniscus just touches the line. Shake and pour about 40 mL of the solution into a fresh sample bottle labelling it 0.50 g L^{-1}.
3. Carefully pipette 25.0 mL into a clean 50 mL volumetric flask. Add distilled water until about 1 cm below the mark etched on the flask neck, cap the flask and shake well. Add more distilled water until the bottom of the meniscus just touches the line. Shake and pour about 40 mL of the solution into a fresh sample bottle labelling it 0.10 g L^{-1} Cu for the 50 and 100 mL flask solutions respectively.

Notes

Investigation 43: Metallurgist needed - copper leaching plant

Notes

3. Using the procedure outlined in step 2 prepare solutions containing 0.050 and 0.010 g/L Cu by diluting the 0.10 g L^{-1} solution. Rinse the volumetric flasks with distilled water if you are reusing them.
4. Perform one more twofold dilution using the 0.050 g L^{-1} solution. You have now prepared a set of standard solutions of decreasing copper content. Using these you can estimate the copper content of unknown solutions by comparing the intensity of the blue colour.
   If you have time, further standard solutions can be made up to enable better colour matching.

The copper ore leach process - part 2

1. By the time you have prepared your set of standard solutions (colour set), the ore sample in the plastic bottle should have had time to settle to the bottom. Carefully lift the container and report your observations, paying particular note to any stratification present.
2. Slowly decant the clear liquid layer into its own sample bottle, being careful not to disturb any solids that have settled to the bottom of the container. Place a piece of white paper behind the rack containing the standard solutions. Compare the colour of the supernatant just decanted and record your estimate of the concentration of the sample solution.
   If possible retrieve some of the larger ore particles and observe them under a microscope to see how the distribution of copper has changed due to leaching.
   The use of the standard solutions demonstrates the principle upon which a spectrophotometer is based. The only major difference between you and a spectrophotometer is the spectrophotometer uses filters to select specific wavelengths of light to measure.

Processing the data

Task 1

1. Take a photograph or draw and colour your set of standard solutions of decreasing copper content. Using these you can estimate the copper content of unknown solutions by comparing the intensity of the blue colour. Record the value obtained.
2. Calculate the percentage of copper in the ore sample.
3. Comment on the richness of your sample. Where does it fit in the typical range of ore copper concentration?
4. Why did the particles settle into layers?
5. Why was the final solution not clear?

Planning the investigation

Task 2

Commercial copper heap leaching has to produce copper that is worth more than the total cost needed to obtain it. The difference between the cost and the value is the company's profit. Extractive metallurgists spend a great deal of time and effort trying to maximise the output (product) while minimising the input (materials, energy and operator costs). Most of this work is based on the fact that copper ions are locked within the minerals within the ore. The question then becomes 'what factors can a metallurgist control that will allow for more efficient release of product (greater quantity, greater speed or both increased quantity and speed)?

Notes

There are several factors that affect chemical reaction rates of this process including:

- the minerals present in the ore;
- the distribution of copper in the ore ;
- the concentration of acid;
- ratio of acid solution to ore;
- the temperature ;
- the pressure of the reaction vessel;
- the particle size of the reactants.

The plant operators cannot affect all of these factors. In this copper heap leaching process some of these cannot be varied easily. The characteristics of the ore for example were determined by nature as the ore body was forming. Pressure is difficult to vary and any pressurised vessel is potentially explosive so most plants operate at ambient pressure.

1. Plan an investigation that explores how varying one of the other factors will improve the speed of production.
2. Write out your proposed procedure, list the chemicals and equipment you need and identify any safety requirements.
3. Check your proposed procedure with your teacher.

Conducting the investigation

Task 2

1. Conduct the investigation collecting and recording the data in a table as you proceed.
2. If time allows, replicate the data collection and average your results.

Processing the data

Task 2

1. Represent your findings in a graphical or pictorial way.
2. Summarise your results and describe any trends in the data.
3. Explain how your results can be used to maximise the plant process.

Evaluating the investigation

Task 2

1. 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.
2. Discuss whether the data are sufficient to support your conclusion about maximising the process.

Extra questions

1. Predict how particle size will affect the rate of leaching.
2. Predict how temperature will affect the rate of leaching.
3. Predict how the concentration of the acid will affect the rate of leaching.
4. Predict how the concentration of copper in the ore sample will affect the rate of leaching.
5. Predict how the distribution of copper within the ore particles affects leaching.
6. What other solutions could you use to dissolve the malachite?
7. What would happen if you used another acid?
8. What do you need to do to ensure successful copper dissolution in an ore that has a lot of dolomite, ($CaMg(CO_3)_2$) present?
9. How does a spectrophotometer work? Use a simple diagram to explain this.

Organic chemistry

The experiments and investigations in the organic chemistry section of *Exploring Chemistry Year 11: Experiments, Investigations and Problems*, provides opportunities for students to explore:

- **elemental carbon and its allotropes, including graphite, diamond and fullerenes, in terms of their significantly different structures and physical properties**
- **the properties of covalent molecular substances**
- **hydrocarbons, including alkanes, alkenes and benzene by making models**
- **molecular models to show the arrangement of atoms and bonding in covalent molecular substances and to help draw structural formulae**
- **the characteristic reactions of alkanes, alkenes and benzene such as combustion, addition reactions for alkenes and substitution reactions for alkanes and benzene**

Carbon - friend or foe?

Carbon, which forms the basis of the biological world, lies at the heart of some of science's most pressing current environmental issues. Carbon's unique properties and rich diversity of compounds certainly ranks it amongst the most fascinating of the elements.

Carbon can form four bonds and has the unusual capacity of bonding strongly to itself creating long chains and ring structures. It can also form strong bonds with other non-metal elements such as hydrogen, oxygen, nitrogen, sulfur and the halogens. Carbon's unique bonding properties have lead to the existence of several million different carbon compounds, including biomolecules, those responsible for maintaining and reproducing life, such as sugars, fats, amino acids and proteins. So diverse and unique is this element, it has its own branch of chemistry, the study of carbon-containing compounds and their properties, called organic chemistry.

Organic chemistry is vital to our understanding of living systems and continues to be significant in contributing to the products of modern life such as plastics, drugs and synthetic fibres. The energy that powers our civilisation is based on organic materials contained in the fossil fuels, coal, oil and natural gas.

Burning fossil fuels produces energy:
Natural gas:
$CH_4 + 2\,O_2 \rightarrow CO_2 + 2\,H_2O$ + Energy
Petrol:
$2\,C_8H_{18} + 25\,O_2 \rightarrow 16\,CO_2 + 18\,H_2O$ + Energy

Burning carbon based fuels

State at S.T.P.	Fuel	% by mass of chemical components (these may vary)	Mass of fuel required to release 1000 kJ of energy by combustion
solid	Black coal from Rockhampton (QLD)	87% carbon (C)	36.6 g
	Black coal from Collie (WA)	73% carbon (C)	50.8 g
	Brown coal from Hazelwood(VIC)	30% carbon (C)	101 g
gas	LPG	60% propane (C_3H_8) 40% butane (C_4H_{10})	20.2 g
	Natural gas	95% methane (CH_4) 2.5% ethane (C_2H_6)	18.6 g
liquid	Heating oil	70% $C_{16}H_{34}$ 30% $C_{17}H_{36}$	21.6 g
	Ethanol	100% C_2H_6O	27.1 g

Adapted with the permission of WestOne Services (Chemistry Unit 2A -SCIENCE857)

Benzene - pleasant smelling but deadly
Prior to the 1920s benzene was used as an after-shave lotion because of its pleasant smell. During its early history it was also used to decaffeinate coffee, in paint strippers, rubber cement, stain removers and hydrocarbon containing products. It was an important industrial solvent. Benzene's toxicity, as a cancer causing material, soon became obvious though, and thankfully its use as a solvent has been discontinued. Other less toxic solvents, especially toluene (methyl benzene) have since replaced it. Toluene has similar physical properties but is not as carcinogenic as benzene.

Experiment 44: Molecular models

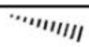
Notes

Organic molecules that have the same molecular formula but different physical structures are called structural isomers

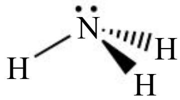
You can purchase petrol with different octane ratings such as 91, 95 and 98. The octane rating of petrol is its anti-knock rating or detonation resistance. ‘Detonation resistance’ is about burning more slowly; overly fast burning (detonation) creates a very fast expansion of gas, pushing unevenly on 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 **structural 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.

This experiment will help you to understand what isomers are and how 2,2,4-trimethylpentane and heptane are named and drawn. You will construct molecular models of, draw structures of and name some simple organic compounds, most of which are derived from natural gas and petroleum refining.

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	Prefix
1	meth-
2	eth-
3	prop-
4	but-
5	pent-
6	hex-
7	hept-
8	oct-

For example a three dimensional molecule like ammonia can be drawn as 

Where one hydrogen is in the plane of the page, one is into the page and one is out of the page.

Equipment

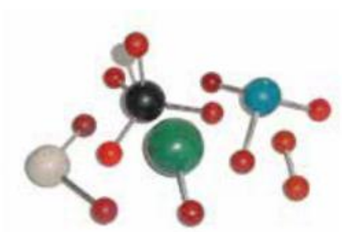
molecular model kit, or polystyrene balls and toothpicks

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 the molecule.



Notes

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 the model.

Dichloroethane [$C_2H_4C\ell_2$]

7. Remove two hydrogen atoms from your model 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 of the two models and label them 1,1-dichloroethane and 1,2-dichloroethane, as appropriate.

Hexane [C_6H_{14}] and Cyclohexane [C_6H_{12}]

11. Construct a model of straight chain hexane.
12. Write the structural formula for straight chain hexane.
13. Write the structural formulas of four additional isomers with formula C_6H_{14}. Two of these isomers have five-carbon chains and two have four-carbon chains. Write the systematic name for these isomers.
14. Arrange the model of 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.
15. Draw a diagram of the structure.

Procedure **B. alkenes**

Ethene [C_2H_4]

1. Construct a model of this molecule.
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 model.

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 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 isomers and write their names.

Experiment 44: Molecular models

Notes

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 the model.

Procedure

C. Alkynes (optional extra)

Ethyne [C_2H_2]

1. Construct a model of this molecule.
2. Draw a diagram of the model.

Processing of results and questions

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

Investigation 45: Combustion of hydrocarbons

Hydrocarbons are used as fuels mostly by burning them in air or pure oxygen to produce heat. Each time you use a Bunsen burner you are using the simplest of the hydrocarbons, methane, as a fuel.

The task

Investigate the products of the combustion reaction of a hydrocarbon. You are required to confirm that the products you predict will form are, indeed, present.

Predicting the products

1. Write a balanced equation for the combustion of methane.
2. Write a list of the physical and chemical properties for each of the products.

Planning the investigation

1. For each product devise a method for collecting a sample for testing from a source where hydrocarbons are burnt such as a burning Bunsen burner, a candle or from some other source.
2. For each product choose one property or more that you could use to conclusively identify the substance.
3. Write a list of equipment you will require for collecting and testing each of the products.
4. Write a description of the procedure you will use to collect and test each product. Be sure to include all safety precautions.

Conducting the investigation

Collect each of the products and conduct the tests to identify them. Write a description of the observations or measurements you made.

Processing the data

1. Explain how your observations or measurements allowed you to conclusively identify each product.
2. Write a description of any difficulties you encountered during the collection or testing of the products.

Evaluating the investigation

Describe how you could change your procedure to improve the collection and testing processes.

Notes

Investigation 46: Reactivity of hydrocarbons

Notes

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.

The tasks

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 (four)

dropper

the representative 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)

potassium permanganate solution [$KMnO_4$] 0.01 mol L^{-1} (4 mL)

sulfuric acid [H_2SO_4] 2 mol L^{-1} (2 mL)

bromine water [Br_2] (5 mL)

Procedure

A. reaction of hydrocarbons with acidified permanganate solution

SAFETY NOTE:
- **Cyclohexane, cyclohexene and toluene are poisonous and must be handled with care.**
- **Do not let these liquids come in contact with your skin and avoid inhaling their vapours.**
- **Carry out the experiment in a fumehood.**

1. Into three separate, labelled test tubes place 1 mL of cyclohexane, cyclohexene and toluene respectively.
2. Into a separate test tube add 4 mL of 0.01 mol L^{-1} $KMnO_4$ and 2 mL of 2 mol L^{-1} H_2SO_4.
3 Add 1 mL of the acidified $KMnO_4$ solution, prepared in procedure #2, 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 a suitable table.

Procedure

B. reaction of hydrocarbons with bromine water

SAFETY NOTE:

- **Bromine water is poisonous and corrosive and must be handled with care.**
- **Do not breathe the vapour given off by the bromine water.**
- **If bromine water comes in contact with your skin, immediately wash the affected area with copious quantities of 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 in each case. Write your observations in a suitable table.

Processing the data

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.

Conclusions

Task 1

Assuming the reactions of these hydrocarbons are typical of the classes of compounds to which they belong, what can you say about the relative reactivities of the three classes?

Task 2

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

Notes

Section 2: Chemical understanding and problem-solving

***Exploring Chemistry Year 11: Experiments, Investigations and Problems* provides opportunities for students to continue developing their chemical understanding and problem-solving skills in chemistry.**

Section 2: Chemical understanding and problem solving involves applying chemical knowledge and understanding of chemical concepts to situations, questions and problems, with a focus on the quantitative aspects of chemistry.


iron
26
Fe
55.845
magnesium
12
Mg
24.305
silver
47
Ag
107.87
barium
56
Ba
137.33
platinum
78
Pt
195.08
zirconium
40
Zr
91.224
mercury
80
Hg
200.59

Measurements in chemistry

Commonly encountered quantities and units used 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 atmosphere millimetres of mercury	kPa atm mm Hg	All sets of units are useful 1 atm = 101.3 kPa = 760 mm Hg
Relative atomic mass (atomic weight)					
Relative molecular mass (molecular weight)	These are all dimensionless quantities				
Relative formula 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 $= 1000\ cm^3$ $= 1 \times 10^{-3}\ m^3$

Set 1: Scientific notation and unit conversions

Quantities and units In chemistry

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 $= 1 \times 10^3$ g

then: 9.213 kg $= 9.213 \times 10^3$ g

$= 9213$ g

Set 1: Scientific notation and unit conversions

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: $\frac{1.8 \times 10^8}{6 \times 10^5} = \frac{1.8 \times 10^{8-5}}{6}$ $= 0.3 \times 10^3$ $= 3 \times 10^2$

Notes

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

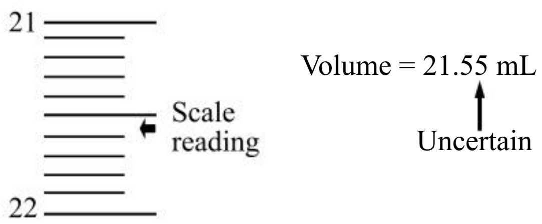
Set 2: Significant figures

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: Significant figures

Rounding off

Rounding is used to reduce the complexity of a number when it is written with more digits than are wanted or justified. The last digit written should give the best approximation of the number as it was before rounding. If the number before rounding is as close to one number as another, the one ending with an even digit is chosen, zero being regarded as even (example 6).

The following examples illustrate rounding to **three significant figures**.

1.	1.294	rounds to	1.29
2.	8.12349	rounds to	8.12
3.	0.01249	rounds to	0.0125
4.	18.951	rounds to	19.0
5.	7.1451	rounds to	7.15
6.	7.145	rounds to	7.14

Rounding in multi-step calculations

When performing calculations requiring several steps, only round to the appropriate number of significant figures after the final step. This avoids possible errors that can accumulate during a calculation if rounding to the strict number of significant figures is carried out at each step.

Notes

Set 2: Exercises

1. How many significant figures are given in each of the following
   (a) 454 (b) 0.3750 (c) 100.10 (d) 6.07×10^{-11} (e) 6000 (f) 0.0003 (g) 14.0 (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.

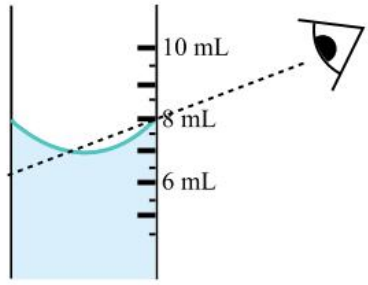
Systematic errors are a result of flaws in the experimental method or apparatus that lead to a result that is always either above or below the true value.

Errors are discussed in detail on page ix and x.

Notes

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.

Set 3: Random and systematic errors

Notes

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 reproducibility 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 estimated 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 work out the enthalpy change (ΔH).

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

Notes

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	

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.

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.

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 sig fig)

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.

Mixtures, kinetic theory and gases

The problem sets in the mixtures, kinetic theory and gases section of *Exploring Chemistry Year 11: Experiments, Investigations and Problems*, provides opportunities for students to explore:

- **pure substances, materials with distinct measurable properties, including melting and boiling points, reactivity, hardness and density**
- **mixtures, materials with properties dependent on the identity and relative amounts of the substances that make up the mixture**
- **how differences in the physical properties of substances in a mixture, including particle size, solubility, density, and boiling point, can be used to separate them**
- **chromatography techniques and how the data collected can be used to determine the composition and purity of substances**
- **the kinetic theory and how it can be used to explain the behaviour of gaseous systems**
- **the behaviour of an ideal gas, including the qualitative relationships between pressure, temperature and volume**

Set 4: Mixtures

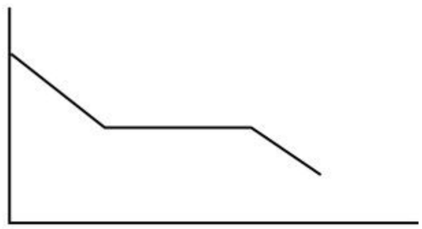
Physical separation techniques are used to purify components or to determine the composition of a mixture. Separation techniques rely on the individual substances having different physical properties. These properties include solubility, level of attraction to a stationary phase, particle size, density, 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 4: Exercises

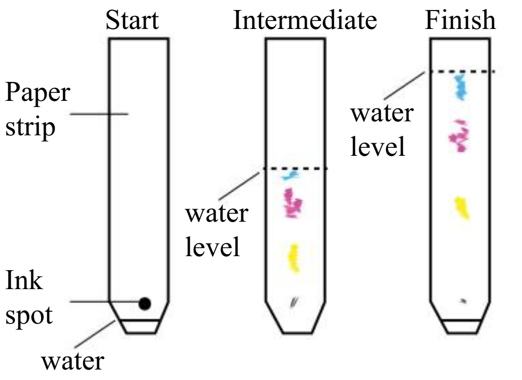
1. Describe using examples the differences between a homogeneous and a heterogeneous mixture.

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.
   (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.

Chromatographic separation of black ink



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.

Notes

Set 5: Kinetic theory

Notes

Set 5: Exercises

1. In Experiment 7 you examined the relationship between the pressure and volume of a gas as described by Robert Boyle. State Boyle's Law, give a mathematical relationship between pressure and volume and sketch a graph to show this relationship.

2. 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.

3. 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.

4. 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?

5. 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?

6. Explain why liquids exert a vapour pressure. Why is the boiling point of a solution higher than the boiling point of its solvent?

Set 6: Macroscopic properties of matter

Set 6: Exercises

Notes

1. Chemists have criteria for classifying matter as solid, liquid or gas, but some substances can fall into more than one category. Research the properties of substances such as glass, bitumen or supercritical carbon dioxide (used to make decaffeinated coffee). Write an argument for and against classifying each material into only one state of matter.

2. Electrolyte imbalances in the blood are closely examined during blood sample analysis. Abnormally high or low values of electrolytes such as sodium, potassium, magnesium or calcium ions can lead to vomiting and dehydration or more severe problems like kidney disease and heart failure. Draw up a table that summarises the different blood electrolytes and their importance, conduct research so you can include their concentrations.

3. Explain the following observations:
   (a) an aerosol container explodes when heated;
   (b) a partly inflated balloon shrivels up when cooled in liquid nitrogen yet returns to its normal shape and volume when allowed to warm to room temperature.

4. 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.

5. 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?

6. **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.

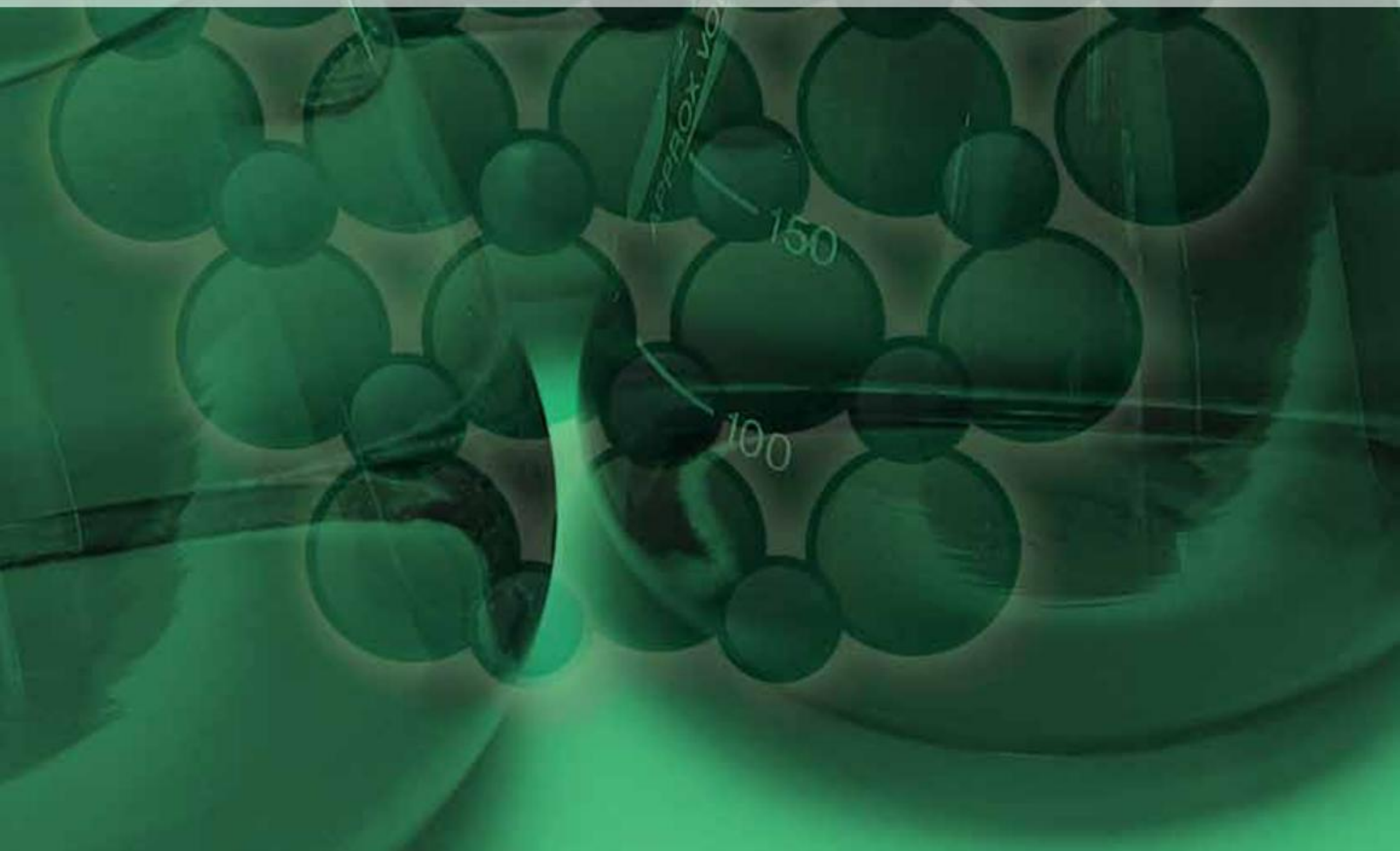
7. **SCUBA Diving:**
   SCUBA diving, a popular underwater pastime, depends on a supply of compressed air to allow breathing under water. Compressed air is air kept under a pressure greater than atmospheric pressure.
   (a) Why is the air in SCUBA tanks kept under pressure?
   (b) List the gases added 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.

8. Charles' Law also describes gas behaviour. Research and describe this law.

Atomic Structure and Bonding

The problem sets in the atomic structure and bonding section of *Exploring Chemistry Year 11: Experiments, Investigations and Problems*, provides opportunities for students to explore:

- **how the atomic model and models of bonding explain the structure and properties of elements and compounds**
- **how atoms can be modelled as a nucleus, surrounded by electrons in distinct energy levels, held together by electrostatic forces of attraction between the nucleus and electrons; the location of electrons within atoms can be represented using electron configurations**
- **analytical techniques such as flame tests and atomic absorption spectroscopy (AAS) and how they can be used to identify elements**
- **how observable properties, including vapour pressure, melting point, boiling point and solubility, can be explained by considering the nature and strength of intermolecular forces within a covalent molecular substance**
- **the type of bonding within ionic, metallic and covalent substances and how it explains their physical properties, including melting and boiling points, conductivity of both electricity and heat and hardness**



Set 7: 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.

Notes

Your curriculum authority may provide a list of elements and common compounds that you must know.

Set 7: 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) ruthenium
   (k) thorium
   (l) astatine
   (m) germanium
   (n) technetium
   (o) barium

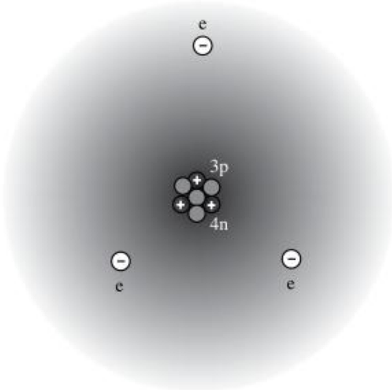
2. Write the names for the following symbols
   (a) Na
   (b) Re
   (c) $C\ell$
   (d) Au
   (e) Zr
   (f) Pu
   (g) Ce
   (h) As
   (i) Fe
   (j) Co
   (k) P
   (l) Kr
   (m) Zn
   (n) K
   (o) Sn

Set 8: Atoms and isotopes

Notes

Set 8: 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 was 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. Copy and 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 :

$$^{A}_{Z}E \quad \text{eg. } ^{63}_{29}Cu$$

For the element copper, identify:

E (element symbol): ________________________________

A (mass number): ________________________________

Z (atomic number): ________________________________

Notes

6. Copy and 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 9: Atomic structure and the Periodic Table

Notes

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 9: Exercises

1. Copy and 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 e^- or gains 3 e^-
8					

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	Atomic number	Number of protons	Number of electrons	Number of neutrons
Example	$^{19}_{9}F$	9	9	9	10
A	$^{16}_{8}X$				
B	$^{17}_{8}X$				
C	$^{16}_{8}X^{2-}$				
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?
   (a) 2, 7, 8, 1
   (b) 1, 8, 8
   (c) 2, 8, 7, 1

Notes

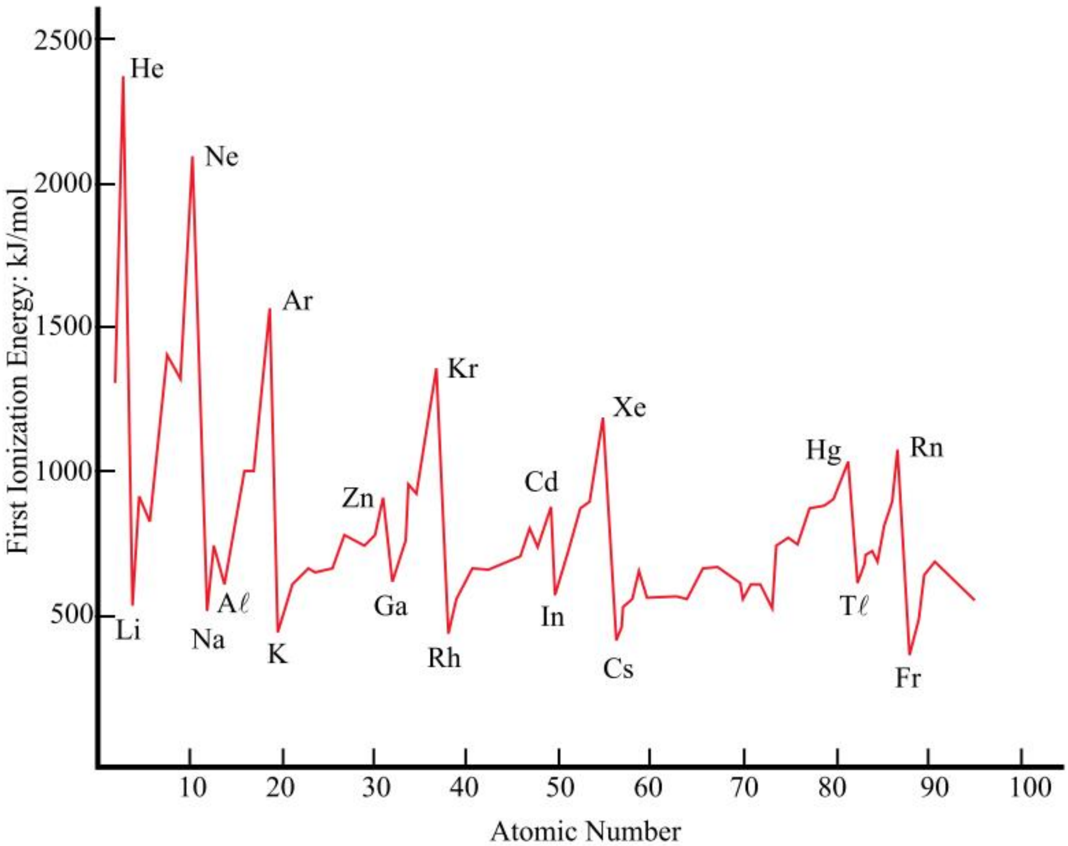
Set 10: Ionisation energy

Notes

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 10: Exercises

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

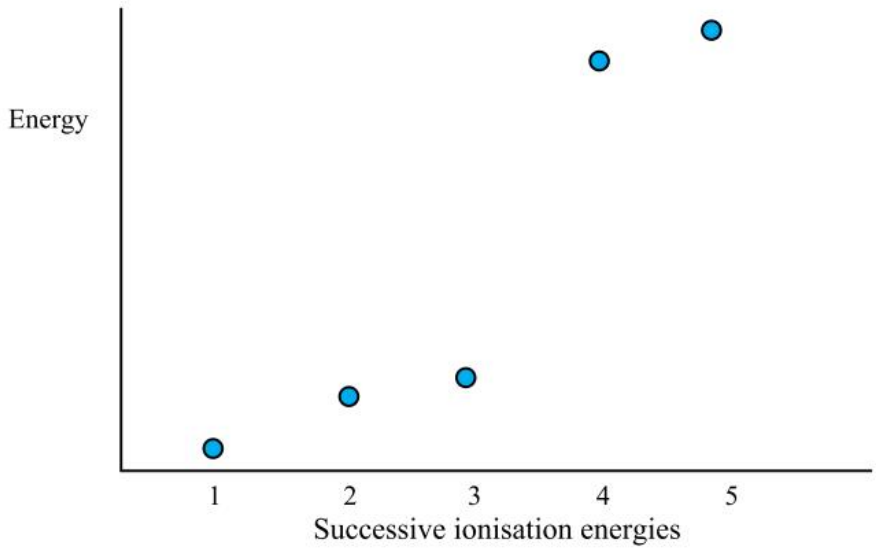
2. Why would you expect an ionised atom to be in the gaseous state?

3. "The first ionisation energy of atoms shows periodicity" Explain what is meant by this statement.

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?

Notes

Set 11: Periodic trends

Notes

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.

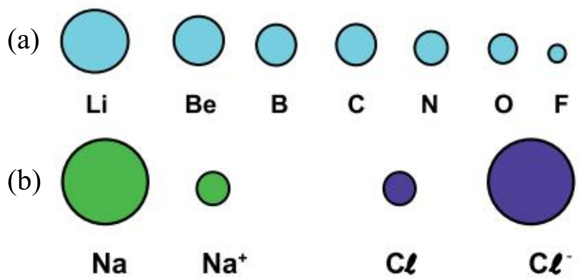
Periodic trends and patterns:
When going from left to right across a Period in the Periodic Table:

- 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 11: 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.

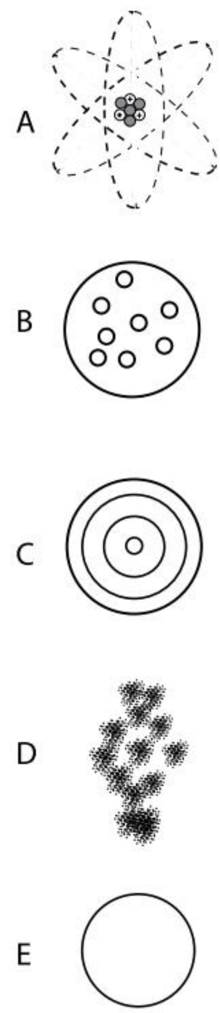
4. A new element is discovered and found to have 2 valence electrons. Write a paragraph describing the physical and chemical properties of this new element. In which Group of the Periodic Table would you place it?

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 12: Properties and structures of atoms

Set 12: Exercises

1. The Atom Timeline: A student searches for images of the structure of the atom and finds many different images, including those on the right. Answer the following questions about the structure of the atom.
   (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 on page 126 is more accurate than structure A?
   (d) Sequence the structures to make a timeline of our understanding of atomic structure. Include the structure on page 126.
   (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. Describe how the following isotopes are used in medical research:
   (a) Technetium-99 (b) Molybdenum-99 (c) Cobalt-60.

4. 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

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

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

Set 13: Compounds and formulae

Notes

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
nona-	9
deca-	10

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*

Notes

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 13: Compounds and formulae

Notes

Set 13: Exercises

1. Name the following molecular substances
   - (a) CO
   - (b) SO_2
   - (c) $PC\ell_5$
   - (d) N_2S
   - (e) P_2Br_4
   - (f) SF_6

2. Write the formula of each of the following molecular compounds.
   - (a) nitrogen monoxide
   - (b) nitrogen dioxide
   - (c) dinitrogen tetroxide
   - (d) sulfur trioxide
   - (e) water
   - (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 14: 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.

Notes

Set 14: 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)
(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. Explain, in terms of the metallic bonding model, each of the following properties of iron. Use labelled diagrams where appropriate:
   (a) high melting point
   (b) electrical conductivity
   (c) thermal conductivity
   (d) malleability

3. Explain, in terms of covalent bonding, each of the following properties of silicon dioxide. Use labelled diagrams where appropriate:
   (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 their
   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
   i. Melting point ii. Electrical conductivity iii. Hardness

5. Explain, in terms of ionic bonding, each of the following properties of sodium chloride. Use labelled diagrams where appropriate:
   (a) high boiling point
   (b) electrical conductivity when solid
   (c) electrical conductivity when liquid/molten
   (d) brittleness
   (e) hardness

Set 15: Uses, properties and structure

Notes

By making observations about materials we may develop an awareness of the different properties of materials. The properties observed include colour, state, hardness, elasticity, conductivity, melting point and boiling point. Many of these properties can be explained by knowledge of the type of bonding that occurs in the material. Therefore, in order to understand the many materials that surround us it is essential to have an understanding of the structure of ionic, metallic, covalent network and covalent molecular substances. This will then enable us to develop an understanding of the relationships between the properties and structures of these substances.

The following sets of questions are designed to get you thinking about the relationship between structure and properties of materials.

Set 15: Exercises

1. Explain the following:
   (a) electricians' screwdrivers have plastic handles
   (b) saucepans have plastic or wooden handles
   (c) plastic microwave dishes cannot be used in a conventional oven
   (d) tools used to etch glass have diamond tips
   (e) lead is used to make sinkers for fishing

2. Lead is used in the making of leadlight windows. What properties of the metal make it suitable for this purpose?

3. Describe the property of gold that makes it suitable for the following uses:
   a) jewellery
   b) coinage
   c) electronic circuits
   d) shields in spacecraft
   e) dental work

4. Classify the following as either metallic, ionic, covalent molecular or covalent network substances.
   (a) water, H_2O
   (b) copper(II) oxide, CuO
   (c) methane, CH_4
   (d) titanium dioxide, TiO_2
   (e) iron, Fe
   (f) silicon dioxide, SiO_2

5. Which of the substances listed in question 4 would conduct an electric current as a
   (a) solid
   (b) liquid
   Explain your choices.

Set 16: Properties and structures of materials

Set 16: Exercises

Notes

1. Research the composition of solder and compare the individual properties of the two 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?
8. Write an article suitable to be published in a weekend newspaper. Your article will discuss the relationship between the properties and structure of the different materials used to make outdoor furniture. In your article consider as many of the following as you feel necessary:
   - types of properties required of a material used to make outdoor furniture;
   - comparison of the physical and chemical properties of common materials such as wood, metal and plastic and discuss the suitability of each to make outdoor furniture;
   - innovations in chemical knowledge that has influenced the availability of wood, metal and plastic;
   - what you think outdoor furniture will be made of in the future;
   - cost;
   - availability.

Set 17: Electron dot diagrams

Notes

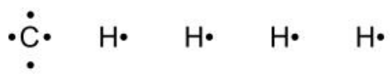
Electron dot diagrams are a useful tool to help you understand the bonding arrangements of atoms. From these diagrams it is possible to determine the shape of molecules and also their polarity. They are quite easy to draw for simple molecules as long as you follow some simple guidelines. Try to satisfy a full outer level of electrons (the octet rule) in all atoms except hydrogen (2 electrons) by:

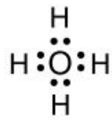
- Determining the total number of valence electrons from all the atoms involved;
- Using a pair of electrons to form a single bond between each pair of bonded atoms; If there are not enough electrons, consider a multiple covalent bond.
- Placing the remaining electrons to satisfy the octet rule. Be sure the electrons are drawn in pairs.

Example 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



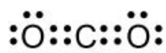


A single bond between each carbon and hydrogen will satisfy all the elements.

Example 2: Carbon dioxide CO_2

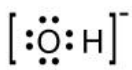
Each oxygen has six outer shell electrons and carbon has four outer shell electrons, giving a total of 16 valence electrons.

- 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.



Example 3: Hydroxide ion OH^-

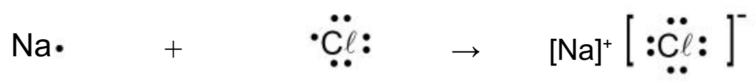
Oxygen has six outer shell electrons and hydrogen has one outer shell electron. As a single negative ion, an extra electron has been gained, giving a total of 8 valence electrons.



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.

Example 4: Ionic substances

Square brackets are used when representing ions in electron dot diagrams of ions and ionic substances.
E.g. in $NaC\ell$, sodium has one valence electron and when forming an ionic bond becomes a positive ion by losing its valence electron.
Chlorine has seven valence electrons and accepts an electron when forming ionic bonds becoming a negative ion.



sodium atom chloride atom sodium chloride

Set 17: Exercises

1. Draw electron dot diagrams for the following atoms and ions:
   (a) Na
   (b) Br
   (c) P
   (d) S^{2-}
   (e) H^+
   (f) N^{3-}

2. Draw electron dot diagrams for the following covalent molecular compounds (The central atom is in **bold** type) :
   (a) $H_2\mathbf{O}$
   (b) $\mathbf{C}H_3C\ell$
   (c) $\mathbf{P}H_3$
   (d) $H_2\mathbf{S}$
   (e) $\mathbf{C}HC\ell_3$
   (f) $H\mathbf{C}N$

3. Draw electron dot diagrams for the following ionic compounds:
   (a) NaOH
   (b) $CaC\ell_2$
   (c) Fe_2O_3
   (d) K_2S
   (e) $Mg(OH)_2$
   (f) $AgNO_3$

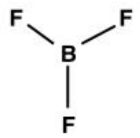
Notes

Set 18: Molecular shape

Notes

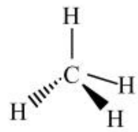
Polarity of a molecule is determined by the electron distribution in the bonds of the molecule and by the shape of a molecule. Shape is determined by the Valence Shell Electron Pair Repulsion Theory (VSEPR Theory). This states that electrons repel each other and will position themselves as far from each other as possible. The shapes you should be familiar with are shown in the diagram below.

Trigonal Planar



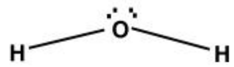
Bond angles for BF_3 are 120°

Tetrahedral



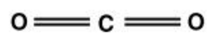
Bond angles for CH_4 are 109.5°

V-shaped



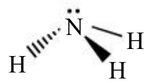
Bond angles for H_2O are 104.5°

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.

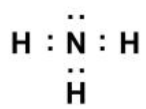
Bonds are polar if the elements bonded have differing electronegativities (electron attracting power).

In the molecule of ammonia (NH_3) in the diagram above for example, 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:



Notes

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 but the shape is symmetrical so the resulting methane molecule is **non-polar.**

Set 18: Exercises

1. Explain why the electronegativity of iodine is much less than for fluorine.
2. $C\ell$-----------------------$C\ell$
   $C\ell$-----------------------H
   $C\ell$-----------------------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
   (b) NH_3
   (c) H_2O
   (d) SO_3
   (e) NO_3^-
   (f) SO_4^{2-}
   (g) CO_3^{2-}
   (h) H_2O_2
   (i) C_2H_2
   (j) $PC\ell_3$

Set 19: Intermolecular forces

Notes

Observable properties of covalent molecular substances can be explained by the type and strength of intermolecular forces present between the molecules.

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 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 dispersion forces. The surface area of the molecules will also impact on the strength of the dispersion forces. The greater the surface area, the greater the 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 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.

Set 19: 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 19.1)
   (f) iodine (I_2)
   (g) hydrogen chloride ($HC\ell$)
   (h) methane (CH_4)

2. List 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, $CH_3C\ell$

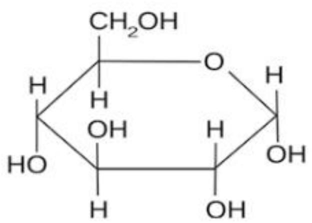


Figure 19.1: Glucose molecule

3. Given the following substances and their boiling points, list 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)
   (b) ethanol (C_2H_5OH)
   (c) ammonia (NH_3)
   (d) hexane (C_6H_{14})
   (e) iodine (I_2)
   (f) glucose ($C_6H_{12}O_6$) (see Figure 19.1)

6. Predict the solubility of each of the following substances in hexane (a non-polar liquid) and explain your decisions.
   (a) methane (CH_4)
   (b) ethanol (C_2H_5OH)
   (c) ammonia (NH_3)
   (d) hexane (C_6H_{14})
   (e) iodine (I_2)
   (f) glucose ($C_6H_{12}O_6$) (see Figure 19.1)

Notes

Set 20: Chromatography

Notes

Set 20: Exercises

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.

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?

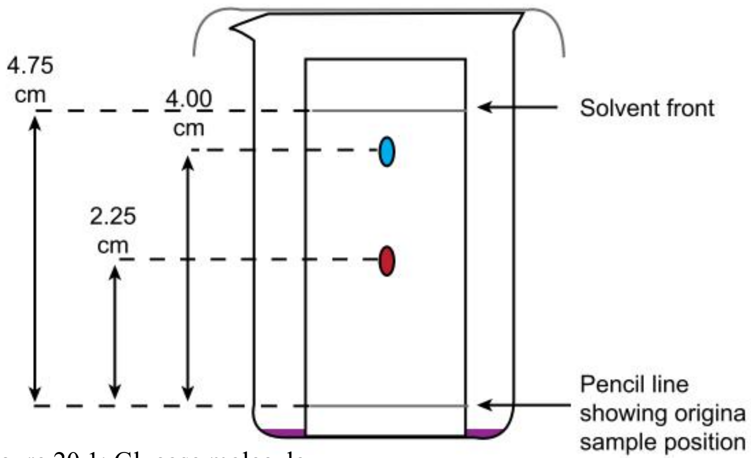


Figure 20.1: Glucose molecule

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 20.1.
   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

Gas chromatography (GC)

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 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?

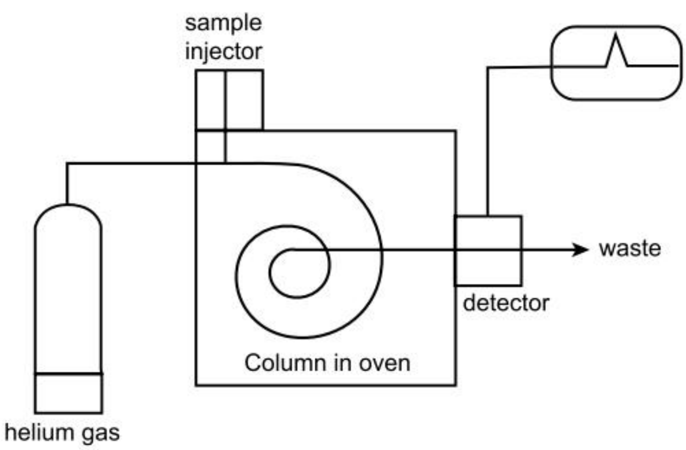
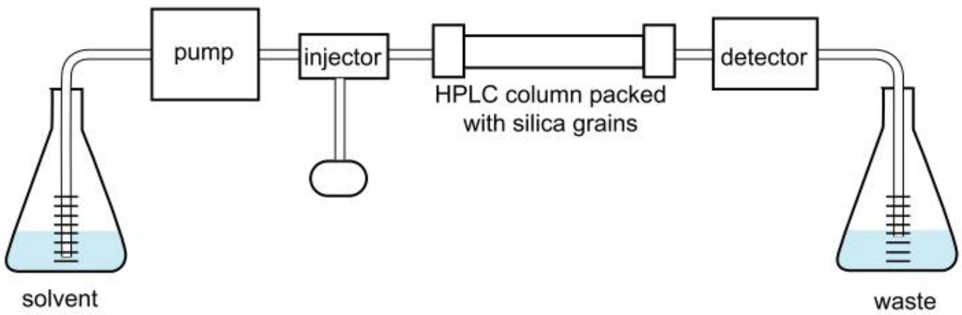


Figure 20.2: Gas chromatography

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.



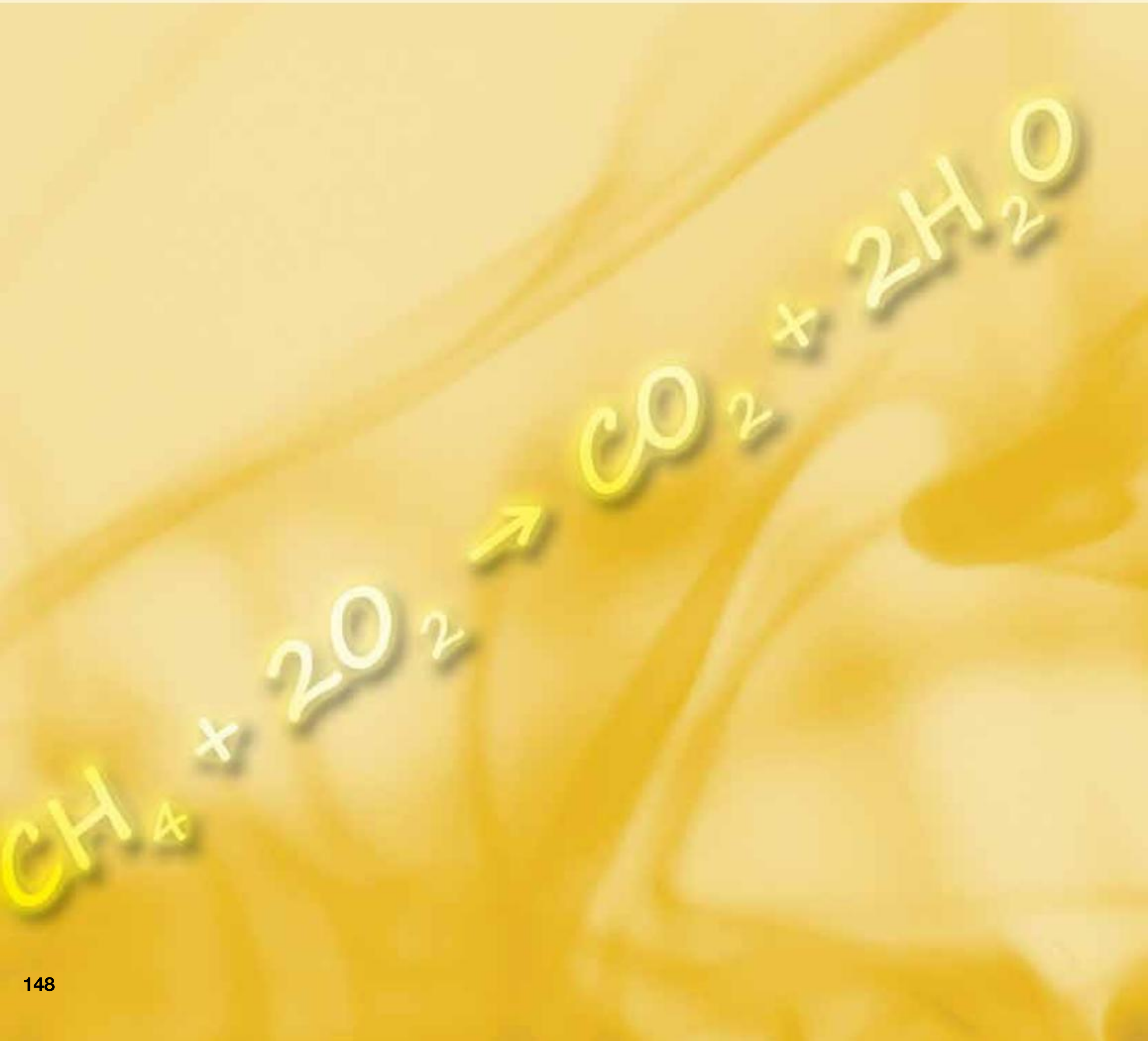
4. As with GC the sample components are identified by their retention time. The retention time depends on the pump pressure, the size and chemical properties of the stationary phase, the solvent and its temperature.
   a) Explain clearly the advantages of using a high pressure pump to force the solvent 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. Give 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. Give 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.

Notes

Chemical reactions and stoichiometry

The problem sets in the chemical reactions and stoichiometry section of *Exploring Chemistry Year 11: Experiments, Investigations and Problems*, provides opportunities for students to explore:

- **different types of chemical reactions, including precipitation reactions and how they can be represented by chemical equations**
- **how the presence of specific ions in solutions can be identified by observing the colour of the solution**
- **chemical reactions quantitatively using the mole concept as it relates mass, moles and molar mass and, with the Law of Conservation of Mass**



Set 21: 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 $1.6605402 \times 10^{-27}$ kg. Using this value, an atom of H-1 will have a mass of 1.0078 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.

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 magnetic field. The charged particles are deflected by the magnetic or electric 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 $C\ell$-35 and $C\ell$-37.

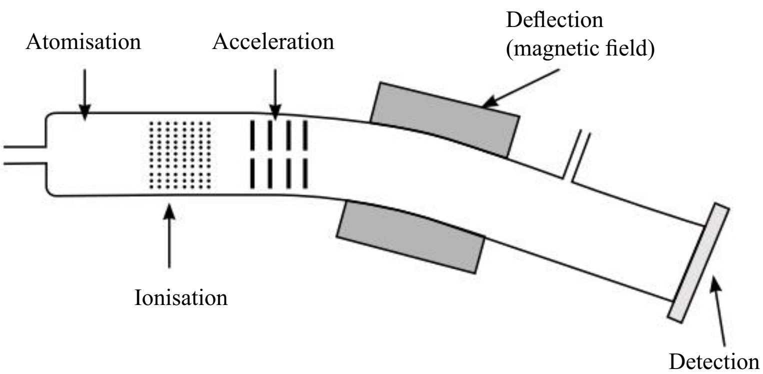
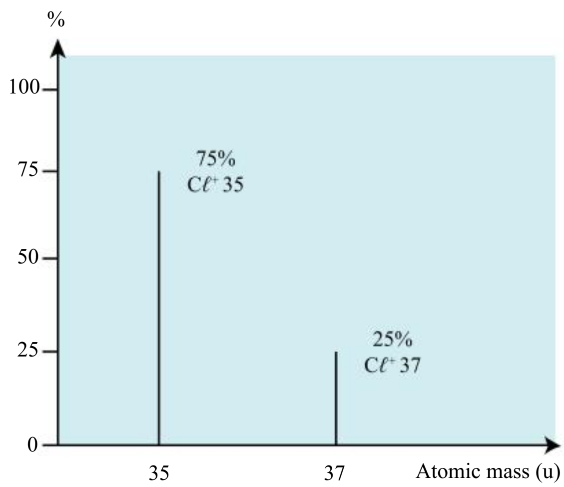


Figure 21.1: Mass spectrometer



Set 21: Relative atomic mass and mass spectroscopy

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.

Examples

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.

$x + y = 100$ equation 1

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

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

The value for y can then be substituted into equation 2

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

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

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

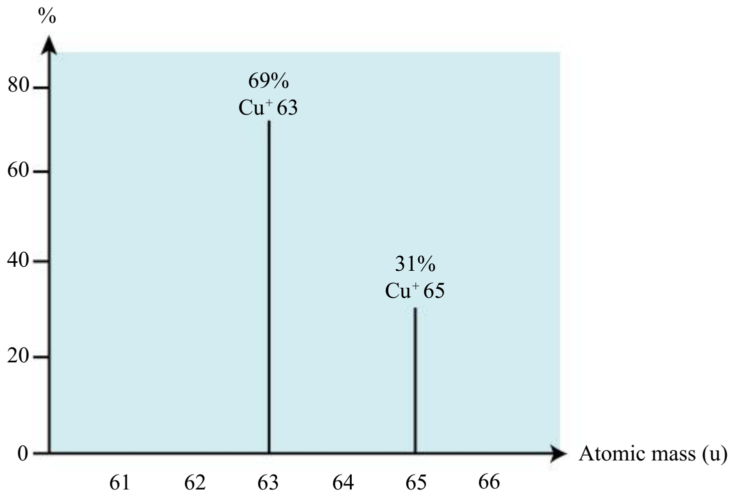
Substitute the value of x into equation 3

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

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.

Notes



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 21: 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.

Set 21: Relative atomic mass and mass spectroscopy

Notes

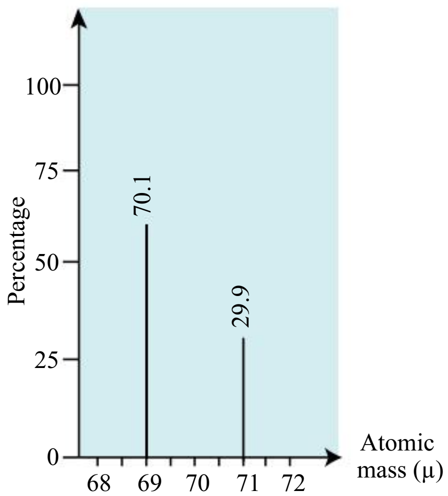
7. Two samples of bromine, collected from different continents were found to have relative atomic masses of 79.9 and 80.9 respectively. What proportion of the stable isotopes of bromine, Br-79 and Br-81, are present in each sample?

8. Refer to the Figure 21.1 on page 149 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

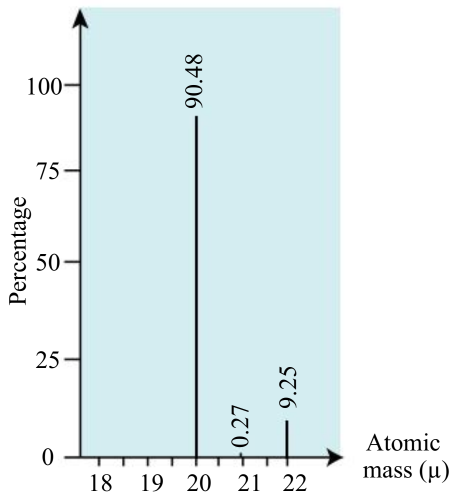
9. Gallium has two isotopes. The following mass spectra was produced. Calculate the relative atomic mass of Gallium. (Ga – 69: 70.1 %; Ga – 71: 29.9%)

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%

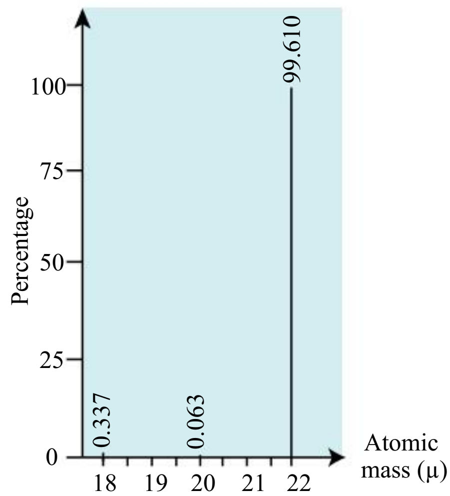
11. The following mass spectra was produced from a sample of argon. Calculate the relative atomic mass of argon.
    36 – 0.337% 38 – 0.063 % 40 – 99.610 %



Spectrum for question 9: Gallium



Spectrum for question 10: Neon



Spectrum for question 11: Argon

Set 22: Molar mass

Notes

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.
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.

Examples

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

1. Magnesium (Mg): M(Mg) = 24.31 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 ($CuSO_4{\cdot}5H_2O$)
$$\begin{aligned} M(CuSO_4{\cdot}5H_2O) &= M(Cu) + M(S) + 4 \times M(O) + 10 \times M(H) + 5 \times M(O) \\ &= (63.55) + (32.01) + (4 \times 16.00) + (10 \times 1.008) + (5 \times 16.00) \\ &= 249.64 \\ &= 2.50 \times 10^2 \text{ g mol}^{-1} \end{aligned}$$

Set 22: Molar mass

Notes

Isotopes

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

Set 22: Exercises

1. Calculate the molar masses of
   (a) potassium hydroxide, KOH
   (b) copper(II) chloride, $CuC\ell_2$
   (c) aluminium chloride, $A\ell C\ell_3$
   (d) calcium hydroxide, $Ca(OH)_2$
   (e) ammonium oxalate, $(NH_4)_2C_2O_4$
   (f) sodium carbonate-l0-water, $Na_2CO_3{\cdot}10H_2O$
   (g) tungsten carbide, WC
   (h) methane, CH_4
   (i) sucrose, $C_{12}H_{22}O_{11}$
   (j) silver chloride
   (k) zinc iodide
   (l) sodium sulfate
   (m) iron(III) sulfate
   (n) sulfur dioxide
   (o) sulfuric acid

2. Carbon dating is possible due to the composition of the C-14 isotope in living organisms.
   Find out:
   (a) how a C-14 atom is similar to and different from a C-12 atom;
   (b) any other isotopes of carbon;
   (c) which isotope of carbon is the most abundant and how you can verify this;
   (d) how carbon dating works.

3. Hydrogen has three isotopes; hydrogen, deuterium and tritium.
   (a) Describe the composition of each of these atoms.
   (b) Where are deuterium and tritium used?

Set 23: 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 6.022×10^{23} particles.

The number of atoms in 12.00 g of carbon-12 is 6.022×10^{23}. This very, very large number, 6.022×10^{23} is called Avogadro's number. Avogadro's number of particles is called a mole. Therefore one mole of any substance will contain 6.022×10^{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 6.022×10^{23} atoms of iron.
- A mole of carbon dioxide molecules (CO_2) consists of 6.022×10^{23} molecules of carbon dioxide.
- A mole of copper(II) sulfate ($CuSO_4$) consists of 6.022×10^{23} formula units of copper(II) sulfate.

Notes

Moles & 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.

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

$$m = nM$$

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

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

Set 23: Moles, particles and mass

Notes

Set 23: 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 24: Interpretation of formulae

Notes

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) $= 8 \times n(Ca_3(PO_4)_2)$
   $= (8)\ (5.00)$
   n (oxygen atoms) $= 40.0$ 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 $n = \frac{m}{M}$ $\quad M(CO_2) = 44.01\ g\ mol^{-1}$

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

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

   $$n(O) = 2 \times n(CO_2) = (2)\ (3.00)$$
   $$n(O) = 6.00\ mol$$

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

   $$n(O) = \frac{m}{M}$$
   $$m(O) = nM = (6.00)\ (16.0)$$
   $$m(O) = 96.0\ g$$

Set 24: Interpretation of formulae

Notes

Set 24: 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 25: Percentage composition

Notes

We are very interested today in knowing the exact composition of many of the products which surround us. When you buy foods, there are consumer panels 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 about goods. Food and medical goods are analysed by chemists to determine the exact percentage compositions of the ingredients. Individual compounds can also be analysed to determine their exact composition.
The percentage composition of a chemical compound specifies the percentage by mass of each of the different elements present in the compound.

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

Examples

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

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

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

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

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

$\overline{100\ \%}$

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

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

Note: The percentages 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.

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

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

$= 33.3\%$

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

Set 25: Percentage composition

Notes

Set 25: Exercises

1. Calculate the percentage by mass of each element in:
   (a) sodium hydroxide
   (b) acetic acid (CH_3COOH)
   (c) copper(II) sulfate-5-water
   (d) potassium phosphate
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?

Set 26: 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.

Notes

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, 25 °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.

$$n = \frac{V}{22.71}$$

$$\begin{aligned} V &= n \times 22.71 \\ &= (2.75)\,(22.71) \\ V &= 62.5 \text{ L} \end{aligned}$$

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

$$n = \frac{V}{22.71}$$

$$\begin{aligned} V &= n \times 22.71 \\ &= (0.300)\,(22.71) \\ V &= 6.81 \text{ L at S.T.P.} \end{aligned}$$

Set 26: Gas volumes

Notes

Set 26: 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 27: Ionic equations

Notes

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:

   $CaCO_3(s) + 2\ H^{+}(aq) \rightarrow Ca^{2+}(aq) + H_2O(\ell) + CO_2(g)$

Set 27: Ionic equations

Notes

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 27: 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 28: Equations and observations

Notes

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.

Example

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 28: Exercises

For the following reactants, write balanced equations and describe the expected observations.

1. Marble chips (calcium carbonate) are added to dilute nitric acid.
2. Magnesium metal is added to dilute hydrochloric acid.
3. Baking soda (sodium hydrogencarbonate) is added to vinegar (acetic acid).
4. Sodium hydroxide solution is added to sulfuric acid solution.
5. Solid sodium hydroxide is added to dilute hydrochloric acid.
6. Dilute sulfuric acid is added to solid cobalt carbonate.
7. Barium chloride solution is added to dilute sulfuric acid.
8. Lead(II) nitrate solution is added to sodium iodide solution.
9. Gold(III) chloride solution is added to solid copper.
10. Sodium metal is added to water.
11. Potassium carbonate solid is added to an excess of dilute nitric acid.
12. Potassium hydroxide solution is added to nitric acid solution.
13. Silver nitrate solution is added to sodium bromide solution.
14. Nickel metal is added to copper nitrate solution.
15. Iron(II) sulfate solution is added to potassium hydroxide solution.

Set 29: Stoichiometry

Notes

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):

   $n(H_2O) = 0.480$

   $m(H_2O) = 0.480\ (18.0)$ $\qquad M(H_2O) = 18.0\ g\ mol^{-1}$

   $= 8.60$ 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

Notes

(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:

$n(HC\ell) = 2 \times n(C\ell_2)$

$= 2 \times \frac{17.8}{70.9}$ $\quad M(C\ell_2) = 70.9\ g\ mol^{-1}$

$n(HC\ell) = 0.502\ mol$

(b) Convert the moles to mass

$n(HC\ell) = 0.502$
$m(HC\ell) = (0.502)(36.458)$ $\quad M(HC\ell) = 36.458\ g\ mol^{-1}$
$= 18.3\ g$

Set 29: 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 moles 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:

$$MgO(s) + 2\ HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2O(\ell)$$

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

Notes

3. When silver nitrate solution is added to a solution of calcium chloride, a white precipitate of silver chloride is produced

$CaC\ell_2(aq) + 2\,AgNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2\,AgC\ell(s)$ or $Ag^+(aq) + C\ell^-(aq) \rightarrow AgC\ell(s)$

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

9. A load of quarried limestone contains 92.0% $CaCO_3$, the remainder being silica (SiO_2). The limestone is to be heated to form quicklime (CaO).

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

(a) What mass of quicklime could be produced by heating 5.00×10^2 kg of the limestone?
(b) What mass of carbon dioxide would be produced?

Set 30: Stoichiometry and gas volumes

Example

Notes

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 30: 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 $2\,Na(s) + 2\,H_2O(\ell) \rightarrow 2\,Na^+(aq) + 2\,OH^-(aq) + H_2(g)$

Set 30: Stoichiometry and gas volumes

Notes

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, and
   (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.

Notes

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.

Solutions and acidity

The problem sets in the solutions and acidity section of *Exploring Chemistry Year 11: Experiments, Investigations and Problems*, provides opportunities for students to explore:

- **the presence of specific ions in solutions by observing the colour of the solution and observing precipitation and acid-base reactions**
- **the solubility of substances in water, including ionic and polar and non-polar molecular substances**
- **the Arrhenius model to explain the behaviour of strong and weak acids and bases in aqueous solutions**
- **indicator colour and the pH scale to classify aqueous solutions as acidic, basic or neutral**
- **patterns of 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**

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Set 31: Solutions

Set 31: Exercises

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 31.1 to classify the salinity of the sample and state if it is suitable for irrigation.

Figure 31.1

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. Refer to Set 27: Ionic equations. Read the explanations and examples. 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.

Notes

Set 32: Solution concentrations

Notes

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 '[ ]' that is, c(NaCℓ) or [NaCℓ] both refer to the concentration of NaCℓ in solution, expressed in mol L^{-1}.

Examples

1. Find the number of moles of sodium chloride in 15.0 mL of a 2.50 mol L^{-1} NaCℓ solution.

$$\begin{aligned} n &= cV \\ &= (2.50)\,(15.0 \times 10^{-3}) \\ n &= 3.75 \times 10^{-2} \text{ mol of NaC}\ell \end{aligned}$$

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:

$$\begin{aligned} n &= \frac{m}{M} \\ &= \frac{5.56}{278.0} \qquad M(FeSO_4{\cdot}7H_2O) = 278.0 \text{ g mol}^{-1} \\ &= 2.00 \times 10^{-2} \text{ mol} \end{aligned}$$

(b) Calculate the concentration of $FeSO_4{\cdot}7H_2O$ solution.

$$\begin{aligned} c &= \frac{n}{V} \\ &= \frac{2.00 \times 10^{-2}}{750 \times 10^{-3}} \\ c &= 2.67 \times 10^{-2} \text{ moI L}^{-1} \end{aligned}$$

Concentration of ions

Ionic substances and some covalent substances, such as strong acids, produce almost 100% ions when dissolved in H_2O. For example

$$NaC\ell(s) \rightarrow Na^+(aq) + C\ell^-(aq)$$

$$HC\ell(g) \rightarrow H^+(aq) + C\ell^-(aq) \qquad \text{or} \qquad HC\ell(g) + H_2O(\ell) \rightarrow H_3O^+(aq) + C\ell^-(aq)$$

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 32: 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

Notes

Set 32: Solution concentrations

Notes

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?

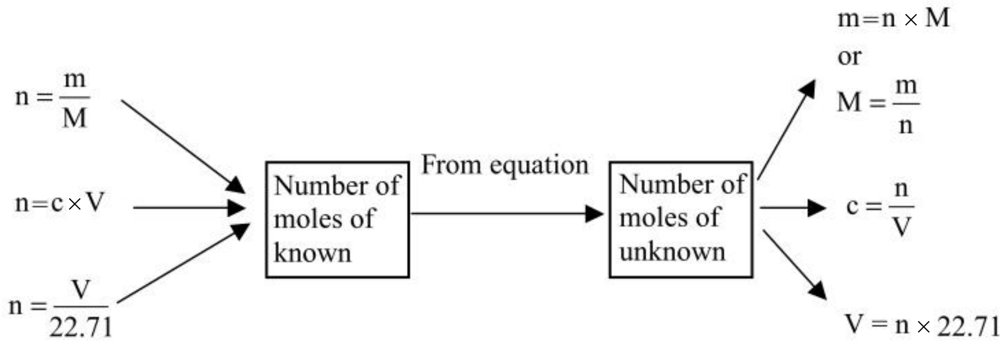
9. What volume of water must be added to 150.0 mL of 1.10 mol L^{-1} sulfuric acid solution to prepare a 0.210 mol L^{-1} solution?

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 33: Reacting masses and gaseous and solution volumes

Notes

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 33: 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

Set 33: Reacting masses and gaseous and solution volumes

Notes

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

10. When white phosphorus is heated in air it burns to form tetraphosphorus decaoxide (P_4O_{10}):

    $$P_4(s) + 5\ O_2(g) \rightarrow P_4O_{10}(s)$$

    If 6.20 g of phosphorus is treated in this way:
    (a) what mass of P_4O_{10} is formed?
    (b) what volume of air (measured at S.T.P.) is required? Assume that air is 20.0% oxygen by volume.

11. A 15.0 g piece of zinc is added to a beaker containing dilute hydrochloric acid and effervescence occurs:

    $$Zn(s) + 2\ HC\ell(aq) \rightarrow ZnC\ell_2(aq) + H_2(g)$$
    or
    $$Zn(s) + 2\ H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$$

    After all reaction has ceased, the piece of zinc is washed, dried and found to have a mass of 2.00 g. Deduce from these results
    (a) the number of moles of hydrochloric acid in the beaker
    (b) the mass of zinc chloride which can be recovered from the solution in the beaker
    (c) the volume of hydrogen gas produced by the reaction at S.T.P.

Notes

12. It has been estimated that 1.00×10^8 tonne of H_2S is released annually into the atmosphere from biogenic sources, especially from sulfate respiration by bacteria. This H_2S is oxidised to SO_2 in a day or so according to the overall process:

$$2\ H_2S(g) + 3\ O_2(g) \rightarrow 2\ SO_2(g) + 2\ H_2O(g)$$

If 8.00 g of oxygen is consumed by a sample of hydrogen sulfide;
(a) how many mole of hydrogen sulfide molecules would react?
(b) what volume at S.T.P. of hydrogen sulfide would react?

13. A technique, known as the Nanaimo system for reducing the NO content of car exhausts involves the injection of a stream of NH_3 into the exhaust vapour. This converts the NO to harmless N_2 and H_2O via the reaction

$$4\ NH_3(g) + 6\ NO(g) \rightarrow 5\ N_2(g) + 6\ H_2O(g)$$

Cars typically emit about 3.00 g of NO per km travelled. How many litres of NH_3, calculated at S.T.P., would be needed to clean up the NO emitted by a car in one year, assuming that it is driven a distance of 1.60×10^4 km?

14. Liquid hydrogen is being used to power an experimental jet engine. In a particular test, in which all of the hydrogen fuel was oxidised, the fuel tank was found to have decreased in mass by 2.00 kg. What volume of oxygen, measured at S.T.P. would be needed in this test run?

15. 1.60 g of hydrogen gas is added to a container with 10.0 L of oxygen gas at S.T.P. The mixture is sparked and a reaction occurs. Find:
(a) the mass of steam formed;
(b) the volume of unused oxygen at S.T.P.
(c) the volume of steam produced at S.T.P.

16. What volume of oxygen is required for the complete combustion of
(a) 1.00 L of hydrogen at S.T.P.?
(b) 2.00×10^2 mL of methane (CH_4) at S.T.P.?

17. When 25.0 mL of a solution of dilute sulfuric acid is mixed with excess barium chloride solution, 0.483 g of precipitate is formed. What volume of 0.0134 mol L^{-1} sodium hydroxide solution is required to neutralise 125 mL of the original acid?

18. A 3.00 g sample of limestone is analysed for purity by reacting it with hydrochloric acid. It is found that 20.0 mL of 2.50 mol L^{-1} $HC\ell$ solution is required for complete reaction. Calculate the percentage of calcium carbonate in the limestone sample.

19. A 0.482 g sample of an alloy of silver and copper was dissolved in nitric acid. The resulting solution required 41.5 mL of 0.0993 mol L^{-1} sodium chloride solution for complete precipitation of the silver ions. Calculate the percentage of silver in the alloy?

Set 34: The pH scale

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:

$$pH = -\log_{10} [H^+] \text{ where pH is a number without units.}$$

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						

Notes

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. Concentration of hydrogen ions is written like this $[H^+]$ using square brackets.

   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:

   $$pH = -\log_{10} [H^+] = \log_{10} (0.0200) = 1.70$$

2. Orange juice has a pH of 4.50. What is the $[H^+]$ in orange juice?

   Calculate the hydrogen ion concentration:

   $$pH = -\log_{10} [H^+]$$
   $$-4.50 = \log_{10} [H^+]$$
   $$[H^+] = 10^{-4.50} = 3.16 \times 10^{-5} \text{ mol L}^{-1}$$

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?

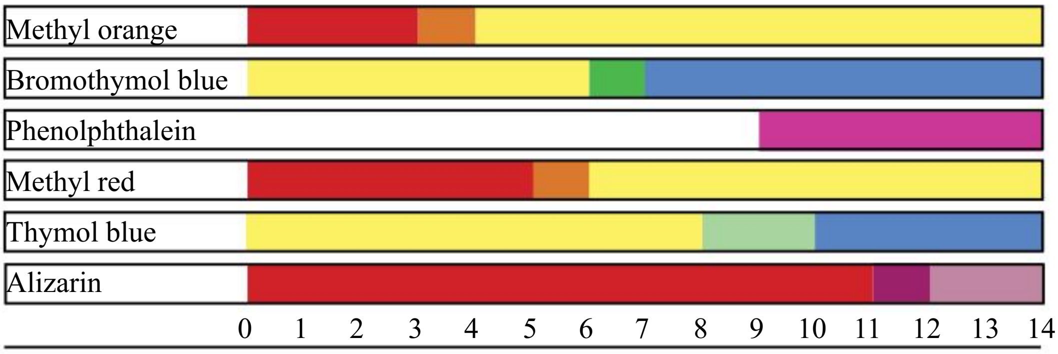
As this is a dilution problem we can use the dilution equation:
$c_1V_1 = c_2V_2$

$0.250 \times 0.210 = 2.00 \times V_2$
$\therefore V_2 = 26.3$ mL
Joan should use 26.3 mL of 2.00 mol L^{-1} acid and make it up to 250 mL with distilled water.

Notes

Set 34: 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. What is 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?

3. Calculate the $[H^+]$ and pH of the following solutions.
   (a) A 0.100 mol L^{-1} $HC\ell$ solution.
   (b) A 0.00500 mol L^{-1} HNO_3 solution.
   (c) A 2.00 mol L^{-1} $HC\ell$ solution

Set 34: The pH scale

Notes

4. Calculate the $[H^+]$ of solutions with the following pH:
   (a) Lemon juice of pH 3.00
   (b) Dish washing solution of pH 11.0 used in a dishwasher.
   (c) Pool acid of pH -1.00 (yes, negative one – pH exists outside 0-14)
   (d) Orange juice of pH 4.56
   (e) Swimming pool water of pH 7.60

5. Before disposal, a hydrochloric acid solution of pH 4.00, used to remove rust from iron parts prior to galvanising, needs to be neutralised. Base is added until the pH is 7.00. By what factor has the hydrogen ion concentration changed?

6. A brick cleaner needs to replenish the hydrochloric acid solution he is using. Using a pH meter he measures the pH of his depleted solution to be 2.00. He decides to add 3.00 L of 3.00 mol L^{-1} $HC\ell$ solution to 2.00 L of his solution. Calculate the final hydrogen ion concentration of the new solution.

7. A laboratory technician wants a low concentration hydrochloric acid solution with a pH of 5.00, for calibrating a conductivity meter. What volume of distilled water must be added to a 25.0 mL sample of hydrochloric acid solution of pH 3.60 to produce the required solution?

8. A chef wishes to clean a coffee machine using citric acid ($C_6H_8O_7$). He requires 100 mL of a citric acid solution of 0.320 mol L^{-1}. Unfortunately his assistant has made up 100 mL of 0.100 mol L^{-1} solution. What mass of citric acid must he add to increase the concentration of this solution to 0.320 mol L^{-1}? Assume the volume of the solution does not change with the addition of the solid.

9. A student prepared a 200 mL mixture made from 100 mL of $HC\ell$ with pH=2 and 100 mL of $HC\ell$ with pH=4. What is the new pH?

10. A water tank contains 15 000 L of bore water of pH 5.50. In an effort to increase the pH 10.0 g of sodium hydroxide was added to the water in the tank. Determine the pH of the resulting solution. Assume the volume does not change on adding the solid.

11. A swimming pool contains 2.00 ML of water at a pH of 7.80 making swimming in it unpleasant. The caretaker needs to reduce the pH to 6.80. What volume of 12.0 mol L^{-1} hydrochloric acid does he need to add? Assume the volume of acid is insignificant compared to the volume of the pool.

Set 35: Solutions 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}), expressed as:

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

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

$$n = cV = (2.50)(0.0150) = 3.75\times10^{-2} \text{ mol of NaOH}$$

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) $c(OH^-) = 2 \times c(Ba(OH)_2)$
$= 2 \times 0.200$
$= 0.400 \text{ mol L}^{-1}$

Notes

Set 35: Solutions of acids and bases

Notes

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:

$$\begin{aligned} n(\text{dilute } HNO_3) &= c(HNO_3)\,(V) \\ &= (0.200)\,(250.0 \times 10^{-3}) \\ &= 5.00 \times 10^{-2} \text{ mol} \end{aligned}$$

(b) Relate the unknown number of moles of concentrated HNO_3 to the known number of moles of dilute HNO_3:

$$\begin{aligned} n\,(\text{concentrated } HNO_3) &= n\,(\text{dilute } HNO_3) \\ (15.6)(V) &= 5.00 \times 10^{-2} \\ V &= \frac{5.00 \times 10^{-2}}{15.6} \\ V &= 3.21 \times 10^{-3} \text{ L} \end{aligned}$$

Set 35: 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}.

Question 8: Sandstone in hydrochloric acid

8. A 0.100 mol L^{-1} solution of hydrochloric acid was required to analyse some sandstone. Calculate the mass of hydrogen chloride gas required to make 300 mL of this solution.

9. Slaked lime (calcium hydroxide) can be used to reduce soil acidity and is sparingly soluble in water. Calculate the concentration in mol L^{-1} of:
   (a) calcium ions;
   (b) hydroxide ions;
   in a solution containing 10.0 mg of calcium hydroxide in 1.00 L of solution.

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.
    (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

Question 10: Hydroponics (photo: Joel Malcolm)

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.

13. Bore water with a hydrogen ion concentration of 3.60×10^{-4} mol L^{-1} can be relatively acidic. This can make it unsuitable for irrigation. Mixing this water with rain water reduces the concentration of hydrogen ions and hence increases the pH.
    (a) Assuming the rain water is neutral, calculate the volume of rain water that should be added to 3.00×10^3 L of bore water to reduce the hydrogen ion concentration to 1.00×10^{-6} mol L^{-1}.
    (b) Is this a practical method of reducing the hydrogen ion concentration? Explain.
    (c) Suggest other methods that could be used to achieve the same result. How does this compare with the addition of rain water?

14. A dilute sulfuric acid solution can be used to dissolve magnesite samples prior to analysis. Magnesite contains mostly magnesium carbonate. Calculate the volume of water that must be added to 1.50×10^2 mL of a 5.50 mol L^{-1} sulfuric acid solution to prepare a 0.500 mol L^{-1} solution for use in the analysis.

Set 36: Acid and base reaction stoichiometry

Notes

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

$$n(Mg) = \frac{m}{M} = \frac{3.00}{24.3} = 0.123\ \text{mol}$$

(c) Calculate the number of moles of hydrogen gas produced

$$n(H_2) = n(Mg) = 0.123\ \text{mol}$$

(d) Calculate the volume of hydrogen gas produced (1 mole of a gas occupies 22.71 L at STP)

$$V(H_2) = n(H_2) \times 22.71 = (0.123) \times 22.71 = 2.79\ \text{L}$$

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.)

$$n(CO_2) = \frac{V(CO_2)}{22.71} = \frac{2}{22.71} = 0.0881\ \text{mol}$$

(c) Calculate the number of moles of calcium carbonate needed

$$n(CaCO_3) = n(CO_2) = 0.0881\ \text{mol}$$

(d) Calculate the mass of calcium carbonate needed

$$m(CaCO_3) = n(CaCO_3) \times M(CaCO_3) = 0.0881 \times 100.09 = 8.81\ \text{g}$$

Set 36: 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.

Notes

Set 37: Aqueous solutions of acids and bases

Set 37: Exercises

Question 2: Marble statue

Question 5: Mortar

Question 6/7: Rust and rust remover

Double arrows $\rightleftharpoons$ are used when writing equilibrium equations such as the formation of weak acids and bases. They indicate that both the forward and reverse reactions are occurring.

1. Write an equation to show how sulfur dioxide in the atmosphere can produce acid rain.

2. In Europe, acid rain destroys the surface of historic marble structures. Marble is almost pure calcium carbonate. Write an equation to illustrate the reaction responsible for this process.

3. Fish are very sensitive to aluminium ions in their aquatic environments. Acid rain causes insoluble aluminium compounds such as aluminium oxide in the soil to dissolve, releasing aluminium ions into ground water as it moves through the soil. This water can end up in streams and lakes where fish populations may die. Write an equation to represent the reaction of aluminium oxide with acid rain.

4. Even in places where no atmospheric sulfur or nitrogen oxides are present, rainwater is still acidic. Explain why this is the case and write an equation to show how the rainwater becomes acidic.

5. Solid concrete and mortar are complex mixtures of hydroxides and carbonates that bind sand and stones together. Typical hydroxides and carbonates are those of calcium and magnesium. Concentrated hydrochloric acid is used to remove mortar that has been accidentally spilt onto brick walls by bricklayers at the time of building. Write equations to show how the acid is able to dissolve these hydroxides and carbonates.

6. Iron objects tend to rust when exposed to moist air. Rust is essentially iron(III) oxide containing some water of crystallisation. Hydrochloric acid is often used to remove the rust. Write an equation to represent this process

7. Rust Buster ® is a colourless solution consisting of 30% phosphoric acid, H_3PO_4, in water. It is used to remove surface rust from iron.
   (a) Write three equations to show how phosphoric acid produces an acid solution.
   (b) Comment on the extent to which each part of the process proceeds towards products.
   (c) Write an equation to show how phosphoric acid removes rust, essentially iron(III) oxide, from the surface of iron.

8. When pure sulfuric acid is mixed with distilled water to make battery acid the process is probably best viewed as a reaction of the sulfuric acid with water to produce hydronium ions. Write equations for this process.

9. Acid soils can be a problem when used to grow certain crops. Suggest how the acidity of the soil can be reduced using naturally occurring materials. Write an equation to show how this can reduce the soil acidity.

Research

1. Acid rain can be produced from atmospheric pollutants and atmospheric carbon dioxide and also from naturally produced organic compounds released into the air by plants. Research the process where acid rain is produced from naturally produced organic compounds. Find out what type of acids are found in this type of acid rain and discuss whether this type of acid rain is likely to be a significant problem.
2. Discuss the role of rainwater and carbon dioxide in the formation of caves in limestone rocks such as in the southwest of Western Australia and the Nullarbor Plain.
3. Research the composition of the digestive juices from your stomach and discuss the use of antacid medications.
4. Research the pH of your skin. What role does this have in maintaining a healthy skin and ultimately what is its effect on your health?
5. Extremophiles are organisms that live in conditions that would not support most living organisms. Two such groups of organisms are acidophiles and alkaliphiles. Research these two types of extremophiles.

Extended answer question

Where applicable, use equations, diagrams and illustrative examples of the chemistry you are describing. Your answer should focus on the relevant chemical content specific to the situations described. Present your answer in a logical and coherent manner.

1. **Hydrochloric acid verses phosphoric acid**

   Concentrated 12.0 mol L^{-1} hydrochloric acid has a hydrogen ion concentration of 12.0 mol L^{-1} whereas 12.0 mol L^{-1} phosphoric acid has a hydrogen ion concentration of around 0.008 mol L^{-1}. This difference in hydrogen ion concentration accounts for some of the differences in their chemical properties and hence their uses.

   Some uses are similar, for example both acids are used to remove rust from the surface of iron. Hydrochloric acid is used to remove rust from iron in preparation for galvanising and phosphoric acid is used to remove rust from iron prior to painting. If the iron treated with hydrochloric acid is washed with water and left in air even for a short time it rapidly turns brown. If iron treated with phosphoric acid is left in air there is no immediate appearance of any brown material on its surface. About a day after it dries, a dark coloured coating becomes apparent.

   A second example of where hydrochloric acid is used extensively is to remove hardened mortar from bricks. Hardened mortar consists of sand stuck together mostly by crystals of calcium and magnesium hydroxides and some calcium and magnesium carbonates. The cleaning process involves application of concentrated hydrochloric acid, then washing it off with water.
   Phosphoric acid cannot be used for removing hardened mortar from bricks.

   (a) Explain fully why the hydrogen ion concentration of 12.0 mol L^{-1} hydrochloric acid is different from the hydrogen ion concentration of 12.0 mol L^{-1} phosphoric acid.

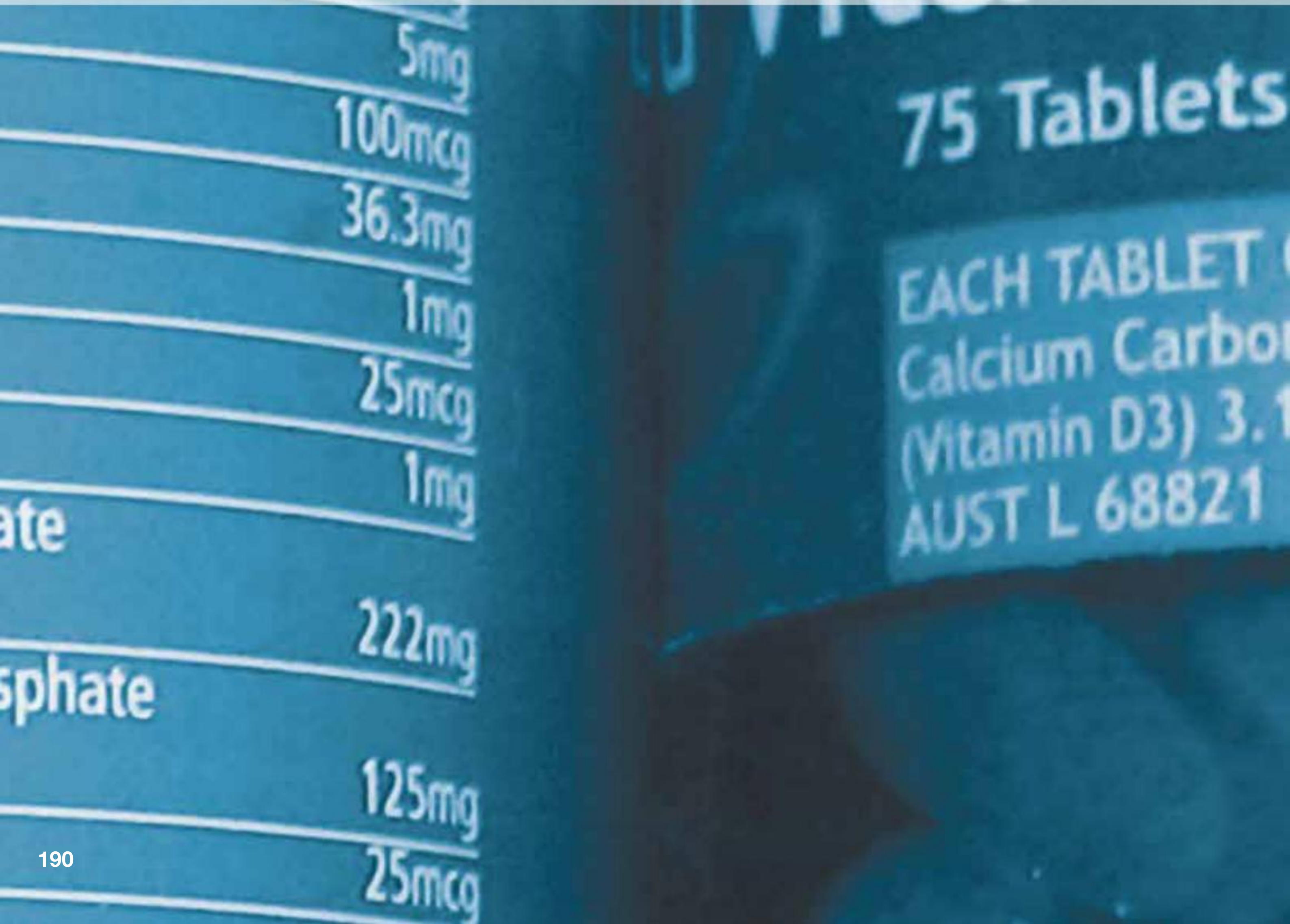
   (b) Fully discuss the chemistry involved in the two examples of the uses of these acids described above. Your discussion should include an explanation of the similarities and differences in the chemistry of each acid when used to remove rust, and an explanation of why hydrochloric acid can be used to remove hardened mortar while phosphoric acid cannot.

Notes

Energy changes and rates of reaction

The problem sets in the energy changes and rates of reaction section of *Exploring Chemistry Year 11: Experiments, Investigations and Problems*, provides opportunities for students to explore:

- **enthalpy changes in chemical reactions and phase changes through observable changes in the temperature of the surroundings and/or the emission of light**
- **endothermic and exothermic reactions**
- **how food and fuels (fossil fuels including coal, oil, petroleum and natural gas and biofuels including biogas, biodiesel and bioethanol) can be compared in terms of their energy output, suitability for purpose, and the nature of products of combustion**
- **the rate of chemical reactions by measuring the rate of formation of products or the depletion of reactants**
- **the effects of concentration, temperature, pressure, the presence of catalysts and surface area on the rate of chemical reactions**



Set 38: Energy changes

Set 38: Exercises

Notes

1. Identify the following reactions as endothermic or 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.

Set 39: 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 39.1 and 39.2 show energy profile diagrams for an exothermic and an endothermic reaction while figure 39.3 shows the distribution of molecular energies in a system.

Figure 39.1
Exothermic reaction

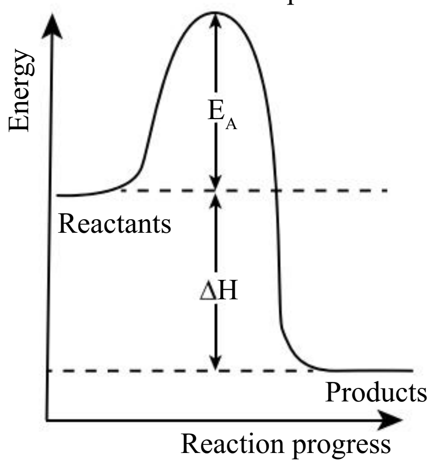


Figure 39.2
Endothermic reaction

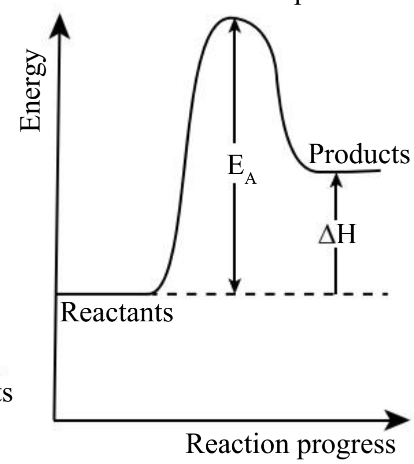
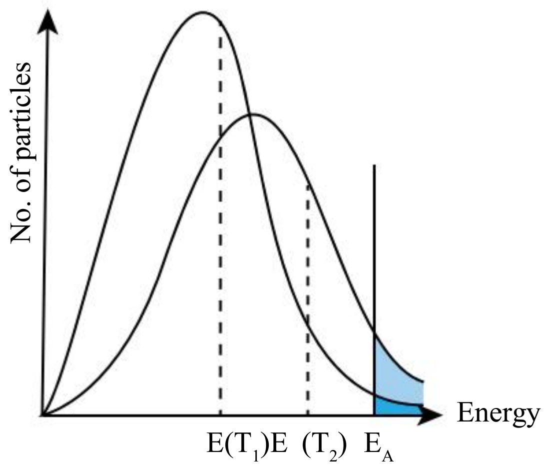


Figure 39.3
Molecular energy distribution



Exothermic reaction:
- Energy released to surroundings
- Surroundings warmer
- ΔH is negative ($-\Delta H$)

Endothermic reaction:
- Energy absorbed from surroundings
- Surroundings cooler
- ΔH is negative ($+\Delta H$)

$E(T)$ = 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

Notes

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 provides more particles available for reaction. As there are more particles, there will be a greater chance of successful collision with hydrogen ions. As there are more successful collisions per second, there will be a greater rate of reaction.

Set 39: Exercises

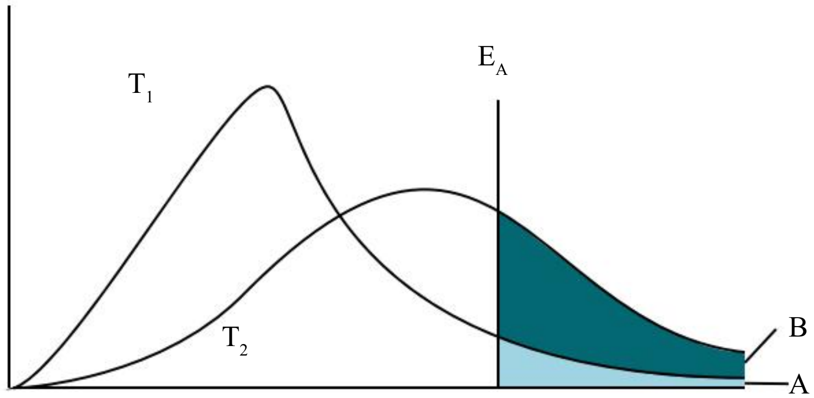
Notes

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.

Set 39: Rates of reaction

Notes

4. A student was experimenting with ways of increasing the reaction rate between marble chips (calcium carbonate) and hydrochloric acid.
   (a) Write a balanced ionic equation for the reaction between calcium carbonate and hydrochloric acid.
   (b) Give 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 a reaction profile diagram for this reaction given that the activation energy for the reaction is 250 kJ.
   (c) Annotate the reaction profile diagram 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.

9. When unleaded petrol was introduced, catalytic converters were added to cars. These converters are designed to control the emissions produced in the combustion of petrol.
   (a) Find out which chemicals are used in catalytic converters. Explain how these converters reduce the amount of noxious gases produced by cars.
   Include equations in your answer.
   (b) Leaded petrol cannot be used in cars with catalytic converters. Explain why.

Organic chemistry

The problem sets in the organic chemistry section of *Exploring Chemistry Year 11: Experiments, Investigations and Problems*, provides opportunities for students to explore:

- **elemental carbon and its allotropes, including graphite, diamond and fullerenes, in terms of their significantly different structures and physical properties**
- **the properties of covalent molecular substances**
- **hydrocarbons, including alkanes, alkenes and benzene by making models**
- **molecular models to show the arrangement of atoms and bonding in covalent molecular substances and to help draw structural formulae**
- **the characteristic reactions of alkanes, alkenes and benzene such as combustion, addition reactions for alkenes and substitution reactions for alkanes and benzene**

Set 40: Naming and drawing hydrocarbons

Notes

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. The rules used here include those revised in 1993.

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.

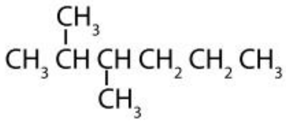
Type of hydrocarbon	Functional group	Suffix
alkane	C – C	-ane
alkene	C = C	-ene
alkyne	C ≡ C	-yne

3. Groups other than alkyl groups can result from addition and substitution reactions. The names of halogen groups are as follows.

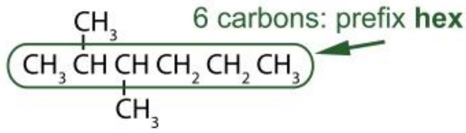
Type of substituted hydrocarbon	Functional group	Name of group
fluorocarbon	-F	fluoro-
chlorocarbon	-Cℓ	chloro-
bromocarbon	-Br	bromo-
iodocarbon	-I	iodo-

The steps for naming hydrocarbons

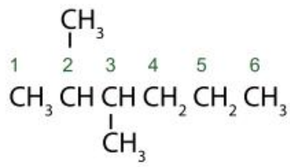
The following structure will be used to illustrate the rules for naming hydrocarbons in each step:



Step 1: Identify the longest continuous carbon chain containing the functional group. This determines the prefix for the name.



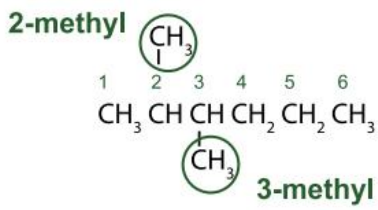
Step 2: Assign a number to each carbon atom in the longest continuous chain starting at the end of the chain that results in the lowest possible number for the position of the functional group.



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.

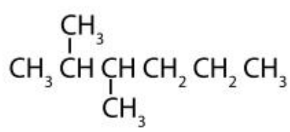


Step 5: Add the names of the groups in front of the chain name, indicating the position of the group using the number of the carbon to which it is attached. If a group occurs more than once, use the prefix di-, tri-, tetra-, etc to indicate the number present and indicate the position of each as described 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



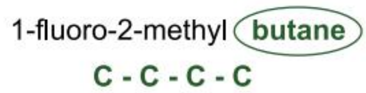
Set 40: Naming and drawing hydrocarbons

Notes

The steps for naming hydrocarbons

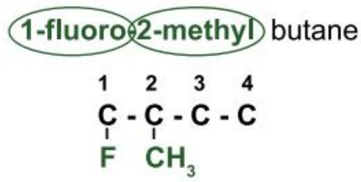
The following compound will be used to illustrate the rules for drawing structural formulas of hydrocarbons: 1-fluoro-2-methylbutane

Step 1: Identify and draw a carbon skeleton of the parent chain.

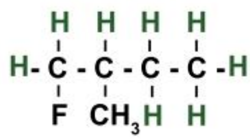


Step 2: Number the carbons in the chain.

Step 3: Add the funtional groups to the appropriate carbons on the chain.

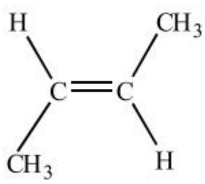


Step 4: Add hydrogen atoms to give each carbon four (4) bonds.

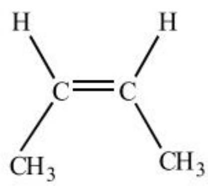


cis-trans isomers

In alkenes, rotation about the double bond is not possible. This results in the existance of *cis-trans* isomers in which the atoms and their bonding is the same but the arrangement of atoms in space can be different. The two different isomers are distinguished by placing the prefix ***trans-*** or ***cis-*** in front of the name, as follows



trans-but-2-ene



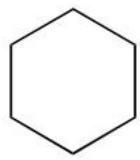
cis-but-2-ene

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.

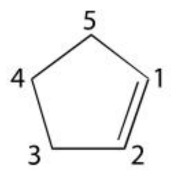
1. The prefix cyclo- is placed in front of the name indicating the number of carbon atoms.

cyclohexane (C_6H_{12})



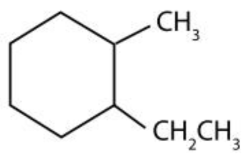
2. The carbon atoms are numbered starting at the functional group and for double bonds the numbering is such that the carbon atoms either side of the multiple bond have the numbers 1 and 2.

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})

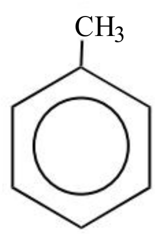


Aromatic compounds based on benzene

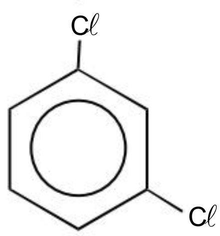
The naming of these compounds is very similar to the alicyclic compounds and because the compounds we will consider are based on the benzene structure, the names are based on the name benzene. This group is called 'aromatic'. For example:

(a) benzene (C_6H_6)

(b) methylbenzene, also known as toluene (C_7H_8)



(c) 1,3-dichlorobenzene ($C_6H_4C\ell_2$)



Notes

Set 40: Naming and drawing hydrocarbons

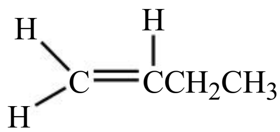
Notes

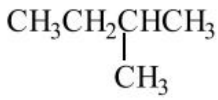
Set 40: Exercises

Petroleum is a complex mixture of hydrocarbons and can include some of the substances in **questions 1 and 2**.

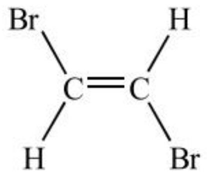
1. Write systematic names for the following compounds.

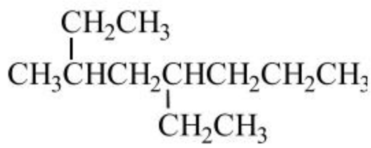
(a) $CH_3CH_2CH_2CH_2CH_3$

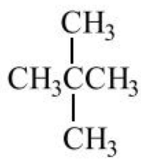
(b) 

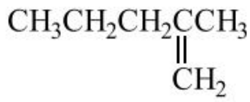
(c) 

(d) $CHC\ell_2CH_2C\ell$

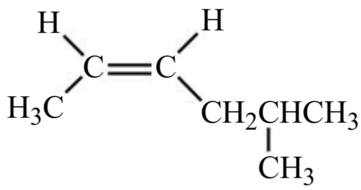
(e) 

(f) 

(g) 

(h) 

(i) $CF_3CH_2CH_2CH_2CH_3$

(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) *trans*-hept-3-ene
   - (g) 1,1-dichloro-*cis*-but-2-ene
   - (h) 5-ethylhept-1-ene

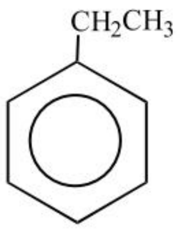
3. Petroleum contains many substances that have the same formula but different structures. Draw the structural isomers and write systematic names for
   - (a) all the isomers of:
     - (i) pentane
     - (ii) pentene
   - (b) four isomers of C_4H_9Br

Notes

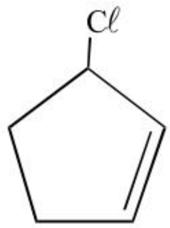
Alicyclic and aromatic compounds are often used as starting materials for important agricultural chemicals including pesticides and herbicides. Examples of substituted alicyclic and aromatic compounds are included in **questions 4 and 5**.

4. Write systematic names for the following compounds.

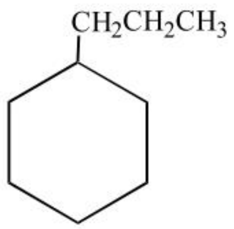
(a)



(b)

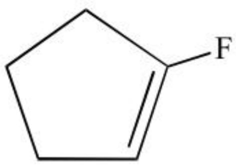


(c)

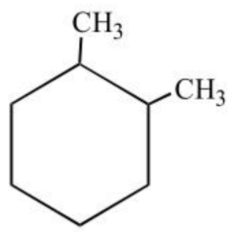


(d)

(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 41: Reactions of hydrocarbons

Notes

The reactions you will deal with in this section are combustion, substitution and addition reactions.

Examples

Combustion of hydrocarbons as shown in the combustion of octane

$2\ C_8H_{18} + 2\ 5O_2 \rightarrow 16\ CO_2 + 18\ H_2O$

Substitution in alkanes as shown in the chlorination of propane

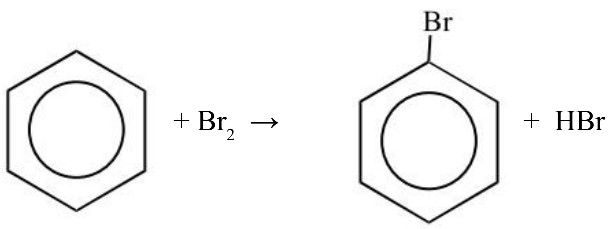
$C_3H_8 + C\ell_2 \rightarrow C_3H_7C\ell + HC\ell$

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, Markovnikoff'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



Propane

Set 41: Exercises

1. Hydrocarbons are mostly used as fuels as they all undergo combustion to produce heat. Write equations for reactions between
   (a) propane and oxygen when ignited.
   (b) ethene and air when ignited.
   (c) benzene and air when ignited
   (d) ethylbenzene and oxygen when ignited

Notes

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

Set 42: Organic chemistry

Notes

Set 42: Exercises

1. Crude oil and natural gas are major sources of hydrocarbons. Research the composition of these non-renewable sources of hydrocarbons, the method and extent of separation currently achieved and the uses to which these separated hydrocarbons are put.

2. One use of the hydrocarbons extracted from crude oil and natural gas is in the production of useful organic products that include solvents, refrigerants and polymers. The more useful but less common components of crude oil and natural gas are alkenes. Discuss why alkenes are more useful as starting materials than the more abundant alkanes and the role that the process of catalytic cracking has in increasing the amounts of alkenes in the products of oil refining.

3. Research the use of ethene as a raw material for the production of useful organic compounds.

4. Research the production, properties and historical uses of CFCs.

CFCs in spray cans

5. Research the impact of CFCs on atmospheric ozone. Include details of the ways in which CFCs are released into the atmosphere and where and how they interact with the ozone. You should also discuss why their interaction with ozone is a major environmental concern.

6. Research methods of removing hydrocarbons from soil contaminated with diesel or lubricating oils.

7. Currently over 90% of hydrocarbons derived from crude oil and natural gas are used as fuels, the rest are used in petrochemical industries for other uses such as solvents and raw materials for the production of other useful substances. There is a view that the use of hydrocarbons as fuels should be reduced and more hydrocarbons reserved for use as raw materials. Research arguments for and against this view.

Extended answer questions

Notes

Where applicable, use equations, diagrams and illustrative examples of the chemistry you are describing. Your answer should focus on the relevant chemical content specific to the situations described, ensure that it is presented in a logical and coherent manner.

1. **Removing and preventing hydrocarbon contamination**.
   As an environmental scientist you are called to a truck repair workshop where local shire environmental officers have detected significant areas of the yard that have been contaminated with hydrocarbons. The main source of these hydrocarbons seems to be from the practice of washing the engines and transmissions of trucks prior to servicing and repairs, as well as washing parts prior to being repaired. The wash down area is located outside the workshop.

   You are engaged to devise strategies, procedures and possibly equipment to:
   (a) remove the contaminating hydrocarbons from the soil, and
   (b) prevent further contamination while at the same time separating the hydrocarbons from the water used for washing so that it can be recycled.

2. **Ground water contamination**
   A council gardener noticed that bore water used to reticulate gardens gradually developed a foul smell over a period of 6 months. The gardens are located approximately 100 m downhill from the local oil recycling plant that deals with used engine oil. As an environmental hydrologist you are engaged by the plant manager and the council to determine the cause of the foul smell in the water.
   The manager of the recycling plant is confident that there has been no leakage from his plant and is as keen as the council to determine the source of the contamination. Your task is to determine whether or not the contamination is from the oil recycling plant.
   (a) Write possible explanations consistent with the observations.
   (b) Design an investigation or a number of investigations to test your theory or theories. Ensure that your design can be conducted safely.
   (c) What recommendations would you make to the client that could reduce or stop the problem?

Answers

Measurement in Chemistry

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) 6.60×10^2 (c) 1.588×10^{-3} (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 reproducible (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.

Mixtures, kinetic theory and gases

Set 4: 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.

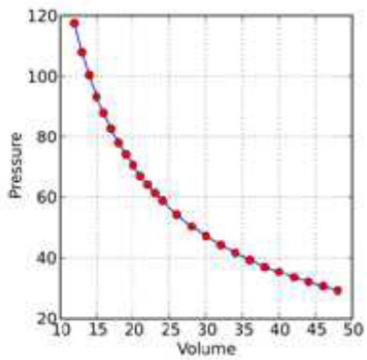
2 (a) x axis - time, y axis - temperature/°C. Freezing when the temperature is constant line is horizontal.
(b) The freezing point of pure ethanol is -114°C labelled on y axis. Pure ethanol contains only one substance - ethanol molecules.
(c) Homogeneous, water.

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 dyes 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 5: 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 6: Macroscopic properties of matter

1 Supercritical carbon dioxide exists at high pressures and behaves as a gas and as a liquid. Whether a substance fits the solid/liquid/gas phase depends on the temperature and pressure it is observed at. There is also some debate about when a liquid is a liquid and when it is a solid see http://www.newscientist.com/article/dn25441-longest-experiment-sees-pitch-drop-after-84year-wait.html#.VOv6YffnUeSo

2 Essential to include Sodium (Na^+), Potassium (K^+), Calcium (Ca^{2+}), Magnesium (Mg^{2+}). Bicarbonate (HCO_3^-)

3 (a) Particles have more kinetic energy and collide with the walls with more energy more frequently causing more pressure rupturing the can.
(b) Air particles when cooled lose kinetic energy, will collide with less energy and less frequently resulting in less pressure. The attractive forces in the balloon overcome the air pressure and the balloon collapses. When returned to room temperature the air pressure forces the balloon to be bigger.

Answers

4 You might need a lot of salt! A pressure cooker is a better way to cook at higher temperature and faster. Much of pasta cooking however is water absorption, which is not so temperature dependant.

5 At higher elevations cooking times would be longer as temperature is lower

6 (a) To keep fish alive and keep an oxidising environment to prevent formation of H_2S
(b) Decreases
(c) Heating of waterways by using water to cool power stations and industrial machinery
(d) Lowers oxygen content. See part a and b

7 (a) Gas flows from high pressure to low. So it must be kept at the some pressure. Decreases volume of the gas - more portable
(b) Shallow SCUBA use only cleaned air. Deep SCUBA use helium added to their air.
(c) Nitrogen narcosis is a dulling of the senses caused by the increased solubility of nitrogen at depth (high pressure). The bends are caused when a diver surfaces quickly and the dissolved nitrogen comes out of solution and causes great pain.

8 Charles' Law describes the relationship between pressure and temperature. This can be explained using the Kinetic Theory. If the volume of a container of gas is increased at a constant temperature the particles have further to travel to strike the walls of the container. This means less frequent collisions resulting in a lower pressure.

Atomic structure and bonding

Set 7: 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) ruthenium Ru (k) thorium Th (l) astatine At
(m) germanium Ge (n) technetium Tc (o) barium Ba

2 (a) Na sodium (b) Re rhenium (c) Cℓ chlorine (d) Au gold
(e) Zr zirconium (f) Pu plutonium (g) Ce caesium (h) As arsenic
(i) Fe iron (j) Co cobalt (k) P phosphorus (I) Kr krypton
(m) Zn zinc (n) K potassium (o) Sn tin

Set 8: Atoms and isotopes

1. (a) label nucleus and electron cloud (b) 3 protons, 4 neutrons and 3 electron (c) Li

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

Symbol	Element	A	Z	Neutrons
${}^{14}_{6}C$	Carbon	14	6	8
${}^{35}_{17}Cl$	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 ${}^{3}_{1}H$

Answers

Set 9: Atomic structure and the Periodic Table

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

	Symbol	Atomic number	Number of protons	Number of electrons	Number of neutrons
Example	${}^{19}_{9}F$	9	9	9	10
A	${}^{16}_{8}O$	8	8	8	8
B	${}^{17}_{8}O$	8	8	8	9
C	${}^{16}_{8}O^{2-}$	8	8	10	8
D	${}^{17}_{9}F$	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

Set 10: Ionisation energy

1. The energy required to remove a mol of electrons in the third shell of a mol 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. On the graph the first ionisation energies of successive elements increase across a Period until Group 18 to then fall for the first element in the next 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 11: Periodic Trends

1. The metallic character increases.
   They have the same number of electrons.
   The first ionisation energy decreases.
   The electronegativity decreases.
   The atomic radius increases.

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 12: 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, the 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, page 126
  (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 us a chemical 'fingerprint' of that material and this fingerprint can be used to positively identify it.

3 (a) Tc - radioactive tracer
  (b) Mo - used to manufacture Tc 99
  (c) Co - as a tracer and to sterilize medical equipment

4 Dalton - atomic theory, Thompson - electron, Rutherford- nucleus and electrons, Bohr- shells, Chadwick- neutron

5 To identify atoms in a sample.

6 Your diagram should include an atomiser, a radiation source, wavelength selector, detector, amplifier and signal processor

Set 13: 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

Answers

Set 14: Bonding and properties

1. (a) covalent molecular (b) metallic
   (c) ionic (d) covalent molecular
   (e) ionic (f) covalent molecular
   (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 bonding present. Requires energy to overcome the bonding.
   (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 are 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 3D 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 are no charge carriers so it does 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 are bound together using strong primary ionic bonds.

Set 15: Uses, properties and structure

1. (a) Plastic handle (covalent molecular substance) does not conduct electricity.
   (b) Plastic or wooden handle (covalent molecular substance) does not conduct heat.
   (c) Conventional ovens rely on conduction. Plastic container is a poor conductor and might melt.
   (d) Diamond is hard due to the strong network of covalent bonds and can be used to write on glass.
   (e) Lead is cheap, low melting point for casting, dense and low reactivity.

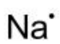
2. Lead is very malleable, soft (and has a relatively low melting point), making it easy to mold.

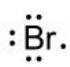
3. Gold: jewellery and coinage: doesn't tarnish, very malleable and has colour. Electric circuits: excellent conductor of electricity and ductile. Shield in spacecraft: excellent reflector. Dental work: non-toxic and unreactive.

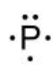
4. (a) Covalent molecular (b) Ionic
   (c) Covalent molecular (d) Ionic
   (e) Metallic (f) Covalent network

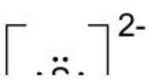
5. (a) Iron (b) Iron, copper(II) oxide and titanium dioxide.

Set 17: Electron dot diagrams

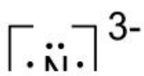
1. (a) 

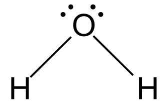
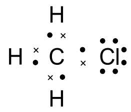
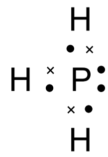
   (b) 

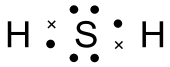
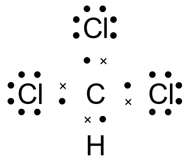
   (c) 

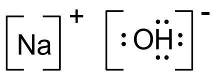
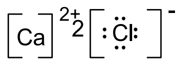
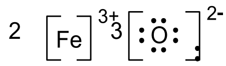
   (d) 

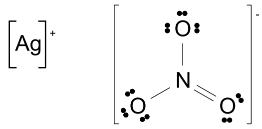
   (e) 

   (f) 

2 (a)  (b)  (c) 

(d)  (e)  (f) 

3 (a)  (b)  (c) 

(d)  (e)  (f) 

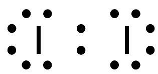
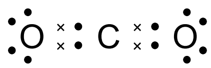
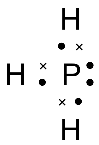
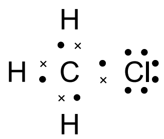
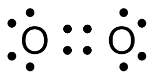
Set 18: Molecular shape

1. The electronegativity of iodine is much less compared with fluorine as the outer electrons 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- Na^+ — Electron transferred / pulled away from sodium by chlorine.
(b) Equal sharing forms a pure covalent bond, unequal sharing forms a polar covalent bond. Transferring of an electron forms an ionic bond.

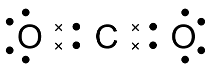
3. (b), (c), (d)

4.

	electron dot diagram	shape	polarity
a		linear	non-polar
b		linear	non -polar
c		pyramidal	polar
d		tetrahedral	Polar
e		linear	non-polar

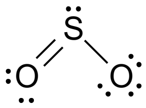
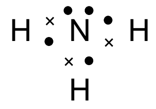
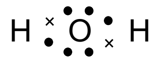
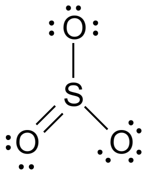
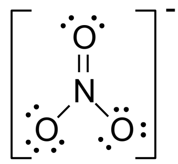
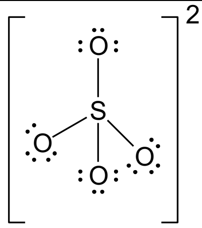
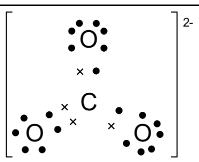
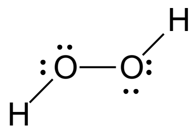
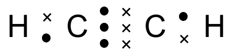
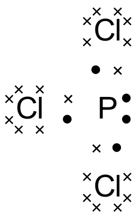
5. The bent and pyramidal shapes will always be polar as they are asymmetrical due to the lone pair of electrons.

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.



Answers

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 19: 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. (a) Boiling points $HC\ell < HBr < HI < HF$ (b) Boiling points $O_2 < CH_3C\ell < NH_3 < H_2O$

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) soluble
(d) insoluble (e) insoluble or slightly soluble (f) soluble

6. (a) soluble (b) insoluble (c) insoluble
(d) soluble (e) soluble (f) insoluble

Set 20: 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

Chemical reactions and stoichiometry

Set 21: Relative atomic mass and mass spectroscopy

1. 35.48 2. 28.11 3. 207.24 4. Another isotope of barium with smaller atomic mass

5. 24.32 6. 72.5% of Cu-63 and 27.5% of Cu-65 7. **Sample A**: 55.0% of Br-79 and 45.0% 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 22: 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 23: 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 24: Interpretation of formulae

1. (a) 5.00 mol (b) 0.100 mol 2. (a) 7.00 mol (b) 0.500 mol

3. (a) 160.0 g (b) 0.467 g (c) 12.0 g

4. (a) 386 g (b) 1.57 g (c) 42.8 g

5. (a) 7.50 mol (b) 2.50 mol (c) 7.50 mol (d) 30.0 mol
(e) 2.50 mol (f) 10.0 mol

6. 448 g 7. 276 g 8. 240 g 9. 600 mol nitrogen atoms 10. 1.77 g

Set 25: 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% Cℓ (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% Aℓ 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% Aℓ, 34.63% H_2O Kaolinite: 20.90% Aℓ, 13.95% H_2O

10. (a) 59.9% Ti (b) 3.12 tonnes

Set 26: 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 27: Ionic equations

1. $KC\ell(s) \rightarrow K^+(aq) + C\ell^-(aq)$
2. $Ba(NO_3)_2(s) \rightarrow Ba^{2+}(aq) + 2\ NO_3^-(aq)$
3. $NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$
4. $H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$
5. $ZnO(s) + 2\ H^+(aq) \rightarrow Zn^{2+}(aq) + H_2O(\ell)$
6. $CaCO_3(s) + 2\ H^+(aq) \rightarrow Ca^{2+}(aq) + CO_2(g) + H_2O(\ell)$
7. $Pb(s) + 2\ Ag^+(aq)$ — $Pb^{2+}(aq) + 2\ Ag(s)$
8. $Ag^+(aq) + C\ell^-(aq) \rightarrow AgC\ell(s)$
9. $CO_2(g) + Ca^{2+}(aq) + 2\ OH^-(aq) \rightarrow CaCO_3(s) + H_2O(\ell)$
10. $2\ H_3PO_4(aq) + 3\ Ag_2CO_3(s) \rightarrow 2\ Ag_3PO_4(s) + 3\ CO_2(g) + 3\ H_2O(\ell)$
11. $H_2S(g) + 2\ Ag^+(aq) \rightarrow Ag_2S(s) + 2\ H^+(aq)$
12. $2\ H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(\ell) + CO_2(g)$
13. $Pb^{2+}(aq) + 2\ I^-(aq) \rightarrow PbI_2(s)$
14. $A\ell(OH)_3(s) + OH^-(aq) \rightarrow [A\ell(OH)_4]^-(aq)$

Set 28: Equations and observations

1. $CaCO_3(s) + 2\ H^+(aq) \rightarrow Ca^{2+}(aq) + CO_2(g) + H_2O(\ell)$
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. $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$ White precipitate formed.

8. $Pb^{2+}(aq) + 2\ I^-(aq) \rightarrow PbI_2(s)$ Yellow precipitate formed.

9. $2\ Au^{3+}(aq) + 3\ Cu(s) \rightarrow 2\ Au(s) + 3\ Cu^{2+}(aq)$
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. $Ni(s) + Cu^{2+}aq)$ — $Ni^{2+}(aq) + Cu(s)$ 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.

Set 29: 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. $2\ KC\ell O_3 \rightarrow 2\ KC\ell + 3\ O_2$ 1.20 mol 6. $2\ HgO \rightarrow 2\ Hg + O_2$ 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 UF_6 = 0.978 g

9. (a) m = CaO = 258 kg (b) m CO_2 = 202 kg

Set 30: 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

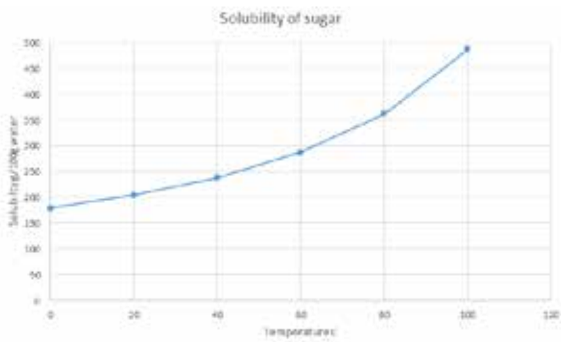
7. $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$ (a) V CO_2 = 1576 L (b) V CH_4 = 1576 L 8. m NaCN = 1.06 tonnes

9. (a) $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$ $2\ C_8H_{18} + 25\ O_2 \rightarrow 16\ CO_2 + 18\ H_2O$
   (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 %

Solutions and acidity

Set 31: Solutions



1. (a) See graph
   (b) Sugar solubility increases rapidly with temperature
   (c) 30°C = 220 g/100g water, 70° = 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) $Cu^{2+}(aq) + 2\ OH^{-}(aq) \rightarrow Cu(OH)_2(s)$
   (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

Answers

Set 32: 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 33: 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) $Na_2CO_3(s) + 2\,HC\ell\,(aq) \rightarrow 2\,NaC\ell\,(aq) + CO_2(g) + H_2O(\ell)$ (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

11. (a) 0.398 mol (b) 27.1 g (c) 4.51 L

12. (a) 0.167 mol (b) 3.79 L

13. 2.42×10^4 L

14. 1.12×10^4 L

15. (a) 14.4 g (b) 0.988 L (c) Nil

16. (a) 0.500 L (b) 0.400 L

17. 1.54 L

18. 83.3 % $CaCO_3$

19. 92.2 % silver

Set 34: 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) $[H^+] = 0.100 \text{ mol L}^{-1}$ pH = 1.00 (b) $[H^+] = 0.00500 \text{ mol L}^{-1}$ pH = 2.30
(c) $[H^+] = 2.00 \text{ mol L}^{-1}$ pH = 0.301

4. (a) Lemon juice $[H^+] = 1.00 \times 10^{-3} \text{ mol L}^{-1}$ (b) Dish washing solution $[H^+] = 1.00 \times 10^{-11} \text{ mol L}^{-1}$
(c) Pool acid $[H^+] = 10.0 \text{ mol L}^{-1}$ (d) Orange juice $[H^+] = 2.75 \times 10^{-5} \text{ mol L}^{-1}$
(e) Swimming pool water $[H^+] = 2.51 \times 10^{-8} \text{ mol L}^{-1}$

5. Concentration changed by a factor of 1000

6. $[H^+] = 1.80 \text{ mol L}^{-1}$

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 35: 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) $4.04 \times 10^{-3} \text{ mol L}^{-1}$

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) $0.000135 \text{ mol L}^{-1}$ (b) $0.000270 \text{ mol L}^{-1}$

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

Answers

Set 36: 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}

Set 37: Acids and bases

1. $SO_2(g) + H_2O(\ell) \rightarrow H_2SO_3(aq) \rightarrow H^+(aq) + HSO_3^-(aq)$
2. $CaCO_3(s) + 2\,H^+(aq) \rightarrow Ca^{2+}(aq) + CO_2(g) + 2\,H_2O(\ell)$
3. $Al_2O_3(s) + 6\,H^+(aq) \rightarrow 2\,Al^{3+}(aq) + 3\,H_2O(\ell)$
4. $CO_2(g) + H_2O(\ell) \rightarrow H_2CO_3(aq) \rightarrow H^+(aq) + HCO_3^-(aq)$
5. $Ca(OH)_2(s) + 2\,H^+(aq) \rightarrow Ca^{2+}(aq) + 2\,H_2O(\ell)$
   $CaCO_3(s) + 2\,H^+(aq) \rightarrow Ca^{2+}(aq) + CO_2(g) + 2\,H_2O(\ell)$
6. $Fe_2O_3(s) + 6\,H^+(aq) \rightarrow 2\,Fe^{3+}(aq) + 3\,H_2O(\ell)$
7. $H_3PO_4(aq) \rightarrow H^+(aq) + H_2PO_4^-(aq)$
   $H_2PO_4^-(aq) \rightarrow H^+(aq) + HPO_4^{2-}(aq)$
   $HPO_4^{2-}(aq) \rightarrow H^+(aq) + PO_4^{3-}(aq)$
8. $H_2SO_4(aq) + H_2O(\ell) \rightarrow H_3O^+(aq) + HSO_4^-(aq)$
   $HSO_4^-(aq) + H_2O(\ell) \rightarrow H_3O^+(aq) + SO_4^{2-}(aq)$
9. Powdered $CaCO_3$
   $CaCO_3(s) + 2\,H^+(aq) \rightarrow Ca^{2+}(aq) + CO_2(g) + 2\,H_2O(\ell)$
10. (a) $NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$
    (b) $A\ell(OH)_3(s) + OH^-(aq) \rightarrow [A\ell(OH)_4]^-(aq)$

Energy changes and rates of reaction

Set 38 Energy changes

1. Exothermic – a, b, e, f, g   Endothermic – c, d
2. (a) 2.80×10^3 kJ (b) 467 kJ (c) 1.70×10^3 kJ (d) 105 g
3. (a) 242 kJ (b) 1210 kJ (c) 16.9 g (d) $2\,H_2O(\ell) + 484\text{ kJ} \rightarrow 2\,H_2(g) + O_2(g)$
4. (a) $2\,H_2(g) + CO(g) \rightarrow CH_3OH(g)$ $\Delta H = -90$ kJ (b) 2232 kJ (c) 7.78 g (d) 198 g
5. (a) They are molecules representing different structural combinations of an element (b) $O_3(g) + NO(g) \rightarrow O_2(g) + NO_2(g)$ + 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 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. $CH_3COONa(aq) \rightarrow CH_3COONa(s)$ + Heat

Set 39: Rates of reactions

1 (b) endothermic as the ΔH is positive (c) (i) $\Delta H = -20$kJ (ii) Ea = 10kJ (d) The ΔH is the same but the E_A 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) E_A – 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 + $N_2(g) + 3\,H_2(g) \rightarrow 2NH_3(g)$
   (b) Increasing temperature increases the number of collisions and particles that have enough energy to react.
   (c) Increasing the pressure increases the concentration so collisions can occur more frequently.
   (d) A catalyst provides an alternative reaction pathway with a lower energy of activation.

4. (a) $CaCO_3(s) + 2\,H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(\ell) + CO_2(g)$
   (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 E_A.
   (c) Decrease the rate of reaction. Acetic acid is a weaker acid so $[H^+]$ available to collide and react is less.

5. (a) $2\,CO(g) + O_2(g) \rightarrow 2\,CO_2(g)$ + 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 E_A.

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.

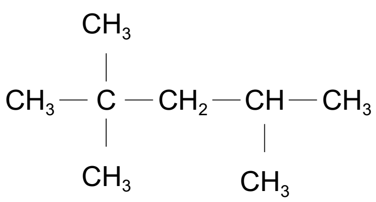
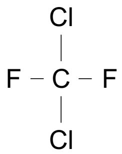
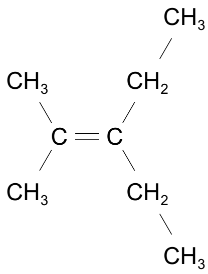
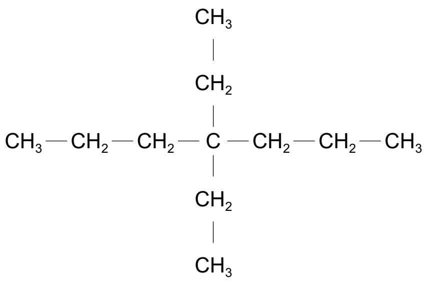
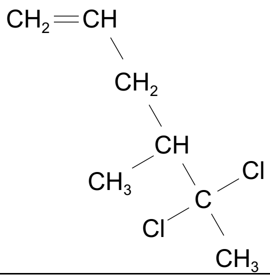
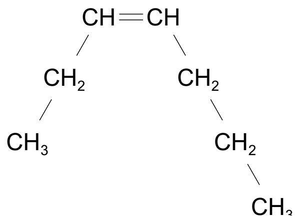
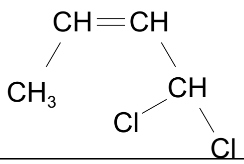
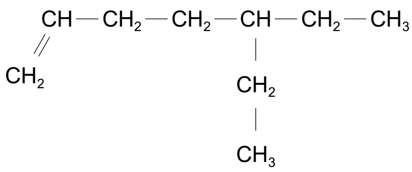
9. (a) Platinum and rhodium and or palladium. $2\ NO \rightarrow N_2 + O_2$ $2\ NO_2 \rightarrow N_2 + 2\ O_2$ $2\ CO + O_2 \rightarrow 2\ CO_2$
(b) Lead in leaded fuel reacts with the metal catalyst in the converter destroying its capacity to function.

Organic chemistry

Set 40: Naming and drawing hydrocarbons

1.
(a) pentane
(b) but-1-ene
(c) 2-methylbutane
(d) 1,1,2-trichloroethane
(e) *trans*-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-*cis*-hex-2-ene

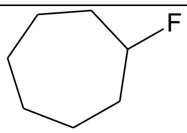
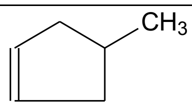
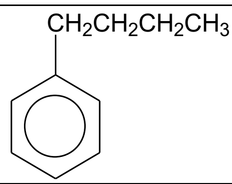
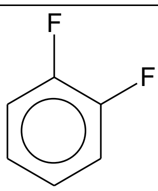
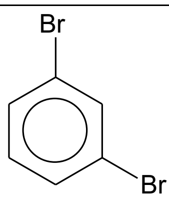
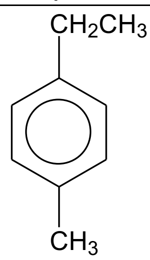
2.

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

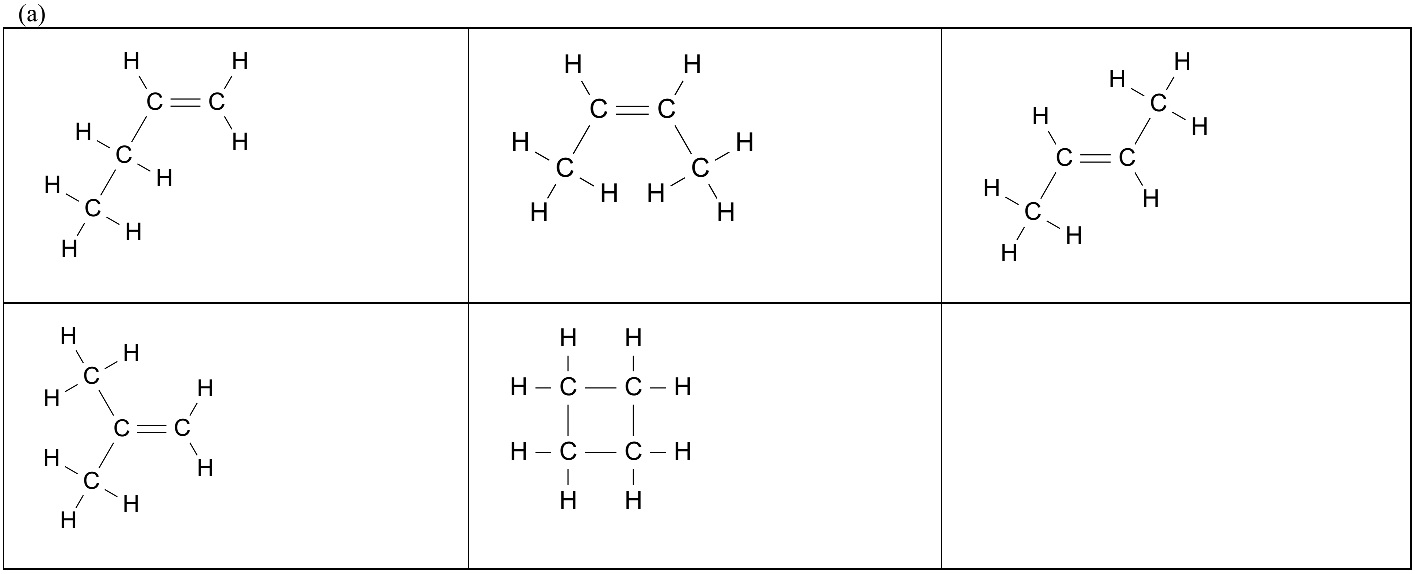
4.
(a) ethylbenzene
(b) 3-chlorocyclopentene
(c) propylcyclohexane
(d) 3-bromo-5-chlorocycloheptene
(e) 1-fluorocyclopentene
(f) 1,2-dimethylhexane

5.

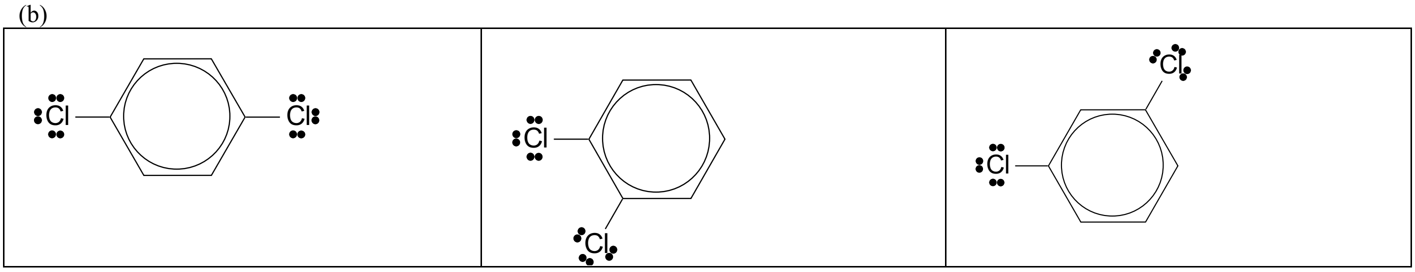
(a)		(b)		(c)	
(d)		(e)		(f)	

Answers

6. (a)

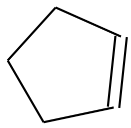
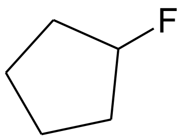


(b)



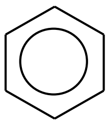
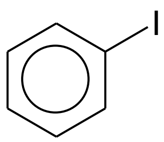
Set 41: Reactions of hydrocarbons

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) $2\ C_6H_5CH_2CH_3 + 21\ O_2 \rightarrow 16\ CO_2 + 10\ H_2O$

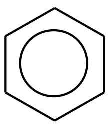
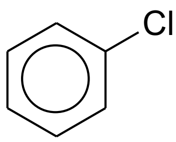
2. (a) $CH_3CH_3 + C\ell_2 \rightarrow CH_3CH_2C\ell + HC\ell$
   (b) $CH_3CH{=}CH_2 + Br_2 \rightarrow CH_3CHBrCH_2Br$
   (c) $CH_3CH{=}CHCH_3 + HC\ell \rightarrow CH_3CHC\ell CH_2CH_3$
   (d) $CH_3CH_2CH{=}CHCH_3 + H_2 \xrightarrow{Pt} CH_3CH_2CH_2CH_2CH_3$
   (e)  + HF → 

   (f) benzene + $Br_2 \rightarrow$ bromobenzene + HBr
   (g) $CH_3CH{=}CH_2 + 2\ C\ell_2 \rightarrow CH_3CHC\ell CH_2C\ell$

3. (a) methane CH_4, fluorine and chlorine — $CH_4 + F_2 \rightarrow CH_3F + HF$
   (b) ethene CH_2CH_2 and hydrogen chloride — $CH_2CH_2 + HC\ell \rightarrow CH_3C\ell CH_2C\ell$
   (c) ethene CH_2CH_2 and chlorine — $CH_2CH_2 + C\ell_2 \rightarrow CH_2C\ell CH_2C\ell$
   (d) propene and hydrogen chloride — $CH_3CHCH_2 + HC\ell \rightarrow CH_3CHC\ell CH_3$
   (e) benzene and iodine

 + I_2 → HI + 

   (f) benzene and chlorine

 + $C\ell_2$ → $HC\ell$ + 

Appendix: Chemical data

Names and symbols of ions

Ion name	Formula
aluminium	$A\ell^{3+}$
ammonium	NH_4^+
barium	Ba^{2+}
bromide	Br^-
caesium	Cs^+
calcium	Ca^{2+}
carbonate	CO_3^{2-}
chloride	$C\ell^-$
chromate	CrO_4^{2-}
chromium(III)	Cr^{3+}
cobalt(II)	Co^{2+}
copper(II)	Cu^{2+}
cyanide	CN^-
dichromate	$Cr_2O_7^{2-}$
dihydrogenphosphate	$H_2PO_4^-$
ethanoate (acetate)	CH_3COO^-
fluoride	F^-
hydrogen	H^+
hydrogencarbonate	HCO_3^-
hydrogenphosphate	HPO_4^{2-}
hydrogensulfate	HSO_4^-
hydroxide	OH^-
iodide	I^-
iron(II)	Fe^{2+}
iron(III)	Fe^{3+}
lead(II)	Pb^{2+}
lithium	Li^+
magnesium	Mg^{2+}
manganese(II)	Mn^{2+}
nickel(II)	Ni^{2+}
nitrate	NO_3^-
nitride	N^{3-}
nitrite	NO_2^-
oxalate	$C_2O_4^{2-}$
oxide	O^{2-}
permanganate	MnO_4^-
phosphate	PO_4^{3-}
potassium	K^+
rubidium	Rb^+
silver	Ag^+
sodium	Na^+
strontium	Sr^{2+}
sulfate	SO_4^{2-}
sulfide	S^{2-}
sulfite	SO_3^{2-}
zinc	Zn^{2+}

write the molecular formulae of commonly encountered molecules that have non – systematic names including

ammonia	NH_3
water	H_2O
hydrogen peroxide	H_2O_2
ethanoic acetic acid	CH_3COOH
hydrochloric acid	$HC\ell$
nitric acid	HNO_3
carbonic acid	H_2CO_3
sulfuric acid	H_2SO_4
sulfurous acid	H_2SO_3
phosphoric acid	H_3PO_4

Appendix: Chemical data

Colours of selected substances

In general, ionic solids have the same colour as that of any coloured ion they contain.
Two colourless ions in general produce a white solid.
Selected exceptions to these two basic rules are noted below.

Ionic Solid	Colour
copper(II) carbonate	green
copper(II) chloride	green
copper(II) oxide	black
copper(II) sulfide	black
lead(II) iodide	yellow
lead(II) sulfide	grey
manganese(IV) oxide	black
silver carbonate	yellow
silver iodide	pale yellow
silver oxide	brown
silver sulfide	black

Other coloured substances

Most gases and liquids are colourless, and most metals are silvery or grey. Selected exceptions to these basic rules are noted below.

Ionic Solid	Colour
copper(s)	salmon pink
gold(s)	yellow
nitrogen dioxide(g)	brown
sulfur(s)	yellow

Coloured halogens

Halogen	Colour of free element
F_2(g)	yellow
$C\ell_2$ (g)	greenish-yellow
Br_2 (ℓ)	red
I_2 (s)	purple

Halogen	Colour of halogen in aqueous solution
$C\ell_2$ (aq)	pale yellow
Br_2 (aq)	orange
I_2 (aq)	brown

Halogen	Colour of halogen in organic solvent
Br_2	red
I_2	purple

Solubility rules - ionic solids in water

Soluble in water

Soluble	Exceptions	
	Insoluble	Slightly soluble
Most chlorides	$AgC\ell$, $Hg_2C\ell_2$	$PbC\ell_2$
Most bromides	AgBr, Hg_2Br_2, $HgBr_2$	$PbBr_2$
Most iodides	AgI, Hg_2I_2, HgI_2, PbI_2	
All nitrates	Nil	
Most sulfates	$SrSO_4$, $BaSO_4$, $HgSO_4$, $PbSO_4$	$CaSO_4$, Ag_2SO_4

Insoluble in water

Insoluble	Exceptions	
	Soluble	Slightly soluble
Most hydroxides	NaOH , KOH, $Ba(OH)_2$	$Ca(OH)_2$, $Sr(OH)_2$
Most carbonates	Na_2CO_3, K_2CO_3, $(NH_4)_2CO_3$	
Most phosphates	Na_3PO_4, K_3PO_4 $(NH_4)_3PO_4$	
Most sulfides	Na_2S, K_2S, $(NH_4)_2S$	

Soluble = more than 0.1 mole dissolves per litre
Slightly soluble = between 0.01 and 0.1 mole dissolves per litre
Insoluble = less than 0.01 mole dissolves per litre

Periodic Table of the Elements

Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8	Group 9	Group 10	Group 11	Group 12	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
1 **H** 1.008																	2 **He** 4.003
3 **Li** 6.968	4 **Be** 9.012			24 **Cr** 52.00	←Atomic number ←Symbol ←Molar mass							5 **B** 10.82	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.31											13 **Aℓ** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cℓ** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.87	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.38	31 **Ga** 69.72	32 **Ge** 72.63	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.96	43 **Tc** (98)	44 **Ru** 101.1	45 **Rh** 102.9	46 **Pd** 106.4	47 **Ag** 107.9	48 **Cd** 112.4	49 **In** 114.8	50 **Sn** 118.7	51 **Sb** 121.8	52 **Te** 127.6	53 **I** 126.9	54 **Xe** 131.3
55 **Cs** 132.9	56 **Ba** 137.3	57 ***La** 138.9	72 **Hf** 178.5	73 **Ta** 180.9	74 **W** 183.9	75 **Re** 186.2	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.1	79 **Au** 197.0	80 **Hg** 200.6	81 **Tℓ** 204.4	82 **Pb** 207.2	83 **Bi** 209.0	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** (226)	89 ****Ac** 227.0	104 **Rf**	105 **Db**	106 **Sg**	107 **Bh**	108 **Hs**	109 **Mt**									

* Lanthanide Series	58 **Ce** 140.1	59 **Pr** 140.9	60 **Nd** 144.2	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 152.0	64 **Gd** 157.3	65 **Tb** 158.9	66 **Dy** 162.5	67 **Ho** 164.9	68 **Er** 167.3	69 **Tm** 168.9	70 **Yb** 173.12	71 **Lu** 175.0
** Actinide Series	90 **Th** 232.0	91 **Pa** 231.0	92 **U** 238.0	93 **Np** 237.0	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lw** (260)

A relative atomic mass in brackets is the mass number of the isotope with the longest half-life