ACADEMIC ASSOCIATES

WACE Study Guide

CHEMISTRY

YR 11 ATAR COURSE

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TO THE STUDENT

The purpose of this guide is to assist students in their preparation for tests and examinations in the new ATAR Chemistry course for Units 1 and 2. The structure of the topics will allow students to use the book throughout the year.

The guide closely adheres to the W.A. School Curriculum and Standards Authority ATAR syllabus. Essential theory is interwoven with revision exercises so that students will be able to actively review core theory and concepts.

Science Understanding

Essential core theory for each topic of Science Understanding is covered clearly and in detail. Illustrations and worked examples are used extensively to assist students in their learning. Throughout each chapter, questions and exercises are integrated with theory to help students clarify and consolidate their understanding of new concepts.

Review questions at the end of each chapter provide a wide range of problems including a more challenging 'for the experts' contextual question. All questions and review exercises have detailed answers to provide students with immediate feedback and a means of enhancing their progress.

Chemistry Calculations

Chemistry calculations are an important part of the Chemistry course and these have been treated in great detail. A comprehensive range of worked examples are included to assist students in developing their skills in this area. The review questions and solutions also provide many opportunities for further independent learning.

Trial Tests

Trial tests for each major topic provide an ideal means of self assessment. The style and structure of these tests is similar to that proposed for the WACE examination. They contain a multiple choice section and a short and extended answer section. The marks allocated for each of the sections also reflect the weightings proposed by the SCSA for the examinations.

Chemistry is a most interesting and an enjoyable science to study. The practical work, in particular, holds a fascination for students. We hope that this study guide will help students to better understand the concepts they will encounter and to achieve greater success in the subject.

Michael Lucarelli and David Proctor

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UNIT 2

INTRODUCTION

SYLLABUS CHECKLIST

***Please note:** The Syllabus checklist below lists only the Science as a Human Endeavour and Science Understanding sections of the Chemistry Year 11 ATAR syllabus as at the time of printing. For a more detailed and current syllabus, please check with your school or the School Curriculum and Standards Authority.*

Unit 1 – Chemical fundamentals: structure, properties and reactions

SCIENCE AS A HUMAN ENDEAVOUR

Properties and structure of atoms

Findings from a range of scientific experiments contributed to the understanding of the atom, enabling scientists, including Dalton, Thomson, Rutherford, Bohr and Chadwick to develop models of atomic structure and make reliable predictions about the mass, charge and location of the sub-atomic particles.

Science Understanding

- elements are represented by symbols
- 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
- the ability of atoms to form chemical bonds can be explained by the arrangement of electrons in the atom and in particular by the stability of the valence electron shell
- the structure of the periodic table is based on the atomic number and the properties of the elements
- the elements of the periodic table show trends across periods and down main groups, including in atomic radii, valencies, 1st ionisation energy and electronegativity as exemplified by groups 1, 2, 13–18 and period 3
- flame tests and atomic absorption spectroscopy (AAS) are analytical techniques that can be used to identify elements; these methods rely on electron transfer between atomic energy levels and are shown by line spectra
- isotopes are atoms of an element with the same number of protons but different numbers of neutrons and are represented in the form ^{A}X (IUPAC) or X-A
- isotopes of an element have the same electron configuration and possess similar chemical properties but have different physical properties
- the relative atomic mass (atomic weight), A_r is the ratio of the average mass of the atom to 1/12 the mass of an atom of ^{12}C; relative atomic masses of the elements are calculated from their isotopic composition
- mass spectrometry involves the ionisation of substances and the separation and detection of the resulting ions; the spectra which are generated can be analysed to determine the isotopic composition of elements and interpreted to determine relative atomic mass.

SCIENCE AS A HUMAN ENDEAVOUR

Properties and structure of materials

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.

Science Understanding

- materials are pure substances with distinct measurable properties, including melting and boiling points, reactivity, hardness and density; or mixtures with properties dependent on the identity and relative amounts of the substances that make up the mixture
- pure substances may be elements or compounds which consist of atoms of two or more elements chemically combined; the formulae of compounds indicate the relative numbers of atoms of each element in the compound
- nanomaterials are substances that contain particles in the size range 1–100 nm and have specific properties relating to the size of these particles which may differ from those of the bulk material
- differences in the physical properties of substances in a mixture, including particle size, solubility, density, and boiling point, can be used to separate them
- the type of bonding within ionic, metallic and covalent substances explains their physical properties, including melting and boiling points, conductivity of both electricity and heat and hardness
- chemical bonds are caused by electrostatic attractions that arise because of the sharing or transfer of electrons between participating atoms; the valency is a measure of the bonding capacity of an atom
- ions are atoms or groups of atoms that are electrically charged due to a loss or gain of electrons; ions are represented by formulae which include the number of constituent atoms and the charge of the ion (for example, O^{2-}, SO_4^{2-})
- ionic bonding can be modelled as a regular arrangement of positively and negatively charged ions in a crystalline lattice with electrostatic forces of attraction between oppositely charged ions
- the ionic bonding model can be used to explain the properties of ionic compounds, including high melting point, brittleness and non-conductivity in the solid state; the ability of ionic compounds to conduct electricity when molten or in aqueous solution can be explained by the breaking of the bonds in the lattice to give mobile ions
- the formulae of ionic compounds can be determined from the charges on the relevant ions (refer to Appendix 2). **See Table 3.1 chapter 3.*
- metallic bonding can be modelled as a regular arrangement of atoms with electrostatic forces of attraction between the nuclei of these atoms and their delocalised electrons that are able to move within the three dimensional lattice
- the metallic bonding model can be used to explain the properties of metals, including malleability, thermal conductivity, generally high melting point and electrical conductivity; covalent bonding can be modelled as the sharing of pairs of electrons resulting in electrostatic forces of attraction between the shared electrons and the nuclei of adjacent atoms
- the properties of covalent network substances, including high melting point, hardness and electrical conductivity, are explained by modelling covalent networks as three-dimensional structures that comprise covalently bonded atoms
- elemental carbon exists as a range of allotropes, including graphite, diamond and fullerenes, with significantly different structures and physical properties
- the properties of covalent molecular substances, including low melting point, can be explained by their structure and the weak intermolecular forces between molecules; their

non-conductivity in the solid and liquid/molten states can be explained by the absence of mobile charged particles in their molecular structure

- molecular formulae represent the number and type of atoms present in the molecules (refer to Appendix 2) **See chapter 3 (3.16)*
- percentage composition of a compound can be calculated from the relative atomic masses of the elements in the compound and the formula of the compound
- hydrocarbons, including alkanes, alkenes and benzene, have different chemical properties that are determined by the nature of the bonding within the molecules
- molecular structural formulae (condensed or showing bonds) can be used to show the arrangement of atoms and bonding in covalent molecular substances
- IUPAC nomenclature is used to name straight and simple branched alkanes and alkenes from C_1- C_8
- alkanes, alkenes and benzene undergo characteristic reactions such as combustion, addition reactions for alkenes and substitution reactions for alkanes and benzene.

SCIENCE AS A HUMAN ENDEAVOUR

Chemical reactions: reactants, products and energy change

- There are differences in the energy output and carbon emissions of fossil fuels (including coal, oil, petroleum and natural gas) and biofuels (including biogas, biodiesel and bioethanol). These differences, together with social, economic, cultural and political values, determine how widely these fuels are used.

Science Understanding

- chemical reactions can be represented by chemical equations; balanced chemical equations indicate the relative numbers of particles (atoms, molecules or ions) that are involved in the reaction
- chemical reactions and phase changes involve enthalpy changes, commonly observable as changes in the temperature of the surroundings and/or the emission of light
- endothermic and exothermic reactions can be explained in terms of the Law of Conservation of Energy and the breaking of existing bonds and forming of new bonds; heat energy released or absorbed by the system to or from the surroundings, can be represented in thermochemical equations
- 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 mole is a precisely defined quantity of matter equal to Avogadro's number of particles
- the mole concept relates mass, moles and molar mass and, with the Law of Conservation of Mass, can be used to calculate the masses of reactants and products in a chemical reaction.

Unit 2 – Molecular interactions and reactions

SCIENCE AS A HUMAN ENDEAVOUR

Intermolecular forces and gases

Chromatographic techniques, including thin layer chromatography (TLC), gas chromatography (GC), and high performance liquid chromatography (HPLC), are used to determine the components of a wide range of mixtures in various settings. The decision to use a particular chromatographic technique depends on a number of factors, including the properties of the substances being separated, the amount of substance available for analysis and the sensitivity of the equipment. Chromatographic techniques have a wide range of analytical and forensic applications, including monitoring air and water pollutants, drug testing of urine and blood samples, and testing for food additives and quality.

Science Understanding

- 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 valence shell electron pair repulsion (VSEPR) theory and Lewis structure diagrams can be used to explain, predict and draw the shapes of molecules
- the polarity of molecules can be explained and predicted using knowledge of molecular shape, understanding of symmetry, and comparison of the electronegativity of atoms involved in the bond formation
- the shape and polarity of molecules can be used to explain and predict the nature and strength of intermolecular forces, including dispersion forces, dipole-dipole forces and hydrogen bonding
- data from chromatography techniques, including thin layer chromatography (TLC), gas chromatography (GC), and high-performance liquid chromatography (HPLC), can be used to determine the composition and purity of substances; the separation of the components is caused by the variation in strength of the interactions between atoms, molecules or ions in the mobile and stationary phases
- the behaviour of an ideal gas, including the qualitative relationships between pressure, temperature and volume, can be explained using the Kinetic Theory
- the mole concept can be used to calculate the mass of substances and volume of gases (at standard temperature and pressure) involved in a chemical reaction.

SCIENCE AS A HUMAN ENDEAVOUR

Aqueous solutions and acidity

The supply of potable drinking water is an extremely important issue for both Australia and countries in the Asian region. Water sourced from groundwater and seawater undergoes a number of purification and treatment processes (such as desalination, chlorination, fluoridation) before it is delivered into the supply system. Chemists regularly monitor drinking water quality to ensure that it meets the regulations for safe levels of solutes. Heavy metal contamination in ground water is monitored to ensure that concentrations are at acceptable levels. Several methods can be used to reduce heavy metal contamination; the method used is influenced by economic and social factors.

Science Understanding

- the unique physical properties of water, including melting point, boiling point, density in solid and liquid phases and surface tension, can be explained by its molecular shape and hydrogen bonding between molecules
- solutions can be classified as saturated, unsaturated or supersaturated; the concentration of a solution is defined as the quantity of solute dissolved in a quantity of solution; this can be

represented in a variety of ways, including by the number of moles of the solute per litre of solution (mol L^{-1}) and the mass of the solute per litre of solution (g L^{-1}) or parts per million (ppm)

- the presence of specific ions in solutions can be identified by observing the colour of the solution, flame tests and observing various chemical reactions, including precipitation and acid-base reactions
- the solubility of substances in water, including ionic and polar and non-polar molecular substances, can be explained by the intermolecular forces, including ion-dipole interactions between species in the substances and water molecules, and is affected by changes in temperature
- the Arrhenius model can be used to explain the behaviour of strong and weak acids and bases in aqueous solutions
- indicator colour and the pH scale are used to classify aqueous solutions as acidic, basic or neutral
- pH is used as a measure of the acidity of solutions and is dependent on the concentration of hydrogen ions in the solution
- 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, allow products and observations to be predicted from reactants; ionic equations represent the reacting species and products in these reactions
- the mole concept can be used to calculate the mass of solute, and solution concentrations and volumes involved in a chemical reaction.

SCIENCE AS A HUMAN ENDEAVOUR

Rates of chemical reactions

Catalysts are used in many industrial processes in order to increase the rates of reactions that would otherwise be uneconomically slow. Catalysts are also used to reduce the emission of pollutants produced by car engines. Motor vehicles have catalytic converters which are used to catalyse reactions that reduce the amount of carbon monoxide, unburnt petrol and nitrogen oxides that are emitted.

Science Understanding

- varying the conditions under which chemical reactions occur can affect the rate of the reaction
- the rate of chemical reactions can be quantified by measuring the rate of formation of products or the depletion of reactants
- collision theory can be used to explain and predict the effects of concentration, temperature, pressure, the presence of catalysts and surface area on the rate of chemical reactions
- the activation energy is the minimum energy required for a chemical reaction to occur and is related to the strength and number of the existing chemical bonds; the magnitude of the activation energy influences the rate of a chemical reaction
- energy profile diagrams, which can include the transition state and catalysed and uncatalysed pathways, can be used to represent the enthalpy changes and activation energy associated with a chemical reaction
- catalysts, including enzymes and metal nanoparticles, affect the rate of certain reactions by providing an alternative reaction pathway with a reduced activation energy, hence increasing the proportion of collisions that lead to a chemical change.

STUDY HINTS AND EXAM TECHNIQUES

Study Strategies

The earlier you start preparing for tests and exams, the more effective your revision will be. To get maximum benefit from your study program you need to establish a plan of attack. In brief, you need to be sure of where, how and what you will be doing to maximise your chances at test and examination time.

You will need to:

- Set up your ideal study area: Uninterrupted privacy and an organised study area are essential.
- Identify your needs: Check your areas of weakness and give them priority.
- Develop a study plan: Study timetables and setting of timelines are essential.
- Stick to the plan: A disciplined study routine will ensure your success.
- Monitor your progress: If your study plan is not working changes may be necessary.
- Your study area: The effectiveness of your study can be greatly improved by creating the right environment.

Your ideal study area should:

- be an area free from interruptions and distractions – preferably a room well away from the family room in order to minimise "visitors" or "passers through"
- be out of hearing range of the television
- have good ventilation so that you don't become drowsy
- have a good sized desk for your study needs
- be well lit so that when you sit at the desk:
  - the main light will shine from behind;
  - a second light, but not as bright, shines across the desk.
  - this second light reduces the shadows which would otherwise tire your eyes quickly.
  - have your study timetable well in view.

To be avoided in your study area: easy chair, stereo, television, phone.

Planning

Identify your needs

- list the areas that need revision for your test (or examination). Use your syllabus or school program to help you.
- prioritise areas of weakness.

Develop a timeline

- check weeks remaining to your test or Semester Examination.
- allocate time available (each week) for revision of Chemistry. Remember that you will have assignments to complete as well as study to do.
- place a sequence of areas to revise on a timeline. It is probably best to follow the sequence of your school's Chemistry program.
- allocate more time to areas in which you are weak.

Following the plan

Force yourself to stick to the plan. If time is short at any stage – prioritise in terms of what is going to be of most benefit to your long term plans.

Have a checklist (e.g. your course outline):

- tick areas you have revised, re-tick each time revised
- get a copy of the syllabus for Chemistry and tick off the course objectives as you revise them. If aiming for excellence you must know ALL areas of the course.

Study notes:

- develop a set of single page, A4, notes. One page per area revised. Try to visualise your pages of study notes – key reminder pictures or words are important tools to trigger your memory.
- you need to have set aside a section of your Chemistry file to store these revision notes. (Some students use small palm cards or record cards to store notes on small sections of the course – these are an excellent alternative.)

Using this study guide:

- read through the section appropriate to what you are revising and **highlight** all important points
- attempt **all** the questions to check your understanding
- trial tests – attempt these prior to a test or examination. To get maximum value from these tests:
  – prepare first
  – set a time
  – observe test conditions.

Monitoring your progress:
The trial tests at the back of this study guide are designed to give you immediate feedback on how your revision is proceeding and how effective it has been. If your study plan is not working, then alter it so that you can achieve your desired results.

EXAMINATION STRATEGIES

Before the examination

The evening prior to an exam should be set aside for light study and a brief review of your revision notes. If you have been provided with a copy of the examination instructions to candidates by your teacher, then it would be a good idea to go over these in detail. Avoid staying up late, as a good night's sleep will prove to be of great value to you the following day.

On the morning of the exam check that you have all that you need; examination number slip, calculator/s fitting the SCSA approved guidelines – preferably with new batteries, pens, soft pencil, eraser, and ruler. Your own watch may also be more convenient for you to use during the examination rather than the clock in the exam room.

Arrive at school with at least 20 minutes to spare. Try to relax by talking to your friends but avoid worrying about the exam, or what you may have forgotten. You will find that work you have learnt will be readily recalled during the exam.

Reading time

This ten minutes of valuable time must be used wisely. It will be natural curiosity to have a quick look at the whole paper but this will be of limited value. First, and most important, you must read the instructions and make sure you understand them. It is then best to read steadily through the questions, beginning with the short and extended answer section of the paper.

It is best to leave the multiple-choice items at this stage since you will not be able to write down an answer even though you may have read through quite a lot of detail (remember - no writing allowed during reading time). This means you will only have to *re-read* through them during the allotted 3 hours.

It is better, then to read through the questions from the beginning and to establish in your own mind the **concepts** involved and how you may tackle the question.

Working through the paper

It is best to work through the paper in the order that it is written. However this is not essential and some students prefer starting with a section they are most confident about.

Pace yourself

It is important to know how much time you have for your exam and allocate it to each section so that you can pace yourself evenly. Suggested times for each section are usually shown on the paper. Your end of year Chemistry exam is likely to have the following sections (check with your teacher).

Section One	Multiple-choice
Section Two	Short Answer
Section Three	Extended Answer

Check the value in marks of all questions and relate it to a time value. For example if your examination paper is a 3 hour paper and there are a total of 180 marks then we have:

180 minutes = 180 marks   1.0 minute = 1 mark

Hence if there are 25 multiple-choice questions worth 50 marks you should not take more than about 50 minutes for this section. A short answer question which is worth say 10 marks should take about 10 minutes. By keeping reasonably close to the time value of a question you will avoid running out of time.

Multiple-choice

Attempt each question as you come to it: Be sure to select an answer, even if unsure – there is no penalty for guessing. It is **wasteful** of your time to read through a question and then leave it because you are not quite sure of the answer. When you come back to it later you will only have to go through it again. If unsure, eliminate the obvious distractors and make an educated guess. Put a mark by the side of the questions so that if you do have time you can come back to reconsider it. If you do run short of time, however, you will at least have a good chance of being correct.

Read each question through carefully: Take note of any key words. Where possible, it is best to think of your answer to a question before you proceed to read through each of the alternatives given. Then look through each of the distractors to see if your answer is there. Don't immediately write this down as your answer but convince yourself that each of the other distractors is incorrect.

As you work through each distractor place a cross (x) next to those that you are sure are incorrect, a question mark (?) next to those you are not quite sure about and circle your answer when you have it. This will save time, particularly if it is a difficult question and you have to come back to it.

Changing your selections: Do this only if you are sure that your initial choice was incorrect. Usually the first choice is the best choice.

Relate to study notes: If you have difficulty with a question there are points of reference you should try to think about to work towards the best possible answer:

(i) Try to identify the syllabus concepts that the question and distractors relate to.
(ii) Try to relate the question to parts of your study notes.

A question can often become more clear when you are able to classify it according to your prior knowledge (i.e. study notes).

Short and extended answer questions

Complete all that you can—come back to others as ideas come flooding back (hopefully). When stuck on any particular question, try to associate the question with a page of your study notes —can these notes be related to the question in any way? Where appropriate chemical equations and diagrams are very useful for demonstrating your chemical literacy.

Calculations

Show all your working for top marks. Be sure to state your answer to the required number of significant figures and in scientific notation.

For stoichiometric calculations the balanced equation is an essential part of the solution to a question. If the equation is not given look for clues as to what the reactants and products might be. Classify the information as to what type of reaction it might be (e.g. *acid* + *base*, *redox*, etc). It is important that you try to produce a balanced equation. It is better to have attempted to write an equation and to answer the remainder of the question than to completely disregard the question.

Before attempting a calculation note down information, produce a **simple flow diagram** if complicated information has been presented. This will help to produce a standard of working clarity that is much more logical to follow (for you and for the person marking **your** exam).

Clearly indicate the **steps or stages** to your calculation (e.g. $n(Na_2CO_3) = 2.5$). Clearly indicate which section of a multiple part question you are answering. If the question has 5 parts then your working must be broken up into 5 clearly labelled parts. **You should try to make the marker's task as easy as possible**, make your setting out clear and logical, and clearly highlight (or underline) your answer.

MEASUREMENT AND QUANTITIES

SI Fundamental (Base) Units

QUANTITY	SYMBOL FOR QUANTITY	UNIT	SYMBOL FOR UNIT
length	l	metre	m
mass	m	kilogram	kg
time	t	second	s
electric current	I	ampere	A
temperature	T	kelvin	K
amount of substance	n	mole	mol

Commonly used units in Chemistry

QUANTITY	COMMONLY USED UNITS AND SYMBOLS		CONVERSIONS (ETC)
amount of substance	mole	mol	1 mole = 6.022×10^{23} particles
concentration (solutions)	mole per litre gram per litre	mol L^{-1} g L^{-1}	mol L^{-1}
electric charge	coulomb	C	
electric current	ampere	A	
electromotive force	volt	V	
energy	joule kilojoule	J kJ	1000 J = 1 kJ
density	gram per litre	g L^{-1}	
mass	gram kilogram tonne	g kg	1 tonne = 1000 kg = 1×10^6 g
molar charge	faraday	F	1 F = 9.649×10^4 C
molar volume	litre per mole	L mol^{-1}	22.47 L (S.T.P. – 0.0°C, 100.0 kPa) 24.47 L (25.0°C, 101.3 kPa)
pressure	kilopascal atmosphere millimetres of Hg	kPa atm mm Hg	1 atm = 101.3 kPa = 760.0 mm Hg
temperature	kelvin degree celsius	K °C	K = °C + 273.15
time	second hour	s h	1 h = 3600 s
volume	litre millilitre cubic centimetre cubic decimetre	L mL cm^3 dm^3	1 L = 1000 mL = 1 dm^3 = 1000 cm^3

Commonly used prefixes

PREFIX	SYMBOL	FACTOR	PREFIX	SYMBOL	FACTOR
tera	T	10^{12}	milli	m	10^{-3}
giga	G	10^{9}	micro	μ	10^{-6}
mega	M	10^{6}	nano	n	10^{-9}
kilo	k	10^{3}	pico	p	10^{-12}

Some Physical Constants used in Chemistry

CONSTANT	SYMBOL	VALUE
Avogadro constant	N	6.022×10^{23} mol^{-1}
electronic charge	q_e	1.602×10^{-19} C
Faraday constant	F	9.649×10^{4} C
molar volume (ideal gas)	V_m	22.47 L (S.T.P.)* 24.47 L (25.00°C and 101.3 kPa)
universal gas constant	R or	8.314 J K^{-1} mol^{-1} 0.08206 L atm K^{-1} mol^{-1}

* *S.T.P. (standard temperature and pressure) = 0.0°C, 100.0 kPa.*

Scientific notation

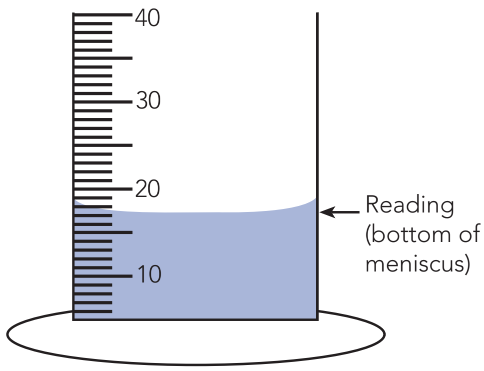
Scientific (or exponential) notation is used to conveniently express very large or very small numbers.Numbers are written so that there is a single non zero digit to the left of the decimal point times the appropriate power of 10.

e.g.

6275	=	6.275×10^{3}	
0.06275	=	6.275×10^{-2}	
0.00014	=	1.4×10^{-4}	
0.000140	=	1.40×10^{-4}	
2500	=	2.5×10^{3}	(if 2 sig. figures)
2500	=	2.500×10^{3}	(if 4 sig. figures)

Significant figures

Significant figures are used to indicate the precision or accuracy of a particular measurement or given data. They are not to be confused with decimal points.



Recording significant figures

When a measurement is made we record all integers of which we are certain plus one other of which there is some uncertainty, e.g. a graduated cylinder is used to determine the volume of liquid as shown at right. The reading indicates a volume somewhere between 17 mL and 18 mL. Where exactly between these points is uncertain. Hence the reading (say 17.4 mL) is recorded to only 3 significant figures.

2 integers known with certainty

uncertain integer

$$\underbrace{\overbrace{17}^{}.4}_{\text{3 significant figures}} \text{ mL}$$

Determining significant figures

In any given data (number) a significant figure is

- any integer
- any zero that follows an integer after a decimal point

e.g.

1.46 , 0.106 , 1.10	all have 3 significant figures
2.100×10^{3} , 1.467 , 0.02600	all have 4 significant figures
8.1 , 0.00026 , 26	all have 2 significant figures

Where a zero occurs at the end of whole number such as 250 and 5000, the correct number of significant figures is uncertain. Use of scientific notation avoids this.

e.g. 250 is written as 2.50×10^2 if it has 3 significant figures

250 is written as 2.5×10^2 if it has 2 significant figures

Using significant figures

When doing calculations, remember that your final value cannot be more precise than your least precise data. During the actual calculation leave all the digits in your calculator. However, your final result should be rounded off to have the same number of significant figures as your least significant data. In nearly all cases 3 significant figures are used.

Rounding off

Numbers often need to be rounded to the correct number of significant figures. To do this:

- note which are the significant digits
- look at the next digit to the right of them, and

if it is > 5	round up
if it is $= 5$	round up if an even number results
if it is < 5	no change

e.g. The following have all been rounded to 3 significant figures.

1.294	rounds to	1.29
0.06437	rounds to	0.0644
21.65	rounds to	21.6
21.75	rounds to	21.8
3.955	rounds to	3.96

Exact numbers

Some numbers used in Chemistry calculations have an exact value and do not affect the number of significant figures in the answer. These are such numbers as:

- the coefficients in chemical equations
  e.g. 2HCl means 2.000 moles of HCl

- conversion ratios
  e.g. 1 kg = 1000.00 g

FORMULAE, IONS, COLOURS, SOLUBILITY

Symbols and names of ions (see table 3.1, page 49).

Formulae of molecular substances

The formulae and names of some important molecular substances

ELEMENTS		COMPOUNDS			
bromine	Br_2	ammonia	NH_3	hypochlorous acid	$HClO$
chlorine	Cl_2	carbon dioxide	CO_2	nitric acid	HNO_3
fluorine	F_2	carbon monoxide	CO	nitrogen dioxide	NO_2
hydrogen	H_2	dinitrogen monoxide (nitrous oxide)	N_2O	nitrogen monoxide (nitric oxide)	NO
iodine	I_2	dinitrogen tetroxide	N_2O_4	phosphoric acid	H_3PO_4
nitrogen	N_2	hydrogen bromide	HBr	sulfur dioxide	SO_2
oxygen	O_2	hydrogen chloride	HCl	sulfur trioxide	SO_3
phosphorus	P_4	hydrogen fluoride	HF	sulfuric acid	H_2SO_4
sulfur	S_8	hydrogen iodide	HI	sulfurous acid	H_2SO_3
		hydrogen peroxide	H_2O_2	water	H_2O
		hydrogen sulfide	H_2S		

Colour of species in aqueous solutions

Students should be able to use the colours of the following ions to infer and describe the products of reactions.

CATION	COLOUR	CATION	COLOUR	ANION	COLOUR	HALOGEN	COLOUR
$Al^{3+}{}_{(aq)}$	colourless	$Mg^{2+}{}_{(aq)}$	colourless	$Br^{-}{}_{(aq)}$	colourless	$Cl_{2(aq)}$	pale yellow
$NH_4^{+}{}_{(aq)}$	colourless	$Mn^{2+}{}_{(aq)}$	colourless*	$Cl^{-}{}_{(aq)}$	colourless	$Br_{2(aq)}$	orange
$Ba^{2+}{}_{(aq)}$	colourless	$Ni^{2+}{}_{(aq)}$	green	$CrO_4^{2-}{}_{(aq)}$	yellow	$I_{2(aq)}$	brown
$Ca^{2+}{}_{(aq)}$	colourless	$K^{+}{}_{(aq)}$	colourless	$Cr_2O_7^{2-}{}_{(aq)}$	orange		
$Cr^{3+}{}_{(aq)}$	deep green	$Ag^{+}{}_{(aq)}$	colourless	$I^{-}{}_{(aq)}$	colourless	**HALOGEN IN ORGANIC SOLVENT**	
$Co^{2+}{}_{(aq)}$	pink	$Na^{+}{}_{(aq)}$	colourless	$MnO_4^{-}{}_{(aq)}$	deep purple	**Halogen**	**Colour**
$Cu^{2+}{}_{(aq)}$	blue	$Sr^{2+}{}_{(aq)}$	colourless	$PO_4^{3-}{}_{(aq)}$	colourless	Br_2	red
$Fe^{2+}{}_{(aq)}$	pale green	$Sn^{2+}{}_{(aq)}$	colourless	$S^{2-}{}_{(aq)}$	colourless	I_2	purple
$Fe^{3+}{}_{(aq)}$	brown	$Zn^{2+}{}_{(aq)}$	colourless			Cl_2	very pale green
$Pb^{2+}{}_{(aq)}$	colourless						

* very pale pink if saturated solution

SOLUBILITY RULES

Students should be able to apply the following solubility rules for ionic solids in water.

Soluble in water

SOLUBLE	EXCEPTIONS	
	Insoluble	**Slightly soluble**
Most chlorides	$AgCl$	$PbCl_2$
Most bromides	$AgBr$	$PbBr_2$
Most iodides	AgI, PbI_2	
All nitrates	No exceptions	
All ethanoates	No exceptions	
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, $Ba(OH)_2$ (NH_4OH and $AgOH$ do not exist)	$Ca(OH)_2$, $Sr(OH)_2$
Most carbonates	Na_2CO_3, K_2CO_3, $(NH_4)_2CO_3$	
Most phosphates	Na_3PO_4, K_3PO_4, $(NH_4)_3PO_4$	
All 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

Colour of precipitates

The colour of the following precipitates are listed for reference only.

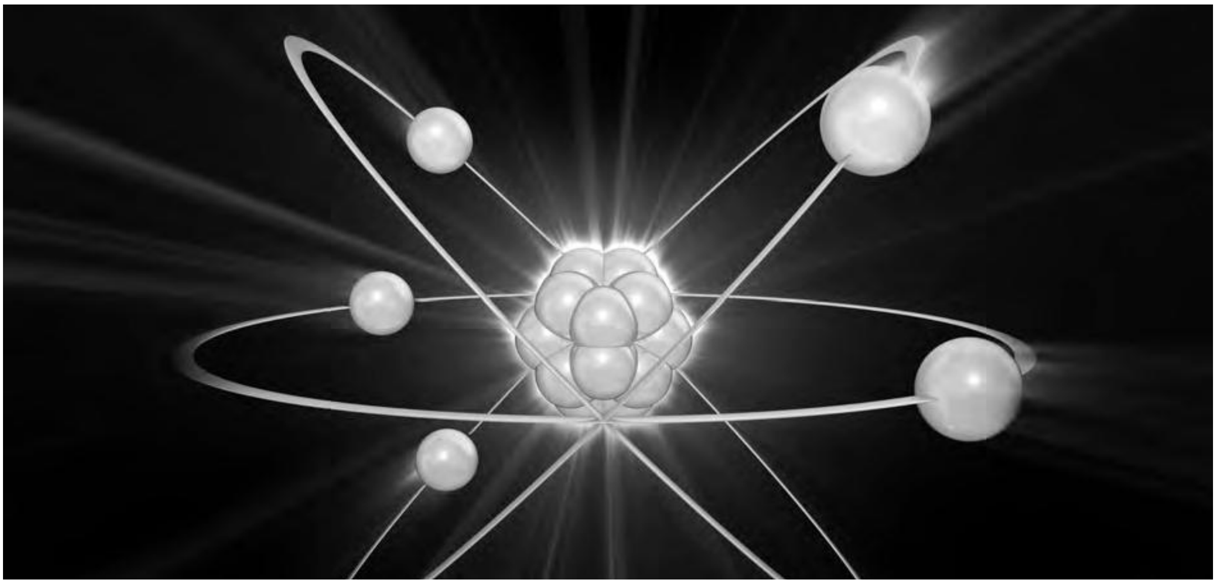
	Cl^-	SO_4^{2-}	OH^-	CO_3^{2-}		Cl^-	SO_4^{2-}	OH^-	CO_3^{2-}
Mg^{2+}	-	-	white	white	Cu^{2+}	-	-	pale blue	green
Ca^{2+}	-	white	white	white	Ag^+	white	white	(Ag_2O) brown ($AgOH$ does not exist)	white
Sr^{2+}	-	white	white	white	Zn^{2+}	-	-	white	white
Ba^{2+}	-	white	-	white	Cd^{2+}	-	-	white	white
Cr^{3+}	-	-	green	- ($CrCO_3$) green	Hg_2^{2+}	white	white/ yellow	(Hg_2O) black	yellow/ brown
Mn^{2+}	-	-	pink	pink	Hg^{2+}	-	-	(HgO) red/yellow	-
Fe^{2+}	-	-	green	grey/green	Al^{3+}	-	-	white	white
Fe^{3+}	-	-	brown	brown	Sn^{2+}	-	-	white	white
Co^{2+}	-	-	red/pink	red	Pb^{2+}	white	white	white	white
Ni^{2+}	-	-	green	green		Na^+ , K^+ , NH^{4+} do not form precipitates.			

CHEMISTRY

UNIT 1

ATOMIC STRUCTURE

1



Topics covered in this chapter:

1.1 STRUCTURE OF THE ATOM

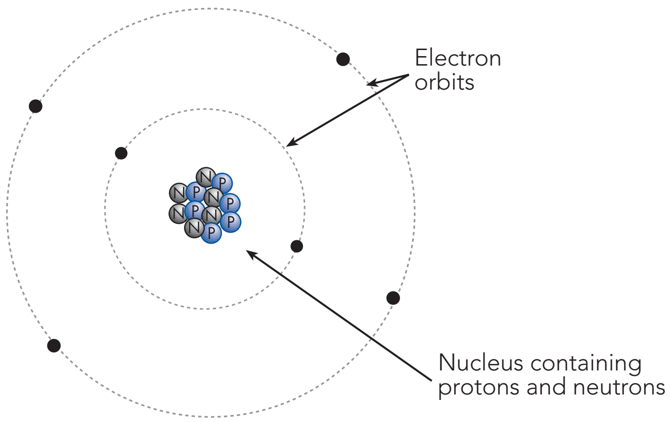


Figure 1.1 *Rutherford-Bohr model of the atom. The atom's mass is concentrated in the tiny nucleus while the electron cloud occupies most of the volume.*

In 1804, John Dalton first proposed the idea that tiny particles called atoms were the fundamental particles of nature. His atomic theory helped explain the experimental data available at that time and laid the foundations to our modern view of matter.

Part of Dalton's theory was that atoms of elements were solid and indivisible. However, work carried out by many other scientists such as Faraday, Thompson, Rutherford and Bohr established that in fact, atoms consist of protons, neutrons and electrons. Their discoveries led to a nuclear model of the atom and the following ideas:

- Atoms consist basically of two regions, that is, a small dense nucleus surrounded by a cloud of electrons.
- The nucleus is positively charged and contains protons and neutrons. The very large majority of the mass of an atom is contained in the nucleus.
- The electrons are negatively charged and have a very small mass. They move rapidly in the region of space around the nucleus creating the effect of an electron cloud. This electron cloud makes up nearly all the volume of the atom.
- Atoms are electrically neutral. Hence the number of protons in any atom is equal to the number of electrons.

Table 1.1 *Properties of protons, neutrons and electrons.*

PARTICLE	LOCATION	MASS (kg)	RELATIVE MASS	RELATIVE CHARGE
proton	nucleus	1.673×10^{-27}	1	+1
neutron	nucleus	1.675×10^{-27}	1	0
electron	electron cloud	9.11×10^{-31}	1/1836	-1

1.2 ATOMIC NUMBER, MASS NUMBER

Atoms can differ in atomic number and mass number. The atomic and mass numbers can be shown with the symbol of the element as follows.

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

X = symbol of the element
A = mass number
Z = atomic number

e.g. ${}^{7}_{3}Li$, ${}^{12}_{6}C$, ${}^{35}_{17}Cl$

- The **atomic number** (Z) of an atom is the number of protons in the nucleus. All atoms of the same element have the same atomic number. For neutral atoms this is also equal to the number of electrons.
- The **mass number** (A) is the total number of protons and neutrons.

Note also A = Z + N where N = No. of neutrons

Worked Example

1.1 Determine the number and type of particles in a neutral atom of ${}^{23}_{11}Na$:

Number of protons = Z = 11

Number of neutrons = (A - Z) = 23 - 11 = 12

Number of electrons = 11 (same as number of protons for a neutral atom).

1.3 ISOTOPES

All the atoms of a given element have the same number of protons but the number of neutrons may vary. Hence different forms of an element may have a different mass number (A). These different forms of an element are called **isotopes**.

Isotopes of a particular element are chemically similar since they have the same number of electrons. This makes them difficult to separate as they only differ slightly in mass and density.

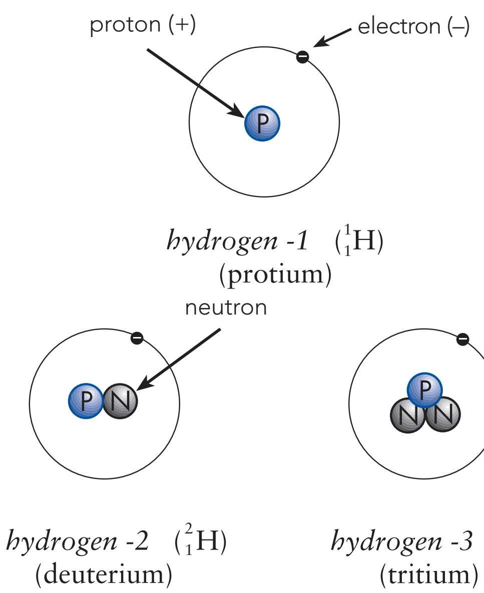


Figure 1.2 *The three isotopes of hydrogen.*

1.4 THE MASS SPECTROMETER

A mass spectrometer can be used to measure the masses and relative concentration of atoms and molecules. It can used to determine the isotopic masses of naturally occurring elements and their relative abundance. This data can also be used to determine the relative atomic mass of that element.

A mass spectrometer consists of several components but essentially it uses a magnetic field in a vacuum to deflect the path of fast moving charged particles. The substance to be analysed is initially vaporised and then ionised by fast moving electrons. The resulting positive ions are then accelerated by an electric field, pass through a velocity selector and then enter the mass spectrometer.

The magnetic field then causes the ions to be deflected and separate depending on their mass and charge. Any difference in charge would give distinctly different results compared to small changes in mass and hence is easily accounted for. Effectively, the mass/charge ratio and relative concentration are recorded on the output chart.

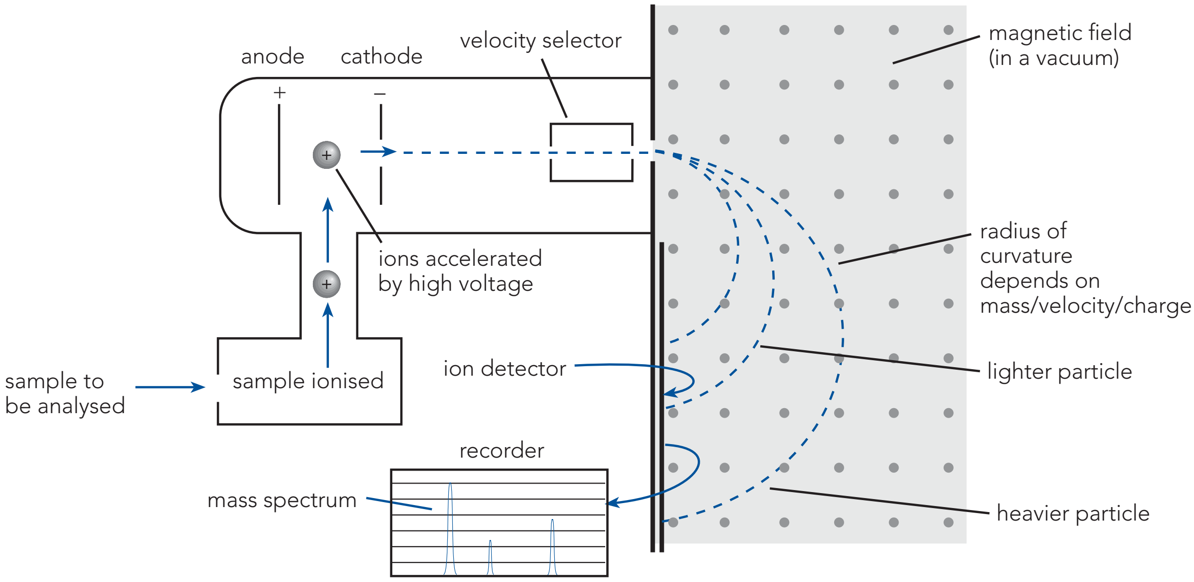


__Figure 1.3__ Simplified view of the mass spectrometer. The sample to be analysed is firstly ionised and then its ions are accelerated. A velocity selector is used if just identifying masses. The mass spectrometer is essentially a strong magnetic field in a vacuum with suitable detectors. The mass spectrum gives distinctive peaks for each isotope.

Mass spectra and relative atomic mass

Typically, a mass spectrometer can be used to measure atomic masses, identify elements and generally help analyse small traces of unknown substances. Importantly, it can be used to determine the relative atomic mass of pure elements by identifying all of its isotopes and their relative abundance.

The mass spectrum of an element shows characteristic peaks for each of its isotopes. The position and height of these peaks indicates the relative mass and relative abundance of each isotope. To determine the percentage abundance of each isotope the peaks are carefully measured and their heights totalled. The height of each peak can then be considered as a fraction, or percentage, of the total height.

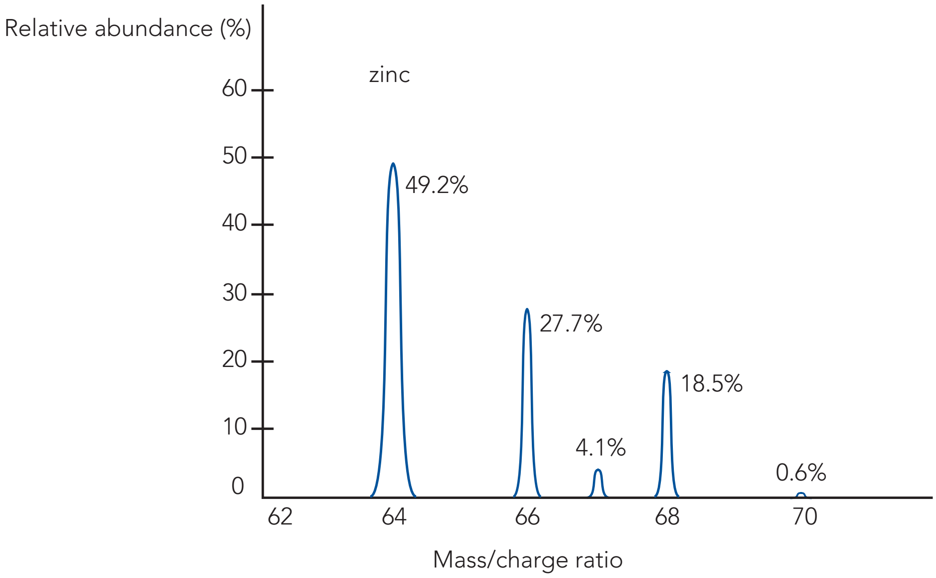


***Figure 1.4** The mass spectrum of Zinc showing the relative abundance of each of its five isotopes. The isotope ^{64}Zn is the most abundant (49.2%) and its relative isotopic mass being 63.93. The relative atomic mass (atomic weight) of the element zinc can be calculated from the data in this spectrum. See worked example below.*

Worked Example

1.2 Data determined from the mass spectrum for zinc is shown below. Use this to calculate the relative atomic mass (Ar) for zinc.

ISOTOPE	ISOTOPIC MASS	ABUNDANCE (%)
^{64}Zn	63.93	49.2
^{66}Zn	65.93	27.7
^{67}Zn	66.93	4.0
^{68}Zn	67.93	18.5
^{70}Zn	69.93	0.6

Ar (Zn) = $\sum$ [isotopic mass x abundance %]/100

= $\sum [(63.93 \times 49.2) + (65.93 \times 27.7) + (66.93 \times 4.0) + (67.93 \times 18.5) + (69.93 \times 0.6)]/100$

= $31.45 + 18.26 + 2.68 + 12.57 + 0.42$

= 65.38

1.3 The two isotopes of Lithium, 6_{Li} and 7_{Li}, have relative isotopic masses of 6.015 and 7.016 respectively. Their abundance in nature is 7.59% and 92.41%. Calculate the relative atomic mass (Ar) for Lithium

Ar (Li) = $\sum$ [isotopic mass × abundance %]/100

= $\sum [(6.015 \times 7.59) + (7.016 \times 92.41)]/100$

= $0.457 + 6.48$

= 6.94

Table 1.2 Isotopes of some different elements.

NAME OF ISOTOPE	SYMBOL	ABUNDANCE IN NATURE %	ATOMIC NUMBER (Z)	MASS NUMBER (A)	NUMBER OF PROTONS	NUMBER OF NEUTRONS	NUMBER OF ELECTRONS IN NEUTRAL ATOM
hydrogen – 1	${}^{1}_{1}H$	99.9885	1	1	1	0	1
hydrogen – 2	${}^{2}_{1}H$	0.0115	1	2	1	1	1
lithium – 6	${}^{6}_{3}Li$	7.59	3	6	3	3	3
lithium – 7	${}^{7}_{3}Li$	92.41	3	7	3	4	3
aluminium – 27	${}^{27}_{13}Al$	100.0	13	27	13	14	13
chlorine – 35	${}^{35}_{17}Cl$	75.76	17	35	17	18	17
chlorine – 37	${}^{37}_{17}Cl$	24.24	17	37	17	20	17
uranium – 234	${}^{234}_{92}U$	0.006	92	234	92	142	92
uranium – 235	${}^{235}_{92}U$	0.720	92	235	92	143	92
uranium – 238	${}^{238}_{92}U$	92.274	92	238	92	146	92

Question 1.1

Indicate the number of protons, neutrons and electrons for the following neutral atoms:

(a) ${}^{7}_{3}Li$ __________ protons, __________ neutrons, __________ electrons.

(b) ${}^{14}_{7}N$ __________ protons, __________ neutrons, __________ electrons.

(c) ${}^{19}_{9}F$ __________ protons, __________ neutrons, __________ electrons.

(d) ${}^{35}_{17}Cl$ __________ protons, __________ neutrons, __________ electrons.

Question 1.2

Complete the following table:

ISOTOPE	NAME OF ISOTOPE	Z ATOMIC NO.	A MASS NO.	NUMBER OF PROTONS	NUMBER OF NEUTRONS
${}^{12}_{6}C$	Carbon – 12	6	12	6	6
${}^{14}_{6}C$	Carbon – 14				
${}^{24}_{12}Mg$					
	Argon – 40	18	40		
${}^{27}_{13}Al$					
	Cobalt – 59	27			

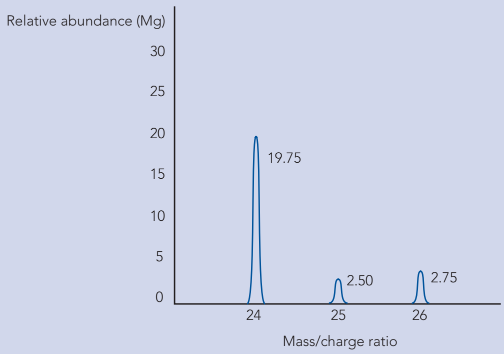
Question 1.3

The relative masses for the two isotopes of bromine and their abundance are:
Bromine-79 : 78.92 and 50.69% Bromine-81 : 80.92 and 49.31%
Use this data to determine the relative atomic mass (A_r) for bromine.

Question 1.4

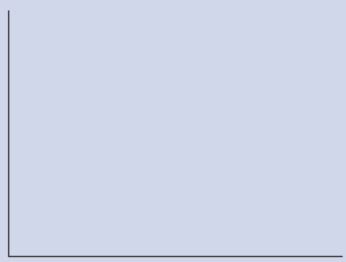
The mass spectrum for magnesium is shown at right. The relative isotopic masses are ^{24}Mg (23.99), ^{25}Mg (24.99) and ^{26}Mg (25.98).

Using the peak heights indicated determine:
(a) Percentage abundance for each isotope.
(b) Relative atomic mass (A_r) for Mg.



Question 1.5

An element X has two isotopes, ^{69}X and ^{71}X.
Draw a mass spectrum for these isotopes on the blank graph grid at right. The relative abundances are 60.1% and 39.9% and the isotopic masses 68.92 and 70.92 respectively.
Determine A_r for element X.



Question 1.6

There are two isotopes of copper; copper-63 and copper-65. Their isotopic masses are 62.93 and 64.93 respectively. The relative atomic mass for copper is 63.55. Use this data to determine the percentage abundance of each isotope of copper.

1.5 ATOMIC STRUCTURE AND LIGHT SPECTRA

The Nature of Light

We are all familiar with the colours of the rainbow and the multitude of colours in a fireworks display. But how are these different colours formed? The way electrons are arranged in atoms, and the ability of all atoms to absorb and emit energy, provides an answer. To better understand colour and spectra we need to briefly consider the nature of light.

Light can be described as being both wave-like and particle-like in nature and makes up a small part of the overall electromagnetic spectrum (see below). Different colours of light, or other radiation, can all be considered as waves of a particular wavelength and frequency. It can also be said that each of the colours of light, or other radiation, are made up of a stream of particles called photons. Photon energies vary; the greater the frequency of the radiation, the greater the individual photon energies.

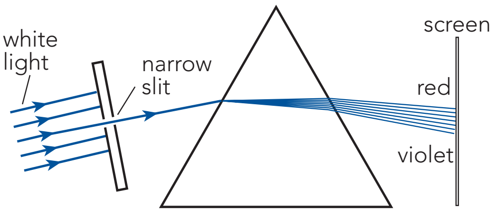


Figure 1.5 *Dispersion of white light by a prism. Light is a small part of the electromagnetic spectrum visible to the eye. The glass prism disperses the light into its component colours creating a continuous emission spectrum.*

The Electromagnetic Spectrum

Although not visible to the eye, there exists other radiation on either side of the visible spectrum. All radiation, like light, can travel through space. Radio waves, microwaves, ultraviolet waves and X-rays are all part of the electromagnetic spectrum and travel at the speed of light (3×10^8 ms^{-1} in a vacuum).

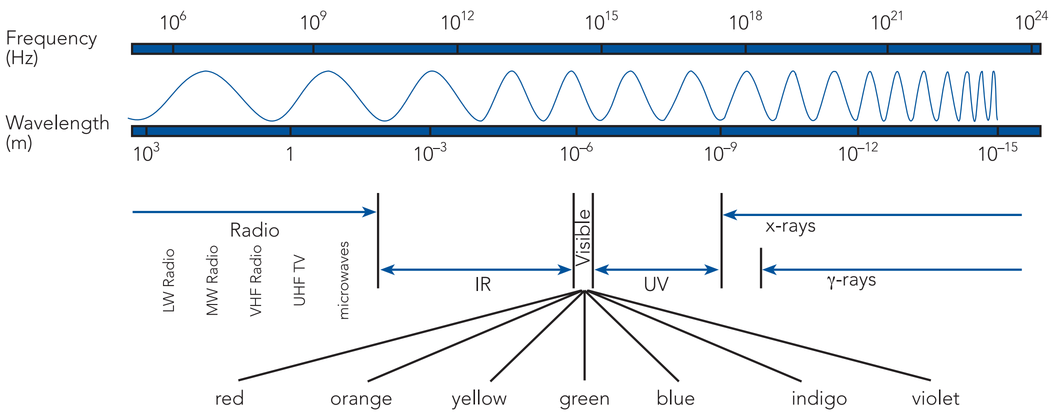


Figure 1.6 *The Electromagnetic Spectrum. All radiation travels at the speed of light. Radio waves have relatively large wavelengths (low frequencies) while X-rays have very short wavelengths (very high frequencies). Photon energies increase with frequency, hence X-Rays can be considered to be made up of high energy photons while radio waves are made up of low energy photons.*

The Hydrogen Spectrum

If an element such as hydrogen is heated by an electric current in a discharge tube, light is emitted. The spectra produced when this light is analysed through a prism is line emission spectra. In the visible range the spectra consists of four distinct coloured lines on a black background. Each of these coloured lines indicates light of a particular frequency, or energy, being given off by the atoms of hydrogen.

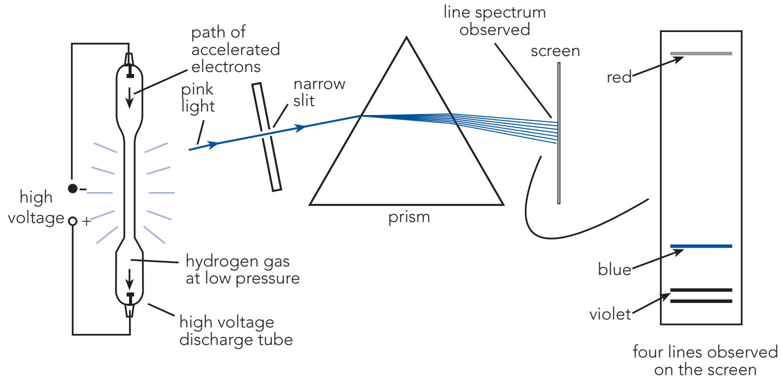


***Figure 1.7** Observing the line emission spectrum for Hydrogen.*

The Bohr model of the atom

In developing his theory, Neils Bohr was particularly keen on being able to explain the line spectra he observed from a hydrogen discharge tube. He proposed that the spectra was due to the movement of excited electrons of the hydrogen atoms falling back to their normal, or stable state, within the atom. The atoms electrons must have being initially excited to higher energy levels by the collisions of the atoms with the electrons produced and accelerated by the discharge tube.

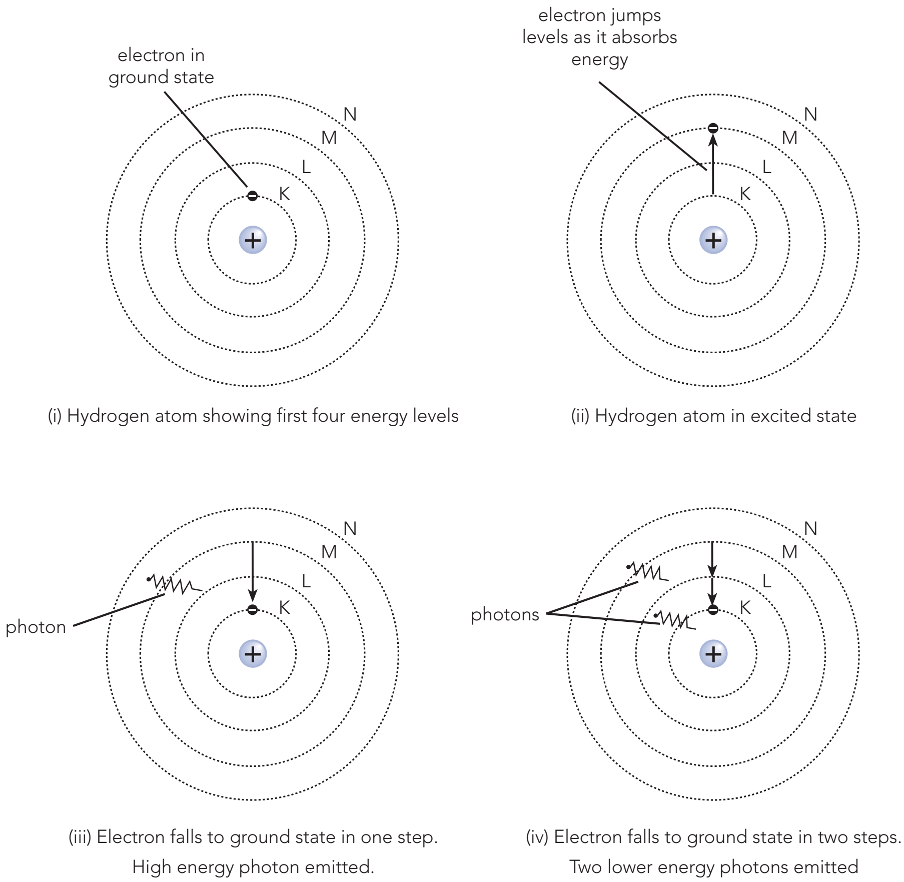


***Figure 1.8** Atomic energy levels and spectra. Atoms can absorb energy through heating, light or electrical discharge. Only specific amounts of energy are absorbed by the electrons as they move to possible higher levels. In the excited state atoms are unstable and electrons return to their ground state within a few nanoseconds. They can fall in one step or more but in each case a photon of light is emitted. The photon energy is the same as that lost by the electron falling. In the case shown above there are three possible photon energies created corresponding to three distinct spectral lines or colours.*

Bohr proposed that the electrons of atoms can only exist in specific energy levels. These levels, also referred to as shells, are denoted 1, 2, 3, ... or K, L, M, ... as we go outward from the nucleus. Electrons in atoms are normally in their lowest possible energy level or ground state. However, they can absorb energy and jump to higher levels in which case the atom is said to be in an excited, but unstable state. The electrons quickly fall back to their ground state, sometimes in more than one step, and emit light of a specific frequency corresponding to the energy jump.

Some possible electron transitions for the hydrogen atom are shown below.

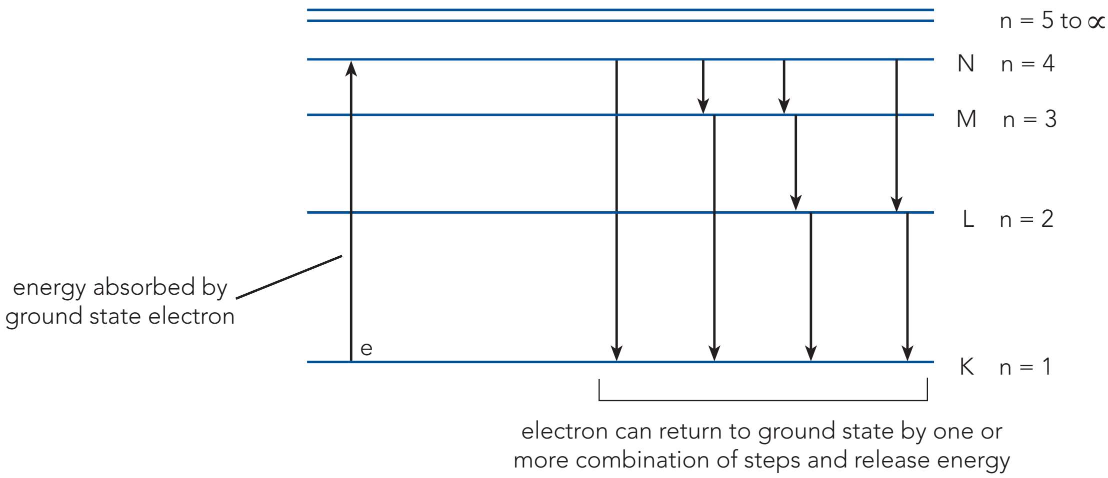


***Figure 1.9** Electron transitions in hydrogen and spectra. When hydrogen atoms absorb energy its electron can jump to any of the available energy levels including leaving the atom altogether (ionisation). In the example above an electron absorbs energy and jumps to the n = 4 level. Since it is not stable at this level it returns to the ground state. It can do so in a variety of steps. When considering a large number of atoms all possible pathways occur and we observe all the characteristic spectral lines for hydrogen. (Note that some will not be in the visible part of the spectrum).*

Question 1.7

(a) Which visible colour is least refracted by a prism? ____________________

(b) Which has the greater photon energy; IR or UV? ____________________

(c) Which is the longest wavelength light; blue or orange? ____________________

(d) How many visible lines are in the hydrogen spectrum? ____________________

Question 1.8

An excited electron in a hydrogen atom can return to its ground state directly or by a sequence of steps. Photons of different energies are emitted depending on the difference in energies between the levels traversed. Determine the number of possible photon energies emitted if an electron is initially excited to the:

(a) third level (n = 3) ____________________

(b) fourth level (n = 4) ____________________

(c) fifth level (n = 5) ____________________

1.6 ELECTRON ARRANGEMENTS IN ATOMS

The way that electrons are arranged in atoms is very important as it determines chemical behaviour. The nucleus is not involved in chemical reactions and is not affected by them.

Although they cannot say exactly where an electron is and how fast it is moving, scientists have established that electrons can only exist in specific **energy levels** within the atom. In 1912 Danish physicist Niels Bohr proposed a theory of the atom which was able to explain more clearly the behaviour of electrons in atoms and was consistent with quantum theory. He proposed that:

- electrons can only exist in specific energy levels
- electrons could be excited from one level to another by specific amounts of energy corresponding to the difference in energy levels
- energy, in the form of photons, is emitted whenever an electron moves from a high energy level to a lower one.

The Bohr theory helped explain the observation of emission and absorption spectra and is still the basis of today's atomic theory. However, a refinement of the Bohr model of the atom based on quantum mechanics (Erwin Schrodinger, 1926) better describes the nature of atoms, particularly multi electron atoms.

This refined model of the atom states that electrons move in regions of space called orbitals rather than in specific orbits. The quantum mechanical model of the atom states that within an atom there exists:

- principal energy levels, or shells, denoted 1 , 2 , 3 , . . . or K , L , M , . . .
- a maximum number of electrons for any level given by the formula $2n^2$, where n is the energy level number
- energy sublevels, or subshells, denoted s , p , d , f , . . .
- orbitals that make up the sublevels: s (1 orbital), p (3 orbitals), d (5 orbitals) and f (7 orbitals)
- each orbital can hold a maximum of 2 electrons.

Table 1.3 *Shells, subshells and orbitals in atoms*

ENERGY LEVEL (FIRST 4)	SHELL SYMBOL	SUBSHELL*	NUMBER OF ORBITALS IN SUBSHELL*	MAXIMUM NUMBER OF ELECTRONS ($2n^2$)
1	K	1s	1	2
2	L	2s 2p	1 3	8
3	M	3s 3p 3d	1 3 5	18
4	N	4s 4p 4d 4f	1 3 5 7	32

* *May not be required for your course.*

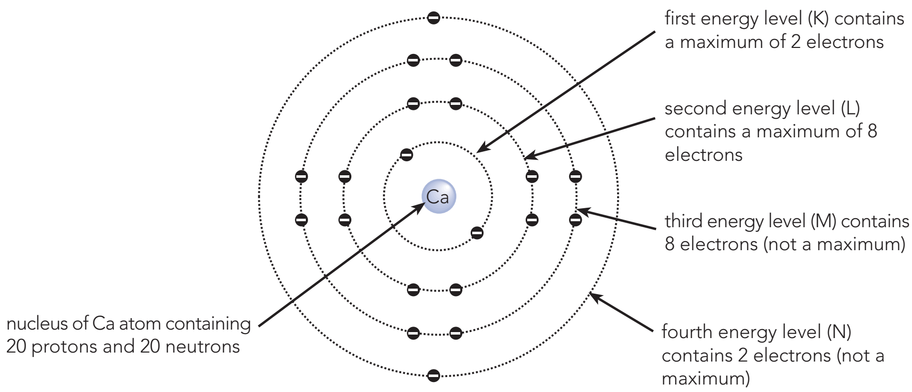


Figure 1.10 Simplified electron arrangement in a calcium atom

Table 1.4 The electron configuration of some of the first 20 elements. The electrons in the outermost shell of an atom are called valence electrons

ATOMIC NUMBER (Z)	ELEMENT	NUMBER OF ELECTRONS	ELECTRON CONFIGURATION			
			K	L	M	N
1	Hydrogen	1	1			
2	Helium	2	2			
3	Lithium	3	2	1		
7	Nitrogen	7	2	5		
10	Neon	10	2	8		
11	Sodium	11	2	8	1	
16	Sulfur	16	2	8	6	
18	Argon	18	2	8	8	
19	Potassium	19	2	8	8	1
20	Calcium	20	2	8	8	2

We can indicate how electrons are arranged in an atom by writing the electron configuration. This can be done simply by indicating the number of electrons that exist in each level of the atom. It is also possible to indicate more precisely in which orbitals all the electrons are by using s p d f notation* as shown below.

Two ways that we can write the electron configuration for a chlorine atom (Z = 17) are as follows:

number of electrons   orbital type

Cl   2 , 8 , 7   or $\underbrace{1s^2}_{\text{1st shell}}$, $\underbrace{2s^2\ 2p^6}_{\text{2nd shell}}$, $\underbrace{3s^2\ 3p^5}_{\text{3rd shell}}$

Similarly for example:

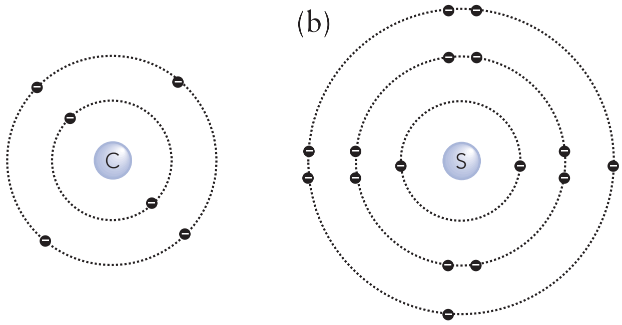
N   2 , 5   or $1s^2$, $2s^2\ 2p^3$
Ne   2 , 8   or $1s^2$, $2s^2\ 2p^6$
Ca   2 , 8 , 8 , 2   or $1s^2$, $2s^2\ 2p^6$, $3s^2\ 3p^6$, $4s^2$

** This form of notation (s, p, d, f) may not be required for your course*

Worked Example

1.4 Show the electron arrangement around: (a) carbon and (b) sulfur

(a) (b)



These are sometimes called electron energy level diagrams. Energy level diagrams are sometimes drawn with electrons in pairs, showing they occupy the same orbital. This is only done if an energy level contains more than 4 electrons. An abbreviated way of writing the electron structure of an atom is to give its electron configuration. For example, as shown in Table 1.4, the electron configuration for sodium is given as 2, 8, 1.

1.5 Write the electron configuration for the two atoms in example 1.4.

(a)	carbon	2 , 4
(b)	sulfur	2 , 8 , 6

1.6 How many electrons are there in the outermost energy level for the atoms nitrogen, fluorine and sodium?

Determine electron configurations first. Hence:

N	2 , 5	∴	5 electrons in outer level
O	2 , 6	∴	6 electrons in outer level
Na	2 , 8 , 1	∴	1 electron in outer level

Question 1.9

Write the electron configuration for the following elements. To help you, use the periodic table to determine the atomic number (Z) of each element and hence the number of electrons for a neutral atom of that element.

(a) Carbon ____________________ (d) Fluorine ____________________

(b) Chlorine ____________________ (e) Calcium ____________________

(c) Magnesium ____________________ (f) Boron ____________________

Question 1.10

How many electrons are there in the outermost energy level of:

(a) Silicon ____________________ (d) Fluorine ____________________

(b) Aluminium ____________________ (e) Chlorine ____________________

(c) Sulfur ____________________ (f) Carbon ____________________

1.7 FLAME TESTS

As we have learnt, the colours we may observe in a flame are due to electron transitions within excited gaseous atoms. A flame test can be used to identity, although only qualitatively, a range of metal ions due to the characteristic colours produced when their salts are burnt. A small amount of the substance being tested is placed on a platinum wire and burnt in a very hot non-luminous Bunsen flame. Typical flame colours are yellow for sodium, red for barium and blue/green for copper.

Only a small number of metal ions can be distinguished in this way as a Bunsen flame is not hot enough to excite the electrons of many atoms. It is also difficult to distinguish between very similar colours. For example, reddish colours are produced by calcium, strontium and lithium compounds. The colours may also be affected by the presence of traces of other ions. However the use of reference flames of known compounds burnt under the same conditions will help. Observation of the flames in a darkened room using a spectroscope will also provide more certainty.

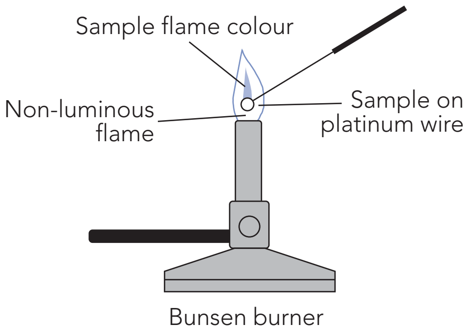


Figure 1.11 *Flame testing. Samples placed in a hot non-luminous flame give characteristic colours for different metallic ions.*

Table 1.5 *Typical flame colours for some metals.*

METAL IONS	FLAME COLOUR
Boron	Bright green
Barium	Pale green
Calcium	Orange/Red
Copper	Blue/Green
Iron	Gold
Lithium	Red/Crimson
Potassium	Violet/Lilac
Sodium	Yellow
Strontium	Deep red

Question 1.11

Use your understanding of atomic structure to explain how different colours are produced when different salts are placed in a hot Bunsen flame.

Question 1.12

Two different salts, sodium chloride and sodium nitrate, each give an intense yellow colour when flame tested. Explain why the same colour is observed.

Question 1.13

Describe a simple test that would distinguish between calcium nitrate and barium chloride salts.

1.8 ATOMIC EMISSION SPECTROMETRY (AES)

As mentioned above the use of a spectroscope for the visual observation of spectra from coloured flames is very useful. However it is largely qualitative and is limited only to the visible spectrum. A more efficient and quantitative method of analysing elements in a flame is the use of the atomic emission spectrometer.

The samples to be analysed are heated to much higher temperatures and the characteristic light is passed through a prism or diffraction grating, much like the spectroscope. However, instead of viewing all the spectra as a whole, a device called a monochromator allows only single wavelengths at a time to pass through. The spectra can then be detected and recorded in various ways.

The intensity of each spectral line can be recorded on film using a spectrograph or more conveniently as a graphic display using a spectrometer. In the spectrograph the darkness on the negative for a particular wavelength indicates intensity and the abundance of the element causing it. In the spectrometer the detected light from the monochromator is converted to an electrical current and then displayed digitally or graphically. Hence the atomic emission spectrometer provides a more practical and quantitative method for identifying most metalloid elements (see Figure 1.12 below).

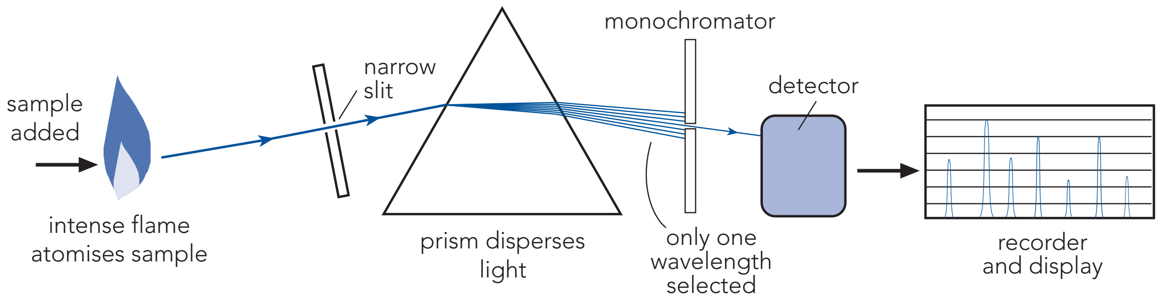


Figure 1.12 *The atomic emission spectrometer. The sample solution is drawn into the flame as a very fine spray and becomes vaporised and atomised. The excited atoms emit light which is characteristic of the elements in the sample. The prism disperses the light into its characteristic spectra with the monochromator allowing each wavelength to be detected and recorded separately.*

1.9 ATOMIC ABSORPTION SPECTROMETRY (AAS)

Absorption Spectra

As we have seen, white light can be dispersed by a prism and a continuous emission spectrum is produced (Figure 1.5). However, if the white light passes through a gas sample before being dispersed, then the resulting spectra will have dark lines present within an otherwise continuous spectrum. The dark lines (really an absence of light) are in the exact position that bright lines would appear in the emission spectrum of the sample gas used. This is called absorption spectra and can be used to identify the gaseous elements through which the light has passed.

A classic example of absorption spectra are the dark lines, called Fraunhofer lines, which appear within the continuous spectrum of sunlight. The cause of the dark lines is the absorption of specific frequencies of radiation as the light from the sun passes through its large gaseous atmosphere. In this way the presence of both hydrogen and helium on the sun were identified.

Atomic Absorption Spectrometer

Both emission and absorption spectroscopy are a useful means of analysing substances. However the accurate measurement of very small concentrations of metallic elements is quite difficult. In 1952 a very sensitive technique, called atomic absorption spectrometry, was developed by Alan Walsh, an Australian scientist of the CSIRO division of Chemical Physics. The technique did not involve the usual absorption from white light as a source, but rather, absorption of light emitted by the element being analysed.

The atomic absorption spectrometer (AAS) was first demonstrated at the Melbourne University in 1954. An essential component of the spectrometer is the hollow cathode lamp whose cathode is coated with the metal being investigated (see Figure 1.13 below). Hence the light which passes through the atomised sample contains the exact wavelengths that can be absorbed by the metal being analysed. The amount of cathode light absorbed by the metal atoms gives a measure of their concentration. Importantly, the detectors are able to distinguish between the cathode light left over after absorbance and the light naturally emitted by the gaseous atoms in the flame returning to their ground state. This is achieved by pulsing (or chopping) the light from the cathode so that the detector can distinguish it from the continuous beam from the flame.

The atomic absorption spectrometer is one of the most important scientific instruments developed in Australia. It provides a sensitive and high speed technique for the measurement of small traces of metals down to a few parts per billion. Today it is an essential analytical tool used in agriculture, mining, industry, hospitals and chemical laboratories.

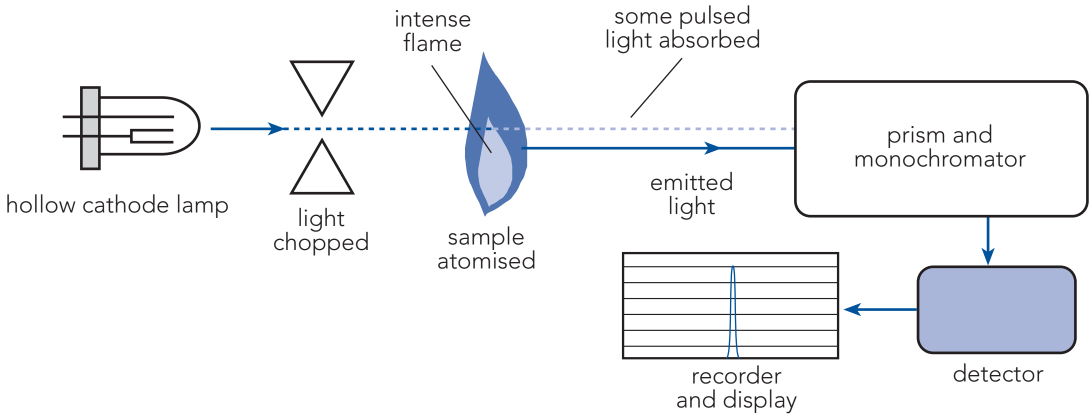


Figure 1.13 *The atomic absorption spectrometer. Light from the hollow cathode lamp has the specific wavelengths that the metal being analysed can absorb. The detector effectively measures the amount of this light which is absorbed and this indicates the concentration of the metal atoms in the sample.*

Question 1.14

Explain the purpose of each of the following in the atomic emission spectrometer.

(a) Prism ______________________________

(b) Monochromator ______________________________

Question 1.15

Explain the reason of each of the following in the atomic absorption spectrometer.

(a) The cathode of the lamp is coated with the metal being analysed.

____

(b) Light from the lamp is pulsed ______________________________

1.10 ELECTRON CONFIGURATIONS AND THE PERIODIC TABLE

The position of an element in the Periodic Table can tell us the number of electrons it has in its outermost energy level. This is important because an element's chemical properties are controlled by how many electrons are in its outermost energy level.

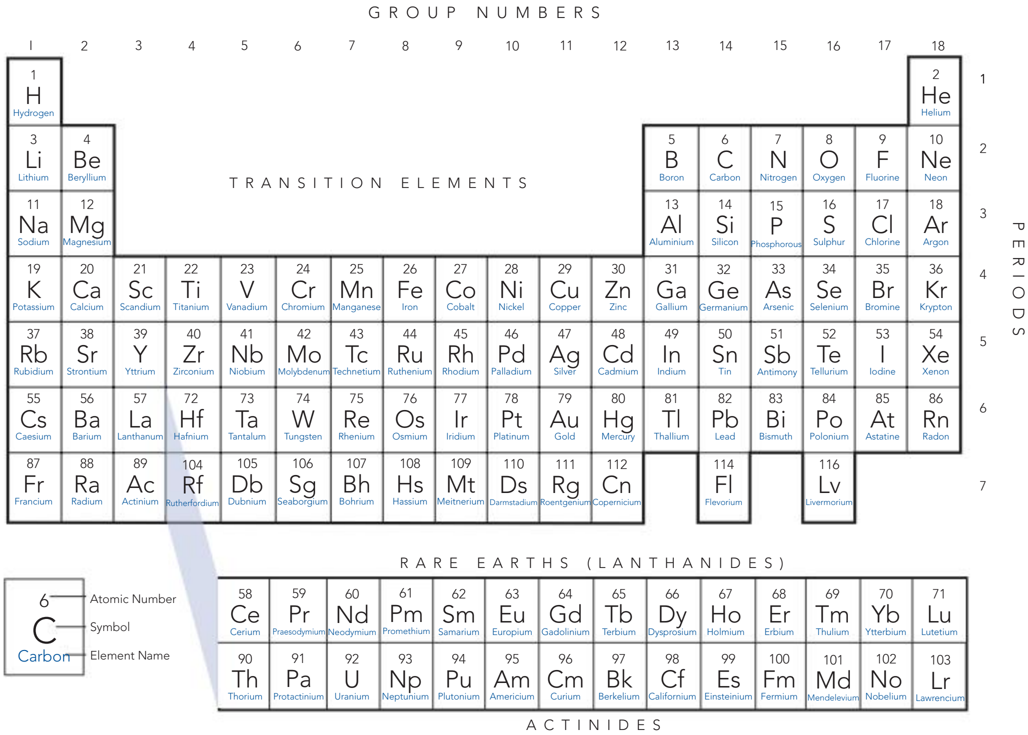


Figure 1.14 Electron configuration and the periodic table. The position of an element on the table is related to its electron configuration

The column number is often referred to as the Group Number. Because they have the same number of electrons in their outermost energy level, referred to as **valence electrons,** elements in the same group have similar chemical properties.

Some groups are known by specific names such as the **alkali** metals (group 1), the **halogens** (group 17) and the **noble gases** (group 18). The horizontal rows are referred to as periods. The elements in period 3 for example (Na to Ar) all have valence electrons in the third energy level.

Question 1.16

Group 1 of the periodic table are called alkali metals. What feature is common about their electronic configuration?

____

Question 1.17

How many valence electrons have atoms from each of the following groups?

(a) Group 2 , the alkali earth metals ______________

(b) Group 17 , the halogens ______________

(c) Group 18 , the noble gases ______________

1.11 FROM ATOMS TO IONS

When elements react the atoms involved must collide. This collision causes the outermost electrons to interact. Atoms involved in a reaction try to get the same electron configuration as the nearest noble gas.

For most elements this means they try to get 8 electrons in their outermost energy level. This is referred to as a stable octet.

Worked Example

1.7 Show how atoms of (a) aluminium and (b) oxygen can get a full outer energy level.

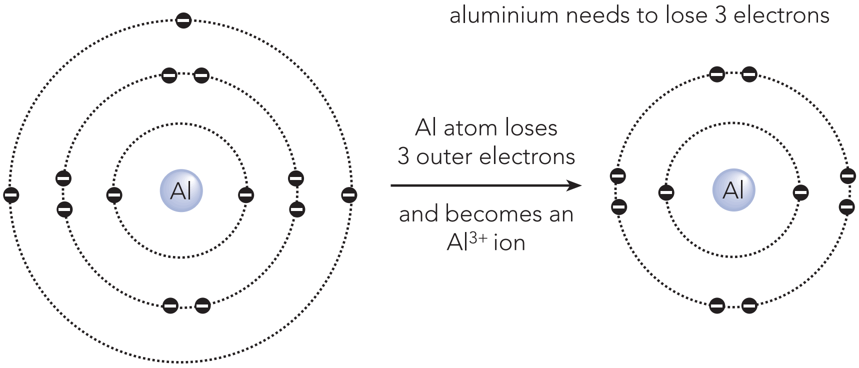
(a) Al : 2 , 8 , 3 Nearest noble gas Ne : 2 , 8

Hence aluminium needs to **lose** 3 electrons.

electron configuration of Al : 2 8 3

nearest noble gas : 2 8 (NOT 2 8 8)

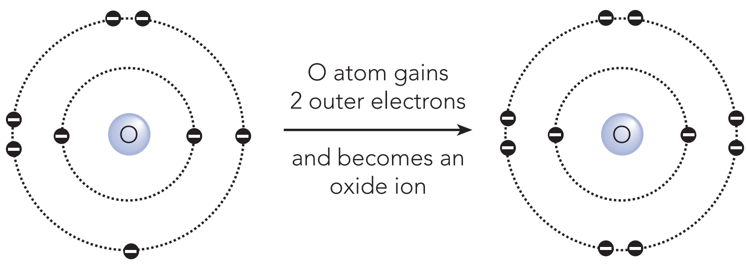
aluminium needs to lose 3 electrons



(b)

electron configuration of O : 2 6

nearest noble gas Ne : 2 8



Question 1.18

Four elements and their electronic configuration are as follows:

Ca : 2 , 8 , 8 , 2 S : 2 , 8 , 6 F : 2 , 7 Ar : 2 , 8 , 8

In a chemical reaction which element is most likely to:

(a) gain 2 e^-? ________________

(b) lose 2 e^-? ________________

(c) not react? ________________

1.12 ELECTRON DOT DIAGRAMS

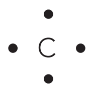
An electron dot diagram is a simple way of showing the valence electrons of an atom. Electrons are represented by a dot (or a cross). A maximum number of 8 electrons in the outermost energy level is possible.

Electrons are placed in four regions (orbitals) around the symbol for the element; a maximum of 2 electrons occur in each orbital.

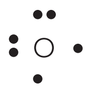
Worked Example

1.8 Draw the electron dot diagrams for magnesium, carbon and oxygen.

Calcium has 2 electrons in its valence shell (2 , 8 , 2). Each of these electrons is situated in separate orbitals.



Carbon has 4 electrons in its valence shell (2 , 4). Each of these electrons is situated in separate orbitals.



Sulfur has 6 electrons in its valence shell (2 , 6). Two orbitals are full (have 2 electrons in them) and two orbitals are partly full (one electron in each).

Question 1.19

Draw the electron dot diagrams for each of the following elements:

(a) Ar (c) B

(b) Cl (d) N

Question 1.20

Which of the following atoms would need to gain the greatest number of electrons to achieve noble gas configuration?

(a) fluorine (b) sulfur (c) nitrogen (d) oxygen

Answer: ____

Question 1.21

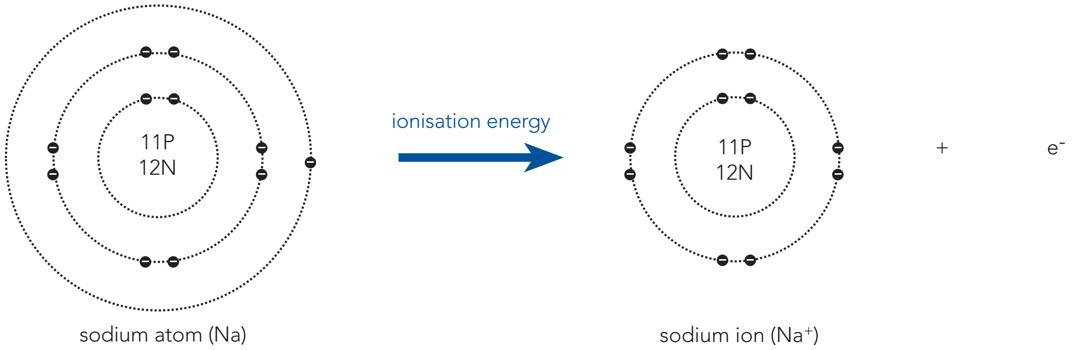
Complete the table below:

SECOND ROW ELEMENT	Li	Be	B	C	N	O	F	Ne
Electron configuration								
Electron dot diagram								

1.13 IONISATION ENERGY

Ionisation energy is the energy required to remove the most loosely held electron from an atom in the gaseous phase. This energy helps overcome the electrostatic attraction between the protons in the nucleus and the electron.

e.g. $Na_{(g)}$ + Energy $\rightarrow$ $Na^{+}_{(g)}$ + e^-



Ionisation energy depends on two main factors; the nuclear charge and the distance that the valence electron is from the nucleus.

- **nuclear charge** - the greater the nuclear charge, that is the number of protons in the nucleus, the greater the ionisation energy. For example the first ionisation energy of chlorine (1251 kJ/mol) is much greater than that for sodium (496 kJ/mol) even though both elements are from the third period. The greater number of protons in the nucleus of the chlorine atom means that its valence electrons are more tightly held. This greater attraction is reduced to some degree by the shielding effect of the extra electrons in a chlorine atom.
- **distance from the nucleus** - the further the electron is from the nucleus the lower is the ionisation energy required. Electrostatic attraction decreases markedly with distance which means that the valence electrons of the larger atoms are not strongly held. The shielding effect of the inner-shell electrons also reduces this attraction. The first ionisation energy of potassium (419 kJ/mole) for example is less than for sodium (496 kJ/mole) even though potassium has a greater nuclear charge.

In relation to the periodic table this means that generally:

- metals have low ionisation energies and ionisation energy decreases as we go down a group.
- non-metals have high ionisation energies since ionisation energy increases as we go across a period (left to right).

This is best illustrated by a diagram as shown below:

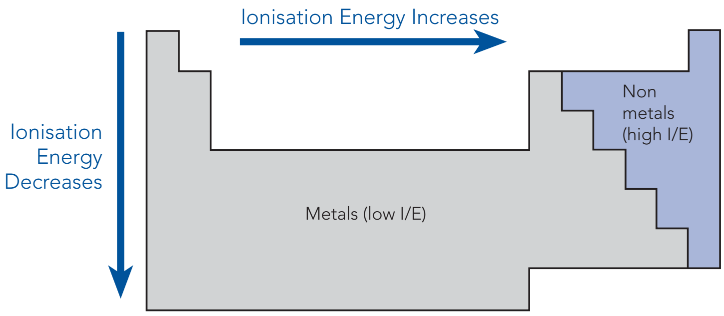


Figure 1.15 Trends in ionisation energies

1.14 REMOVAL OF SUCCESSIVE ELECTRONS FROM ATOMS

The first ionisation energy refers to the energy required to remove only the first electron from a gaseous atom. More electrons can be removed, but it becomes increasingly difficult since the electron is now being removed from a positively charged ion. If an electron is being removed from a lower energy level the energy required will be much greater.

Table 1.5 *Successive ionisation energies for carbon*

PROCESS	ENERGY REQUIRED kJ mol^{-1}	
$Si_{(g)}$ + energy → $Si^{+}_{(g)}$ + e^{-}	787	(1st ionisation energy)
$Si^{+}_{(g)}$ + energy → $Si^{2+}_{(g)}$ + e^{-}	1577	(2nd ionisation energy)
$Si^{2+}_{(g)}$ + energy → $Si^{3+}_{(g)}$ + e^{-}	3232	(3rd ionisation energy)
$Si^{3+}_{(g)}$ + energy → $Si^{4+}_{(g)}$ + e^{-}	4356	(4th ionisation energy)
$Si^{4+}_{(g)}$ + energy → $Si^{5+}_{(g)}$ + e^{-}	16090	(5th ionisation energy)
$Si^{5+}_{(g)}$ + energy → $Si^{6+}_{(g)}$ + e^{-}	19810	(6th ionisation energy)

Question 1.22

Using their position on the periodic table, arrange the following elements in order of *increasing* first ionisation energy:

(a) C , N , Si increasing order of I/E is ____________________

(b) Na , Mg , K increasing order of I/E is ____________________

Question 1.23

(a) Referring to Table 1.5, which ionisation energy represents a large difference compared to the previous one? Why?

____

____

(b) How many electrons is this element likely to lose when forming an ion?

____

Question 1.24

The successive ionisation energies of an element X (in kJ mol^{-1}) are 738, 1451, 7733, 10540, 13630.

(a) How many valence electrons has element X? ____________________

(b) Indicate the valency of its ion. ____________________

(c) By comparing data with Table 1.5 do you suppose X is a metal or non metal? Why?

____

1.10 TRENDS AND THE PERIODIC TABLE

Elements in the periodic table are listed in order of increasing atomic number but in such a manner as to create vertical groups with similar chemical properties. As we have seen the position of an element on the periodic table is also related to its electron configuration.

As we examine the physical and chemical properties of the elements of a particular group we find many similarities but also marked differences. These differences appear to change in a regular pattern or trend as we go down the group. Similarly trends in the properties of the elements can also be observed as we go across a period of the periodic table.

Trends down a group

- Atomic size increases

  As we go down a group there is an extra shell of electrons and the atomic radii become larger.

- Ionisation energy decreases

  As the atom gets larger the outer electrons are further away from the nucleus and less energy is required to remove the most loosely bound electron from the atom. **Electronegativity** (the electron attracting ability of atoms) also decreases.

- Metallic nature increases

  Metals typically give up their electrons easily. Hence as we go down a group elements become more metallic. This can be pronounced. Group 14 for example, has the non metal carbon at the top and the metal lead at the bottom.

Trends across a period

- Atomic size decreases

  As we go across a period there are a greater number of protons attracting the electrons. Hence the electrons move closer to the nucleus. This means that an atom of argon, for example, is smaller than an atom of sodium even though the argon atom has a much greater mass.

- Ionisation energy increases

  As we go across a period there are a greater number of protons in the nucleus (greater nuclear charge or core charge) and the electrons are more strongly held. Hence more energy is required to remove an electron from the atom. Ionisation energy and electronegativity both increase as we go across a period. Note that inert gases are not considered to be electronegative.

- Metallic nature decreases

  Electrons are more strongly held to the atom as we go across a period. Hence elements become less metallic in nature. Magnesium for example is less metallic than sodium and chlorine is a non metal.

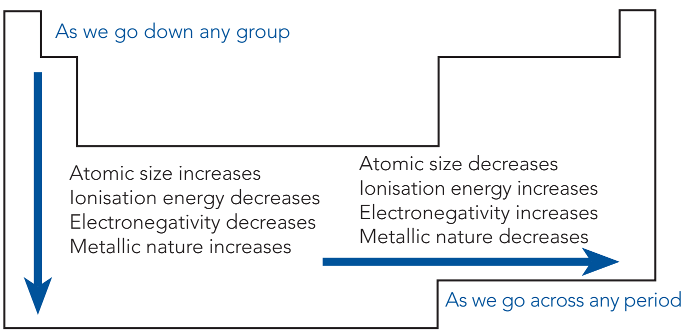


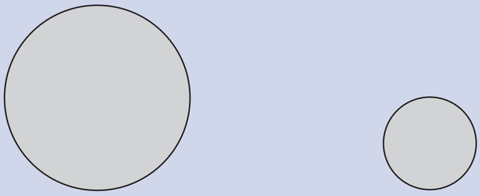
Figure 1.16 Trends in the periodic table

Question 1.25

The atomic radius of sodium ($\approx 157 \times 10^{-12}$ m) is significantly greater than that of the much heavier atom chlorine ($\approx 99 \times 10^{-12}$ m). Explain the reason for this.

____

____



Na $\approx 157 \times 10^{-12}$ m   Cl $\approx 99 \times 10^{-12}$ m

Question 1.26

Which element is likely to be more electronegative, sulfur or oxygen? Explain why.

____

____

Question 1.27

Which element of each of the following pairs would you expect to have the greater metallic character?

(a) Ca or Zn ________________   (c) Sr or Cs ________________

(b) K or Na ________________   (d) Al or Ca ________________

Question 1.28

Metals such as magnesium and calcium typically give up their electrons more easily than nonmetals. Referring to their position on the periodic table and periodic table trends, explain which of these two metals is more metallic.

____

____

Question 1.29

Three elements from the third group of the periodic table are aluminium, silicon, and phosphorus. Arrange these elements in order of increasing:

(a) atomic size ________________

(b) ionisation energy ________________

(c) metallic nature ________________

(d) electronegativity ________________

REVIEW QUESTIONS

Chapter 1: Atomic Structure

1. Complete the following table to summarise the basic structure of the atom.

PARTICLE	LOCATION	RELATIVE MASS (compared to a proton)	RELATIVE CHARGE

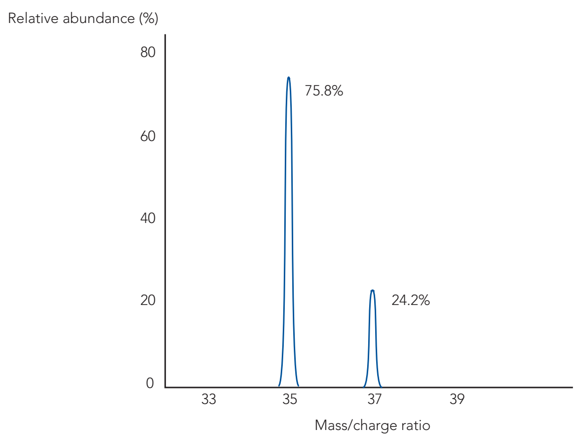
2. State the number of protons, neutrons and electrons in each of the following neutral atoms.

   (a) $^{7}_{3}Li$ (b) $^{23}_{11}Na$ (c) $^{37}_{17}Cl$ (d) $^{56}_{26}Fe$

3. The element boron has 6 known isotopes as listed below. The most abundant boron isotope is boron-11 (80.1%) with the remainder being mostly boron-10 (19.9%). Complete the following table to show the number of protons and neutrons that each isotope possesses.

ISOTOPE	SYMBOL	NUMBER OF PROTONS	NUMBER OF NEUTRONS
Boron-8			
Boron-10			
Boron-11			
Boron-12			
Boron-13			
Boron-14			

4. The graph below shows the relative abundance of the two isotopes of chlorine, ^{35}**Cl** and ^{37}**Cl**. Their relative isotopic masses are 34.97 and 36.97 respectively. Using the graph and this data determine the relative atomic mass (A_r) for chlorine.



5. The two naturally occurring isotopes of rubidium are ^{85}Rb and ^{87}Rb. Their relative isotopic masses are 84.91 and 86.91. The relative atomic mass for rubidium is 85.468. Determine the percentage abundance of each isotope of rubidium.

6. The electromagnetic spectrum is made up of many types of electromagnetic radiation including visible light, X-rays, microwaves, UV light and radio waves. Place these radiation types in increasing order of wavelength.

7. When the light emitted by a discharge tube containing hydrogen gas is analysed with a spectrometer distinct coloured lines are visible. In terms of atomic structure briefly explain how these spectral colours are produced.

8. Flame tests can be used to help identify the presence of some metal ions due to their characteristic flame colours. Briefly explain the following.

   (a) The origin of the characteristic colours for different metal ions in the flame
   (b) The limitations of simple flame tests.

9. Fluorine has a ground state electron configuration of 2, 7

   In a similar manner write the ground state electron configurations for the following neutral elements:

   (a) He (b) Al (c) Na (d) Mg (e) C (f) Ca

10. How many valence electrons are there in an atom of each of the elements described below?

    (a) An alkali metal from the 5th period.
    (b) The element from group 2 in period 3.
    (c) A noble gas from period 4.
    (d) The element from period 4 that belongs to group 17.
    (e) The element from the third period that belongs to group 16.

11. (a) What is true about the electron configuration of all elements belonging to the same group on the periodic table?
    (b) Apart from helium, what outer energy level electron configuration do all noble gases have?
    (c) What is (chemically) important about this electron configuration?
    (d) When atoms become ions, what general rule appears to be followed?

12. Match the following chemical species to the correct electron configurations.

Al^{3+}	2
S^{2-}	2, 6
O	2, 8
He	2, 4
C	2, 8, 8

13. The electron configuration of a chemical species is 2, 8, 6. Which of the following statements **might** be true about this species?

(a) It belongs to group 14 in the 3rd period
(b) It is a non-metal with a valency of +2.
(c) It can gain two electrons to become more stable.
(d) Ar^{2-} would possess this electron configuration (if it existed).
(e) S^{2-} would possess this electron configuration (if it existed).

14. Why do elements belonging to group 2 tend to form ions with a valency of +2?

15. Write the electron configuration of each of the following:

(a) Mg^{2+} (d) Ar

(b) Cl^{-} (e) H

(c) S^{2-} (f) Be^{2+}

16. Draw electron dot diagrams for the following atoms:

(a) Ca (c) O

(b) F (d) Ne

17. Define the term **ionisation energy**.

18. The following table shows the first ionisation energy for some of the elements belonging to groups 17 and 18.

GROUP 17 ELEMENT	IONISATION ENERGY OF GROUP 17 ELEMENT
Fluorine	1680 kJ
Chlorine	
Bromine	1140 kJ
Iodine	1010 kJ

GROUP 18 ELEMENT	IONISATION ENERGY OF GROUP 18 ELEMENT
Neon	2080 kJ
Argon	1520 kJ
Krypton	
Xenon	1170 kJ

(a) Complete the table by estimating the ionisation energies for Cl and Kr.

(b) What trend in ionisation energy of elements down a group, does this table show?

(c) Write an equation (including energy) for the first ionisation of fluorine.

19. (a) Which element has the lowest first ionisation energy of all elements in the third period?

(b) Which element has the second highest first ionisation energy of all the elements belonging to the 2nd period?

(c) For any period, which element has the highest first ionisation energy?

20. (a) Why does first ionisation energy decrease down a group in the periodic table?

(b) Why does first ionisation energy increase from left to right across a period in the periodic table?

21. Explain the following trend in successive ionisation energies for any individual element:

1st ionisation energy < 2nd ionisation energy < 3rd ionisation energy, etc.

22. Complete the following table by predicting the valency and the group number of each of the elements listed:

NUMBER OF IONISATIONS AND IONISATION ENERGY IN kJ						GROUP NUMBER	VALENCY
	1st	2nd	3rd	4th	5th		
X	578	1817	2745	11577	14842		
Y	590	1145	4912	6490	8153		
Z	1086	2256	4620	6223	37831		

23. If you were told that elements X, Y and Z (question 22) belonged to different periods, which do you think belongs to:

(i) period 2? (ii) period 3? (iii) period 4?

Justify your answers.

FOR THE EXPERTS

The Carbon-14 Isotope

24. Carbon has several isotopes, the three most abundant are listed below. The most common carbon isotope is carbon-12 (98.93%) with the remainder being mostly carbon-13 (1.07%). The relative isotopic masses of the ^{12}C and ^{13}C are 12.000 and 13.003 respectively.

Very small traces of carbon-14 occur in the atmosphere. Carbon-14 is continually produced in the atmosphere due to the action of neutrons with nitrogen-14 atoms. However, carbon-14 decays back to nitrogen-14 at the same rate as its formation leaving a constant small presence of this isotope in the air.

Carbon-14 is present in all living things through the formation of carbon dioxide which is then absorbed during photosynthesis. The resulting sugars and starches are eaten by animals. When an organism dies, the presence of the carbon-14 isotope is reduced since it is no longer being replaced. The reduced presence of carbon-14 in dead organisms can be used to date them. The half-life of C-14 is 5715 years.

(a) Complete the following table to show the number of protons and the number of neutrons for each isotope.

ISOTOPE	SYMBOL	NUMBER OF PROTONS	NUMBER OF NEUTRONS
Carbon-12			
Carbon-13			
Carbon-14			

(b) The relative atomic mass for carbon (atomic weight) is given as 12.01 (see periodic table). Show, using the data above, how this value can be verified.

(c) The processes for the formation and decay of carbon-14 in the atmosphere are as follows:

$$^{14}_{7}N + ^{1}_{0}n \rightarrow ^{14}_{6}C + X$$

$$^{14}_{6}C \rightarrow ^{14}_{7}N + Y$$

What is the likely identity and nature of particles X and Y?

(d) Carbon dioxide molecules in the atmosphere may contain carbon-14 atoms (about 1 in 10^{12}) instead of carbon-12 atoms.

(i) In what way would the properties of the gases made up of these different molecules vary?
(ii) In what way would they be the same?

MATERIALS

2

Topics covered in this chapter:

2.1 ELEMENTS, COMPOUNDS AND MIXTURES

A substance may be classified as either an element, a compound or a mixture. Pure substances (elements and compounds) are uniform in composition while mixtures may be either homogeneous or heterogeneous.

An **element** is a pure substance that is made up of only one kind of atom. There are only 90 naturally occurring elements (see periodic table on the inside cover). Of the first 92 elements to uranium, technetium and promethium are synthetically prepared. Elements beyond uranium are also artificially produced and referred to as transuranium elements. The most abundant elements in the Earth's crust are oxygen (46.1%), silicon (28.2%), aluminium (8.2%) and iron (5.6%).

A **compound** is a pure substance that is made up of two or more different elements chemically combined. There are literally millions of different compounds. Some, like water and carbon dioxide, consist of simple small molecules while others, such as organic compounds, can be quite large and complex.

A **mixture** is an impure substance made up of two or more pure substances. The components that make up a mixture can be easily separated by physical separation techniques such as filtration and distillation. Homogeneous mixtures such as solutions have a uniform composition. Heterogeneous mixtures such as concrete have variable composition.

Question 2.1

Choose from the following descriptions and examples to complete the table below.

Descriptions

- uniform composition throughout
- cannot be separated into simpler substances
- two or more different elements chemically combined
- non-uniform composition

Examples: *sea water, sugar, gold, limestone, rock, water, air, cement, oxygen, solder, rubber, salt, brass, aluminium, petrol.*

MATTER	
Pure substances Constant composition throughout. Unique set of physical properties.	**Mixtures (impure substances)** Variable composition. Constituents can be separated by simple physical methods.
Elements • Description ____________ • Examples ____________	**Homogeneous** • Description ____________ • Examples ____________
Compounds • Description ____________ • Examples ____________	**Heterogeneous** • Description ____________ • Examples ____________

Question 2.2

(a) A compound and a solution are both homogeneous substances. How are they different?

____

____

(b) Is a solution different to a mixture? Explain.

____

____

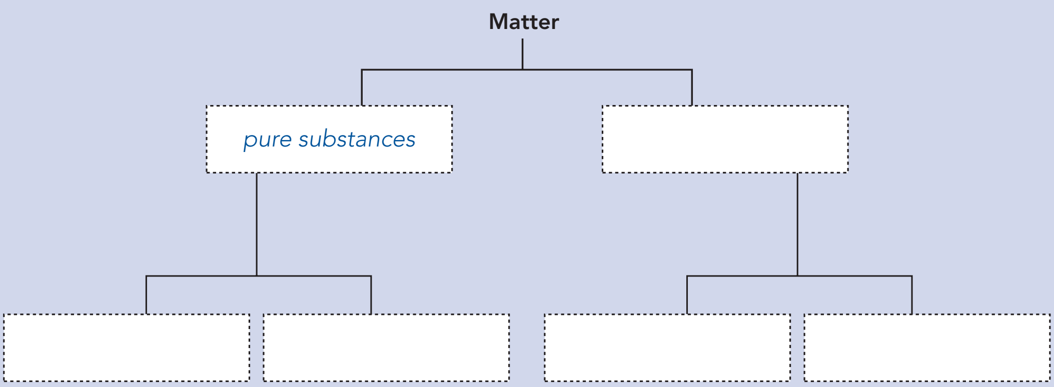
(c) A salt was heated strongly in a crucible until it became molten. An electric current was then passed through it using inert electrodes. A silvery metal formed at one electrode and a yellow/green gas was given off at the other. Is the salt an element or a compound? Explain.

____

____

Question 2.3

Use the information from the table in Question 2.1 to complete the classification diagram below. It should include elements, compounds, mixtures, heterogeneous, homogeneous, solutions and pure substances. The first box has been completed for you.



Question 2.4

List the following substances as either elements, compounds or mixtures.

Pure water, sea water, sugar, oxygen gas, carbon dioxide gas, zinc, vinegar, air, gold.

- Elements: ____
- Compounds: ____
- Mixtures: ____

2.2 SEPARATION OF MIXTURES

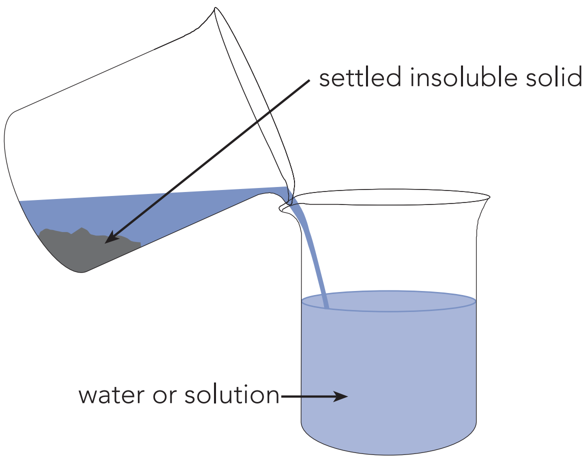
Most materials are found as mixtures in nature and not as pure substances. Mixtures can be separated into their pure constituents using physical separation techniques such as decantation, filtration, evaporation, recrystallisation, distillation and chromatography.

The separation technique used depends on the properties of the constituents in the mixture.

Decantation

Decantation is the pouring off of a liquid from a settled solid. It can be used to separate liquid from undissolved solid. Separation is possible due to low solubility and density.

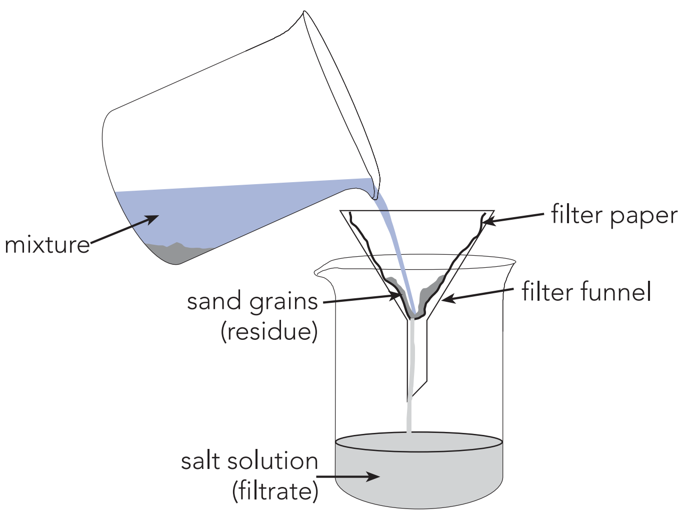
Common examples include the separation of:
- sand from water
- heavy (dense) precipitate from solution
- mercury from nitric acid solution.



Filtration

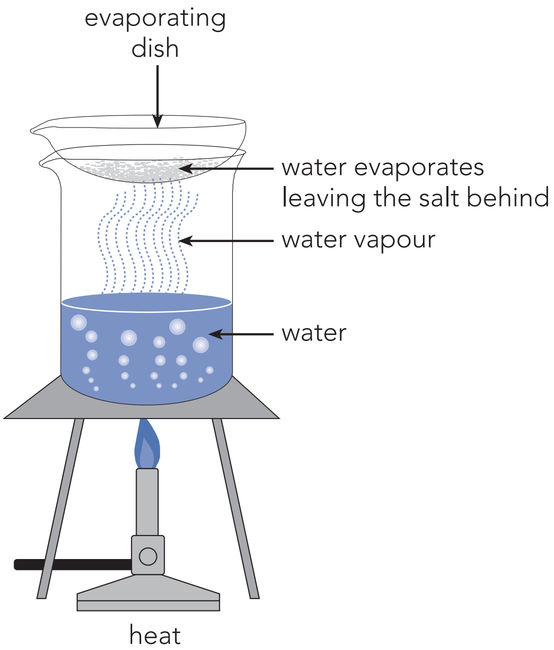
Filtration is the separation of undissolved solids from a liquid or solution using filters. It can be used to separate insoluble solids from soluble ones. Separation is possible due to differences in solubility.

Common examples include the separation of:
- sand from sea water
- charcoal from salt
- calcium carbonate from sodium carbonate.



Evaporation

Evaporation is the recovery of a dissolved solid from a solution. A dissolved solid, such as salt in sea water, cannot be separated by filtration. However, by evaporating the solvent (water) we can recover the solute (salt). Separation is possible because the salt cannot evaporate.

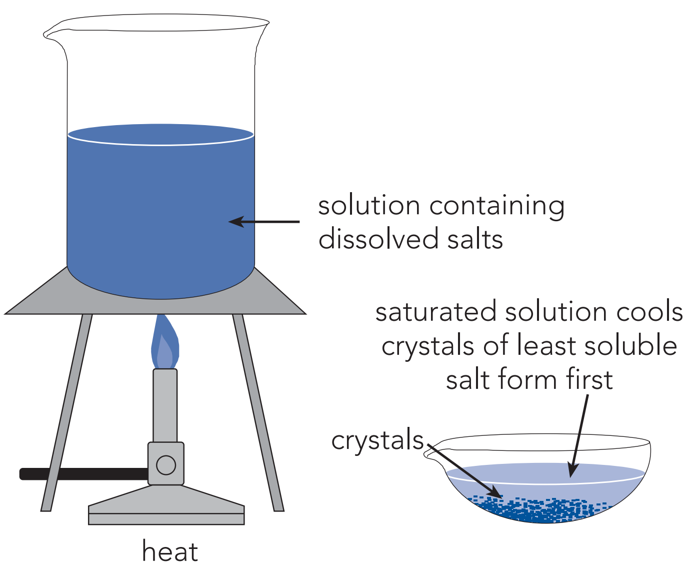


Crystallisation

Crystallisation is the recovery of dissolved solids as pure crystals from a saturated solution. It is very useful for separating different soluble salts from solutions. This separation is possible due to differences in solubility.

Common examples include the separation of:

- pure copper (II) sulfate crystals from solution
- potassium nitrate from sodium chloride
- table salt from baking soda.



Simple Distillation

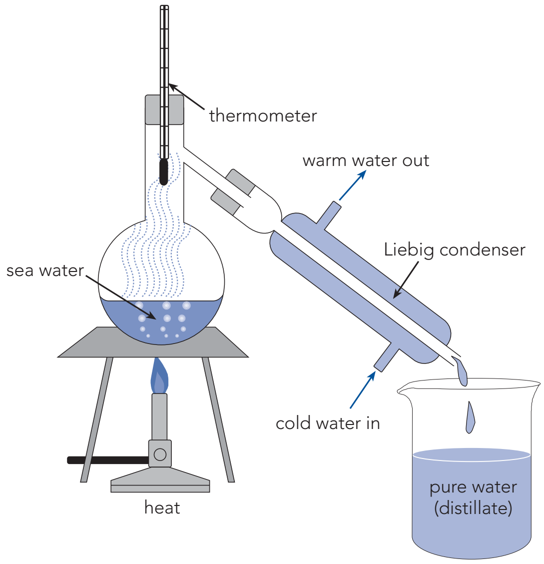
This is the recovery of a liquid from solution by means of evaporation and condensation. Separation is possible since dissolved solids cannot evaporate.

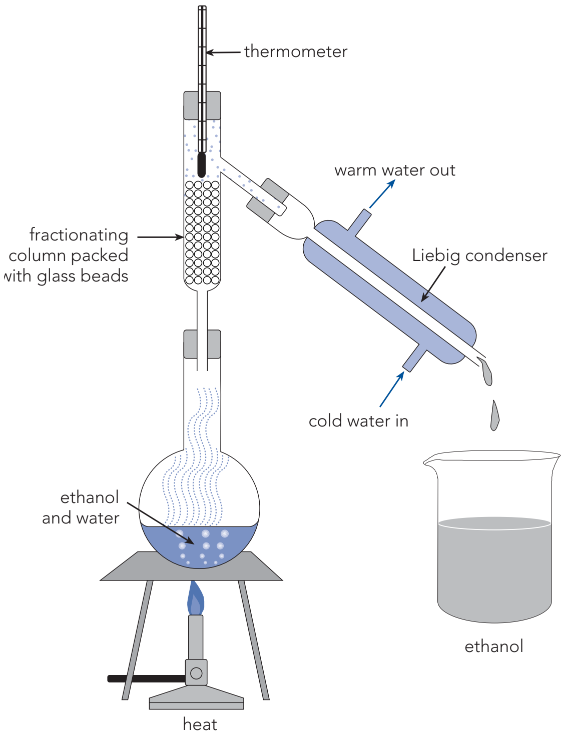
Common examples include the separation of:
- pure water from sea water
- pure water from ink.

Fractional Distillation

This is the separation of two or more liquids from a mixture. A fractionating column is added to the distilling flask. Separation is possible due to differences in the boiling point of the liquids.

Common examples include the separation of:
- alcohol from water
- alcohol from wine
- petrol, kerosene and oil from crude oil.





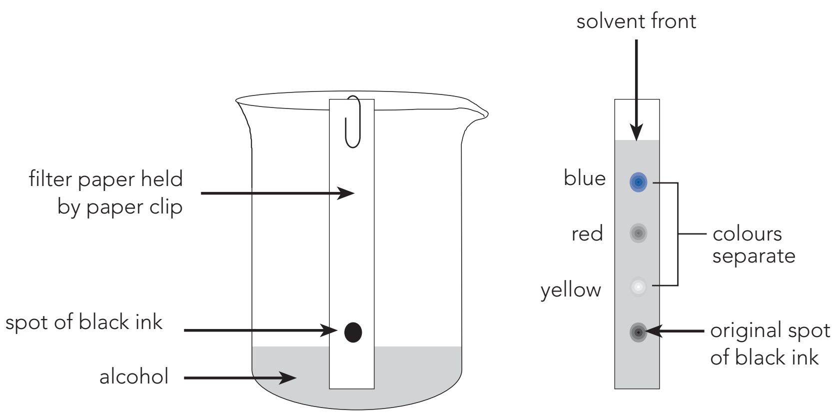
Chromatography

Chromatography is the separation of small amounts of dissolved solids such as the different components that make up black ink.

This method is most useful in identifying unknown substances. Separation is possible due to differences in solubility and selective adsorption.

Common examples of paper chromatography include the separation of:
- coloured dyes from black ink
- different plant pigments
- different food colourings.

* *Chromatography techniques are covered in much more detail in Chapter 7.*



Question 2.5

Which techniques should be used to separate

(i) pure water from salty bore water? ______________________

(ii) salt from salty bore water? ______________________

(iii) oxygen from liquid air? ______________________

(iv) calcium carbonate (insoluble) from water? ______________________

(v) calcium chloride (soluble) from calcium sulfate (slightly soluble)?

____

Question 2.6

(a) What is the purpose of running water through a Liebig condenser?

____

(b) Why is the inlet of the water lower than the outlet?

____

(c) How is fractional distillation different to simple distillation?

____

Question 2.7

Sugar, sand, iron filings and some dry leaves are mixed together thoroughly. Suggest a procedure for separating and recovering each of these substances.

____

Question 2.8

Separation of the components of a mixture is possible due to differences in the properties of the components. For example salt can be separated from a salt/sand mixture due to a difference in the solubility in water. For each of the following list the property and/or difference that is utilised in separating the first mentioned component (bold) from the second.

(i) salt from water ______________________

(ii) water from salt ______________________

(iii) sand from water ______________________

(iv) charcoal from salt ______________________

(v) KNO_3 from NaCl ______________________

(vi) blue dye from ink ______________________

(vii) alcohol from wine ______________________

(viii) salt from sugar ______________________

2.3 NANOMATERIALS

Nanoparticles

The physical and chemical properties of bulk materials are usually expected to remain the same whatever their size. However, if the particle size of a substance does become very small, then their properties can markedly change. This is particularly noticeable if the particle size approaches the nano scale.

Nano particles are defined as those between 1 and 100 nm. Typical examples are the C_{60} molecule, the buckyball, which is just over 1nm in diameter and nanotubes which are 1.3 nm in diameter but up to millimetres in length.

This range of particles is often considered as being a bridge between the macro world of bulk materials and the atomic and molecular world. They exhibit novel and potentially very useful properties. Through nanotechnology this has led to the development of many new substances, structures and devices.

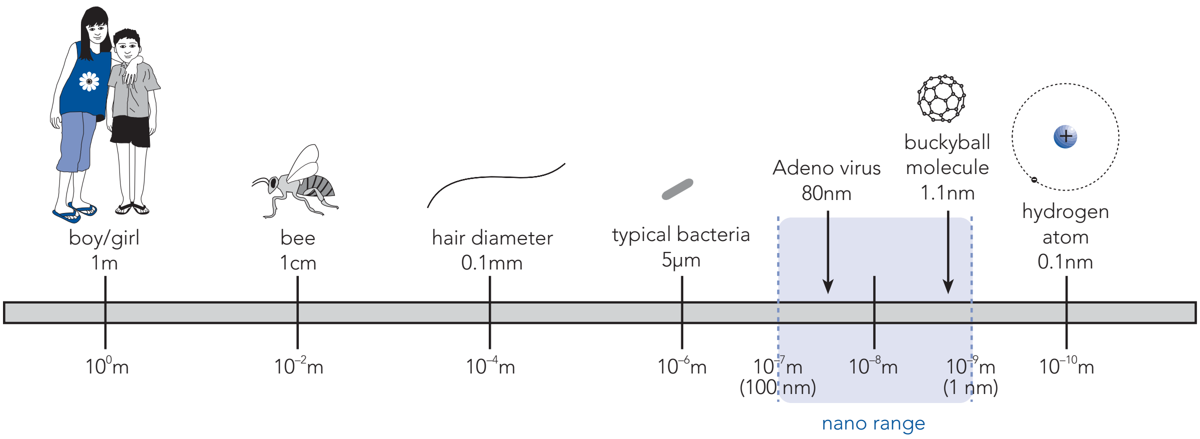


Figure 2.1 The nanoparticle range. The C_{60} molecule is just over one nanometer in size and one of the smallest nanoparticles. A nanotube is a similar size in diameter but can be very long, up to 1mm or so.

Question 2.9

(a) How many nanometers are there in each of the following

(i) 1.0 m ________________

(ii) 1.0 mm ________________

(iii) 1.0 μm ________________

(b) A soccer ball is approximately 22 cm in diameter. In comparison, the van der Waals diameter of a C_{60} molecule or buckyball is approximately 1.1 nm.

Calculate the factor by which the soccer ball is larger in terms of:

(i) diameter ____

(ii) volume ____

Note: for a sphere $V = \frac{4}{3}\pi r^3$

The effect of surface area

As particles become smaller their relative surface area to volume increases. This affects the properties of materials since much of their chemistry occurs at the surface. While for particles larger than, say a micron, the percentage of their atoms at the surface is very small; for particles in the nano range this becomes quite significant. The increased surface area leads to materials with some unusual properties and generally more reactive in nature.

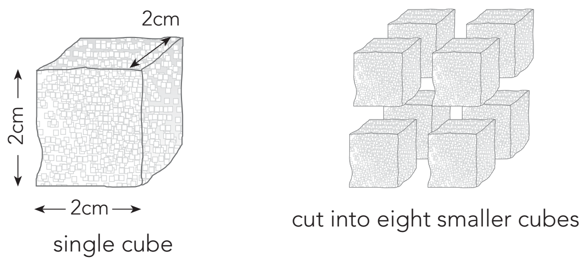


***Figure 2.2** The effect of particle size and relative surface area. When a sugar cube is cut into eight smaller cubes the total surface area doubles. This of course will increase the rate of dissolution. How much greater would the total surface area be if we could cut the sugar cube into cubes of only 1 mm sides? What if they were only 1 nm sides?*

Worked Example

2.1 A sugar cube measuring 2.0 cm × 2.0 cm × 2.0 cm as shown in Figure 2.2 is cut into 8 smaller and equal cubes. Assuming no loss of material calculate (a) the surface area of the original cube, (b) the total surface area of the 8 smaller cubes and (c) the factor by which the surface area changed.

(a)	Surface area of large cube	=	(6) × (2 × 2)	=	24 cm^2
(b)	Surface area of 8 smaller cubes	=	(8) × (6) × (1 × 1)	=	48 cm^2
(c)	Change in surface area	=	48/24	=	2

The surface area has doubled.

Question 2.10

Suppose the sugar cube shown above in Figure 2.2 could be further subdivided into very small cubes with 1.0 mm sides. Assume no loss of material.

(a) Determine how many 1.0 mm cubes there would be. ___________________

(b) Calculate the total surface area of the smaller cubes. ___________________

(c) By what factor does the total surface area differ from the original?

____

(d) What would this factor be if the cubes were nano cubes with sides of 1.0 nm only?

____

____

2.4 NANOTECHNOLOGY

As we have seen earlier nanoparticles have relatively high surface areas and hence a large percentage of their atoms or molecules can more readily interact. Their small particle size also results in unexpected and unique properties which are quite different to those of bulk materials. For example gold nanoparticles melt at a much lower temperature and appear red in solution. Some others such as zinc oxide nanoparticles are great absorbers of ultraviolet radiation; clay nanoparticles impart greater strength to polymers and titanium dioxide nanoparticles can create self-cleaning surfaces.

Nanotechnology makes use of these special properties of nanoparticles in creating many new products and techniques that are used in everyday life and industry. These include the use of such things as special coatings and smart surfaces, antimicrobial dressings and filters, stronger materials utilising nanoparticle additives and automotive catalytic converters.

Nanotechnology is a very large field of study and developments are ongoing in such areas as – drug therapy delivery systems, antimicrobial filters and dressings, nanocomposites offering lighter and stronger materials, computer chips and mass storage systems, solar cells and more efficient fuels and batteries . The possibilities are huge and diverse.

Below we take a brief look at a few of the more common and current examples which are the result of nanotechnology.

Sunscreens

A very common example of nanotechnology in action is the use of zinc oxide nanoparticles in some sunscreen lotions. These particles are excellent absorbers of ultra-violet light but unlike their bulk counterparts they create a transparent layer. This transparency is due to the fact that the zinc oxide nanoparticles being used are much smaller than the wavelength of visible light. Hence visible light is able to pass between the nanoparticles with very little being reflected.

The resulting sunscreens are very popular since they are very effective and importantly, also less visible. They are also cheaper to produce since smaller amounts of zinc oxide are actually needed.

Nanoparticles of zinc oxide and titanium dioxide are used in a similar way in paints and plastics. They absorb ultraviolet light and so help prevent the early breakdown of these materials when exposed to the sun.

Special surfaces and coatings

Nanoparticles can be used in a variety of ways to create special surfaces or alter the surfaces of materials such as fabrics in a useful way. Self-cleaning surfaces and stain resistant fabrics are two common examples.

A self-cleaning surface, on say window glass, can be achieved by coating it with a thin layer, typically 40 nm thick, of titanium dioxide nanoparticles. Several properties of these particles are helpful in achieving the desirable features of this surface.

Firstly, as with the invisible sunscreens discussed above, the particles used are so small as to reflect very little visible light. This makes the self-cleaning film transparent. Also, the coating causes photocatalysis to occur which assists in the chemical breakdown of any organic material on the glass. In addition, when water wets the glass the nanoparticles prevent water from forming droplets but rather a thin layer which flows and washes off the dirt.

Fabrics which are stain resistant or water repellent are now a reality through the use of polymer coatings containing suitable nanoparticles such as carbon nanotubes. The effect of the coatings is to produce very tiny nano sized bumps or hairs on the surface of the fabrics fibres. These prevent the water from soaking through but instead rolling off and taking any dirt with it.

This is sometimes referred to as the "lotus effect" since the leaves of lotus plants have similar characteristics. Manufacturers that produce a range of these fabrics have taken the cue from nature in using nanoparticles to mimic particular plant surfaces. Importantly, due to nanoparticles being so small, they do not affect the look or feel of the fabrics.

Nanocomposite Materials

Materials can be made lighter and stronger through the addition of suitable nanoparticles or simply by arranging the atoms of the element to form desirable nanostructures. These materials differ from normal composites in that the reinforcing phase, that is the nanoparticles added, have a high aspect ratio (much greater in length than width) and very high surface area to volume ratio.

This means that the addition of only small amounts of nanoparticles, say as little as one percent by weight, can markedly affect the properties of the resulting material. The concentration, type and shape of nanoparticles used can differently affect the final properties of the material created. This includes their melting point, strength, stiffness, durability, thermal and electrical conductivity, catalytic action and so on.

Hence a wide variety of useful nanomaterials are now available. Car manufacturers for example use carbon nanocomposites to create bumpers that are much lighter and more resistant to scratches and dents. A nanoflex alloy (carbon and iron) is widely used for the manufacture of sporting goods due to its high strength, hardness and formality. Special purpose composites are continually being developed such as sponge like silica surfaces which can trap toxic metals like lead and mercury in water.

A selection of nano-consumer products and the nanoparticles used is shown below.

PRODUCT DESCRIPTION	NANOPARTICLE/S USED
Antibacterial towels, garments	Ag
Auto paint finish	Ceramic
Auto sealants	SiO_2
Carbon fibre tennis racquet	C nanotubes
Ceramic filters	Ceramics, Ag
Computer processor chip	Cu, Si
Flash memory stick	Si
Fuel catalyst	CeO
Fullerene face cream	C_{60}
Golf clubs	C, Fe, Ti
Racing bicycles alloy frame	Al_2O_3, C
Stain resistant clothing	C nanotubes
Self cleaning window spray	TiO_2
Sunscreen lotion	ZnO

Table 2.1 *A selection of nano-consumer products and the nanoparticles used. There are literally thousands of new and improved products which are the result of nanotechnology. A great variety of materials are used as nanoparticles. A convenient website which lists the myriad of nano-consumer products on the market and the companies producing them is* **www.nanotechproject.org/cpi.**

Medical Applications

Nanotechnology in the medical field has resulted in many important and often lifesaving applications. These applications, in such areas as medical research, diagnostics and treatment, are both numerous and diverse. They range from the simple use of antibacterial bandages to the use of sophisticated diagnostic tools for the better imaging and screening of patients.

Some other commonly used applications include cancer therapy, controlled drug delivery, sunscreens, antioxidants, bone growth, dental ceramics and bio-composites.

The most exciting application in biomedicine is the use of nanoparticles for targeted drug delivery. This is an area of active research with different methods being tested to find, for example, a way that nanoparticles can be used to specifically target cancer cells.

It may be possible to encapsulate drugs within suitable nanoparticles so that they can more effectively be delivered to diseased cells without affecting healthy cells. In this way it may be possible to target undesirable bodies such as virus cells or penetrate through difficult areas such as the blood-brain barrier in order to reach and treat brain tumors.

Safety Issues and Regulations

The use of nanoparticles has provided a great number of new and very useful products. However since their properties are often markedly different from those of bulk materials their use poses potential hazards to health and the environment. Nanoparticles are generally very reactive and catalytic. Their unique properties and in particular the way they react with living cells and the environment is under continual analysis.

The CSRIO carries out ongoing research into the health, safety and environmental aspects of using nanotechnologies in manufacturing. Their research aims to understand and minimise possible risks. For example, in its study of the use of nanoparticles in sunscreens its research poses such questions as: Do zinc oxide nanoparticles penetrate the skin? If so, to what extent and to what effect? What are any long term health effects and what is the long term fate of the nanoparticles in the environment? Research worldwide is undertaken by many scientific and industry bodies in order to better understand the benefits and risks associated with nanotechnology.

Regulations also exist, both internationally and nationally, to regulate the use of nanonaterials to avoid potential risks to health and safety. In Australia for example, the Therapeutic Goods Administration (TGA) regulates the use of many products including sunscreens. Only approved ingredients that have been assessed for safety can be included in these products. The TGA also requires that the efficacy of each product is tested to determine the sun protection factor (SPF) which is printed on the label.

Question 2.11

Zinc oxide nanoparticles are commonly used in the manufacture of some sunscreen lotions. Describe two properties of these nanoparticles that make them ideal for this use.

Question 2.12

Stain resistant and water repelling fabrics are produced by thinly coating their surfaces with a polymer layer containing carbon nanotubes.

(a) Describe the kind of surface this creates at the nanoscale level.

(b) How does this surface prevent staining?

(c) Why is this feature often referred to as the lotus effect?

Question 2.13

A thin coating of titanium dioxide (TiO_2) nanoparticles applied to glass creates a surface which wets more easily and prevents the formation of water droplets. The coating may also be photo-catalytic.

Briefly explain how this helps the glass surface to be self-cleaning.

Question 2.14

Nanocomposite materials can be made stronger by the addition of very small amounts of suitable nanoparticles, such as nanotubes, or ceramic platelets. Briefly discuss two useful properties of the nanoparticles used and how these help in creating stronger materials.

REVIEW QUESTIONS

Chapter 2: Materials

1. Classify each of the following as either element, compound or mixture.

(a)	air	(b)	water	(c)	salty water
(d)	oxygen gas	(e)	brass	(f)	table salt

2. Mixtures can be homogeneous or heterogeneous. Give an example of each.

3. Mixtures can be separated by physical techniques such as decantation, filtration, crystallisation and distillation.

   Name the most suitable method to carry out the following separations.
   (a) sand and rock from a muddy slurry
   (b) alcohol from an alcohol/water solution
   (c) calcium carbonate (insoluble) from sodium chloride (soluble)
   (d) salt from sea water
   (e) water from sea water

4. Select from the following the most correct term for the descriptions that follow.

   evaporation, filtrate, element, solution, filtration, distillation, crystallisation, dissolution, mixture, decantation, solvent, residue

   (a) undissolved solid trapped by filter paper
   (b) the pouring off of a liquid, gently, from sediment
   (c) its constituents can be separated by simple physical techniques
   (d) a substance which cannot be split into simpler substances
   (e) clear solution remaining after filtration
   (f) the formation of crystals by cooling a saturated solution
   (g) technique for separating solids with very different solubilities
   (h) process by which a heated liquid changes to vapour
   (i) means by which alcohol is separated from wine
   (j) a homogeneous mixture

5. Using the values given in Figure 2.1 determine how many times larger the first mentioned particle is in each of the following pairs.
   (a) A buckyball compared to a hydrogen atom
   (b) A typical bacteria compared to the Adeno virus
   (c) A bee compared to a buckyball.

6. Nanoparticles are generally more reactive than the bulk materials of the same substance. Briefly explain the most likely reason for this.

7. The wavelength of visible light ranges from approximately 0.4 μm (violet light) to 0.7 μm (red light).
   (a) Convert these units to nanometres (nm).
   (b) Nanoparticles used in sunscreen lotions are typically, on average, 40 nm in size. Compare this size to that of the average wavelength of visible light, say 0.55 μm.
   (c) Explain why sunscreen lotions using nanoparticles are transparent.

8. The approximate diameters for three different particles are given below.

Sand grain particle	1 mm
Dust particle	1 µm
Carbon-60 buckyball	1 nm

(a) Assuming each of these particles to be spherical, calculate their surface area and volume and complete the table below.

For a sphere: Surface Area $= 4\pi r^2$ Volume $= 4/3\ \pi r^3$

PARTICLE	SURFACE AREA (m^2)	VOLUME (m^3)	SURFACE AREA / VOLUME RATIO
Sand grain			
Dust particle			
Carbon-60 buckyball			

(b) How does the surface area to volume ratio change as we consider smaller particles?

(c) Briefly explain why this is significant in terms of chemical properties.

9. Traditional sunscreens using bulk particles of zinc oxide appear white when applied to our skin. By comparison, sunscreen lotions using nanoparticles of zinc oxide are transparent when applied.

(a) What is the essential purpose of the sunscreens?
(b) Explain clearly why the sunscreens appear different
(c) The sunscreen containing nanoparticles is still effective even though transparent in appearance. Why is this?

10. Fabrics which are coated with a very thin polymer layer containing carbon nanotubes become stain resistant. Briefly answer the following.
(a) Describe the nature of the surface created by the polymer coating
(b) How does this help to make the fabric stain resistant?
(c) The look and feel of the fabric is not affected by the polymer coating. Why not?

11. There are a great variety of nanoparticles used in nanotechnology. Name the nanoparticle most associated with the following uses.
(a) Antibacterial bandages
(b) Carbon fibre tennis racquets
(c) Self-cleaning window spray
(d) Computer processor chip
(e) Auto sealants.

FOR THE EXPERTS

Some important substances used in the kitchen are as follows:

- **Table salt:** sodium chloride ($NaCl$). Used with food for seasoning and taste.
- **Sugar:** sucrose ($C_{12}H_{22}O_{11}$). Used as a sweetener in drinks and in the preparation of many foods particularly biscuits and cakes.
- **Baking soda:** sodium hydrogencarbonate powder ($NaHCO_3$). Also called carb-soda or sodium bicarbonate. An important ingredient of baking powder.
- **Baking powder:** a powder consisting of sodium hydrogencarbonate (baking soda), powdered acids such as tartaric acid and starch powder. Baking powder is a leavening agent. During baking it releases large amounts of carbon dioxide which are trapped in the flour and cause the mixture to rise.
- **Vinegar:** approximately 4% ethanoic (acetic) acid – (CH_3COOH), 96% water and very small amounts of natural colouring/flavours. Used mostly for seasoning foods.
- **Aluminium foil:** aluminium (Al). Used in cooking preparation and storage.

12. (a) Which substance/s listed above is/are:
    (i) pure substances
    (ii) mixtures
    (iii) elements
    (iv) compounds.

    (b) Some salt is accidentally dissolved in a bowl of water.
    (i) Is this a physical or chemical change?
    (ii) Suggest a method for recovering the salt.

    (c) If some baking soda is added to vinegar a fizzy reaction is observed. The word equation for the reaction is:

    sodium hydrogencarbonate + ethanoic acid → sodium ethanoate + carbon dioxide + water

    (i) Is this a physical or chemical change? Explain.
    (ii) Give a balanced equation for the reaction.

    (d) Baking powder can help to make cakes rise due to the reaction of two of its ingredients, baking soda and tartaric acid.
    (i) What is produced by the reaction of these two ingredients to make cakes rise?
    (ii) Is this a chemical or physical reaction?
    (iii) (Difficult) What prevents baking powder from reacting while stored for long periods in the pantry?

    (Clue – the powdered starch helps to prevent this but is itself not the reason.)

BONDING

3

Topics covered in this chapter:

3.1 BONDING – MAKING ATOMS STICK TOGETHER

All atoms try to gain maximum stability or lowest energy. The noble gases are very stable and our model of bonding assumes that atoms try to gain maximum stability by getting the same valence electron configuration as their nearest noble gas.

All noble gases have 8 electrons in their valence shell (2 for He). Atoms can achieve a stable octet in their valence shell by either losing their valence electrons, gaining extra valence electrons or sharing some of their valence electrons with another atom.

The three main types of chemical bonds between atoms are:

Metallic bonds

These form *between metal atoms*. The valence electrons of these atoms are very mobile or delocalised. Metallic bonds result from the attraction between these delocalised electrons and the positively charged metallic ions.

Ionic bonding

These form *between metal and non metal atoms*. A transfer of electrons between these atoms creates positive and negative ions, each having an inert gas electron configuration. Ionic bonds result from the strong attraction between these oppositely charged ions.

Covalent bonds

These form *between non metal atoms*. Electrons are shared in a common bond so that each atom can achieve an inert gas configuration. Covalent bonds result from the strong electrostatic attraction between the shared electrons and the protons of adjacent atoms.

3.2 IONIC BONDING

Ionic bonds occur between metals and non metals. Common table salt (NaCl) is a typical example of an ionic solid. It consists of sodium and chloride ions arranged in a regular lattice. These ions are formed by the *transfer of electrons* from the metal atom sodium, to the non metal atom chlorine.

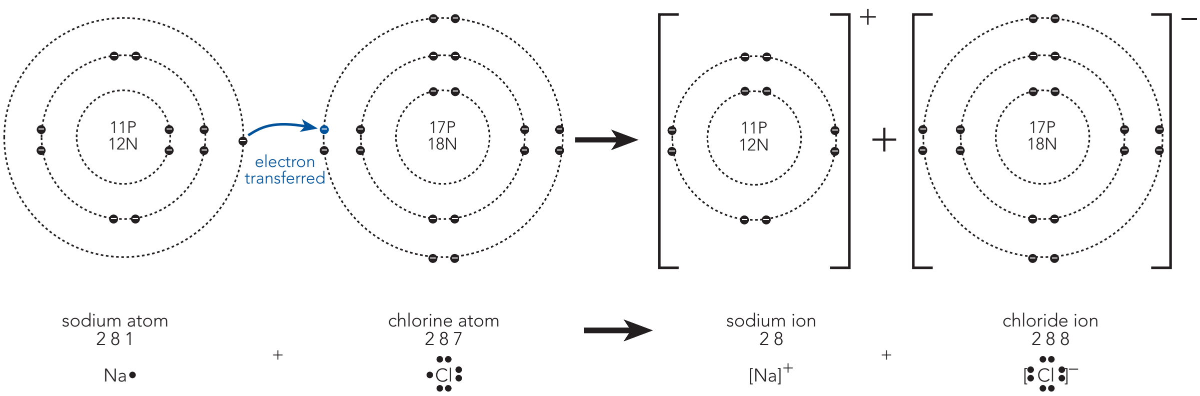


***Figure 3.1** Electron transfer between sodium and chlorine. Each atom achieves an inert gas structure.*

3.3 IONIC SOLIDS

Solid sodium chloride consists of an infinite array of sodium and chloride ions. They are held together by strong electrostatic attraction. Each Na^+ ion is surrounded by six Cl^- ions while each Cl^- ion is surrounded by six Na^+ ions.

The nature of the ionic lattice structure leads to the following physical properties:

- **high melting and boiling points**
  Large amounts of energy are needed to melt ionic solids like sodium chloride. This energy is required to overcome the strong electrostatic forces between the oppositely charged ions.

- **brittleness**
  Ionic solids are very hard, brittle and difficult to scratch due to the strong electrostatic forces. They are not malleable like metals because if layers of ions are forced to slide over each other repulsion occurs between like charges.

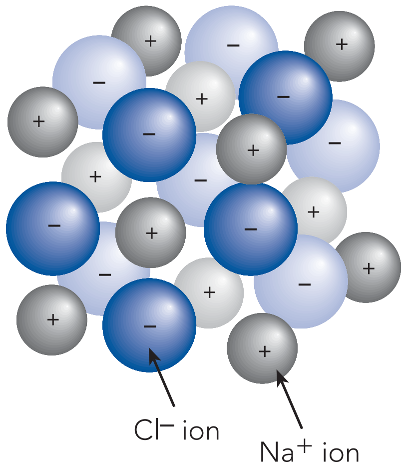


Figure 3.2 A crystal lattice of sodium chloride.

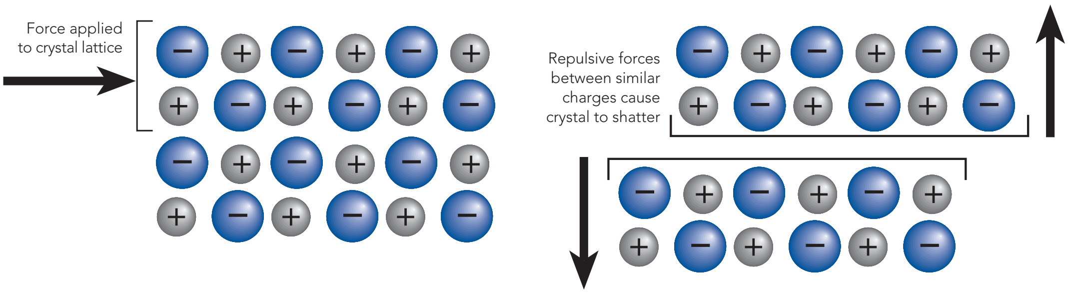


Figure 3.3 The brittleness of ionic solids.

- **good conductivity when molten or in aqueous solution**
  Ionic solids cannot conduct electricity since all the ions are in fixed positions and are not free to move. When molten however, the ions are mobile and will conduct a current. If the ionic solid is soluble in water they create a conducting solution since the ions are free to migrate through the water.

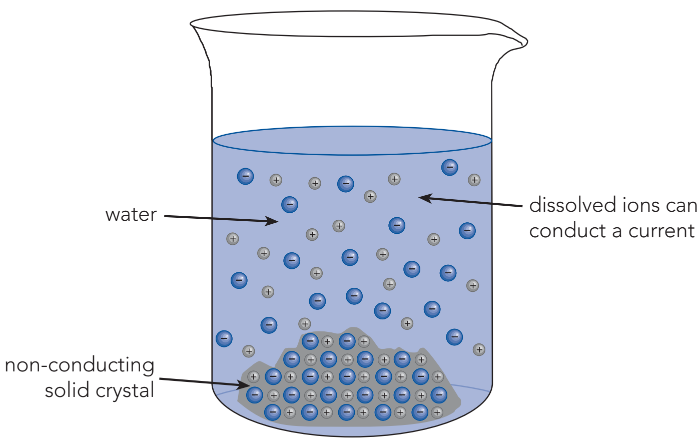


Figure 3.4 Ionic substances form conducting solutions

Question 3.1

Complete the following table:

ELEMENT	ELECTRON DOT DIAGRAM	TO FORM IONS THIS ATOM TENDS TO	ION FORMED	ION HAS SAME ELECTRON CONFIGURATION AS
sodium	Na •	lose 1 e^-	Na^+	Ne
magnesium				
nitrogen				
sulfur				
bromine				
calcium				
potassium				
lithium				

3.4 IONS

Ions form when atoms gain or lose electrons and achieve an inert gas structure (refer also to 1.11, page 18). Metals atoms generally have a low electronegativity and tend to lose electrons and form positive ions. Positive ions are also referred to as cations.

e.g. K^+ : potassium ion — It forms when a potassium atom loses 1 electron.
Mg^{2+} : magnesium ion — It forms when a magnesium atom loses 2 electrons.

Non-metal atoms generally have high electronegativity and tend to gain electrons and form negative ions. Negative ions are also referred to as anions.

e.g. Cl^- : chloride ion — It forms when a chlorine atom gains 1 electron.
S^{2-} : sulfide ion — It forms when a sulfur atom gains 2 electrons.

When a combination of several atoms forms an ion it is referred to as a polyatomic ion. The formula for a polyatomic ion indicates the number of atoms of each element present in the ion.

e.g. CO_3^{2-} : carbonate ion — This is a group of four atoms carrying an overall –2 charge.
NH_4^+ : ammonium ion — This is a group of five atoms carrying an overall +1 charge.

3.5 VALENCIES

The charge on an ion is called its valency. A table of common valencies is shown below:

Table 3.1 Valencies of some simple and polyatomic ions

POSITIVE IONS (CATIONS)		
1+	caesium	Cs^+
	hydrogen	H^+
	lithium	Li^+
	potassium	K^+
	rubidium	Rb^+
	silver	Ag^+
	sodium	Na^+
2+	barium	Ba^{2+}
	calcium	Ca^{2+}
	cobalt (II)	Co^{2+}
	copper (II)	Cu^{2+}
	iron (II)	Fe^{2+}
	lead (II)	Pb^{2+}
	magnesium	Mg^{2+}
	manganese (II)	Mn^{2+}
	nickel (II)	Ni^{2+}
	strontium	Sr^{2+}
	zinc	Zn^{2+}
+3	aluminium	Al^{3+}
	chromium (III)	Cr^{3+}
	iron(III)	Fe^{3+}
+4	lead (IV)	Pb^{4+}
	tin (IV)	Sn^{4+}
POLYATOMIC IONS		
+1	ammonium	NH_4^+

NEGATIVE IONS (ANIONS)		
−1	bromide	Br^-
	chloride	Cl^-
	fluoride	F^-
	iodide	I^-
−2	oxide	O^{2-}
	sulfide	S^{2-}
−3	nitride	N^{3-}
POLYATOMIC IONS		
−1	cyanide	CN^-
	dihydrogenphosphate	$H_2PO_4^-$
	ethanoate (acetate)	CH_3COO^-
	hydrogencarbonate	HCO_3^-
	hydrogensulfate	HSO_4^-
	hydroxide	OH^-
	nitrate	NO_3^-
	nitrite	NO_2^-
	permanganate	MnO_4^-
-2	carbonate	CO_3^{2-}
	chromate	CrO_4^{2-}
	dichromate	$Cr_2O_7^{2-}$
	hydrogenphosphate	HPO_4^{2-}
	oxalate	$C_2O_4^{2-}$
	sulfate	SO_4^{2-}
	sulfite	SO_3^{2-}
−3	phosphate	PO_4^{3-}

Question 3.2

Complete the following:

(a) Positive ions are also called ______________________________

(b) An example of a polyatomic ion with a positive valency is __________________

(c) An element that has both a positive and negative valency is ________________

Question 3.3

Complete the following:

(a) K • + :F̤̈ • ⟶ [ ] + [ ]

potassium atom + fluorine atom

2 , 8 , 8 , 1   2 , 7

(b) • Mg • + :S̈ • ⟶ [ ] + [ ]

magnesium atom + sulfur atom

2, 8, 2   2, 8, 6

Question 3.4

A crystal lattice of sodium chloride is made up of an equal number of sodium ions, Na^{+}, and chloride ions, Cl^{-} (See Figure 3.2). Determine how many:

(a) chloride ions surround each sodium ion ______________________

(b) sodium ions surround each chloride ion ______________________

Question 3.5

Explain the nature of the force holding an ionic crystal together.

____

Question 3.6

Explain why ionic substances such as magnesium chloride can conduct electricity in the molten state but not in the solid state.

____

____

Question 3.7

A sodium oxide crystal is formed from the reaction of sodium metal with oxygen gas.

(a) Which atoms lose electrons in this process? ______________________

(b) Name the ions which exist within this crystal. ______________________

(c) What is the ratio of positive to negative ions in this crystal? ______________

(d) Is the crystal formed positive, negative or neutral? ______________________

3.6 METALLIC BONDING

Metals are characterised by low ionisation energy which indicates that their valence electrons are not strongly held. In metallic structures these electrons become delocalised and not attached to any particular atom. This results in the atoms achieving a stable octet structure, becoming positive ions and attracted to the mobile delocalised electrons.

Metallic bonds, then, are due to the mutual attraction between the delocalised valence electrons and the positively charged metal ions (cations).

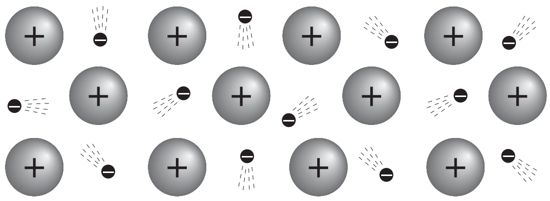


Figure 3.5 Metallic bonding. Metals consist of positive metallic ions surrounded by a sea of delocalised mobile electrons.

3.7 PROPERTIES OF METALS EXPLAINED

The properties of metals can be generally explained by the metallic bonding model outlined above. These properties include:

- **high electrical conductivity**
  Since the valence electrons in metals are highly mobile, any applied voltage will cause a flow of charge.

- **high thermal conductivity**
  When a substance is heated the particles vibrate more rapidly. In a metal lattice the delocalised electrons readily transfer this energy as they move through the lattice.

- **malleable and ductile**
  This means that metals can be hammered into sheets or drawn into wires. This is possible since metallic bonds are non-directional and layers of positive ions can simply slip over each other. The electrostatic forces between the positive ions and valence electrons still operate.

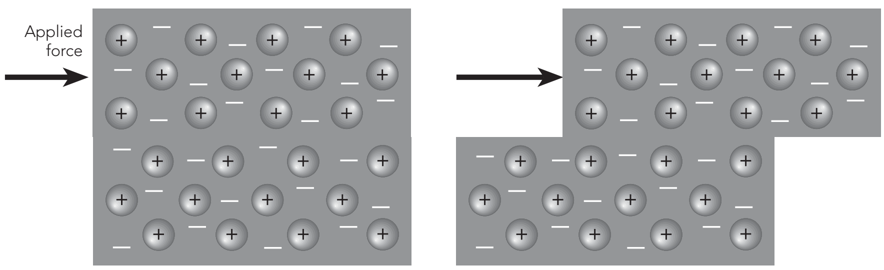


Figure 3.6 Malleability of metals. Metals can be bent, shaped and drawn into wires. Layers of atoms can slide past each other without breaking the metallic bond due to its non-directional nature.

- **wide range of melting points and boiling points**
  The strong electrostatic bonding means that melting and boiling points are high. The greater the number of valence electrons the stronger the bond. Hence M.P. and B.P. are higher for (say) group 2 elements than group 1 elements.

- **relatively high density**
  Strong electrostatic bonding and close packing of the ions means that metals are generally dense. This increases with the number of valence electrons per atom.

Question 3.8

How can metal atoms achieve a full outer energy level?

Question 3.9

In metallic bonding the atoms become stable by becoming positive ions. How are they able to form a strongly bonding crystal lattice?

Question 3.10

Explain in terms of their bonding structure why:

(a) metals are good conductors of electricity.

(b) metals are malleable and ductile.

Question 3.11

Sodium and magnesium are both metals and have similar sized atoms. However magnesium has a much higher melting point (650°C) than sodium (98°C).

Explain this difference in terms of the metallic bonding model.

Question 3.12

Alloy metals are formed by melting together one metal (copper) with small amounts of another metal (zinc). The alloy formed (brass, in this case) is less malleable than the original copper.

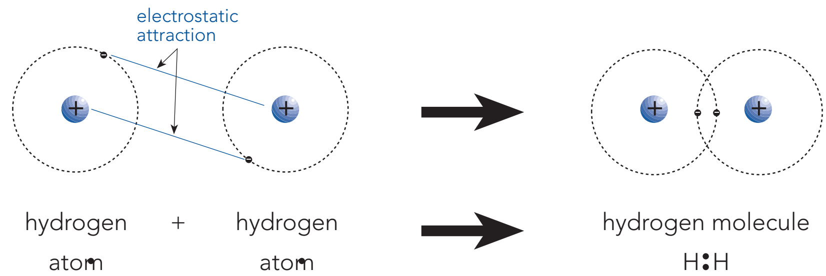
Suggest why this is so in terms of the metallic bonding model.

3.8 COVALENT BONDING

Covalent bonds form between non metal atoms. That is, they form between atoms having similar, but high electronegativity. This means that neither atom can gain electrons from the other but instead the electrons are shared. In this way each atom is able to achieve a stable octet configuration.

A covalent bond consists of a shared pair of electrons. It results from the strong electrostatic attraction between the shared electrons and the positive nuclei of each atom.

e.g. 1 Hydrogen molecule



e.g. 2 Oxygen molecule

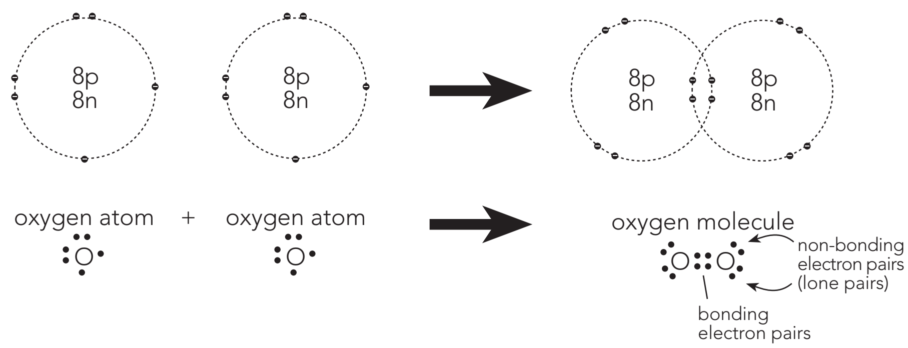
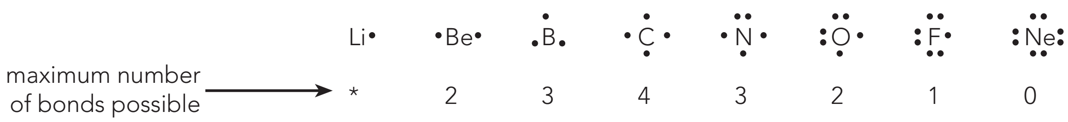


Figure 3.7 Formation of covalent bonds. By sharing electrons the atoms bond and reach a more stable state.

3.9 ELECTRON DOT DIAGRAMS FOR MOLECULES

An electron dot diagram shows how the valence electrons of an atom are distributed during bonding. The electrons available are arranged in such a way that each atom has a share of eight valence electrons. This is known as the octet rule. Hydrogen atoms, of course, share in only two electrons. The number of bonds an atom can form depends upon the number of valence electrons it has. The valence electrons for the period 2 elements are as follows.



*Lithium cannot form covalent bonds

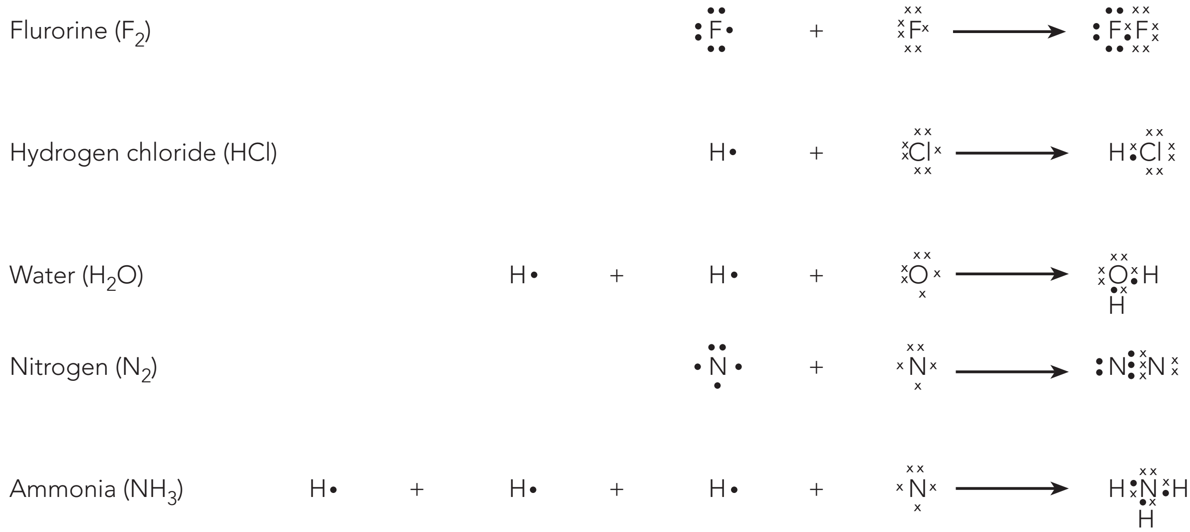
When drawing electron dot diagrams for molecules the main steps are as follows.

- Draw an electron dot diagram for each atoms first.
- Bonds can form between unpaired electrons (bonding electrons).
- Where more than two atoms are involved, the atom with the greatest number of unpaired electrons becomes the central atom.
- When complete, each atom is surrounded by an octet* of electrons and all electrons are shown as pairs.

** except for hydrogen which has 2, and some non octet molecules*

Worked Example

5.1 Electron dot diagrams for some simple molecules are shown below:



Note that electrons can be represented by either dots (•) or crosses (×) when doing electron dot diagrams. This does not mean that there are different types of electrons but it helps to identify from which atom they came.

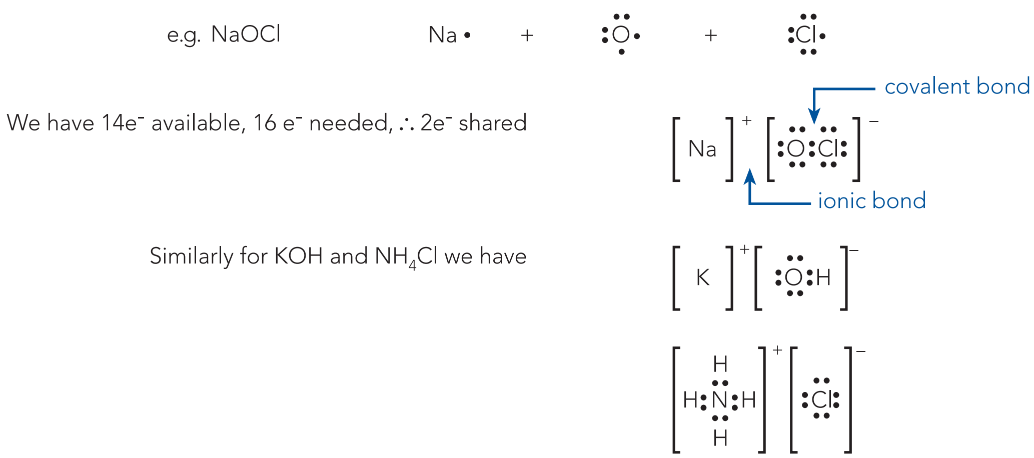
3.10 NON OCTET MOLECULES

Although very useful, the octet rule does have limitations and does not always apply. Some important exceptions to the octet rule are:

- Be and B being small atoms form molecules with only 2 and 3 electrons pairs respectively. e.g. $BeCl_2$, BF_3.
- larger atoms such as P and S can accommodate 5 or 6 bonding pairs. e.g. PCl_5, SF_6.

3.11 COMPOUNDS WITH BOTH IONIC AND COVALENT BONDS

Compounds containing polyatomic ions exhibit both ionic and covalent bonding. Three typical examples are KOH, NH_4Cl and NaOCl. The electron dot diagrams for the polyatomic ions are determined by the same procedure shown before.



Question 3.13

Why do non metal atoms share pairs of valence electrons?

Question 3.14

Explain why elements in group 15 such as nitrogen, can form three covalent bonds but elements in group 17, such as fluorine, can only form one.

Question 3.15

Draw the electron dot diagrams for the following molecules:

(a) Cl_2 (b) Br_2

(c) HCN (d) CO_2

3.12 COVALENT MOLECULAR SUBSTANCES

Atoms making up a molecule such as chlorine gas are very strongly held together by a covalent bond. However the force of attraction between two neighbouring chlorine molecules (intermolecular force) is very weak.

This explains why many molecular substances such as hydrogen, nitrogen and carbon dioxide are gases at normal temperature.

Intermolecular forces vary greatly between different molecules due to such factors as molecular size, shape or polarity. These factors will be dealt with in greater detail in Chapter 7.

Although there is a great variety in the properties of molecular substances, we can say that in general they have:

- low melting / boiling points
- do not conduct electricity (solid or liquid)
- are soft and easily scratched
- low solubility, forming non conducting solutions (except where molecules react with water forming ions and a conducting solution).

3.13 COVALENT NETWORK LATTICES

Substances such as diamond (C) and silica (SiO_2), consist of a three dimensional network structure wherein all the atoms are linked together by strong covalent bonds. These covalent bonds are continuous throughout the structure and result in very hard substances with high boiling points. Covalent lattices may either be:

- **three dimensional network solids**

e.g.	diamond	C	• very hard and brittle
	silicon	Si	• non-conductors
	silicon dioxide	SiO_2	• very high m.p. & b.p.
	silicon carbide	SiC	• inert, insoluble

- **two dimensional networks solids**

e.g.	graphite	C	• soft and slippery
			• very high m.p. & b.p.
			• conduct electricity
			• inert, insoluble

The properties listed above are the result of the strong covalent bonding giving strength and high melting points. The highly localised electrons prevent the conduction of electricity except for graphite which has one delocalised electron for each carbon atom.

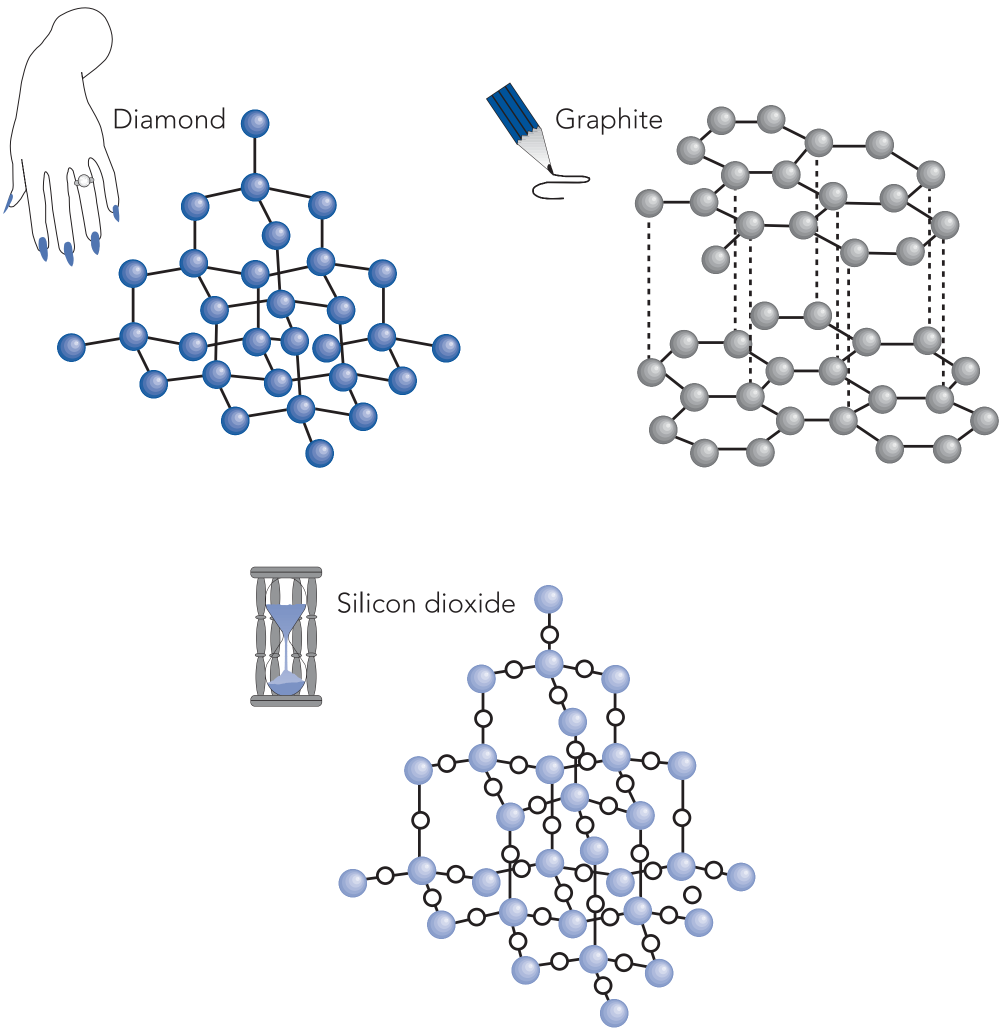


Figure 3.8 The structure of covalent network solids.

3.14 ALLOTROPES OF CARBON

There are a number of elemental forms, or allotropes, of carbon. Each has a different structure, resulting in markedly different properties. *Diamond* and *graphite* are two of the most common allotropes of carbon and their properties have already been outlined.

Another form, *amorphous carbon*, is made up of very small particles of graphite. Soot, lampblack and charcoal are examples of amorphous carbon. Important uses of this form of carbon include activated carbon filters, printing ink, toner cartridges and an additive for rubber products.

A more recently discovered allotrope of carbon, the *fullerenes*, are made up of cage-like molecules. These molecules, synthesised for the first time in 1985, consist of a large number of carbon atoms covalently bonded together as a series of pentagons and hexagons. They usually form a spherical or cylindrical hollow shape.

The most stable and interesting fullerene is the C_{60} molecule, which is made up of 60 carbon atoms arranged as twelve pentagons, each surrounded by a hexagon. Each carbon atom is covalently bonded to three other atoms. The result is a strong structure with one delocalised electron for each atom in the molecule. This feature allows fullerenes, like graphite, to conduct electricity.

Fullerenes were named after the American architect, Buckminster Fuller, who was noted for his design of geodesic domes of high stability. The C_{60} molecule was given the name buckminster fullerene but is usually referred to as *buckyball* for short. Other stable caged molecules of carbon, some with over 600 atoms, have also being found to exist or manufactured. Their shapes vary widely and include the C_{70} molecule shaped like a rugby ball, hemispheres and cylindrical nanotubes.

Carbon nanotubes have an interesting structure, best described as a graphite layer rolled into a tube. They are just over one nanometre in diameter but can be very long. They can be manufactured for specific purposes, be closed or open ended and designed to carry other molecules such as gases or drugs. Nanotubes are very strong and lightweight and can act as conductors or insulators depending on design. Their many uses include fibres for nanocomposite materials, medical applications, water purification, special surfaces and computer technology. *(See also previous chapter on materials).*

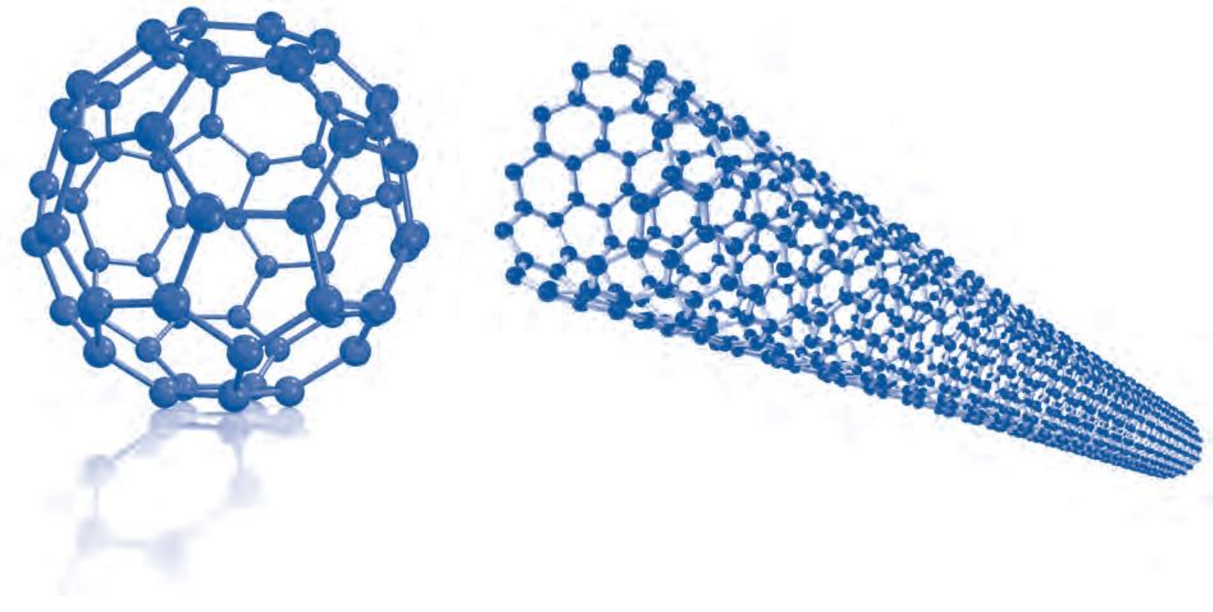


Figure 3.9 *Two common fullerenes, the C_{60} molecule, left, and a carbon nanotube. In both cases each carbon atom is covalently bonded to three other carbon atoms. There is one delocalised electron for each atom in their molecules. The C_{60} buckyball has a pattern of alternating pentagons and hexagons similar to a soccer ball. The nanotube is made up only of hexagons and can be considered as a graphite sheet rolled into a cylinder.*

3.15 THE STRUCTURE OF SOLIDS – SUMMARY

The majority of solids are crystalline substances, that is, they consist of particles that are arranged in some regular pattern or lattice.

The properties of the solid depend upon:

- the types of particles within the lattice.
- the way the particles are arranged.
- the nature of the forces between the particles.

In this chapter we have dealt with the four types of bonding structures. These are metallic solids, ionic solids, covalent molecular solids, covalent network solids. Their structure and properties are summarised below in Table 3.2:

Table 3.2 The four types of bonding structures.

BONDING STRUCTURE	METALLIC	IONIC	COVALENT MOLECULAR*		COVALENT NETWORK	
			Polar*	Non-polar*	3-d lattice	2-d lattice
Examples	Mg , Al ,Au	NaCl , CuO , $Mg(NO_3)_2$	NH_3 , H_2O , H_2S , CH_3Cl	He , H_2 , O_2 , CO_2	diamond , SiO_2 , SiC	graphite
Particles within the lattice	positive metal ions in a 'sea of electrons'	positive and negative ions	polar molecules	non-polar molecules	non-metal atoms	non-metal atoms
Forces between particles	strong electrostatic attraction	strong electrostatic attraction	dispersion forces, dipole-dipole forces & hydrogen bonding	dispersion forces only	strong covalent bonds	strong covalent bonds *within* layers - weak dispersion forces *between* layers
Melting point	mostly high	high	medium/ low	low	very high	very high
Electrical conductivity	solid - yes liquid - yes	solid - no liquid - yes solution - yes	solid - no liquid - no	solid - no liquid - no	no	yes
Solubility	no	mostly soluble in polar solvents	mostly soluble in polar solvents	mostly soluble in non-polar solvents	no	no
Hardness and malleability	malleable ductile	hard brittle	soft or slightly brittle	mostly soft	very hard brittle	soft brittle

** Covalent molecular substances and intermolecular forces will be treated in more detail in Chapter 7.*

Question 3.16

Carbon dioxide and chlorine are typical covalent molecular substances. Explain why

(a) they do not conduct electricity.

(b) they exist as gases at normal everyday temperatures.

Question 3.17

Silicon dioxide and diamond are typical covalent network substances. Explain why:

(a) they do not conduct electricity.

(b) they exist as solids at normal everyday temperatures.

Question 3.18

Most covalent network solids cannot conduct electricity. However graphite, a two dimensional network solid, can conduct electricity. Explain how this is possible. (Hint: the structure of graphite shown in Figure 3.8 may help you.)

Question 3.19

Fullerenes are a recently discovered allotrope of carbon.

(a) Explain what is meant by an allotrope.

(b) Name three other allotropes of carbon.

3.16 WRITING CORRECT FORMULA

Covalent molecular substances

These are composed of non-metal atoms covalently bonded to other non-metal atoms. The formula and name indicates the number and type of atoms present in a molecule of that substance.

Rules for naming

- The element that is closer to the left side of the periodic table is named first and it keeps its normal name. (If both elements are in the same group – the element further towards the bottom is named first.)
- The element closer to the right hand side of the periodic table is named second and has the end of its name changed to end in – ide.
- Prefixes are used to indicate if more than one atom of that element is present in the molecule.
- If there is only one atom of the first element, the prefix mono is not used.
- Molecules composed of one element only, get the name of that element.

mono	=	1
di	=	2
tri	=	3
tetra	=	4
penta	=	5
hexa	=	6
hepta	=	7
octa	=	8
nona	=	9
deca	=	10

e.g.

2 nitrogen atoms

N_2O_4 **di**nitrogen **tetra**oxide

4 oxide atoms

1 phosphorus atom

PCl_3 phosphorus **tri**chloride

3 chlorine atoms

Question 3.20

Name the following molecules that are made of non metal atoms only:

(a) CO ______________________ (b) CO_2 ______________________

(c) Br_2 ______________________ (d) P_2O_5 ______________________

(e) SO_2 ______________________ (f) SO_3 ______________________

Question 3.21

Write the correct formula for the following covalent molecules:

(a) chlorine _________ (b) nitrogen dioxide _________

(c) carbon tetrachloride _________ (d) sulfur trioxide _________

(e) oxygen dichloride _________ (f) dinitrogen pentaoxide _________

Ionic substances

These are composed of metal ions bonded to non-metal ions. These substances form immense lattices with countless numbers of ions. The formula of ionic substances provides the ratio of ions present, NOT the actual number of ions present.

Rules for writing formulae for ionic substances

- The positive ion (usually a metal ion) is written first.
- The negative ion (non-metal ion) is written second.
- The number of each ion is adjusted so that total positive charge equals the total negative charge. Subscripts are used to indicate the number of each ion present.

Worked Examples

3.2 Write the correct formula for calcium bromide.

Step 1 Write the formula for each ion, including valency. : Ca^{2+} Br^{-}

Step 2 Determine how many of each ion is needed to get the total positive charge to balance the total negative charge. : 1 Ca^{2+} ion balances 2 Br^{-} ions

Step 3 Re-write this information so that the numbers of each ion needed are written as subscripts. (Note: the number 1 is not written.) : $CaBr_2$

3.3 Write the correct formula for aluminium nitrate.

Step 1 Al^{3+} NO_3^{-}

Step 2 1 Al^{3+} ion balances 3 NO_3^{-} ions

Step 3 Correct formula is $Al(NO_3)_3$

3.3 – a quick method

Write the correct formula for aluminium oxide.

As usual, firstly write the formula of each ion. A quick method is to then simply cross (swap) the valency numbers.

Step 1 Al^{3+} O^{2-}

Step 2 Al O gives Al_2O_3

Step 3 Correct formula is Al_2O_3

Question 3.22

Combine the following ions to give the correct formula:

	Cl^{-}	O^{2-}	N^{3-}	OH^{-}	SO_4^{2-}
Na^{+}					
Mg^{2+}					
Fe^{3+}					

Rules for naming ionic compounds

- The metal, or positive ion is named first and it keeps its normal name.
- The non-metal, or negative ion is named second and has the end of its name written as … ide, . … ate or … ite (refer to valency table of ions)
- If the metal ion has several valencies possible, indicate its valency using roman numerals.

Worked Example

3.4 Name the following ionic compounds:

a) $AlCl_3$ Correct name is aluminium chloride

b) $Mg_3(PO_4)_2$ Correct name is magnesium phosphate

c) CuO Correct name is copper (II) oxide

d) FeO Correct name is iron (II) oxide

e) Fe_2O_3 Correct name is iron (III) oxide
(iron has valencies of +2 and +3; this compound contains Fe^{3+})

Question 3.23

Write the correct formula for each of the following ionic compounds.

barium chloride		zinc oxide	
aluminium bromide		zinc sulfate	
aluminium nitride		barium ethanoate	
caesium sulfide		calcium hydroxide	
copper (II) sulfite		copper (II) hydroxide	
zinc phosphate		ammonium chloride	
cobalt fluoride		ammonium hydrogensulfate	
lead (II) oxide		chromium (III) oxide	

Question 3.24

Name the following ionic compounds.

(a) $NaBr$ ____________ (b) $Fe(NO_3)_3$ ____________

(c) $FeSO_4$ ____________ (d) $(NH_4)_2SO_4$ ____________

(e) BaO ____________ (f) $CuCO_3$ ____________

(g) $KHCO_3$ ____________ (h) $Cu(OH)_2$ ____________

(i) $FeCO_3$ ____________ (j) $Ba_3(PO_4)_2$ ____________

REVIEW QUESTIONS

Chapter 3: Bonding

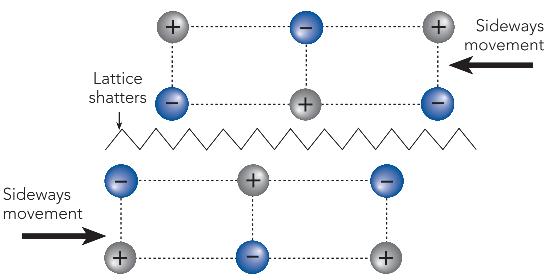
1. State which of the following elements; Mg, N, Cl, S, Li, Ca, F tend to get a full outer energy level by:

   (a) losing electrons
   (b) gaining electrons
   (c) sharing electrons

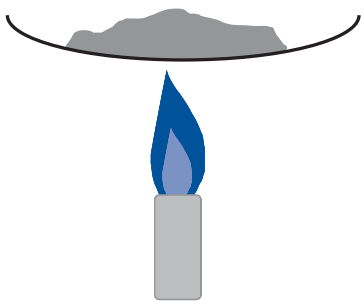
2. Why do the elements belonging to group I and II of the periodic table tend to form positive ions?

3. Each of the following diagrams illustrate a property typical of ionic substances. Use the nature of the ionic bond to explain the reason for each property.

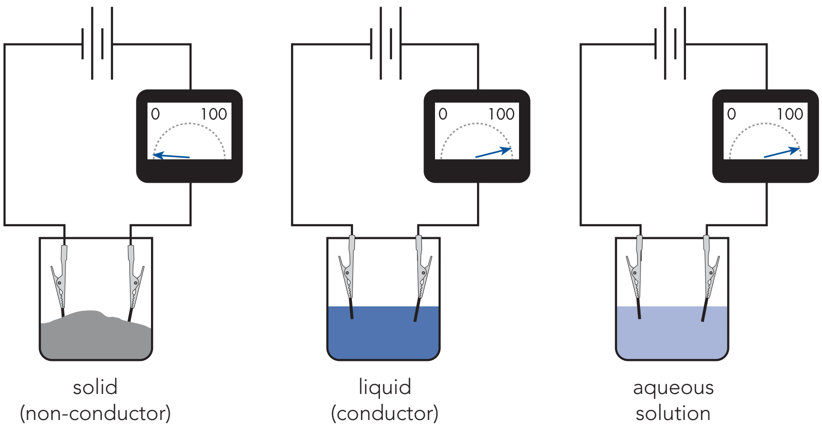
   (i) brittle



   (ii) high melting point



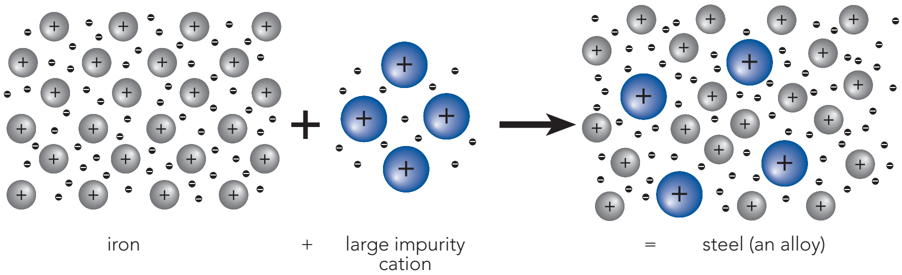
   (iii) conduct electricity only in solution or as a liquid



4. Explain why metals are such good conductors of electricity.

5. Potassium and calcium are both metallic elements made up of similar sized atoms. Which would you expect to have a higher melting point? In terms of metallic bonding, explain why.

6. Steel is an alloy formed by melting together iron with small amounts of other elements. The following diagram shows the formation of a steel alloy which results from the addition of larger 'inpurity' atoms.



Referring to the diagram above explain why such an alloy as that shown is likely to be much less malleable than pure iron.

7. A chlorine molecule forms when two chlorine atoms are covalently bonding together.
   (a) What force actually holds the atoms together?
   (b) How many electrons are shared?
   (c) How many non-bonded pairs of electrons are there?

8. Draw electron dot diagrams for the following molecules:
   (a) I_2 (b) PCl_3 (c) CH_4 (d) C_2H_2

9. Some molecules, such as BF_3 and PCl_5, do not conform to the octet rule. Draw their electron dot diagrams.

10. Draw the electron dot diagrams for the following polyatomic ions:
    (a) SO_4^{2-} (b) NO_3^- (c) H_3O^+

11. For each of the following compounds name the type of bonding involved:
    (a) NaCl (d) HNO_3
    (b) Pb (e) SO_2
    (c) $PbCl_2$ (f) NH_4NO_3

12. Graphite and diamond are allotropes of carbon.

Properties of diamond
Very hard, brittle, has a very high melting point and is a poor conductor of electricity.

Properties of graphite
Soft, excellent lubricant, very high melting point and is a good conductor of electricity.

(a) Briefly explain why both diamond and graphite have very high melting points.
(b) Explain why these two forms of carbon have such differing properties of electrical conductivity and hardness.
(c) Use its structure and bonding to explain why graphite is able to act as a lubricant.

13. Classify the following substances by placing them into their correct classification:

carbon dioxide, silicon dioxide, magnesium oxide, dinitrogen tetraoxide, iron, lithium bromide, mercury, diamond (carbon), fluorine, lead (II) nitrate

IONIC	METALLIC	COVALENT MOLECULAR	COVALENT NETWORK

14. There are many different oxides of nitrogen. Give the correct name for each of the following:

(a) NO (b) NO_2 (c) N_2O

(d) N_2O_3 (e) N_2O_4 (f) N_2O_5

15. Write the correct formula for:

(a) phosphorus pentachloride

(b) diphosphorus pentaoxide

(c) nitrogen trifluoride

(d) carbon tetrafluoride

(e) sulfur dichloride

(f) dihydrogen monoxide

16. Write the correct formula for each of the following:

(a) copper (II) sulfite

(b) cobalt (II) nitrate

(c) manganese (II) oxide

(d) chromium (III) chloride

(e) aluminium phosphate

(f) calcium hydrogenphosphate

(g) sodium dihydrogenphosphate

(h) sodium hydrogenphosphate

(i) iron (II) hydroxide

(j) potassium permanganate

(k) iron (II) oxide

(l) sodium oxalate.

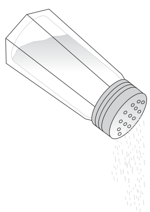
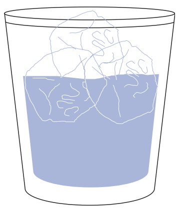
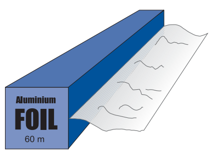
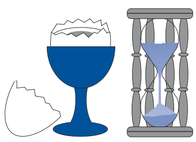
FOR THE EXPERTS

17. Many substances found commonly in the kitchen include:

- table salt (sodium chloride)
- ice (frozen water)
- aluminum
- sand (silicon dioxide)

These are all crystalline substances. However the bonding structures within them vary greatly.

Complete the following table to show the differences in their bonding structures.

SUBSTANCE	TYPE OF LATTICE STRUCTURE	PARTICLES WITHIN THE LATTICE	SIMPLE DIAGRAM OF LATTICE
salt 			
ice 			
aluminium 			
sand 			

CARBON CHEMISTRY

4

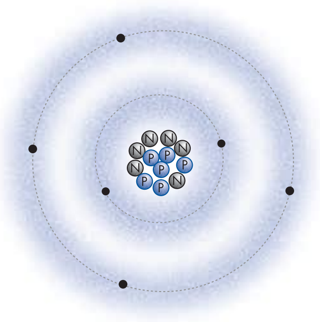
Topics covered in this chapter:

4.1 COVALENT BONDING AND CARBON

Diamond (C)

CO_2

Petrol (C_7H_{16} and C_8H_{18})



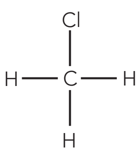
There are literally millions and millions of compounds that have carbon as their basic building block. Why is carbon able to form such a large and diverse range of compounds? The answer lies with the electron configuration and subsequent bonding capacity of carbon.

Electron configuration: 2 4

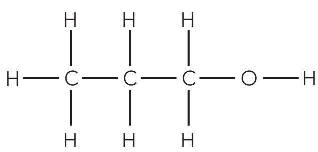
To get a full outer shell, a carbon atom can:

1. Form four single bonds with four other atoms.

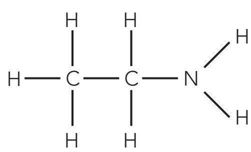
e.g.



chloromethane



propan-1-ol

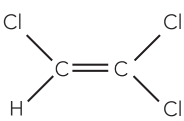


ethanamine

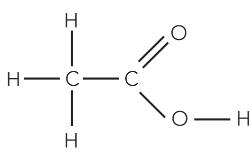
Carbon commonly forms single bonds with other carbon atoms, hydrogen, oxygen, nitrogen and the halogens.

2. Form two single bonds and a double bond with three other atoms.

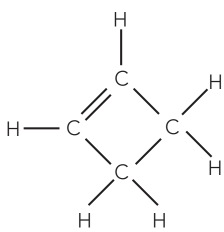
e.g.



trichloroethene



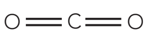
ethanoic acid



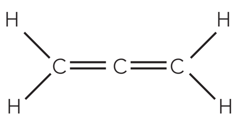
cyclobutene

3. Form two double bonds with two other atoms.

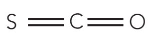
e.g.



carbon dioxide



propandiene

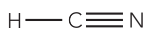


carbonyl sulfide

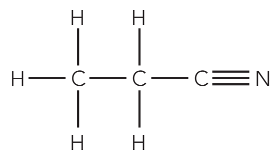
4. **Form a triple bond and a single bond with two other atoms.**

e.g.

ethyne



hydrogen cyanide



propannitrile

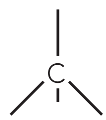
These four different methods of obeying the octet rule lead to carbon forming such a variety of compounds.

4.2 SHAPE OF CARBON COMPOUNDS

The shape of molecules will be covered in Chapter 7, however it is important to be aware of the shape of carbon compounds.

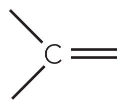
1. Carbon atom involved in four single bonds.

   Shape: tetrahedral Bond angle: 109.5°



2. Carbon atom involved in one double and two single bonds.

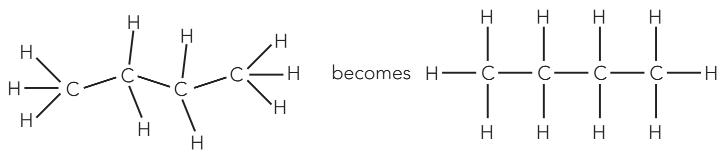
   Shape: triangular planar Bond angle: 120°

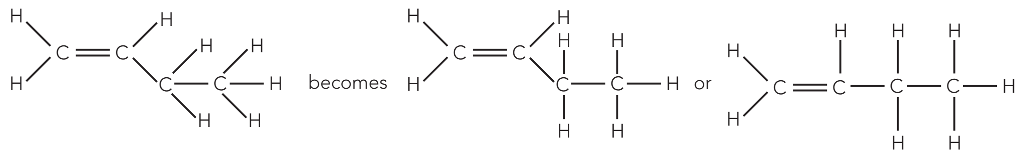


3. Carbon atom involved in one triple and one single bond.

   Shape: linear Bond angle: 180°

Representing three dimensional figures on two dimensional paper is sometimes confusing. The general procedure followed in drawing the 3-D shape on paper is to draw 109.5° bond angle as 90°. Two examples are given below:





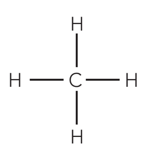
4.3 HYDROCARBONS

Hydrocarbons contain hydrogen and carbon only. They are classified according to the nature of the bonding between carbon atoms and named according to the number of carbon atoms in the longest chain.

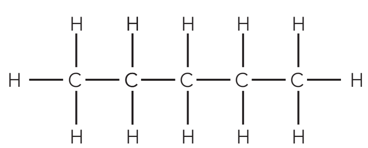
Saturated hydrocarbons

Alkanes – only single bonds exist between carbon atoms.

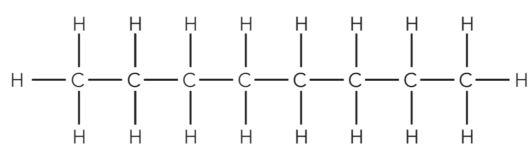
e.g.



methane



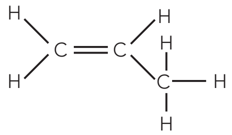
pentane



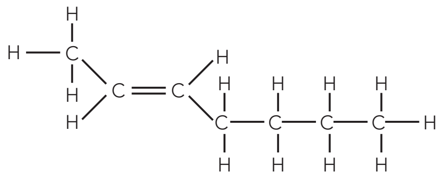
octane

Unsaturated hydrocarbons

Alkenes – carbon chain contains at least one double covalent bond between a pair of carbon atoms.

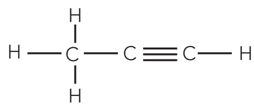


propene

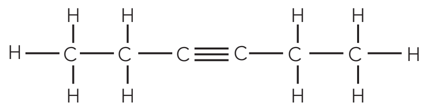


hept-2-ene

Alkynes – carbon chain contains at least one triple covalent bond between a pair of carbon atoms.



but-1-yne



hex-3-yne

Alkyl Groups – are monovalent groups derived from alkanes, i.e. they are alkanes that are missing a hydrogen atom. They attach to other carbon chains at the point where they are missing the hydrogen atom.

Question 4.1

Carbon has the electron structure 2, 4 or $1s^2$ $2s^2$ $2p^2$. How can carbon get the same electron configuration as its nearest noble gas (neon)?

____

Question 4.2

(a) Carbon often forms single covalent bonds with other carbon atoms, hydrogen

_______________, _______________, _______________.

(b) Carbon often forms double bonds with other ______ atoms and oxygen.

(c) Carbon often forms triple bonds with other _____ atoms and ____.

Naming Simple Organic Compounds using IUPAC Rules

1. Identify the longest continuous carbon chain. If the compound contains a double or triple bond the longest chain must contain this feature.
2. Number the chain from the end which gives the double or triple bond the lowest possible number. If there are no double or triple bonds, number so that alkyl groups or halides get the lowest possible number.
3. Number all attached groups using the order from #2.
4. Use the appropriate prefix or suffix to name the attached groups (alkyl groups and halides) and multiple bonds.
5. Use one word to name the compound. The name of each group is preceded by a numeral indicating its attachment position in the main chain. Alphabetical order is used when more than one group is involved.
6. Numerals are separated from words by a hyphen and from other numerals by a comma.

*NB: **Alkyl groups** are alkanes that are missing a hydrogen atom. They attach to other carbon chains at the point where they are missing the hydrogen atom.*

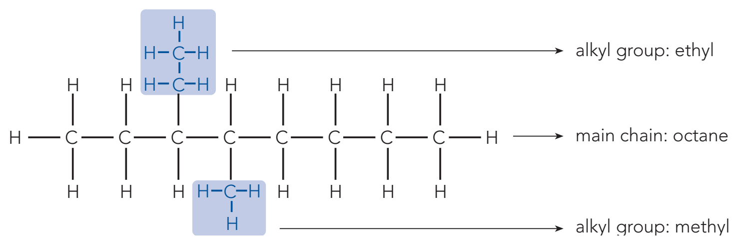
Question 4.3

Complete the following table to show the prefixes that are used for indicating the number of carbon atoms in a chain.

NUMBER OF C ATOMS IN CHAIN	ALKANE	ALKENE	ALKYNE	ALKYL GROUP
suffix	**–ane**	**–ene**	**–yne**	**–yl**
1	methane	–	–	methyl
2	ethane	ethene	ethyne	ethyl
3	propane			
4	butane			
5	pentane			
6	hexane			
7	heptane			
8	octane			

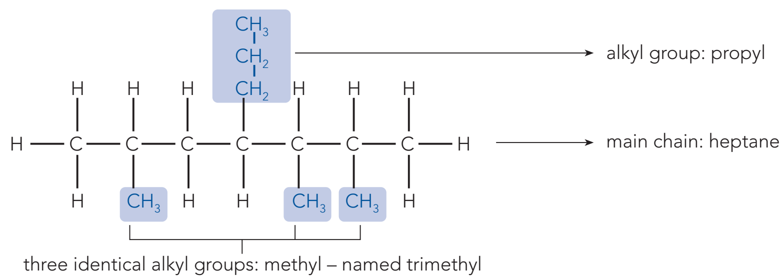
Examples of alkyl groups

e.g. 1 This molecule has two alkyl groups attached.



correct name : 3-ethyl-4-methyloctane

e.g. 2



correct name: 2,3,6-trimethyl-4-propylheptane

NOTE:

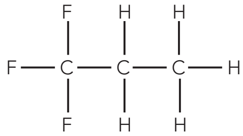
- When more than one of a functional group or alkyl group is attached to a carbon chain the following prefixes are used: 2 – di ; 3 – tri ; 4 – tetra ; 5 – penta ; 6 – hexa
- When these prefixes are used, the numbers indicating the position of each attachment must also be used.
- These prefixes (di, tri, tetra, etc) are not considered as names when determining alphabetical order.

4.4 HALOALKANES

A group of elements that are often part of organic molecules are the halogens. When a halogen is part of an organic molecule they are named using the following prefixes:

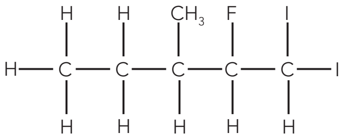
F	: fluoro ____________	Cl	: chloro ____________
Br	: bromo ____________	I	: iodo ____________

e.g. 1



1,1,1–trifluoropropane

e.g. 2

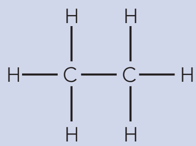


2–fluoro-1,1-diiodo-3-methylpentane

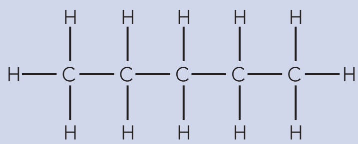
Question 4.4

Use IUPAC rules to name the following organic compounds:

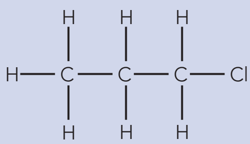
(a)



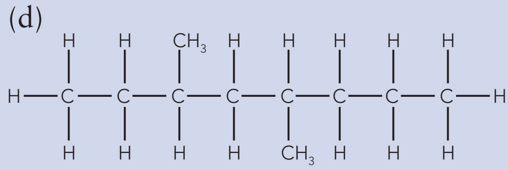
(b)



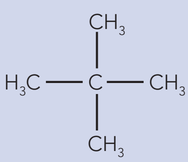
(c)



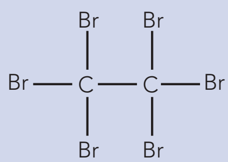
(d)



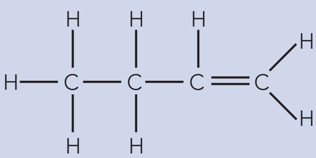
(e)



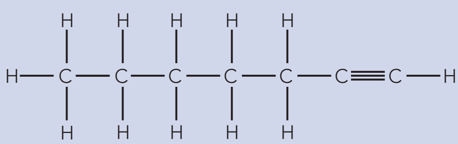
(f)



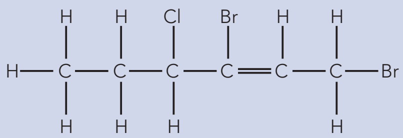
(g)



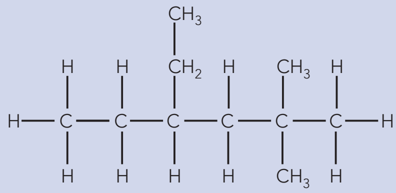
(h)



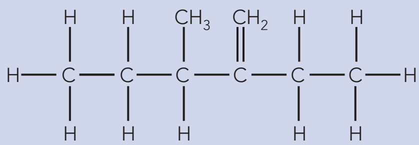
(i)



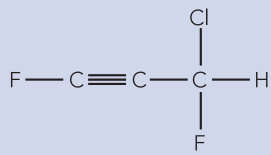
(j)



(k)



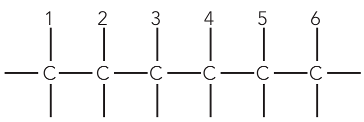
(l)

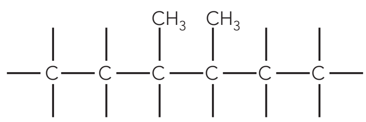


4.5 STRUCTURAL FORMULA

A structural formula gives a 2-dimensional representation of the organic molecule. It provides important information about where each different atom is attached to a carbon chain. A molecular formula provides the correct number of atoms of each element in a compound, a structural formula provides that **plus** the position of each atom of each element.

e.g. 1 Draw the structural formula for 3,4-dimethylhexane.

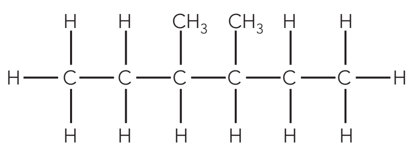




(i) Identify the longest chain by reading the end of the name – hexane.

(ii) Identify **what** is attached to the longest chain and **where** it is attached – two methyls at carbons numbered 3 & 4.

Complete the diagram by drawing the remaining hydrogens.



Question 4.5

Draw the structural formula of each of the following:

(a) 1,1,1-trichloroethane

(b) 5,6,6,7-tetramethyloct-3-ene

(c) 4-chloro-2,3-dimethyloctane (d) 3- chloro-2,3-difluorobut-1-ene

Question 4.6

Name the hydrocarbon family that would be represented by the following general formulae:

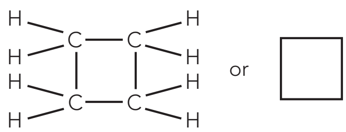
(a) C_nH_{2n} ______________________ (b) C_nH_{2n+2} ______________________

(c) C_nH_{2n-2} ______________________

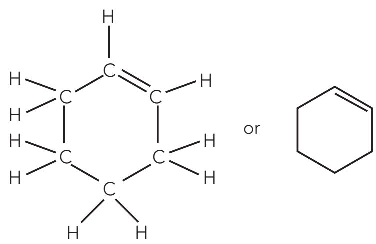
4.6 CYCLIC HYDROCARBONS

Carbon can form compounds in which the carbon atoms are in rings. These can be cycloalkanes or cycloalkenes. The names for cyclic hydrocarbons are the same as for chain hydrocarbons except for the addition of the prefix "cyclo".

e.g. 2 Structural formula for cyclobutane:



e.g. 3 Structural formula for cyclohexene:



NOTE: the double bond is not numbered when naming, it is assumed that it is given the lowest possible number, the next number would be the other side of the double bond.

Question 4.7

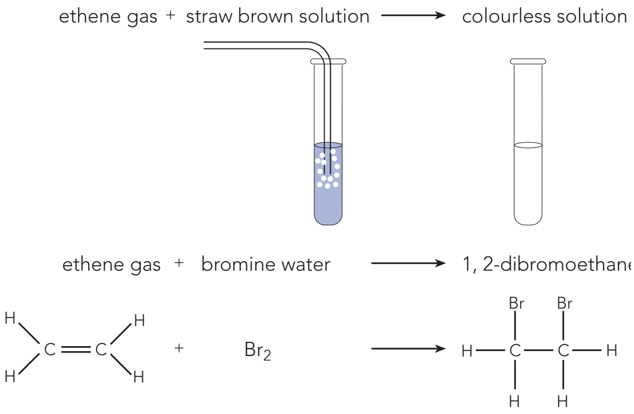
Draw the structural formula for each of the following:

(a) cyclopentene (b) cyclopropane

4.7 REACTIONS OF HYDROCARBONS

Addition reactions

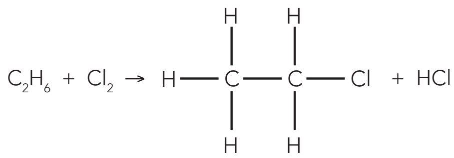
Addition reactions occur when alkenes and alkynes have their double or triple bond broken and elements such as halogens added into the carbon chain.



Substitution reactions

Substitution reactions occur when alkanes and aromatic compounds react such that a hydrogen atom is removed and another element such as a halogen is substituted into its position on the hydrocarbon chain.

e.g.



Alkanes and aromatic compounds are relatively inert compounds. Substitution reactions need to occur at moderately high temperatures (250°C) or with UV light added and in the presence of a catalyst.

Combustion

All hydrocarbon compounds can be burnt in oxygen. If complete combustion occurs, then the only products are carbon dioxide and water vapour. The production of energy by burning organic fuels (fossil fuels) is the major form of energy production on Earth.

Question 4.8

Write balanced equations for the following reactions and classify each as ADDITION, SUBSTITUTION, or COMBUSTION.

(a) $C_4H_{10} + F_2$ + UV light ____________________

Classification: ____________________

(b) $C_2H_4 + Cl_2$ ____________________

Classification: ____________________

(c) propene + H_2 ____________________

Classification: ____________________

(d) $C_7H_{16} + O_2$ ____________________

Classification: ____________________

(e) pentane + iodine water + UV light ____________________

Classification: ____________________

(f) $C_2H_2 + O_2$ ____________________

Classification: ____________________

Question 4.9

Provide experimental evidence that supports the hypothesis that "alkenes and alkynes are more reactive than similar sized alkanes because of the presence of the multiple bond".

___

___

Question 4.10

Complete the following table that highlights the major hydrocarbon constituents of the following fossil fuels:

FUEL	MAJOR CONSTITUENTS
Natural Gas	
Liquefied Petroleum Gas (LPG)	
Liquefied Natural Gas (LNG)	
Kerosene	
Petrol	

Question 4.11

State what reactants could be used to make the following compounds:

(a) 1,2-dichloropentane ___

(b) 2,3-dibromobutane ___

(c) 2-fluoroheptane ___

(d) 1,1,2,2-tetraiodoethane ___

Question 4.12

Write the equation for the following reactions. Show the structural formula of all organic compounds.

(a) ethene + Cl_2

___

(b) hexane + Br_2

___

(c) but-1-ene + HBr

___

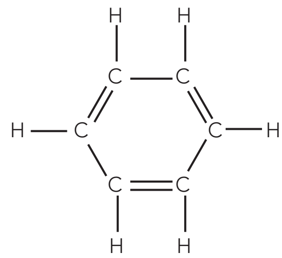
(d) $C_8H_{18} + O_2$

___

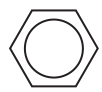
4.8 BENZENE – C_6H_6

Benzene is a cyclic compound having 6 carbon atoms in a ring and 3 double bonds within the ring.

e.g.



The behaviour of the double bonds is quite unique as compared to bonds in other hydrocarbons. The double bonds are shared equally over the size carbon atoms. The bonds are of equal strength and this makes the benzene molecule (or benzene ring) quite stable. Benzene is said to have two resonant structures:

 and  These are both possible and are usually represented as: 

Unlike other alkenes, benzene tends to undergo substitution rather than addition reactions. The substitution reactions require a catalyst and/or additional energy.

e.g. 1 $C_6H_6 + Br_2 \xrightarrow{\text{catalyst}} C_6H_5Br + HBr$
bromobenzene

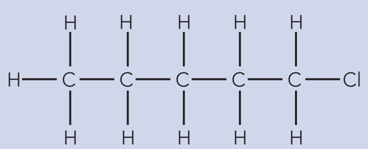
e.g. 2 $C_6H_6 + Cl_2 \xrightarrow{\text{catalyst}} C_6H_5Cl + HCl$
chlorobenzene

4.9 ISOMERISM

Structural isomers are molecules that have the SAME molecular formula but have different properties. They have different chemical properties because they have different arrangements of atoms on the molecule, i.e. they have different structures.

Question 4.13

5 isomers of the molecule $C_5H_{11}Cl$ are drawn below. Name them.

(a) 

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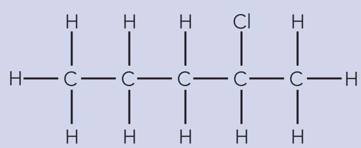
(b) 

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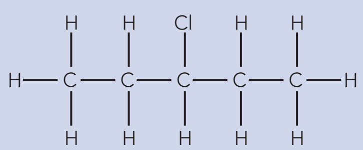
(c) 

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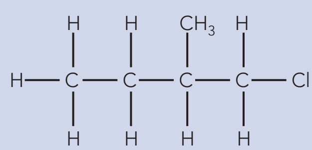
(d) 

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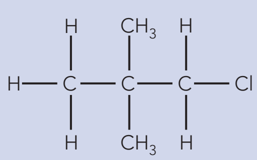
(e) 

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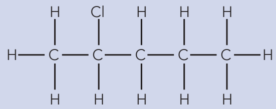
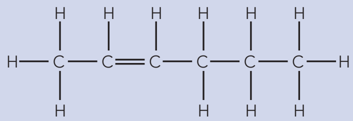
Question 4.14

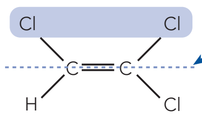
Is  another isomer of $C_5H_{11}Cl$? Explain your answer.

Question 4.15

Are  and  isomers? Explain.

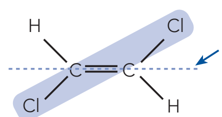
Geometric isomers occur only in alkenes. The existence of the double bond stops the rotation of the chain at that position and so molecules with the same structural formula can have different geometries and hence different properties.

e.g. 3 1,2-dichloroethene has two geometric isomers:



imaginary axis through double bond

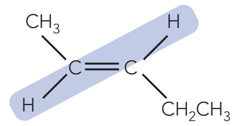
cis-1,2-dichloroethene
"cis" refers to the two Cl atoms being on the same side of the axis splitting the double bond



imaginary axis through double bond

trans-1,2-dichloroethene
"trans" refers to the two Cl atoms being across the axis splitting the double bond

e.g. 4 Name the following geometric isomer:

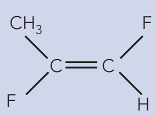


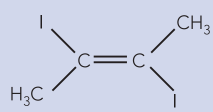
trans pent-2-ene

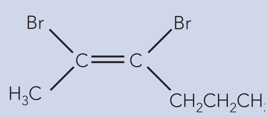
STEP 1: Name the compound as normal
STEP 2: Identify species that are on the same side of or across the double bond
STEP 3: Place the appropriate prefix (cis or trans) in front of the name.

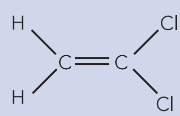
Question 4.16

Name the following compounds using the prefixes "cis" or "trans" where appropriate.

(a) 

(b) 

(c) 

(d) 

Question 4.17

Draw the cis and trans isomers of each of the following:

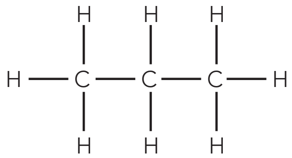
(a) hept-2-ene (b) 1,2-dichloropropene

REVIEW QUESTIONS

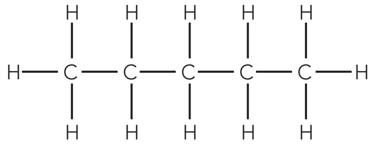
Chapter 4: Carbon Chemistry

1. Use IUPAC rules to name the following hydrocarbons:

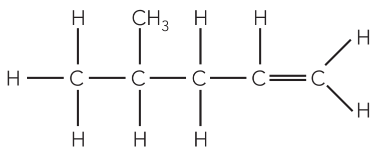
(a)



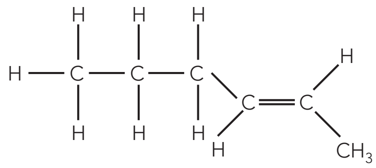
(b)



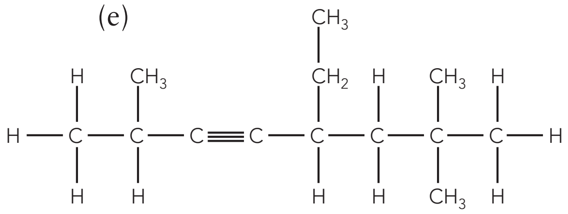
(c)



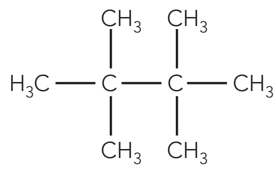
(d)



(e)

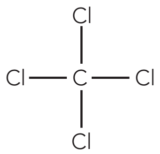


(f)

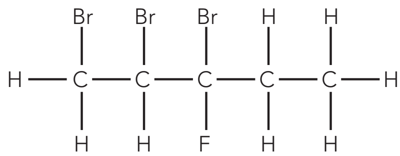


2. Name the following haloalkanes:

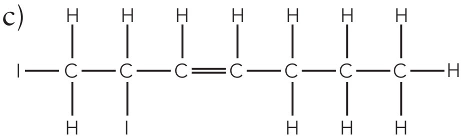
(a)



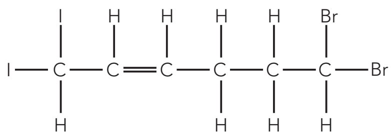
(b)



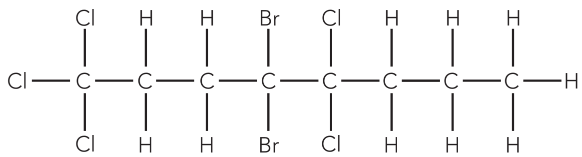
(c)



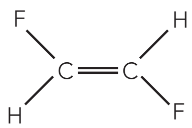
(d)



(e)

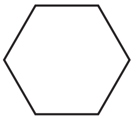


(f)

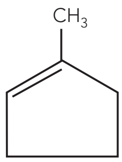


3. Name the following:

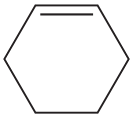
(a)



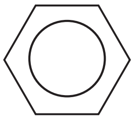
(b)

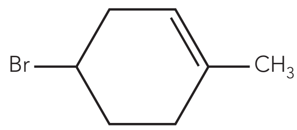


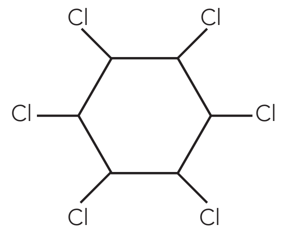
(c)

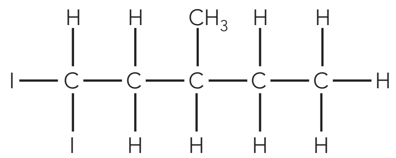


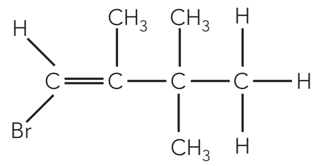
(d)

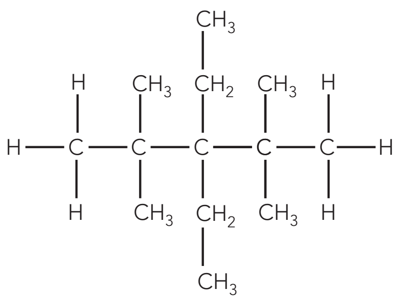


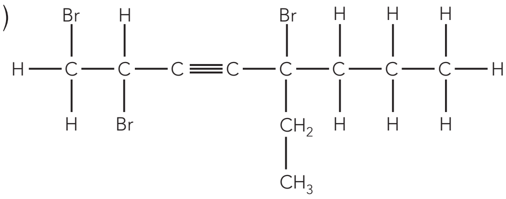
(e) 

(f) 

(g) 

(h) 

(i) 

(j) 

4. Draw the structural formula for each of the following compounds:

(a) 2,3-dimethylhexane
(b) 5-ethyl-1-chlorohept-3-ene
(c) 1,4-difluoro-2-methylcyclopentane
(d) 3,6-diiodocyclohexene
(e) 7-bromo-3,5-diethyloct-l-yne

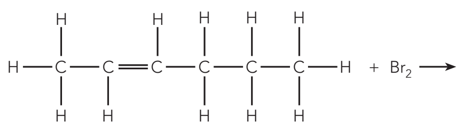
5. Name the products if the following react:

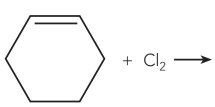
(a) 2,3-dimethylpentane + oxygen gas
(b) 1-butene + chlorine gas
(c) hex-2-yne + bromine water
(d) heptane + fluorine gas
(e) ethene + hydrogen chloride gas

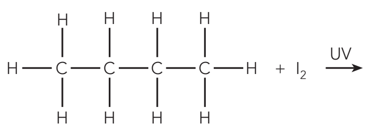
6. Write balanced equations for each of the following reactions:

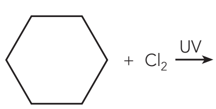
(a) $C_5H_{12} + O_2 \rightarrow$

(b) $C_8H_{16} + O_2 \rightarrow$

(c) 

(d) 

(e) 

(f) 

7. Name the reactants that could be used to produce:

   (a) 1,2-dibromocycloheptane only
   (b) HCl and 2-chloropentane
   (c) 2-bromopropane only
   (d) 2-bromopropane and one other product
   (e) 1,1,2,2-tetraiodoethane

8. (a) Three common fossil fuels are methane, butane and octane.
   Write a balanced equation for the complete combustion of each of these fuels.
   (b) What is common to each of these three reactions?
   (c) Which of the three fuels requires the most oxygen per atoms of carbon involved?

9. Each of the following is incorrectly named. Draw them and give the correct IUPAC name OR state what error has been made.

   (a) hex-4-ene
   (b) 1,1,1-trichloromethane
   (c) 2-butylpent-1-ene
   (d) 1,1-dichlorobut-1-yne
   (e) 2,2-dichlorocyclohex-1-ene
   (f) 2,3-difluorocyclopentane
   (g) cis 1,1-dichloroethene
   (h) 4,6-diiodocyclohexene

10. Draw the following geometric isomers:

   (a) cis 1,2-diiodoethene
   (b) trans hex-3-ene
   (c) cis chloro-1,2-difluoroethene
   (d) trans 4,5-dimethyloct-4-ene

11. 1,1,1-trichloroethane is a common solvent used in a range of products including some brands of correction fluid.

   (a) Draw the structural formula for the compound.
   (b) It can be produced via a three stage process with stage one involving an addition reaction and stages 2 and 3 as substitution reactions.
   Write possible reactions for formation of 1,1,1-trichloroethane.

FOR THE EXPERTS

NOTE: The organic compounds mentioned in the following information are not in your syllabus. However, you are expected to apply your knowledge of the syllabus to suggest possible answers to each question.

12. When petrol is ignited in the cylinder of a car engine, there is a rapid rise in pressure. This pressure pushes the cylinder downwards. Ideally the combustion process should produce a smooth increase in the rise in pressure otherwise the efficiency of the engine decreases.

    In pre-1986 engines, tetraethyl lead (TEL) was added to the petrol to help control the pressure increase in the combustion process. Since 1986, additives such as aromatics, MTBE, ETBE and ethanol have replaced TEL.

(a) Given that Pb is the central atom in TEL, draw a possible structural formula for TEL.

(b) Suggest a reason why TEL is no longer added to petrol.

(c) MTBE (methyl tertiary butyl ether or $C_5H_{12}O$) is essentially a molecule of methylpropane bonded to an oxygen atom that is bonded to a methane molecule. Given that the methylpropane and methane molecules must lose a hydrogen atom before they can join to the oxygen atom, draw the structural formula for MTBE.

    ETBE (ethyl tertiary butyl ether or $C_6H_{14}O$) has a very similar structure to MTBE, except that ethane is joined to the oxygen rather than methane.

(d) Draw the structural formula for ETBE.

(e) Write balanced chemical equations for the combustion of MTBE and ETBE.

    Petrol also contains xylenes, which are isomers of dimethyl benzene, and olefins, which are alkenes.

(f) Draw the structural formula and give the IUPAC name of a xylene and an olefin containing 8 carbon atoms.

(g) Write balanced equations for the combustion of the xylene and the olefin drawn in part (f).

CHEMICAL REACTIONS AND ENERGY CHANGE

5

Topics covered in this chapter:

5.1 PHYSICAL AND CHEMICAL CHANGES

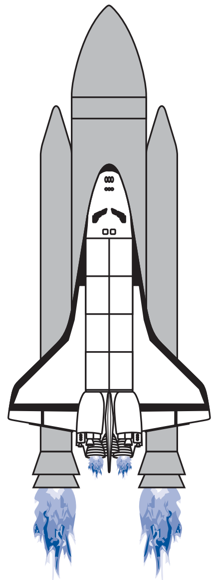
Physical changes are those that only involve a change in a physical property of a substance such as a change of state. Ice melting or salt dissolving in water are examples of physical changes. In each case no new chemical substance is formed.

Chemical changes however involve the formation of a new substance with different chemical properties. Burning paper and rusting iron are two common examples of chemical change. In both cases new chemical substances are produced.

We know that a chemical change, that is, a chemical reaction has occurred by the fact that:

- new substance(s) form
- energy is taken in or given out
- change is not easily reversed.

(a) physical change – ice cubes melting
$H_2O_{(s)}$ + heat $\rightarrow$ $H_2O_{(l)}$



(b) chemical change – hydrogen fuel burning in a rocket
$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$ + heat

Figure 5.1 Physical and chemical changes.

Question 5.1

Indicate whether the following are physical or chemical changes.

(a) candle wax melting ______________________

(b) boiling an egg ______________________

(c) frying an egg ______________________

(d) clouds forming in air ______________________

(e) petrol burning ______________________

(f) grape juice fermenting ______________________

(g) wood cut to sawdust ______________________

(h) digesting ice cream ______________________

5.2 WRITING EQUATIONS

A chemical reaction can be represented by an equation. This can be a simple word equation or a balanced equation using symbols and formulae.

For example the burning of magnesium in oxygen gas can be represented as follows:

word equation	magnesium	+	oxygen	→	magnesium oxide
balanced equation	$2Mg_{(s)}$	+	$O_{2(g)}$	→	$2MgO_{(s)}$

To write a balanced chemical equation:

1. **Write the correct formula** for each of the reactants (on the left) and the products (on the right). You may find it helpful to write the word equation first. Once correct formula is written it *must not* be changed.

2. **Balance the equation** so that the number of atoms of each element on the reactant side is the same as on the product side. Do this by altering the coefficients as necessary (the numbers in front of each formula or symbol). Never change the actual formula.

The **Law of Conservation of Mass** always applies when we balance a chemical equation. This Law states that the total mass of the reactants in a chemical reaction is equal to the total mass of the products. This simply means that the number of atoms of each element involved must remain constant.

Worked Example

5.1 Write a balanced equation for the combustion of propane gas (C_3H_8) in air.

Step 1: word equation
propane + oxygen → carbon dioxide + water

Step 2: unbalanced equation showing correct formulae
$C_3H_{8(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$

Step 3: balancing the equation

- there are insufficient C atoms on the right so the coefficient 3 is placed in front of CO_2
  $C_3H_{8(g)} + O_{2(g)} \rightarrow 3CO_{2(g)} + H_2O_{(g)}$

- there are insufficient H atoms on the right so a 4 is placed in front of H_2O
  $C_3H_{8(g)} + O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$

- there are insufficient O atoms on the left so a 5 is placed in front of the O_2
  $C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$

Check: the equation now shows
$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$

left side (reactants)	right side (products)
3 atoms of C	3 atoms of C
8 atoms of H	8 atoms of H
10 atoms of O	10 atoms of O

Hence equation is balanced and obeys the Law of Conservation of Mass.

5.3 EXOTHERMIC AND ENDOTHERMIC REACTIONS

Chemical reactions are nearly always associated with the release or absorption of energy. This is due to differences in the bonding energy of reactants and products.

Exothermic reactions are those that release energy to the surroundings. For example when petrol burns, energy is released. This causes the surroundings to become hot.

Endothermic reactions are those that absorb energy from the surroundings. For example the melting of ice absorbs energy from its surroundings. Hence if ice is placed in a glass of water it will cool it.

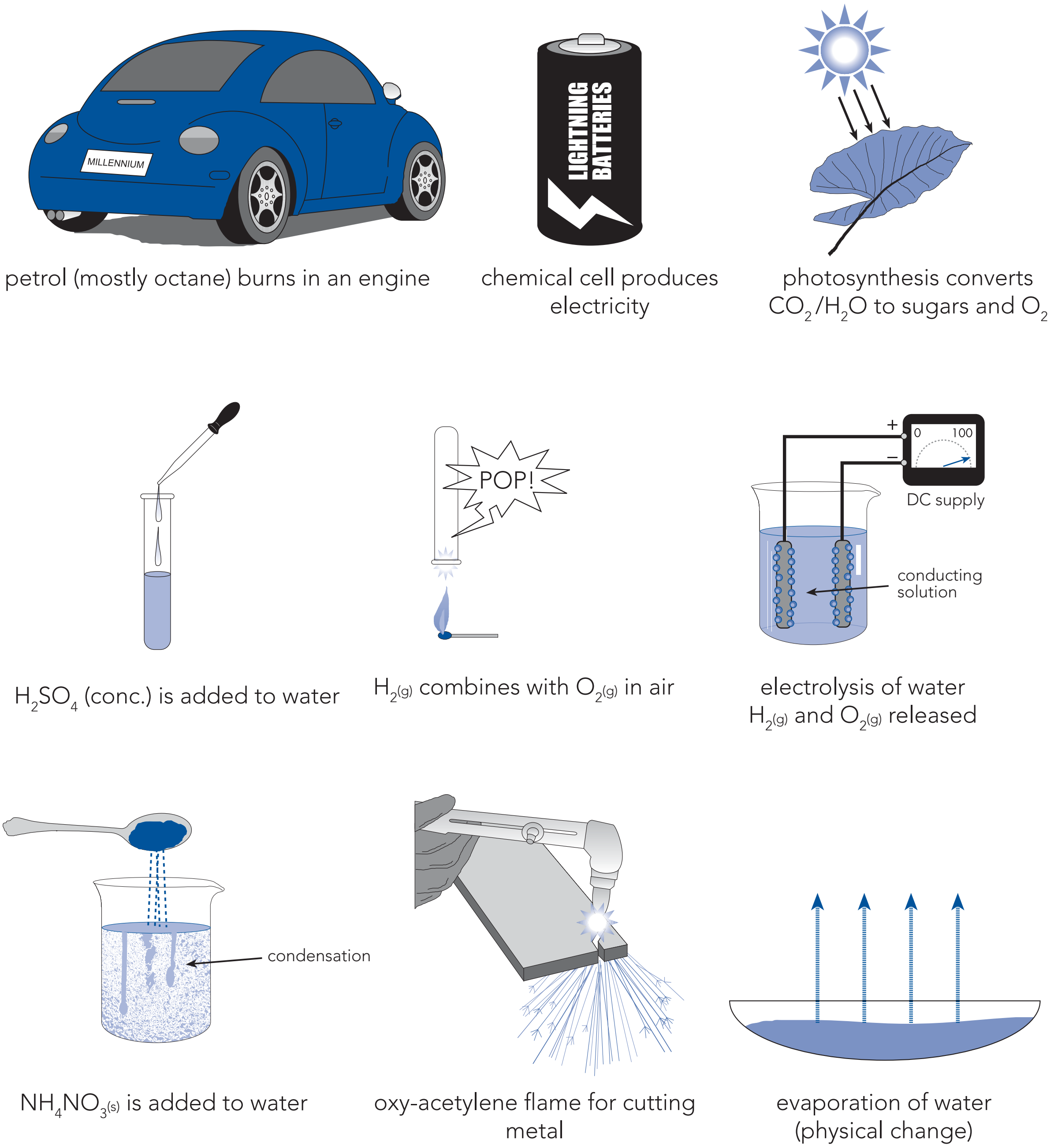


Figure 5.2 Chemical reactions involve energy changes.

Question 5.2

Complete the following table using all the examples shown in Figure 5.1. The first one is done for you.

REACTION EXAMPLE	ENERGY ABSORBED/ RELEASED TO SURROUNDINGS	REACTION TYPE EXOTHERMIC/ ENDOTHERMIC
combustion of petrol (octane)	*released*	*exothermic*
chemical cell produces electricity		
photosynthesis converts CO_2/H_2O to sugars and O_2		
H_2SO_4 (conc.) is added to water		
$H_{2(g)}$ combines with $O_{2(g)}$ in air		
electrolysis of water; $H_{2(g)}$ and $O_{2(g)}$ released		
$NH_4NO_{3(s)}$ is added to water		
oxy-acetylene flame for cutting metal		
evaporation of water (physical change)		

5.4 CONSERVATION OF ENERGY

During a chemical reaction energy may be transformed from one form to another but the total amount of energy remains constant. This is the **Law of Conservation of Energy.**

Within a substance chemical energy exists as kinetic energy (energy of motion) and potential energy (energy of position).

Kinetic energy (energy of motion)

Gaseous molecules can have kinetic energy due to three different types of motion:

(a) **vibrational:** atoms move away and towards each other.

(b) **rotational:** molecule rotates about its centre.

(c) **translational:** molecule moves from place to place.

Note that atoms only have translational motion.

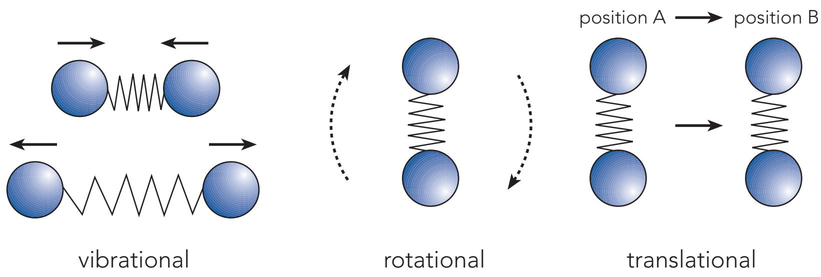


Figure 5.3 Types of motion for a gaseous diatomic molecule.

Potential energy (energy of position)

Potential energy in a chemical substance is due to the forces of attraction (or repulsion) that exist between the positively charged nuclei of atoms and negatively charged electrons. Several forces may act on any one charged particle at the same time (see Figure 5.4).

When two atoms of hydrogen, for example, are close together, the total attractive forces overcome the total repulsive forces and a molecule is formed. Potential energy is given up (exothermic reaction).

If the two atoms of a hydrogen molecule were to be pulled apart, the separated atoms will have gained potential energy (endothermic reaction).

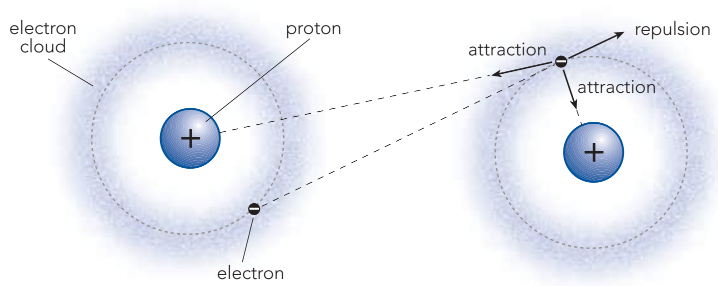


Figure 5.4 Forces of attraction and repulsion between nearby hydrogen atoms. The proton and electron of each atom are under the influence of attractive and repulsive forces. If the total attractive forces overcome the total repulsive forces then a molecule is formed.

Question 5.3

(a) The diagram above (Figure 5.4) shows a simplified view of two atoms of hydrogen near each other. The forces acting on one of the electrons are shown. Draw all the forces acting on the other three particles.

(b) In the situation illustrated above the attractive forces are greater than the repulsive forces. The two atoms move closer together and form a chemical bond.

(i) Is this an endothermic or exothermic process? ______________________

(ii) Has potential energy increased or decreased?

____

(iii) How is the temperature of the surroundings affected? Explain your answer.

____

5.5 HEAT CONTENT (ENTHALPY)

The stored chemical energy of a chemical substance is referred to as its heat content or enthalpy (H). The enthalpy change (ΔH) during a reaction is referred to as the heat of reaction.

Energy changes which occur during reactions can be illustrated graphically as shown below.

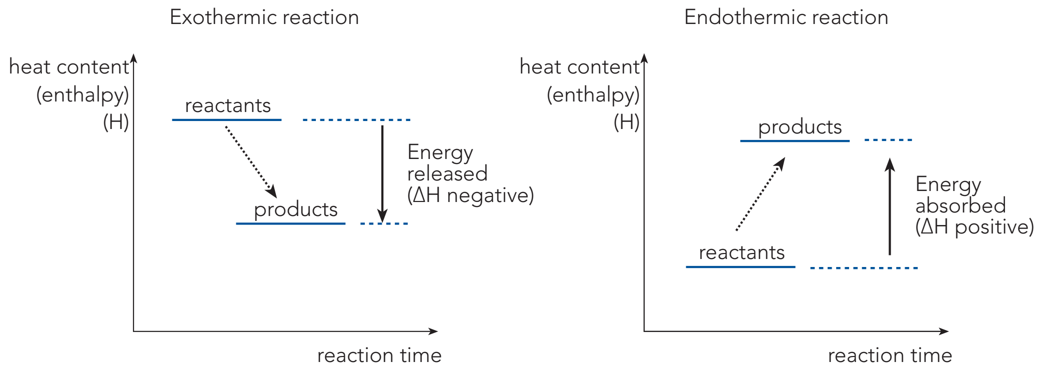


***Figure 5.5** Energy changes in exothermic and endothermic reactions.*

5.6 INDICATING HEAT OF REACTION

The change in enthalpy for a particular reaction can be written as part of the chemical equation for that reaction or simply stated as a ΔH value.

e.g. $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)} \quad \Delta H = -393\ kJ$

or $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)} + 393\ kJ$

In either case we are indicating that 393 kJ of energy are released by the combustion of 1 mole of carbon.

Question 5.4

Complete the following:

(i) ΔH = heat content of ________________ – heat content of ________________

(ii) For exothermic reactions:

- ΔH is ________________
- The enthalpy (H) of the products is ___________ than that of the reactants.
- Heat is _______________ (to/from) the surroundings.

(iii) Similarly for endothermic reactions:

- ____
- ____
- ____

Question 5.5

For the following reactions indicate the ΔH value and the effect on the temperature of the surroundings:

(a) $H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)} + 185\ kJ$

(b) $C_{(s)} + H_2O_{(g)} + 132\ kJ \rightarrow H_{2(g)} + CO_{(g)}$

(c) $O_{2(g)} + ½O_{2(g)} + 143\ kJ \rightarrow O_{3(g)}$

(d) $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)} + 92\ kJ$

(a) ΔH = ___________ , temperature of surroundings will ________________

(b) ΔH = ___________ , temperature of surroundings will ________________

(c) ΔH = ___________ , temperature of surroundings will ________________

(d) ΔH = ___________ , temperature of surroundings will ________________

Question 5.6

Write balanced equations for the following reactions and include heat of reaction:

(a) When 1 mole of sulfur is burnt in oxygen gas to produce sulfur dioxide, 297 kJ of energy are released.

____

(b) When 1 mole of copper (II) carbonate is converted to its oxide 46.0 kJ of heat are absorbed.

____

Question 5.7

Consider the following reactions:

(i) $H_2O_{(l)} \rightarrow H_2O_{(g)}$ $\quad \Delta H = +44\ kJ$

(ii) $H_{2(g)} + ½\ O_{2(g)} \rightarrow H_2O_{(g)}$ $\quad \Delta H = -242\ kJ$

(ii) $H_{2(g)} + ½\ O_{2(g)} \rightarrow H_2O_{(l)}$ $\quad \Delta H = -286\ kJ$

(a) The heat of reaction for (ii) is much greater than (i). Explain in terms of bonding changes involved.

____

(b) Reaction (ii) and (iii) involve the same substances yet the ΔH is different. Explain.

____

5.7 FOSSIL FUELS

Fossil fuels such as coal, oil and natural gas were formed millions of years ago from the remains and decay of plants and animals. They all contain carbon and release energy when burned. Unfortunately their burning also results in the release of carbon dioxide and other unwanted substances into the air.

Hence the efficiency with which different fuels burn and the relative amount of carbon emissions they produce are important considerations in how they are used. Further to this fossil fuels are a non-renewable resource which will eventually run out or be in very limited supply.

FUEL	HEATING VALUE MJ/kg	g of C / MJ
Wood – dry pine	20	25
Coal – anthracite	34	27
Coal – low bituminous	29	26
Natural gas	54	14
Petrol	47	18
Fuel oil	41	21
Ethanol	30	18
Hydrogen	142	0

***Table 5.1** Heating values and carbon emissions of some common fuels. Values given (room temp) are typical only as the composition of some of the fuels listed can vary widely. For example natural gas is essentially methane but can contain small quantities of ethane, say 3%, and smaller quantities of other hydrocarbons and inert compounds. Coal in particular varies greatly from low grade brown coal to anthracite.*

Coal

Coal is formed from the remains of plants of ancient forests buried millions of years ago. The decay process initially produces peat. However over time, the weight of more overburden and geothermal heat combine to produce different forms of coal such as lignite, black coal and anthracite. Anthracite produces the most amount of energy but due to its limited supply it is mainly used for processing metals.

The most common use of coal is as an energy source for power stations to generate electricity. World wide, it is by far the major energy source for this purpose. Unfortunately the burning of coal can produce unwanted pollutants as well as greenhouse gases. The emission of pollutants is minimised to regulated limits, as set by the EPA, by the use of filters and scrubbing systems.

The greenhouse gases emitted, water and carbon dioxide, form part of our atmosphere and are not considered as pollutants. However the level of carbon dioxide in the air has significantly increased over the past century or so and it is important that its emission is reduced.

One way that the environmental impact of burning coal may be reduced is its conversion to synthetic natural gas (syngas), a fuel which has a lower carbon footprint. Clean coal technologies are also being developed to reduce the emissions of pollutants and greenhouse gases. For example, it may be possible to develop a practical way to capture carbon dioxide emissions from power stations and store them safely away from the atmosphere.

Oil and gas

Petroleum and natural gas are formed from the remains of tiny marine plants and animals which settled on the sea floor and became covered by silt. The action of pressure, geothermal heat and anaerobic bacteria gradually changed the complex plant and animal molecules into simpler hydrocarbons, oil and gas. The oil and gas produced in this way gradually seeped into large reservoirs of porous rock where they were trapped by impervious layers of rock.

Natural gas and petroleum are often found together. However large reserves of mainly natural gas have been found such as those on the North-West shelf in Western Australia. Natural gas contains mainly methane (CH_4) together with some heavier hydrocarbons and impurities such as water, carbon dioxide, hydrogen sulfide and nitrogen.

Natural gas generally needs little treatment before it is either directly piped to consumers or exported as liquefied natural gas (LNG). However some of the heavier hydrocarbons such as ethane (C_2H_6) are sometimes separated as valuable by-products. Ethane is an important raw material for the petrochemical industry while propane (C_3H_8) and butane (C_4H_{10}) can be bottled as liquefied petroleum gas (LPG).

Petroleum, or crude oil, contains a great number of hydrocarbon compounds which can be separated by fractional distillation due to their different boiling points. In an oil refinery the crude oil is heated and the different "fractions" are collected at different points in fractionating tower. These "fractions" include gases at the top of the tower as they have the lowest boiling point, then petrol a little lower down the tower, followed by kerosene, oil and bitumen residue at the bottom.

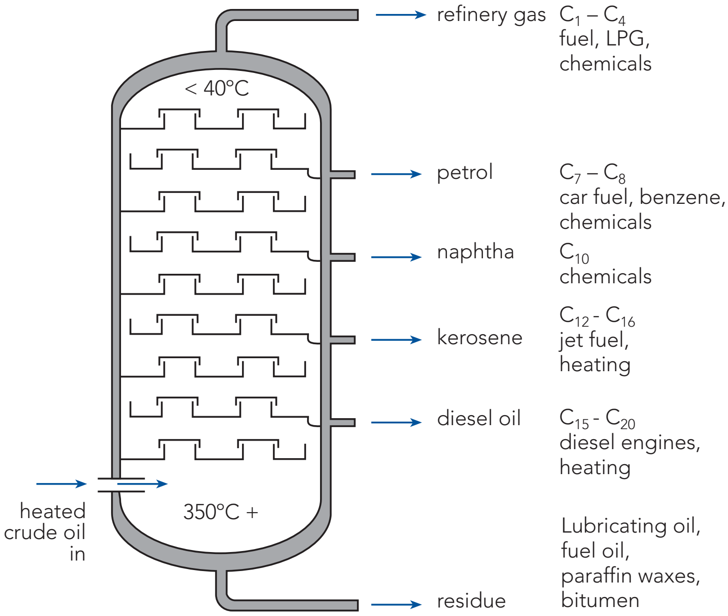


***Figure 5.6** Fractional distillation of crude oil. The oil is heated to over 350 °C and then enters the fractionating tower, mostly as gas. High boiling point hydrocarbons condense first while remaining vapours rise up the tower through bubble caps. As they cool and condense different fractions are collected from the many horizontal trays.*

Energy from fossil fuels

Fossil fuels consist predominantly of hydrocarbons and when they are burnt the products are carbon dioxide, water and energy. The energy released, or ΔH, varies for the different hydrocarbon molecules as shown by the examples below. The energy released per unit mass (not shown) is greatest when the hydrogen/carbon ratio is highest, that is, when the number of hydrogen atoms per carbon atom is greatest.

The combustion reactions for some common hydrocarbons are shown below. The ΔH values given are all for 25 °C and 1 atm.

Methane $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$ $\Delta H = -891$ kJ mol^{-1}

Ethene $C_2H_{4(g)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 2H_2O_{(l)}$ $\Delta H = -1411$ kJ mol^{-1}

Ethane $2C_2H_{6(g)} + 7O_{2(g)} \rightarrow 4CO_{2(g)} + 6H_2O_{(l)}$ $\Delta H = -1561$ kJ mol^{-1}

Butane $2C_4H_{10(g)} + 13O_{2(g)} \rightarrow 8CO_{2(g)} + 10H_2O_{(l)}$ $\Delta H = -2878$ kJ mol^{-1}

Heptane $C_7H_{16(l)} + 11O_{2(g)} \rightarrow 7CO_{2(g)} + 8H_2O_{(l)}$ $\Delta H = -4817$ kJ mol^{-1}

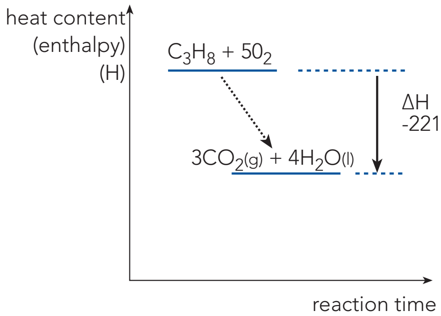
Worked Example

5.2 **Bottled gas (LPG), such as that used for barbecues, is mostly propane. The heat of reaction for propane at 25 °C and 1 atm is 2219 kJ mol^{-1}.**

(a) **Write the equation for the combustion of propane including the heat of reaction.**
(b) **Draw a simplified enthalpy diagram for the reaction showing the change in ΔH.**
(c) **Calculate the energy produced per kilogram of propane.**

(a) $C_3H_{8(g)} \quad + \quad 5O_{2(g)} \quad \rightarrow \quad 3CO_{2(g)} \quad + \quad 4H_2O_{(l)} \quad + \quad 2219\ kJ$

(b)



$M(C_3H_8)$ = $(3)(12.01) + (8)(1.008)$ = 44.09 g mol^{-1}

Moles of C_3H_8 in 1.0 kg = 1000/44.09 = 22.68 mol

Energy produced per kg = 2219 x 22.68 = 50327 kJ

Energy per kilogram of propane = 50.3 MJ

Question 5.8

Consider the combustion of methane (CH_4) as shown in the equation above and answer the following.

(a) During the reaction six chemical bonds are broken: 4 C-H bonds and 2 O=O bonds. Name and list the bonds that were formed during the reaction.

____

____

(b) In total which bonds involve more energy, those that are broken or those that are formed? Explain your answer.

____

(c) Calculate the energy produced per kg of methane.

____

(d) Compare this result with the value that was calculated for propane in worked example 5.2.

____

Question 5.9

Compare the combustion of methane (CH_4) and butane (C_4H_{10}) as shown in the equations listed on page 92 to answer the following.

(a) Determine the moles of carbon dioxide produced per mole of fuel burnt in each case.

____

____

(b) Determine the moles of carbon dioxide produced per MJ of energy produced in each case.

____

____

(c) Which fuel gives the most energy for the same quantity of carbon dioxide produced? Give a possible reason for this.

____

____

5.8 ALTERNATIVE ENERGY SOURCES

Fossil fuels provide the majority of the energy used in the world today. However, they are a limited resource and their combustion can contribute negatively to the environment. Hence a greater use of alternatives is being made. These include the use of nuclear energy and renewable sources such as solar energy, wind power, hydro-electric power and wave energy.

New energy sources are continually being sought and investigated. For example, just off Garden Island in WA, underwater wave power converters are being trialled. These underwater actuators are designed to capture the power of underwater waves and use it to produce electricity and desalinated water. If effective, the major bonus would be the use of a readily available energy source without greenhouse gas emissions.

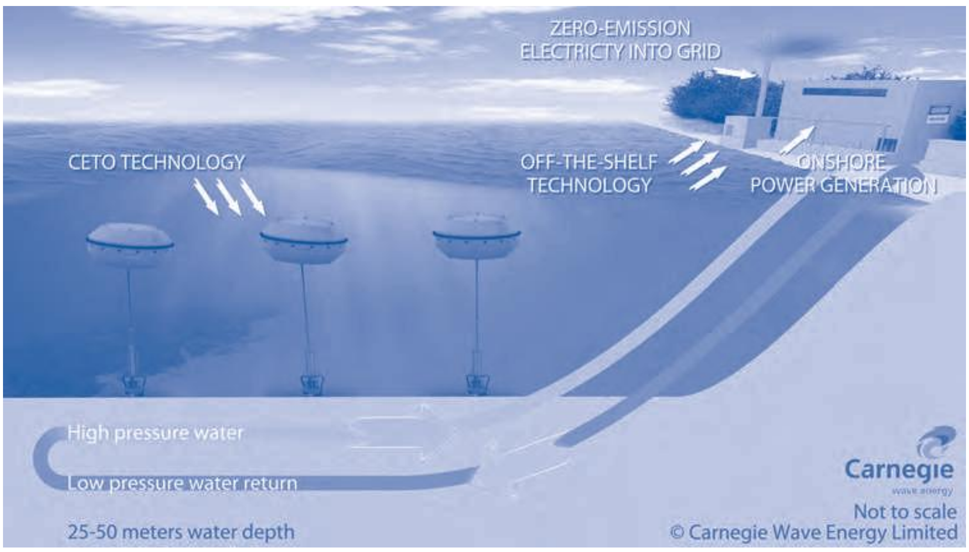


Diagram courtesy of Carnegie Wave Energy Ltd.

Biofuels

Alternatives to fossil fuels are also being developed and used. These include biofuels such as biogas, biodiesel and bioethanol which all offer the benefit of a reduced greenhouse effect. However their use is still not significant due to many factors such as limited availability and cost of production.

Biogas for example is sourced from the processing of organic waste such as manure, sewage and green waste. The fermentation and breakdown of these biodegradable materials produces methane gas and a little carbon dioxide. The use of biogas provides what is essentially free energy and it also has the advantage of removing methane from the environment. However it is only available in relatively small quantities.

Biodiesel blends and ethanol fuels however, are now becoming more readily available. The feedstock for these fuels include tallow, used cooking oil, oil seeds, wheat stubble and sugar cane waste.

Ethanol fuel can be readily produced from sugar cane, wheat, corn or maize. However these are also important food resources. Ethanol as a motor fuel is used mainly as a blend with petrol as this means engines need little if any modifications. The United States and Brazil are by far the largest producers of ethanol fuel for cars.

Major advancements in the use of biofuels will depend on finding new, low cost, and efficient ways of converting available biomass. A local example of this is the research that the CSIRO is currently undertaking in the use of algae for biofuels. Algae are not a competing food resource and it may be possible to produce it much more cheaply than alternate land crops used for biofuels. Algae can be fermented and processed to produce different biofuels such as biogas (methane), bioethanol and biodiesel.

REVIEW QUESTIONS

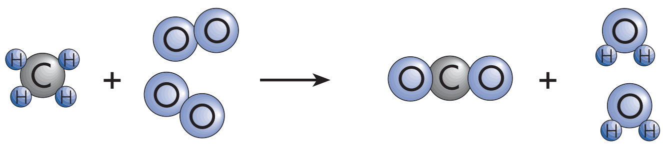
Chapter 5: Chemical Reactions and Energy Changes

1. Classify each of the following as either physical or chemical changes.
   (a) sugar dissolving in a coffee drink
   (b) steam condensing and forming water droplets
   (c) hydrogen gas burning and forming water droplets
   (d) clothing fabric losing their colour and becoming faded
   (e) iron nail rusting in moist air
   (f) meat cooking in a barbeque

2. When a candle is lit a pool of liquid wax forms near the wick and a flame is produced.
   (a) Has there been a physical change? Explain.
   (b) Has there been a chemical change? Explain.

3. Methane gas, often called natural gas, is used in many homes for heating. When it burns it combines with oxygen gas in the air and forms carbon dioxide gas and water vapour.

   A visual representation of the reaction is show below.



   (a) Write a word equation for the reaction.

   (b) Give a balanced molecular equation.

4. Determine appropriate coefficients to balance the following unbalanced equations.

   (a) $\_\_CO_{(g)} + \_\_O_{2(g)} \rightarrow \_\_CO_{2(g)}$

   (b) $\_\_C_3H_{8(g)} + \_\_O_{2(g)} \rightarrow \_\_CO_{2(g)} + H_2O_{(g)}$

   (c) $\_\_Ca_{(s)} + \_\_HCl_{(aq)} \rightarrow \_\_CaCl_{2(aq)} + \_\_H_{2(g)}$

   (d) $\_\_Ca(HCO_3)_{2(s)} \rightarrow \_\_CaCO_{3(s)} + \_\_CO_{2(g)} + H_2O_{(g)}$

   (e) $\_\_Fe_2O_{3(s)} + \_\_CO_{(g)} \rightarrow \_\_Fe_{(s)} + \_\_CO_{2(g)}$

   (f) $\_\_H_2O_{2(l)} \rightarrow \_\_O_{2(g)} + \_\_H_2O_{(l)}$

   (g) $\_\_Al_{(s)} + \_\_HCl_{(aq)} \rightarrow \_\_AlCl_{3(aq)} + \_\_H_{2(g)}$

   (h) $\_\_KHCO_{3(s)} \rightarrow \_\_K_2CO_{3(s)} + \_\_H_2O_{(l)} + CO_{2(g)}$

5. Write balanced equations for the following reactions.

(a) hydrogen + oxygen → water

(b) carbon monoxide + oxygen → carbon dioxide

(c) sodium carbonate + calcium chloride → calcium carbonate + sodium chloride

(d) potassium hydroxide + sulfuric acid → potassium sulfate + water

(e) magnesium oxide + phosphoric acid → magnesium phosphate + water

(f) copper(II)hydroxide + hydrochloric acid → copper(II) chloride + water

(g) sodium hydroxide + phosphoric acid → sodium phosphate + water

(h) ammonium hydrogencarbonate + nitric acid → ammonium nitrate + water + carbon dioxide

6. Complete the following table by indicating whether the processes are exothermic or endothermic and whether the final enthalpy (H) is higher or lower.

PHYSICAL OR CHEMICAL CHANGE	PROCESS IS EXOTHERMIC / ENDOTHERMIC	ENTHALPY OF PRODUCTS (H) IS HIGHER / LOWER
propane gas (barbecue gas) is burnt		
ice is placed in water and melts		
the two atoms making up an oxygen molecule are separated		
solid carbon dioxide (dry ice) sublimes to its gaseous form		

7. (a) What is meant by the enthalpy of a substance?

(b) Under what circumstances will the enthalpy of a system decrease?

8. Determine the ΔH value (including + or –) for each of the following:

(a) $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)} + 393\ kJ$

(b) $C_{(s)} + H_2O_{(g)} + 132\ kJ \rightarrow H_{2(g)} + CO_{(g)}$

(c) $H_2O_{(l)} + 44\ kJ \rightarrow H_2O_{(g)}$

9. Write balanced equations for the following reactions including the heat of reaction:

   (a) When 1 mole of hydrogen gas combines with 1 mole of chlorine gas to produce hydrogen chloride gas, 185 kJ of heat is produced.

   (b) 1 mole of oxygen gas (O_2) combines with more O_2 to form 1 mole of ozone (O_3) by absorbing 143 kJ of heat.

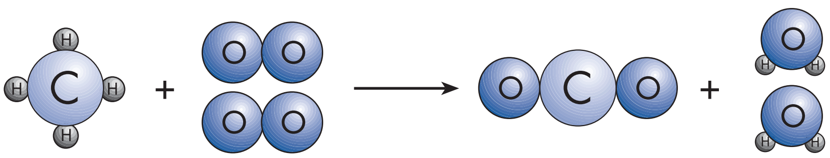
10. Consider the two following changes:

   (a) $H_2O_{(g)} \rightarrow H_2O_{(l)}$ $\Delta H - 44$ kJ

   (b) $H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(l)}$ $\Delta H - 286$ kJ

   In both cases $H_2O_{(l)}$ is produced. Explain the great difference in energy released in each case.

11. The combustion of methane gas is illustrated below.



$$CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(g)}$$

   (a) Determine how many bonds must be broken to burn 1 mole of methane. List these.

   (b) Determine how many bonds form when 1 mole of methane is burnt. List these.

   (c) Since this process is highly exothermic which bonds, those of the reactants or those of the products, are likely to be stronger?

12. The combustion of methane, as shown in the previous question, occurs much more rapidly than the similar combustion of larger hydrocarbon molecules such as octane. Give a likely reason for this.

13. Acetylene (C_2H_2) produces an intense flame when burnt with pure oxygen and is widely used for welding. The heat of reaction for this gas, at 25 °C and 1 atm, is 1301 kJ mol^{-1}.

   (a) Write the equation for the combustion of acetylene.

   (b) Name and list the bonds that were broken and those that were formed during the reaction.

   (c) In total which bonds involve more energy, those that are broken or those that are formed? Explain your answer.

14. Two important fuels used for motor cars are petrol and auto gas (LPG). Petrol is a blend of hexane and octane while auto gas is a blend of propane and butane. For the purposes of this question we will consider the fuels as being octane and butane. Their heats of combustion are 5471 kJ mol^{-1} for octane and 2878 kJ mol^{-1} for butane.

In order to compare these two fuels answer the following:

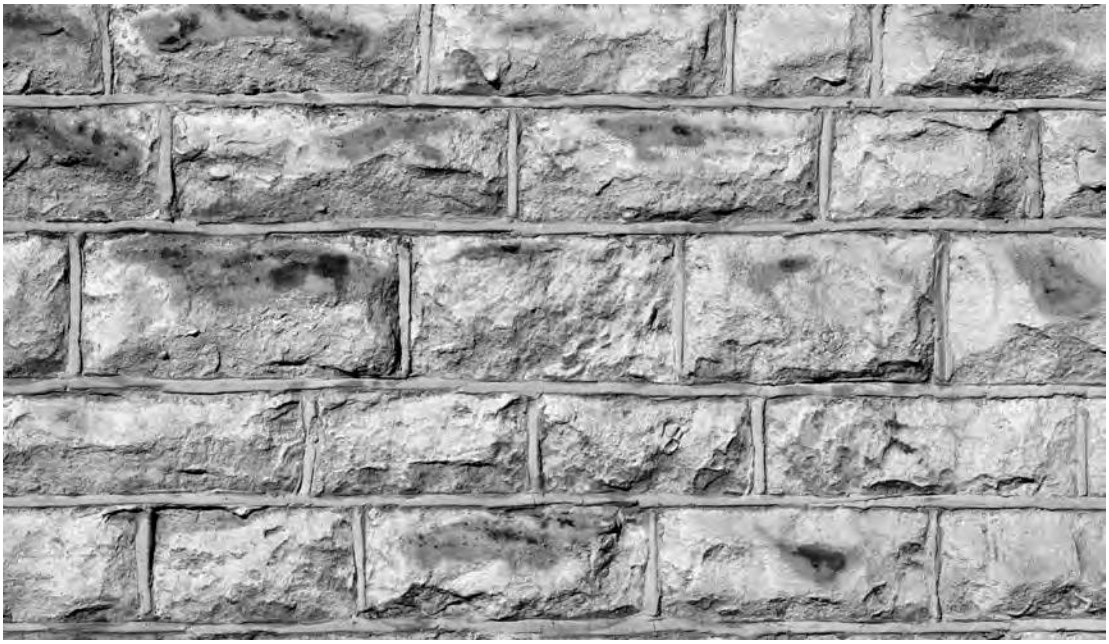
(a) Write the equation for the combustion of butane and octane.

(b) Determine the energy produced per kg of each fuel.

(c) Determine the moles of carbon dioxide produced per MJ of energy produced in each case.

(d) Use your results to compare these fuels.

FOR THE EXPERTS



Limestone is an important building material in many ways. It is used in the construction of roads and buildings but also to produce lime, a principal ingredient of mortar and cement.

- Limestone is essentially calcium carbonate ($CaCO_3$). When it is heated in a hot furnace to about 1500°C it decomposes to form lime (CaO) and carbon dioxide gas (CO_2).
- Lime produced by this process, also called quicklime, can be converted to calcium hydroxide [$Ca(OH)_2$], also called hydrated lime, by the addition of water.

Mortar is a mixture of cement, lime, sand and water. Cement is produced by heating limestone with clay.

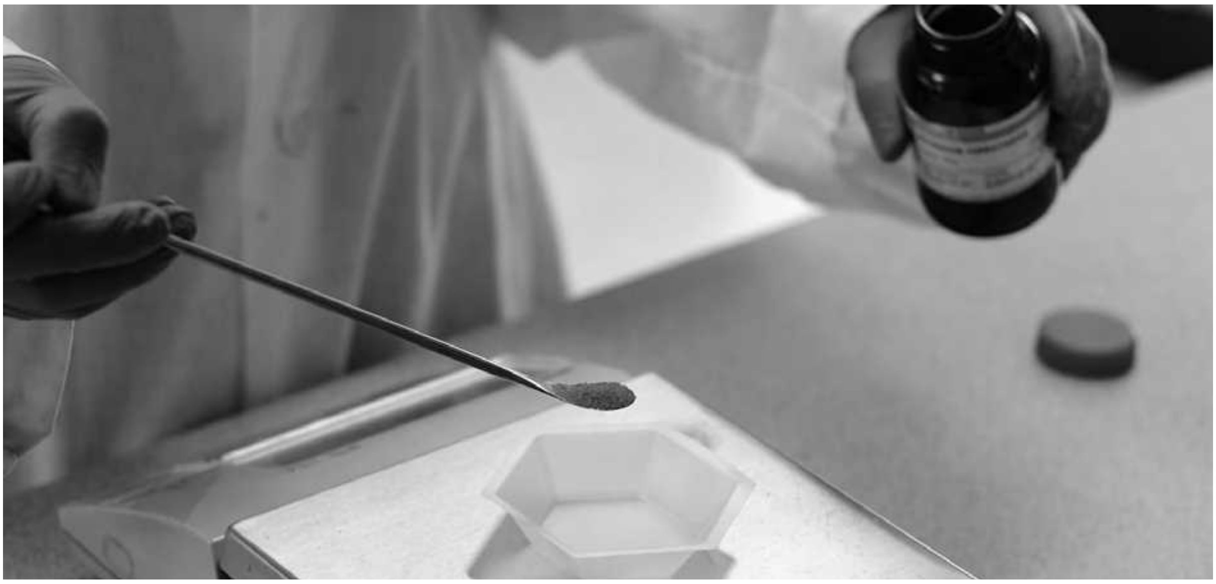
15.

(a) The decomposition of calcium carbonate to lime and carbon dioxide requires 179 kJ/mole of heat.
  (i) Write a balanced equation for this reaction including enthalpy changes.
  (ii) Is this an endothermic or exothermic reaction? Briefly explain your answer in terms of bonds broken/formed during the reaction.

(b) The formation of hydrated lime from lime releases 66 kJ/mole.
  (i) Write a balanced equation for this reaction including enthalpy changes.
  (ii) Is this an endothermic or exothermic reaction? Briefly explain your answer in terms of bonds broken/formed during the reaction.

CHEMISTRY CALCULATIONS

6



Topics covered in this chapter:

6.1 RELATIVE MASSES

Relative atomic mass (A_r)

The mass of individual atoms is very small and hence it is more convenient to use relative masses. The **relative atomic mass** (**A_r**) of an element is the average mass of one atom of it compared to $^1/_{12}$ of the mass of an atom of carbon-12. The mass of an atom of carbon-12 is taken as being exactly twelve (12.0).

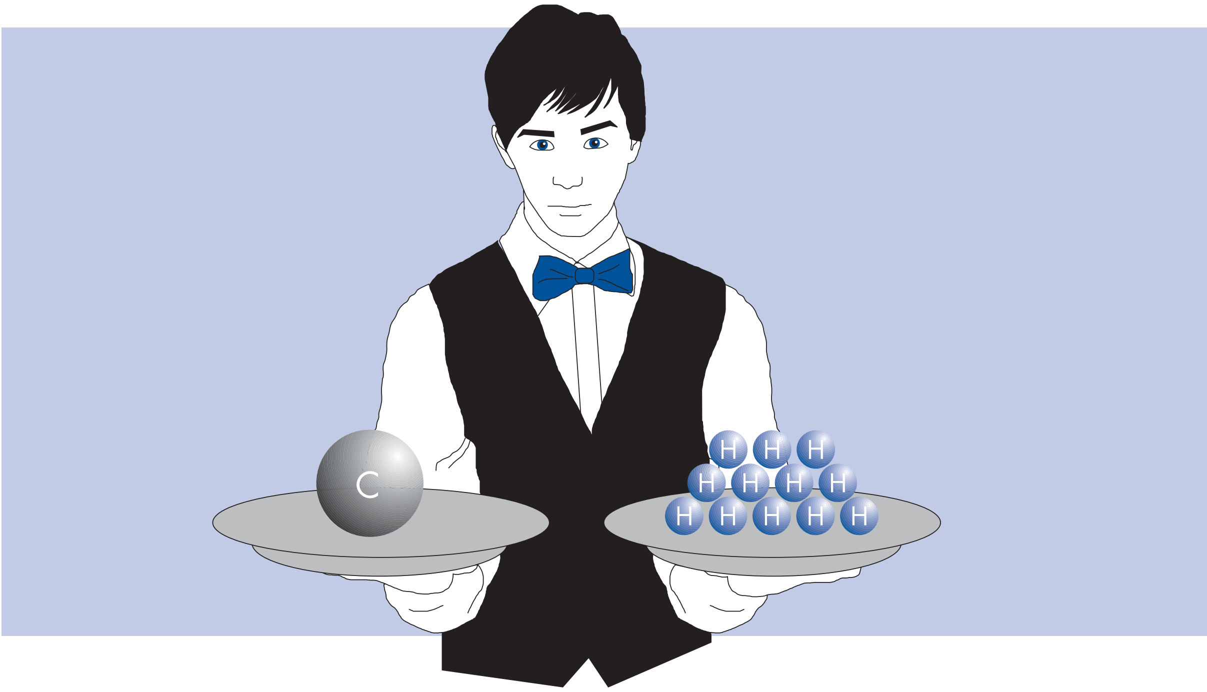


Figure 6.1 Relative masses. The mass of 1 carbon atom is equal to the mass of 12 hydrogen atoms. Relative atomic masses have no units because they are ratios.

Worked Example

6.1 The relative atomic mass (A_r) for carbon atoms is 12.0.

i) The mass of magnesium atoms is found to be approximately twice that of carbon atoms. What would be the likely A_r for Mg?

A_r (Mg) ≈ (2) (12.0) ≈ 24.0

ii) It is found that the mass of 3 helium atoms is exactly the same as 1 carbon atom. What should be the A_r for He?

A_r (He) = 1/3 (12.0) = 4.0

Table 6.1 Relative atomic masses (A_r) of some elements

METALS			NON METALS		
Element	**Symbol**	**A_r**	**Element**	**Symbol**	**A_r**
Aluminium	Al	26.98	Argon	Ar	39.95
Barium	Ba	137.3	Boron	B	10.82
Calcium	Ca	40.08	Bromine	Br	79.90
Chromium	Cr	52.00	Carbon	C	12.01
Copper	Cu	63.55	Chlorine	Cl	35.45
Gold	Au	197.0	Fluorine	F	19.00
Iron	Fe	55.85	Helium	He	4.003
Lead	Pb	207.2	Hydrogen	H	1.008
Magnesium	Mg	24.31	Krypton	Kr	83.80
Mercury	Hg	200.6	Iodine	I	126.9
Potassium	K	39.10	Neon	Ne	20.18
Silver	Ag	107.9	Nitrogen	N	14.01
Sodium	Na	22.99	Oxygen	O	16.00
Tin	Sn	118.7	Phosphorus	P	30.97
Titanium	Ti	47.87	Silicon	Si	28.09
Zinc	Zn	65.38	Sulfur	S	32.07

Question 6.1

Use information from Table 6.1 to answer the following:

(a) Approximately how many times heavier than a carbon-12 atom is an atom of:

(i) Silver? ________________ (ii) Krypton? ________________

(b) Approximately how many times heavier than a hydrogen atom is an atom of:

(i) Carbon? ________________ (ii) Iron? ________________

(c) Approximately how many atoms of aluminium have the same mass as 1 atom of silver?

____

(d) Which metal atom is approximately 40 times heavier than hydrogen?

____

Relative molecular mass (M_r)

The atoms of non-metals, such as those listed in Table 6.1, are able to combine to form molecules. The mass of these molecules can also be compared to the mass of carbon-12 atoms to find their relative molecular mass.

The relative molecular mass (M_r) of an element or a compound is the mass of one of its molecules compared to $^1/_{12}$ of the mass of an atom of carbon-12.

Relative formula mass (M_r)

Ionic compounds such as sodium chloride (table salt) are not made up of discrete molecules and hence it would be incorrect to refer to their relative molecular mass. Instead, for ionic substances, we refer to their relative formula mass (M_r). You will notice the same symbol, M_r, is used and it is defined in the same way as relative molecular mass.

Calculating relative masses

The relative molecular mass (or relative formula mass) can be calculated by simply adding the relative masses of all the atoms shown in the formula.

Worked Examples

6.2 i) Find the M_r for H_2O.

$$M_r\,(H_2O) = 2\,(1.008) + 16.00 = 18.016$$

ii) Find the M_r for H_2SO_4.

$$M_r\,(H_2SO_4) = 2\,(1.008) + 32.07 + 4\,(16.00) = 98.086$$

iii) Find the M_r for $Zn(NO_3)_2$.

$$M_r\left(Zn(NO_3)_2\right) = 65.38 + 2\,(14.01) + 6\,(16.00) = 189.4$$

Question 6.2

Determine the M_r for the following:

(a) $M_r\,(SO_2)$ = ______________________

(b) $M_r\,(CuSO_4)$ = ______________________

(c) $M_r\left(Ca(OH)_2\right)$ = ______________________

(d) $M_r\left(Al_2(SO_4)_3\right)$ = ______________________

(e) $M_r\left(CH_3COOH\right)$ = ______________________

Question 6.3

Write down the formulae for the following and determine the relative molecular (or formula) mass.

		Formula	Mass
(a)	Carbon dioxide	______________	______________
(b)	Magnesium hydroxide	______________	______________
(c)	Chlorine gas	______________	______________
(d)	Ammonia gas	______________	______________
(e)	Iron (III) sulfate	______________	______________

Percentage composition

The proportion (or percentage) by mass of each of the elements in a compound can be determined from the formula by comparing relative masses.

$$\textbf{Percentage (\%) by mass} = \frac{\textbf{total } A_r \textbf{ for that element}}{M_r \textbf{ for the compound}} \times 100 \textbf{ (of an element)}$$

Worked Examples

6.3 1) Determine the percentage by mass of carbon in carbon dioxide.

STEP 1 Find M_r for carbon dioxide.

$$M_r\ (CO_2) = 12.01 + 2(16.00) = 44.01$$

STEP 2 % by mass of carbon

$$= \frac{12.01}{44.01} \times 100 = 27.3\%$$

2) Determine the percentage by mass for each element in silver nitrate

$$M_r\left(AgNO_3\right) = 107.9 + 14.01 + 3(16.00) = 169.91$$

Hence $\%\ Ag = \dfrac{107.9}{169.91} \times 100 = 63.50\%$

$$\%\ N = \frac{14.01}{169.91} = 8.25\%$$

$$\%\ O = \frac{48.0}{169.91} = 28.25\%$$

Question 6.4

Determine the percentage by mass of each element within the following compounds:

(a) carbon monoxide

(b) sodium chloride

(c) calcium hydroxide

Question 6.5

Determine the percentage by mass of water in the following hydrated compounds:

(a) sodium carbonate-10-water

(b) barium chloride-2-water

(c) magnesium sulfate-7-water

Question 6.6

Most metals are recovered from minerals which are found in ores. The minerals are actually compounds of the desired metal. Calculate the percentage by mass of the metallic element in the following minerals:

(a) aluminium in alumina (aluminium oxide)

(b) calcium in pure limestone (calcium carbonate)

(c) iron in haematite (iron (III) oxide)

6.2 THE MOLE

Moles

Chemical reactions involve very large numbers of extremely small particles. These particles, such as atoms, molecules and ions are simply too small to count or weigh individually.

The mole (n) is a number that chemists use to count large numbers of particles. Just as we can refer to a dozen eggs (12), or a ream of paper (500 sheets), chemists speak of a mole of atoms (6.022×10^{23} atoms).

The mole is the amount of substance containing Avogadro's number (N_A) of particles. This number, named after the Italian scientist Amadeo Avogadro is approximately 602,200,000,000,000,000,000,000 or more simply 6.022×10^{23}.

Chemists don't actually count out a mole of atoms but instead do so by weighing. Just as a bank teller would weigh out $100 of 5¢ coins, a chemist can weigh out a mole of, say, aluminium atoms by weighing out its relative atomic mass in grams (27 g).

Counting very tiny particles is impractical but we can achieve the same thing by weighing substances. The relative atomic mass (A_r) or relative molecular mass (M_r), measured in grams, always contains 1 mole of particles.

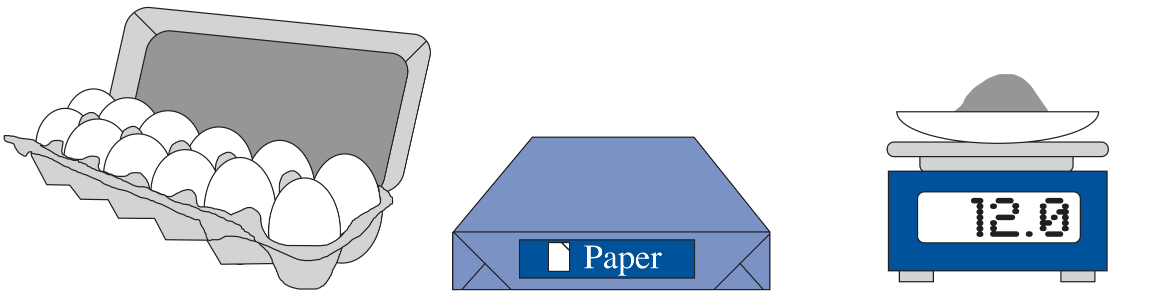


Figure 6.2 *Counting particles. The mole (n) is simply a convenient number when dealing with chemical reactions involving large numbers of extremely small particles. If we weigh out exactly 12.0 g of carbon-12 it will contain exactly 1 mol of carbon atoms (6.022×10^{23}).*

Calculations – moles and particles

The number of particles that make up one mole of a substance is 6.022×10^{23}. Hence if we wish to convert the number of particles to moles we use the following relationship.

$$n = \frac{N}{6.022 \times 10^{23}}$$

n = moles of particles
N = number of particles

Worked Examples

6.4 i) How many moles of ammonia molecules are there in 3.60×10^{24} molecules of this substance?

$$n\,(NH_3) = \frac{N}{6.022 \times 10^{23}} = \frac{3.60 \times 10^{24}}{6.022 \times 10^{23}} = 5.98 \text{ mol of } NH_3 \text{ molecules}$$

ii) How many atoms of aluminium are there in 12.0 moles of this substance?

$$n = \frac{N}{6.022 \times 10^{23}} \qquad 12.0 = \frac{N}{6.022 \times 10^{23}}$$

$\therefore$ N = $(12.0)\,(6.022 \times 10^{23})$ = 7.23×10^{24} atoms of aluminium

iii) How many atoms of a) carbon and b) oxygen are there in 3.0 moles of carbon dioxide molecules (CO_2)?

3.0 mol of CO_2 molecules contain:

→ 3.0 mol of C atoms
and → 6.0 mol of O atoms

$\therefore$ a) N (C atoms) = $(3.0)\,(6.022 \times 10^{23})$ = 1.81×10^{24} atoms of carbon

b) N (O atoms) = $(6.0)\,(6.022 \times 10^{23})$ = 3.61×10^{24} atoms of oxygen

iv) Calculate the number of a) Ca^{2+} ions; b) NO_3^- ions and c) oxygen atoms contained in 4.0 mol of calcium nitrate ($Ca(NO_3)_2$).

4.0 mol of $Ca(NO_3)_2$ formula units contain:

→ 4.0 mol of Ca^{2+} ions
→ 8.0 mol of NO_3^- ions
and → 24.0 mol of O atoms

$\therefore$ a) N (Ca^{+2} ions) = $(4.0)\,(6.022 \times 10^{23})$ = 2.41×10^{24} Ca^{+2} ions

b) N (NO_3^- ions) = $(8.0)\,(6.022 \times 10^{23})$ = 4.82×10^{24} NO_3^- ions

c) N (O atoms) = $(24.0)\,(6.022 \times 10^{23})$ = 1.45×10^{25} O atoms

Question 6.7

How many atoms of hydrogen are there in:

(a) 1 mol of hydrogen atoms? ______________________

(b) 1 mol of hydrogen molecules? ______________________

(c) 3 mol of water molecules? ______________________

(d) 0.2 mol of ammonia molecules? ______________________

(e) 4 mol of glucose molecules ($C_6H_{12}O_6$)? ______________________

Question 6.8

You win lotto and are paid out your winnings of 1 million dollars in five cent coins (imagine that!).

(a) How many moles of five cent coins will you have? ____________________

(b) If your winnings had been 1 mole of five cent coins, how many dollars would you have?

____

____

Molar mass (M)

The mass of one mole of a substance (M) is called its molar mass. This mass is simply the relative atomic mass (A_r), or relative molecular mass (M_r), expressed in grams.

e.g. i) molar mass of sodium = mass of 1 mol of Na atoms
= 22.99 g mol^{-1}

ii) molar mass of carbon dioxide = mass of 1 mol of CO_2 molecules
= 44.01 g mol^{-1}

or, more simply

A_r (Na) = 22.99 ∴ M (Na) = 22.99 g mol^{-1}

M_r (CO_2) = 44.01 ∴ M (CO_2) = 44.01 g mol^{-1}

Calculating molar mass

Molar mass can be calculated in the same way that we calculated relative molecular mass. Simply add the individual masses of the atoms in the formula.

Worked Examples

6.5 i) Find the molar mass for sulfur dioxide.

M (SO_2) = 32.07 + 2(16.00) = 64.07 g mol^{-1}

ii) Find the molar mass for zinc carbonate.

M ($ZnCO_3$) = 65.38 + 12.01 + 3(16.00) = 125.39 g mol^{-1}

iii) Find the molar mass for iron (III) sulfate.

M[$Fe_2(SO_4)_3$] = 2(55.85) + 3(32.07) + 12(16.00) = 399.91 g mol^{-1}

Calculations – mass to moles

The number of moles of a substance in a given mass of that substance can be calculated using the following relationship:

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

n = number of moles
m = mass of substance
M = molar mass of substance

Worked Examples

6.6 i) Calculate the number of moles of sodium chloride in 200 g of this substance.

$$M\ (NaCl) = 22.99 + 35.45 = 58.44\ g\ mol^{-1}$$

$$n\ (NaCl) = \frac{m}{M} = \frac{200}{58.44} = 3.42\ mol$$

ii) Calculate the number of moles in 10.0 g of aluminium sulfate.

$$M\ (Al_2(SO_4)_3) = (2 \times 26.98) + (3 \times 32.07) + (12 \times 16.00) = 342.17\ g\ mol^{-1}$$

$$n\ (Al_2(SO_4)_3) = \frac{10}{342.17} = 2.92 \times 10^{-2}\ mol$$

Calculations – moles to mass

We can rearrange the moles / mass formula to convert moles to mass.

$$\mathbf{n = \frac{m}{M}\ \text{becomes} \rightarrow m = n \times M}$$

n = number of moles
m = mass of substance
M = molar mass of substance

Worked Examples

6.7 i) Calculate the mass of 5.0 mol of ammonia gas.

$$M\ (NH_3) = 14.01 + 3(1.008) = 17.03\ g\ mol^{-1}$$

$$\therefore m\ (NH_3) = n\ .\ M = (5.0)(17.03) = 85.2\ g$$

ii) Calculate the mass of 0.2 mol of ethanoic acid (CH_3COOH).

$$M\ (CH_3COOH) = (2 \times 12.01) + (4 \times 1.008) + (2 \times 16.00) = 60.05\ g\ mol^{-1}$$

$$\therefore m\ (CH_3COOH) = n\ .\ M = (0.20)(60.05) = 12.01\ g$$

Question 6.9

Calculate the number of moles of:

(a) sodium in 100.0 g of sodium

(b) water in 30.0 g of water

(c) methane in 4.00 g of methane

Question 6.10

Convert to moles each of the following:

(a) 10.0 g of sodium chloride

(b) 2.4×10^{-2} g of hydrogen gas

(c) 1000 g of octane (C_8H_{18})

Question 6.11

Calculate the mass of each of the following:

(a) 2.0 mol calcium hydroxide

(b) 10.0 mol of water

(c) 0.40 mol of ammonium carbonate

Question 6.12

Which of the following contains the greatest number of molecules? (Show working).

(a) 100 g of ammonia gas

(b) 200 g of oxygen gas

Moles / mass – constituents

The formula of any given substance indicates the number of atoms (or ions) present in a molecule (or formula unit) of that substance.

e.g. 1 molecule of H_2CO_3 contains:

→ 2 atoms of hydrogen
→ 1 atom of carbon
→ 3 atoms of oxygen

We can use the same logic when considering moles of a substance. Remember, a mole is just a number (6.022×10^{23}).

e.g. 1 mole of H_2CO_3 contains:

→ 2 mol of hydrogen atoms
→ 1 mol of carbon atoms
→ 3 mol of oxygen atoms

Similarly we can say:

e.g. 4.0 mol of $Zn(NO_3)_2$ contains:

4.0 mol of Zn atoms, 8.0 mol of NO_3^- ions
8.0 mol of N atoms, 24.0 mol of O atoms

Question 6.13

Complete the following:

(a) Six molecules of sulfuric acid (H_2SO_4) contain __________ atoms of hydrogen, __________ atoms of sulfur and __________ atoms of oxygen.

(b) Four mole of ethanol molecules (C_2H_5OH) contain __________ mole of carbon atoms, __________ mole of hydrogen atoms and __________ mole of oxygen atoms.

Question 6.14

A sample of calcium hydroxide has a mass of 150.0 g. For this sample, determine the:

(a) moles of $Ca(OH)_2$ ______________________________

(b) moles of OH^- ions ______________________________

(c) moles of O atoms ______________________________

(d) mass of Ca atoms ______________________________

Calculations – Number of particles / moles / mass

The following overview of mole relationships is useful in dealing with calculations such as those shown below.

number of particles N $\xleftrightarrow{n = \frac{N}{6.022 \times 10^{23}}}$ moles n $\xleftrightarrow{n = \frac{m}{M}}$ mass m

Worked Examples

6.8 1) In a laboratory experiment a balloon is filled with 0.40 g of hydrogen gas. Calculate the number of hydrogen molecules in the balloon.

STEP 1 From the diagram above we can see that we should firstly convert mass to moles.

i.e. mass $\rightarrow$ moles $\rightarrow$ number of particles

$$n(H_2) = \frac{m}{M} = \frac{0.40}{(2)(1.008)} = 0.198 \text{ mol}$$

STEP 2 Number of particles = $(0.198)(6.022 \times 10^{23})$ = molecules of H_2 gas = 1.19×10^{23}

2) A sugar cube has a mass of 1.50 g. How many molecules of sucrose ($C_{12}H_{22}O_{11}$) would be contained within this cube?

$$M(C_{12}H_{22}O_{11}) = 12(12.01) + 22(1.008) + 11(16.00) = 342.3 \text{ g mol}^{-1}$$

$$n(C_{12}H_{22}O_{11}) = \frac{m}{M} = \frac{1.50}{342.3} = 4.38 \times 10^{-3} \text{ mol}$$

$$\therefore N(C_{12}H_{22}O_{11}) = (n)(6.022 \times 10^{23}) = 2.64 \times 10^{21} \text{ molecules}$$

3) A small piece of aluminium foil is thought to contain 3.01×10^{22} atoms of aluminium. Calculate its mass.

$$n = \frac{N}{6.022 \times 10^{23}} = \frac{3.01 \times 10^{22}}{6.022 \times 10^{23}} = 0.050 \text{ mol}$$

$$\therefore m = n \,.\, M = (0.050)(26.98) = 1.349 \text{ g}$$

Question 6.15

(a) A drop of water was found to have a mass of 8.60×10^{-2} g. Determine the number of water molecules in this drop.

____

(b) A lead sinker used for fishing has a mass of 10.0 g. How many lead atoms does it contain?

____

6.3 REACTIONS, EQUATIONS AND STOICHIOMETRY

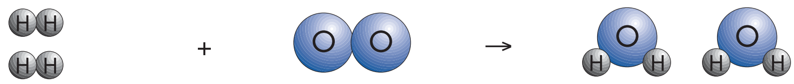
Chemical Reactions and Equations

Chemical reactions can be represented by chemical equations if all the substances involved are known. A balanced equation will tell us in what proportion (mole ratio) the reactants and products are involved.

For example we can represent the combustion of hydrogen with oxygen as follows:

word equation	hydrogen gas	+	oxygen gas	→	water vapour
molecular equation	$2H_{2(g)}$	+	$O_{2(g)}$	→	$2H_2O_{(g)}$

visual



We can see from the visual representation that the molecules of hydrogen and oxygen actually break up into atoms to form new molecules of water. We can also see that the number of atoms on the left (reactants) and the right (products) are the same showing that the Law of Conservation of Mass applies.

Most importantly the balanced equation gives us the mole relationship (mole ratio) for a particular chemical reaction. In this case the equation tells us that:

	2 mol of H_2 gas	reacts with	1 mol of O_2 gas	to give →	2 mol of H_2O gas
∴ mole ratio	2	:	1	:	2

The **mole ratio** for a particular chemical reaction is always the same no matter what actual amounts are involved. This allows us to calculate how many moles of one substance are involved if we know the molar amount of another substance in the reaction.

Calculations from Equations - 4 Easy Steps

STEP 1	Write a balanced equation.	Also indicate all known and unknown quantities.
STEP 2	Convert known quantities to moles.	As required use: $n = \frac{m}{M}$; $n = \frac{V}{22.71}$ (only for STP) ; $n = cV$. Sometimes you are given known quantities in moles and no conversion is necessary.
STEP 3	Determine moles of unknown.	Using the mole ratio from the balanced equation we have: $\therefore n_{(unknown)} = n_{(known)} \times$ mole ratio
STEP 4	Convert to required units.	The moles of the unknown you have calculated may need to be changed to mass, volume, concentration etc. Use formulae shown in step 2 as necessary.

Worked Example (mole/mole)

6.9 Propane gas burns with oxygen to form carbon dioxide gas and water vapour. If 15.0 mol of propane are burnt how many mol of carbon dioxide are produced?

STEP 1	Write a balanced equation	$C_3H_{8(g)}$	+ $5O_{2(g)}$ →	$3CO_{2(g)}$	+	$4H_2O_{(g)}$
	Indicate known/unknown	15.0 mol *(known)*		? *(unknown)*		

STEP 2 Convert known to moles: Already given

STEP 3 Determine moles of unknown: $n(CO_2) = (15.0) \times \left(\frac{3}{1}\right)$ ← *mole ratio from equation*

$= 45.0$ mol of CO_2

STEP 4 Convert to required units: OK – only moles required
Hence 45.0 mol of CO_2 produced.

Worked Example (mass/mole)

6.10 Oxygen gas can be produced from potassium chlorate as shown in the equation below. Calculate the number of moles of oxygen that would be produced from 20.0 g of potassium chlorate.

STEP 1.	Write a balanced equation	$2KClO_{3\,(s)}$	→ $2KCl_{(s)}$	+	$3O_{2(g)}$
	Indicate known/unknown	20.0 g *(known)*			? *(unknown)*

STEP 2. Convert known to moles: $n(KClO_3) = \frac{m}{M} = \frac{20.0}{122.55} = 0.163$ mol

STEP 3. Determine moles of unknown: $n(O_2) = 0.163 \times \left(\frac{3}{2}\right)$ ← *mole ratio from equation*

$= 0.245$ mol

STEP 4. Convert to required units: OK – only moles required
Hence 0.245 mol of O_2 produced.

Worked Example (mole/mass)

6.11 When sodium hydrogencarbonate is heated it decomposes to sodium carbonate, carbon dioxide and water. Calculate the mass of sodium carbonate produced from the decomposition of 5.50 mol of sodium hydrogencarbonate.

STEP 1.	Write a balanced equation	$2NaHCO_{3\,(s)}$ →	$Na_2CO_{3(s)}$	+	$CO_{2(g)}$	+	$H_2O_{(l)}$
	Indicate known/unknown	5.50 mol *(known)*	? *(unknown)*				

STEP 2. Convert known to moles: Already given

STEP 3. Determine moles of unknown: $n(Na_2CO_3) = 5.50 \times \left(\frac{1}{2}\right)$ ← *mole ratio from equation*

$= 2.75$ mol of Na_2CO_3

STEP 4 Convert to required units $m(Na_2CO_3) = (2.75)(105.99) = 291.47$

Hence 2.91×10^2 g of Na_2CO_3 are produced.

Worked Example (mass/mass)

6.12 Determine the mass of lead (II) iodide precipitated by adding excess potassium iodide to 15.0 g of lead (II) nitrate in solution.

STEP 1 Write a balanced equation $Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \rightarrow PbI_{2(s)} + 2KNO_{3(aq)}$

Indicate known/unknown: 15.0 g *(known)* ; ? *(unknown)*

STEP 2 Convert known to moles $n(Pb(NO_3)_2) = \frac{15.09}{331.2} = 0.0453$ mol

STEP 3 Determine moles of unknown $n(PbI_2) = (0.0453) \times \left(\frac{1}{1}\right)$ *mole ratio from equation*

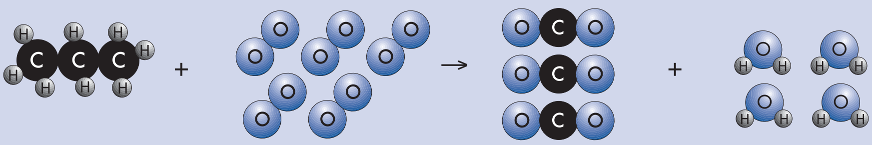
$= 0.0453$ mol

STEP 4 Convert to required units $m(PbI_2) = (0.0453)(461.0) = 20.9$ g

Hence 20.9 g of lead (II) iodide are precipitated.

Question 6.16

The reaction between propane gas and oxygen gas to produce carbon dioxide gas and water vapour is illustrated below. Complete each line as indicated.



a) Word equation

__________ + __________ → __________ + __________

b) Balanced equation

__________ + __________ → __________ + __________

c) Mole ratio

$n(C_3H_8) : n(O_2) : n(CO_2) : n(H_2O) =$ _____ : _____ : _____ : _____

d) Is the Law of Conservation of Mass obeyed? Explain.

____

e) If 15.0 mol of CO_2 gas are produced in this way how many mol of water are also produced?

____

Question 6.17

When butane gas, C_4H_{10}, (commonly used as a cigarette lighter fuel) burns in air, the gases produced are carbon dioxide and water vapour. Complete all the steps indicated below to determine the mass of water vapour produced by burning 10.0 g of butane gas.

1. Write a balanced equation.

____

2. Convert known to moles.

____

3. Determine moles of unknown.

____

4. Convert to required units.

____

Hence ____

Calculations involving gas volumes

For this part of the course we will only deal with gases at STP. The molar volume of all gases is taken as 22.71 L at STP. Hence for any mol/volume conversions use V = (n)(22.71). It is also possible to deal with gases under different conditions using PV = nRT.

When only gas volumes are involved in a calculation from an equation we can simply use volume ratios instead of converting to moles. Since the molar volume is the same for all gases it follows that in chemical reactions involving gases the following applies:

	mole ratio	=	volume ratio		
For example:	$2SO_{2(g)}$	+	$O_{2(g)}$	→	$2SO_{3(g)}$
	2 mol		1 mol		2 mol
	2 volumes		1 volume		2 volumes

Worked Example (volume/volume)

6.13 Determine the volume of oxygen consumed when 5.0 L of natural gas (methane) is burnt. Assume all volumes are measured at STP.

Step 1	Write a balanced equation	$CH_{4(g)}$	+	$2O_{2(g)}$	→	$CO_{2(g)}$	+	$2H_2O_{(g)}$
	Indicate known/unknown	5.0 L *(known)*		? *(unknown)*				

Step 2 Determine mole ratio $\dfrac{n\,(O_2)}{n\,(CH_4)} = \dfrac{2}{1}$

STEP 3 Hence volume ratio $= \frac{2}{1}$

STEP 4 Determine volume of unknown $\text{Vol}(O_2) = (5.0)\left(\frac{2}{1}\right)$ ← *mole ratio from equation*

$= 10.0 \text{ L}$

Hence 10.0 L of O_2 gas at STP are consumed.

Worked Example (mass/volume)

6.14 Determine the volume of ammonia gas at STP produced when 15.0 g of hydrogen gas reacts with excess nitrogen gas.

STEP 1 Write a balanced equation $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$

Indicate known/unknown: $3H_2$: 15.0 g *(known)*; $2NH_3$: ? *(unknown)*

STEP 2 Convert known to moles $n(H_2) = \frac{15.0 \text{ g}}{2.016} = 7.44 \text{ mol}$

STEP 3 Determine moles of unknown $n(NH_3) = (7.44) \times \left(\frac{2}{3}\right)$ ← *mole ratio from equation*

$= 4.96 \text{ mol}$

STEP 4 Convert to required units $V(NH_3) = (4.96)(22.71)$

$= 112.6 \text{ L at STP}$

Hence 1.13×10^2 L of NH_3 gas are produced at STP.

Worked Example (volume/mass)

6.15 Determine the mass of carbon dioxide produced when 10.0 L of acetylene gas at STP (C_2H_2) is completely burnt in air.

STEP 1 Write a balanced equation $2C_2H_{2(g)} + 5O_{2(g)} \rightarrow 4CO_{2(g)} + 2H_2O_{(g)}$

Indicate known/unknown: $2C_2H_2$: 10.0 L *(known)*; $4CO_2$: ? *(unknown)*

STEP 2 Convert known to moles $n(C_2H_2) = \frac{10.0 \text{ L}}{22.71} = 0.440 \text{ mol}$

STEP 3 Determine moles of unknown $n(CO_2) = (0.440) \times \left(\frac{4}{2}\right)$ ← *mole ratio from equation*

$= 0.880 \text{ mol}$

STEP 4 Convert to required units $m(CO_2) = (0.880)(44.01)$

$= 38.8 \text{ g}$

Hence 38.8 g of CO_2 gas are produced.

Question 6.18

Carbon monoxide gas burns readily in oxygen gas to produce carbon dioxide gas.

(a) Give a balanced equation for the reaction.

____

(b) Give the mole ratios involved in the reaction.

$n(CO) : n(O_2) : n(CO_2) =$ ______ : ______ : ______

(c) Give the volume ratios involved in the reaction.

$Vol(CO) : Vol(O_2) : Vol(CO_2) =$ ______ : ______ : ______

(d) If 300 ml of CO gas are burnt determine the volume of O_2 gas required.

____

Calculations involving solutions

When different solutions are combined precipitates may form. Calculations often involve finding the mass of precipitates and concentrations of the remaining ions. Refer also to Chapter 9 on concentrations of solutions.

Worked Example

6.16 Excess silver nitrate solution is added to 200.0 ml of 0.35 mol L^{-1} sodium chloride solution and a precipitate forms. Assuming all available chloride ions have reacted and that the total new volume of the solution is 300.0 mL determine: (a) the mass of precipitate formed; (b) the concentration of the Na^+ ion in the final solution.

(a)

STEP 1	Write a balanced equation	$AgNO_{3(aq)}$ +	$NaCl_{(aq)}$	$\rightarrow$	$AgCl_{(s)}$ +	$NaNO_{3(aq)}$
	Indicate known/unknown	excess	200 mL 0.35 mol L^{-1} *(known)*		? *(unknown)*	

STEP 2 Convert known to moles

$n(NaCl) = (0.35)(0.200) = 7.00 \times 10^{-2}$ mol

$\therefore n(Na^+) = n(Cl^-) = 7.00 \times 10^{-2}$ mol

STEP 3 Determine moles of unknown

$n(AgCl) = (7.00 \times 10^{-2}) \times \left(\frac{1}{1}\right)$ ← *mole ratio from equation*

STEP 4 Convert to required units

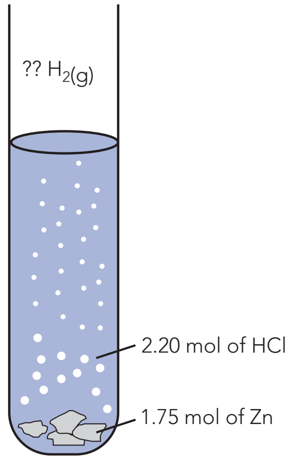
$m(AgCl) = (7.00 \times 10^{-2})(143.35) = 10.03$ g

Hence mass of AgCl precipitated is 10.03 g.

(b) To find the final concentration of the Na^+ ions in solution we simply divide the moles of Na^+ by the total volume of the solution (L).

$$c(Na^+) = \frac{n}{V} = \frac{7.00 \times 10^{-2}\ \text{mol}}{0.300\ \text{L}} = 0.233\ \text{mol L}^{-1}$$

Calculations involving a limiting reagent



The limiting reagent is the reactant that is consumed first. It is the reactant that runs out and when there is none of it left, the reaction stops.

(1) Recognising limiting reagent questions

Zn	+	2HCl	→	$ZnCl_2$	+	H_2
1.75 mol		2.20 mol				??

There is enough information provided to calculate the number of moles of two reactants.

(2) Identifying the limiting reagent

Divide the number of moles of each reactant by coefficient for that reactant in the balanced equation. The smallest answer is the limiting reagent.

Zn: $= \frac{1.75}{1} = 1.75$ HCl: $\frac{2.20}{2} = 1.10$ HCl is the limiting reagent.

(3) Only the limiting reagent is used to calculate n(unknown)

$n(H_2)$ produced $= ½ \times n(HCl)$ used
$= 1.10$

Worked Example

6.17 Determine the mass of $Ca(NO_3)_2$ produced when 5.40 g of $CaCO_3$ reacts with 6.30 g of HNO_3

STEP 1: Write a balanced equation and identify knowns and unknowns

$CaCO_3 + 2HNO_3 \rightarrow Ca(NO_3)_2 + CO_2 + H_2O$
5.40 g 6.30 g ? g

STEP 2: Calculate n (known)

$n(CaCO_3) = \frac{m}{M} = \frac{5.40}{100.09} = 0.0540$

$n(HNO_3) = \frac{m}{M} = \frac{6.30}{63.018} = 0.100$

STEP 2-1: Identify the limiting reagent by finding the smallest number of moles when each reactant is divided by the coefficient used to balance it in the balanced equation.

$\frac{n(CaCO_3)}{1} = \frac{0.0540}{1} = 0.540$

$\frac{n(HNO_3)}{2} = \frac{0.100}{2} = 0.500$ $\therefore$ HNO_3 is the limiting reagent (LR)

STEP 3: Use the LR to determine n (unknown)

$n(Ca(NO_3)_2) = ½\, n(HNO_3)$

$= ½ \times 0.100 = 0.0500$

STEP 4: Convert to required units

$m(Ca(NO_3)_2) = n.M$

$= 0.0500 \times 164.1$

$= m(Ca(NO_3)_2) = 8.21$ g

REVIEW QUESTIONS

Chapter 6: Chemistry Calculations

RELATIVE MASS

1. For each of the following sets of elements select the heaviest atom:
   (a) phosphorus, potassium, titanium
   (b) neon, argon, krypton
   (c) iodine, chlorine, fluorine

2. Determine approximately:
   (a) how many atoms of helium would weigh as much as one atom of oxygen;
   (b) how many atoms of boron would weigh as much as one atom of silver;
   (c) how many atoms of hydrogen would weigh as much as one atom of carbon.

3. Determine the relative molecular mass (M_r) for each of the following molecular compounds:
   (a) carbon dioxide
   (b) oxygen gas
   (c) sulfur dioxide
   (d) ammonia gas

4. Determine the relative formula mass (M_r) for the following:
   (a) calcium nitrate
   (b) sodium hydroxide
   (c) ammonium carbonate
   (d) iron (III) chloride

5. Determine which of the following molecules:
   NH_3 , C_3H_8 , H_2O , CO , C_2H_5OH , CH_3COOH
   (a) is the heaviest;
   (b) has approximately the same mass as 5 carbon atoms;
   (c) has approximately the same mass as 17 hydrogen atoms.

PERCENTAGE COMPOSITION

6. Determine the percentage by mass of each element in the following compounds:
   (a) aluminium oxide
   (b) water
   (c) sodium chloride
   (d) propane (C_3H_8)

7. Determine the percentage by mass of water in the following hydrated compounds:
   (a) calcium sulfate-2-water
   (b) iron (III) chloride-3-water

8. Some common fuels burnt for energy are:
   - natural gas (methane – CH_4)
   - bottled gas (propane – C_3H_8)
   - acetylene gas (ethyne – C_2H_2)
   - petrol (mostly octane – C_8H_{18})

   Determine which fuel has the:
   (a) highest percentage by mass of carbon
   (b) highest percentage by mass of hydrogen.

Moles / particles

9. Calculate the number of moles of:
   (a) sodium atoms in 6.022×10^{22} atoms of sodium
   (b) copper atoms in 2.41×10^{24} atoms of copper
   (c) oxygen molecules in 1.20×10^{23} molecules of oxygen.

10. Calculate the number of:
   (a) calcium atoms in 0.25 mol of calcium
   (b) carbon monoxide molecules in 5.50 mol of carbon monoxide
   (c) hydrogen atoms in 4.0 mol of water molecules
   (d) oxygen atoms in 1.25 mol of nitric acid.

11. Which of the following contains the greatest number of atoms?
   (a) 3 mol of hydrogen gas
   (b) 8 mol of lead metal
   (c) 2 mol of carbon dioxide gas
   (d) 1 mol of ammonium nitrate

12. It was found that 36 g of carbon contained three times as many atoms as 108 g of silver.
   (a) Which element has the heaviest atoms?
   (b) By what factor are they heavier?

Moles / mass / particles

13. Determine the molar mass of:
   (a) chlorine gas
   (b) carbon dioxide
   (c) ammonia gas
   (d) sulfur trioxide gas

14. Determine the molar mass of :
   (a) zinc oxide
   (b) magnesium hydrogencarbonate
   (c) aluminium nitrate
   (d) sodium carbonate-10-water

15. Explain the difference between relative molecular mass (M_r) and molar mass (M).

16. Calculate the number of moles of:
   (a) carbon dioxide in 440 g of carbon dioxide
   (b) sulfuric acid in 49.0 g of sulfuric acid
   (c) propane in 8.50 kg of propane C_3H_8.

17. Convert to moles.
   (a) 500 g of water
   (b) 2.15 g of silver chloride
   (c) 100.0 g of ethanoic acid (CH_3COOH)

18. Which contains the greatest number of molecules, 20.0 g of oxygen gas or 20.0 g of carbon dioxide gas?

19. Some common gases are ammonia, carbon dioxide, carbon monoxide, chlorine, hydrogen, nitrogen, oxygen and sulfur dioxide.

Calculate the molar mass of each gas and then list them from lightest to heaviest:

20. Determine the number of moles of hydrogen in each of the following:
  (a) 1.0 mol of hydrogen gas
  (b) 4.0 mol of ammonia gas
  (c) 2.0×10^{-2} mol of ethanoic acid (CH_3COOH)

21. Determine the number of moles of each element in:
  (a) 1.50 mol of nitric acid
  (b) 50.0 g of hydrochloric acid
  (c) 100.0 g of glucose ($C_6H_{12}O_6$).

22. Determine the number of moles of each ion in the following ionic compounds:
  (a) 2.00 g of lead (II) chloride
  (b) 2.00×10^3 g of iron (III) oxide
  (c) 5.25×10^{-3} g of aluminium sulfate

23. Determine how many atoms in total are contained in:
  (a) 10.0 g of water
  (b) 10.0 g of hydrogen gas
  (c) 20.0 g of sulfur dioxide gas

24. Determine the number of molecules in:
  (a) 90.0 g of chlorine gas
  (b) 12.0 g of hydrogen gas
  (c) 12.0 g of ethyne gas (C_2H_2)

25. Determine the mass of:
  (a) 6.022×10^{23} molecules of ammonia gas
  (b) 3.011×10^{24} atoms of sodium
  (c) an ionic compound containing 4.0×10^{24} Mg^{2+} ions and 8.0×10^{24} OH^- ions

CALCULATIONS BASED ON EQUATIONS

Mole / mole

26. Ethane (C_2H_6) burns in oxygen to produce carbon dioxide gas and water vapour.
  (a) Give a balanced equation for the reaction.
  (b) What do the coefficients in front of each formula indicate?
  (c) How many moles of carbon dioxide gas would be produced when 25.0 mol of ethane burns?

27. Octane, the main constituent of petrol is combined with oxygen gas as follows:

$$2C_8H_{18(g)} + 25O_{2(g)} \rightarrow 16CO_{2(g)} + 18H_2O_{(g)}$$

Assuming 0.25 mol of octane are burnt, calculate:

  (a) the number of moles of oxygen gas required.
  (b) the total number of moles of products formed.

28. Magnesium burns in air to form magnesium oxide. Determine how many moles of oxygen gas would be required to produce 1.65×10^{-3} mol of magnesium oxide.

Mole / mass or mass / mole

29. Determine the mass of water vapour produced from the complete combustion of 1.50×10^{-2} mol of butane (C_4H_{10}).

30. Determine the mass of copper (II) carbonate required to produce 5.46 mol of carbon dioxide from the following reaction:

$$CuCO_{3(s)} \xrightarrow{heat} CuO_{(s)} + CO_{2(g)}$$

31. Calcium carbonate reacts with hydrochloric acid as shown below:

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

Determine:
(a) the mass of carbon dioxide that would be produced from 2.40×10^{-3} mol of calcium carbonate.
(b) the mass of hydrochloric acid needed to produce 1.50 mol of water.
(c) the mass of calcium chloride produced from 6.40×10^{-2} mol of hydrochloric acid.

32. Magnesium metal reacts with hydrochloric acid to produce magnesium chloride and hydrogen gas.
(a) Give a balanced equation for the reaction.
(b) Calculate the number of moles of hydrochloric acid necessary to consume 4.26 g of this metal.
(c) How many moles of magnesium chloride would be produced?

33. During photosynthesis, green plants are able to convert carbon dioxide and water to glucose and oxygen.

$$6CO_{2(g)} + 6H_2O_{(l)} \xrightarrow{sunlight} C_6H_{12}O_{6(aq)} + 6O_{2(g)}$$

(a) In order for plants to produce 1.00 g of glucose, how many moles of carbon dioxide would be required?
(b) How many moles of oxygen would be produced?

Mass/mass

34. Calculate the mass of silver metal produced when 14.5 g of copper are dissolved in silver nitrate according to the following reaction:

$$2Ag^{+}_{(aq)} + Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2Ag_{(s)}$$

35. A typical barbecue gas cylinder contains 9.00 kg of compressed propane gas (C_3H_8).

(a) Write a balanced equation for the reaction of propane gas when it burns in air to form carbon dioxide gas and water vapour.
(b) Assuming all the gas in the cylinder is burnt completely as per your equation in (a), calculate the total mass of products released into the air.

36. Solid ammonium carbonate decomposes as follows:

$$(NH_4)_2CO_{3(s)} \xrightarrow{heat} 2NH_{3(g)} + CO_{2(g)} + H_2O_{(g)}$$

(a) If a 6.42 g sample of ammonium carbonate is completely decomposed, determine the mass of ammonia gas produced.
(b) What should be the total mass of the products?

37. Chalcopyrite ($CuFeS_2$) is an important mineral from which copper metal can be recovered. One of the first steps in the recovery of the copper is to roast it in air to remove the iron and some of the sulfur. The reaction that occurs is:

$$2CuFeS_{2(s)} + 3O_{2(s)} \rightarrow 2CuS_{(l)} + 2FeO_{(l)} + 2SO_{2(g)}$$

(a) For each tonne of chalcopyrite mineral completely reacted as above, calculate the mass of copper sulfide produced.
(b) Determine the mass of copper which could be recovered from the copper sulfide.

[Hint: Perhaps consider % composition]

Gases

38. One of the steps in the manufacture of nitric acid involves the oxidation of ammonia gas as follows:

$$4NH_{3(g)} + 5O_{2(g)} \rightarrow 4NO_{(g)} + 6H_2O_{(g)}$$

(a) If 200 L of ammonia gas are reacted, determine the volume of oxygen gas required.
(b) To produce 10 kL of nitrogen monoxide, what volume of ammonia would be required? (Assume STP for all volumes.)

39. Ethyne gas (C_2H_2) is burnt completely in oxygen.
(a) Write a balanced equation for the reaction.
(b) If 10.0 L of ethyne was burnt, determine the volume of oxygen required. (Assume all gas volumes are at STP.)

40. The fertiliser commonly called sulfate of ammonia $(NH_4)_2SO_4$ is manufactured using ammonia and sulfuric acid. The equation is as follows:

$$2NH_{3(g)} + H_2SO_{4\,(aq)} \rightarrow (NH_4)_2SO_{4(aq)}$$

(a) Calculate the mass of ammonium sulfate that could be produced from 25.0 kL of ammonia gas at STP.
(b) What mass of sulfuric acid would be required?

41. Aluminium metal reacts with hydrochloric acid to produce aluminium chloride and hydrogen gas. If a student wanted to produce 500 mL of this gas at STP, determine the mass of aluminium required.

Solutions

42. Excess sulfuric acid solution was added to 240 mL of 0.150 mol L^{-1} lead (II) nitrate solution and a precipitate forms. Assuming all the lead (II) ions reacted, calculate the mass of precipitate formed.

43. A solution containing 12.2 g of sodium carbonate has 150.0 mL of 1.20 mol L^{-1} nitric acid added to it. The reaction which occurs is as follows:

$$Na_2CO_{3(aq)} + 2HNO_{3(aq)} \rightarrow 2NaNO_{3\,(aq)} + CO_{2(g)} + H_2O_{(l)}$$

(a) Assuming all the nitric acid is consumed, determine the mass of sodium carbonate that reacted.
(b) What further volume of nitric acid solution will be needed to completely react all the sodium carbonate?

44. An excess quantity of zinc metal is placed in a solution containing 125 mL of concentrated nitric acid (6.0 mol L^{-1}). The equation for the reaction is:

$$Zn_{(s)} + 2NO_3^-{}_{(aq)} + 4H^+{}_{(aq)} \rightarrow Zn^{2+}{}_{(aq)} + 2NO_{2(g)} + 2H_2O_{(l)}$$

Assuming all the nitric acid reacted with the zinc as shown in the equation above, determine:
(a) mass of zinc that reacted
(b) mass of zinc nitrate that could be recovered from solution.

DIFFICULT

45. A solution containing 200 mL of 1.0 mol L^{-1} sodium carbonate solution is added to 200 mL of 1.0 mol L^{-1} barium chloride solution.
(a) Will a precipitate form? If so, write an equation for the reaction.
(b) Calculate the mass of any precipitate
(c) Determine the concentration of any remaining ions in solution.
[Hint: Check solubility table]

FOR THE EXPERTS

An essential element for plant growth is nitrogen. Plants use this element to make proteins and other organic molecules needed for good growth. Plants are not able to absorb nitrogen from the air even though the air is made up of 78% nitrogen, but acquire it from soluble nitrogen compounds in the soil.

Some common fertilisers that are used in the home garden are:

- ammonium sulfate, $(NH_4)_2SO_4$ (called sulfate of ammonia)
- ammonium nitrate, NH_4NO_3 (called nitrate of ammonia)
- ammonium phosphate, $(NH_4)_3PO_4$ (called potash)
- diaminomethanal, $CO(NH_2)_2$ (called urea).

46.

(a) Calculate the percentage by mass of nitrogen in each of the fertilisers listed above.

(b) Rita is applying some sulfate of ammonia to her front lawn. The directions on the bag say to use 45 grams per square metre of lawn.

(i) If her rectangular lawn is approximately 4.5 m by 12.0 m, what mass of fertiliser should she use?

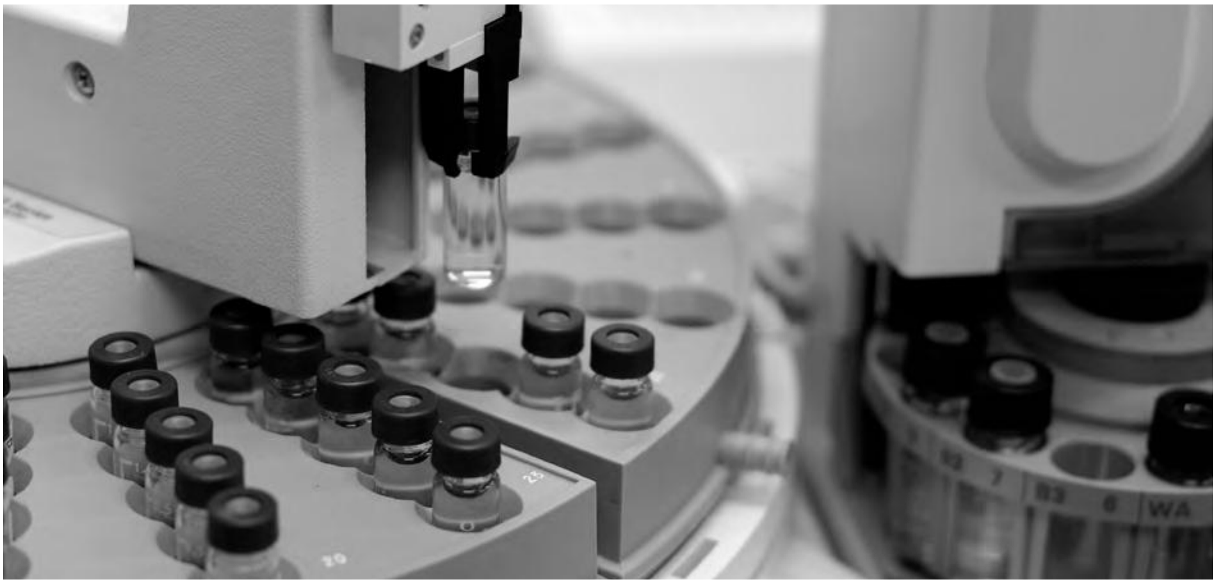
(ii) What total mass of nitrogen would be in this amount of fertiliser?

CHEMISTRY

UNIT 2

INTERMOLECULAR FORCES

7



Topics covered in this chapter:

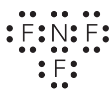
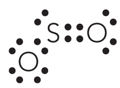
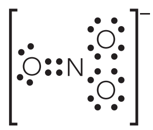
7.1 Drawing electron dot diagrams
7.2 Polarity and shapes of molecules
7.3 Intermolecular forces
7.4 Chromatography

Revision required: *This section requires a good understanding of covalent bonding. You may need to revise Chapter 3 of this study guide, in particular the sections on electron dot diagrams, the covalent bond and covalent molecular substances.*

7.1 DRAWING ELECTRON DOT DIAGRAMS

The electron dot diagrams for simple molecules are usually easy to draw. However, the more 'complex' examples can often involve a long and tedious procedure of trial and error. It is possible however to determine the electron dot diagram by a systematic process as shown below.

Table 7.1 *A systematic approach to drawing electron dot diagrams.*

PROCESS	NF_3	SO_2	CO_2	NO_3^-
1. Determine total number of valence electrons available.	5 + 7 + 7 + 7 = 26	6 + 6 + 6 = 18	4 + 6 + 6 = 16	5 + 6 + 6 + 6 + 1 = 24
2. Determine number of electrons required (octet rule).	4 (atoms) × 8 = 32	3 (atoms) × 8 = 24	3 (atoms) × 8 = 24	4 (atoms) × 8 = 32
3. Difference between 1 and 2 indicates the number of electrons to be shared.	32 – 26 = 6	24 – 18 = 6	24 – 16 = 8	32 – 24 = 8
4. Hence, number of bonds required (2 electrons = 1 bond).	3 3 single bonds	3 1 single bond 1 double bond	4 2 double bonds	4 2 single bonds 1 double bond
5. Hence, electron dot diagram. Note: • lone atom is usually central atom. • double bonds as necessary to achieve total number of bonds required.				

Question 7.1

Sketch electron dot diagrams for the following:

(a) NH_3

(b) Cl_2O

(c) CCl_4

(d) CH_3OH

Question 7.2

Use the systematic process shown in Table 2.3 to determine the electron dot diagram for the following:

(a) N_2

(b) HCN

(c) CS_2

(d) SO_3

(e) SO_3^{2-}

(f) ClO_3^-

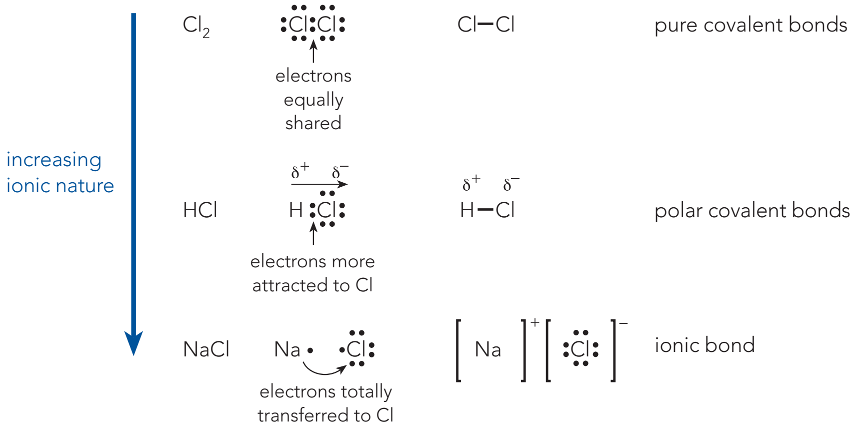
7.2 POLARITY AND SHAPES OF MOLECULES

Polarity of bonds

When a bond forms between atoms that are different, an unequal sharing of electrons takes place. This is because the atoms have different electron attracting abilities (electronegativity) and hence a polar bond results.

The nature of a bond depends on the differences (or otherwise) of the electronegativities of the atoms (see Table 7.2). Bonds may be purely covalent, polar covalent or ionic.

e.g.



Electronegativity

The electronegativity of an atom is a relative measure of its electron attracting ability in a bond. When considering the elements in the periodic table we can see the following trends:

- Electronegativities are greater for non-metals. Note that noble gases are not considered electronegative.
- Electronegativities decrease as we go down any particular group.

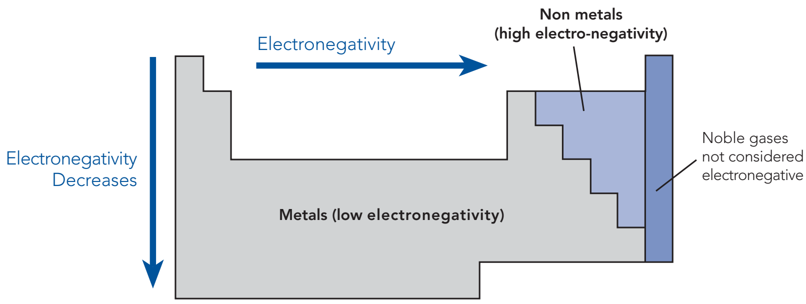


Figure 7.1 *Trends in electronegativity.*

Table 7.2 *Electronegativity values for some elements.*

H							He
2.1							–
Li	Be	B	C	N	O	F	Ne
1.0	1.5	2.0	2.5	3.0	3.5	4.0	–
Na	Mg	Al	Si	P	S	Cl	Ar
0.9	1.2	1.5	1.8	2.1	2.5	3.2	–

Shapes of molecules - VSEPR theory

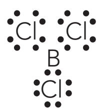
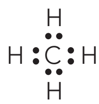
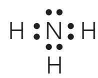
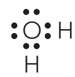
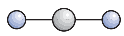
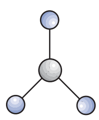
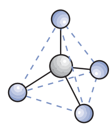
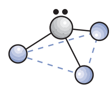
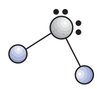
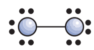
The shape of a molecule is determined by the influence that the electron pairs around the central atom have on each other. The Valence Shell Electron Pair Repulsion (VSEPR) theory states that:

- pairs of outer shell electrons in atoms form charge clouds which are roughly spherical in shape
- these charge clouds repel each other and so are positioned as far apart as possible. This includes both bonding and non bonding pairs of electrons.

Hence to determine the shape of a molecule we need to:

- draw the electron dot structure
- check on how many charge clouds (electron pairs) are around the central atom - this will tell us how they are arranged (i.e. basic shape)
- check if there are lone pairs of electrons around the central atom as this will affect shape of the molecule.

Table 7.3 The shapes of simple molecules.

EXAMPLE	$BeCl_2$	BCl_3	CH_4	NH_3	H_2O	Cl_2
Electron dot diagram						
No. of pairs of electrons around central atom	2	3	4	4	4	4
Non-bonding pairs of electrons	–	–	–	1	2	3
Shape diagram						
Bond angle	180°	120°	109°	107°	104.5°	180°
Shape name	linear	triangular planar	tetrahedral or trigonal planar	pyramidal	bent (v shaped)	linear

Question 7.3

Sketch electron dot diagrams for:

(a) BeF_2

(b) PCl_3

(c) C_2H_4

(d) C_3H_8

Question 7.4

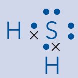
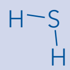
Use the information in Table 7.3 to arrange the following bonds in order of increasing polarity (i.e. least polar first):

H–F , O–H , F–F , N–H , B–H , N–O , C–O , C–S

Answer: ____

Question 7.5

Draw electron dot diagrams for the following molecules and ions. Also draw and name the shape in each case. The first is done for you.

Example	Electron dot diagram	Central atom electron pairs lone pairs	Shape diagram	Shape name
(a) H_2S		4 charge clouds 2 lone pairs		bent shape
(b) HBr				
(c) SF_2				
(d) BH_3				
(e) SO_4^{2-}				
(f) SO_3^{2-}				
(g) CS_2				
(h) CH_2Br_2				
(i) C_2H_6				
(j) C_2Cl_4				

Polar molecules

A molecule may contain polar bonds but this does not necessarily mean it is itself polar. This depends on the symmetry, or otherwise, of the polar bonds within the molecule. To determine whether a molecule is polar you need to:

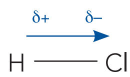
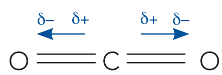
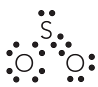
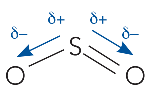
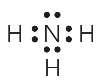
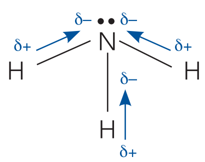
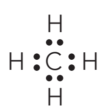
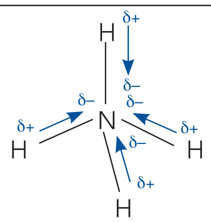
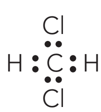
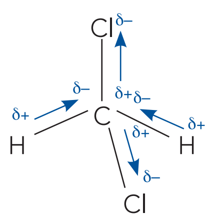
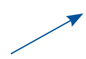
- draw an electron dot diagram of the molecule
- determine the polarity (or otherwise) of each bond.

If the two atoms involved in a covalent bond are the same then the bond is non polar. If the two are different then the bond will be polar since they will attract the bonding electrons differently. The atom with the highest electronegativity will attract the bonding electrons the most. The polar bond created, or dipole, can be indicated with arrow pointing to the negative end ($\overset{\delta+\quad\delta-}{\longrightarrow}$).

- determine from the symmetry (or otherwise) whether an overall molecular dipole exists.

Symmetrical dipoles cancel out ($\therefore$ non-polar). Asymmetrical dipoles combine like vectors to give an overall dipole.

Worked Example 7.1

EXAMPLE	ELECTRON DOT DIAGRAM	MOLECULAR SHAPE SHOWING POLARITY OF BONDS	OVERALL MOLECULAR DIPOLE	POLARITY OF MOLECULE
HCl				polar
CO_2			zero	non-polar
SO_2				polar
NH_3				polar
CH_4			zero	non-polar
CH_2Cl_2			 direction midway between 2 Cl atoms	polar

Question 7.6

Complete the table below and determine the polarity, or otherwise, of the examples given.

Example	Electron dot diagram	Molecular shape showing bonds	Overall molecular dipole	Polarity of molecule
1. H_2O				
2. N_2				
3. BF_3				
4. NF_3				
5. CS_2				
6. CCl_4				

Covalent molecular substances – writing correct formula

These substances are composed of non-metal atoms covalently bonded to other non-metal atoms. The formula and name indicates the number and type of atoms present in a molecule of that substance.

mono	=	1
di	=	2
tri	=	3
tetra	=	4
penta	=	5
hexa	=	6
hepta	=	7
octa	=	8
nona	=	9
deca	=	10

Rules for naming

- The element that is closer to the left side of the periodic table is named first and it keeps its normal name. (If both elements are in the same group – the element further towards the bottom is named first.)
- The element closer to the right hand side of the periodic table is named second and has the end of its name changed to end in – ide.
- Prefixes are used to indicate if more than one atom of that element is present in the molecule.
- If there is only one atom of the first element, the prefix mono is not used.
- Molecules composed of one element only, get the name of that element.

e.g.

2 nitrogen atoms

N_2O_4 **di**nitrogen **tetra**oxide

4 oxygen atoms

1 phosphorus atom

PCl_3 phosphorus **tri**chloride

3 chlorine atoms

7.3 INTERMOLECULAR FORCES

The attractive forces that exist between molecules in covalent molecular substances are very weak. This explains why these substances have such low melting and boiling points.

There are no strong electrostatic forces between the molecules (they are neutral), but weak intermolecular forces do exist. These intermolecular forces, however, must not be confused with the strong covalent bonds existing *within* the molecule.

The three types of intermolecular forces are *dispersion forces*, *dipole-dipole forces* and *hydrogen bonding*. They are collectively referred to as **van der Waals forces**.

Dispersion forces

These are the weakest type of intermolecular force and act between molecules. They become significant when they are the only forces of attraction such as is the case between atoms of noble gases and between non polar molecules. They also become significant as molecules become larger.

Dispersion forces are due to temporary dipoles created when molecules come close together. These dipoles arise from the uneven spread, or dispersion, of the constantly moving electrons.

The strength of dispersion forces is greatest for:

- **molecules with a greater number of electrons**

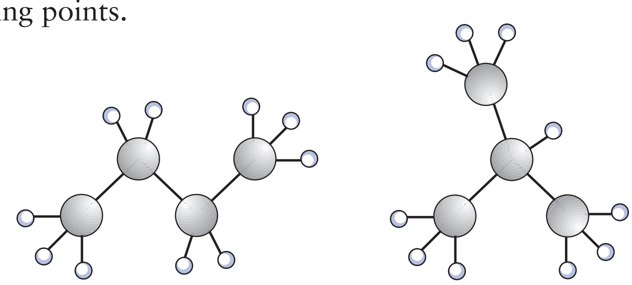
  For example, the melting and boiling points of the noble gases increases with their atomic mass and number of electrons. This is also the case for such molecules as the halogens and the alkanes.

- **molecules of linear, rather than compact shape**

  Where molecules are linear, there is more surface interaction than exists between similar sized but more compact molecules. For example, the isomers butane and methyl propane have the same molecular mass but different melting and boiling points.

***Table 7.4** Melting and boiling points of the inert gases.*

NOBLE GAS	NUMBER OF ELECTRONS	MELTING POINT °C	BOILING POINT °C
He	2	−270	−269
Ne	10	−249	−246
Ar	18	−189	−186
Kr	36	−157	−152
Xe	54	−112	−108



Butane (C_4H_{10})
MP –138°C, BP –0.5°C

Methyl propane (C_4H_{10})
MP –160°C, BP –12°C

***Figure 7.2** Structure of butane and methyl propane*

Dipole-dipole forces

Polar molecules such as HCl and HBr have an electron arrangement which causes one end of the molecules to be positively charged while the other end is negatively charged. They are referred to as dipoles. Dipole-dipole attraction occurs between polar molecules and is a significant intermolecular force for small molecules. This causes their melting and boiling points to be significantly higher than similar sized non polar molecules.

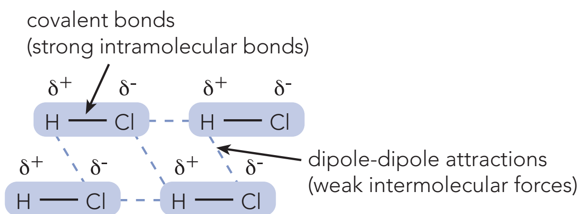


***Figure 7.3** Dipole-dipole forces between HCl molecules*

Hydrogen bonding

Hydrogen bonding is an extreme form of dipole-dipole attraction. It occurs between molecules which have hydrogen covalently bonded to N, O, or F atoms. The hydrogen bond is between the hydrogen atom on one molecule and a N, O or F atom on a *neighbouring* molecule.

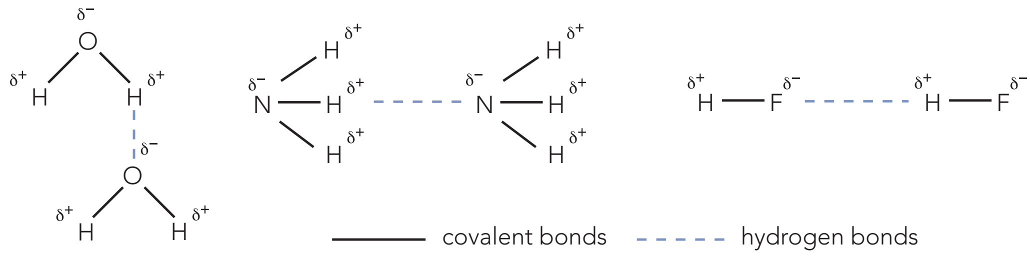


Figure 7.4 *Hydrogen bonding in H_2O, NH_3 and HF.*

Important points to note:

- hydrogen bonding must not be confused with covalent bonding. It is only an intermolecular force although its strength is greater than dispersion forces or weak dipole-dipole forces.
- it only occurs *between* molecules which have H atoms covalently bonded to N, O or F atoms within a single molecule.
- it strongly influences the properties of substances like H_2O, NH_3, HF and many organic molecules. For example, the boiling points of these substances is much higher than would otherwise be predicted (see graph).

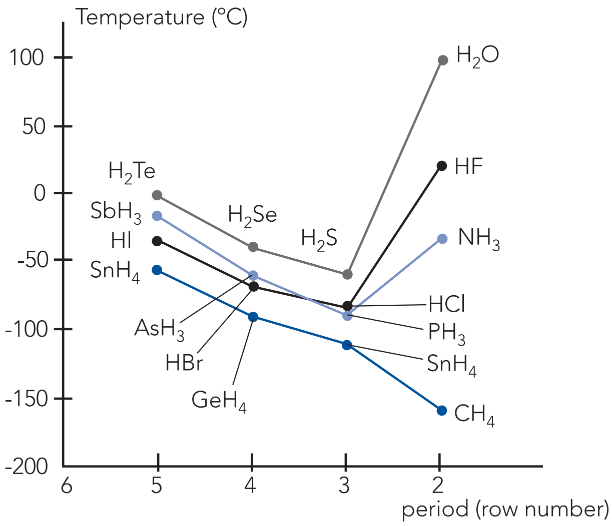


Figure 7.5 *Trends in the boiling points of groups 14, 15, 16 and 17 hydrides. The unexpectedly high boiling points of H_2O, HF and NH_3 are due to hydrogen bonding. The boiling point of CH_4 follows the expected trend since there is no hydrogen bonding*

Intermolecular forces – points to note:

Not as strong as ionic, covalent or metallic bonding but are significant attractive forces between all molecules in the solid and liquid states.

- hydrogen bonding > *stronger than* > weak dipole-dipole forces > *stronger than* > dispersion forces
- Dispersion forces are present between **all** molecules:
  – depend on molecular size and shape
  – are the **only** forces between non polar molecules.

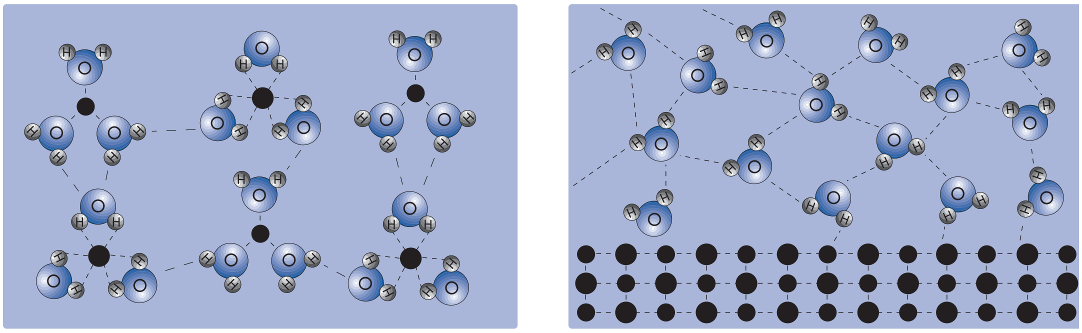
- Dipole-dipole forces are usually stronger than dispersion forces and are particularly significant in small polar molecules.
- Hydrogen bonding is the strongest of the van der Waals forces but is weak compared to covalent bonding. It explains the high melting and boiling points of H_2O, NH_3, HF and organic molecules such as alcohols.

Solubility of substances

When a substance (solute) dissolves in another (solvent) a solution results.

solute + solvent ⟶ solution

For this process to occur, there must exist a significant force of attraction between the solute and solvent particles. This force must overcome the force of attraction between the solute-solute and solvent-solvent particles. For a substance to dissolve in water for example the forces of attraction between the solute particles and the water molecules are enough to separate the solute particles from each other and allow them to spread throughout the water.



(a) Soluble substance
The solute particles are attracted to the water particles strongly enough to cause the solute particles to separate from each other and then spread uniformly throughout the water.

(b) Insoluble substance
The attraction between solute particles and water particles is not strong enough to separate solute particles from each other or it is not enough to force them to spread water molecules apart so that a homogeneous mixture can form.

Figure 7.6 Solubility depends on the strength of attraction between particles. These attractions may be solute-solute, solvent-solvent or solute-solvent.

Predicting solubility using bonding

We can make general predicitions about the solubility of a substance if we know the type of bonding that exists in both the solute and the solvent. Where the type of bonding is similar the solute will dissolve, where the bonding is different it will not. This general rule is often expressed as **like dissolves like**. Although there are exceptions the following is helpful in determining if dissolution occurs.

polar solvents will dissolve polar solutes.

e.g. Ammonia (NH_3), ethanol (CH_3CH_2OH) and sugar ($C_{12}H_{22}O_{11}$) are all polar molecules and readily dissolve in water which is also a polar molecule.
Methane (CH_4), hydrogen (H_2) and petrol ($\approx C_8H_{18}$) are all non polar molecules and do not dissolve in water.

non-polar solvents will dissolve non-polar solutes.

e.g. Oil and wax, which are non-polar substances, will readily dissolve in kerosene which is made up of non-polar molecules. Oil and wax however, do not dissolve in water.

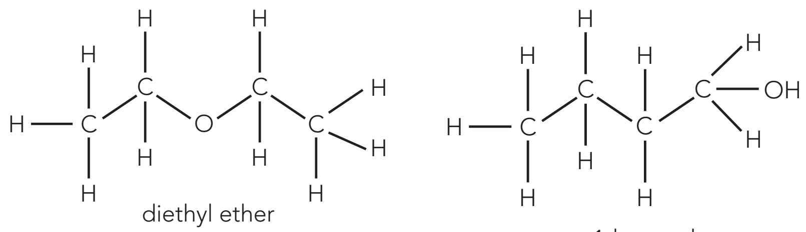
Table 7.5

SUBSTANCE TYPE	GENERAL TREND IN SOLUBILITY IN WATER
Metallic	Insoluble
Covalent Network	Insoluble
Ionic	Many are soluble (see page 178 for more detail)
Covalent Molecular	Polar molecules tend to be soluble. Non-polar molecules tend to be insoluble.

Vapour Pressure

For any liquid, a percentage of molecules that reach the surface of the liquid have sufficient kinetic energy to leave the liquid and form a vapour. The greater the number of particles escaping, the greater the vapour pressure.

The vapour pressure is dependent on the temperature of the liquid. The vapour pressure of one covalent molecular substance compared to another, at a set temperature is dependent on the intermolecular forces exhibited.



diethyl ether

1-butanol

Because of the -OH group attached to 1-butanol, it has much stronger intermolecular forces than diethyl ether. If measured at the same temperature, 1-butanol would have a lower vapour pressure than diethyl ether.

Question 7.7

Which of the following pairs of liquids would have the higher vapour pressure at 20 C? Give a brief reason for your choice.

(a) pentane ($CH_3CH_2CH_2CH_2CH_3$) or 1-pentanol ($CH_3CH_2CH_2CH_2CH_2OH$)

____

____

(b) water (H_2O) or ethanol (CH_3CH_2OH)

____

____

(c) pentane ($CH_3CH_2CH_2CH_2CH_3$) or heptane ($CH_3CH_2CH_2CH_2CH_2CH_2CH_3$)

____

____

Question 7.8

Predict what effect adding salt (NaCl) would have on the vapour pressure of water. Give a brief reason for your answer.

____

____

____

____

7.4 CHROMATOGRAPHY

Chromatography is a method of chemical analysis used to separate and then identify the components of mixtures. It is very useful when separating mixtures containing very small quantities of components to be identified.

There are 3 main parts to any analysis involving chromatography:

- **Mixture:** contains elements/compounds to be separated

- **Stationary Phase:** mixture is placed onto the stationary phase. Components in the mixture are attracted to stationary phase to varying degrees because of intermolecular forces involved.

- **Mobile Phase:** a specially selected solvent that will travel through the stationary phase. The strength of intermolecular forces between components in the mixture and the mobile phase determines how far and how rapidly the components in the mixture will move away from their original position.

The intermolecular forces between compounds to be identified and the compounds in the stationary phase and the mobile phase to a large extent determine the results of a chromatograph. Typically these intermolecular forces are dispersion forces, dipole-dipole forces and hydrogen bonding (i.e. van der Waals forces).

Paper Chromatography

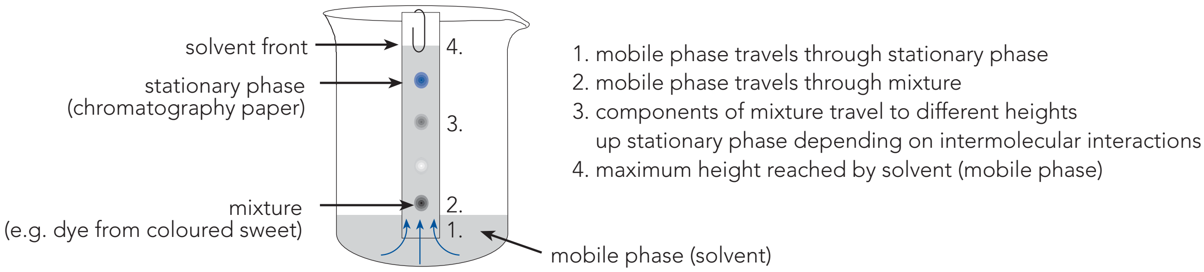


***Figure 7.7** Separating colours in food dye*

Question 7.9

State what would happen if the components in the mixture where not soluble in the mobile phase solvent.

____

____

Question 7.10

In an experiment it was found that component A travelled further up the stationary phase than component B. Explain what this indicates about the intermolecular interactions between these two components and the stationary phase.

____

____

____

Thin Layer Chromatography (TLC)

Thin layer chromatography is very similar to paper chromatography. The difference is that the stationary phase is coated onto an inert, rigid material such as glass or aluminium. Silica and alumina (Al_2O_3) are common materials used as the stationary phase.

Question 7.11

What advantages are there in using stationary phase coated onto an inert, rigid material rather than using paper?

____

____

____

High Performance Liquid Chromatography (HPLC)

In HPLC, the mixture to be analysed and the liquid mobile phase are pumped at high pressure through the stationary phase. Intermolecular forces between the compounds in the mixture and the stationary phase affect the speed at which these compounds travel through the stationary phase. This leads to the separation of the compounds in the mixture which in turn allows for further analysis of each compound and the determination of the relative percentage composition of the mixture.

Gas Chromatography

In gas chromatography the mobile phase is injected into an inert carrier gas, such as He, N_2 or Ar, before travelling through the stationary phase.

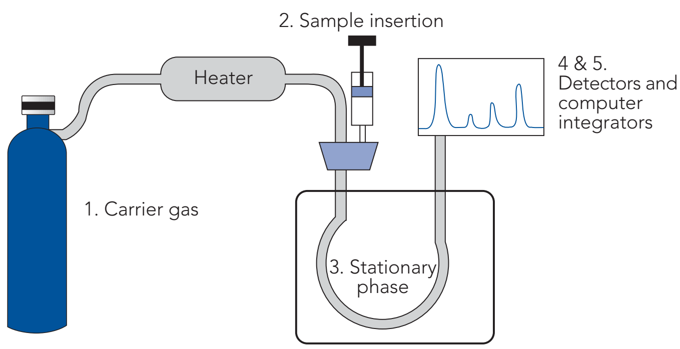


Figure 7.8 Schematic of a gas chromatograph

There are five key components of a gas chromatograph:

1. **Carrier gas:** very pure H_2, He, N_2 or Ar.

2. **Sample insertion:** very small samples of mixture to be analysed injected into carrier gas. Samples might be 0.2 mL in volume.

3. **Stationary phase:** a column or coil containing the stationary phase which is a very viscous liquid such as methyl silicone. Separates gas components by virtue of solubility or boiling point and affinity of components for the molecules in the stationary phase.

4 & 5. **Detectors and Computer Integrators:** to analyse data, position and height of peaks compared to reference data to identify compounds and give relative (percentage) composition.

Gas chromatography is used when the mixture to be analysed is a gas, or is easily converted to a gas. HPLC is preferred when the mixture is less volatile, the boiling point of the compounds are too high to be suitable for use in GC.

The data produced by the detectors in both GC and HPLC is very similar:

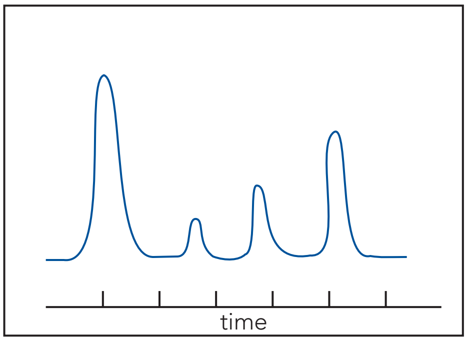


Figure 7.10 *Trace for GC or HPLC*

- Information about intermolecular forces given by position of peaks.
- Relative abundance of molecules present given by height of peak and/or area under curve.

Question 7.12

What type of chromatography would you recommend to be used to analyse the following mixtures:

(a) Coloured inks used to make a black marker pen.

____

(b) Testing the quantity of ethanol in E-10 petrol.

____

(c) Composition of a fruit juice.

____

REVIEW QUESTIONS

Chapter 7: Intermolecular Forces

1. Draw the electron dot diagram for each of the following:

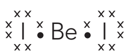
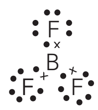
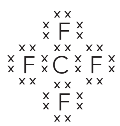
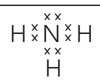
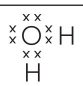
(a) H_2O (b) OH^-

(c) CO_2 (d) NO_3^-

(e) HCO_3^- (f) SO_4^{2-}

(g) CH_3COOH (h) K_2CO_3

2. Predict the shape of each molecule mentioned by completing the table below:

ELECTRON DOT DIAGRAM	CENTRAL ATOM	NUMBER OF ELECTRON PAIRS ON CENTRAL ATOM	NUMBER OF LONE PAIRS ON CENTRAL ATOM	DIAGRAM AND NAME OF SHAPE
				
				
				
				
				
				

3. Use the general trends in electronegativity provided by the periodic table to rank the following covalent bonds in order showing increasing degree of ionic character.

F–O ; F–Si ; F–F ; F–C

4. Which of the following molecules contain bonds that are polar?

CO_2 ; NO_2 ; Cl_2 ; HCl ; CCl_4 ; H_2O ; Br_2 ; N_2

5. Complete the following table to indicate which of the following molecules are polar.

MOLECULE	DOES IT CONTAIN POLAR BONDS?	MOLECULAR SHAPE	IS THE MOLECULE POLAR?
CH_4			
CCl_2F_2			
H_2S			
HI			
PH_3			
$AsBr_3$			

6. Carbon and nitrogen are neighbours on the periodic table and so have similar electron configurations. Both form covalent molecular compounds with other non-metal elements.

   Explain why pure carbon has a melting point of approximately 3550°C while pure nitrogen has a melting point of –210°C.

7. Dry ice, or solid carbon dioxide sublimes (goes from a solid to gas) at a temperature of –78.5°C, whereas another group 14 element's oxide (SiO_2 or silica) has a melting point of 1610°C.

   Use your knowledge of bonding in solids to explain this large difference in melting points.

8. Graphite and diamond are allotropes of carbon.

   Properties of diamond: *very hard, brittle, has a very high melting point and is a poor conductor of electricity.*

   Properties of graphite: *soft, excellent lubricant, very high melting point and is a good conductor of electricity.*

   (a) Briefly explain why both diamond and graphite have very high melting points.

   (b) Explain why these two forms of carbon have such differing properties of electrical conductivity and hardness.

   (c) Use its structure and bonding to explain why graphite is able to act as a lubricant.

9. Name the intermolecular forces that would exist between molecules of the following gases. If there is more than one type, only name the most significant.

   (a) H_2 (b) N_2 (c) CO_2 (d) CO

   (e) HCl (f) H_2S (g) SO_2 (h) Cl_2O

   (i) Ar (j) He

10. H_2O has a boiling point of 100°C, HCl has a boiling point of –84.9°C and F_2 has a boiling point of –188°C.

(a) Name the van der Waals forces that exist between molecules of the three substances.

(b) What special circumstances would need to exist for dipole-dipole interactions to be classified as hydrogen bonds?

(c) If the only intermolecular forces present between the three substances were dispersion forces, which would probably have the lower boiling point – give a reason for your answer.

11. In the following pairs, the first substance mentioned has the higher boiling point. In each case give possible reasons why.

(a) H_2Te and H_2S (b) H_2S and Ar

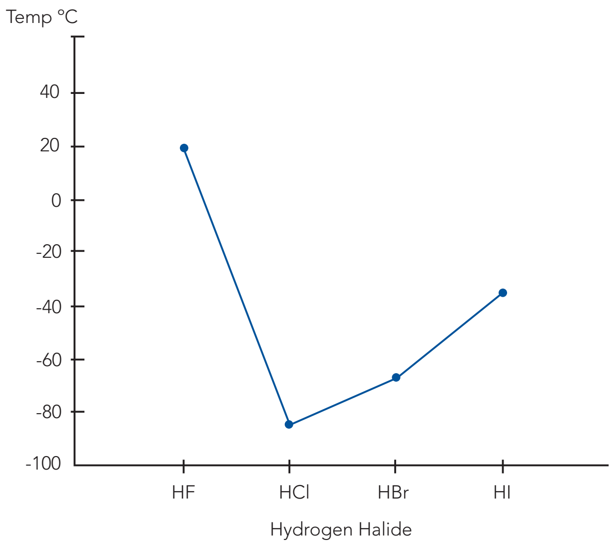
(c) H_2O and H_2S (d) H_2O and CH_4

12. Explain why HCl is more soluble in water than Cl_2.

13. The relative boiling points of the hydrogen halides is given in the graph below.

(a) Explain why the melting point of HF does not follow the trend of the other hydrogen halides.

(b) Why is the melting point of HI greater than the melting point of HBr?



14. The boiling points of water and iodine are 100°C and 184°C respectively. The most significant intermolecular forces in water are hydrogen bonds while in iodine only dispersion forces exist. Explain how it is possible for a substance exhibiting dispersion forces only, to have a higher boiling point than a substance having hydrogen bonds between molecules.

15. Classify each of the following solvents as polar or non-polar:

    Petrol, ammonia, ethanol, kerosene, turpentine, methylated spirits

16. Classify the following solutes as ionic, polar or non-polar:

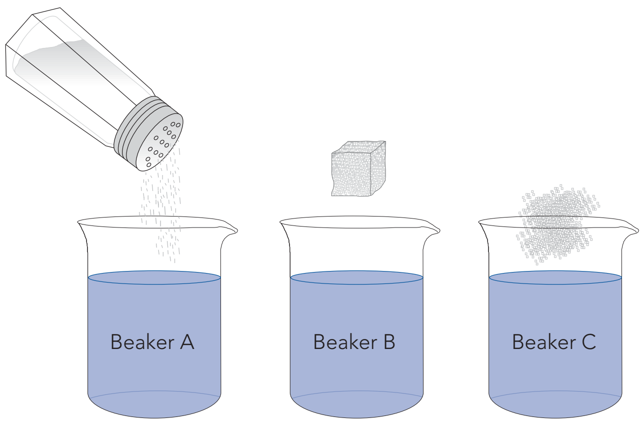
    Sugar, copper (II) sulfate, ammonium chloride, ammonia, sodium hydroxide, carbon, sulfur, hydrogen chloride, hydrogen sulfide, methane, silver chloride.

17. Three beakers of water, labelled A, B and C, were placed onto a bench. Sodium chloride was dissolved in beaker A, sugar was dissolved in beaker B, and iodine crystals were dropped into beaker C (very little of the iodine dissolved).

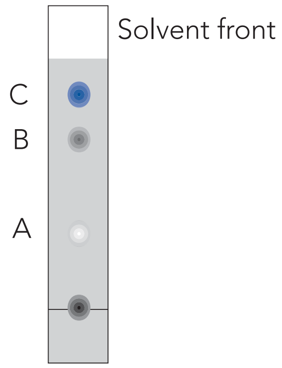
    (a) Indicate the forces of attraction that exist within the water in each beaker once the solid has been added to the water.

    (b) In which beaker are the forces of attraction between solid particles not significantly decreased by the attraction between the water molecules and the solid particles?

    (c) Draw a diagram to show the interactions between the particles in salt and water that cause the salt to dissolve in the water.



18. Silica gel which exhibits dipole-dipole forces of attraction was used as the stationary phase for the TLC chromatograph shown below. What information does this chromatograph tell you about the intermolecular forces exhibited by molecules A, B and C if the mobile phase was hexane?



FOR THE EXPERTS

19. Cholesterol is a waxy, non-polar substance found in many foods. Almost all cells in the body can make cholesterol but the liver is especially efficient at producing and distributing it.

Lipoproteins are complex molecules that have a lipid section that is non-polar and a protein section that is polar. Lipoproteins enable cholesterol to be transported around the body in the blood.

(a) Given that blood is 92% water, comment on the solubility of cholesterol in blood. Explain your answer.

(b) Lipoproteins are soluble in blood and can dissolve cholesterol. Explain how this is possible.

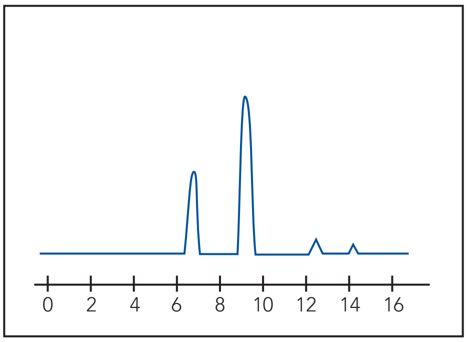
Lipoproteins are classified by their density, the greater the percentage of protein the higher the density, the greater the percentage of lipid the lower the density.

HDL	High Density Lipoproteins
IDL	Intermediate Density Lipoproteins
LDL	Low Density Lipoproteins
VLDL	Very Low Density Lipoproteins

LDLs are primarily responsible for transporting cholesterol around the body, too many LDLs in our diet can result in cholesterol being deposited on artery walls.

HDLs tend to remove excess cholesterol from the blood stream and reduce the amount of cholesterol deposited on artery walls. It is suggested by Health authorities that the ratio of LDL to HDL in the blood be between 3.2 and 3.6 and this represents an average coronary risk factor.

Analysis of lipoproteins in the blood of a person using HPLC produced the following results



(c) Given that the two larger peaks represent LDL and HDL and the stationary phase was highly polar, predict which of the two peaks was the LDL peak and which was the HDL peak. Explain your answer.

(d) Comment on the relative amounts of LDL and HDL in the person's blood.

8 KINETIC THEORY AND GASES

Topics covered in this chapter:

8.1 STATES OF MATTER

Matter can exist as either solid, liquid or gas under normal conditions. The properties of these three states of matter (sometimes called phases) can best be explained by the Kinetic Theory of Matter.

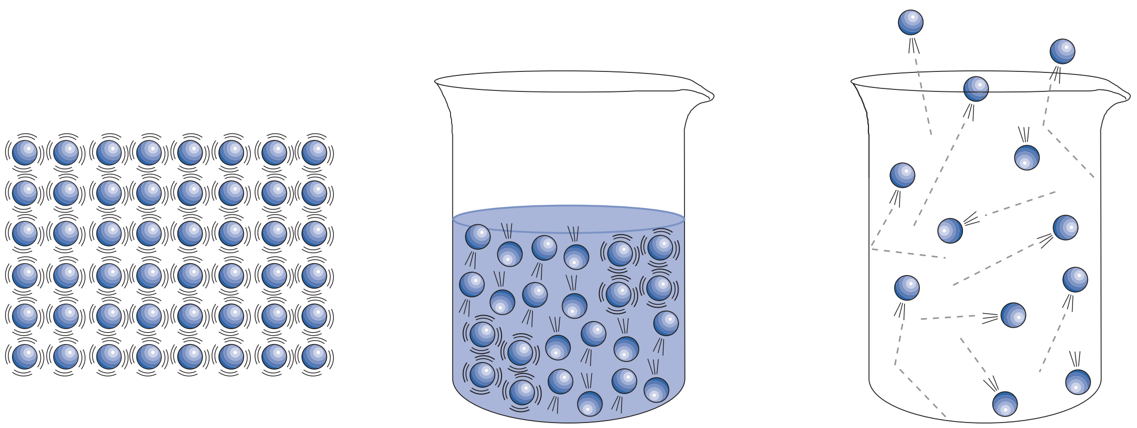
Some of the basic assumptions of the Kinetic Theory are:

- all matter is made up of extremely small particles (e.g. atoms, molecules, ions)
- these particles are in constant random motion
- the collisions between these particles are perfectly elastic
- mutual attractive forces exist between these particles but they are insignificant in the gaseous state.

In solids, the particles are held close together in a regular pattern by strong forces of attraction. The particles vibrate about fixed positions.

In liquids, the vibrating particles have sufficient energy to move from their fixed positions to other parts within the liquid. Liquids take the shape of their container.

In gases, the particles have sufficient kinetic energy to escape the attractive forces from other particles. Gases take up the complete volume of a container.



(a) **Solids:** particles vibrate about fixed positions. Solids have a definite shape.

(b) **Liquids:** particles are free to move within the liquid. Some particles escape the surface to form a vapour.

(c) **Gases:** particles move freely in all directions. Gases are compressible since there is a lot of space between particles.

***Figure 8.1** Kinetic theory - solids, liquids and gases*

DID YOU KNOW?

The average distance between molecules in the air, mostly N_2 and O_2, is approximately 3.4×10^{-9} m. This may seem very close but typically you could fit another ten molecules in between them in all directions. When a gas is liquefied the volume is reduced by a factor of one thousand or so.

Question 8.1

Use the descriptions of solids, liquids and gases illustrated in Figure 8.1 to help you complete the table.

	SOLID	LIQUID	GAS
Shape	fixed	that of container to level surface	takes up complete volume of container
Volume			
Density			
Ease of compression			
Ease of diffusion			
Ease of flow			

Question 8.2

Use the kinetic theory to explain:

(a) Why gases can be easily compressed.

____

(b) Why solids have a relatively high density.

____

(c) Why liquids and gases flow easily but solids do not.

____

Question 8.3

In which state/s of matter:

(a) do particles have the least restricted movement? _______________

(b) are particles most influenced by attractive forces? _______________

(c) do particles move mostly in straight lines? _______________

8.2 IDEAL GASES AND THE KINETIC THEORY

Gases tend to behave very similarly in terms of temperature, pressure and volume, especially if stored at temperature and pressures well away from those needed to cause a phase change. Due to these similarities in behaviour, gases are often treated as ideal gases where one set of rules describes the behaviour of many gases.

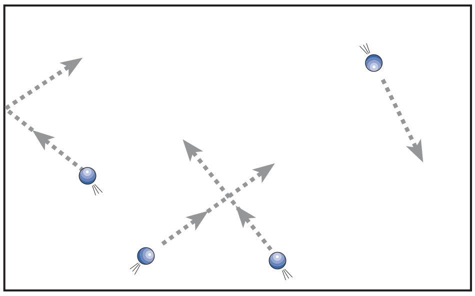
To be considered as an ideal gas the following assumptions are made about a gas:

- The molecules in the gas are in a state of constant, random, straight line motion (Brownian Motion).
- The size of an individual gas molecule is insignificant compared to the distance of separation between molecules in the gas.
- All collisions involving gas molecules are perfectly elastic meaning that there is no overall loss of kinetic energy during collisions.
- Gas molecules do not experience attractive or repulsive forces (intermolecular forces). The total energy of a sample of gas is the total kinetic energy of molecules in the gas.
- The absolute temperature of a gas is directly proportional to its average kinetic energy.
- All ideal gases have the same average kinetic energy at the same temperature.

Gases are considered as non-ideal when intermolecular forces affect their behaviour, which happens when the gas is at temperatures and pressures close to those required for a phase change.

molecules experience
no intermolecular forces

molecules move in straight lines
unless they collide with each other
or container walls



gases are mainly empty space –
size of molecules is insignificant
compared to distance of separation

collisions are perfectly elastic
with no loss of kinetic energy

Question 8.4

Explain why it is very easy to compress the gas in a bicycle pump, even when the air outlet is blocked.

Question 8.5

Would your answer to the previous question change if the gas in the pump was changed from air (76% N_2, 21% O_2) to helium? Explain.

Question 8.6

Explain why an inflated party balloon will often deflate even though it is properly tied.

Question 8.7

An over ripe banana can often be smelt throughout the room it is kept in. Use the kinetic theory to explain how this odour can spread around a room.

8.3 GAS TEMPERATURE

Temperature is a measure of the average kinetic energy (KE) of all of the molecules in a gas.

$$KE = \frac{1}{2} mv^2$$

KE = kinetic energy
m = mass of molecules
v = speed of molecule

When a gas is cooled, on average its molecules will be moving more slowly, the average kinetic energy of molecules is lower. Theoretically, if a gas is cooled sufficiently, the molecules will cease to move. This is the lowest temperature that can be reached and is called **absolute zero**. The Kelvin, or absolute temperature scale has this temperature as its zero point.

O K = absolute zero = -273.15°C

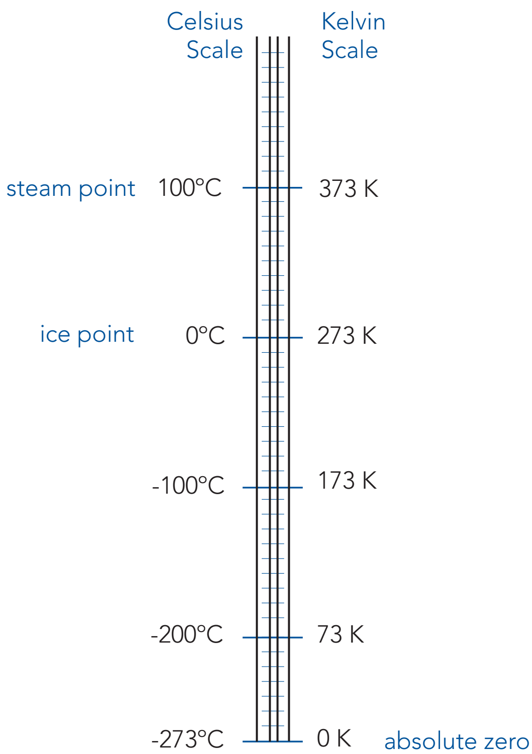
8.4 TEMPERATURE SCALES

To establish a temperature scale on a thermometer it is important to choose two easily reproducible temperature conditions called fixed points. In the Celsius scale these points are:

- lower fixed point
  - melting point of pure ice
  - 0°C (zero degrees Celsius)

- upper fixed point
  - boiling point of pure water at standard atmospheric pressure (100 kPa)
  - 100°C (100 degrees Celsius)

To calibrate an unmarked thermometer, it is necessary to place the thermometer firstly in melting ice and mark the 0°C point, and then in boiling water and mark the 100°C point. The rest of the scale can then be marked by dividing the space between the points into 100 equal divisions. Each division is a degree.



In the Kelvin or absolute scale, the zero point is set at absolute zero (–273°C) with kelvin (K) divisions that are the same size as those on the Celsius scale. Absolute zero is the temperature at which a substance has the lowest possible internal energy. This occurs at –273.15°C.

Note: Kelvin temperatures are not expressed as degrees, but simply kelvin.

K = °C + 273

Question 8.8

Complete the following table of temperature data by converting to either °C or K as appropriate. The first one is done for you.

DESCRIPTION	TEMP (°C)	TEMP (K)
Boiling point of helium	-269	4
Boiling point of oxygen	-183	
Melting point of mercury		234
Melting point of ice	0	
Normal laboratory temperature	25	
Normal body temperature		310
Boiling point of water		373
Melting point of aluminium	660	
Melting point of NaCl		1024
Melting point of tungsten	3422	

8.5 HEAT AND TEMPERATURE

Heat and temperature are often confused as being the same thing. However they have distinctly different meanings.

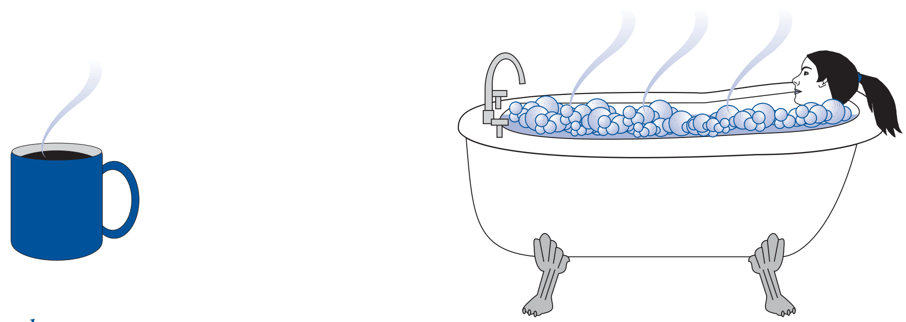


Figure 8.2 Heat and temperature

- Heat refers to the energy which is transferred from one body to another due to a difference in temperature. It is a measure of the increase in total internal energy (KE + PE). The units are Joules.

- Temperature is used to describe how hot something is. The greater the average kinetic energy of particles in a body the greater the temperature. Temperature scales (e.g. Celsius) are used to indicate the degree of hotness.

When matter is heated it can store energy as kinetic energy (temperature increases) or as potential energy (phase change only). An increase in kinetic energy ($KE = \frac{1}{2} mv^2$) means that the constituent particles are moving faster or vibrating more rapidly. An increase in potential energy will cause the particles to become more separated from each other. This occurs because the particles have sufficient energy to overcome the attractive forces that normally hold them together. A phase change occurs.

Question 8.9

The kinetic energy of a particle is given by KE = ½ mv_2 (m = mass, v = speed).

(a) A He atom has twice the mass of a H_2 molecule.

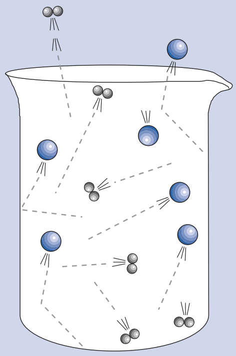
Which would have the greatest KE if:

i) they were both moving at 500 ms^{-1}?

____

ii) the He atom slowed down to 250 ms^{-1}?

____



(b) In a mixture of gases, such as air, the average KE is the same for all particles at any given temperature. Does this mean that oxygen and nitrogen molecules travel at the same speed in the air? Explain.

____

____

____

8.6 KINETIC ENERGY DISTRIBUTION

The particles in a substance are in constant motion and are involved in frequent collisions. This means that there is a range of kinetic energies at any given temperature.

The distribution of molecular kinetic energies of a sample of gas is shown below. Although the kinetic energy of an individual molecule can change due to collisions, the average KE and range of KE for the sample does not change at a fixed temperature. If the temperature is increased (T_2 is greater than T_1 in Figure 8.3) there is a greater range of KE and higher average kinetic energy.

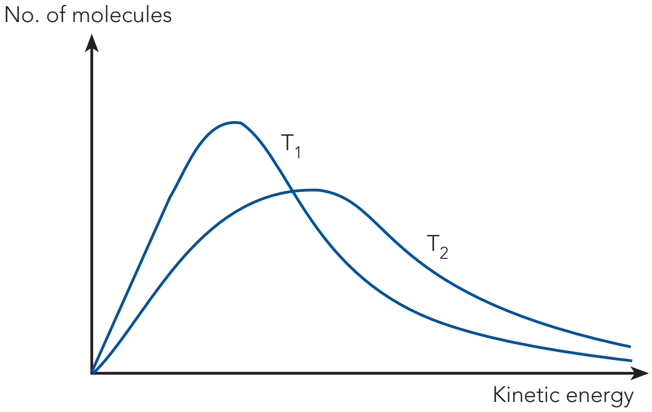
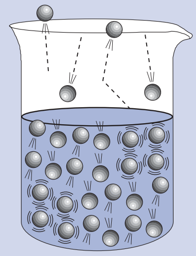


Figure 8.3 Distribution of kinetic energy of particles in a substance at different temperatures.

Question 8.10

(a) What does the area under the kinetic energy distribution graph represent? (Figure 8.3)

(b) At which temperature (T_1 or T_2) is the relative number of particles with high KE the lowest?



Question 8.11

Evaporation of liquids occurs to some extent at all temperatures.

Use the idea of kinetic energy distribution to explain:

(a) Why evaporation can occur?

(b) What happens to the average kinetic energy of the remaining molecules?

(c) What effect this has on the temperature of the remaining liquid?

8.7 VAPOUR PRESSURE AND BOILING

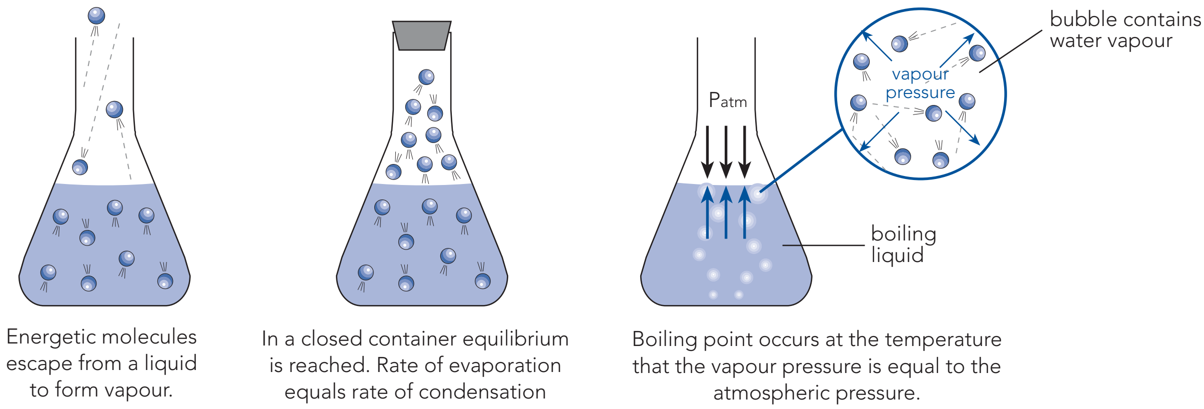


Figure 8.4 *Vapour pressure and boiling point.*

When a liquid is placed in a container some of the more energetic molecules of the substance evaporate into the air to form vapour. The particles of vapour continually move about, some falling back into the liquid while most escape if in an open container.

In a closed container an equilibrium is reached when the rate of evaporation is equal to the rate of condensation. At this point the pressure exerted by the vapour is called the vapour pressure. Vapour pressure increases with an increase in temperature. This is the result of the particles having greater kinetic energy and escaping more freely from the liquid. Boiling point is reached when the vapour pressure of a liquid is exactly that of the atmospheric pressure.

8.8 FACTORS AFFECTING VAPOUR PRESSURE

Vapour pressure is due to the molecules that have escaped from the liquid surface a liquid colliding with the walls of the container. The pressure exerted by these molecules will vary with temperature as well as being different for different substances.

SUBSTANCE	VAPOUR PRESSURE (KPA) AT VARIOUS TEMPERATURES			
	0°C	25°C	50°C	100°C
Water	0.61	3.2	12.3	101.3
Ethanol	1.63	7.86	29.6	225.7
Mercury	0.00002	0.00025	0.0017	0.036

Table 8.1 Variations of vapour pressure with temperatures for some common substances.

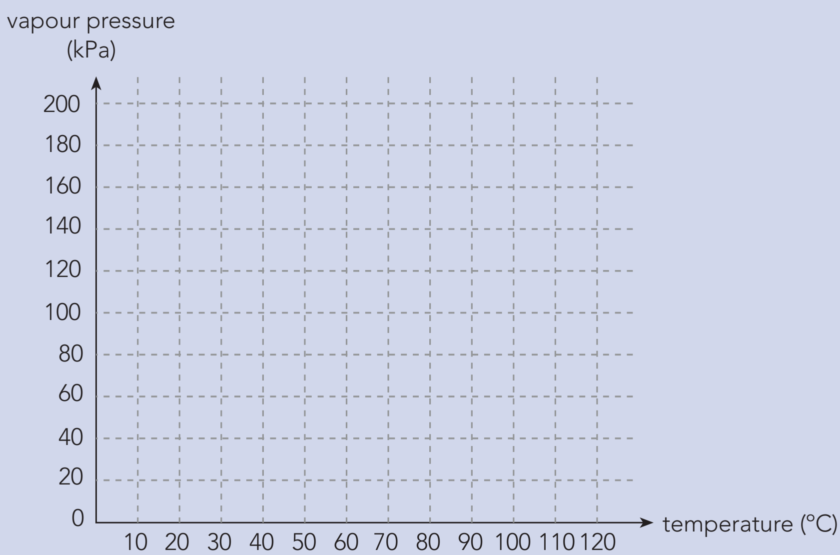
Vapour pressure increases with temperature since there are a greater number of particles with sufficient energy to escape from the liquid phase (see Figure 8.4). These molecules are also moving more rapidly and hence colliding more often with greater force.

The variation in vapour pressure for different substances is due to differences in intermolecular forces. Liquids, like alcohol for example, have a relatively high vapour pressure since the forces of attraction between neighbouring alcohol molecules (intermolecular forces) are rather weak. Such liquids are called volatile liquids and also have a low boiling point.

Question 8.12

Graph the following data to show the variation in vapour pressure with temperature for ethanol and water.

TEMPERATURE °C		0	10	20	30	40	50	60	70	80	90	100	120
VAPOUR PRESSURE (kPa)	water	0.6	1.2	2.3	4.2	7.4	12	20	31	47	70	101	200
	ethanol	1.6	3.2	5.9	11	18	30	47	72	110	160		



Question 8.13

Use your graph to predict:

(a) the boiling point of ethanol at sea level (100 kPa). ______________________

(b) the boiling point of water on a high mountain where the atmospheric pressure is 60 kPa.

____

Question 8.14

Liquids such as ethanol have a greater vapour pressure than water for any given temperature (see your graph from question 8.12). Comment on the strength of the intermolecular forces in ethanol compared to water.

____

____

Question 8.15

The normal boiling point of water is 100°C. However, a pressure cooker is designed to allow water to boil at 140°C. How is this possible?

____

____

8.9 GAS PRESSURE

Gas pressure is caused by molecules striking the container walls, it is a measure of the force of and the number of collisions the molecules have with the container walls.

The SI unit for pressure is the **pascal (Pa)** or newton per square metre (Nm^{-2})

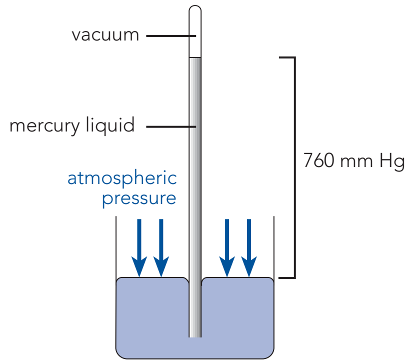
1 pascal pressure is created when the gas molecules strike the container walls with a total force of 1 newton for every square metre of container wall.

In chemistry it is common to quote gas pressure in kilo pascals (kPa) whereas in meteorology the hecto pascal (hPa) is used.

1.00 kPa = 10 hPa = 1000 Pa

Two non-SI units for pressure that are commonly used are:

- atmosphere (atm).
- millimetres of mercury (mmHg).



1.00 atm = 760 mmHg = 101.3 kPa

8.10 MOLAR VOLUME OF GASES

The molar volume of gas is simply the volume occupied by 1 mole of gas. At Standard Temperature and Pressure or STP (0.0°C and 100 kPa) the molar volume of an ideal gas is taken to be 22.71 L.

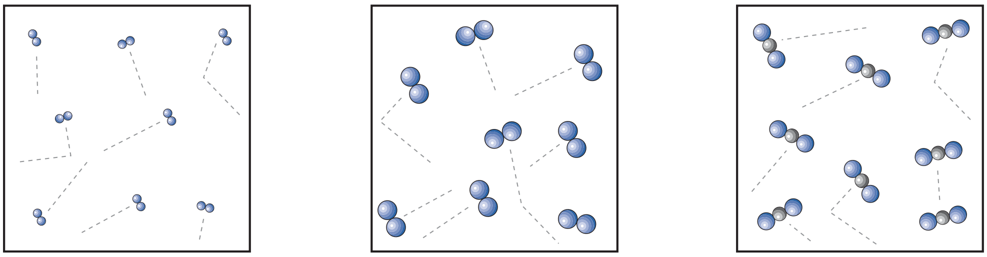


Figure 8.5 *The molar volume of gases at STP. For an ideal gas this is 22.71 L. In reality the molar volume of some gases is slightly different due to the effect of intermolecular forces. Note that the number of molecules is the same in each one.*

$$\text{Number of moles of a gas at STP} = \frac{\text{Volume of gas at STP}}{22.71}$$

$$\mathbf{n\ (gas\ at\ STP) = \frac{V\ (gas\ at\ STP)}{22.71}}$$

Worked Examples

8.1 MOLES FROM VOLUME

Calculate the number of moles of hydrogen gas in the 294 litres at STP.

$$n(H_2) = \frac{V\ (H_2\text{ at STP})}{22.71} = \frac{294}{22.71}$$

$$n(H_2) = 12.9$$

8.2 VOLUME OF MASS

Determine the volume of carbon dioxide produced at STP when 125 grams of dry ice sublimes, $CO_2(s) \rightarrow CO_2(g)$.

$$N(CO_2) = \frac{m}{M} = \frac{125}{(12.01 + 32.00)} = 2.84$$

$$V(CO_2)\text{ at STP} = n \times 22.71 = 2.84 \times 22.71$$

$$V(CO_2)\text{ at STP} = 64.5\text{ L}$$

8.3 VOLUME FROM MASS

What mass of propane gas (C_3H_8) would be required to fill a gas bottle 4640 L of propane at STP is compressed to fill the gas bottle.

$$n(C_3H_8) = \frac{V}{22.71} = \frac{4640}{22.71} = 204$$

$$m(C_3H_8) = n.M = 204 \times (36.03 + 8.064)$$

$$m(C_3H_8) = 9.00 \times 10^3 \text{ g}$$

Question 8.16

Calculate the number of moles of chlorine gas produce when a swimming pool chlorinator generates 32.5 litres of the gas at STP.

Question 8.17

Calculate the volume of nitrogen gas produced when 695 g of liquid nitrogen is heated until it reaches STP.

Question 8.18

The Hindenberg class airship was 244 m long and required approximately 18 tonnes of hydrogen gas to fill its 16 gas cells. What volume of hydrogen gas does this airship hold when at STP?

8.11 THE GAS LAWS

In a gas, particles are relatively far apart and constantly on the move. They occupy the total volume of any closed container and exert a pressure which is due to their collisions with the container walls.

If the temperature is increased, the average kinetic energy of the gas particles also increases, and hence their velocity.

For any given mass of gas, the pressure, volume and temperature are all related. A change in any one of these conditions will cause a change in one or both of the others. To study gas behaviour one condition is always kept constant.

Pressure-temperature relationship

When a gas is placed in a container of *fixed volume* and heated, its pressure rises as shown graphically below. This increase in pressure is due to the higher velocity of the gas particles which now collide with the container walls more forcefully and more often.

Experimental results show that *the pressure of a fixed mass of gas is directly proportional to its absolute temperature provided that its volume is kept constant.*

An important inference from this relationship is the concept of an absolute zero temperature. As a gas is cooled its pressure is reduced, and would eventually be zero at –273.15 °C. In reality, most gases become liquids well before this temperature is reached.

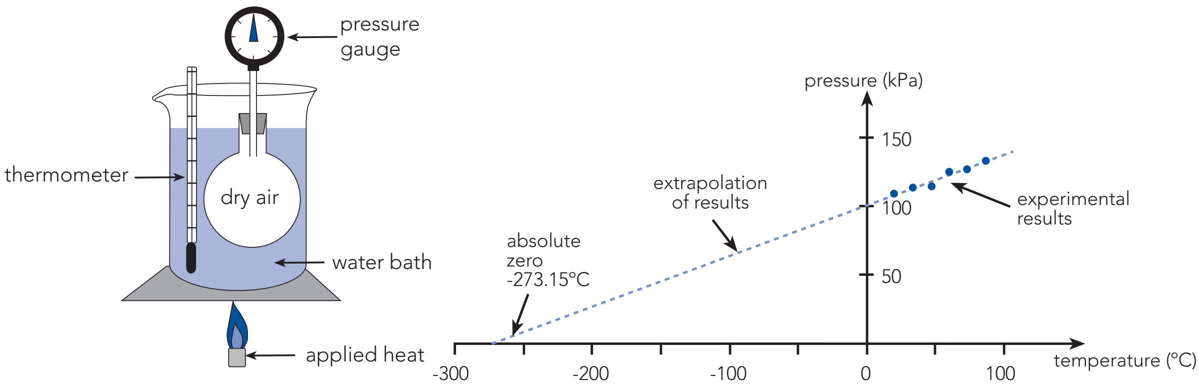
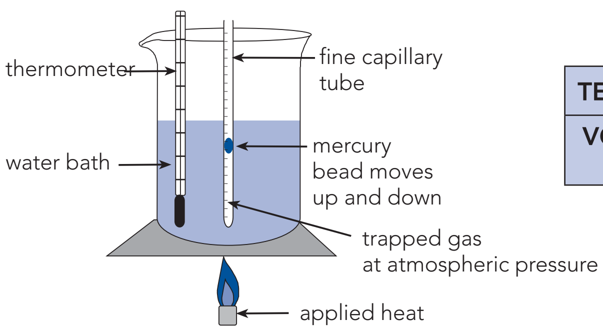


***Figure 8.6** The pressure of a gas is directly proportional to its absolute temperature. Extrapolation of graphed results indicates that a zero pressure would exist at –273.15 °C.*

Volume-temperature relationship

If a gas is heated at constant pressure its volume increases. As before, the increased temperature causes the gas particles to move more rapidly. Collisions become more frequent and more forceful. If the pressure is to remain constant the volume taken up will increase. Experimental results show that *the volume of a fixed mass of gas is directly proportional to its absolute temperature provided its pressure is kept constant*. This is often referred to as Charles' Law.



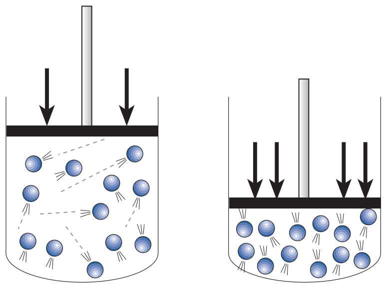
TEMP (°C)	0	20	40	60	80	100
VOLUME (mL)	3.50	3.76	4.00	4.27	4.52	4.78

***Figure 8.7** Charles' law investigation. Some typical results are shown.*

Pressure-volume relationship

When a gas is forced into a smaller volume its pressure increases. This increase in pressure is due to the greater number of collisions between the gas particles and the container walls.

Experimental results show that *the pressure of a fixed mass of gas is inversely proportional to its volume provided its temperature is kept constant*. This is often referred to as Boyle's Law.



PRESSURE (kPa)	100	200	300	400	500	600
VOLUME (L)	4.00	2.00	1.33	1.00	0.80	0.67

Figure 8.8 *Kinetic theory explanation of Boyle's Law. In confined spaces, gas particles collide with container walls more often, hence increasing the pressure. Some typical results are shown.*

Question 8.19

Use the kinetic theory of gases to explain why the volume of a gas increases when the temperature is increased. Assume pressure is kept constant.

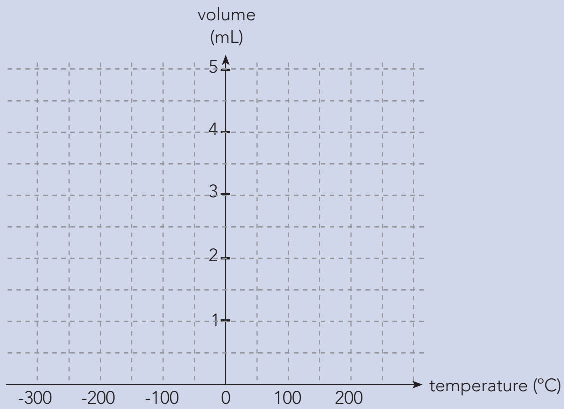
____

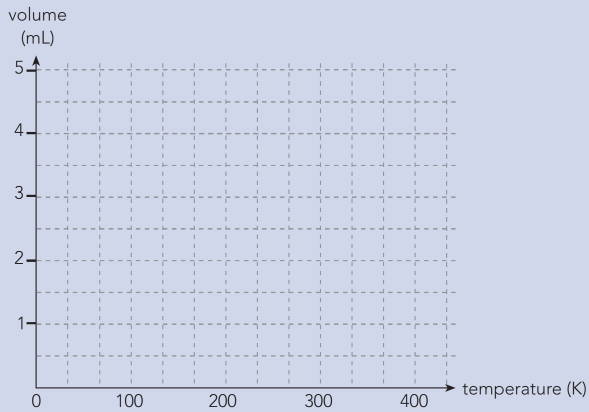
____

Question 8.20

Graph the data given in Figure 8.7 as:

i) V versus T (°C)   ii) V versus T (K)





Question 8.21

(a) Extrapolate the first graph to a volume of zero. What temperature is indicated?

____

(b) What is the significance of this temperature?

____

REVIEW QUESTIONS

Chapter 8 Kinetic Theory and Gases

1. Substances can exist as solids, liquids and gases. In which of these phases would you expect the particles of the substance:

   (a) to be closest together?
   (b) to have the least influence on each other?
   (c) to be able to flow?
   (d) to vibrate about a fixed position?
   (e) to occupy minimal space compared to distance between particles?

2. In terms of the kinetic theory, explain why:

   (a) solids do not flow
   (b) gases take up all of the volume of their container
   (c) when heated substances expand.

3. (a) Under what conditions do gases tend not to behave as ideal gases?
   (b) Why would these conditions lead to a change in the behaviour of the gas particles?

4. A sample of H_2 gas was stored under the same temperature and pressure conditions as a sample of N_2 gas. Does this mean that H_2 particles must have the same average velocity as the N_2 particles? Justify your answer

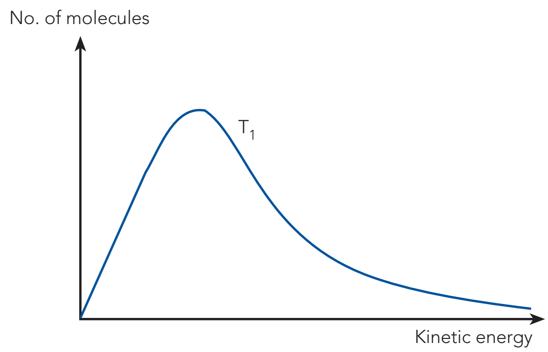
5. Convert the following temperatures as indicated.

   (a) °C to K
   -150°C, -20°C, 0.0°C, 100°C
   (b) K to °C
   100 K, 200 K, 273 K, 500 K

6. A sample of gas experiences a temperature change that doubles the average kinetic energy of the particles in the gas. Which temperature change would account for this? Explain your answer.

   A: 135°C → 270°C B: -123°C → 27°C

7. The graph below shows the distribution of molecular kinetic energies of a sample of air molecules (mostly nitrogen and oxygen molecules).



   (a) What does the peak of the graph indicate?
   (b) How would the average kinetic energy of N_2 and O_2 molecules compare?

(c) Which molecules, N_2 or O_2, on average, are likely to be moving the fastest? Why?
(d) If the air was heated to a new temperature (T_2) how would the shape of the graph change? Copy the graph and then also draw the approximate shape for a higher temperature (T_2).

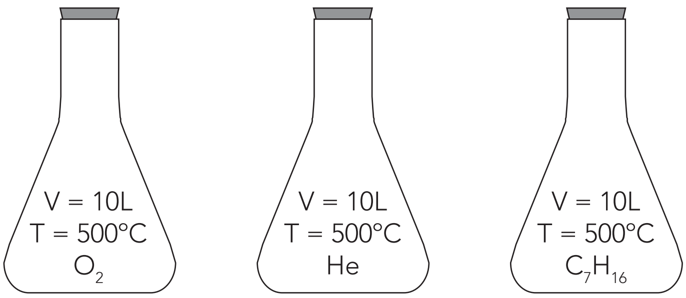
8. Under standard conditions the boiling point of methanol is 65°C and that for ethanol is 78°C. From this information which of these alcohols

(a) can be said to be more volatile?
(b) would have the highest vapour pressure at 20°C?
(c) has the strongest intermolecular forces?
(d) is likely to be more difficult to freeze?

9. Use the ideas of Kinetic Theory as it applies to gases and explain why:

(a) gases are easily compressed
(b) when we blow up a football it becomes hard and difficult to squash
(c) the pressure in car tyres increases on hot days.

10. Three identical sealed containers hold equimolar amounts of three different gases, O_2, He and C_7H_{16}, at the same temperature.



(a) In which container are the particles hitting the container walls most often?
(b) In which container are the particles hitting the container walls with the greatest average velocity?
(c) In which container would the gas pressure be the greatest?

11. In an experiment the boiling point of water was measured under different conditions as follows:

(i) pure water boiling in an open pot in the kitchen
(ii) salty water boiling in an open pot in the kitchen
(ii) pure water boiling in an open pot at the top of a high mountain.

Which situation would give:

(a) the highest boiling point? Why?
(b) the lowest boiling point? Why?

12. Matthew decides to see the effect of placing a small inflated party balloon in a refrigerator for a few minutes.

(a) What change is likely to occur to the balloon?
(b) Use Kinetic Theory to explain why this would occur.

13. The molar volume of real gases varies slightly from what might be expected of an ideal gas. Some examples are listed.

GAS	FORMULA	MOLAR VOLUME ($L\ mol^{-1}$) AT STP
hydrogen	H_2	22.72
methane	CH_4	22.67
oxygen	O_2	22.68
carbon dioxide	CO_2	22.55
*ideal gas	–	(22.71)

(a) What is meant by an ideal gas?
(b) Suggest a reason as to why real gases behave slightly differently to an ideal gas.
(c) Does the data above show any apparent trend in the variation of molar volume? If so, what is the trend?

14. Determine the number of moles of each of the following gases represented by the volume given at STP.

(a) 450 L of F_2
(b) 1.93×10^4 L of He
(c) 32 ml of CH_4
(d) 150 L of O_2

15. Calculate the volume that the following gases would occupy at STP.

(a) 2.32 moles of hydrogen
(b) 1.84×10^5 moles of ethane (C_2H_6)
(c) 72.5 g of chlorine

16. Rather than carry a pump, some cyclists carry small cannisters of highly compressed CO_2 to pump up flat tyres they may experience on a long ride.

The tyre of a typical road bike is 678 mm in diameter and contains a tube that is 23 mm in diameter. The inflated volume of such a tyre is 1.00 Litres.

On a 20.0ºC day a cyclist might typically pump the tyre up to a pressure that requires 7.5 times more gas than at standard conditions. Would a 16.0 g CO_2 canister contain enough gas to inflate this tyre?

FOR THE EXPERTS

An understanding of the relationships between pressure, volume and temperature of gases is essential for anyone involved in scuba diving. From calculating the time that a person can remain under water to examining the interaction between gases under pressure and our body's physiology, an understanding of the behaviour of gases is essential.

To calculate the amount of time a diver can remain under water for using a particular gas cylinder can be calculated using the following formula:

$$\text{Diving Time} = \frac{(\text{Initial Cylinder Pressure - Reserve Pressure}) \times \text{Cylinder Volume}}{(\text{Divers Breathing Rate} \times \text{Average Water Pressure During Dive})}$$

- The reserve pressure is the quantity of gas that needs to be left in the gas cylinder to allow for any emergencies that may occur during the dive.
- The average water pressure during the dive is measured in atmospheres. Pressure increases by 1 atm for every 10m of water depth, i.e. at 0 metres pressure is 1 atm; at 10 metres depth, the pressure is 2 atm; and at 20 metres the pressure is 3 atm.
- The breathing rate is determined in litres per minute.
- The diving time is measured in minutes.

17. A scuba diver wanted to be able to dive for 45 minutes to be able to explore a shipwreck that was at an average depth of 15 m. The diver was to use an 18.0 litre gas cylinder that was filled to a pressure of 210 atmospheres with a standard air mixture of 79% N_2 and 21% O_2. The diver had established that their breathing rate for the dive would be 22 litres of gas per minute and that they would require reserve pressure of 50 atm.

(a) Use the Diving Time formula to determine if the scuba diver will have sufficient gas in their tank to be able to dive for 45 minutes.

(b) Assuming that the gases in the cylinder behaved as ideal gases and that it contained 21% O_2, predict the mass of oxygen gas that a the cylinder would contain when filled to a pressure of 210 atmosphere at a temperature of 273 K.

AQUEOUS SOLUTIONS

9

Topics covered in this chapter:

9.1 SOLUTIONS

A solution is a homogeneous mixture formed when one substance, the solute, dissolves in another substance, the solvent. Solvents are usually liquids while solutes may be solids, liquids or gases. Hence, typically, solutions may be solid in liquid, liquid in liquid, gas in liquid or gas in gas.

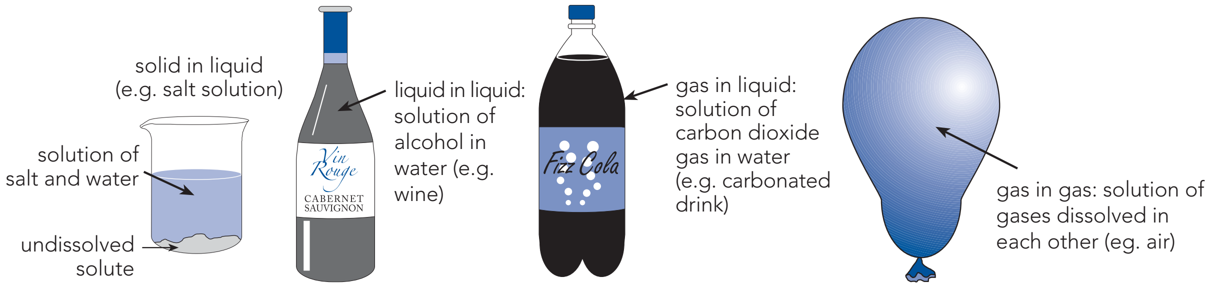


Figure 9.1 Types of solutions

The most important solvent on Earth is water and an understanding of aqueous solutions is essential to any study of living things, life processes, environmental change and many aspects of mining and industry. Solutions may be unsaturated, saturated or supersaturated.

Unsaturated Solution: solvent can dissolve more solute. An unsaturated solution can be recognised by the fact that no excess solute is evident.

Saturated Solution: at its current temperature, the solvent cannot dissolve any more solute.

Supersaturated Solution: if a saturated solution is heated it can dissolve more solute. In some circumstances, when the solution is cooled, the extra solute remains dissolved and the solution contains more dissolved solute than expected.

Table 9.1

TYPE OF SOLUTION	BEHAVIOUR IF EXTRA SOLUTE IS ADDED
Unsaturated	Any extra solute dissolves
Saturated	Assuming it is more dense than water, the extra solute added will sink to the bottom of the solution. There is no overall change in the amount of solute dissolved
Supersaturated	Solute added sinks to the bottom without dissolving. Some extra solute may also come out of the solution as the solution is supersaturated.

Question 9.1

Classify the four solutions shown in Figure 9.1 as unsaturated, saturated or supersaturated.

Question 9.2

Complete the following table.

SOLUTION	SOLUTE	SOLVENT	SOLUTION TYPE
sea water	*salt*	*water*	*solid in liquid*
carbonated drink			
air			
ammonia solution			
brass alloy			
deep sea diver's gas			
vinegar			

Question 9.3

In Figure 9.1 a salt water solution is illustrated.

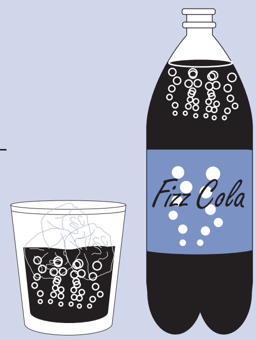
(a) Is it a saturated solution? ____________

(b) How do you know this? ____

(c) How could more salt be made to dissolve? _______________________________

Question 9.4

When a cool drink bottle is opened a lot of bubbles appear.



(a) What are the bubbles made of? ___________________

(b) Why do the bubbles appear only when the bottle is opened?

____

(c) Why does the drink go flat more quickly if it is warm?

____

____

(d) How do you think the following affect the solubility of gases?

(i) temperature ____

(ii) pressure ____

9.2 FORMING AQUEOUS SOLUTIONS

For a solute to be dissolved by a solvent the solute-solvent attractions must be sufficiently strong to counteract the solute-solute attractions and allow the solute particles to separate and be dispersed throughout the solvent. (See page 136 of Chapter 7 for more detail.)

Because of its V-shape and small size, water is a very good solvent for many ionic substances and polar covalent molecular substances.

When a substance dissolves in water an **aqueous solution** is formed.

e.g. 1: $NaCl_{(s)} \rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$

NB: as an ionic solid, NaCl dissociates when forming aqueous solutions.

e.g. 2: $CH_3CH_2OH_{(l)} \rightarrow CH_3CH_2OH_{(aq)}$

NB: covalent molecular substances stay as complete molecules when forming aqueous solutions.

e.g. 3: $HCl_{(g)} \rightarrow H^{+}_{(aq)} + Cl^{-}_{(aq)}$

NB: as a strong acid, HCl completely ionises when forming aqueous solutions.

Question 9.5

Write the equation to represent what occurs when the following substances dissolve in water to form an aqueous solution.

(a) Potassium nitrate

____

(b) Calcium chloride

____

(c) Aluminium sulfate

____

(d) Sucrose ($C_{12}H_{22}O_{11}$)

____

(e) The strong acid, HNO_3

____

9.3 ELECTROLYTES

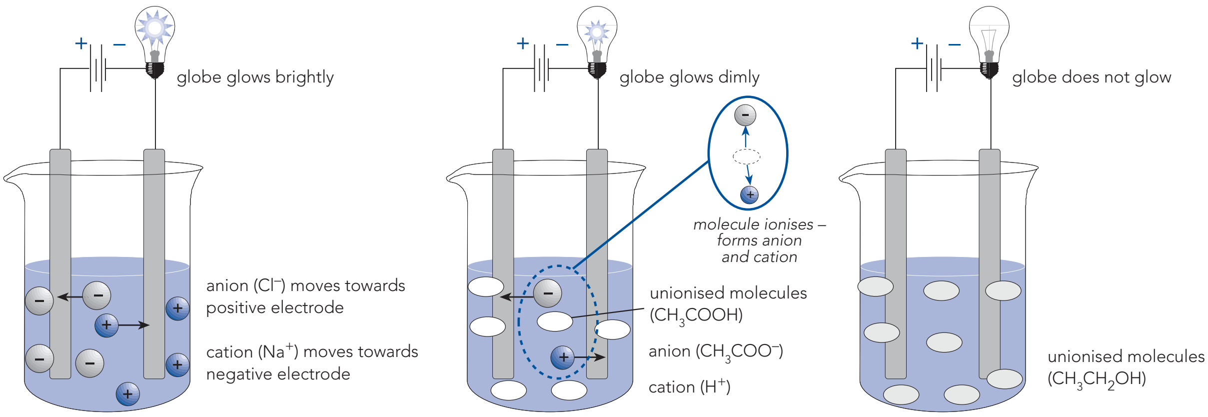
Any aqueous solution that conducts electricity must contain ions. These conducting solutions are called **electrolytes.**

Dissociation: ionic solids are composed of oppositely charged ions held together by strong forces of electrostatic attraction. When dissolved in water the ions separate or dissociate.

e.g. $Mg(NO_3)_{2(s)} \rightarrow Mg^{2+}{}_{(aq)} + 2NO_3^-{}_{(aq)}$

Ionisation: is the process in which ions are formed. An example of ionisation is when an acid dissolves in water.

e.g. $HCl_{(g)} \rightarrow H^+{}_{(aq)} + Cl^-{}_{(aq)}$



i) $NaCl_{(aq)}$ is a strong electrolyte as it is completely dissociated

ii) $CH_3COOH_{(aq)}$ is a weak electrolyte as it only partly ionises

iii) $CH_3CH_2OH_{(l)}$ is a non-electrolyte as it does not ionise

Figure 9.2 Electrolytes and non-electrolytes

Strong electrolytes: good conductors of electricity, for example aqueous solution containing ionic substances or strong acids.

Weak electrolytes: poor conductors of electricity, for example aqueous solutions of weak acids and weak bases (that are not ionic).

By convention, equations that show the formation of a strong electrolyte use a single arrow to indicate the process goes to completion and that the original substance has completely dissociated (or ionised).

e.g. $NaCl_{(s)} \rightarrow Na^+{}_{(aq)} + Cl^-{}_{(aq)}$

Equations showing the formation of a weak electrolyte use a double arrow to indicate that at any one time only a small proportion of molecules have ionised.

e.g. $CH_3COOH_{(aq)} \rightleftarrows CH_3COO^-{}_{(aq)} + H^+{}_{(aq)}$

Question 9.6

Refer to Figure 9.3 and for each situation list the species present in solution. List in order of abundance, most abundant first. Assume all solutions are 1.0 mol L^{-1}.

(i) $NaCl_{(aq)}$ ______________________

(ii) $CH_3COOH_{(aq)}$ ______________________

(iii) $CH_3CH_2OH_{(l)}$ ______________________

Question 9.7

Explain the differences in conductivity of the three solutions illustrated in Figure 9.3.

____

____

Question 9.8

All ionic compounds completely dissociate when they dissolve in water. Even compounds that are only slightly soluble fully dissociate to the extent that they are soluble. Complete the following dissociation equations:

(a) $NaCl_{(s)}$ ______________________

(b) $(NH_4)_2SO_{4(s)}$ ______________________

(c) $BaSO_{4(s)}$ ______________________

Question 9.9

Some covalent molecular compounds are able to either partially or fully ionise in aqueous solutions. Complete the following ionisation equations:

(a) $HCl_{(aq)}$ ______________________

(b) $H_2SO_{4(aq)}$ ______________________

(c) $HNO_{3(aq)}$ ______________________

(d) $CH_3COOH_{(aq)} \rightleftarrows$ ______________________

(e) $NH_{3(aq)} \rightleftarrows$ ______________________

Question 9.10

How is dissociation different to ionisation?

____

____

____

____

Question 9.11

Complete the following table and allocate each of the substances to their correct position:

$NaCl$, $AgCl$, CH_3COOH, H_2O, K_2CO_3, NH_3, CH_3CH_2OH (ethanol), $BaSO_4$, H_3PO_4, CH_4, $NaCH_3COO$, $C_{12}H_{22}O_{11}$ (sugar) , H_2CO_3

ELECTROLYTE STRENGTH	BRIEF DESCRIPTION	EXAMPLES
Strong		
Weak		
Non		

Question 9.12

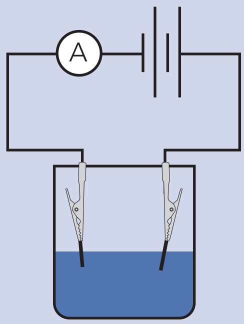
Aqueous solutions of the following are tested for electrical conductivity. Classify the substances as good conductors, weak conductors or non-conductors:

$HCl_{(aq)}$, CH_3OH (methanol), KCH_3COO, $NH_{3(aq)}$, H_2CO_3, $AgCl$, $CaCl_2$, $CaSO_4$, $C_6H_{12}O_6$ (glucose), C_8H_{18} (octane)

(i) good conductors ______________________________

(ii) poor conductors ______________________________

(iii) non-conductors ______________________________



Question 9.13

If ethanoic acid and potassium hydroxide solutions were mixed in the correct stoichiometric ratio so that all reactants were consumed, classify the resulting solution as a strong, weak or non-electrolyte. Explain your answer.

$$CH_3COOH + KOH \rightarrow$$

____

____

____

____

____

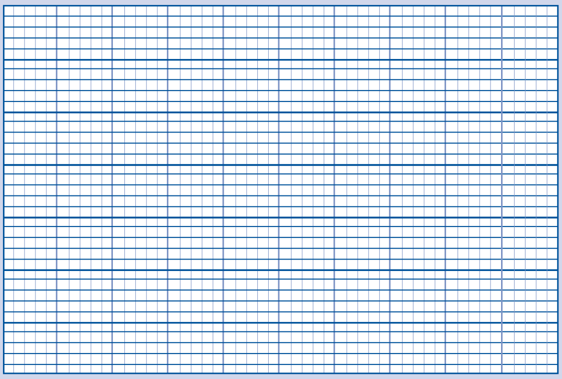
9.4 TEMPERATURE AND SOLUBILITY

Temperature changes usually have a marked effect on the solubility of substances dissolved in water. The data in Question 9.14 illustrates some typical effects.

Question 9.14

Use the following experimental data to plot the solubility curves for the substances listed. The gases CO_2 and O_2 may be omitted from the graph for simplicity.

TEMPERATURE °C		0	20	40	60	80
Solubility gL^{-1} (grams of solute per litre of solution)	NaCl	357	360	366	373	384
	NH_4Cl	294	372	458	552	656
	Sugar	1790	2040	2380	2870	3620
	NH_3	900	530	320	170	65
	CO_2	3.4	1.7	1.0	1.6	–
	O_2	0.07	0.04	0.03	0.02	0.01



Question 9.15

Use the data and graph in the previous question to answer the following.

(a) How is the solubility of solids and gases affected by temperature?

(i) solids ______________________________

(ii) gases ______________________________

(b) A cup of hot coffee (200 mL at 50°C) has 500 g of sugar poured into it. Will all the sugar dissolve? Explain. (Assume the volume of liquid does not change significantly).

____

(c) 100 g of NH_4Cl is dissolved into hot water and the resulting 250 mL of solution is allowed to cool.

(i) At what temperature will crystals appear? ______________

(ii) What type of solution exists at this point? ______________

(iii) What type of solution exists at temperatures below this point if no crystals form?

____

9.5 CHANGES TO BOILING POINT AND FREEZING POINT

The physical properties of a liquid solvent, such as water, are changed when a solute is added. The dissolved solute particles affect the action of the intermolecular forces of the solvent particles. This reduces vapour pressure (which increases boiling point) and also makes freezing more difficult. The table below shows the effect on boiling points (B.P.) and freezing points (F.P.) as solutes are dissolved in water.

MASS OF SOLUTE ADDED TO 1.00 L OF WATER	CONCENTRATION OF SOLUTE PARTICLES MOL L^{-1}	FREEZING POINT °C	BOILING POINT °C
pure water	–	0.0	100.0
100 g sugar ($C_{12}H_{22}O_{11}$)	0.29	–0.5	100.1
200 g sugar ($C_{12}H_{22}O_{11}$)	0.58	–1.1	100.3
100 g of salt (NaCl)	3.42	–6.4	101.8
200 g of salt (NaCl)	6.84	–12.7	103.5
200 g of magnesium fluoride (MgF_2)	9.63	–17.9	104.9

Table 9.2 Predicted boiling points and freezing points of salt and sugar solutions.

Question 9.16

Use the information in Table 9.2 to answer the following questions:

(a) In general, how does the addition of a solute affect the:

(i) freezing point of a solution? ____________________

(ii) boiling point of a solution? ____________________

(b) On which concentration measurement (g L^{-1} of solute or mol L^{-1} of ions in solution) does the variation in F.P. and B.P. appear to be directly related? Explain.

____

(c) Suggest a reason why salt is used to melt ice on frozen roads.

____

(d) Ethylene glycol is often added to the water in car radiators. In cold countries it is usually referred to as “antifreeze” while in warm countries like Australia it is referred to as a coolant. Explain the purpose and effect of this substance in each case.

____

(e) If salty water is boiled, it boils at slightly over 100°C. As the water evaporates it is found that the boiling temperature continues to rise slightly. Explain.

____

(f) Boiling point occurs when the vapour pressure of a liquid is equal to the atmospheric pressure. Use the data in Table 9.2 to explain whether a solute increases or decreases the vapour pressure of a liquid.

____

9.6 CONCENTRATION OF SOLUTIONS

The concentration of a solution refers to the relative amount of one substance (the solute) dissolved in another (the solvent). The actual units used to express this concentration vary to suit the particular situation. Some common ways of expressing concentration of a solution are:

Grams per litre ($g\ L^{-1}$)

$$c = \frac{\text{mass of solute (g)}}{\text{volume of solution (L)}}$$

This unit is useful as it allows for a quick calculation of the mass of solute in a given amount of solution. E.g. the sucrose concentration in cola drinks is usually 0.106 g mL^{-1}. Hence a 375 mL can of this drink contains 39.8 g of sugar.

Moles per litre ($mol\ L^{-1}$)

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

This is the most commonly used method of expressing concentration. It is sometimes referred to as the molarity of a solution.

c = concentration ($mol\ L^{-1}$) n = moles of solute (mol) V = volume of solution (L)

Parts per million (ppm)

This unit is often used for very dilute solutions such as the concentration of salt in tap water or impurities/additives in a product.

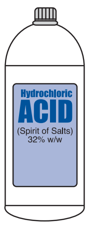
Natural mineral waters typical analysis (ppm)

hydrogencarbonates	380
calcium	145
magnesium	25
chloride	20
sodium	10

$$ppm = \frac{\text{mg of solute}}{\text{kg of solution}}$$

Percentage composition

This unit is effectively the mass of solute (g) per 100 g of solution.

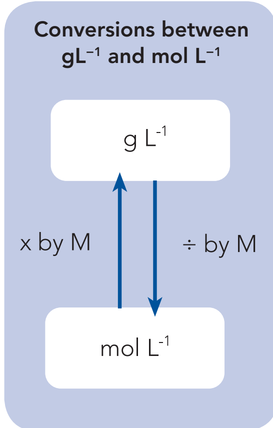


$$\%\ \text{by mass} = \frac{\text{mass of solute (g)}}{\text{mass of solution (g)}} \times 100$$

For example concentrated hydrochloric acid is labelled as 32% w/w. This means that 1.00 kg of this acid solution would contain 320 g of hydrochloric acid.

Worked Example 9.1

A solution of sodium carbonate is made up by dissolving 4.65 g of this substance to make up 250.0 mL of solution. Calculate the concentration of this solution in:
(a) moles per litre; (b) grams per litre.



(a) Firstly determine moles of solute.

$$n = \frac{m}{M} = \frac{4.65}{105.99} = 0.0439 \text{ mol}$$

Hence find concentration of Na_2CO_3 solution.

$$c = \frac{n}{V} = \frac{0.0439}{0.250} = 0.175 \text{ mol L}^{-1}$$

(b) To find concentration in g L^{-1}.

$$c = \frac{\text{mass of solute}}{\text{volume of solution}} = \frac{4.65 \text{ g}}{0.250 \text{ L}} = 18.6 \text{ g L}^{-1}$$

Worked Example 9.2

(a) The concentration of ethanoic (acetic) acid (CH_3COOH) in a sample of vinegar is 38.5 g L^{-1}. What is the concentration in mol L^{-1}?
(b) A salt solution (NaCl) has a concentration of 0.150 mol L^{-1}. What is the concentration in g L^{-1}?

(a) c (CH_3COOH) in vinegar = 38.5 g L^{-1}

$$\textit{or} \quad = \frac{38.5}{60.05} = 0.641 \text{ mol L}^{-1}$$

(b) c (NaCl) in saline solution = 0.150 mol L^{-1}

$$\textit{or} \quad = (0.150)\,(58.44) = 8.77 \text{ g L}^{-1}$$

Worked Example 9.3

The typical analysis for different spring waters varies with their source. One particular brand indicates total dissolved solids of only 80 ppm. What mass of solids would be contained in a 1.50 L bottle of this water? (Assume 1 L of spring water weighs 1 kg.)

c (solids) = 80 ppm = 80 mg per kg
∴ mass of solids = (80 mg) (1.5) = 120 mg = 0.12 g

Worked Example 9.4

A particular brand of beer contains 2.5% (w/w) alcohol (ethanol, CH_3CH_2OH). Determine:
(a) the ethanol concentration in mol L^{-1}
(b) the mass of ethanol in a 200.0 mL glass of this beer.

Assume density of the beer is 1.00 g mL^{-1}.

(a) 2.5% = 2.5 g / 100 g beer

= 25 g / 1000 mL beer (density = 1.00 g mL^{-1})

$$= \frac{25}{46.07} \quad (M_{ethanol} = 46.07 \text{ g mol}^{-1}) = 0.543 \text{ mol L}^{-1}$$

(b) mass of ethanol = 2.5 g / 100 g ∴ = 5.0 g in 200 mL

Question 9.17

A solution of NaOH has a concentration of 1.75 mol L^{-1}.

(a) How many moles of NaOH would 5.0 L of this solution contain?

____

(b) What mass of NaOH would be needed to make up 250 mL of this solution?

____

Question 9.18

A 2.50 kg sample of sea water was evaporated to dryness and 86.5 g of solids remained. Further analysis showed that 73.4 g of the solids were sodium chloride (NaCl). The density of the sea water was also determined to be 1.03 g mL^{-1}.

(a) Calculate the concentration of the solids in:

(i) % by mass ____

(ii) ppm ____

(b) Calculate the concentration of NaCl in sea water in:

(i) ppm ____

(ii) moles per litre ____

Hint: You may need to use density = mass/volume to find the volume of the 2.50 kg sample of sea water.

Question 9.19

Determine the concentration in g L^{-1} and mol L^{-1} of the solute in each of the following solutions.

(a) 20.0 g of potassium nitrate (KNO_3) in 250 mL of solution.

____

(b) 87.7 g of alcohol (ethanol – CH_3CH_2OH) in a 750 mL bottle of wine.

____

(c) 75.6 g of ethanoic acid (CH_3COOH) in a 2.0 L container of vinegar.

____

(d) 44.5 g of sucrose ($C_{12}H_{22}O_{11}$) in a 375 mL can of soft drink.

____

9.7 SOLUBILITY OF IONIC COMPOUNDS

Ionic compounds display a wide variety of solubilities when placed in water. To improve the accuracy of our descriptions of a substance's solubility, the following terms are used:

SOLUBILITY (mol L^{-1})	TERM USED
more than 0.1 moles of solute dissolve per litre of solution	soluble
between 0.001 and 0.1 moles of solute dissolve per litre of solution	slightly soluble
less than 0.001 moles of solute dissolve per litre of solution	insoluble

The solubility of ionic compounds is summarised in the following tables:

SOLUBLE ANIONS (−ve ions)	EXCEPTIONS	
	Insoluble	Sparingly Soluble
most chlorides	$AgCl$	$PbCl_2$
most bromides	$AgBr$	$PbBr_2$
most iodides	AgI, PbI_2	
all nitrates	no exceptions	
all ethanoates	no exceptions	
most sulfates	$SrSO_4$, $BaSO_4$, $HgSO_4$, $PbSO_4$	$CaSO_4$, Ag_2SO_4

INSOLUBLE ANIONS	EXCEPTIONS	
	Soluble	Slightly Soluble
most hydroxides	$NaOH$, KOH, $Ba(OH)_2$, (NH_4OH and $AgOH$ do not exist)	$Ca(OH)_2$, $Sr(OH)_2$
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, $(NH_4)_2S$	

Colours of Metal Ions and their Precipitates

Ionic solids tend to have the same colour as that of any coloured ion they contain. Solutions will also take on the colour of these ions. Precipitates formed from two colourless ions are generally white.

Coloured Ions

ION	COLOUR
Cr^{3+}	deep green
Co^{2+}	pink
Cu^{2+}	blue
Fe^{2+}	pale green
Fe^{3+}	pale brown
Mn^{2+}	very pale pink
Ni^{2+}	green
CrO_4^{2-}	yellow
$Cr_2O_7^{2-}$	orange
MnO_4^{2-}	purple

Exceptions to the Rules (Solids with colours that do not fit the rules)

IONIC SOLID	COLOUR
$CuCO_3$	green
$CuCl_2$	green
CuO	black
CuS	black
PbI_2	yellow
PbS_2	grey
MnO_2	black
Ag_2CO_3	yellow
AgI	pale yellow
Ag_2O	brown
Ag_2S	black

Question 9.20

State if the following ionic salts would be classified as soluble, slightly soluble or insoluble:

$Zn(NO_3)_2$, $AlCl_3$, $(NH_4)_3PO_4$, AgI , $PbCl_2$, $MgSO_4$, $Ca(OH)_2$, $CaSO_4$, $BaSO_4$, Na_2CO_3 , AgCl , $PbSO_4$, Ag_2SO_4 , $NaNO_3$, NaCl , $Mg(OH)_2$, $PbBr_2$, AgBr

SOLUBLE	SLIGHTLY SOLUBLE	INSOLUBLE

9.8 PRECIPITATION REACTIONS

It is important to remember that when ionic solids dissolve in water, the ions dissociate. When different ionic solutions are mixed together a precipitate may form if the appropriate combination of ions is created. For example:

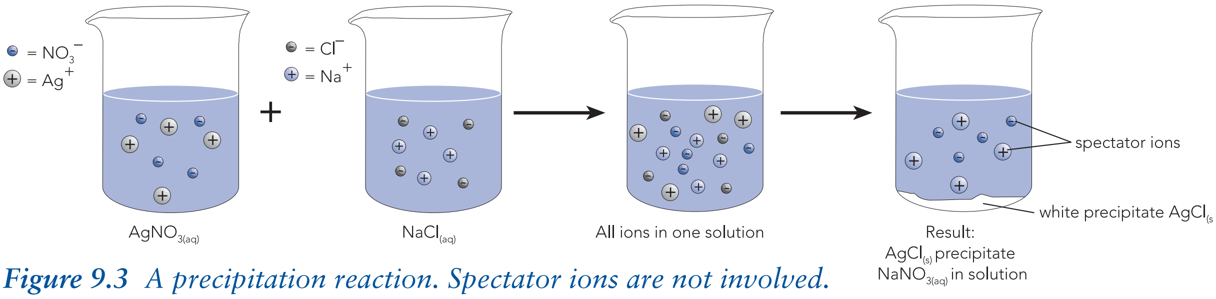


Figure 9.3 A precipitation reaction. Spectator ions are not involved.

Worked Example 9.5

AgNO$_3$ solution is mixed with a NaCl solution. Determine if a precipitate will form.

STEP 1 : List all the ions present and swap them to form possible new pairs.

$$\left.\begin{array}{lll} NaCl_{(aq)} & \rightarrow & Na^+_{(aq)} + Cl^-_{(aq)} \rightarrow \\ AgNO_{3(aq)} & \rightarrow & Ag^+_{(aq)} + NO_3^-{}_{(aq)} \rightarrow \end{array}\right\} \begin{array}{l} NaNO_3\ (?) \\ AgCl\ (?) \end{array}$$

New pairs possible

STEP 2 : Check the solubility of the new pairs of ions from solubility table.

$NaNO_3$ = soluble

$AgCl$ = insoluble (precipitate forms)

STEP 3 : Use an equation to summarise the change.
Only include the ions that form the precipitate.

$$Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}$$

STEP 4 : Check that all particles and changes are balanced and that state variables have been included.

Worked Example 9.6

Determine if a precipitate forms when equal volumes of a 0.25 mol L^{-1} solutions of KCl and Pb(NO$_3$)$_2$ are mixed.

STEP 1 :

$KCl_{(aq)} \rightarrow K^+_{(aq)} + Cl^-_{(aq)} \rightarrow PbCl_2\ (?)$

$Pb(NO_3)_2 \rightarrow Pb^+_{(aq)} + 2NO_3^-{}_{(aq)} \rightarrow KNO_3\ (?)$

STEP 2 :

KNO_3 = soluble

$PbCl_2$ = insoluble (concentration greater than 0.1 mol L^{-1})

STEP 3 & 4 : $Pb^{2+}_{(aq)} + 2Cl^-_{(aq)} \rightarrow PbCl_{2(s)}$

With practice, only the balanced ionic equation needs to be written down.

Question 9.21

Determine if a precipitate will form when each of the following pairs of 0.1 mol L^{-1} solutions are mixed. Write the basic ionic equation for the cases where a precipitate forms.

(a) $Na_2SO_{4(aq)}$ + $Al(NO_3)_{3(aq)}$

STEPS 1 & 2 : ______________________________

STEPS 3 & 4 : ______________________________

(b) $Zn(NO_3)_{2(aq)}$ + $NaOH_{(aq)}$

STEPS 1 & 2 : ______________________________

STEPS 3 & 4 : ______________________________

(c) $CaCl_{2(aq)}$ + $AgNO_{3(aq)}$

STEPS 1 & 2 : ______________________________

STEPS 3 & 4 : ______________________________

(d) $BaCl_{2(aq)}$ + $(NH_4)_2SO_{4(aq)}$

STEPS 1 & 2 : ______________________________

STEPS 3 & 4 : ______________________________

(e) $Pb(NO_3)_{2(aq)}$ + $NaI_{(aq)}$

STEPS 1 & 2 : ______________________________

STEPS 3 & 4 : ______________________________

Question 9.22

Use the solubility rules to determine if a precipitate will form for any combination of ions shown below. Assume at least 0.1 mol L^{-1} solutions of each ion. Where a precipitate will form give the formula and colour of the precipitate as per the example completed in the table.

	K^+	Pb^{2+}	NH_4^+	Ba^{2+}	Ag^+
Cl^-		$PbCl_{2(s)}$ white			
I^-					
SO_4^{2-}					
CO_3^{2-}					

9.9 ACCESS TO DRINKING WATER

A fundamental need of any population is the access to a reliable and high quality source of fresh water. For the majority of the twentieth century it was commonly considered that water catchment from dams in the Darling Scarp would provide sufficient water for Perth's population well into the twenty first century. This view was based on long-term data that showed that these dams had an annual inflow of high quality fresh water averaging nearly 400 gigalitres which would be sufficient for a population of just under 2 million people.

In times of reduced inflow because of drought, ground water supplies were used to supplement water from dams. In the past as much as 40% of the water supply has been from ground water but currently only 10% comes from ground water in the driest years.

Private and business users have also drawn considerable unmetered quantities of water from underground sources. The long-term sustainability of these sources is also affected by the reduced rainfall.

Two changes have occurred in the past 30 years that have altered beliefs and subsequent planning for the water supply to the south west region:

(i) The population of the region has grown at a far greater rate than planners in the 1960's and 70's expected;
(ii) The dramatic decrease in rainfall and has seen the average inflow to the dams in the past decade reduce to just over 100 GL annually.

Three important projects to reduce the increasing stress on Perth's water supply have been:
(i) Two desalination plants have been constructed that use reverse osmosis to purify sea water. These plants now provide just under half of Perth's water supply.
(ii) Recycling waste water. Uses of recycled waste water include irrigating crops; irrigating sports grounds, golf courses, parks; maintaining wetlands; and in dust suppression around construction sites.
(iii) Ground water replenishment, where waste water is recycled by adding it to the aquifers which are the source of our ground water. This water can then be reused as part of our water supply.

Currently about 13.5% of waste water is recycled as per points (ii) and (iii), by 2030 the Water Corporation plans to recycle 30% of all waste water.

Monitoring Ground Water Quality

As ground water plays an important part in our water supply, it is important to monitor its quality. Heavy metal pollution is of particular concern for water drawn from areas that may be affected by domestic discharge/waste from large populations, mining, agriculture, industry and/or water recycling. The heavy metals that are of a concern in drinking water are in the dissolved state, i.e. they occur in the water as ions. Heavy metals of concern include Al, Cd, Cr, Pb, Hg and Zn.

HEAVY METAL	RECOMMENDED MAXIMUM CONCENTRATION	HEALTH ISSUES	TREATMENT OF DRINKING WATER
Aluminium	0.2 mg/L	Considerable evidence that it is neurotoxic	Flocculation and filtration
Cadmium	0.002 mg/L	Long term exposure can cause kidney dysfunction, may be carcinogenic	Addition of lime and $FeCl_3$
Chromium	0.05 mg/L	Carcinogenic, may cause mutations and chromosome aberrations	Addition of lime, ion exchange, reverse osmosis
Lead	0.01 mg/L	Cumulative poison, affects central nervous system, cause kidney damage, interfere with red blood cell and bone formation	Addition of lime
Mercury	0.001 mg/L	Affects kidneys and central nervous system	Use of granulated activated carbon
Zinc	3 mg/L	Not usually a health issue for water supplies.	Addition of alum or lime

Accurate identification of heavy metals in ground water is usually carried out using atomic absorption spectroscopy.

Question 9.23

Lime (CaO) is added to drinking water to reduce the concentration of several heavy metals that are present as ions.

When dissolved in water, lime forms Ca^{2+} ions and OH^{-} ions. Write an ionic equation to represent this process. Use an equation to help explain how the addition of lime would help reduce the concentration of Cd^{2+} ions in water.

Question 9.24

A ground water sample had a small amount of dilute sodium sulfate solution added to it which caused the formation of a white precipitate. This precipitate was collected and a subsequent flame test produced a deep red flame colour.

(a) Name the metal ion responsible for these results and write the ionic equation for the precipitate formation.

(b) Are there any other metal ions that may cause similar or conflicting results? Explain why you chose the metal ion given in your answer to part (a) rather than ions that may cause conflicting results.

REVIEW QUESTIONS

Chapter 9: Aqueous Solutions

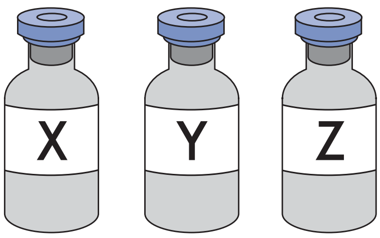
1. A student was given a clear, colourless solution of NaCl. Explain what steps the student could take to determine if the solution was unsaturated, saturated or supersaturated.

2. Complete the following table by placing the substances listed into the correct cells.
You may need to refer to your solubility tables as well as Chapter 5 on bonding.

$NaCl$, HCl , Zn , SiO_2 , Fe , CH_3COOH , Na_2CO_3 , $C_{12}H_{22}O_{11}$ (sugar) , $Cu(NO_3)_2$, Au , olive oil , ammonia , kerosene , ethanol , diamond , petrol

	SUBSTANCE TYPE			
	Covalent Network	Covalent Molecular	Metallic	Ionic
SOLUBLE				
INSOLUBLE				

3. A student was given three bottles labelled X, Y and Z and was told that the bottles contained 0.1 mol L^{-1} ethanoic acid, 0.1 mol L^{-1} sulfuric acid or distilled water.



(a) What single test (no chemical reaction involved) could be used to identify the chemical in each bottle?

(b) Explain how the test results indicate which liquid is in each bottle.

4. Write equations to show what happens when each of the substances listed is dissolved in water:

(a)	hydrogen chloride gas	(e)	sodium hydroxide
(b)	magnesium sulfate	(f)	ammonia gas
(c)	ethanoic acid (CH_3COOH)	(g)	potassium iodide
(d)	phosphoric acid	(h)	aluminium sulfate

5. It has been observed that ammonia and ethanoic acid solutions are weak conductors of electricity whereas the solution formed when they are mixed is a good conductor. Explain these observations.

6. A 50 000 litre swimming pool requires a "total alkalinity" of between 120 and 160 ppm. The "total alkalinity" is managed by adding Na_2CO_3 to the water. Would adding 5.50 kg of Na_2CO_3 be enough to create a "total alkalinity" reading within the accepted range if the initial reading was 0 ppm? (Assuming that 1.0 L of the pool water weighs 1.0 kg.)

7. The nitrogen (N) content of a sample of water from the Swan River was found to be 1.0 mg L^{-1}. Convert this to a concentration measured in mol L^{-1}.

8. A sample of river water was found to have a salt (NaCl) content of 20.0 g L^{-1} while a sample of lake water was found to have a salt content of 0.500 mol L^{-1}. How much salt could be extracted from 750 L of each type of water?

9. Write ionic equations for the following reactions. If no precipitate forms write "no reaction". Assume that all solutions are 0.1 mol L^{-1}.

(a) $AgNO_{3(aq)} + NaBr_{(aq)}$ (b) $FeCl_{3(aq)} + NaOH_{(aq)}$

(c) $Na_3PO_{4(aq)} + Cu(NO_3)_{2(aq)}$ (d) $Cr_2(SO_4)_{3(aq)} + (NH_4)_2CO_{3(aq)}$

(e) $Pb(NO_3)_{2(aq)} + KI_{(aq)}$ (f) $SrCl_{2(aq)} + Fe_2(SO_4)_{3(aq)}$

10. Name the metal ions that responsible for the following observations:

(a) Solutions of this metal ion did not from precipates when mixed with solutions of potassium carbonate, potassium sulfate and potassium hydroxide. When a the salt of this ion was placed in a bunsen burner flame an yellow colour was produced.

(b) A blue solution produced a green precipitate when mixed with a potassium carbonate solution.

(c) When a nitrate solution of this metal ion was mixed with a sodium chloride solution no precipitate was formed but a white precipitate was form when the nitrate solution was mixed with a sodium sulfate solution. When placed in a bunsen burner flame, the nitrate salt of this metal produced a pale green flame.

11. Calculate the concentration, in mol L^{-1}, of each of the following solutions:

(a) 15.2 g of sodium carbonate was dissolved in enough water to make 375 mL of solution.

(b) 2.50 kg of calcium chloride was added to a swimming pool that contained 45.0 m^3 of water.

(c) 9.47 g of hydrogen chloride was dissolved in enough water to create 0.250 L of solution.

12. A 0.150 kg sample of sludge from the bottom of a river was found to contain 1.026×10^{-4} g of mercury. Determine the concentration of the mercury in the sludge in parts per million.

13. The label on a brand of toothpaste states that it contains 0.32% w/w sodium fluoride. What mass of sodium fluoride is contained in 0.750 g of toothpaste?

14. Water is essential to life and provides a medium for many chemical reactions. Since it is a good solvent, the water around us in natural aquatic systems is not pure. It contains varying amounts of dissolved substances. For each of the following substances explain their importance or significance if dissolved in groundwater. The first one is done for you.

(i) Dissolved carbon dioxide.

Means that rainwater is slightly acidic – dissolves underground limestone – causes cave formation. Also, dissolved CO_2 *allows photosynthesis in aquatic systems.*

(ii) Dissolved oxygen.

____

____

(iii) Dissolved nitrogen and phosphorus compounds.

____

____

(iv) Dissolved salts due to rising water table.

____

____

(v) Dissolved salts of heavy metals such as mercury.

____

____

(vi) Excess dissolved nutrients from waste water, detergents and fertilisers.

____

____

FOR THE EXPERTS

15. Many Australian soils are lacking in nutrients such as phosphorus and nitrogen. Gardeners and farmers often apply fertilisers to improve the soil quality.

(a) The label on a container of soluble garden fertiliser stated that it had a 4.5% phosphorus content.

Determine the P concentration (in mol L^{-1}) produced when two scoopfuls, or 15 g, of this fertiliser was dissolved in 9.0 L of water.

(b) Superphosphate is regularly used on farms to improve soils before planting legumes. Superphosphate is a mixture of the soluble compounds calcium dihydrogenphosphate and calcium sulfate dihydrate.

(i) Write the equation showing what occurs when superphosphate is dissolved in water.

(ii) A manufacturer claims that their superphosphate is 8.8% P, of which 95% is useable by plants.

Calculate the amount of phosphorus that would be made available to plants if a farmer spread 10 tonne of superphosphate onto the paddocks (1 tonne = 1000 kg).

ACIDS AND BASES

10

Topics covered in this chapter:

10.1 WATER – A VERY WEAK ELECTROLYTE

Water is a very weak electrolyte, the ionisation of water can be shown as:

$$H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$$

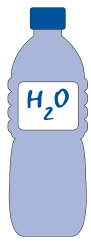
This occurs to such a small extent that in any aqueous solution the product of the H^+ ion concentration and the OH^- concentration is 1.00×10^{-14}.

i.e. $[H^+]\,.\,[OH^-] = 1.00 \times 10^{-14}$

$$\left[\begin{array}{l}\text{concentration of}\\ \text{the hydrogen}\\ \text{ion in mol L}^{-1}\end{array}\right] \times \left[\begin{array}{l}\text{concentration of}\\ \text{the hydroxide}\\ \text{ion in mol L}^{-1}\end{array}\right] = \left[\begin{array}{l}\text{ionisation constant}\\ (K_W)\text{ for water at }25^{\circ}\text{C}\end{array}\right]$$

Because the concentration of H^+ ions and OH^- ions is so low, the electrical conductivity of pure water is extremely low.

A **neutral** solution will have equal concentrations of H^+ ions and OH^- ions.



i.e. at 25°C , $[OH^-] = [H^+]$

If $[H^+]\,.\,[OH^-] = 1.00 \times 10^{-14}$ then for a neutral solution ,

$[H^+]\,.\,[OH^-] = \sqrt{(1.00 \times 10^{-14})} = 1.00 \times 10^{-7}$ mol L^{-1}

An **acidic** solution will have $[H^+] > [OH^-]$,

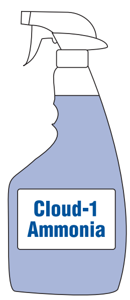
or $[H^+] > 1.00 \times 10^{-7}$ mol L^{-1}

and $[OH^-] < 1.00 \times 10^{-7}$ mol L^{-1}

A **basic** solution will have $[OH^-] > [H^+]$,

or $[OH^-] > 1.00 \times 10^{-7}$ mol L^{-1}

and $[H^+] < 1.00 \times 10^{-7}$ mol L^{-1}



Worked Example

10.1 An environmental chemist was investigating the effect of air pollution on the hydrogen ion concentration in lake water. The chemist found the $[H^+]$ to be 6.12×10^{-6} mol L^{-1}. Is the lake water acidic or basic?

$[H^+]$ in lake water = 6.12×10^{-6} mol L^{-1}

$[H^+]$ in neutral water = 1.00×10^{-7} mol L^{-1}

$[H^+]$ in lake water > 1.00×10^{-7} mol L^{-1} $\therefore$ lake water is acidic.

10.2 Ammonium nitrate (NH_4NO_3) is a commonly used fertiliser that increases the nitrogen content of the soil. When rain dissolves the ammonium nitrate the following occurs:

$$NH_4NO_{3\,(s)} \rightarrow NH_4^+{}_{(aq)} + NO_3^-{}_{(aq)}$$

$$NH_4^+{}_{(aq)} + H_2O_{(l)} \rightleftharpoons NH_{3(aq)} + H_3O^+{}_{(aq)}$$

Will this cause the soil to become acidic or basic?

The ammonium ion reacts with water (hydrolysis) producing the weak base ammonia and hydronium ions.

$$NH_4^+{}_{(aq)} + H_2O_{(l)} \xrightarrow{H^+} NH_{3(aq)} + H_3O^+{}_{(aq)}$$

The increased presence of $H_3O^+{}_{(aq)}$ makes the soil acidic.

10.2 THEORIES ON ACIDS AND BASES

Early descriptions of acids and bases (pre 19th Century):

1. Acids taste sour, bases bitter.	4. Acids and bases can affect the colour of some vegetable extracts.
2. Bases feel slippery or soapy.	5. Acids react with some metals to form a salt.
3. Acids react with bases to form salts.	6. Acids and bases are corrosive.

Arrhenius Theory (1880)

1. When dissolved in water, acids form H^+ ions:

   e.g. $HCl_{(g)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$

   The H^+ ion is responsible for the solution's acid properties:

2. When dissolved in water, bases form OH^- ions.

   e.g. $NaOH_{(s)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$

   The OH^- ion is responsible for the solution's basic properties.

3. When acids are added to bases in the correct stoichiometric ratio, **neutralisation** occurs:

   i.e. $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

Brønsted-Lowry Theory (1923)

1. In a reaction, the substance losing or donating a H^+ ion (called a proton) is an acid.

2. The substance gaining or accepting the proton (H^+ ion) is a base.

loses a proton to form

e.g. 1 $CH_3COOH_{(aq)} + NH_{3(aq)} \rightarrow CH_3COO^-_{(aq)} + NH_4^+{}_{(aq)}$

proton donation   gains a proton to form

CH_3COOH acting as an acid, NH_3 as a base.

donates a proton to form

e.g. 2 $HF_{(aq)} + H_2O_{(l)} \rightleftharpoons F^-_{(aq)} + H_3O^+_{(aq)}$

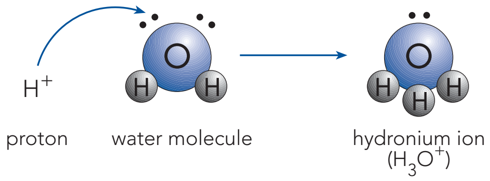
proton donation   accepts a proton to form

HF acting as an acid, H_2O as a base.

H^+ – A closer look!

The simplest atom, hydrogen, consists of a proton surrounded by an electron. Hence the symbol **H^+** can be used to represent a **proton.**

Protons are very small and readily attach themselves to the negative end of say, polar substances. In water they combine with H_2O to form the **hydronium ion, H_3O^+**.



Question 10.1

Use greater than (>), less than (<) or equal signs (=) to complete the following table:

ACIDIC SOLUTIONS	BASIC SOLUTIONS
[H^+] .... [OH^-]	[H^+] .... [OH^-]
[H^+] .... 1.0×10^{-7} mol L^{-1}	[H^+] .... 1.0×10^{-7} mol L^{-1}
[OH^-] .... 1.0×10^{-7} mol L^{-1}	[OH^-] .... 1.0×10^{-7} mol L^{-1}

Question 10.2

Use Arrhenius' Theory to show what happens when the following substances are dissolved in water. State if the substance is acidic or basic.

(a) $HNO_{3(aq)} \rightarrow$, ____________________

(b) $HClO_{(aq)} \rightarrow$, ____________________

(c) $KOH_{(s)} \rightarrow$, ____________________

(d) $Ba(OH)_{2(s)} \rightarrow$, ____________________

Question 10.3

Identify the acid and base in each of the following reactions:

(a) $CH_3COOH_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$

(b) $HSO_4^-{}_{(aq)} + CO_3^{2-}{}_{(aq)} \rightleftharpoons SO_4^{2-}{}_{(aq)} + HCO_3^-{}_{(aq)}$

(c) $H_3PO_{4(aq)} + 3NaOH_{(aq)} \rightleftharpoons Na_3PO_{4(aq)} + 3H_2O_{(l)}$

(d) $NH_{3(aq)} + HCl_{(aq)} \rightleftharpoons NH_4^+{}_{(aq)} + Cl^-_{(aq)}$

(e) $PO_4^{3-}{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HPO_4^{2-}{}_{(aq)} + OH^-_{(aq)}$

Question 10.4

Use the Brønsted-Lowry Theory to show how water is capable of reacting as an acid and as a base.

ACID: $H_2O_{(l)}$ + ______________________________

BASE: $H_2O_{(l)}$ + ______________________________

Question 10.5

A hydrangea plant that normally produces blue flowers can often be made to produce pink flowers by adding hydrated lime ($Ca(OH)_2$) to the soil. Will the hydrated lime make the soil acidic or basic? Use an equation to support your answer.

____

____

____

Question 10.6

Sulfur is a common impurity in fossil fuels such as petrol and coal. The combustion of these fuels produces the acidic oxides SO_2 and SO_3. Write equations to illustrate why these oxides are considered acidic in aqueous solution.

____

____

10.3 IDENTIFYING SOLUTIONS AS ACIDIC OR BASIC

The acidic or basic nature of a solution can be identified by addition of an indicator. The indicator will change colour depending on the hydrogen ion concentration of the solution. This colour change tends not to occur at an exact $[H^+]$ but rather over a range of concentrations.

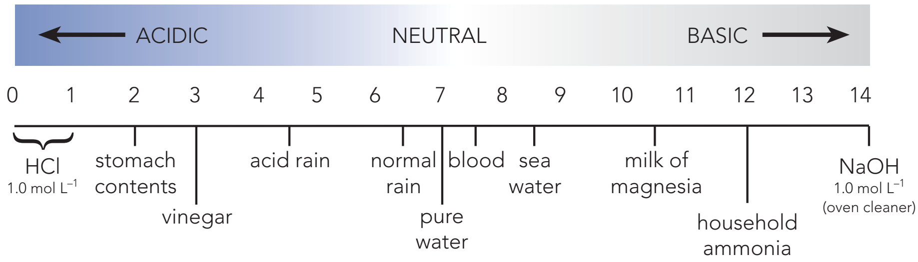
INDICATOR	COLOUR IN ACIDS	COLOUR IN BASES	$[H^+]$ RANGE OVER WHICH COLOUR CHANGE OCCURS MEASURED IN Mol L^{-1}	pH RANGE OVER WHICH COLOUR CHANGE OCCURS
Litmus	Red	Yellow	1×10^{-5} - 1×10^{-8}	5 - 8
Methyl orange	Red	Yellow	7.9×10^{-4} - 4×10^{-5}	3.1 - 4.4
Phenolphthalein	Colourless	Pink	5×10^{-9} - 1×10^{-10}	8.3 - 10
Bromothymol blue	Yellow	Blue	1×10^{-6} - 2.5×10^{-8}	6 - 7.6

pH NUMBERS

pH numbers are usually used to help describe solutions of low acid/base concentration – such as waterways, shampoos, body fluids, soil conditions, etc.

$$\mathbf{pH = -log\,[H^+]}$$

pH	<	7	acidic
pH	=	7	neutral
pH	>	7	basic



Worked Examples (for experts only)

10.3 A bottle of HCl has a hydrogen ion content 5.65×10^{-5} mol L^{-1}. Calculate its pH.

$$\begin{aligned} pH &= -\log [H^+] \\ &= -\log (5.65 \times 10^{-5}) \\ pH &= 4.25 \end{aligned}$$

10.4 A sample of pond water has a pH of 7.95, calculate the hydrogen ion concentration of this water.

$$\begin{aligned} pH &= -\log [H^+] \\ [H^+] &= \text{inv. log } (-pH) \\ &= \text{inv. log } (-7.95) \\ [H^+] &= 1.12 \times 10^{-8} \text{ mol L}^{-1} \end{aligned}$$

on some calculators there is no inv. log, use 10^x where x = (-pH)

$[H^+] = 10^{-pH} = 10^{-7.95}$

$[H^+] = 1.12 \times 10^{-8}$ mol L^{-1}

Question 10.7

A vinegar solution has a hydrogen ion content of 4.16×10^{-4} mol L^{-1}.

(a) Comment on the ethanoic acid concentration of the vinegar.
(HINT: compare $[CH_3COOH]$ to $[H^+]$ and explain the difference.)

(b) Calculate the pH of this solution.

Question 10.8

A detergent has a pH of 9.00.

(a) Calculate the hydrogen ion content of the detergent.

____

____

(b) Would solutions of this detergent be acidic, basic or neutral? Explain.

____

____

10.4 THE STRENGTH OF ACIDS AND BASES

Revision

- Covalent molecular substances **ionise** when water causes them to break up to form ions.
- Ionic substances **dissociate** when water causes their ions to separate.
- The extent to which a substance ionises or dissociates when dissolved in water can be determined by measuring its electrical conductivity.

Strong acids

Completely ionise in water to produce H^+ (or H_3O^+) ions.

e.g. $HNO_{3(aq)} \rightarrow H^+_{(aq)} + NO_3^-{}_{(aq)}$

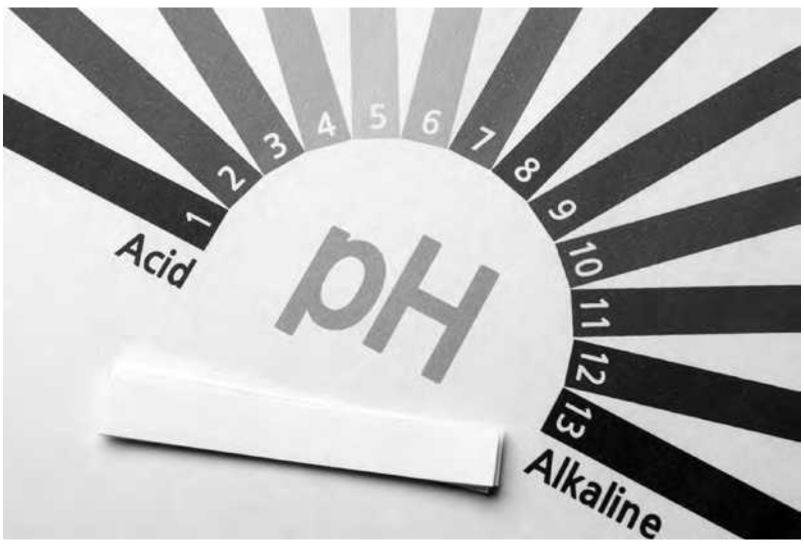
or $HNO_{3(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + NO_3^-{}_{(aq)}$

Weak acids

Partially **ionise** in water to produce H^+ (or H_3O^+) ions. (Partially meaning that only a small percentage of the molecules will break up to form ions.)

e.g. $H_2CO_{3(aq)} \rightleftharpoons H^+_{(aq)} + HCO_3^-{}_{(aq)}$

or $H_2CO_{3(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HCO_3^-{}_{(aq)}$



Strong bases

Completely **dissociate** in water to produce OH^- ions.

e.g. $KOH_{(aq)} \rightarrow K^+_{(aq)} + OH^-_{(aq)}$

Weak bases

Partially **dissociate** in water to produce OH^- ions.

e.g. $Na_2CO_{3(aq)} + H_2O_{(l)} \rightarrow 2Na^+_{(aq)} + CO_3^{2-}{}_{(aq)} + H_2O_{(l)}$

and $CO_3^{2-}{}_{(aq)} + H_2O_{(l)} \rightarrow HCO_3^-{}_{(aq)} + OH^-_{(aq)}$

(Being ionic it completely dissociates but only a small quantity of OH^- ions produced).

Ammonia is a weak base that is covalent molecular and so partially ionises.

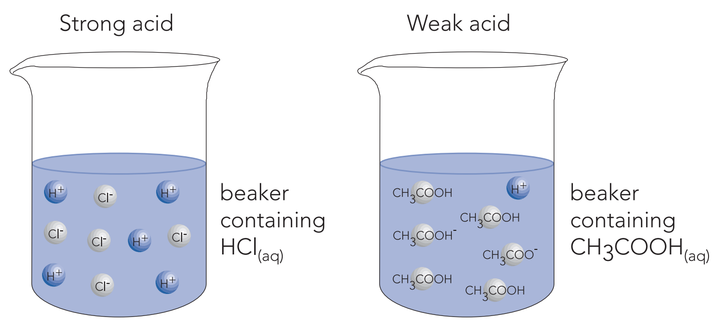
e.g. $NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-_{(aq)}$

Equation conventions

A single arrow " → " indicates complete ionisation or dissociation.
A double arrow " ⇌ " indicates partial ionisation or dissociation.

STRONG ACIDS	WEAK ACIDS
HCl	most other acids
H_2SO_4	H_2CO_3
HNO_3	H_3PO_4
HBr	CH_3COOH
HI	HF

STRONG BASES	WEAK BASES
Group I and II oxides and hydroxides	most other bases
	NH_3
	Na_2CO_3
	$KHCO_3$



Complete ionisation of HCl molecules (only ions exist in solution)

Partial ionisation of CH_3COOH molecules (mainly molecules and only a few ions exist in solution)

BASIC OXIDES	AMPHOTERIC OXIDES	NEUTRAL OXIDES	ACIDIC OXIDES
Na_2O , K_2O	Al_2O_3 , Cr_2O_3	H_2O	CO_2 , NO_2
CaO , MgO	ZnO , PbO_2	CO	P_4O_{10} , SO_2

A very general rule is that oxides involving ionic bonding tend to be basic whereas oxides involving covalent bonding tend to be acidic. (If you have studied electronegativity, elements with an electronegativity less than 1.3 tend to form basic oxides, elements with an electronegativity greater than 2.0 tend to form acidic oxides and those with an electronegativity between 1.3 and 2.0 tend to form amphoteric oxides.) Metal oxides tend to be basic and non-metal oxides tend to be acidic.

An **amphoteric** substance can behave as an acid or a base.

e.g. 1 $Al(OH)_3$ acting as an acid:

$Al(OH)_{3(s)} + OH^-_{(aq)} \rightarrow [Al(OH)_4]^-_{(aq)}$

e.g. 2 Al_2O_3 acting as an acid:

$Al_2O_{3(s)} + 2OH^-_{(aq)} + 3H_2O_{(l)} \rightarrow 2[Al(OH)_4]^-_{(aq)}$

e.g. 3 $Al(OH)_3$ acting as a base:

$$Al(OH)_{3(s)} + 3H^+_{(aq)} \rightarrow Al^{3+}_{(aq)} + 3H_2O_{(l)}$$

e.g. 4 Al_2O_3 acting as a base:

$$Al_2O_{3(s)} + 6H^+_{(aq)} \rightarrow 2Al^{3+}_{(aq)} + 3H_2O_{(l)}$$

Aqueous Solutions of Polyprotic Acids

Some acids have more than one hydrogen atom per molecule that can be released when forming an aqueous solution. H_2SO_4 has the potential to release two hydrogen ions in an aqueous solution.

First ionisation: $H_2SO_4 \rightarrow H^+_{(aq)} + HSO_4^-{}_{(aq)}$

As H_2SO_4 is a strong acid, practically all H_2SO_4 molecules ionise.

Second ionisation: $HSO_4^-{}_{(aq)} \rightleftharpoons H^+_{(aq)} + SO_4^{2-}{}_{(aq)}$

Few H^+ and SO_4^{2-} ions form in the second ionisation as HSO_4^- is a weak acid.

Other polyprotic acids you may use in a chemistry laboratory are phosphoric acid (H_3PO_4), carbonic acid (H_2CO_3) and oxalic acid ($C_2H_2O_4$ or HOOCCOOH).

Each mole of diprotic acid (H_2SO_4 and H_2CO_3) requires two moles of OH^- for neutralisation whereas each mole of tripotic acid (H_3PO_4) requires three moles of OH^- for neutralisation.

Question 10.9

Write equations to show what occurs when the following acidic substances are dissolved in water:

(a) $HBr \rightarrow$ ____________________

(b) $H_3PO_4 \rightleftharpoons$ __________ $\rightleftharpoons$ __________ $\rightleftharpoons$ __________

(c) $SO_2 + H_2O \rightleftharpoons H_2SO_3 \rightleftharpoons$ ____________________

(d) $SO_3 + H_2O \rightarrow H_2SO_4 \rightarrow$ __________ $\rightleftharpoons$ __________

Question 10.10

Write equations to show what happens when the following basic substances are dissolved in water:

(a) $Na_2O_{(s)} + H_2O_{(l)} \rightarrow$ ____________________

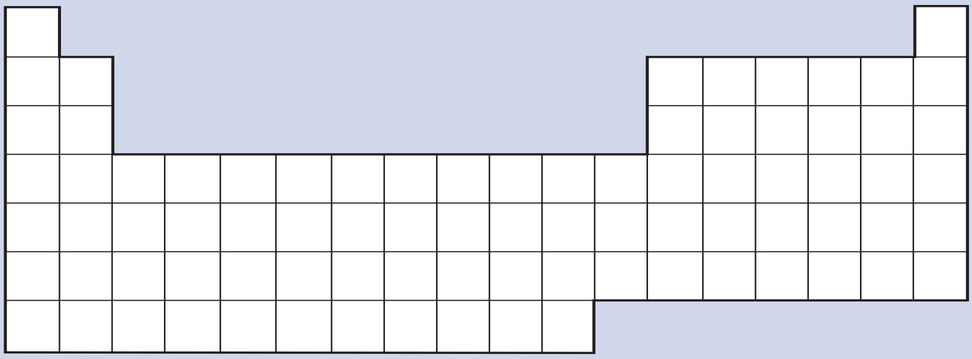
(b) $NH_{3(g)} + H_2O_{(l)} \rightleftharpoons$ ____________________

(c) $PO_4^{3-}{}_{(aq)} + H_2O_{(l)} \rightleftharpoons$ ____________________

(d) $KHCO_{3(s)} + H_2O \rightarrow$ ____________________

$\rightleftharpoons$ ____________________

Question 10.11



Use the periodic table above to:

- illustrate trends about acids and bases given by the periodic table
- indicate the position of elements that form neutral oxides and those that form amphoteric oxides and hydroxides.

10.5 GENERAL REACTIONS OF ACIDS

It is always preferable to write equations that include only those species involved in change and not include "spectator" ions. In the general reactions given here, the complete equation will also be given to allow you to compare the information given by each equation.

A second example is always given for you to complete.

1. acid + base → salt + water

e.g. 1. $HCl_{(aq)} + KOH_{(aq)} \rightarrow KCl_{(aq)} + H_2O_{(l)}$

or $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

e.g. 2. $HNO_{3(aq)} + Ca(OH)_{2(aq)} \rightarrow$ ____________________

or ____________________ → ____________________

2. acid + reactive metal → salt + hydrogen gas

e.g. 1. $H_2SO_{4(aq)} + Fe_{(s)} \rightarrow FeSO_{4(aq)} + H_{2(g)}$

or $2H^+_{(aq)} + Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + H_{2(g)}$

e.g. 2. $HCl_{(aq)} + Al_{(s)} \rightarrow$ ____________________

or ____________________ → ____________________

3. acid + metal oxide → salt + water

(NB: metal oxides are basic)

e.g. 1. $2HNO_{3(aq)} + MgO_{(s)} \rightarrow Mg(NO_3)_{2(aq)} + H_2O_{(l)}$

or $2H^+_{(aq)} + MgO_{(s)} \rightarrow Mg^{2+}_{(aq)} + H_2O_{(l)}$

e.g. 2. $HCl_{(aq)} + Al_2O_{3(s)} \rightarrow$ ____________________

or ____________________ → ____________________

4. acid + carbonate → salt + carbon dioxide + water

e.g. 1. $H_2SO_{4(aq)} + CaCO_{3(s)} \rightarrow CaSO_{4(s)} + CO_{2(g)} + H_2O_{(l)}$

or $2H^+_{(aq)} + CaCO_{3(s)} + SO_4^{2-}{}_{(aq)} \rightarrow CaSO_{4(s)} + CO_{2(g)} + H_2O_{(l)}$

e.g. 2. $HNO_{3(aq)} + K_2CO_{3(aq)} \rightarrow$ ____________________

or ____________________ → ____________________

5. acid + hydrogen carbonate → salt + carbon dioxide + water

e.g. 1. $HNO_{3(aq)} + KHCO_{3(aq)} \rightarrow KNO_{3(aq)} + CO_{2(g)} + H_2O_{(l)}$

or $H^+_{(aq)} + HCO_3^-{}_{(aq)} \rightarrow CO_{2(g)} + H_2O_{(l)}$

e.g. 2. $H_2SO_{4(aq)} + Mg(HCO_3)_{2(aq)} \rightarrow$ ____________________

or ____________________ → ____________________

6. (strong) acid + sulfite → salt + sulfur dioxide + water

e.g. 1. $2HCl_{(aq)} + Na_2SO_{3(aq)} \rightarrow 2NaCl_{(aq)} + SO_{2(g)} + H_2O_{(l)}$

or $2H^+_{(aq)} + SO_3^{2-}{}_{(aq)} \rightarrow SO_{2(g)} + H_2O_{(l)}$

e.g. 2. $H_2SO_{4(aq)} + K_2SO_{3(aq)} \rightarrow$ ____________________

or ____________________ → ____________________

7. (strong) acid + sulfide → salt + hydrogen sulfide

e.g. 1. $HNO_{3(aq)} + Na_2S_{(aq)} \rightarrow 2NaNO_{3(aq)} + H_2S_{(g)}$

or $2H^+_{(aq)} + S^{2-}{}_{(aq)} \rightarrow$

e.g. 2. $HCl_{(aq)} + K_2S_{(aq)} \rightarrow$ ____________________

or ____________________ → ____________________

8. ammonium salts + hydroxide → salt + ammonia + water

e.g. 1. $NH_4Cl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + NH_{3(aq)} + H_2O_{(l)}$

or $NH_4^+{}_{(aq)} + OH^-{}_{(aq)} \rightarrow NH_{3(aq)} + H_2O_{(l)}$

e.g. 2. $NH_4NO_{3(aq)} + KOH_{(aq)} \rightarrow$ ____________________

or ____________________ → ____________________

9. non-metal oxide + base → salt + water

(NB: non-metal oxides are acidic)

e.g. 1. $SO_{2(g)} + 2NaOH_{(aq)} \rightarrow Na_2SO_{3(aq)} + H_2O_{(l)}$

or $SO_{2(g)} + 2OH^-{}_{(aq)} \rightarrow SO_3^{2-}{}_{(aq)} + H_2O_{(l)}$

e.g. 2. $CO_{2(g)} + KOH_{(aq)} \rightarrow$ ________

or ________ $\rightarrow$ ________

10. amphoteric metal + base → complex ion + hydrogen gas

e.g. 1. $2Al_{(s)} + 2NaOH_{(aq)} + 6H_2O_{(l)} \rightarrow 2NaAl(OH)_{4(aq)} + 3H_{2(g)}$

or $2Al_{(s)} + 2OH^-_{(aq)} + 6H_2O_{(l)} \rightarrow 2Al(OH)_4{}^-{}_{(aq)} + 3H_{2(g)}$

e.g. 2. $Zn_{(s)} + KOH_{(aq)} + 2H_2O_{(l)} \rightarrow$ ________

or ________ $\rightarrow$ ________

11. amphoteric metal oxide + base → complex ion

e.g. 1. $Al_2O_{3(s)} + 2NaOH_{(aq)} + 3H_2O_{(l)} \rightarrow 2NaAl(OH)_{4(aq)}$

or $Al_2O_{3(s)} + 2OH^-_{(aq)} + 3H_2O_{(l)} \rightarrow 2Al(OH)_4{}^-{}_{(aq)}$

e.g. 2. $ZnO_{(s)} + KOH_{(aq)} + 2H_2O_{(l)} \rightarrow$ ________

or ________ $\rightarrow$ ________

12. amphoteric metal hydroxide + base → complex ion

e.g. 1. $Zn(OH)_{2(s)} + 2NaOH_{(aq)} \rightarrow Na_2Zn(OH)_{4(aq)}$

or $Zn(OH)_{2(s)} + 2OH^-_{(aq)} \rightarrow Zn(OH)_4{}^{2-}{}_{(aq)}$

e.g. 2. $Al(OH)_{3(s)} + KOH_{(aq)} \rightarrow$ ________

or ________ $\rightarrow$ ________

Complex ions produced from amphoteric substances

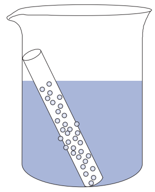
Al, Al_2O_3, $Al(OH)_3$	+	OH^-	→	$Al(OH)_4{}^-{}_{(aq)}$	tetrahydroxoaluminate ion (or aluminate ion) (colourless)
Zn, ZnO, $Zn(OH)_2$	+	OH^-	→	$Zn(OH)_4{}^{2-}{}_{(aq)}$	tetrahydroxozincate ion (or zincate ion) (colourless)
Cr, Cr_2O_3, $Cr(OH)_3$	+	OH^-	→	$Cr(OH)_4{}^-{}_{(aq)}$	tetrahydroxochromate ion (or chromite ion) (green)

(NB: Amphoteric substances are extra to syllabus requirements).

Why no spectators?

An equation is a very useful summary of what has happened in a chemical reaction. When writing an equation, think of it as describing the changes that occurred in the reaction. If a chemical species has not changed then it does not need to be included in the equation.

$CaCO_{3(s)}$ is dissolved by the H^+ ion to form Ca^{2+} ions, $CO_{2(g)}$ and $H_2O_{(l)}$. The Cl^- plays no part.


Chalk in acid

or: $CaCO_{3(s)} + 2H^+_{(aq)} \rightarrow Ca^{2+}_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

Question 10.12

Write the balanced equation for each of the following. Where appropriate write the formula for the reacting species – not the entire formula – i.e. use $H^+_{(aq)}$ instead of $HCl_{(aq)}$.

(a) $HCl_{(aq)} + Mg(OH)_{2(s)}$ ____________________

(b) $Al_{(s)} + HCl_{(aq)}$ ____________________

(c) $SO_{3(g)} + H_2O_{(l)}$ ____________________

(d) $Al_{(s)} + NaOH_{(aq)}$ ____________________

(e) $SO_{3(g)} + NaOH_{(aq)}$ ____________________

(f) $Al(OH)_{3(s)} + NaOH_{(aq)}$ ____________________

(g) $H_2SO_{4(aq)} + Ca(OH)_{2(aq)}$ ____________________

(h) $ZnCO_{3(s)} + H_3PO_{4(aq)}$ ____________________

(i) $H_3PO_{4(aq)} + Al(OH)_{3(s)}$ ____________________

(j) $K_2CO_{3(aq)} + H_2SO_{4(aq)}$ ____________________

(k) ammonia solution + nitric acid

____

(l) ammonia solution + sulfuric acid

____

(m) iron (III) hydroxide + hydrochloric acid

____

(n) sulfur trioxide gas + calcium hydroxide solution

____

(o) barium oxide + nitric acid

____

(p) barium hydroxide solution + sulfuric acid

____

Question 10.13

Explain what an amphoteric substance is.

____

____

____

Question 10.14

Complete and balance the following equations that demonstrate the amphoteric nature of zinc, aluminium and chromium.

(a) $Zn_{(s)} + H^+_{(aq)} \rightarrow$ ____________________

(b) $Al_{(s)} + OH^-_{(aq)} + H_2O_{(l)} \rightarrow$ ____________________

(c) $Cr(OH)_{3(s)} + OH^-_{(aq)} \rightarrow$ ____________________

(d) $Cr(OH)_{3(s)} + H^+_{(aq)} \rightarrow$ ____________________

(e) $Cr_2O_{3(s)} + H^+_{(aq)} \rightarrow$ ____________________

(f) $Cr_2O_{3(s)} + OH^-_{(aq)} + H_2O_{(l)} \rightarrow$ ____________________

10.6 ACID RAIN

A modern, industrial society has many processes that cause acidic oxides (non-metal oxides) to be released into the atmosphere. The extraction of metals from their ores and the burning of fossil fuels to release stored chemical energy are essential for maintaining lifestyles enjoyed in Australia and many other countries. However they both have associated problems that cause major environmental damage that need to be better managed.

Both processes release vast quantities of non-metal oxides into the atmosphere. These non-metal oxides can alter the pH of the water they dissolve in.

An important step in the production of carbon-steel (the most commonly used metal alloy) is to direct pure oxygen over molten cast iron. This process reduces the amount of carbon and sulfur in the steel.

i.e. $C + O_2 \rightarrow CO_2$

$S + O_2 \rightarrow SO_2$

The CO_2 and SO_2 are released into the atmosphere.

Question 10.15

Write balanced equations to show what happens when the CO_2 and SO_2 dissolve in moisture in the atmosphere.

(i) CO_2 + H_2O ______________________________

(ii) SO_2 + H_2O ______________________________

Question 10.16

Explain what effect these gases would have on the pH of rainwater.

____

Question 10.17

Many famous statues and structures are made of limestone or marble which are predominantly calcium carbonate.

Describe what will happen when rainwater containing CO_2 and SO_2 (acid rain) falls on such structures. It is important that, as a student of chemistry, your description includes chemical equations to summarise the changes occurring.

____

____

____

____

____

____

The combustion of fossil fuels, such as petrol and diesel, also releases large amounts of CO_2 and SO_2 into the atmosphere. In a car engine the temperatures are high enough to cause the normally inert N_2 to react with O_2 to form nitrogen oxides.

$$\text{e.g. } N_2 + 2O_2 + \text{energy} \rightarrow 2NO_2$$

These nitrogen oxides are soluble in water.

Question 10.18

What effect will dissolved NO_2 have on the pH of rain water?

___

Question 10.19

Write a balanced equation for NO_2 dissolving in water.

___

Catalytic converters

Catalytic converters are part of a car's exhaust system and they consist of a honeycomb ceramic mesh that is coated with aluminium that contains less than 2 g of precious metals such as platinum, palladium and rhodium. These precious metals are the catalysts. The mesh has a huge surface area, in excess of 20 000 m^2, yet it is smaller than an average household toaster. The catalytic converter operates at temperatures of around 450°C. At this temperature the platinum and palladium turn CO and unburnt hydrocarbons into CO_2 and H_2O. Rhodium converts oxides of nitrogen into N_2 and H_2O.

One manufacturer claims that their converter changes nitrogen into ammonia which then reacts with nitrogen oxides to produce nitrogen gas and water.

Question 10.20

Write the balanced equation for the reaction between nitrogen dioxide and ammonia.

___

___

REVIEW QUESTIONS

Chapter 10: Acids and Bases

1. The addition of ammonium chloride to water can be described by the following equation:

$$NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$$

(a) Name an acid in this reaction.
(b) Use the appropriate theory to explain why the substance chosen is acting as an acid.

2. If applicable, identify the acid and the base in each of the following reactions:

(a) $HS^-(aq) + CO_3^{2-}(aq) \rightleftharpoons S^{2-}(aq) + HCO_3^-(aq)$

(b) $HPO_4^{2-}(aq) + H_2O(l) \rightleftharpoons H_2PO^-_4(aq) + OH^-(aq)$

(c) $CH_3NH_2(aq) + CH_3COOH(aq) \rightleftharpoons CH_3NH_3^+(aq) + CH_3COO^-(aq)$

(d) $2H_2O(g) + Mg(s) \rightarrow Mg(OH)_2(s) + H_2(g)$

3. Aluminium and zinc are amphoteric metals (see note page 201).
(a) Write an equation to show zinc reacting with a base.
(b) Write an equation to show aluminium hydroxide reacting with an acid.
(c) You have been given two measuring cups of antacid, each containing 30 mL of a thick, white suspension. Explain how you could conduct a chemical test to identify which cup contained magnesium hydroxide and which contained aluminium hydroxide. Clearly state what observations you would expect to make for each suspension.

4. Write balanced chemical equations to show what occurs when the following chemicals are mixed. (Only those chemical species that react should be included in the equations.) Write ionic equations where applicable.

(a) $HCl(g) + H_2O(l)$
(b) $NH_3(g) + H_2O(l)$
(c) $HNO_3(aq) + NaOH(s)$
(d) $CH_3COOH(aq) + Na_2CO_3(s)$
(e) $H_2CO_3(aq) + KOH(aq)$
(f) $MgCO_3(s) + HCl$

5. Write balanced chemical equations to show what occurs when the following chemicals are mixed. (Only those chemical species that react should be included in the equations.) Write ionic equations where applicable.

(a) sulfuric acid solution + magnesium oxide (solid)
(b) hydrochloric acid + sodium ethanoate solution
(c) phosphoric acid solution + iron(III) hydroxide (solid)
(d) sulfur dioxide + sodium hydroxide solution
(e) a small quantity of carbon dioxide + calcium hydroxide solution
(f) an ammonia solution is mixed with vinegar solution

6. Calculate the pH of the following solutions:

(a) 0.001 mol L^{-1} HCl
(b) 1×10^{-5} mol L^{-1} HNO_3
(c) 1×10^{-3} mol L^{-1} NaOH
(d) 3×10^{-2} mol L^{-1} HCl
(e) 7.5×10^{-6} mol L^{-1} $Ca(OH)_2$

7. Copy and complete the following table by calculating the $[H^+]$ and $[OH^-]$ of the household materials listed:

MATERIAL	pH	$[H^+]$	$[OH^-]$
vinegar	3.00		
toothpaste	6.80		
oven cleaner	13.5		
window cleaner	9.75		

8. A student was given 3 test tubes. One tube contained hydrochloric acid, one contained nitric acid while the third contained ethanoic acid. The three acid solutions were of equal concentration. When several drops of universal indicator were added to each test tube, the first two went red while the third turned yellow. Explain why the indicator turned a different colour in CH_3COOH.

9. (a) Classroom acid spillage kits often contain sodium carbonate powder. Explain what makes Na_2CO_3 suitable for use in these kits.
(b) Explain the steps that could be used to clean up a spilt, concentrated sodium hydroxide solution.

10. (a) Use some of the following terms to explain the difference between the concepts, acid strength and acid concentration.

dissociate, ionise, strong electrolyte, weak electrolyte, non-electrolyte, moles, litres, volume, solubility

(b) Give an example of a solution that you have used in your chemistry laboratory that would be an example of:

(i) a dilute solution of a strong base;
(ii) a dilute solution of a weak acid;
(iii) a concentrated solution of a strong acid.

11. Give a simple test that could be used to differentiate between the following substances (give a clear statement of what observations would be expected).

(a) 1.00 mol L^{-1} solutions of barium hydroxide and ammonia.
(b) Sodium ethanoate and sodium chloride powers.
(c) 2.00 mol L^{-1} solutions of HCl and H_2SO_4.

12. Calculate the mass of carbon dioxide gas produced when 62.4 mL of a 0.250 mol L^{-1} potassium carbonate solution is added to a solution that contained an excess of phosporic acid.

Unbalanced molecular equation:

$K_2CO_3 + H_3PO_4 \rightarrow K_3PO_4 + CO_2 + H_2O$

13. Calculate the volume of 5.75 mol L^{-1} H_2SO_4 that must be used to neutralise a 7.99 L of 2.77 mol L^{-1} solution of NaOH.

14. Determine the mass of calcium chloride that would be produced if 8.24 mL of 0.330 mol L^{-1} calcium hydroxide was mixed with excess hydrochloric acid.

15. An excess of magnesium ribbon was placed in 50.0 mL of 1.75 mol L^{-1} HNO_3 solution.
   (a) Calculate the decrease in mass of the magnesium ribbon because of the reaction with the nitric acid.
   (b) Calculate the volume of hydrogen gas that would be produced if the gas was measured at STP.

16. Complete the following general equations:

(a) acid + base → ___________ + ___________

(b) acid + reactive metal → ___________ + ___________

(c) acid + metal oxide → __________ + ___________

(d) acid + carbonate → ___________ + ___________ + ___________

(e) acid + hydrogen carbonate → ___________ + ___________ + ________

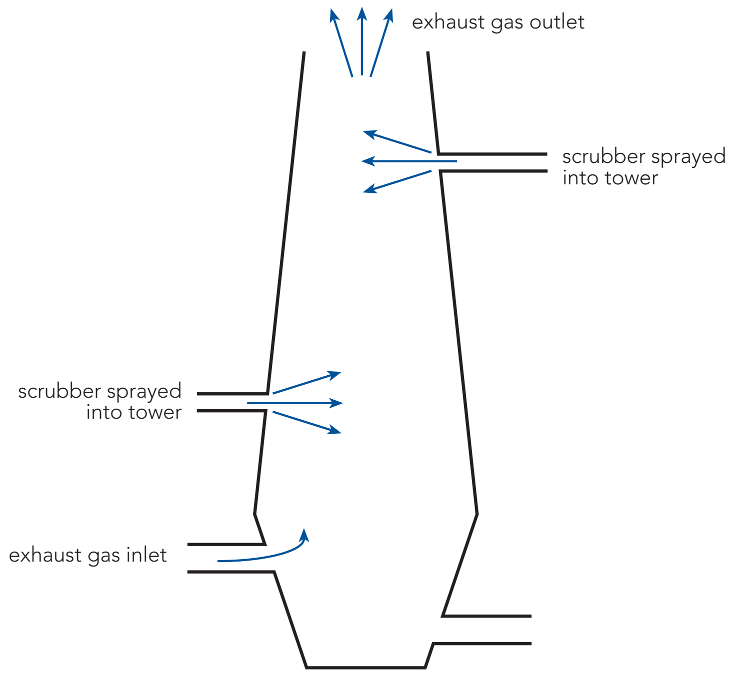
(f) (strong) acid + metal sulfite → __________ + __________ + __________

(g) (strong) acid + metal sulfide → __________ + __________

(h) non-metal oxide + base → __________ + __________

FOR THE EXPERTS

17. CO_2 and SO_2 are common exhaust gases in power generation and metal extraction processes. One method of reducing the emission of these gases is by using scrubbing towers. In these towers NaOH, Na_2CO_3 and $CaCO_3$ are mixed in water and then sprayed into the tower as the exhaust gases pass through.



Write equations for CO_2 and SO_2 dissolving in water.

18. Write equations to show how the NaOH can remove the CO_2 from the exhaust gas.

19. Write an equation to show how $CaCO_3$ and Na_2CO_3 can remove SO_2 from the exhaust gas.

11

REACTION RATES

Topics covered in this chapter:

11.1 REACTION RATE

Some chemical reactions can be very slow, such as the rusting of nails, while others, like the formation of precipitates, are extremely fast. Chemists are particularly keen to know what makes some reactions go faster than others as this would allow them to favourably control important chemical reactions.

For example, it would be an advantage to slow down reactions such as the formation of rust and speed up reactions such as the setting of glue.

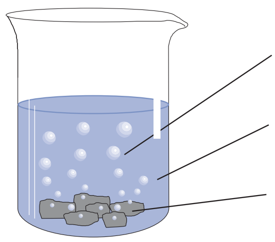
Question 11.1

List the following chemical reactions under appropriate headings:

rusting of iron, ***burning*** of H_2 gas, ***precipitate*** formation, ***fermentation*** of sugar to wine, ***setting*** of glue, ***digestion*** of food, ***neutralisation*** of NaOH with HCl, ***photosynthesis*** of sugars in plants, dynamite ***explosion***, ***weathering*** of rocks.

VERY FAST	SLOW	VERY SLOW

Figure 11.1 *Hydrochloric acid (2.0 mol L^{-1}) added to granulated zinc. When the acid is added to the zinc there is initially a fairly rapid reaction producing hydrogen gas.*



Several changes occur during a chemical reaction and any of them can be used to measure reaction rate. Consider the reaction illustrated at right where hydrochloric acid has been added to granulated zinc:

$$Zn_{(s)} + 2H^+_{(aq)} \rightarrow Zn^{2+}_{(aq)} + H_{2(g)}$$

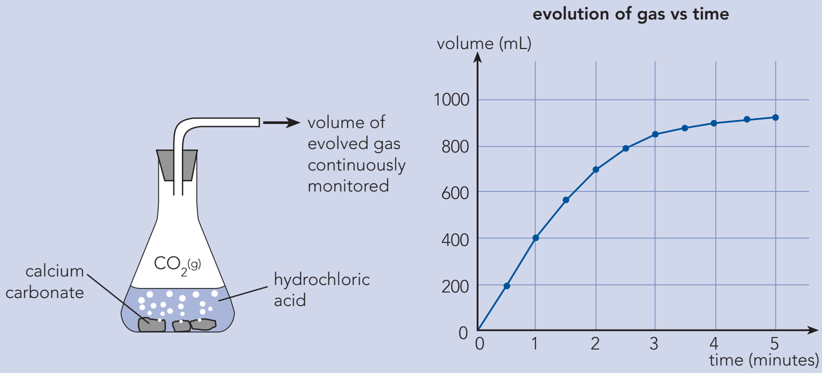
Changes that take place during the reaction include:

- mass of zinc decreases;
- the H^+ concentration decreases;
- the Zn^{2+} concentration increases;
- the volume of $H_{2(g)}$ increases. The rate of reaction in the above situation can be determined by observing the rate of any of the changes described above. In general we can say,

The rate of reaction, at some instant, is the rate at which reactants are used up or, alternatively, the rate at which products are formed.

Question 11.2

Consider the reaction illustrated below. A quantity of concentrated hydrochloric acid (6.0 mol L^{-1}) is added to calcium carbonate (marble chips) and as the reaction proceeds the volume of carbon dioxide gas produced is monitored. The results are shown graphically.



(a) Describe the reaction you would observe once the acid is added to the calcium carbonate.

___

(b) Write a balanced equation (ionic).

___

(c) In what way would you expect the reaction to differ over a period of five minutes?

___

(d) There are several measurable changes that take place during the reaction. How do the following change? (The first one is done for you.)

- the mass of $CaCO_3$ *decreases*
- the $H^+_{(aq)}$ concentration ___
- the $Ca^{2+}_{(aq)}$ concentration ___
- the volume of $CO_{2(g)}$ ___

(e) From the graph we can see that the rate at which carbon dioxide gas (CO_2) is evolved decreases over time. Suggest a reason for this.

___

(f) Use the graph to estimate how much greater the reaction rate is during the first minute of the reaction when compared with the third minute of the reaction.

___

(g) Suggest three factors that would have influenced the reaction rate of this reaction.

(i)___ (ii)___ (iii)___

Factors affecting reaction rate

The rate of a chemical reaction is affected by the:

- nature of the reactants
- concentration of the reactants
- state of subdivision of the reactants
- temperature
- presence of catalyst.

In order to explain why these five factors affect reaction rate it is best to look at what happens at a molecular level and collision theory.

11.2 COLLISION THEORY

Simply writing the equation for a reaction **does not** explain what happens at a molecular (or particle) level. Consider the combustion of hydrogen gas as illustrated below:

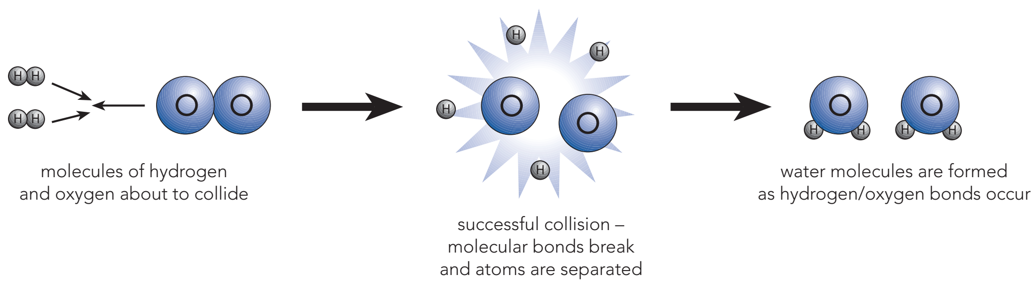


Figure 11.2 *The combustion of hydrogen gas.*

We can see that in order for a reaction to take place there must be successful collisions between reacting particles. The success of any collision is improved if:

- the reacting particles have an appropriate collision orientation
- the reacting particles collide with sufficient energy.

Ultimately the greater the rate of successful collisions, the greater the reaction rate.

Question 11.3

In the reaction illustrated above (Figure 11.2):

(a) how many bonds break? __________

List these. ______________________

(b) how many bonds form? __________

List these. ______________________

Question 11.4

What is meant by a successful collision between reacting particles?

____

____

11.3 THE NATURE OF REACTANTS

Reaction rate depends on the nature of the reactants involved. In general,

- **ionic reactions** are rapid as they do not involve the breaking of bonds or electron transfer.

e.g. 1. $Ba^{2+}{}_{(aq)} + SO_4{}^{2-}{}_{(aq)} \rightarrow BaSO_{4(s)}$

2. $Ag^{+}{}_{(aq)} + Cl^{-}{}_{(aq)} \rightarrow AgCl_{(s)}$

Both reactions are almost instantaneous as no bonds have to break and there is a strong electrostatic attraction between oppositely charged ions.

- **molecular reactions** are slow since they involve bond breaking and bond rearrangement. Collisions are often unsuccessful at room temperature as there is insufficient activation energy.

e.g. 3. $C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$

4. $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

Both of these reactions are very slow (almost undetectable) at room temperature as they involve the breaking of strong covalent bonds.

Question 11.5

Which of the molecular reactions (e.g. 3 or e.g. 4) is likely to be the slowest? Why?

____

11.4 THE CONCENTRATION OF REACTANTS

Reaction rate is affected by the concentration of reactants in either the **solution** or **gaseous** phase. Collision theory tells us that for a reaction to take place the reacting particles must collide. Hence if the number of particles per unit volume is increased the number of collisions will also increase. Typically, if the pressure of two reacting gases is doubled then the reaction rate between them is increased fourfold. This occurs since the number of collisions increases by a factor of four.

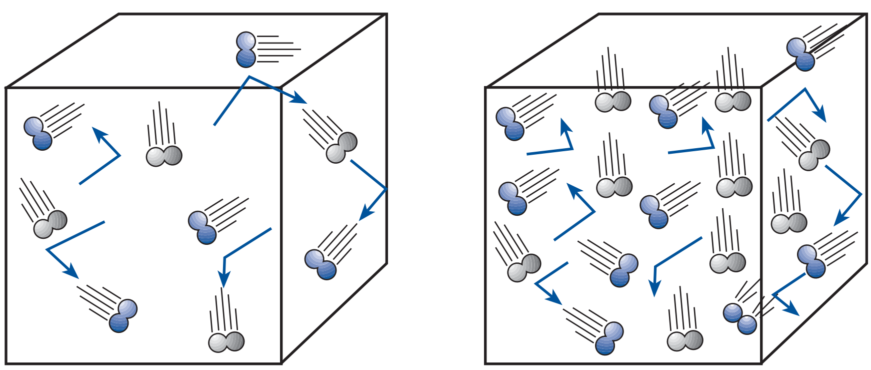


Figure 11.3 *The number of collisions per second depends upon the concentration of each of the reactant particles. An increase in the rate of collisions means an increase in the rate of successful collisions and hence the rate of reaction.*

11.5 THE STATE OF SUBDIVISION OF REACTANTS

A lump of sugar dissolves far more slowly in your coffee than does fine grained sugar. This observation can be readily explained in terms of collision theory. The fine grained sugar has a greater surface area than the solid lump of sugar and hence there is an increased frequency of collisions with water molecules. This results in an increased reaction rate.

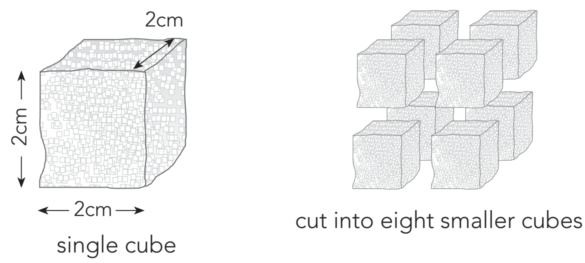


***Figure 11.4** The state of subdivision increases the surface area of a reactant. See question 11.7.*

11.6 TEMPERATURE

Reaction rate is greatly affected by temperature. As temperature increases so does the velocity of the reacting particles. This has a twofold effect on reaction rate.

- The number of collisions per unit time increases. This increases the number of successful collisions and hence reaction rate.
- The force, and therefore likely success, of each collision increases. More importantly, the proportion of particles with sufficient activation energy increases with temperature and hence the reaction rate increases markedly. (See kinetic energy distribution graph – fig. 11.7).

Question 11.6

In the experiment illustrated in Figure 11.1, the reaction between the acid and the zinc was fastest at the beginning. Use collision theory to explain why.

____

____

Question 11.7

In Figure 11.4 a solid cube measuring 2 cm × 2 cm × 2 cm is cut into 8 smaller and equal cubes.

(a) Calculate the surface area of the large cube. ____________________

(b) Calculate the total surface area of the eight smaller cubes.

____

(c) Comment.

____

____

11.7 CATALYSTS

The presence of some substances called catalysts helps to speed up a reaction. These substances are themselves not consumed.

The actual mechanism by which catalysts work is often complex and not always understood. However, it is evident that catalysts provide an alternate and easier reaction pathway. They effectively lower the activation energy. Catalysts are involved in many important chemical reactions such as the following:

e.g.

1. Platinum in the H_2/O_2 fuel cell. $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$
2. Iron in the Haber process. $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$
3. Vanadium pentoxide in the Contact Process. $2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$

Question 11.8

Catalysts are usually in the form of powders or fine wire mesh. Why is this?

____

____

____

____

Nanoparticles are being used increasingly as catalysts because they have a very large surface area to volume ratio (reactions occur at the surface of reactants because it is the surface areas that "meet" on collision) and they can be custom made to suit a specific purpose. These particles range in size from 1 to 100 nanometres (1×10^{-7} to 1×10^{-9}m).

Nanoparticle catalysts can be divided into four groups depending on composition:

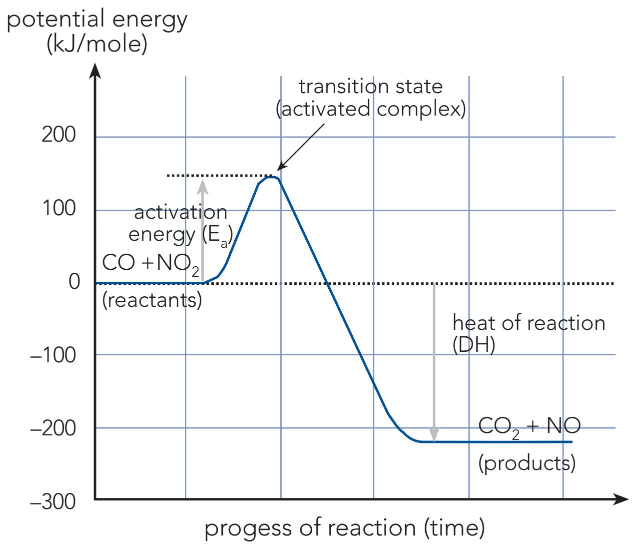
- Platinum or palladium based
- Gold based
- Based on other metals such as Cu and Rh
- Nanoparticles dispersed in poymers especially as carbon nanotubes
- Nanoparticle catalysts are used in vehicle catalytic converters, fuel cells, reduction of nitrogen oxide pollutants, petroleum refining, production of biofuels and in food processing.

Enzymes are biological catalysts that target very specific chemical reactions in the body and enable rapid reaction rates at the relatively low temperature of the human body. Generally they are proteins. Examples of enzymes:

- **Urease:** the first enzyme to be crystallised, responsible for the break down of urea into NH_3 and CO_2 in simple living things.
- **Pepsin:** produced in the stomach and enables the digestion of proteins
- **Catalase:** break down of H_2O_2 into O_2 and H_2O
- **Lactase:** break down of the lactose in milk
- **Lipases:** responsible for the break down of lipids (fats and oils)
- **Sucrases:** break down of sucrose to the simpler sugars, fructose and glucose
- **Amylase:** changes starches (a carbohydrate) into complex sugars which are then broken down by sucrases.

11.8 POTENTIAL ENERGY DIAGRAMS

These diagrams can be used to represent the energy changes that occur during a chemical reaction. The potential energy diagram for a typical exothermic reaction is illustrated below.



- **Activation energy** (E_a)**:** This is the energy that colliding particles must have in order to form an activated complex. Reactions requiring a high activation energy do not proceed easily (or at all) at room temperature. This explains why, for example, a mixture of hydrogen and oxygen in a test tube will not react unless exposed to a flame.
- **Transition state:** Reactant particles which collide with sufficient energy form an activated complex and are in a transition state. The activated complex is unstable and quickly breaks down to either form products or re-form the reactants.
- **Heat of reaction** (ΔH)**:** This is the difference in potential energy (enthalpy) between the reactants and products.

Figure 11.5 *Energy profile diagram for the reaction* $CO_{(g)} + NO_{2(g)} \rightarrow CO_{2(g)} + NO_{(g)} + 225$ *kJ. Note that reactants go through a transition state (or activated complex) before forming products. Activation energy* (E_a) ≈ *135 kJ mol*$^{-1}$*. Heat of reaction (ΔH) ≈ -225 kJ mol*$^{-1}$ *(exothermic).*

11.9 ACTIVATION ENERGY – EFFECT OF CATALYSTS

Since a catalyst essentially provides an alternate, and easier pathway for a reaction, it effectively lowers the activation energy. This is true for both the forward and reverse reaction.

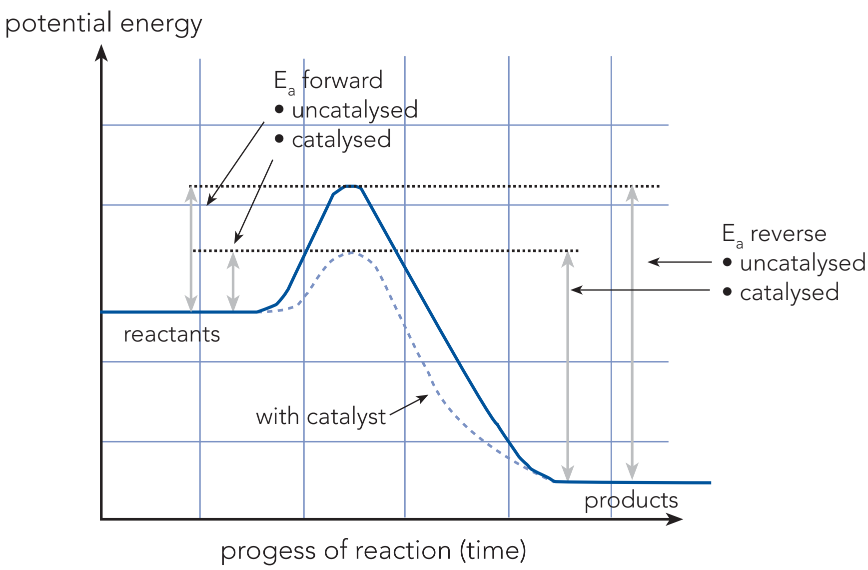
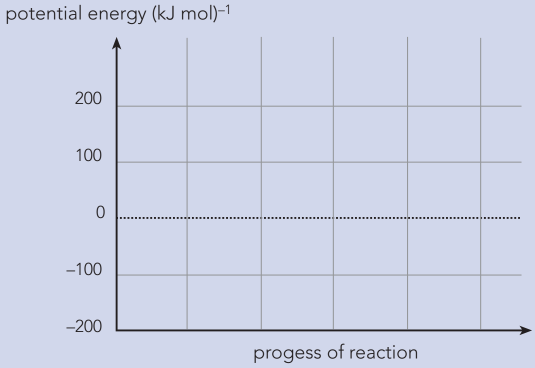


Figure 11.6 *The effect of catalyst on reaction pathway. The activation energy is lowered for both forward and reverse reactions.*

Question 11.9

Ethene can be produced by heating ethane in the presence of a catalyst. The equation can be represented as follows:

$$C_2H_{6(g)} \rightarrow C_2H_{4(g)} + H_{2(g)} \qquad \Delta H = +120 \text{ kJ mol}^{-1}$$



On the axes shown:

(a) draw an energy profile diagram for the uncatalysed reaction if the activation energy is 180 kJ mol^{-1}.

(b) show a likely catalysed pathway for the reaction on the same diagram.

Question 11.10

Draw the likely shape of the potential energy diagrams for the following cases:

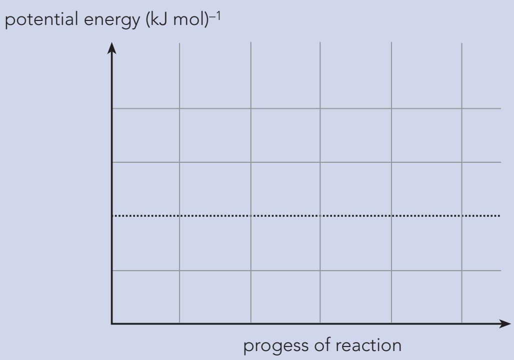
(a) White phosphorus ignites spontaneously when exposed to air on a warm day.

(b) Hydrogen gas can burn in air fiercely and produces a lot of heat. However a flame is required to start it.

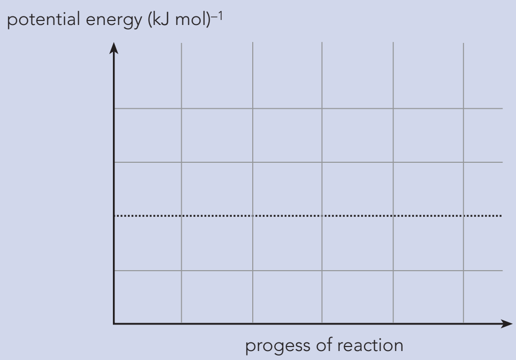
(c) $2H_2O_{2(l)} \rightarrow 2H_2O_{(l)} + O_{2(g)}$ $\qquad E_a = 75$ kJ , $\Delta H = -98$ kJ

(d) $NO_{2(g)} + O_{2(g)} \rightarrow NO_{(g)} + O_{3(g)}$ $\qquad E_a = 210$ kJ , $\Delta H = +200$ kJ

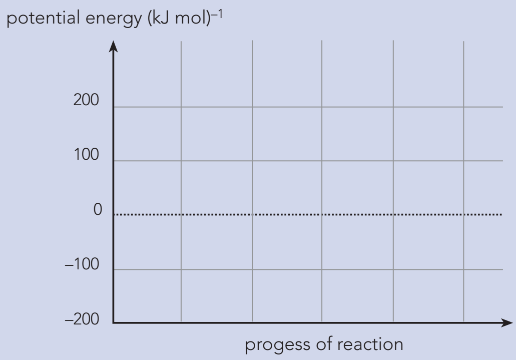
(a)



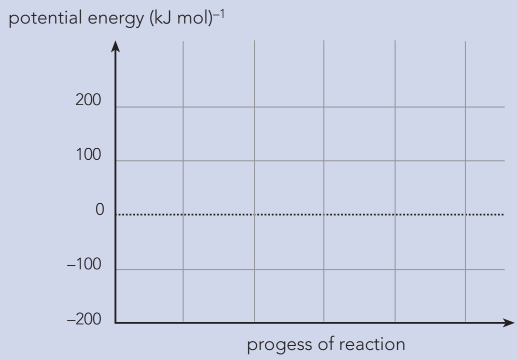
(b)



(c)



(d)



11.10 COLLISION THEORY – KINETIC ENERGY DISTRIBUTION

The temperature of any substance is related to the average kinetic energy of all its particles. However not all the particles (at any particular instant) are travelling at the same speed. The diagram over the page shows the typical distribution of molecular kinetic energies at two different temperatures. If we consider a reaction with a fairly high activation energy then very few particles have sufficient energy to undergo a reaction (shaded area). A small change in temperature however, can markedly increase the proportion of particles with energy greater than activation energy (E_a) as can be seen by a large increase in total area to the right of E_a.

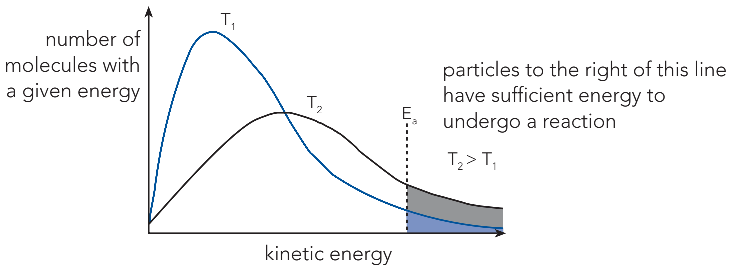


Figure 11.7 Kinetic energy distribution of particles at different temperatures.

Question 11.11

(a) What is the significance of the areas under the graphs in Figure 11.7? (Hint: the areas are equal.)

____

(b) What does the heavily shaded portion to the right of E_a represent in Figure 11.7?

____

(c) An increase in temperature of only 10°C can markedly increase reaction rate. Use the graph in Figure 11.7 to explain the **main** reason for this.

____

____

____

REVIEW QUESTIONS

Chapter 11: Reaction Rates

1. In order to produce H_2 gas a strip of $Al_{(s)}$ was placed into 500.0 mL of a 3.50 mol L^{-1} solution of caustic soda.

$$2Al_{(s)} + 2NaOH_{(aq)} + 6H_2O_{(l)} \rightarrow 2Na[Al(OH_4]_{(aq)} + 3H_{2(g)}$$

State what effect **each** of the following changes will have on the **rate** of this reaction and briefly explain why the change occurs.

(a) Another 100 mL of 3.50 mol L^{-1} NaOH was added as the aluminium was placed into the original caustic soda solution.
(b) The aluminium strip was cut into smaller pieces.
(c) The reaction vessel was placed into a trough of cold water.
(d) 25.0 mL of 5.00 mol L^{-1} NaOH was added to the mixture.

2. The following are typical reactions that can be carried out in a laboratory:

(i) $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

(ii) $C_6H_{12}O_{6(aq)} + 6O_{2(g)} \rightarrow 6CO_{2(g)} + 6H_2O_{(l)}$

(iii) $Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}$

Which of these reactions would you expect to be:

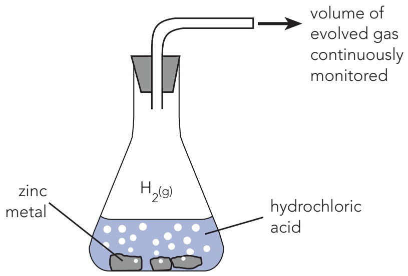
(a) the most rapid? Explain.

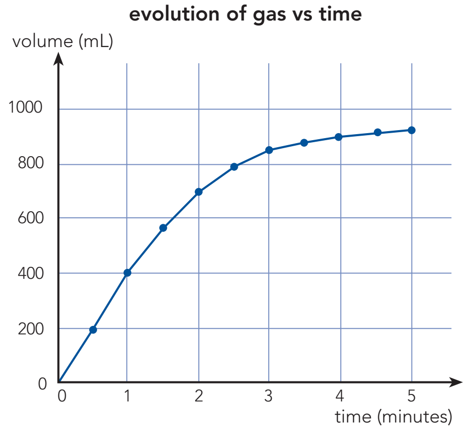
(b) the slowest? Explain.

3. For gases and solutions, an increase in the concentration of reactants leads to an increase in the rate of reaction.

Explain this observation fully in terms of **collision theory**.

4. A quantity of concentrated hydrochloric acid (6.0 mol L^{-1}) is added to zinc granules and as the reaction proceeds the volume of hydrogen gas produced is monitored. The results are shown graphically:





(a) Writee a balanced ionic equation for the reaction.

(b) The reaction rate between zinc and hydrochloric acid was much greater at first. Explain.

(c) Estimate the time it would have taken to produce 400 mL of H_2 gas if the concentration of the HCl acid used had been 2.0 mol L^{-1}.

(d) Suggest three factors that would have influenced the reaction rate of this reaction as it proceeded.

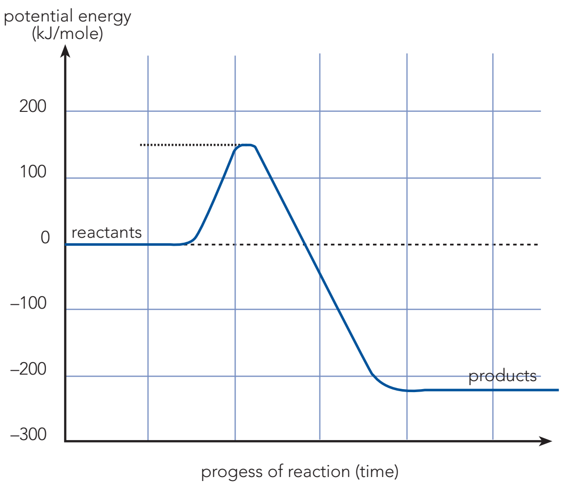
5. Sugar cubes added to your coffee dissolve much more slowly than sugar granules. Explain clearly why this occurs.

6. Explain the following observations in terms of collision theory:
   (a) Explosions sometimes occur in flour mills due to flour dust.
   (b) A mixture of hydrogen and oxygen gas in a test tube at room temperature will only explode if a flame is applied.
   (c) Foods such as meats and cheese are refrigerated to prevent them from spoiling.
   (d) A fire will burn more brightly if it is fanned by a breeze.

7. Reaction rate is greatly affected by temperature. Explain this observation fully in terms of **collision theory**. (Hint: An increase in temperature increases the velocity and kinetic energy of reacting particles.)

8. The presence of some substances called catalysts helps to speed up chemical reactions. These substances are themselves not consumed but simply provide an alternate and easier pathway for the reaction. Name the catalysts involved in:
   (a) the decomposition of H_2O_2
   (b) catalytic converters in cars
   (c) biochemical reactions (general name only).

9. Use the diagram below to answer the following.



   (a) Is this an exothermic or endothermic reaction?
   (b) What is the activation energy for the forward reaction?
   (c) What is the heat of reaction for the reverse reaction?
   (d) What effect would a catalyst have on:
      (i) the heat of reaction;
      (ii) the activation energy for the reverse reaction?

10. Explain the meaning of the following terms:

(a) activation energy

(b) transition state

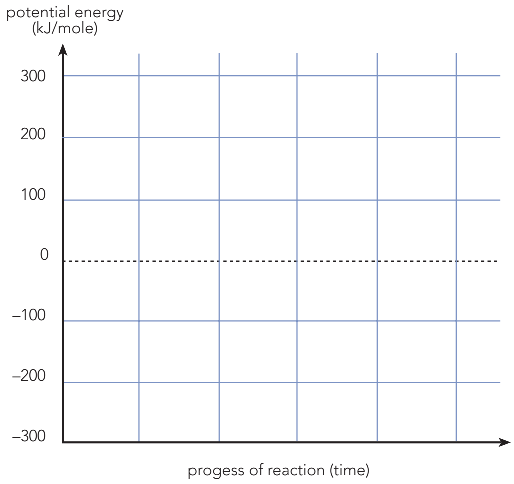
(c) heat of reaction.

11. (a) On the axes shown draw a potential energy diagram for the following reaction:

$A + B \rightarrow C + D$

$\Delta H = 150$ kJ

$E_a = 200$ kJ



(b) Show a possible catalysed pathway on the diagram and label it.

(c) For your catalysed pathway determine:

(i) ΔH = ____________________________

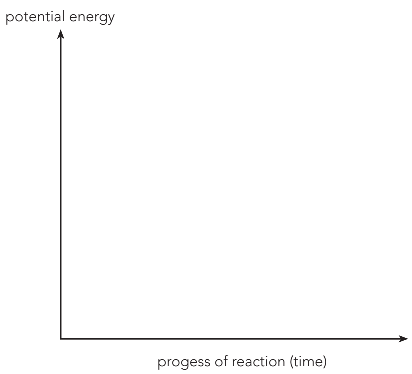
(ii) E_a = ____________________________

12. Draw the likely shape of the potential energy diagrams for the following cases:

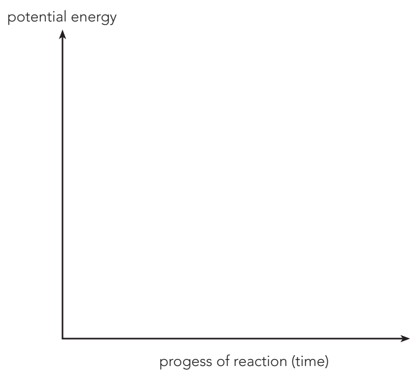
(a) $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$ $E_a = 320$ kJ , $\Delta H = -890$ kJ

(b) $C_2H_{6(g)} \rightarrow C_2H_{4(g)} + H_{2(g)}$ $E_a = 180$ kJ , $\Delta H = +120$ kJ

(a)



(b)



FOR THE EXPERTS

The combustion of hydrocarbons to produce energy is essential for the current lifestyle enjoyed by people living in developed countries. Whether those hydrocarbons be in the form of wood, coal, petroleum products or natural gas, their combustion provides the energy that is essential for a vast range of human endeavours. It is very difficult to think of any activity we carry out in our daily life that at some stage has not involved the energy produced by the combustion of hydrocarbons.

The burning of hydrocarbons, however, comes with a range of environmental costs. One of these is the production of large quantities of toxic carbon monoxide and nitrogen oxides by motor vehicles. Carbon monoxide is produced when hydrocarbons are burnt in a limited supply of oxygen and nitrogen oxides are produced when normally inert nitrogen gas is mixed with oxygen gas at high temperatures, both conditions occurring in a car engine.

13.

(a) Write the equations for the production of CO and NO by the engine of a motor vehicle.

To reduce levels of emission of these toxic pollutants car exhaust systems are fitted with catalytic converters. These catalytic converters commonly contain minute particles of Pd and Rh coated onto the surface of a ceramic support that has a honeycomb structure. The Pd is responsible for catalyzing the exothermic reaction between CO and O_2 to produce CO_2 whereas the Rh catalyst speeds up the reaction between the pollutants CO and NO to produce CO_2 and N_2.

(b) Write the equations for the reactions catalyzed by Pd and Rh.

(c) Explain why Pd and Rh are used in the catalytic converter.

(d) Explain why the Pd and Rh are coated onto the surface of the ceramic support material and why the structure has a honeycombed shape.

(e) From a chemical point of view, explain why the catalytic converter in a motor vehicle should last the lifetime of the vehicle.

TRIAL TEST 1: ATOMIC STRUCTURE

Time allowed:	**70 minutes**	**Section 1 –** Multiple Choice	20 marks
Total marks:	**80**	**Section 2 –** Short & Extended Answer	60 marks

SECTION 1 – MULTIPLE CHOICE (20 MARKS)

1. The mass number of an atom is:
   (a) equal to the number of particles in the nucleus of that atom.
   (b) unique and identifies what element the atom is.
   (c) always greater than the mass number of the element that immediately precedes it in the periodic table.
   (d) the same as the mass number of all other isotopes of that atom.

2. A species has 4 neutrons, 2 electrons and 3 protons. This species could be:
   (a) An ion of an isotope of beryllium.
   (b) A neutral atom of an isotope of boron.
   (c) An ion of an isotope of lithium.
   (d) An ion of an isotope of carbon.

3. Which of the following is the correct electron configuration for the ion Al^{3+}?

   (a) 2,3.
   (b) 2,5.
   (c) 2,8.
   (d) 2,8,3.

4. Which of the following has the greatest number of electrons?

   (a) $^{14}_{7}N^{3-}$ (b) $^{19}_{9}F^{-}$ (c) $^{27}_{13}Al^{3+}$ (d) $^{23}_{11}Na$

5. Which of the following are isotopes of the same element?

Atom P:	36 protons	38 neutrons	37 electrons
Atom Q:	37 protons	38 neutrons	36 electrons
Atom R:	A = 76	Z = 38	
Atom S:	A = 80	its neutral atom contains 37 electrons	

   (a) Q and S
   (b) Q and R
   (c) P and S
   (d) R and S

6. Which of the following are listed in increasing order of ionisation energy?
   (a) Li, Na, K
   (b) F, O, N
   (c) Ne, Ar, Kr
   (d) B, C, N

7. An atoms electron is excited from its ground state (n=1) to the third energy level (n=3). As the electron returns to its ground state it gives up energy in the form of photons. How many different photon energies are possible in this case?
   (a) 2
   (b) 3
   (c) 4
   (d) 6

8. A mass spectrometer was used to determine the relative concentration of the three isotopes of an element X. The mass spectrum for the element contained three distinctive peaks with the following approximate heights; ^{24}X was 8.0 cm, ^{25}X was 1.0 cm and ^{26}X also 1.0 cm. From this information it can be concluded that the relative atomic mass (A_r) of element X is closest to:
   (a) 24.0
   (b) 24.5
   (c) 25.0
   (d) 26

9. Flame tests are carried out by placing small samples of metal salts in a Bunsen flame and observing the colours produced. Which of the following is correct regarding flame tests?
   (a) The sample is placed in the flame using clean copper wire.
   (b) Sodium salts burn with a bright green flame.
   (c) The colours are due to electron transitions in the atomised sample.
   (d) Flame tests can be used to identify the majority of metal ions.

10. Which of the following is not a feature of the atomic absorption spectrometer (AAS)?
   (a) A bright Bunsen flame.
   (b) A hollow cathode lamp.
   (c) A source of pulsed light.
   (d) A prism and monochromator.

SECTION 2 – SHORT AND EXTENDED ANSWER (60 MARKS)

Answer each question in the space provided.

11. (a) One of the isotopes of chlorine is symbolised as $^{37}_{17}Cl$. This indicates that for this isotope:

   (i) the mass number is ________________

   (ii) the atomic number is ________________

   (iii) the number of neutrons in the nucleus is ________________

   (iv) the number of electrons in a neutral atom would be ____________

   (b) The electron configuration for sodium can be written as 2,8,1. Similarly, give the electron configuration for each of the following:

   (i) a nitrogen atom ____

   (ii) a magnesium ion ____

   (iii) a sulfide ion ____

   (iv) a potassium atom ____

[8 marks]

12. (a) What are valence electrons?

____

(b) State the number of valence electrons that each of the following atoms have.

(i) Be _______ (ii) Ar________ (iii) Na ________ (iv) Al ________

[6 marks]

13. The structure of the periodic table is closely linked to the electron configuration of the elements and is helpful in predicting their likely properties.

(a) What is common to the valence electrons of all elements in the third period of the periodic table?

____

(b) What is common about the number of valence electrons of all alkali elements?

____

(c) Why do elements belonging to the same group tend to have very similar chemical properties?

____

____

(d) Briefly explain why Group 18 elements, the noble or inert gases, are all unreactive.

____

____

(e) What link do chemists make between the electron structure of the inert gases and the electron interactions of all other elements?

____

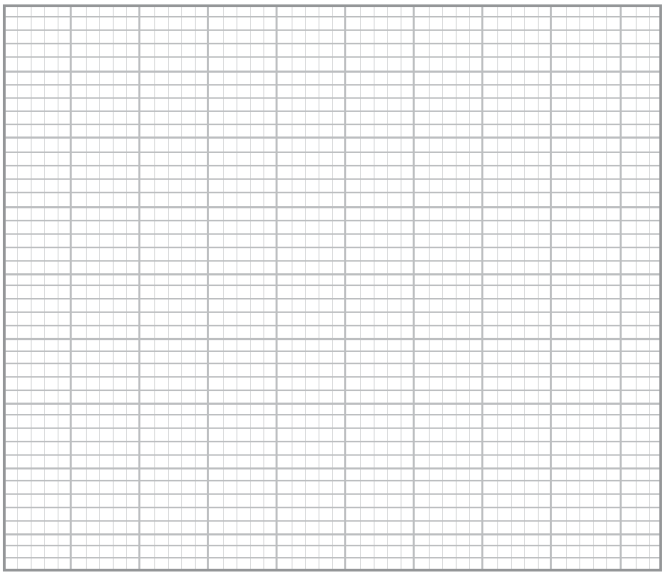
____

[10 marks]

14. Data determined from the mass spectrum for lead is shown below.

ISOTOPE	ISOTOPIC MASS	ABUNDANCE (%)
^{204}Pb	203.97	1.4
^{206}Pb	205.97	24.1
^{207}Pb	206.98	22.1
^{208}Pb	207.98	52.4

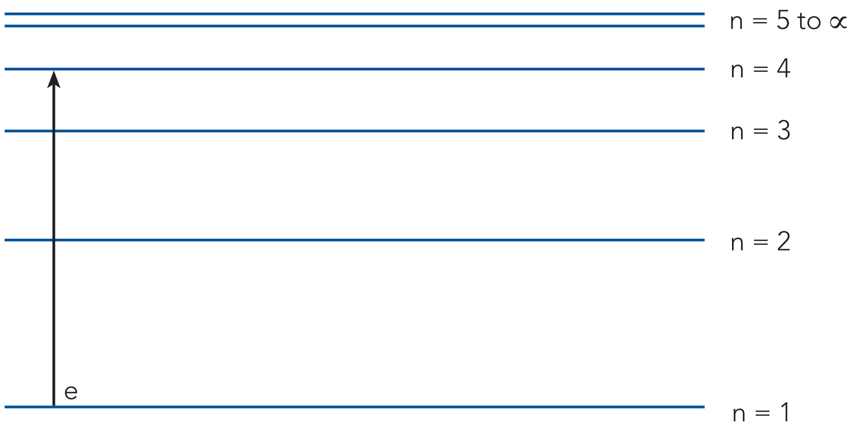
(a) Sketch the mass spectrum you would expect for these isotopes of lead on the blank graph grid below.



(b) Calculate the relative atomic mass (Ar) for the element lead.

[12 marks]

15. The electron of an atom has moved from its ground state (n = 1) to an excited state (n = 4) as shown in the diagram below.



(a) Has energy being absorbed or emitted by the atom?

(b) The excited electron can return to its ground state by one or several different steps. Indicate each of the possible transitions on the diagram above.

(c) Which transition would involve the largest wavelength photon?

(d) For this atom the transitions n = 4 to n = 2 and n = 3 to n = 2 both result in photons of visible light. If the two colours are red and green which transition would have resulted in the green?

[12 marks]

16. The atomic absorption spectrometer (AAS) developed in Australia in the 1950s, is a very useful and sensitive instrument for the analysis of small traces of metal contaminants. Briefly explain each of the following.

(a) Why the sample being analysed is placed in an intense flame.

(b) Why the light from the cathode lamp is passed through the flame.

(c) How the detector can distinguish light from the lamp from that emitted by the flame.

(d) What measurement indicates the presence, if any, and concentration of a target metal in the sample.

[12 marks]

END OF TEST (80 MARKS)

TRIAL TEST 2:
MATERIALS AND BONDING

Time allowed:	**60 minutes**	**Section 1 –** Multiple Choice	20 marks
Total marks:	**80**	**Section 2 –** Short & Extended Answer	60 marks

SECTION 1 – MULTIPLE CHOICE (20 MARKS)

1. Which of the following contains only pure substances?
   (a) water, zinc, copper (II) sulfate, iron (II) oxide
   (b) neon, sodium chloride solution, sugar solution, argon
   (c) gold, magnesium oxide, petrol, chlorine
   (d) copper, aluminium oxide, oxygen, air

2. Which of the following is INCORRECT?
   (a) Elements and compounds are pure substances.
   (b) Mixtures can be separated by physical means.
   (c) Mixtures may be homogeneous or heterogeneous.
   (d) Sublimation refers to the phase change of a solid to a liquid.

3. In order to separate and recover sugar crystals from sand the order of the different steps is most likely to be:
   (a) dissolution, evaporation, filtration, crystallisation
   (b) dissolution, filtration, evaporation, crystallisation
   (c) dissolution, distillation, crystallisation, evaporation
   (d) filtration, evaporation, crystallisation.

4. Nanoparticles are defined as being those between:
   (a) 1.0 nm and 1.0 μm
   (b) 1.0×10^{-9} m and 1.0×10^{-10} m
   (c) 1.0×10^{-8} m and 1.0×10^{-9} m
   (d) 1.0×10^{-7} m and 1.0×10^{-9} m

5. Nanoparticles exhibit markedly different properties to those of their larger counterparts in bulk materials. The most likely reason for this is:
   (a) they are all made up of covalently bonded carbon atoms
   (b) they all have a large surface area to volume ratio
   (c) they all have a very low aspect ratio
   (d) they are all smaller than the wavelength of visible light.

6. Which of the following statements is INCORRECT for an ionic solid?
   (a) The substance will have a high melting point because of the strong electrostatic attraction between oppositely charged ions.
   (b) When heated sufficiently charged particles can move more freely and allow the passage of an electric current through the substance.
   (c) When the ions in the crystal lattice are forced to move, electrostatic repulsion tends to make the solid shatter.
   (d) When dissolved in water, the ionic lattice breaks up and makes electrons available to allow the passage of an electric current through the solution.

7. Which of the following statements is correct?
   (a) Atoms from group 17 on the periodic table share electrons with atoms from group 1 to form a covalent bond.
   (b) Atoms with nearly empty outer energy levels tend to share valence electrons so as to get a full outer energy level.
   (c) Substances with covalent bonds have more electrons and are harder to melt.
   (d) Atoms with nearly full outer energy levels can share valence electrons with other similar atoms to attain a full outer energy level.

8. Which of the following statements about metals is INCORRECT?
   (a) Metals are malleable.
   (b) With few exceptions, metals are silvery grey in colour.
   (c) Due to the strength of the metallic bond, ions in the metallic lattice are very close giving all metals high densities (when compared to water).
   (d) An ionic substance can be formed by reacting a metallic substance with a non-metallic substance.

9. Which of the following statements is true about the allotropes of carbon?
   (a) All are good conductors due to delocalised electrons.
   (b) All contain covalently bonded carbon atoms.
   (c) All their carbon atoms are bonded to three other carbon atoms.
   (d) All their structures are fairly similar.

10. Two elements X and Y have the following electron configurations:

    X: 2 , 8 , 2   Y: 2 , 7

    Which of the following is the most likely result if the two elements combine?

   (a) An ionic substance with the formula XY_2
   (b) An ionic substance with the formula X_2Y
   (c) A covalent substance with the formula XY_2
   (d) A covalent substance with the formula X_2Y_7

SECTION 2 – SHORT AND EXTENDED ANSWER (60 MARKS)

Answer each question in the space provided.

11. (a) Classify each of the following as either element, compound or mixture:
sugar , steam , air , oxygen , ethanol , vinegar , copper (II) oxide , potassium , limestone

element: ______________________________

compound: ______________________________

mixture: ______________________________

(b) Which of the above are:

(i) pure substances? ______________________________

(ii) heterogeneous mixtures? ______________________________

(iii) cannot be separated into simpler substances? ______________________________

[6 marks]

12. Pure substances can be separated from mixtures by using physical separation techniques. Separation is possible due to the substances having different properties. Complete the table by naming the technique for each example and the property/s involved.

SUBSTANCE TO BE RECOVERED	SEPARATION TECHNIQUE/S	PROPERTY/S MAKING SEPARATION POSSIBLE
salt from a salt/sand/water mixture		
water from a salt/water mixture		
$KCl_{(s)}$ from a $KCl_{(aq)}$ + $CuCl_{2(aq)}$ mixture		
alcohol from a alcohol/ water mixture		
plant pigment from plant colouring		

[10 marks]

13. For each of the following name the nanoparticle/s used and briefly explain their effect.

(a) Sunscreen lotion ______________________________

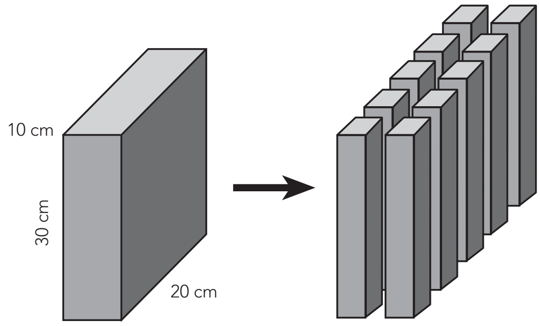
____

(b) Antibacterial bandages ______________________________

____

[6 marks]

14. A rectangular block of firewood is shown below. It measures approximately 30 cm high, 20 cm wide and 10cm thick. The block is split into ten equal pieces of 30 cm by 5.0 cm by 4.0 cm.



Wood block 30 cm x 20 cm x 10 cm   Ten pieces each 30 cm x 5 cm x 4 cm

(a) Assuming no loss of wood when it is split determine,

(i) The initial surface area of the block

____

____

(ii) The final total surface area of the ten pieces

____

____

(iii) The factor by which the surface area changed

____

____

(b) It was found that the small pieces of wood burn much more quickly than a large block. In terms of your results above explain clearly why this is the case.

____

____

[10 marks]

15. Covalent bonds form readily between atoms that have almost full valence shells. An example is the formation of a fluorine molecule from two fluorine atoms.

(a) Use electron dot diagrams to illustrate and explain how this formation occurs.

(b) Explain why a bond forms between the two atoms.

(c) Use the nature of this type of bond to explain why molecular solids are poor conductors of electricity.

(d) The type of bond that would form between fluorine and sodium atoms would be of quite a different nature. What type of bond would form and give a reason?

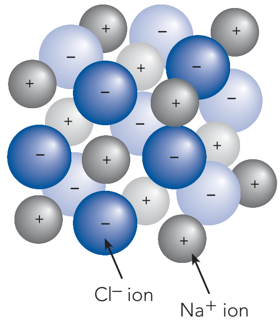
[12 marks]

16. Classify the following substances by placing them into their correct classification in the table below: *ammonia, sulfur dioxide, silicon dioxide, calcium oxide, potassium, lithium bromide, water, diamond.*

METALLIC	IONIC	COVALENT MOLECULAR	COVALENT NETWORK

[4 marks]

17. A solid ionic compound consists of an infinite array of oppositely charged ions. A typical crystal lattice, such as that of sodium chloride is shown. With the aid of diagrams and electron dot diagrams where necessary clearly explain:



(a) the formation of an ionic bond between sodium and chlorine atoms.

(b) why ionic solids are hard and brittle.

(c) why ionic solids do not conduct electricity.

[12 marks]

END OF TEST (80 MARKS)

TRIAL TEST 3:
CARBON CHEMISTRY

Time allowed: **70 minutes** | **Section 1 –** Multiple Choice | 20 marks
Total marks: **80** | **Section 2 –** Short & Extended Answer | 60 marks

SECTION 1 – MULTIPLE CHOICE (20 MARKS)

1. Which of the following contains only saturated hydrocarbons?
   (a) $CH_3CHCHCH_3$, CH_3CH_2OH, C_2H_4
   (b) C_2H_4, C_2H_2, C_2H_6
   (c) CH_3CH_3, C_2H_6, $CH_3CH_2CH_2CH_2CH_3$
   (d) C_7H_{14}, $CH_3CH_2CH_3$, C_2H_6

2. Carbon has the electron configuration of 2, 4 and a single carbon atom is able to form:
   (a) 3 single covalent bonds with hydrogen atoms and 1 single covalent bond with oxygen
   (b) 3 single covalent bonds with chlorine atoms and 1 double covalent bond with sulfur
   (c) 2 single covalent bonds with hydrogen atoms and one double covalent bond with an iodine atom
   (d) 1 single covalent bond with a hydrogen atom and a triple covalent bond with a sulfur atom

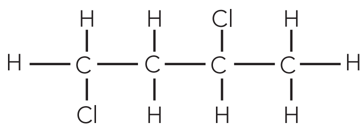
3. Which of the following names is acceptable by IUPAC rules?

   (a) pent-3-ene (b) trans-1,1-dichloroethene
   (c) 1, 5-dibromocyclopentyne (d) cis-2,3-dibromohex-2-ene

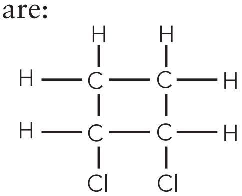
4. What reactants could be used to produce 2-chlorobutane?

   (a) Cl_2 and cyclobutene (b) Cl_2 and $CH_3CHCH_2CH_3$
   (c) Cl_2 and $CH_3CH_2CH_2CH_3$ (d) Cl_2 and C_4H_6

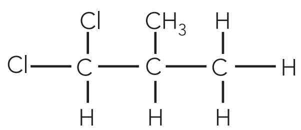
5. The structural formulae for 4 compounds are:



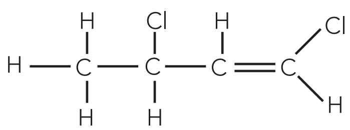
I



II



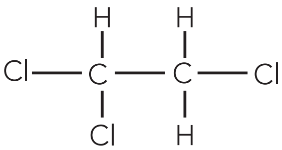
III

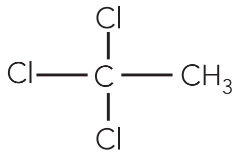


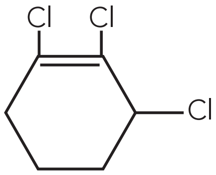
IV

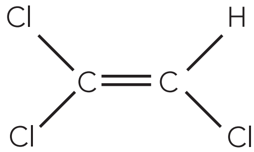
Which of the structures are isomers?
(a) I and II (b) I and III (c) II, III and IV (d) III and IV

6. 1,1,2 – trichloroethene is a common organic solvent. Which one of the following structures corresponds to this solvent?

(a) 

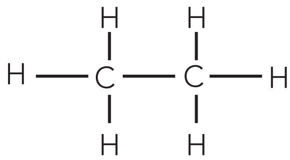
(b) 

(c) 

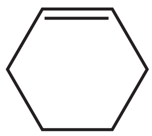
(d) 

7. Which of the following compounds readily undergo addition reactions?

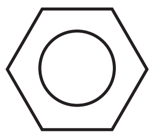
I



II



III



IV

(a) I only
(b) II and IV only
(c) I and III only
(d) I, III and IV only

8. When $CH_3CH_2CH_2CH_2Cl$ burns in excess oxygen:
(a) For every molecule of oxygen consumed, a molecule of carbon dioxide is produced
(b) Carbon dioxide and water are the (only) products
(c) Carbon monoxide is produced because of the presence of the Cl
(d) An exothermic reaction occurs with at least three products formed.

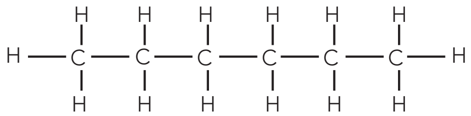
9. When propyne reacts with bromine, a possible product is:
(a) HBr
(b) $CH_3CHBrCH_2Br$
(c) 2,3-dibromopropane
(d) $CHBr_2CBr_2CH_3$

10. Which one of the following formulae could represent two substances which are cis-trans geometric isomers?
(a) $(CH_3)_2C{=}C(CH_3)_2$
(b) $CH_3CH{=}CHCH_3$
(c) CH_2CH_2
(d) $CH{\equiv}CH$

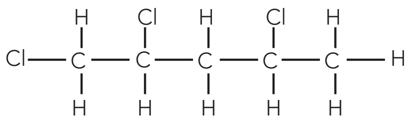
SECTION 2 – SHORT AND EXTENDED ANSWER (60 MARKS)

Answer each question in the space provided.

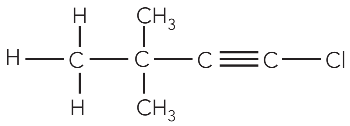
11. Use the IUPAC rules to name the following compounds:



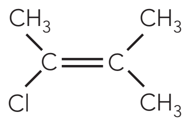
(a) ______________________________



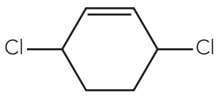
(b) ______________________________



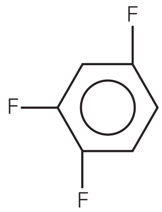
(c) ______________________________



(d) ______________________________



(e) ______________________________



(f) ______________________________

[12 marks]

12. Draw the structural formula for each of the following:

(a) 1,3,5 – triiodohept-3-ene

(b) 1,4-dimethylcyclohexene

(c) 1,1-difluorobut-1-ene

(d) methylbenzene (toluene)

(e) hexachlorobenzene

(f) 1,7,7,7-tetrabromo-3,4-diethylhept-1-yne

[12 marks]

13. Write the balanced equation and give observations when the following reactants are mixed:

(a) cyclohexane + bromine water in the presence of a suitable catalyst

equation ______________________________

observation ______________________________

(b) cyclohexene + iodine water

equation ______________________________

observation ______________________________

(c) a spark is put to a mixture of heptane and oxygen gas

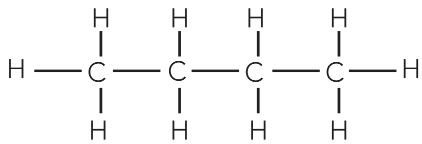
equation ______________________________

observation ______________________________

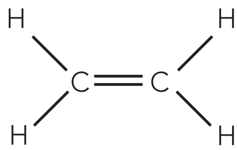
[9 marks]

14. Complete and balance the following equations:

(a)

 $+ Br_2 \xrightarrow{\text{catalyst}}$

(b)

 $+ HCl \longrightarrow$

(c) $C_6H_6 + F_2 \rightarrow$

(d) $CH_3CCCH_3 + I_2 \rightarrow$

(e) $C_3H_6 + O_2 \rightarrow$

[10 marks]

15. Draw the structural formula and name 5 structural isomers with the formula C_5H_{10}

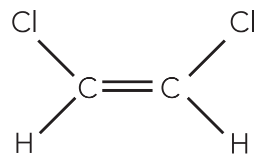
(a) ______________________ (b) ______________________

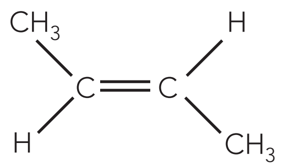
(c) ______________________ (d) ______________________

(e) ______________________

[10 marks]

16. (a) Use IUPAC rules to name the following compounds:





(i) ______________________ (ii) ______________________

(b) Draw the geometric isomers of $C_3H_4I_2$

(i) (ii)

(c) Draw a structural isomer of $C_3H_4I_2$ that is **not** a geometric isomer.

[7 marks]

END OF TEST (80 MARKS)

TRIAL TEST 4: CHEMICAL REACTIONS AND ENERGY CHANGE

Time allowed:	**70 minutes**	**Section 1 –** Multiple Choice	20 marks
Total marks:	**80**	**Section 2 –** Short & Extended Answer	60 marks

SECTION 1 – MULTIPLE CHOICE (20 MARKS)

1. Which of the following involves a chemical change?
   (a) Freezing pure water
   (b) Boiling an egg
   (c) Melting wax
   (d) Chopping firewood.

2. Which of the following chemical equations is correctly balanced?
   (a) $H_{2(g)} + O_{2(g)} \rightarrow H_2O_{(g)}$
   (b) $Fe_2O_{3(s)} + CO_{(g)} \rightarrow 2Fe_{(g)} + CO_{2(s)}$
   (c) $Ca(OH)_{2(s)} \rightarrow CaO_{(s)} + H_2O_{(g)}$
   (d) $H_2O_{2(l)} \rightarrow H_2O_{(l)} + O_{2(g)}$

3. Exothermic reactions always result in:
   (a) more products than reactants.
   (b) a decrease in enthalpy.
   (c) a decrease in temperature of the surroundings.
   (d) a decrease in the activation energy barrier.

4. Which of the following reactions is endothermic?

   (a) $H_{(g)} + H_{(g)} \rightarrow H_{2(g)}$

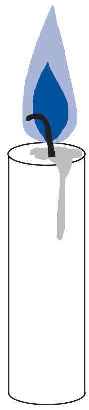
   (b) $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

   (c) $H_2O_{(l)} \rightarrow H_2O_{(g)}$

   (d) $C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$

5. Which of the following statements is FALSE?
   (a) Within a substance, chemical energy exists as kinetic energy and potential energy.
   (b) Potential energy in a substance is due to the forces of attraction and repulsion between charged particles.
   (c) Molecules can store potential energy through translational motion.
   (d) Molecules can store kinetic energy through rotational and vibrational motion.

6. When you light a candle a pool of liquid wax forms near the wick and a flame is produced as shown. It would be true to say that:
   (a) the wax burning results in a decrease in enthalpy value
   (b) only chemical reactions are taking place
   (c) heat is given off so the overall ΔH value for the process is positive
   (d) the wax melting is an exothermic reaction.



7. The ΔH for a particular reaction has a positive value. This would indicate that:
   (a) heat is given up to the surroundings during the reaction
   (b) the reaction is endothermic
   (c) the ΔH value can be included on the right hand side of a balanced equation
   (d) there are more bonds created than are formed.

8. Which of the following are all non-renewable sources of energy?
   (a) Natural gas, fuel oil, hydroelectric power
   (b) Coal, biomass to produce ethanol, oil
   (c) Crude oil, wood, natural gas
   (d) Natural gas, anthracite coal, bottled gas (LPG).

9. Worldwide, which of the following is the most widely used resource for power generation?
   (a) Coal
   (b) Natural gas
   (c) Nuclear energy
   (d) Solar energy and tidal power.

10. The two hydrocarbon fuels, methane and butane, can be combusted as shown below. Their heats of reaction at 25°C and 1 atm are also shown.

Methane $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$ ΔH = - 891 kJ mol^{-1}

Butane $2C_4H_{10(g)} + 13O_{2(g)} \rightarrow 8CO_{2(g)} + 10H_2O_{(l)}$ ΔH = - 2878 kJ mol^{-1}

From this information it would be true to say:
   (a) $n(CO_2)/n(\text{fuel})$ for butane is equal to 8
   (b) $n(CO_2)$/kJ of energy is greatest for methane
   (c) energy/n(fuel) is approximately 1440 kJ for butane
   (d) energy/kg(fuel) is approximately 55 MJ for methane.

SECTION 2 – SHORT AND EXTENDED ANSWER (60 MARKS)

Answer each question in the space provided.

11. Indicate whether each of the following are exothermic or endothermic reactions.

(a) $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ Answer: ______________________

(b) $I_{2(g)} \rightarrow I_{2(s)}$ Answer: ______________________

(c) $O_{2(g)} \rightarrow 2O_{(g)}$ Answer: ______________________

(d) $NH_4NO_{3(s)} \rightarrow NH_4NO_{3\,(aq)}$ Answer: ______________________

[4 marks]

12. The combustion of hydrogen gas can be represented as follows:

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)} + \text{heat}$$

(a) Which bonds are broken in this reaction? List them.

____

[2 marks]

(b) Which bonds are formed in this reaction? List them.

____

[2 marks]

(c) Which process involved more energy; bond breaking or bond forming? Explain.

____

[3 marks]

(d) Which bonds are likely to be the strongest, those of the reactants or those of the products? Explain.

____

[3 marks]

13. Bottled gas, such as that used for BBQs, is compressed propane gas (C_3H_8). When propane burns in air it not only produces heat but also carbon dioxide gas and water vapour.

(a) Is the burning of propane a physical or chemical change? Explain.

____

____

(b) Write a word equation for the reaction.

(c) Write a balanced chemical equation for the reaction.

(d) Petrol burnt in cars is mostly octane (C_8H_{18}) and burns in a similar way to propane to produce CO_2 and H_2O.

Write a balanced chemical equation for this.

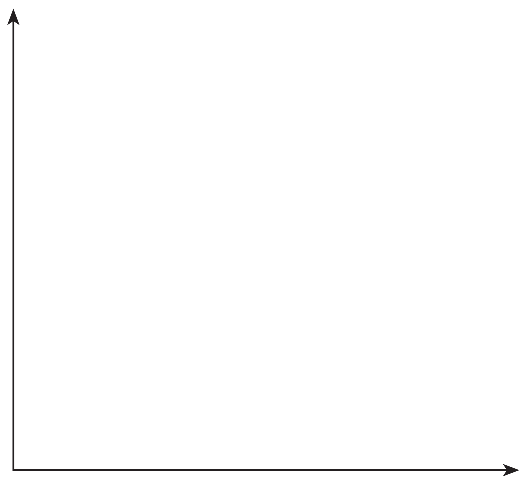
[10 marks]

14.

(a) Ethanol (C_2H_6O) is used increasingly as a fuel for motor cars. Its heat of combustion at 25 °C and 1 atm is given as $\Delta H = -1367\ kJ\ mol^{-1}$. Write a balanced equation for the combustion of ethanol including the heat of formation.

[2 marks]

(b) Using the axes below, draw an enthalpy diagram for the reaction.
Fully label the axes and indicate the reactants, products and the change in ΔH.



[8 marks]

(c) Calculate the energy produced by burning one kilogram of ethanol.

[4 marks]

(d) Calculate the mass of carbon dioxide produced in this case.

[4 marks]

15. Three important fuels are methane, propane and octane. They are, respectively, the main components of natural gas, LPG, and gasoline. They are listed in the table below together with their heats of combustion, ΔH. Complete the following in order to compare these three fuels.

(a) Write balanced equations for the combustion of these fuels including the heat of formation.

[6 marks]

(b) From your equations determine how much energy is produced per kg of fuel in each case. Also determine the ratios of moles of CO_2 per mole of fuel and moles of CO_2 per MJ of energy. Use your values to complete the table.

FUEL	ΔH kJ mol^{-1} 25 °C , 1 atm.	MJ of energy/kg of fuel	$n(CO_2)/n(fuel)$	$n(CO_2)$/MJ of energy
Methane	-891			
Propane	-2219			
Octane	-5471			

[9 marks]

(c) Natural gas is widely used for home heating and some power generation. Use your table of results to suggest reasons for this.

[2 marks]

(d) Suggest a reason why natural gas is not used as bottled gas or as a motor fuel.

[1 mark]

END OF TEST (80 MARKS)

TRIAL TEST 5: CHEMISTRY CALCULATIONS

Time allowed:	**70 minutes**	**Section 1 –** Multiple Choice	20 marks
Total marks:	**80**	**Section 2 –** Short & Extended Answer	60 marks

SECTION 1 – MULTIPLE CHOICE (20 MARKS)

1. The relative atomic mass (A_r)of helium is 4.00. This information indicates that:
   (a) the mass of a helium atom would be 4.00 g
   (b) the mass of a helium atom would be exactly one third of that of a carbon-12 atom
   (c) the mass of a helium atom would be exactly 4 times that of a carbon-12 atom
   (d) the mass of a helium and carbon atoms cannot be compared.

2. The formula for carbon dioxide gas is CO_2. This means that for a molecule of this gas:

   (a) there is twice as much oxygen by mass as there is carbon

   (b) the percentage of carbon by mass is 33.3%
   (c) the percentage of carbon by mass is given by $\frac{12.01}{44.01} \times \frac{100}{1}$

   (d) the percentage of oxygen by mass is given by $\frac{16.0}{44.01} \times \frac{100}{1}$

3. Which of the following contains the greatest number of particles?

   (a) 6.022×10^{23} atoms of zinc
   (b) 0.500 moles of water molecules
   (c) 4.00 g of hydrogen gas
   (d) 22.47 L of oxygen gas

4. An amount of 2.0 mol of CH_3COOH molecules would contain:

   (a) 2.0 mol of C atoms in total
   (b) 3.0 mol of H atoms in total
   (c) 4.0 mol of O atoms in total
   (d) 8.0 mol of atoms in total

5. Consider the following word equation:

   aluminium sulfate + calcium hydroxide → aluminium hydroxide + calcium sulfate

   If it was converted to symbol form and then balanced, the coefficient in front of calcium hydroxide would be:

   (a) 1
   (b) 2
   (c) 3
   (d) 4

6. During the reaction between iron and oxygen, 32 g of iron was converted to iron (III) oxide.

$Fe + O_2 \rightarrow Fe_2O_3$ *(not balanced)*

The mass of oxygen consumed would be:

(a) $\frac{32}{55.85} \times \frac{4}{1} \times \frac{32}{3}$ g

(b) $\frac{32}{55.85} \times \frac{32}{3}$ g

(c) $\frac{32 \times 4}{55.85} \times \frac{32 \times 3}{1}$ g

(d) $\frac{32}{55.85} \times \frac{32}{1} \times \frac{3}{4}$ g

7. Copper reacts with nitric acid according to the following equation:

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

The mass of which product would be increased by the greatest percentage if the mass of copper consumed was doubled?

(a) $Cu(NO_3)_2$
(b) NO
(c) H_2O
(d) none of the above

8. For the reaction between a potassium hydroxide solution and a magnesium iodide solution, which of the following is INCORRECT?

(a) The mass of hydroxide ions in the solution must remain constant to ensure the Law of Conservation of Mass is obeyed.
(b) The mass of the magnesium hydroxide produced will be less than combined mass of the reactants.
(c) The equation representing this reaction is best written as:
$Mg^{2+}{}_{(aq)} + 2OH^-{}_{(aq)} \rightarrow Mg(OH)_{2(s)}$
(d) The addition of some $HCl_{(aq)}$ would reduce the quantity of precipitate on the bottom of the reaction vessel.

9. The production of nitrogen gas can be represented by the following equation:

$$4NH_{3(g)} + 3O_{2(g)} \rightarrow 2N_{2(g)} + 6H_2O_{(l)}$$

If 26.7 g of ammonia reacts with excess O_2, the volume of N_2 produced at STP (100 kPa and 0.0°C) would be:

(a) 1.59 L (b) 17.8 L (c) 35.6 L (d) 178 L

10. Ammonia can be produced from nitrogen and hydrogen as indicated by the following equation:

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

If 3.75 L of hydrogen are completely reacted with nitrogen in this way, what volume of ammonia would be produced? Assume all volumes are at STP.

(a) 1.25 L (b) 2.50 L (c) 3.75 L (d) 7.50 L

SECTION 2 – SHORT AND EXTENDED ANSWER (60 MARKS)

Answer each question in the space provided.

11. (a) The relative atomic mass (A_r) for helium is 4.00 and for oxygen is 16.0. What specifically does this information tell us about the atoms of these elements?

(b) Calculate the relative molecular mass for sulfuric acid and hydrochloric acid.

(c) Which of the acids above has the greatest percentage by mass of hydrogen? What is this percentage?

[8 marks]

12. For each of the following determine the number of:

(a) moles of gold atoms in 6.022×10^{25} atoms of gold

(b) lead atoms in 0.25 mol of lead

(c) molecules of water in 2.5 mol of water

(d) molecules of sulfur dioxide in 64.0 g of this gas

(e) moles of chloride ions in 250 mL of 0.0200 mol L^{-1} sodium chloride solution

[10 marks]

13. Calculate the volume of hydrogen chloride gas produced when 23.5 L of ammonia gas at STP (100.0 kPa and 0.0°C) reacts with excess chlorine gas. The hydrogen chloride gas is collected and stored at STP conditions.

$$2NH_{3\,(g)} + 3Cl_{2(g)} \rightarrow N_{2(g)} + 6HCl_{(g)}$$

[8 marks]

14. Concentrated sulfuric acid can be used to oxidise hydrogen sulfide gas to produce pure sulfur and water.

(a) Write the balanced equation for this reaction.

(b) Calculate the mass of hydrogen sulfide gas used to recover 37.6 g of sulfur when only 97.6% of the sulfur is recovered.

[8 marks]

15. Excess sodium sulfate solution is added to 150 mL of 2.0 mol L^{-1} barium chloride solution.

(a) Write the balanced equation for this reaction.

(b) Assuming all available barium ions have reacted determine the mass of precipitate formed.

(c) If the final volume of the two combined solutions is 400 mL, determine the final chloride ion concentration.

[10 marks]

16. (a) Methane gas (CH_4) and propane gas (C_3H_8) are commonly used fuels. Give a balanced equation for the combustion of each fuel with oxygen gas. Assume only CO_2 and H_2O are produced in each case.

(b) Determine which fuel has the highest percentage by mass of carbon.

(c) Determine which fuel produces the greatest volume of gaseous products (at say STP) per gram of fuel burnt.

[16 marks]

END OF TEST (80 MARKS)

TRIAL TEST 6: INTERMOLECULAR FORCES

Time allowed:	**70 minutes**	**Section 1 –** Multiple Choice	20 marks
Total marks:	**80**	**Section 2 –** Short & Extended Answer	60 marks

SECTION 1 – MULTIPLE CHOICE (20 MARKS)

1. Which of the following compounds would contain polar covalent bonds?
   (a) NaCl
   (b) S_8
   (c) Fe_2O_3
   (d) NH_3

2. Choose the list below which shows the bonds arranged in order of increasing polarity.
   (a) H-H, H-N, H-F
   (b) N-H, P-H, As-H
   (c) H-Cl, H-C, H-Cl
   (d) Cl-Cl, H-C, H-Cl

3. Which of the following lists contains molecules that are likely to be the most soluble in water?
   (a) HCl, C_3H_8, N_2
   (b) Cl_2, H_2, CCl_4
   (c) H_2S, NH_3, HF
   (d) HBr, C_6H_{14}, F_2

4. Which of the following statements about chromatography is incorrect?
   (a) Chromatography can be used to separate and identify very small quantities of pollutants in river water.
   (b) In gas chromatography the molecules in the mixture are separated according to the strength of intermolecular forces between and the different molecules in the stationary phase.
   (c) In high performance liquid chromatography the molecules in the mobile phase must be strongly attracted to molecules in the stationary phase to ensure complete separation and identification of the molecules in the mobile phase.
   (d) Two reasons why thin layer chromatography is preferred over paper chromatography are sensitivity and destruction of cellulose by corrosive agents.

5. Which of the following statements is most correct?
   (a) $BeCl_2$ cannot exist as a molecule because Be does not have 8 electrons in its outermost shell after it has bonded with two chlorine atoms.
   (b) NH_3 is a polar molecule that is pyramidal in shape as it has one lone pair of electrons.
   (c) CH_4 is a polar molecule because carbon and hydrogen have different electronegativities.
   (d) NH_3 is likely to be a less soluble in hexane than it is in water because hexane is a non-polar molecule.

6. Choose the row in the table below that correctly lists examples of the main intermolecular force exhibited by the covalent molecular substances listed.

	Hydrogen Bonding	**Dipole-Dipole**	**Dispersion**
(a)	H_2Te	NH_3	F_2
(b)	H_2S	HF	HCl
(c)	H_2O	HCl	C_2H_6
(d)	H_2O	CCl_4	CH_3CH_3

7. Which of the following covalent molecular compounds is both a polar molecule and tetrahedral in shape?
(a) CH_3Cl
(b) CCl_4
(c) H_2O
(d) CH_2CH_2

8. Ammonia has a boiling point of – 33 °C while phosphine (PH_3) has a boiling point of -88 °C. The difference between these boiling points is most likely due to the fact that:
(a) ammonia molecules form hydrogen bonds whereas phosphine molecules do not.
(b) the dispersion forces between phosphine molecules are stronger than those between ammonia molecules.
(c) the dispersion forces between ammonia molecules are stronger than those between phosphine molecules.
(d) phosphine molecules form hydrogen bonds whereas ammonia molecules do not.

9. Choose the group that is different to the other three.

(a)	CCl_4	H_3O^+	H_2
(b)	CH_4	NH_4^+	H_2S
(c)	NH_4^+	NH_3	HCl
(d)	CF_4	PCl_3	BeF_2

10. Choose the group that has the pure substances correctly listed in order of increasing melting point.

(a)	NH_3	CH_4	SiO_2	W
(b)	Cl_2	CO_2	Na	C_7H_{16}
(c)	CCl_4	$CaCO_3$	$Al_2(CO_3)_3$	H_2CO_3
(d)	NH_3	C_7H_{16}	Na	SiO_2

SECTION 2 – SHORT AND EXTENDED ANSWER (60 MARKS)

Answer each question in the space provided.

11. Name a substance whose properties match those described.

(a) A highly polar molecular substance that is very soluble in water. Individual molecules are described as having a pyramidal shape.

___

(b) Combines with another element sharing one pair of valence electrons. When combined with hydrogen it forms molecules that exhibit hydrogen bonding and are acidic.

___

(c) Has a very high first ionisation energy, a very low boiling point and does not form bonds with the two elements that have atomic numbers either one smaller or one greater than its own.

___

[6 marks]

12. Draw the electron dot diagrams (Lewis structures) for the following:

(a) Cl_2 (b) NH_3

(c) CO_2 (d) NO_3^-

(e) SO_4^{2-} (f) C_2H_4

[12 marks]

13. Complete the following table by stating the shape and polarity of each molecule

NAME	SHAPE	POLAR OR NON-POLAR?
chloroform ($CHCl_3$)		
hydrogen sulfide		
phosphorus trichloride		
dichlorine monoxide		

[8 marks]

14. Explain why H_2O has a **higher** boiling point than H_2S but H_2S has a **lower** boiling point than H_2Se.

[6 marks]

15. Explain the following trend in boiling points: $CH_4 < C_2H_6 < C_3H_8$

[6 marks]

16. Explain why polar molecular compounds are often soluble in water but non-polar molecular compounds are rarely soluble in water. Include a discussion on the intermolecular processes involved in dissolving.

[6 marks]

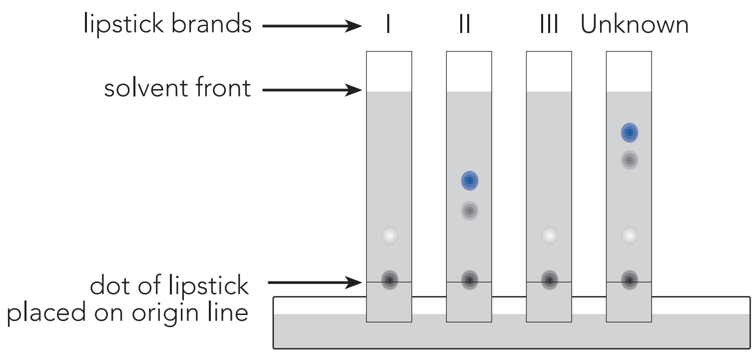
17. (a) Explain why CO is a polar molecule while CO_2 is a non-polar molecule.

[4 marks]

(b) At a temperature of -100° C and a pressure of 100 kPa, CO is a gas and CO_2 is a solid. In terms of intermolecular forces, suggest a reason why the non-polar CO_2 is a solid while the polar CO is a gas.

[2 marks]

18. A student was given the task to analyse three brands of lipstick to see which one matched an unknown brand. The student conducted a paper chromatography analysis as summarised below.



The mobile phase used was a 50% ethanol/50% acetone mixture.

(a) Explain why the origin line for placing the dot of lipstick was drawn in pencil and not pen.

[2 marks]

(b) Suggest why the mobile phase was an ethanol/acetone mixture rather that water.

[2 marks]

(c) Explain why the different colours (compounds) travel to different heights up the chromatography paper.

[2 marks]

(d) Which of the lipstick brands is the unknown brand likely to be? Give a brief justification for your answer.

[2 marks]

(e) In the forensic analysis of these lipsticks, HPLC is preferred over gas chromatography. What does this suggest about the boiling points of the coloured compounds being tested.

[2 marks]

END OF TEST (80 MARKS)

TRIAL TEST 7: KINETIC THEORY, GASES

Time allowed:	**70 minutes**	**Section 1 –** Multiple Choice	20 marks
Total marks:	**80**	**Section 2 –** Short & Extended Answer	60 marks

SECTION 1 – MULTIPLE CHOICE (20 MARKS)

1. Which one of the following statements about gases is FALSE?
   (a) The volume of gas particles is negligible compared to the total volume taken up by the gas.
   (b) Collisions between gas molecules are perfectly elastic and there is no overall loss of kinetic energy.
   (c) As the pressure of a gas is increased, the average kinetic energy of the gas molecules also increases.
   (d) There is a lot of space between gas molecules.

2. A real gas will differ from an ideal gas when:
   (a) The pressure has been increased so that the gas molecules become compressed and take up less space.
   (b) The temperature has been increased so that molecular collisions will have sufficient energy to allow intermolecular forces to exist.
   (c) Temperature and pressure conditions are increased to alter the average kinetic energy of the gas molecules.
   (d) The temperature of the gas has been decreased to the extent that intermolecular forces are experienced between molecules.

3. When a fully inflated and tied party balloon is placed in the refrigerator:
   (a) its volume will decrease because the gas molecules collide with each other less often.
   (b) its volume will decrease because the gas molecules on average have less kinetic energy.
   (c) its volume will decrease because the rubber molecules move more slowly and are now held in fixed positions.
   (d) its volume will decrease because the air inside the refrigerator is colder and so exerts less pressure on the balloon.

4. Dry air is made up of mostly nitrogen (~78%), oxygen (~21%) and argon (~1%). A particular sample of dry air is at 20.0°C. It would be true to say that:
   (a) the molecules of all of these gases, on average, have the same velocity.
   (b) the molecules of all these gases, on average, have the same kinetic energy.
   (c) at this temperature the oxygen molecules, on average, would be travelling the fastest.
   (d) argon molecules would have the greatest velocity as there are less of them.

5. The vapour pressure of liquids can vary greatly due to different factors. Which of the following does NOT affect vapour pressure?
   (a) Atmospheric pressure above the liquid.
   (b) The temperature of the liquid.
   (c) The intermolecular forces within the liquid.
   (d) The presence of dissolved solids in the liquid.

6. A small inflated party balloon had its volume reduced by placing some books on top of it. Assuming there was no change in the temperature of the air in the balloon it would be correct to say that:
   (a) the pressure in the balloon would have increased due to the reduced volume.
   (b) the pressure in the balloon would not have changed because there was no change in temperature.
   (c) the pressure in the balloon would have decreased as the air particles are closer together.
   (d) the air particles in the balloon would be moving about more rapidly.

7. Which of the following statements about the behaviour of gas is FALSE?
   (a) A gas that had its temperature doubled from 30˚C to 60˚C would experience a doubling of its volume if pressure was kept constant.
   (b) The pressure of a fixed mass of gas would treble if its volume was reduced from 9.0 L to 3.0 L at a constant temperature.
   (c) 32.0 g of oxygen gas at STP would occupy the same volume as 44.01 g of carbon dioxide gas at STP.
   (d) A gas that had its temperature decreased would exert less pressure if its volume was kept constant.

8. What volume would the contents of a 4.00 kg cylinder of compressed oxygen occupy if measured at STP? (The 4.00 kg refers to the mass of the oxygen in the cylinder).

   (a) $\frac{4000}{32.00} \times 22.71$

   (b) 4.00×22.71

   (c) $\frac{22.71}{4.00} \times 32.00$

   (d) $\frac{22.71}{4000} \times 32.00$

9. Which of the following statements about a 6.00 g gas sample at STP is correct.
   (a) The volume occupied by the gas is dependent on the temperature and pressure conditions and not on the chemical composition of the gas.
   (b) If the mass of the gas was increased to 12.00 g the pressure would double if temperature and volume were kept constant.
   (c) 6.00 g of oxygen gas would occupy a greater volume than 6.00 g of nitrogen gas at STP.
   (d) If the gas had a volume of 3.41 L it would be argon.

10. What volume of oxygen gas, measured at STP, would be required to completely react with 3.56 g of hydrogen gas?

    The unbalanced equation for the reaction is:

    $H_2 + O_2 \rightarrow H_2O$

    (a) 3.53 L
    (b) 80.0 L
    (c) 20.1 L
    (d) 40.0 L

SECTION 2 – SHORT AND EXTENDED ANSWER (60 MARKS)

Answer each question in the space provided beneath the question. Show all working where calculations are involved.

11. Use the kinetic theory to explain why:

(a) gases are easily compressed.

(b) you can smell perfume from an opened bottle even though it is some distance away from you.

(c) the pressure in a balloon increases if we pump air into it.

(d) the pressure in a car tyre increases on a hot day

[12 marks]

12. (a) A small canister of compressed gas is used to inflate a fishing balloon. The balloons volume changes from 0.50 L to 12.0 L with a negligible change in temperature. Use the kinetic theory to explain how the volume of the balloon has changed without a temperature change.

(b) An identical canister is used to inflate a bicycle tyre but the very rigid rubber of the tyre limits any change to volume. Explain the change that must have occurred inside the tyre in terms of the kinetic theory.

[8 marks]

13. (a) Describe an experiment that could be used to predict the temperature at which the volume of an ideal gas would be zero (i.e. absolute zero). The description needs to be a brief description (or labelled diagram) of the equipment used and what variables need to be measured.

[6 marks]

(b) Draw a qualitative graph (no values required) of the results of the experiment and explain how the graph can be used to predict the value for absolute zero (labels on the axes need to be included).

[6 marks]

(c) According to the kinetic theory, what is the significance of absolute zero?

____

____

____

[4 marks]

14. Atmospheric pressure becomes less at higher altitudes as shown below. Use the data to help you answer the following:

3000 m	70.1 kPa
2000 m	79.5 kPa
1000 m	89.9 kPa
Sea level	101.3 kPa

(a) A helium filled party balloon is accidentally released into the air and rises into the atmosphere. Assuming there is little change in temperature predict what will happen to the size of the balloon as it rises and use kinetic theory to explain why.

____

(b) In actual fact the temperature of the air is also lower at higher altitudes. Use the kinetic theory to explain how this will affect your prediction in (a) and why.

____

____

(c) Some mountain climbers at about 2000 m above sea level have set up camp and are about to do some cooking. To their surprise they found that eggs boiled in an open pot of boiling water were not cooked.

(i) Explain this observation in terms of your understanding of vapour pressure and boiling point.

____

____

(ii) Suggest how the climbers may be able to boil eggs successfully.

____

____

[12 marks]

15. (a) Convert the following temperatures as indicated.

i) 115 K = __________ ˚C

ii) 293˚C = __________ K

iii) -54˚C = __________ K

(b) What volume would a 2.95 g same of methane gas (CH_4) occupy at STP?

(c) A school classroom is 10.0 m long, 7.00 m wide with 2.80 m high ceilings. If air is 21% oxygen gas, calculate the mass of oxygen gas in the room at STP.

[12 marks]

END OF TEST (80 MARKS)

TRIAL TEST 8:
SOLUTIONS AND SOLUBILITY

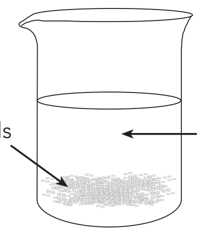
Time allowed:	**70 minutes**	**Section 1 –** Multiple Choice	20 marks
Total marks:	**80**	**Section 2 –** Short & Extended Answer	60 marks

SECTION 1 – MULTIPLE CHOICE (20 MARKS)

1. A solution may best be described as:
   (a) a pure substance of constant composition.
   (b) a homogeneous mixture of uniform composition.
   (c) a substance which can be purified by filtration.
   (d) a heterogeneous mixture of variable composition.

2. Gases most easily dissolve in water at:
   (a) low temperature and low pressure.
   (b) low temperature and high pressure.
   (c) high temperature and low pressure.
   (d) high temperature and high pressure.

3. A quantity of sodium chloride is added to water in a beaker and stirred briskly for several minutes. The mixture is allowed to settle and appears as a clear colourless liquid with some white crystals at the bottom as illustrated. It would be true to say that:



   (a) the clear liquid is a saturated solution.
   (b) a little more water is needed in order to produce a saturated solution.
   (c) the clear liquid is a supersaturated solution.
   (d) cooling the solution would make it more concentrated.

4. Adding some salt to water will affect its properties. It would be true to say that the salt would:
   (a) raise the water's vapour pressure.
   (b) raise the water's boiling point.
   (c) raise the water's freezing point.
   (d) lower the water's boiling point.

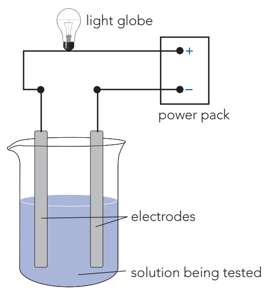
5. Approximately equal volumes of the two solutions in the following pairs of solutions listed below are mixed together. All solutions are 0.1 mol L^{-1}. A precipitate will form in:
   (a) (i) and (ii) only.
   (b) (ii) and (iii) only.
   (c) (i), (ii) and (iii) only.
   (d) (i), (ii) and (iv) only.

(i)	$Pb(NO_3)_{2(aq)}$	and	$Na_2CO_{3(aq)}$
(ii)	$NaCl_{(aq)}$	and	$AgNO_{3(aq)}$
(iii)	$BaCl_{2(aq)}$	and	$Na_2SO_{4(aq)}$
(iv)	$NH_4NO_{3(aq)}$	and	$KCl_{(aq)}$
(v)	$MgCl_{2(aq)}$	and	$(NH_4)_2SO_{4(aq)}$

6. Which of the following equations correctly expresses the dissolving of barium chloride in water?

(a) $BaCl_{2(s)} + H_2O \rightarrow BaCl_{2(aq)}$
(b) $BaCl_{2(s)} \rightarrow Ba^{2+}{}_{(aq)} + 2Cl^-{}_{(aq)}$
(c) $BaCl_{2(s)} + H_2O \rightarrow BaCl_{(aq)} + Cl^-{}_{(aq)}$
(d) $BaCl_{2(s)} \rightarrow BaCl_{2(aq)}$

7. Conductivity tests were carried out using the apparatus shown at right.



Test A - beaker contains distilled water. When hydrogen chloride gas is bubbled through it the globe begins to glow brightly.

Test B - beaker contains ethanol. When hydrogen chloride gas is bubbled through it the globe does not glow.

Which of the following best explains these observations?

(a) Electrons are able to flow through an aqueous solution of $HCl_{(g)}$ but not through an ethanol solution of $HCl_{(g)}$.
(b) Distilled water is a good conductor of electricity whereas ethanol is not.
(c) Water molecules are more easily ionised than ethanol molecules.
(d) $HCl_{(g)}$ dissolves in water to form charged particles whereas in ethanol it remains as neutral molecules.

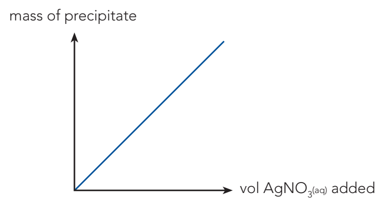
8. Conductivity apparatus similar to that in question 7 is used to test the conductivity of the following solutions. In which case will the globe glow brightest?

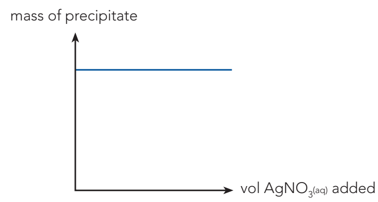
(a) 300 mL of 1.0 mol L^{-1} $HCl_{(aq)}$
(b) 100 mL of 2.0 mol L^{-1} $H_2CO_{3(aq)}$
(c) 300 mL of 1.0 mol L^{-1} $H_2CO_{3(aq)}$
(d) 200 mL of 2.0 mol L^{-1} $CH_3COOH_{(aq)}$

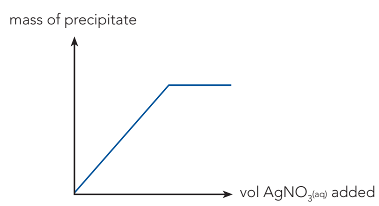
9. If 2.0 mg of sodium chloride crystals were dissolved in 100 mL of distilled water which of the following would most correctly describe the solution? Assume density of the resulting solution to be 1.0 g mL^{-1}.

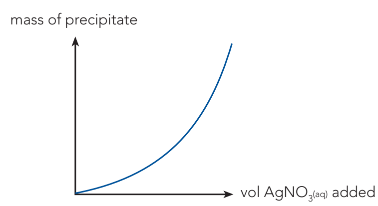
(a) A 2.0% salt solution.
(b) A 20 g L^{-1} salt solution.
(c) A 0.348 mol L^{-1} salt solution.
(d) A 20 ppm salt solution.

10. Silver nitrate solution was added in excess to a sodium chloride solution. The mixture was constantly stirred as the silver nitrate solution was added. Select which of the following graphs most likely represents the mass of precipitate formed as the mixing occurred.

(a)


(b)


(c)


(d)


SECTION 2 – SHORT AND EXTENDED ANSWER (60 MARKS)

Answer each question in the space provided beneath the question. Show all working where calculations are involved.

11. (a) Dissociation and ionisation are different processes. Explain what occurs in each case.

(i) Dissociation

___

___

(ii) Ionisation

___

___

[4 marks]

(b) Give a balanced equation to illustrate each of the above processes.

(i) Dissociation

___

___

(ii) Ionisation

___

___

[4 marks]

12. Write balanced **ionic** equations and describe what would be observed when:

(a) sodium sulfate solution is mixed with barium hydroxide solution.

equation: ___

observation: ___

(b) some calcium hydroxide powder is added to water.

equation: ___

observation: ___

(c) a few drops of silver nitrate are added to household tap water.

equation: ___

observation: ___

[9 marks]

13. (a) Briefly describe what is meant by an **electrolyte**.

____

____

(b) List each of the following as either strong, weak or non-electrolytes.

$C_{12}H_{22}O_{11}$ (sugar), K_2CO_3, $NaCH_3COO$, H_3PO_4 , $MgSO_4$,
H_2O, NH_3 , H_2CO_3 , CH_3CH_2OH (ethanol)

(i) strong electrolytes: ______________________________

(ii) weak electrolytes: ______________________________

(iii) non-electrolytes: ______________________________

[5 marks]

14. Clearly explain each of the following:

(a) What is meant by a saturated solution.

____

____

(b) Why salty water is more difficult to freeze than pure water.

____

____

(c) Why a salt solution can conduct an electric current but a sugar solution cannot.

____

____

(d) Why a 1.0 mol L^{-1} solution of $MgCl_2$ is a better conductor than a 1.0 mol L^{-1} solution of NH_4Cl.

____

____

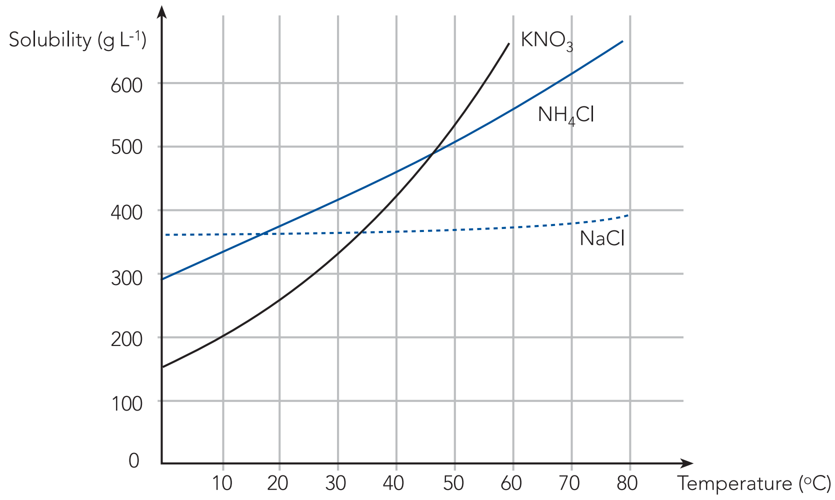
(e) Why the addition of salt to water reduces its vapour pressure and increases its boiling point.

____

____

[10 marks]

15. Use the graph of solubility shown to answer the following. Show all working.



(a) What is the maximum amount of NaCl that can be dissolved in 5.0 L of NaCl solution at 20°C ?

(b) 100 g of NH_4Cl were dissolved in warm water to make up 250 mL of solution. To what temperature must this solution be cooled for crystals to form?

(c) 100 g of KNO_3 crystals are dissolved in distilled water at 50°C to make up 200 mL of solution. The solution is then allowed to cool to 10°C. What mass of KNO_3 crystals will form?

[12 marks]

16. (a) A solution of barium chloride is made up by dissolving 2.55 g of this substance in distilled water and making up to 250.0 mL. Calculate the concentration of this solution in:

(i) grams per litre

(ii) moles per litre

[6 marks]

(b) A particular brand of beer contains 5.0% (w/w) alcohol (C_2H_5OH). Calculate the mass of alcohol that would be contained in a 330 mL bottle of this beer.

[4 marks]

(c) A 5.00 kg sample of sea water was evaporated to dryness and 155.5 g of solids remained.

Calculate the concentration of the solids in:

(i) % by mass

(ii) ppm

[6 marks]

END OF TEST (80 MARKS)

TRIAL TEST 9: ACIDS AND BASES

Time allowed:	**70 minutes**	**Section 1 –** Multiple Choice	20 marks
Total marks:	**80**	**Section 2 –** Short & Extended Answer	60 marks

SECTION 1 – MULTIPLE CHOICE (20 MARKS)

1. Which of the following equations shows the first reactant listed acting as an acid according to the Arrhenius theory?

   (a) $2\ Na_{(s)} + 2\ H_2O_{(aq)} \rightarrow 2\ Na^+_{(aq)} + 2\ OH^-_{(aq)} + H_{2(g)}$

   (b) $H_2PO_4^-{}_{(aq)} \rightarrow HPO_4^{2-}{}_{(aq)} + H^+_{(aq)}$

   (c) $CH_3NH_{2(aq)} + CH_3COOH_{(aq)} \rightarrow CH_3NH_3^+{}_{(aq)} + CH_3COO^-_{(aq)}$

   (d) $HS_{(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow S^{2-}{}_{(aq)} + HCO_3^-{}_{(aq)}$

2. When a solution is formed by dissolving 1 mole of phosphoric acid in 1 litre of water, which of the following would be present in the greatest concentration?
   (a) H_3PO_4 molecules.
   (b) $H_2PO^-_4$ ions.
   (c) HPO_4^{2-} ions.
   (d) PO_4^{3-} ions.

3. A student conducted a set of experiments to determine if a substance was an acid or a base. Which of the following results would support the finding that the substance was an acid?
   I Turned red litmus blue.
   II Produced bubbles of a colourless odourless gas when mixed with sodium hydrogen carbonate.
   III Had a pH of 5.2.
   IV Produced a white precipitate when mixed with a solution of barium nitrate.

   (a) All of I, II, III and IV.
   (b) II, III and IV only.
   (c) III and IV only.
   (d) II and III only.

4. Strong bases are:
   (a) Substances that react completely with acids to form H_2O.
   (b) Recommended for use when cleaning up spills of strong acids.
   (c) Oxides and hydroxides of group I and group II elements.
   (d) Able to complete dissociate when placed in water to form positive metal ions and negative hydroxide ions.

5. Which of the following statements is correct?
   (a) Sea water is slightly basic as it has a pH slightly less than 7.
   (b) Sodium carbonate is a weak base because it is only slightly soluble in water.
   (c) Barium oxide can react with acids and bases and is called an amphoteric oxide.
   (d) Rain water that turns blue litmus red has a $[H^+] > [OH^-]$.

6. Which of the following would be the best conductor of an electric current?
   (a) 0.1 mol L^{-1} HNO_3.
   (b) 0.1 mol L^{-1} NH_3.
   (c) 0.1 mol L^{-1} H_3PO_4.
   (d) 0.1 mol L^{-1} CH_3COOH.

7. Which of the following lists contains an acid, a base and a neutral substance?
   (a) LiOH , H_2O , $CaCO_3$.
   (b) $Ca(OH)_2$, NaCl , KNO_3.
   (c) CO_2 , NH_3 , Na_2CO_3.
   (d) $MgCl_2$, MgO , NO_2.

8. A 0.1 mol L^{-1} HNO_3 solution has a pH of about 1 while a 0.1 mol L^{-1} CH_3COOH solution has a pH of about 3. Which of the following provides the best explanation for this?
   (a) As HNO_3 donates protons more readily than CH_3COOH, a greater percentage of HNO_3 molecules than CH_3COOH molecules will react with water to form H_3O^+ ions.
   (b) A CH_3COOH molecule ionises to yield more H^+ ions than HNO_3.
   (c) Each HNO_3 molecule ionises completely but each CH_3COOH molecule is only partially ionised.
   (d) HNO_3 is ionic and the ions dissociate completely in aqueous solution, whereas CH_3COOH is molecular and only partially ionises in aqueous solution.

9. Which of the following provides the best explanation of why $Cr(OH)_3$ is called amphoteric?
   (a) $Cr(OH)_3$ is a stronger acid than it is a base.
   (b) $Cr(OH)_3$ is not soluble in water.
   (c) $Cr(OH)_3$ has a pH of 7.0.
   (d) $Cr(OH)_3$ dissolves in $HCl_{(aq)}$ and in $NaOH_{(aq)}$.

10. Which one of the following solutions would have the lowest pH?
   (a) 0.1 mol L^{-1} NaOH
   (b) 0.1 mol L^{-1} CH_3COOH
   (c) 0.1 mol L^{-1} KCl
   (d) 0.1 mol L^{-1} Na_2CO_3

SECTION 2 – SHORT AND EXTENDED ANSWER (60 MARKS)

Answer each question in the space provided beneath the question. Show all working where calculations are involved.

11. Using the Arrhenius theory, classifying the following as acids or bases when they dissolve in water. Write equations to support your answer.

(a) $NaOH$

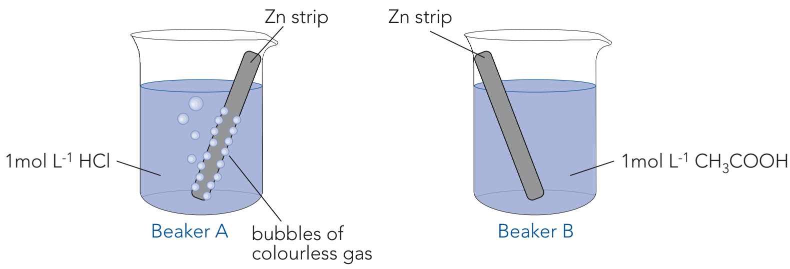
(b) CH_3COOH

(c) $Ba(OH)_2$

(d) H_2S

[8 marks]

12.



(a) Complete the diagram for Beaker B so as to show any differences to what is happening in Beaker A.

[1 mark]

(b) Explain the reason for your answer to part (a).

[3 marks]

13. (a) Complete the following general equations:

(i) Acid + Metal Hydroxide → ______________________

(ii) Acid + Metal Carbonate → ______________________

(iii) Acid + Metal Oxide → ______________________

[3 marks]

(b) Write balanced, ionic equations for each of the following:

(i) $CaCO_{3(s)} + HNO_{3(aq)}$ ______________________

(ii) $Mg(OH)_{2(s)} + H_2SO_{4(aq)}$ ______________________

(iii) $NaHCO_{3(aq)} + H_3PO_{4(aq)}$ ______________________

(iv) $SO_{2(g)} + H_2O_{(l)}$ ______________________

(v) ammonium nitrate solution + sodium hydroxide solution →

____

(vi) sulfuric acid solution + magnesium → ______________________

(vii) solid sodium oxide + hydrochloric acid solution →

____

(viii) ammonia + hydrochloric acid → ______________________

[16 marks]

14. Complete the following table by adding the formula from the following list to the correct column. *N.B. each formula may be placed in more than one column and assume that each has been dissolved in water.*

$NaCl$, HBr, $Ca(OH)_2$, $Ca(NO_3)_2$, SO_2, Na_2O

TURNS RED LITMUS BLUE	TURNS GREEN IN UNIVERSAL INDICATOR	HAS A PH GREATER THAN 7	HAS A PH LESS THAN 7

[8 marks]

15. (a) Rank the following solutions in order of increasing pH:

1 mol L^{-1} Na_2CO_3, 1 mol L^{-1} HNO_3, 1 mol L^{-1} $NaCl$
1 mol L^{-1} KOH, 1 mol L^{-1} CH_3COOH, 1×10^{-6} mol L^{-1} HCl

1. ______________ 2. ______________ 3. ______________
(lowest pH)

4. ______________ 5. ______________ 6. ______________
(highest pH)

[6 marks]

(b) The addition of ammonium nitrate fertiliser causes some paddocks to gradually become acidic. What steps could a farmer take to increase the pH of acid soils?

[2 marks]

16. (a) Use equations to explain why sulfuric acid is considered to be a polyprotic acid.

[3 marks]

(b) Briefly describe an experiment that could be used to compare the strengths of a group of acids.

[4 marks]

(c) Aluminium hydroxide is classified as an amphoteric hydroxide. Use equations to show this.

[3 marks]

(d) With the aid of an equation, explain why water is considered to be a weak electrolyte.

[3 marks]

END OF TEST (80 MARKS)

TRIAL TEST 10: REACTION RATES

Time allowed: 70 minutes
Total marks: 80

Section 1 – Multiple Choice 20 marks
Section 2 – Short & Extended Answer 60 marks

SECTION 1 – MULTIPLE CHOICE (20 MARKS)

1. The production of ammonia can be represented by the following equation:

$$N_{2(g)} + 3\,H_{2(g)} \rightarrow 2\,NH_{3(g)} + \text{energy}$$

Which of the following changes would increase the rate of production of ammonia?
(a) Decreasing the temperature.
(b) Addition of $HCl_{(g)}$.
(c) Increasing the volume of the reaction container.
(d) Increasing the pressure.

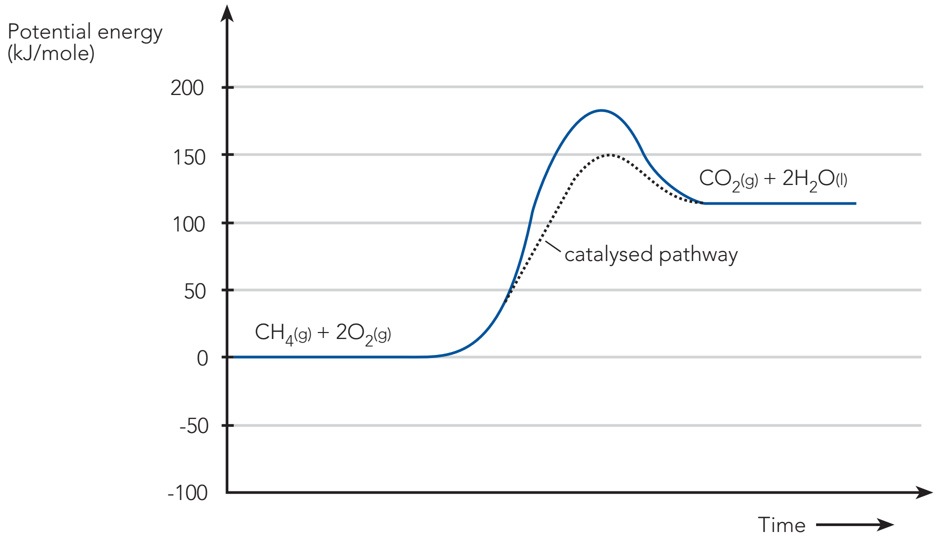
2. The rate chemical reactions generally decreases with time because:
(a) Reactant concentrations decrease with time.
(b) Heat is lost to the surroundings.
(c) A catalyst is required to maintain a constant reaction rate.
(d) The fraction of reactant particles with energies greater than the activation energy decreases with time.

3. A strip of zinc is being dissolved by a solution of hydrochloric acid. The addition of a piece of copper causes an increase in the number of bubbles of H_2 gas. The most probable reason for this is:
(a) The copper has increased the reaction surface area.
(b) The copper increases the concentration of metal available for reaction.
(c) Copper is more reactive than the zinc.
(d) Copper acts as a catalyst.

4. The effect of a catalyst on a chemical reaction is to:
(a) decrease the activation energy.
(b) increase the activation energy.
(c) increase the potential energy of the reactants.
(d) decrease the ΔH.

5. The reaction rate between two gaseous reactants will increase if there is a rise in temperature. The main reason for this is:
(a) an easier reaction pathway is possible.
(b) an increase in pressure inside the reaction vessel.
(c) an increase in collisions between product molecules.
(d) an increase in the proportion of molecules with sufficient activation energy.

Questions 6 and 7 refer to the following diagram.



6. For an uncatalysed pathway, the activation energy and ΔH of the above reaction would be approximately:
(a) 150 kJ , 120 kJ
(b) 180 kJ , 120 kJ
(c) 180 kJ , –60 kJ
(d) –180 kJ , –120 kJ

7. For the reverse reaction using a catalyst the activation energy and ΔH would be approximately:
(a) 180 kJ , –120 kJ
(b) 60 kJ , –120 kJ
(c) 30 kJ , 120 kJ
(d) 30 kJ , –120 kJ

8. Hydrogen gas is produced by reacting zinc metal with concentrated hydrochloric acid. The equation for the reaction is as follows:

$$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

Which of the following **could not** be used to monitor the rate of reaction?
(a) Mass of zinc remaining.
(b) Total mass of reactants and products in a closed container.
(c) Volume of hydrogen gas produced.
(d) Concentration of hydrogen ions in solution.

9. Carbon dioxide is produced by reacting marble chips with dilute hydrochloric acid. The marble chips are placed in a flask and *completely covered* by the acid to begin the reaction. The equation for the reaction is as follows:

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

Which of the following **would not** increase the initial reaction rate?
(a) Adding more acid.
(b) Agitating and stirring the mixture.
(c) Crushing the marble chips.
(d) Adding a catalyst.

10. Nitrogen monoxide is a colourless gas that reacts with oxygen gas to produce the brown gas nitrogen dioxide.

$$NO_{(g)} + 2\ O_{2(g)} \rightarrow 2\ NO_{2(g)}$$

The concentration of NO_2 would increase if:
(a) The temperature was decreased.
(b) The concentration of the O_2 was decreased.
(c) The pressure on the system was decreased.
(d) The size of the reaction vessel was decreased.

SECTION 2 – SHORT AND EXTENDED ANSWER (60 MARKS)

Answer each question in the space provided beneath the question. Show all working where calculations are involved.

11. The car engine is designed to burn petrol is oxygen to produce carbon dioxide gas, water vapour and large amounts of energy to move the car. It is very important that the reaction is very rapid.

(a) If petrol is given the formula $C_7H_{16(l)}$, write the balanced equation for the combustion of petrol.

[4 marks]

Use the collision theory to explain the reason for the following modifications to car engines:

(b) Fuel injection systems cause the petrol to be sprayed into the combustion chamber as a very fine mist.

[4 marks]

(c) Turbo chargers use exhaust gases to drive a turbine which pumps air into the combustion chamber at much higher pressure than normal.

[4 marks]

12. Hydrogen peroxide disproportionates to form oxygen gas and water according to the following equation:

$$2\ H_2O_{2(l)} \rightarrow O_{2(g)} + 2\ H_2O_{(l)}$$

In the human body, H_2O_2 is a very harmful bi-product of normal metabolic processes. The body produces the enzyme, catalase to increase the rate at which H_2O_2 breaks down to water and oxygen gas.

Use the Collision Theory to explain the action of catalase. Include an energy profile diagram in your answer.

[8 marks]

13. Explain why increasing the temperature of a system leads to an increased reaction rate.

[6 marks]

14. Give a clear explanation for each of the following:

(a) Granulated sugar will dissolve more quickly into a cup of coffee than will an equivalent amount of sugar in the form of cubes.

______

______

(b) Hydrogen gas mixed with air in a test tube will not burn or explode unless a flame is brought near it.

______

______

(c) Food stored in a refrigerator will keep longer.

______

______

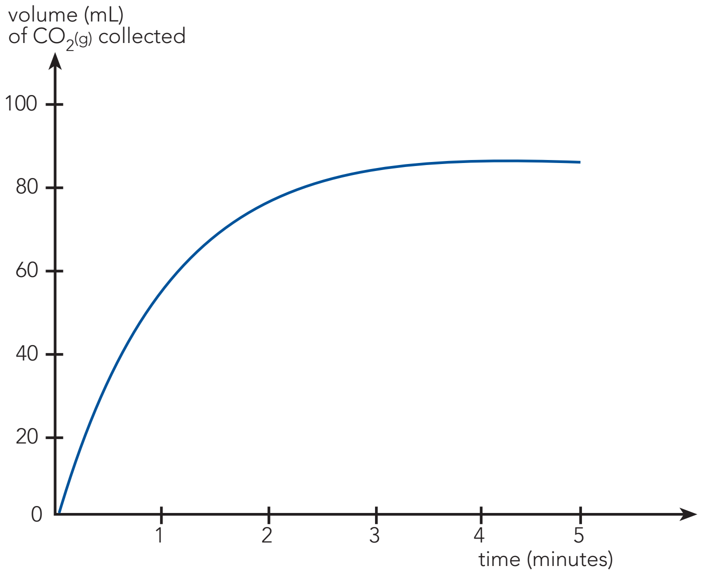
(d) Magnesium ribbon will burn much more brightly in pure oxygen than in air.

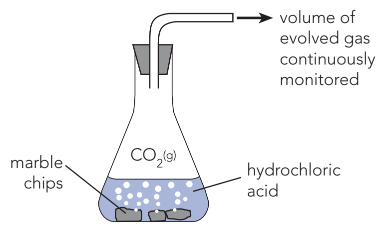
______

______

[16 marks]

15. Carbon dioxide gas is generated as shown below by reacting marble chips (calcium carbonate) with 1.0 mol L^{-1} hydrochloric acid. As the reaction proceeds the volume of CO_2 gas produced was measured at regular intervals. The results are shown graphically below:





(a) Write a balanced equation for the reaction.

______

(b) Use the graph to determine when the reaction rate is greatest.

(c) Explain why the rate of reaction changes as the reaction proceeds.

(d) Explain what happens in the reaction after the first 4 minutes.

(e) What other quantity could have been measured and graphed in order to monitor the rate of this reaction?

(f) This reaction was carried out using 1.0 mol L^{-1} HCl acid. If the experiment was repeated using 2.0 mol L^{-1} HCl acid,

(i) How would the reaction be different?

(ii) Why?

(iii) Indicate on the graph a likely set of results.

[12 marks]

END OF TEST (80 MARKS)

ANSWERS TO CHAPTER AND REVIEW QUESTIONS

CHP 1: ATOMIC STRUCTURE

Chapter Questions

1.1

(a) $^{7}_{3}Li$ 3 protons, 4 neutrons, 3 electrons.
(b) $^{14}_{7}N$ 7 protons, 7 neutrons, 7 electrons.
(c) $^{19}_{9}F$ 9 protons, 10 neutrons, 9 electrons.
(d) $^{35}_{17}Cl$ 17 protons, 18 neutrons, 17 electrons

1.2

ISOTOPE	NAME OF ISOTOPE	Z ATOMIC NO.	A MASS NO.	NO. OF PROTONS	NO. OF NEUTRONS
$^{12}_{6}C$	Carbon – 12	6	12	6	6
$^{14}_{6}C$	Carbon – 14	6	14	6	8
$^{24}_{12}Mg$	Magnesium – 24	12	24	12	12
$^{40}_{18}Ar$	Argon – 40	18	40	18	22
$^{27}_{13}Al$	Aluminium – 27	13	27	13	14
$^{59}_{27}Co$	Cobalt – 59	27	59	27	32

1.3

Ar (Br) = $\sum$ (isotopic mass × abundance %)/100

= $\sum$ ((78.92 × 50.69) + (80.92 × 49.31))/100

= 79.91

1.4

(a) Total peak heights = 19.75 + 2.50 + 2.75
= 25.00

Abundance %; ^{24}Mg = (19.75/25.00) × 100
= 79.0%

^{25}Mg = (2.50/25.00) × 100 = 10.0%

^{26}Mg = (2.75/25.00) × 100 = 11.0%

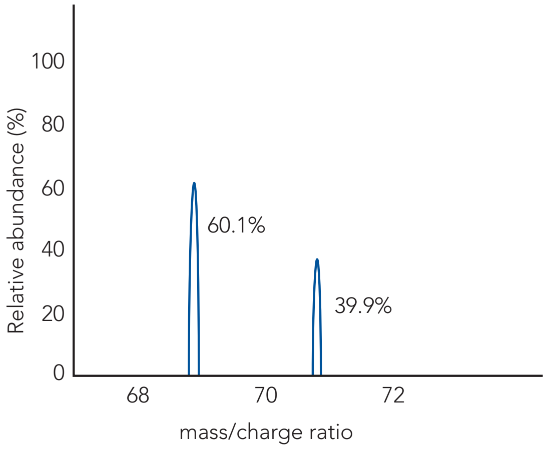
(b) Ar (Br) = $\sum$ (isotopic mass × abundance %)/100

= $\sum$ ((23.99 × 79.00) + (24.99 × 10.0) + (25.98 × 11.0))/100

= 24.31

1.5

(a)



(b) Ar (X) = $\sum$ (isotopic mass × abundance %)/100

= $\sum$ ((68.92 × 60.1) + (70.92 × 39.9))/100

= 69.72

1.6

Let % abundance of ^{63}Cu be x.
Hence % abundance of ^{65}Cu will be (100-x)
Now
Ar (Cu) = $\sum$ (isotopic mass x abundance %)/100
= 63.55
So $\sum$ ((62.93)(x) + (64.93)(100 - x))/100
= 63.55
62.93x + 6493 - 64.93x = 6355
2.00x = 138
Hence % abundance of ^{63}Cu is = 69 %
and % abundance of ^{65}Cu is = 31 %

1.7

(a) Red (b) UV (c) Orange (d) 4

1.8

(a) 3 (b) 6 (c) 10

1.9

(a) C 2,4 (d) F 2,7
(b) Cl 2,8,7 (e) Ca 2,8,8,2
(c) Mg 2,8,2 (f) B 2,3

1.10

(a) Si 4 (d) F 7
(b) Al 3 (e) Cl 7
(c) S 6 (f) C 4

1.11

The energy levels of different metallic ions are unique since their atomic structures are also all different. Hence the electron transitions possible, and the resulting colours visible, are unique to each metallic ion.

1.12

The colours observed are dependent only on the metallic ion. In this case they are both sodium salts and a yellowish flame will result.

1.13

Place a small sample of one of the salts in a colourless Bunsen flame. A reddish colour would indicate a calcium salt while a pale green colour a barium salt.

1.14

(a) The prism disperses the light into its component colours.

(b) The monochromator allows only one colour (or wavelength) to pass through to the detector at a time.

1.15

(a) The metal coating of the cathode lamp allows it to produce light containing the specific wavelength that can be absorbed by the metal sample being analysed.

(b) The light from the lamp is pulsed so that the detector can distinguish it from the light emitted by the flame containing the sample. In this way the detector can determine if any of the lamp light has being absorbed and if so to what extent.

1.16

One electron in outer level

1.17

(a) 2 (b) 7 (c) 8

1.18

(a) S (b) Ca (c) Ar

1.19

(a) :Ar: (8 dots) *(c)* •B• (3 dots)

(b) :Cl• (7 dots) *(d)* •N• (5 dots)

1.20 *Nitrogen*

1.21

Li	*2 , 1*	*Li* •	*N*	*2 , 5*	•N• (5 dots)
Be	*2 , 2*	• *Be* •	*O*	*2 , 6*	:O• (6 dots)
B	*2 , 3*	•B• (3 dots)	*F*	*2 , 7*	:F• (7 dots)
C	*2 , 4*	•C• (4 dots)	*Ne*	*2 , 8*	:Ne: (8 dots)

1.22

(a) Si, C, N. (b) K, Na, Mg.

1.23

(a) 5th ionisation energy → electron is being removed from a lower energy level..

(b) 4 electrons

1.24

(a) 2 (b) +2

(c) Metal

- relatively low 1st and 2nd ionisation energies

- has only 2 valence electrons.

1.25

An atom of chlorine has a much greater number of protons than an atom of sodium. This means that the electrons of the chlorine atom are much more strongly attracted by the nucleus and a smaller radius for the chlorine atom results.

1.26

Oxygen will be more electronegative than sulfur since electronegativity decreases as we go down a group of the periodic table. This can also be explained by the fact that the outer level of a sulfur atom is further out than that for oxygen and the attraction for electrons would be weaker.

1.27 *(a) Ca (b) K (c) Cs (d) Ca*

1.28

Calcium is more metallic than magnesium since it is lower in group 2 of the periodic table and hence has a lower ionisation energy. This indicates that electrons are more easily removed from calcium and a more metallic character results.

1.29

(a) P, Si, Al (c) P, Si, Al

(b) Al, Si, P (d) Al. Si, P

Review Questions

1.

PARTICLE	PROBABLE LOCATION	RELATIVE MASS	RELATIVE CHARGE
proton	nucleus	1	+1
neutron	nucleus	1	0
electron	electron cloud	$\frac{1}{1836}$	−1

2. *(a) 3 protons, 4 neutrons, 3 electrons*
   (b) 11 protons, 12 neutrons, 11 electrons
   (c) 17 protons, 20 neutrons, 17 electrons
   (d) 26 protons, 30 neutrons, 26 electrons

3.

ISOTOPE	SYMBOL	NO. OF PROTONS	NO. OF NEUTRONS
Boron-8	${}^{8}_{5}B$	5	3
Boron-10	${}^{10}_{5}B$	5	5
Boron-11	${}^{11}_{5}B$	5	6
Boron-12	${}^{12}_{5}B$	5	7
Boron-13	${}^{13}_{5}B$	5	8
Boron-14	${}^{14}_{5}B$	5	9

4.
A_r (Cl) = ∑ (isotopic mass × abundance %)/100

= ∑ ((34.97 × 75.8) + (36.97 × 24.2))/100

= 35.45

5. *Let % abundance of ^{85}Rb be x. Hence % abundance of ^{87}Rb will be (100-x).*
Now
A_r (Rb) = ∑ (isotopic mass × abundance %)/100
= 85.468
So
∑ ((84.91)(x) + (86.91)(100 - x))/100
= 85.468
84.91x + 8691 - 86.91x
= 8546.8
2.00x = 144.4
Hence % abundance of ^{85}Rb is = 72.2%
and % abundance of ^{87}Rb is = 27.8%

6. *Order of increasing wavelength:*
X-rays, UV light, visible light, microwaves, radio waves.

7. *In the discharge tube the electrons of the hydrogen atoms are being continually excited from their stable state to higher energy levels. As the electrons return to lower energy levels they emit photons whose energies are equal to the difference between the two levels involved. Since only specific transitions are possible then only photons of specific energy are emitted. These correspond to different wavelengths of radiation, some of which are in the visible light range.*

8.
(a) Flame colours are the result of visible radiation emitted by the excited electrons falling to their original ground state within the metallic ions. Different atoms have different possible energy levels and hence any energy level transitions are unique. This results in different colours being observed.
(b) Flame tests are only qualitative. Colours may be masked by impurities. A Bunsen burner flame is not suitable for many metals due to its relatively low temperature.

9.
(a) 2 *(d) 2, 8, 2*
(b) 2, 8, 3 *(e) 2, 4*
(c) 2, 8, 1 *(f) 2, 8, 8, 2*

10.
(a) 1 *(d) 7*
(b) 2 *(e) 6*
(c) 8

11.
(a) They have the same valence electron configuration.
(b) 8
(c) It is chemically stable or inert.
(d) They try to get the same configuration as their nearest noble gas.

12.
Al^{3+} *2*
S^{2-} *2,6*
O *2,8*
He *2,4*
C *2,8,8*

13. *(c)*

14. *They have 2 loosely held electrons in their outermost energy level.*

15.
(a) Mg^{2+} *2, 8*
(b) Cl^{-} *2, 8, 8*
(c) S^{2-} *2, 8, 8*
(d) Ar *2, 8, 8*
(e) H *1*
(f) Be^{2+} *2*

16.
(a) • Ca • *(c)* :Ö •
(b) :F̈ • *(d)* :N̈e:

17. *Energy required to remove the most loosely held electron from an atom in the gaseous state.*

18.
(a) Approximate ionisation energies:
Cl: 1300 kJ ; Kr: 1320 kJ
(b) Ionisation energy decreases down a group.
(c) $F_{(g)}$ + 1680 kJ → $F^{+}_{(g)}$ + e^{-}

19.
(a) Na
(b) F
(c) The noble gas of that period.

20.
(a) Increased distance between nucleus and

outermost electrons plus increase in number of electrons between nucleus and outermost electron (shielding).

(b) *Although the outermost electrons are all on the same principal energy level the nuclear charge increases. This means that the electrons are more strongly attracted to the nucleus.*

21. *As each electron is removed the distance between the nucleus and the "outermost" electron decreases and so electrostatic attraction increases. Also, electrons are being removed from an increasingly positive ion rather than a neutral atom.*

22.

	GROUP NUMBER	VALENCY
X	3	+3
Y	2	+2
Z	4	+4 or −4

23.
i) Z ; ii) X ; iii) Y : magnitude of the 5th ionisation energy indicates the relative energy level – the higher the 5th ionisation energy the lower the energy level.

24.
(a) ***For the Experts***

ISOTOPE	SYMBOL	NUMBER OF PROTONS	NUMBER OF NEUTRONS
Carbon-12	$^{12}_{6}C$	6	6
Carbon-13	$^{13}_{6}C$	6	7
Carbon-14	$^{14}_{6}C$	6	8

(b) A_r *(carbon)* $= \frac{98.89}{100}(12.000) + \frac{1.11}{100}(13.003)$

$= 12.01$

(c) *X is* $^{1}_{1}P$ *or* $^{1}_{1}H$

Y is $^{0}_{-1}e$ *(Beta particle)*

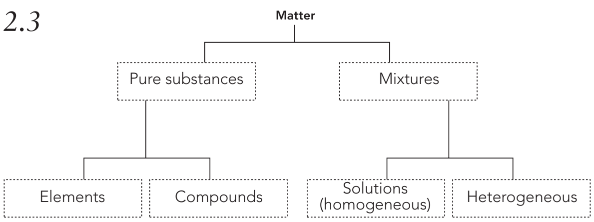
(d)
(i) *Carbon-14 atoms would be slightly heavier and radioactive.*
(ii) *Both would undergo the same chemistry because both have the same number (i.e. 4) valence electrons.*

CHP 2: MATERIALS

Chapter Questions

2.1 *Elements: cannot be separated into simpler substances.*
: gold, oxygen, aluminium.
Compounds: two or more elements chemically combined.
: sugar, water, rubber, salt.
Homogeneous mixture: uniform composition throughout (solutions)
: sea water, air, solder, brass, petrol.
Heterogeneous mixture: non uniform composition.
: limestone, rock, cement.

2.2
(a) *A solution is a mixture and its constituents are not chemically combined. A compound is pure.*
(b) *Mixtures are either homogeneous or heterogeneous. A solution is a homogeneous mixture.*
(c) *The salt is a compound since it can be broken into simpler substances.*

2.3


2.4
- *Elements: oxygen gas, zinc, gold*
- *Compounds: pure water, sugar, carbon dioxide*
- *Mixtures: sea water, vinegar, air*

2.5
(i) *distillation*
(ii) *evaporation/crystallisation*
(iii) *fractional distillation*
(iv) *filtration*
(v) *filtration and fractional crystallisation.*

2.6
(a) *Cooling water causes the vapour to condense quickly.*
(b) *Allows water jacket to fill, increases cooling effect and minimises water use.*
(c) *Simple distillation separates a single liquid from a solution containing only dissolved salts, whereas fractional distillation can separate two or more different liquids from solution.*

2.7 *Place the mixture in water after first removing the iron filings with a magnet. The leaves will float and can be removed with a small sieve or floated off and filtered. The sand can* *be recovered by filtration and the sugar by recrystallisation of the filtrate.*

2.8

(i) salt from water: salt will crystallise from solution.
(ii) water from salt: water evaporates, salt does not.
(iii) sand from water: sand is insoluble.
(iv) charcoal from salt: salt is soluble, charcoal is not.
(v) KNO_3 from NaCl: Differences in solubility, NaCl crystallises first.
(vi) blue dye from ink: selective adsorption of different components of ink.
(vii) alcohol from wine: different boiling points allow fractional distillation.
(viii) salt from sugar: differences in solubility.

2.9

(a) (i) 10^9 (ii) 10^6 (ii) 1000
(b) (i) 0 .22 m / 1.1×10^{-9} m = 2.0×10^8 times larger
(ii) Volume of a sphere is proportional to its radius cubed.
Hence volume ratio = $(0.11)^3$ / $0.55 \times 10^{-9})^3$ = $(2.0 \times 10^8)^3$ = 8.0×10^{24}
Hence volume of the soccer ball is 8.0×10^{24} times larger than that of a buckyball.

2.10

(a) 2 cm = 20 mm hence
N° of cubes with 1 mm sides = $20 \times 20 \times 20$ = 8000

(b) Total surface area = (8000) × (6) × (0.1 cm × 0.1 cm)
= 480 cm^2

(c) Change in surface area = 480/24 = 20
The surface area is 20 times greater.

(d) For nanocubes of 1.0 nm sides; 2 cm = 2.0×10^7 nm
N° of cubes = $(2 \times 10^7) \times (2 \times 10^7) \times (2 \times 10^7)$ = 8×10^{21}
S.A. = $(8 \times 10^{21}) \times (6) \times (1.0 \times 10^{-7}cm \times 1.0 \times 10^{-7}cm)$
= 48×10^7 cm^2
Change in S.A. = $48 \times 10^7/24 = 2 \times 10^7$ times

2.11

ZnO nanoparticles are good absorbers of harmful UV rays present in sunlight. They are also smaller than the wavelength of visible light. Hence light is able to pass between the particles creating a transparent effect.

2.12

(a) The carbon nanotubes create a surface with nano-size bumps and low surface energy.
(b) The higher surface energy of water prevents it from wetting the fabrics surface. Instead the water will bead into droplets, roll over the nano-sized bumps of the fabric, and carry any dirt along with it.
(c) Water and dirt also roll off the leaves of lotus plants. The nano-sized bumps on its surface prevent water from soaking it. This was the inspiration for the development of stain resistant and waterproof fabrics.

2.13

If water droplets form and then dry on glass, they leave water marks due to small amounts of dirt or impurities in them. The effect of the TiO_2 layer is to prevent droplets forming and so water runs off carrying any dirt along with it. The photo-catalytic properties of the TiO_2 layer also help to break down any organic material on the glass.

2.14

Nanoparticles used in nanocomposite materials have very high surface area to volume ratios. This markedly increases the interaction and bonding which occurs between the atoms and particles in the material. The added nanoparticles may also have a high aspect ratio, that is, they are thin and long. This further helps to bond and tie a great number of particles together.

Review Questions

1. *(a) mixture (d) element*
(b) compound (e) mixture
(c) mixture (f) compound.

2. ***homogeneous mixture** e.g. salt solution, coffee drink, cordial.*
***heterogeneous mixture** e.g. concrete mix, sand and salt mixture, fruit cake mix.*

3. *(a) decantation and filtration*
(b) fractional distillation
(c) dissolution and filtration
(d) evaporation/crystallisation
(e) evaporation/distillation.

4. *(a) residue (f) crystallisation*
(b) decantation (g) filtration/crystallisation
(c) mixture (h) evaporation
(d) element (i) distillation
(e) filtrate (j) solution.

5.
(a) 1.1 nm / 0.1 nm = 11
A buckyball is 11 times larger than a hydrogen atom.
(b) 5 μm / 80 nm = 62.5 ≈ 60
A typical bacteria is about 60 times larger than the Adeno virus.

(c) $1\ cm / 1.1\ nm = 0.9 \times 10^7 \approx 10^7$
A bee is about 10,000,000 times larger than a buckyball.

6. *When the particles of a substance become smaller their surface area to volume ratio increases. Hence, for nanoparticles this means that a significant percentage of their atoms are at or near the surface where they can more readily interact.*

7.
(a) $0.4\ \mu m = 400\ nm$, $0.7\ \mu m = 700\ nm$
(b) $0.55\ \mu m / 40\ nm = 550\ nm / 40\ nm = 13.8$
Hence visible light waves are about 14 times larger than the 40 nm nanoparticles.
(c) *The 40 nm particles are too small to reflect visible light and are not visible. The light is either scattered, absorbed, or passes between the particles.*

8.
(a) *For the sand grain particle*
$SA = 4\pi r^2 = 4\pi(0.5 \times 10^{-3})^2 = 3.14 \times 10^{-6}\ m^2$
$Volume = 4/3\ \pi r^3 = 4/3\ \pi(0.5 \times 10^{-3})^3$
$= 0.52 \times 10^{-9}\ m^3$
Similarly for the others as given below:

PARTICLE	SURFACE AREA (m^2)	VOLUME (m^3)	SURFACE AREA / VOLUME RATIO
Sand grain	$3.14 \times 10^{-6}\ m^2$	$0.52 \times 10^{-9}\ m^3$	6.0×10^3
Dust particle	$3.14 \times 10^{-12}\ m^2$	$0.52 \times 10^{-18}\ m^3$	6.0×10^6
Carbon-60 buckyball	$3.14 \times 10^{-18}\ m^2$	$0.52 \times 10^{-27} m^3$	6.0×10^9

(b) *The surface area to volume ratio is much greater for smaller particles?*
(c) *The greater surface area allows more contact and interaction between the atoms of adjoining surfaces.*
This leads to unique chemical properties and often greater reactivity.

9.
(a) *Sunscreens absorb harmful ultraviolet radiation present in sunlight.*
(b) *Bulk particles of zinc oxide are good reflectors of visible light and so appear white. Nanoparticles are much smaller than the wavelength of light and do not readily reflect it. Hence they are not visible.*
(c) *The nanoparticles of zinc oxide are still effective as they retain their excellent UV-light absorbing capacity. They also scatter or absorb visible light and so create a transparent appearance.*

10.
(a) *The carbon nanotubes attach themselves to the fabric fibres and create a surface with nano-sized bumps and low surface energy.*
(b) *This surface prevents water from penetrating or adhering to it.*
(c) *The water forms droplets and rolls off carrying any dirt away with it.*

11.
(a) *Silver* (b) *Carbon nanotubes*
(c) *Titanium dioxide* (d) *Silicon*
(e) *Silicon dioxide*

For the Experts

12.
(a) (i) *pure substance – table salt, sugar, baking soda*
(ii) *mixture – baking powder, vinegar*
(iii) *elements – aluminium foil*
(iv) *compounds – table salt, sugar, baking soda*
(b) (i) *physical change*
(ii) *allow water to evaporate, salt will crystallise.*
(c) (i) *chemical change because new substances are formed*
(ii) $NaHCO_3(aq) + CH_3COOH(aq) \rightarrow NaCH_3COO(aq) + CO_2(g) + H_2O(l)$
(d) (i) *Carbon dioxide gas.*
(ii) *Chemical reaction.*
(iii) *The baking soda and tartaric acid are both in powder form and cannot react unless moisture is present. The starch helps to keep the baking powder dry by absorbing any moisture absorbed from the air.*

CHP 3: BONDING

Chapter Questions

3.1

ELEMENT	ELECTRON DOT DIAGRAM	TO FORM IONS THIS ATOM TENDS TO	ION FORMED	ION HAS SAME ELECTRON CONFIGURATION AS
sodium	Na •	lose 1 e–	Na^+	Ne
magnesium	• Mg •	lose 2 e–	Mg^{2+}	Ne
nitrogen	•• • N • •	gain 3 e–	N^{3-}	Ne
sulfur	•• •• S • •	gain 2 e–	S^{2-}	Ar
bromine	•• •• Br • ••	gain 1 e–	Br^-	Kr
calcium	• Ca •	lose 2 e–	Ca^{2+}	Ar
potassium	K •	lose 1 e–	K^+	Ar
lithium	Li •	lose 1 e–	Li^+	He

3.2
(a) Cations (b) NH_4^+ (ammonium ion)
(c) Hydrogen

3.3
(a) $\underset{2,\,8,\,8}{[K]^+} + \underset{2,\,8}{[:\ddot{\underset{..}{F}}:]^-}$ (b) $\underset{2,\,8}{[Mg]^{2+}} + \underset{2,\,8,\,8}{[:\ddot{\underset{..}{S}}:]^{2-}}$

3.4 (a) 6 (b) 6

3.5 Electrostatic attraction between oppositely charged ions.

3.6 In the molten state, the ions that make up ionic substances such as magnesium chloride are able to move about the liquid freely and hence can conduct electricity (charge movement). In the solid phase, these ions can only vibrate about fixed positions and hence cannot conduct electricity.

3.7
(a) Na atoms lose electrons
(b) Na^+ and O^{2-} ions
(c) 2 : 1
(d) neutral

3.8 By losing their valence electrons.

3.9
Metallic bonds result from the electrostatic attraction between the positively charged metal ions and the delocalised and negatively charged electrons.

3.10
(a) Since the valence electrons in metallic solids are loosely held and very mobile, any applied voltage will cause a flow of charge.
(b) Metallic bonds are non directional and layers of positive ions can simply slip over each other.

3.11
Metallic magnesium has twice as many delocalised electrons which are attracted to doubly positive ions. Hence electrostatic attraction is much greater in Mg than Na.

3.12
The presence of the slightly bigger zinc atoms among the copper atoms causes irregularities in the metallic crystal formed. Layers of metal ions are no longer so regular and cannot slip over each other. This reduces malleability.

3.13
The valence electrons of non metal atoms are very strongly held. Hence these electronegative atoms share their electrons and form covalent bonds.

3.14
Group 15 elements have 3 unpaired electrons and hence can form 3 covalent bonds. Group 17 elements have only 1 unpaired electron.

3.15

(a) $:\ddot{\underset{..}{Cl}}:\ddot{\underset{..}{Cl}}:$ (b) $:\ddot{\underset{..}{Br}}:\ddot{\underset{..}{Br}}:$

(c) $H:C\,\vdots\vdots\,N:$ (d) $\ddot{O}::C::\ddot{O}$

3.16
(a) CO_2 and Cl_2 are substances made up of non polar molecules and cannot carry a charge.
(b) The intermolecular forces between the molecules are very weak. At normal temperatures, the molecules have sufficient energy to move about independently in a gaseous state.

3.17
(a) Covalent network substances like diamond are made up of atoms strongly linked together by covalent bonds. There are no free atoms or electrons to carry charge.
(b) The strongly bonded atoms form a network and a solid at normal temperatures.

3.18
Graphite is made up of atoms which are strongly covalently bonded in two dimensional layers. Within each layer there are delocalised electrons which are able to conduct electricity.

3.19
(a) Different forms (atomic arrangements) of the same element.
(b) Graphite, carbon nanotubes, diamond, amorphous carbon.

3.20
(a) Carbon monoxide
(b) Carbon dioxide
(c) Bromine
(d) Diphosphorus pentoxide
(e) Sulfur dioxide
(f) Sulfur trioxide

3.21
(a) Cl_2 (b) NO_2 (c) CCl_4
(d) SO_3 (e) OCl_2 (f) N_2O_5

3.22

	Cl^-	O^{2-}	N^{3-}	OH^-	SO_4^{2-}
Na^+	$NaCl$	Na_2O	Na_3N	$NaOH$	Na_2SO_4
Mg^{2+}	$MgCl_2$	MgO	Mg_3N_2	$Mg(OH)_2$	$MgSO_4$
Fe^{3+}	$FeCl_3$	Fe_2O_3	FeN	$Fe(OH)_3$	$Fe_2(SO_4)_3$

3.23

barium chloride	$BaCl_2$
aluminium bromide	$AlBr_3$
aluminium nitride	AlN
caesium sulfide	Cs_2S
copper (II) sulfite	$CuSO_3$
zinc phosphate	$Zn_3(PO_4)_2$
cobalt fluoride	CoF_2
lead (II) oxide	PbO
zinc oxide	ZnO
zinc sulfate	$ZnSO_4$
barium ethanoate	$Ba(CH_3COO)_2$
calcium hydroxide	$Ca(OH)_2$
copper (II) hydroxide	$Cu(OH)_2$
ammonium chloride	NH_4Cl
ammonium hydrogensulfate	NH_4HSO_4
chromium (III) oxide	Cr_2O_3

3.24
(a) *Sodium bromide*
(b) *Iron (III) nitrate*
(c) *Iron (II) sulfate*
(d) *Ammonium sulfate*
(e) *Barium oxide*
(f) *Copper (II) carbonate*
(g) *Potassium hydrogencarbonate*
(h) *Copper (II) hydroxide*
(i) *Iron (II) carbonate*
(j) *Barium phosphate*

Review Questions

1.
(a) *Mg, Li, Ca*
(b) *N, Cl, S, F*
(c) *N, Cl, S, F*

2. *Group I and II elements have valence electrons which are not strongly held. When these electrons are lost, to say highly electronegative non metals, positive ions are formed.*

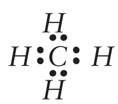
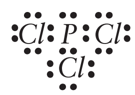
3.
(i) *Brittle – Ionic substances are made up of oppositely charged ions held together by strong electrostatic forces. Layers of these ions are unable to move over other layers without causing separation and hence fracture.*
(ii) *High M.P. – As above, strong electrostatic forces prevent the oppositely charged ions from separating unless they have sufficiently high kinetic energy (high temperature).*
(iii) *Conduction – When solid, ionic substances cannot conduct electricity since all ions are in fixed lattice positions. When liquid or in solution all the individual ions are free to carry charge as they are no longer part of a lattice.*

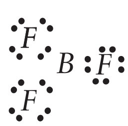
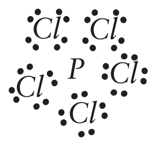
4. *Metals, essentially, are an array of positive ions in a sea of delocalised electrons. These electrons are easily influenced by any applied voltage and a flow of charge results.*

5. *Calcium would have a higher melting point than potassium. Calcium metal has twice as many delocalised electrons compared to potassium. These are attracted to doubly charged positive ions (Ca^{2+}). Hence electrostatic attractions in a calcium metal lattice would be much greater resulting in a higher M.P.*

6. *In the iron lattice it is possible for layers of positive ions to slide over each other within the 'sea' of electrons. However the presence of the larger cations in the lattice causes the layers to be less able to slide past each other. The larger impurity atoms act as locking pins to free movement. Hence less malleable.*

7.
(a) *Electrostatic attraction between the nuclei of the two atoms and the covalently bonded electrons.*
(b) *Two electrons are shared, one from each atom.*
(c) *Six pairs of non-bonded electrons, three on each atom.*

8.
(a) 
(c) 
(b) 
(d) 

9. (a) 
(b) 

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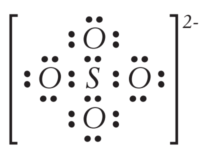
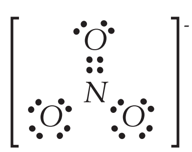
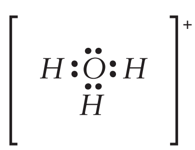
(a) 
(b) 
(c) 

11.
(a) NaCl – ionic
(b) Pb – metallic
(c) $PbCl_2$ – ionic
(d) HNO_3 – covalent
(e) SO_2 – covalent
(f) NH_4NO_3 – covalent and ionic

12.
(a) Bonding in both diamond and graphite is strong covalent bonding. High temperatures are needed to break these bonds and cause melting to occur.
(b) Diamond is a poor conductor because all valence electrons are localised in strong covalent bonds. This also explains the hardness of diamond. Graphite is a good conductor because one electron per atom is delocalised and capable of moving throughout the lattice. The softness of graphite is explained in 12(c).
(c) Carbon atoms in graphite bond in such a way as to form sheets of carbon atoms. The attractive force between these sheets is weak and they are able to slide over each other. This gives graphite lubricant properties.

13.
Ionic – magnesium oxide, lithium bromide, lead (II) nitrate.
Metallic – iron, mercury
Covalent molecular – carbon dioxide, dinitrogen tetroxide, fluorine
Covalent network – silicon dioxide, diamond

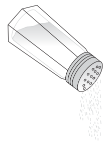
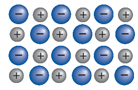
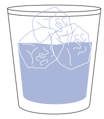
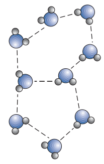
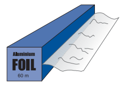
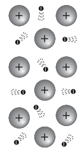
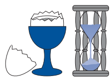
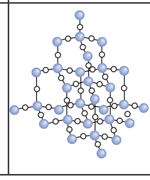
14.
(a) nitrogen monoxide (also called nitric oxide)
(b) nitrogen dioxide
(c) dinitrogen monoxide (also called nitrous oxide)
(d) dinitrogen trioxide
(e) dinitrogen tetraoxide
(f) dinitrogen pentoxide

15.
(a) PCl_5 (d) CF_4
(b) P_2O_5 (e) SCl_2
(c) NF_3 (f) H_2O

16.
(a) $CuSO_3$ (g) NaH_2PO_4
(b) $Co(NO_3)_2$ (h) Na_2HPO_4
(c) MnO (i) $Fe(OH)_2$
(d) $CrCl_3$ (j) $KMnO_4$
(e) $AlPO_4$ (k) FeO
(f) $CaHPO_4$ (l) $Na_2C_2O_4$

For the Experts

17.

SUBSTANCE	TYPE OF LATTICE (OR BONDING) STRUCTURE	PARTICLES WITHIN THE LATTICE	SIMPLE DIAGRAM OF LATTICE
salt 	ionic	ions Na^+, Cl^-	
ice 	covalent molecular	molecules	
aluminium 	metallic	positive ions and electrons	
sand 	covalent network	atoms	

CHP 4: CARBON CHEMISTRY

Chapter Questions

4.1
Sharing 4 of its electrons with 4 electrons from other atoms.

4.2
(a) Oxygen, nitrogen and the halogens
(b) Carbon
(c) Carbon, nitrogen

4.3

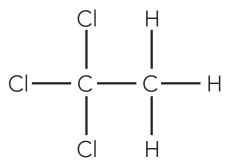
NUMBER OF C ATOMS IN CHAIN	ALKANE	ALKENE	ALKYNE	ALKYL GROUP
suffix	**-ane**	**-ene**	**-yne**	**-yl**
1	methane	-	-	methyl
2	ethane	ethene	ethyne	ethyl
3	propane	propene	propyne	propyl
4	butane	butene	butyne	butyl
5	petane	pentene	pentyne	pentyl
6	hexane	hexene	hexyne	hexyl
7	heptane	heptene	heptyne	heptyl
8	octane	octene	octyne	octyl

4.4
(a) ethane
(b) pentane

(c) *l-chloropropane*
(d) *3,5-dimethyloctane*
(e) *dimethylpropane*
(f) *hexabromoethane*
(g) *but-l-ene*
(h) *hept-l-yne*
(i) *1,3-dibromo-4-chlorohex-2-ene*
(j) *4-ethyl-2,2-dimethylhexane*
(k) *2-ethyl-3-methylpent-l-ene*
(l) *3-chloro-1.3-difluoropropyne*

4.5

(a)



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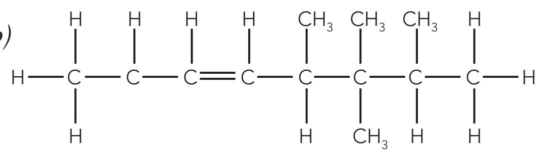
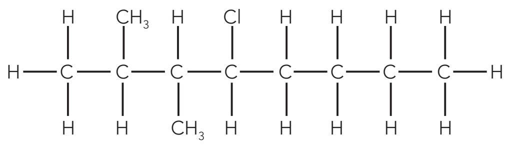


(c)



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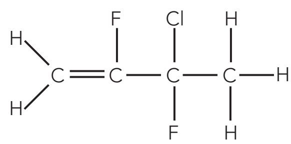
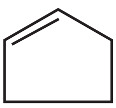
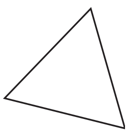


4.6

(a) *alkenes (or cycloalkanes)* (b) *alkanes*
(c) *alkynes*

4.7

(a)  (b) 

4.8

(a) $C_4H_{10} + F_2 +$ *UV light* $\rightarrow C_4H_9F + HF$
substitution
(b) $C_2H_4 + Cl_2 \rightarrow C_2H_2Cl_2$
addition
(c) $C_3H_6 + H_2 \rightarrow C_3H_8$ *(propane)*
addition
(d) $C_7H_{16} + 11O_2 \rightarrow 7CO_2 + 8H_2O$
combustion
(e) $C_5H_{12} + I_2 +$ *UV light* $\rightarrow C_5H_{11}I + HI$
substitution
(f) $2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O$
combustion

4.9

Alkenes and alkynes react with halogens without the need of a catalyst or additional energy whereas alkanes will only react with halogens when additional energy and/or a catalyst is used.

The reactions of cyclohexane and cyclohexene are examples of this.

4.10

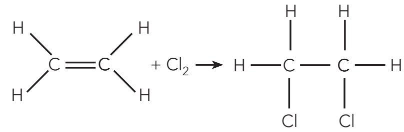
Natural Gas – methane
LPG – propane
LNG – methane
Kerosene – $C_{12}H_{26}$ *to* $C_{16}H_{34}$
Petrol – heptane/octane

4.11

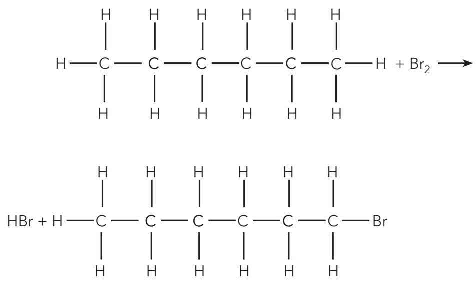
(a) *1-pentene +* Cl_2 (b) *2-butene +* Br_2
(c) *heptane +* F_2 (d) *ethyne +* I_2

4.12

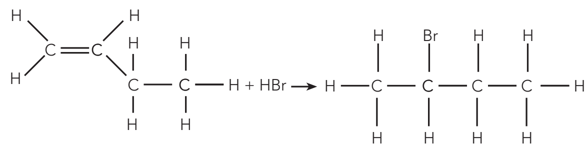
(a)



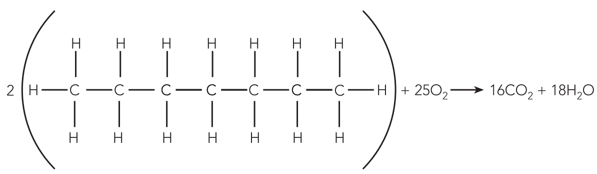
(b)



(c)



(d)



4.13

(a) *1-chloropentane*
(b) *2-chloropentane*
(c) *3-chloropentane*
(d) *1-chloro-2-methylbutane*
(e) *1-chloro-2,2-dimethyl propane*

4.14

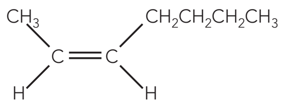
No – it is an alternate way of drawing 2-chloropentane.

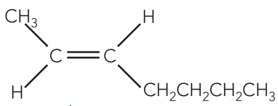
4.15

Yes – they both have the molecular formula C_6H_{12} *but have different structures.*

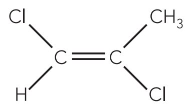
4.16
(a) trans 1,2-difluoropropene
(b) trans 2,3-diiodobut-2-ene
(c) cis 2,3-dibromohex-2-ene
(d) 1,1-dichloroethene

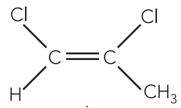
4.17
(a)


cis


trans

(b)


trans


cis

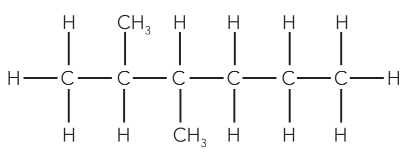
Review Questions

1. (a) propane
   (b) pentane
   (c) 4-methylpent-1-ene
   (d) trans hex-2-ene
   (e) 5-ethyl-2,7,7-trimethyloct-3-yne
   (f) tetramethylbutane

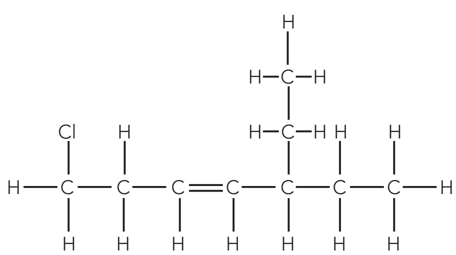
2. (a) tetrachloromethane
   (b) 1,2,3 – tribromo-3-fluoropentane
   (c) 1,2-diiodohept-3-ene
   (d) 6,6-dibromo-1,1-diiodohex-2-ene
   (e) 4,4-dibromo-1,1,1,5,5-pentachlorooctane
   (f) trans 1,2-difluoroethene

3. (a) cyclohexane
   (b) 1-methylcyclopentene
   (c) cyclohexene
   (d) benzene
   (e) 4-bromo-1-methylcyclohexene
   (f) 1,2,3,4,5,6-hexachlorocyclohexane
   (g) 1,1-diiodo-3-methylpentane
   (h) 1-bromo-2,3,3-trimethylbut-1-ene
   (i) 3,3-diethyl-2,2,4,4-tetramethyl pentane
   (j) 1,2,5-tribromo-5-ethyloct-3-ene

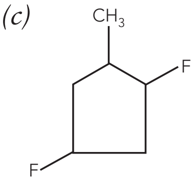
4. (a)



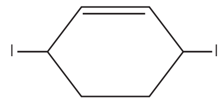
(b)



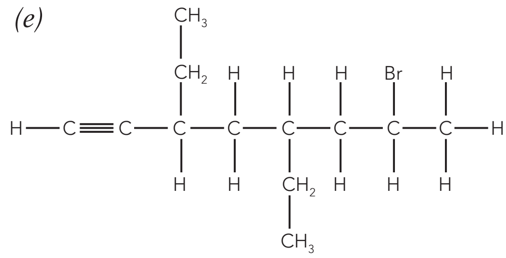
(c)



(d)

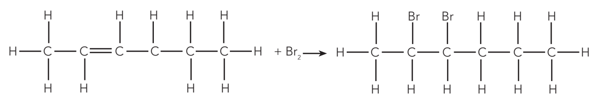


(e)

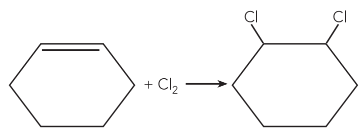


5.
(a) carbon dioxide and water
(b) 1,2-dichlorobutane
(c) 2,3-dibromohex-2-ene OR 2,2,3,3-tetrabromohexane
(d) fluoroheptane + hydrogen fluoride (more information needed to number where the fluorine attaches)
(e) chloroethane

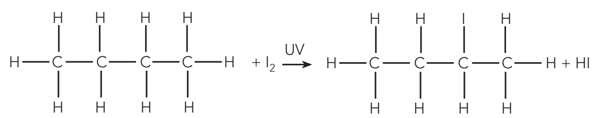
6. (a) $C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$
   (b) $C_8H_{16} + 12O_2 \rightarrow 8CO_2 + 8H_2O$
   (c)



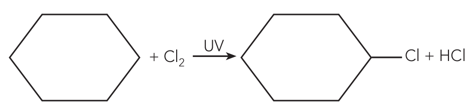
(d)



(e)



(f)



7. (a) Br_2 + cycloheptene
   (b) Cl_2 + pentane
   (c) HBr + propene
   (d) Br_2 + propane
   (e) I_2 + ethyne

8.

(a) $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$

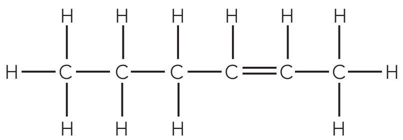
$2C_4H_{10(g)} + 13O_{2(g)} \rightarrow 8CO_{2(g)} + 10H_2O_{(g)}$

$2C_8H_{18(l)} + 25O_{2(g)} \rightarrow 16CO_{2(g)} + 18H_2O_{(g)}$

(b) *Each reaction produces carbon dioxide, water and energy.*

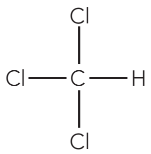
(c) *Methane requires the most oxygen per atom of carbon. (Incidentally, methane also produces the most energy per unit mass of the three fuels).*

9. (a)



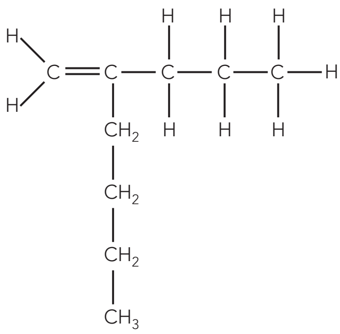
hex-2-ene

(b)



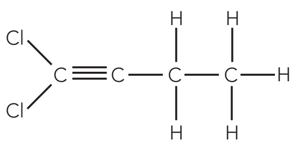
trichloromethane

(c)



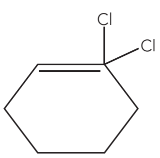
2-propylhex-1-ene

(d)



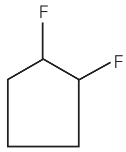
First carbon has 5 bonds which is not possible.

(e)



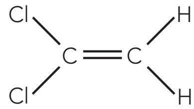
Carbon with Cl attached has 5 bonds – not possible.

(f)



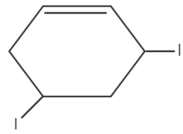
1,2-difluorocyclopentane

(g)



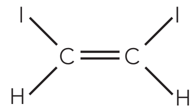
1,1-dichloroethene – not a geometric isomer

(h)

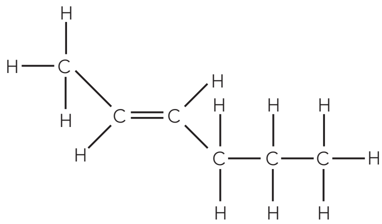


3,5-diiodocyclohexene

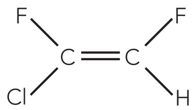
10.(a)



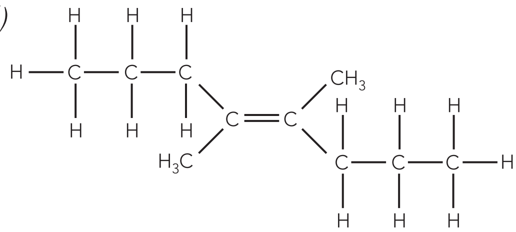
(b)



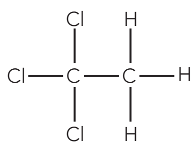
(c)



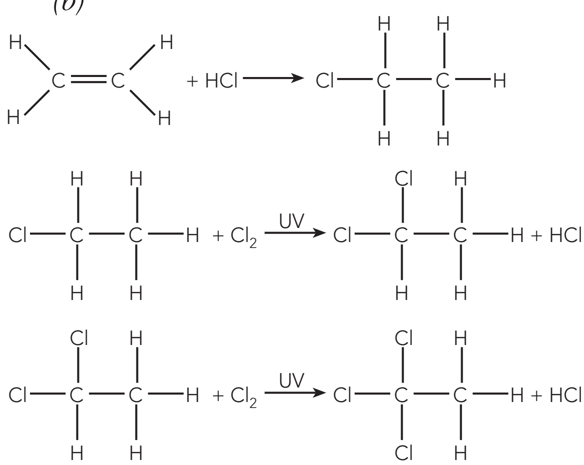
(d)



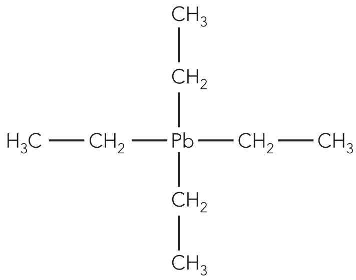
11. (a)



(b)

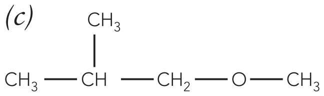


12. *(a)*

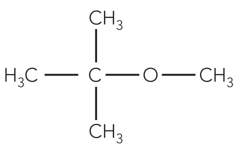


(b) *The combustion of tetraethyl lead introduced Pb as a pollutant from motor vehicles.*

(c)

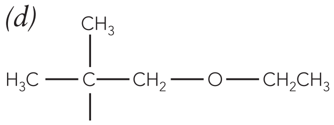


one possible answer

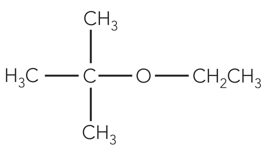


is another possible answer

(d)



one possible answer

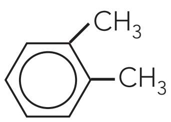


is another possible answer

(e) $2C_5H_{12}O + 15O_2 \rightarrow 10CO_2 + 12H_2O$

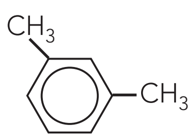
$C_6H_{14}O + 9O_2 \rightarrow 6CO_2 + 7H_2O$

(f) xylene:



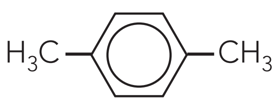
1,2-dimethylbenzene

OR



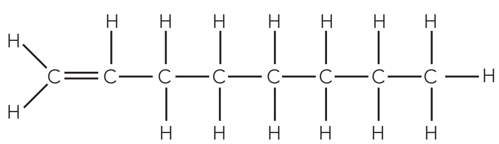
1,3-dimethylbenzene

OR



1,4-dimethylbenzene

olefin:



oct-l-ene is one of many possible answers.

(g) $2C_8H_{10} + 21O_2 \rightarrow 16CO_2 + 10H_2O$

$C_8H_{16} + 12O_2 \rightarrow 8CO_2 + 8H_2O$

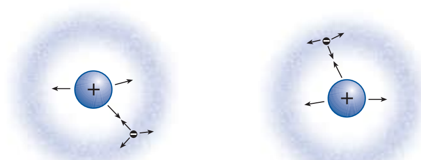
CHP 5: CHEMICAL REACTIONS AND ENERGY CHANGE

Chapter Questions

5.1 *(a) physical*
(b) chemical
(c) chemical
(d) physical
(e) chemical
(f) chemical
(g) physical
(h) chemical.

5.2
(i) released/exothermic
(ii) released/exothermic
(iii) absorbed/endothermic
(iv) released/exothermic
(v) released/exothermic
(vi) absorbed/endothermic
(vii) absorbed/endothermic
(viii) released/exothermic
(ix) absorbed/endothermic

5.3
(a)



(b)
(i) exothermic *(ii) decreased*
(iii) temperature increases because (potential energy) is given up

5.4
(i) ΔH *= H products – H reactants*
(ii)
- ΔH *is -(ve)*
- H_P *is less than* H_R
- *given up to or released to*

(iii)
- ΔH *is +(ve)*
- H_P *is greater than* H_R
- *absorbed from*

5.5
(a) ΔH *is –(ve), temp increases*
(b) ΔH *is +(ve), temp decreases*
(c) ΔH *is +(ve), temp decreases*
(d) ΔH *is –(ve), temp increases*

5.6
(a) $S(s) + O_2(g) \rightarrow SO_2(g) + 297\ kJ$
(b) $CuCO_3(s) + 46.0\ kJ \rightarrow CuO(s) + CO_2(g)$

5.7
(a) *Reaction (i) is a physical change and only relatively weak physical forces are involved.*
(b) *In (ii)* H_2O *is gaseous while in (iii) it is a liquid. The difference in* ΔH *represents latent heat of vaporisation.*

5.8

(a) Six bonds form: 2 C=O bonds and 4 O–H bonds

(b) The ΔH for the reaction is negative. This means a greater amount of energy was given up in the formation of new bonds than was used in breaking up the bonds of the reactants.

(c) $M(CH_4) = 12.01 + (4)(1.008) = 16.04\ g\ mol^{-1}$

Moles of CH_4 in 1.0 kg = 1000/16.04 = 62.34 mol

Energy produced per kg = - 891 × 62.34 = - 55548 kJ

Energy per kilogram of methane = 55.5 MJ

(d) The energy/kg for methane is higher by approximately 10% (55.5/50.3 = 1.10)

5.9

(a) Methane $n(CO_2)/n(CH_4) = 1$
Butane $n(CO_2)/n(CH_4) = 4$

(b) Methane $n(CO_2)$/1 MJ of energy $= 1 \times 10^6/891 \times 10^3 = 1.12$

Butane $n(CO_2)$/1 MJ of energy $= 4 \times 1\ 10^6/2878 \times 10^3 = 1.39$

(c) Methane gives more energy for equal quantities of CO_2. Methane has a higher hydrogen/carbon ratio than butane.

Review Questions

1.

(a) physical (d) chemical
(b) physical (e) chemical
(c) chemical (f) chemical.

2.

(a) Yes. Wax melting is a physical change.
(b) Yes. Wax burning is a chemical change.

3.

(a) methane + oxygen gas → carbon dioxide gas + water vapour.
(b) $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$

4.

(a) $2CO_{(g)} + O_{2(g)} \rightarrow 2CO_{2(g)}$
(b) $C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$
(c) $Ca_{(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_{2(g)}$
(d) $Ca(HCO_3)_{2(s)} \rightarrow CaCO_{3(s)} + CO_{2(g)} + H_2O_{(g)}$
(e) $Fe_2O_{3(s)} + 3CO_{(g)} \rightarrow 2Fe_{(s)} + 3CO_{2(g)}$
(f) $2H_2O_{2(aq)} \rightarrow O_{2\,(g)} + 2H_2O_{\,(l)}$
(g) $2Al_{\,(s)} + 6HCl_{(aq)} \rightarrow 2AlCl_{3\,(aq)} + 3H_{2\,(g)}$
(h) $2KHCO_{3(aq)} \rightarrow K_2CO_{3(aq)} + H_2O_{\,(l)} + CO_{2\,(g)}$

5.

(a) $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$
(b) $2CO_{(g)} + O_{2(g)} \rightarrow 2CO_{2(g)}$
(c) $Na_2CO_{3(aq)} + CaCl_{2(aq)} \rightarrow CaCO_{3(s)} + 2NaCl_{\,(aq)}$
(d) $2KOH_{(aq)} + H_2SO_{4(aq)} \rightarrow K_2SO_{4(aq)} + 2H_2O_{(l)}$
(e) $3MgO_{(s)} + 2H_3PO_{4(aq)} \rightarrow Mg_3(PO_4)_{2(s)} + 3H_2O_{(l)}$
(f) $Cu(OH)_{2(s)} + 2HCl_{(aq)} \rightarrow CuCl_{2\,(aq)} + 2H_2O_{(l)}$
(g) $3NaOH_{(aq)} + H_3PO_{4(aq)} \rightarrow Na_3PO_{4(aq)} + 3H_2O_{(l)}$
(h) $NH_4HCO_{3\,(aq)} + HNO_{3(aq)} \rightarrow NH_4NO_{3\,(aq)} + H_2O_{(l)} + CO_{2(g)}$

6.

PHYSICAL OR CHEMICAL CHANGE	PROCESS IS EXOTHERMIC / ENDOTHERMIC	ENTHALPY OF PRODUCTS (H) IS HIGHER / LOWER
propane gas (barbecue gas) is burnt	exothermic	lower
ice is placed in water and melts	endothermic	higher
the two atoms making up an oxygen molecule are separated	endothermic	higher
solid carbon dioxide (dry ice) sublimes to its gaseous form	endothermic	higher

7.

(a) enthalpy is the stored chemical energy of a substance (heat content).
(b) enthalpy decreases if heat is given up to the surroundings (exothermic reaction).

8.

(a) ΔH = -393 kJ
(b) ΔH = +132 kJ
(c) ΔH = +44 kJ

9.

(a) $H_{2(g)} + C1_{2(g)} \rightarrow 2HCl_{(g)} + 185\ kJ$
(b) $O_{2(g)} + ½O_{2(g)} + 143\ kJ \rightarrow O_{3(g)}$

10. Reaction (a) only involves a physical change. Reaction (b) is a chemical change involving the breaking and forming of very strong chemical bonds.

11.

(a) 4 C-H bonds, 2 O=O bonds
total 6 bonds
(b) 2 C=O bonds, 4 O-H bonds
total 6 bonds
(c) bonds of the products

12. *There are a greater number of bonds that have to be broken when larger hydrocarbon molecules are combusted. Hence octane burns more slowly than methane.*

13.
(a) $2C_2H_{2(g)} + 5O_{2(g)} \rightarrow 4CO_{2(g)} + 2H_2O_{(l)} + 2602\ kJ$
(b) *Eleven bonds break: 2 C≡C bonds, 4 C-H bonds and 5 O=O bonds*
Twelve bonds form: 8 C=O bonds, and 4 O-H bonds
(c) *The ΔH for the reaction is negative. This means a greater amount of energy was given up when new bonds were formed than was used in breaking up the bonds of the reactants.*

14.
(a) $2C_8H_{18(l)} + 25O_{2(g)} \rightarrow 16CO_{2(g)} + 18H_2O_{(l)}$
$\Delta H = -5471\ kJ\ mol^{-1}$

$2C_4H_{10(g)} + 13O_{2(g)} \rightarrow 8CO_{2(g)} + 10H_2O_{(l)}$
$\Delta H = -2878\ kJ\ mol^{-1}$
(b) $M(C_8H_{18}) = (8)(12.01) + (18)(1.008) = 114.22\ g\ mol^{-1}$

Moles of C_8H_{18} *in 1.0 kg* = 1000/114.22 = 8.755 mol

Energy produced per kg = - 5471 × 8.755 = - 47899 kJ

Energy per kilogram for octane = 47.9 MJ

Similarly *for butane the value is* = 49.5 MJ

(c) *Mole ratios*

Octane $n(CO_2)/n(C_8H_{18}) = 8$

Butane $n(CO_2)/n(CH_4) = 4$

$n(CO_2)$*/MJ of energy*

Octane $n(CO_2)$*/1 MJ of energy*
$= 8 \times 1 \times 10^6/5471 \times 10^3 = 1.46$

Butane $n(CO_2)$*/1 MJ of energy*
$= 4 \times 1 \times 10^6/2878 \times 10^3 = 1.39$

(d) *Butane gives more energy per kg than octane and its combustion emits less carbon dioxide. However the differences are fairly small and other factors influence their everyday use.*

For the Experts

15.
(a)
(i) $CaCO_{3(s)} + 179\ kJ \rightarrow CaO_{(s)} + CO_{2(g)}$.
(ii) *Endothermic reaction. The bonds broken are much stronger than the bonds formed during the reaction. The energy required to break the bonds is greater than the energy given up by the formation of the bonds of the products.*
(b)
(i) $CaO_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)} + 66\ kJ$
(ii) *Exothermic reaction. In this case bonds formed are stronger than bonds broken.*

CHP 6: CHEMISTRY CALCULATIONS

Chapter Questions

6.1
(a) (i) 107.9 / 12.01 ≈ 9
(ii) 83.80 / 12.01 ≈ 7
(b) (i) ~ 12
(ii) 55.85 / 1.008 ≈ 55.4
(c) 107.9 / 26.98 ~ 4
(d) Ca

6.2
(a) $M_r(SO_2) = 32.06 + 2(16.0) = 64.06$
(b) $M_r(CuSO_4) = 159.6$
(c) $M_r(Ca(OH)_2) = 74.10$
(d) $M_r(Al_2(SO_4)_3) = 342.14$
(e) $M_r(CH_3COOH) = 60.05$

6.3
(a) CO_2, $M_r = 44.01$
(b) $Mg(OH)_2$, $M_r = 58.32$
(c) Cl_2, $M_r = 70.90$
(d) NH_3, $M_r = 17.03$
(e) $Fe_2(SO_4)_3$, $M_r = 399.91$

6.4
(a) $M_r(CO) = 12.01 + 16.0 = 28.01$

$$\%\ C = \frac{12.01}{28.01} \times 100 = 42.9\%$$

$$\%\ O = \frac{16.0}{28.01} \times 100 = 57.1\%$$

(b) $M_r(NaCl) = 58.44$
% Na = 39.3%
% Cl = 60.7%

(c) $M_r(Ca(OH)_2 = 74.1$
% Ca = 54.1%
% O = 43.2%
%H = 2.7%

6.5
(a) $M_r(Na_2CO_3 . 10H_2O) = 286.15$

$$\%\ H_2O = \frac{180.16}{286.15} \times 100 = 62.96\%$$

(b) M_r ($BaCl_2 . 2H_2O$) = 244.23

$$\% H_2O = \frac{36.03}{244.23} \times 100 = 14.75\%$$

(c) M_r ($MgSO_4 . 7H_2O$) = 246.48

$$\% H_2O = \frac{126.11}{246.48} \times 100 = 51.17\%$$

6.6

(a) M_r (Al_2O_3) = 101.96
% Al = 52.9%

(b) M_r ($CaCO_3$)= 100.09
% Ca = 40.04%

(c) M_r (Fe_2O_3) = 159.7
% Fe = 69.9%

6.7

(a) 6.022×10^{23} atoms of H

(b) 1 mol H_2 contains 2 mol H atoms
= 1.204×10^{24} atoms of H

(c) 3 mol H_2O contains 6 mol H atoms
= 3.613×10^{24} atoms of H

(d) 0.2 mol NH_3 contains 0.6 mol H atoms
= 3.613×10^{23} atoms of H

(e) 4 mol $C_6H_{12}O_6$ contains 48 mol H atoms = 2.89×10^{25} atoms of H

6.8

(a) Number of 5¢ coins in $1,000,000.00

$$= \frac{1.0 \times 10^6}{0.05} = 2.0 \times 10^7 \text{ coins}$$

(20 million coins)

$$n = \frac{N}{6.022 \times 10^{23}} = \frac{2.0 \times 10^7}{6.022 \times 10^{23}}$$

$= 3.32 \times 10^{-17}$ mol

(b) $$1 \text{ mol } 5¢ \text{ coins} = \frac{6.022 \times 10^{23}}{20} \text{ dollars}$$

= $\$3.01 \times 10^{22}$ dollars

6.9

(a) $n\ (Na) = \dfrac{m}{M} = \dfrac{100}{22.99} = 4.35$ mol

(b) $n\ (H_2O) = \dfrac{m}{M} = \dfrac{30}{18.016} = 1.66$ mol

(c) $n\ (CH_4) = \dfrac{m}{M} = \dfrac{4.00}{16.04} = 0.249$ mol

6.10

(a) $n\ (NaCl) = \dfrac{m}{M} = \dfrac{10.0}{58.44} = 0.171$ mol

(b) $n\ (H_2) = \dfrac{m}{M} = \dfrac{2.4 \times 10^{-2}}{2.016}$

$= 1.19 \times 10^{-2}$ mol

(c) $n\ (C_8H_{18}) = \dfrac{m}{M} = \dfrac{1000}{114.22} = 8.76$ mol

6.11

(a) M ($Ca(OH)_2$) = 74.1 g mol^{-1}
m ($Ca(OH)_2$) = n .M = (2.0) (74.1) = 148.2 g

(b) M (H_2O) = 18.016 g mol^{-1}
m (H_2O) = (10) (18.016) = 180.2 g

(c) M ($(NH_4)_2CO_3$) = 96.09 g mol^{-1}
m ($(NH_4)_2CO_3$) = (0.40) (92.06) = 38.44 g

6.12

(a) n (NH_3) molecules = $\dfrac{100}{17.034}$ = 5.87 mol of NH_3

(b) n (O_2) molecules = $\dfrac{200}{32.0}$ = 6.25 mol of O_2

Answer: 200 g of oxygen gas has the most molecules.

6.13

(a) twelve, six, twenty four (b) 8 , 24 , 4

6.14

(a) $n\ (Ca(OH)_2) = \dfrac{150.0}{74.1} = 2.024$ mol

(b) $n\ (OH^-) = (2)(2.024) = 4.048$ mol

(c) n (O) = (2)(2.024) = 4.048 mol

(d) m (Ca) = 81.12 g

6.15

(a) $n\ (H_2O) = \dfrac{n}{M} = \dfrac{8.60 \times 10^{-2}}{18.016}$

$= 4.77 \times 10^{-3}$ mol

N (H_2O) molecules = $(n)(6.022 \times 10^{23})$
$= 2.87 \times 10^{21}$

(b) $n\ (Pb) = \dfrac{n}{M} = \dfrac{10.0}{207.2} = 4.83 \times 10^{-2}$ mol

N (Pb) atoms = $(n)(6.022 \times 10^{23})$
$= 2.91 \times 10^{22}$ atoms

6.16

(a) propane gas + oxygen gas → carbon dioxide gas + water vapour

(b) $C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3\ CO_{2(g)} + 4\ H_2O_{(g)}$

(c) 1:5:3:4

(d) Yes. The number of atoms of each element is the same on both sides of the equation.

(e) n (H_2O) =(15.0)(4/3)= 20.0 mol

6.17

1. $2C_4H_{10(g)} + 13O_{2(g)} \rightarrow 8\ CO_{2(g)} + 10\ H_2O_{(g)}$
2. n (C_4H_{10}) =10.0/58.12 = 0.172 mol
3. n(H_2O) = (0.172)(10/2) = 0.860 mol
4. m(H_2O) = (0.860)(18.016) = 15.5 g

6.18

(a) $2CO_{(g)} + O_{2(g)} \rightarrow 2\ CO_{2(g)}$

(b) 2:1:2

(c) 2:1:2

(d) $n\ (O_2)$ required = (300)(1/2) = 150 mL

Review Questions

RELATIVE MASSES

1. (a) titanium (b) krypton (c) iodine
2. (a) 4 (b) 10 (c) 12
3. (a) 44.01 (b) 32.0 (c) 64.06 (d) 17.03

4.

(a) $M_r\ (Ca(NO_3)_2)$ = 164.1

(b) $M_r\ (NaOH)$ = 40.0

(c) $M_r\ ((NH_4)_2CO_3)$ = 96.09

(d) $M_r\ (FeCl_3)$ = 162.2

5.

(a) CH_3COOH is heaviest

(b) CH_3COOH

(c) NH_3

PERCENTAGE COMPOSITION

6.

(a) $M_r\ (Al_2O_3)$ = 101.96
% Al = 52.9%, % O = 47.1%

(b) $M_r\ (H_2O)$ = 18.016
% H = 11.2%, % O = 88.8%

(c) $M_r\ (NaCl)$ = 58.44
% Na = 39.3%, % Cl = 60.7%

(d) $M_r\ (C_3H_8)$ = 44.09
% C = 81.7%, % H = 18.3%

7.

(a) $M_r\ (CaSO_4 \cdot 2H_2O)$ = 172.17

$$\%\ H_2O = \frac{36.032}{172.17} \times 100 = 20.9\%$$

(b) $M_r\ (FeCl_3 \cdot 3H_2O)$ = 216.25

$$\%\ H_2O = \frac{54.048}{216.25} \times 100 = 25.0\%$$

8.

(a) C_2H_2 – highest % C by mass

(b) CH_4 – highest % H by mass

MOLES / PARTICLES

9.

(a) $n\ (Na) = \dfrac{6.022 \times 10^{22}}{6.022 \times 10^{23}} = 0.10$ mol

(b) $n\ (Cu) = \dfrac{2.41 \times 10^{24}}{6.022 \times 10^{23}} = 4.0$ mol

(c) $n\ (O_2) = \dfrac{1.20 \times 10^{23}}{6.022 \times 10^{23}} = 0.199$ mol

10.

(a) n (Ca) = $(0.25)(6.022 \times 10^{23}) = 1.5 \times 10^{23}$ atoms of Ca

(b) n (CO) = $(5.50)(6.022 \times 10^{23}) = 3.31 \times 10^{24}$ molecules of CO

(c) n (H) = $(2)(4.0)(6.022 \times 10^{23}) = 4.82 \times 10^{24}$ atoms of H

(d) 1.25 mol HNO_3 contains 3 × 1.25 mol of O atoms = $(3)(1.25)(6.022 \times 10^{23}) = 2.26 \times 10^{24}$ atoms of O

11.

(a) 3 mol H_2 → 6 mol of atoms

(b) 8 mol Pb → 8 mol of atoms

(c) 2 mol CO_2 → 6 mol of atoms

(d) 1 mol NH_4NO_3 → 9 mol of atoms
Answer – 1 mol NH_4NO_3 contains greatest number of atoms.

12.

(a) Silver (b) $\dfrac{108}{36} = 3$

MOLES / MASS / PARTICLES

13.

(a) $M\ (Cl_2)$ = 70.90 g mol^{-1}

(b) $M\ (CO_2)$ = 44.01 g mol^{-1}

(c) $M\ (NH_3)$ = 17.03 g mol^{-1}

(d) $M\ (SO_3)$ = 80.06 g mol^{-1}

14.

(a) M (ZnO) = 81.38 g mol^{-1}

(b) $M\ (Mg(HCO_3)_2)$ = 146.3 g mol^{-1}

(c) $M\ (Al(NO_3)_3)$ = 213.01 g mol^{-1}

(d) $M\ (Na_2CO_3 \cdot 10H_2O)$ = 286.15 g mol^{-1}

15.

Relative molecular mass (M_r) of a molecule is its mass compared to 1/12 of the mass of a carbon 12 atom (there are no units). The molar mass however is the actual mass (in grams) of 1 mole of molecules of a substance.

16.

(a) $n\ (CO_2) = \dfrac{m}{M} = \dfrac{440}{44.01} = 10.0$ mol

(b) $n\ (H_2SO_4) = \dfrac{m}{M} = \dfrac{49.0}{98.086} = 0.50$ mol

(c) $n\ (C_3H_8) = \dfrac{m}{M} = \dfrac{8500}{44.09} = 192.8$ mol

17.

(a) $n\ (H_2O) = \dfrac{m}{M} = \dfrac{500}{18.016} = 27.8$ mol

(b) n (AgCl) $= \dfrac{m}{M} = \dfrac{2.15}{143.35} = 1.50 \times 10^{-2}$ mol

(c) $n\ (CH_3COOH) = \dfrac{m}{M} = \dfrac{100}{60.05} = 1.66$ mol

18.

$n\,(O_2) = \dfrac{20.0}{32} = 0.625\ mol$

$n\,(CO_2) = \dfrac{20}{44.0} = 0.455\ mol$

Answer – O_2 contains greatest number of molecules

19.

H_2 (2.016 g mol^{-1}), NH_3 (17.03 g mol^{-1}), CO (28.01 g mol^{-1}), N_2 (28.02 g mol^{-1}), O_2 (32.0 g mol^{-1}), CO_2 (44.0 g mol^{-1}), SO_3 (80.07 g mol^{-1})

20.

(a) 1.0 mol H_2 gas → contains 2.0 mol H

(b) 4.0 mol NH_3 gas → contains 12.0 mol H

(c) 2.0×10^{-2} mol CH_3COOH → contains 8.0×10^{-2} mol H

21.

(a) 1.50 mol HNO_3 contains
→ 1.50 mol H
→ 1.50 mol N
→ 4.50 mol O

(b) $50.0\ g\ HCl = \dfrac{50.0}{36.46}\ mol = 1.37\ mol$

∴ contains 1.37 mol H and 1.37 mol Cl

(c) $100.0\ g\ C_6H_{12}O_6 = \dfrac{100.0}{180.156}\ mol = 0.555\ mol$

∴ contains 6 × 0.555 mol of C = 3.33 mol
contains 6 × 0.555 mol of O = 3.33 mol
contains 12 × 0.555 mol of H = 6.66 mol

22.

(a) $n\,(PbCl_2) = \dfrac{n}{M} = \dfrac{2.00}{278.1} = 7.19 \times 10^{-3}\ mol$

$n\,(Pb^{2+}) = 7.19 \times 10^{-3}\ mol$

$n\,(Cl^-) = (2)\,(7.19 \times 10^{-3}) = 1.44 \times 10^{-2}\ mol$

(b) $n\,(Fe_2O_3) = \dfrac{2000}{159.7} = 12.5\ mol$

$n\,(Fe^{3+}) = (2)\,(12.5) = 25.0\ mol$

$n\,(O^{2-}) = (3)\,(12.5) = 37.5\ mol$

(c) $n\,(Al_2(SO_4)_3) = \dfrac{5.25 \times 10^{-3}}{342.17} = 1.53 \times 10^{-5}\ mol$

$n\,(Al^{+3}) = (2)\,(1.53 \times 10^{-5}) = 3.06 \times 10^{-5}\ mol$

$n\,(SO_4^{-2}) = (3)\,(1.53 \times 10^{-5}) = 4.60 \times 10^{-5}\ mol$

23.

(a) $n\,(H_2O) = \dfrac{10.0}{18.016} = 0.555\ mol$

∴ Number of atoms = $(3)(0.555)(6.022 \times 10^{23})$
$= 1.00 \times 10^{24}$ atoms

(b) $n\,(H_2) = \dfrac{10.0}{2.016} = 4.96\ mol$

∴ Number of atoms = $(2)(4.96)(6.022 \times 10^{23})$
$= 5.97 \times 10^{24}$ atoms

(c) $n\,(SO_2) = \dfrac{20.0}{64.07} = 0.312\ mol$

∴ Number of atoms = $(3)(0.312)(6.022 \times 10^{23})$
$= 5.64 \times 10^{23}$ atoms

24.

(a) $n\,(Cl_2) = \dfrac{90.0}{70.90} = 1.269\ mol$

∴ Number of molecules of Cl_2 = $(1.269)\,(6.022 \times 10^{23})$
$= 7.64 \times 10^{23}$ molecules

(b) $n\,(H_2) = \dfrac{12.0}{2.016} = 5.95\ mol$

∴ Number of molecules of H_2 = $(5.95)(6.022 \times 10^{23})$
$= 3.58 \times 10^{24}$ molecules

(c) $n\,(C_2H_2) = \dfrac{12.0}{26.04} = 0.461\ mol$

∴ Number of molecules of C_2H_2
$= (0.461)\,(6.022 \times 10^{23})$
$= 2.78 \times 10^{23}$ molecules

25.

(a) $n\,(NH_3) = \dfrac{N}{6.022 \times 10^{23}} = \dfrac{6.022 \times 10^{23}}{6.022 \times 10^{23}}$
$= 1.0\ mol$

∴ $m\,(NH_3) = n\,.\,M = (1.00)\,(17.03) = 17.03\ g$

(b) $n\,(Na) = \dfrac{3.01 \times 10^{24}}{6.022 \times 10^{23}} = 5.00\ mol$

∴ $m\,(Na) = (5.00)\,(22.99) = 114.91\ g$

(c) $n\,(Mg(OH)_2) = \dfrac{4.0 \times 10^{24}}{6.022 \times 10^{23}} = 6.64\ mol$

∴ $m\,(Mg(OH)_2) = (6.64)\,(58.326) = 387.4\ g$

CALCULATIONS BASED ON EQUATIONS

mole/mole

26.

(a) $2C_2H_{6\,(g)} + 7O_{2(g)} \rightarrow 4CO_{2(g)} + 6H_2O_{\,(g)}$

(b) coefficients indicate mole ratios.

(c) $n(CO_2) = (25)(4/2) = 50\ mol.$

27.

(a) $n(O_2) = (0.25)(25/2) = 3.125\ mol$

(b) $n(CO_2) = (0.25)(16/2) = 2.0\ mol$
$n(H_2O) = (0.25)(18/2) = 2.25\ mol$

$\therefore$ *total n(products) = 4.25 mol.*

28.

$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

? 1.65×10^{-3} mol

$n(O_2) = (1.65 \times 10^{-3})(1/2)$

$= 8.25 \times 10^{-4}$ mol.

mole/mass or mass mole

29.

$2C_4H_{10(g)} + 13O_{2(g)} \rightarrow 8CO_{2(g)} + 10H_2O_{(g)}$

1.50×10^{-2} mol

$n(H_2O) = (1.50 \times 10^{-2})(10/2) = 7.50 \times 10^{-2}$ mol

$\therefore m(H_2O) = (7.50 \times 10^{-2})(18.016) = 1.35$ g

30.

$n(CuCO_3) = (5.46)(1/1) = 5.46$ mol

$\therefore m(CuCO_3) = (5.46)(123.55) = 675$ g

31.

(a) $n(CO_2) = (2.4 \times 10^{-3})(1/1) = 2.4 \times 10^{-3}$ mol

$m(CO_2) = (2.4 \times 10^{-3})(44.01) = 0.106$ g

(b) $n(HCl) = (1.50)(2/1) = 3.00$ mol

$m(HCl) = (3.00)(36.458) = 109$ g

(c) $n(CaCl_2) = (6.40 \times 10^{-2})(1/2) = 3.20 \times 10^{-2}$ mol

$m(CaCl_2) = (3.20 \times 10^{-2})(110.98) = 3.55$ g

32.

(a) $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2\,(aq)} + H_{2(g)}$

4.26 g ?

(b) $n(Mg) = \dfrac{4.26}{24.31} = 0.175$ mol

$\therefore n(HCl) = (0.175)(2/1) = 0.350$ mol

(c) $n(MgCl_2) = (0.175)(1/1) = 0.175$ mol

33.

(a) $n(C_6H_{12}O_6) = \dfrac{1.0}{180.2} = 5.55 \times 10^{-3}$ mol

$n(CO_2) = (5.55 \times 10^{-3})(6/1) = 3.33 \times 10^{-2}$ mol

(b) $n(O_2) = (5.55 \times 10^{-3})(6/1) = 3.33 \times 10^{-2}$ mol

mass/mass

34.

$n(Cu) = \dfrac{14.5}{63.55} = 0.228$ mol

$\therefore n(Ag) = (0.228)(2/1) = 0.456$ mol

$\therefore m(Ag) = (0.456)(107.9) = 49.2$ g

35.

(a) $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

9.0 kg ?

(b) $n(C_3H_8) = \dfrac{9000}{44.09} = 204.1$ mol

$\therefore n(CO_2) = (204.1)(3/1) = 612.4$

$\therefore m(CO_2) = (612.4)(44.01) = 26951$ g

also $n(H_2O) = (204.1)(4/1) = 816.5$

$\therefore m(H_2O) = (816.5)(18.016) = 14710$ g

$\therefore$ mass of products = 41661 g

= 41.7 kg

36.

(a) $n((NH_4)_2CO_3) = \dfrac{6.42}{96.094} = 6.681 \times 10^{-2}$ mol

$\therefore n(NH_3) = (6.681 \times 10^{-2})(2/1) = 0.1336$ mol

$\therefore m(NH_3) = (0.1336)(17.034) = 2.28$ g

(b) mass of products = mass of reactants = 6.42 g

37.

(a) $n(CuFeS_2) = \dfrac{1.0 \times 10^3 \times 10^3\ g}{183.5} = 5448$ mol

$n(CuS) = (5448)(1/1) = 5448$ mol

$\therefore m(CuS) = (5448)(95.61) = 5.209 \times 10^5$ g

≈ 0.521 tonne

(b) $mCu = (5.21 \times 10^5)\left(\dfrac{63.55}{95.61}\right) = 3.462 \times 10^5$ g

gases

38.

(a) Vol (O_2) needed = (200)(5/4) = 250 L

(b) Vol (NH_3) needed = (10,000)(4/4) = 10,000 L

39.

(a) $2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O$

(b) Vol (O_2) needed = (10.0)(5/2) = 25.0 L

40.

(a) $n(NH_3) = \dfrac{25000}{22.71} = 1.101 \times 10^3$ mol

$\therefore n((NH_4)_2SO_4) = (1.101 \times 10^3)(1/2)$

= 550 mol

$\therefore m((NH_4)_2SO_4) = (550)(132.15)$

$= 7.27 \times 10^4$ g

(b) $n(H_2SO_4) = (1.1136 \times 10^3)(1/2) = 550$ mol

$m(H_2SO_4) = (550)(98.086) = 5.40 \times 10^4$ g

41.

$2Al_{(s)} + 6HCl \rightarrow 2AlCl_3 + 3H_{2\,(g)}$

500 mL

$n(H_2) = \dfrac{0.500}{22.71} = 2.22 \times 10^{-2}$ mol

$\therefore n(Al) = (0.0222)(2/3) = 1.468 \times 10^{-2}$ mol

$\therefore m(Al) = (1.468 \times 10^{-2})(26.98) = 0.396$ g

solutions

42.

$Pb(NO_3)_{2\,(aq)} + H_2SO_{4\,(aq)} \rightarrow PbSO_{4\,(s)} + 2HNO_{3\,(aq)}$

240 ml ?

0.150 mol L^{-1}

$n(Pb(NO_3)_2) = (0.150)(0.240) = 0.036$ mol

$\therefore m(PbSO_4) = (0.036)(303.27) = 10.9$ g

43.

(a) $n(HNO_3) = (1.20)(0.150) = 0.180$ mol
$n(Na_2CO_3)$ consumed $= (0.180)(1/2)$
$= 0.090$ mol
$\therefore$ $m(Na_2CO_3)$ consumed $= (0.090)(105.99)$
$= 9.54$ g

(b) by proportion – to react all 12.2 g Na_2CO_3 we would need a total volume of :
$(150 \text{ mL}) \times \dfrac{12.2}{9.54} = 192$ mL of HNO_3

$\therefore$ require a further 42 mL of HNO_3

44.

(a) $Zn(s) + 2NO_3^-(aq) + 4H^+(aq) \rightarrow Zn^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$
? 125 mL
6.0 mol L^{-1}
$n(H^+) = (6.0)(0.125) = 0.750$ mol
$\therefore n(Zn) = (0.750)(1/4) = 0.1875$ mol
$m(Zn) = (0.1875)(65.38) = 12.3$ g

(b) $n(Zn(NO_3)_2) = (0.1875)(1/1) = 0.1875$ mol
$m(Zn(NO_3)_2) = (0.1875)(189.4) = 35.5$ g

45.

(a) Yes, $BaCO_3$ will form.
$Na_2CO_{3\,(aq)} + BaCl_{2\,(aq)} \rightarrow BaCO_{3\,(s)} + 2NaCl_{(aq)}$
200 mL 200 mL
1.0 mol L^{-1} 1.0 mol L^{-1}

(b) $n(Na_2CO_3) = n(BaCl_2) = (1.0)(0.200) = 0.200$ mol
Hence no limiting reagent
$\therefore n(BaCO_3) = (0.200)\ (1/1) = 0.200$ mol
$\therefore m(BaCO_3) = (0.200)(197.3) = 39.5$ g

(c) $c(Ba^{2+}) = c(CO_3^{2-}) = 0$ All consumed

$c(Na^+) = \dfrac{\text{total moles}}{\text{total vol of solutions}}$

$= \dfrac{(0.200)(2)}{0.400} = 1.0$ mol L^{-1}

also $c(Cl^{-1}) = \dfrac{(0.200)(2)}{0.400} = 1.0$ mol L^{-1}

For the Experts

46.

(a) $M_r((NH_4)_2\,SO_4) = 132.15$

$\%\ N = \dfrac{28.02}{132.15} \times 100 = 21.2\%$

- $M_r(NH_4NO_3) = 80.5$

$\%\ N = \dfrac{28.02}{80.05} \times 100 = 35.0\%$

- $M_r((NH_4)_3PO_4) = 149.10$

$\%\ N = \dfrac{42.03}{149.10} \times 100 = 28.2\%$

- $M_r(CO(NH_2)_2) = 60.06$

$\%\ N = \dfrac{28.02}{60.06} \times 100 = 46.7\%$

Urea, $CO(NH_2)_2$, has the highest % of N by weight.

(b)

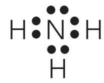
(i) Area of lawn $= (4.5)(12.0) = 54\ m^2$ approximately.
$\therefore$ mass of fertiliser needed $= (54)(145)$
$= 2430$ g
$= 2.43$ kg

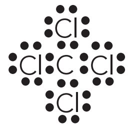
(ii) mass of N in 2.43 kg of $(NH_4)_2SO_4$
$= 2.43 \times \dfrac{21.2}{100}$

$= 0.515$ kg

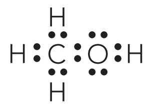
CHP 7: INTERMOLECULAR FORCES

Chapter Questions

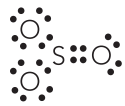
7.1 (a) 

(c) 

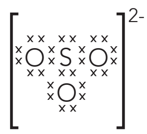
(b) 

(d) 

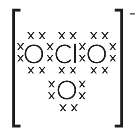
7.2 (a) 

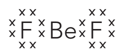
(d) 

(b) 

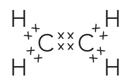
(e) 

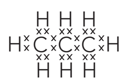
(c) 

(f) 

7.3 (a) 

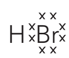
(b) 

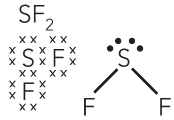
(c) 

(d) 

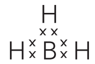
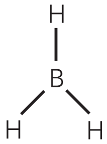
7.4 F-F , C-S , B-H , N-O , N-H , C-O , O-H , H-F

7.5 (a) done

(b) HBr
 H–Br linear

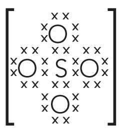
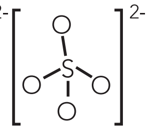
(c) SF_2
 bent

(d) BH_3

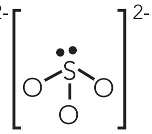
triangular planar

(e) SO_4^{2-}

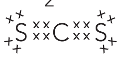
tetrahedral

(f) SO_3^{2-}

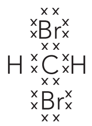
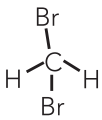
pyramidal

(g) CS_2



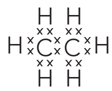
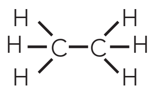

linear

(h) CH_2Br_2

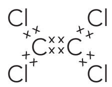
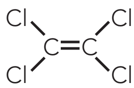
tetrahedral

(i) C_2H_6

tetrahedral

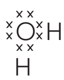
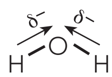
(j) C_2Cl_4

triangular planar

7.6

1.

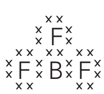
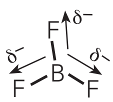




polar

2.

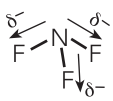

non-polar

3.

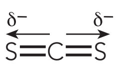
non-polar

4.

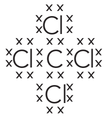
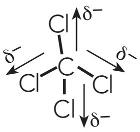

polar

5.

non-polar

6.

non-polar

7.7

(a) *Pentane as it has dispersion forces only whereas 1-pentanol has hydrogen bonding. The weaker the intermolecular forces the greater the vapour pressure.*

(b) *Ethanol, water has stronger hydrogen bonding than ethanol.*

(c) *Pentane, both have dispersion forces but as the smaller molecule pentane will have weaker dispersion forces.*

7.8

As non-volatile solute, the presence of Na^+ *and* Cl^- *ions at the surface will reduce the number of water molecules that at the surface of the liquid and so reduce vapour pressure.*

7.9

They would not be attracted to the molecules in the mobile phase and therefore not separate and move to different height up the stationary phase.

7.10

Molecules in component B have greater attraction for molecules in the stationary phase and do not move as far in the same time period.

7.11

Greater choice in the nature of the solid phase allows for greater range of mixtures to be separated and for greater degree of separation to occur in a shorter time. Paper not suitable for some corrosive materials.

7.12

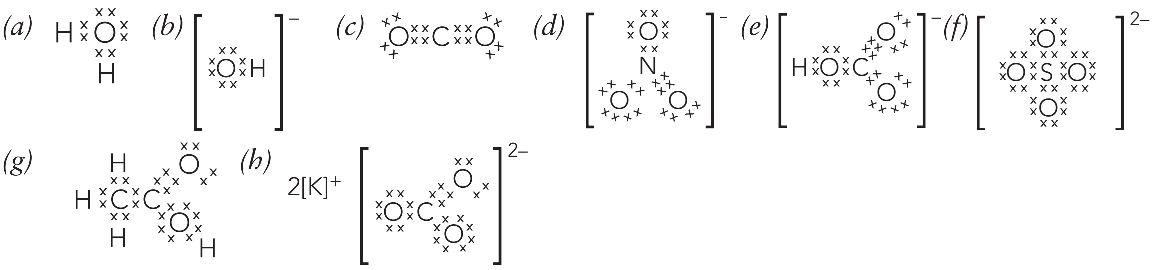
(a) *Paper or thin layer chromatography*

(b) *Gas chromatography*

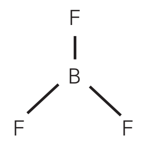
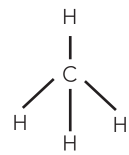
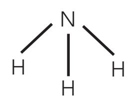
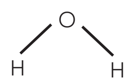
(c) *High Performance Liquid Chromatography*

Review Questions

1.



2.

	CENTRAL ATOM	NUMBER OF ELECTRON PAIRS	NUMBER OF LONE PAIRS	DIAGRAM AND NAME OF SHAPE
HCl	–	–	–	H—Cl linear
BeI_2	Be	2	0	I — Be — I linear
BF_3	B	3	0	 triangular planar
CF_4	C	4	0	 tetrahedral
NH_3	N	4	1	 pyramidal
H_2O	O	4	2	 bent

3. *F – F ; F – O ; F – C ; F – Si*

4. CO_2 , NO_2 , HCl , CCl_4 , H_2O

5.

MOLECULE	DOES IT CONTAIN POLAR BONDS?	MOLECULAR SHAPE	IS THE MOLECULE POLAR?
CH_4	Yes	tetrahedral	No
CCl_2F_2	Yes	tetrahedral	Yes
H_2S	Yes	bent	Yes
HI	Yes	linear	Yes
PH_3	Yes	pyramidal	Yes
$AsBr_3$	Yes	pyramidal	Yes

6. *The bonding holding the atoms to each other throughout the crystal lattice in carbon is covalent. This strong bonding gives carbon such a high MP. The attractive force holding the molecules to other molecules throughout the N_2 crystal lattice is weak dispersion force. The very low strength of this attraction causes N_2 to have a very low MP.*

7. *In dry ice the intermolecular forces are weak whereas in SiO_2 the bonding throughout the lattice is very strong.*

8.
(a) *Bonding in both diamond and graphite is strong covalent bonding. High temperatures are needed to break these bonds and cause melting to occur.*
(b) *Diamond is a poor conductor because all valence electrons are localised in strong covalent bonds. This also explains the hardness of diamond. Graphite is a good conductor because one electron per atom is delocalised and capable of moving throughout the lattice.*
(c) *Carbon atoms in graphite bond in such a way as to form sheets of carbon atoms. The attractive force between these sheets is weak and they are able to slide over each other. This gives graphite lubricant properties.*

9. *(a) dispersion (b) dispersion*
(c) dispersion (d) dipole-dipole
(e) dipole-dipole (f) dipole-dipole
(g) dipole-dipole (h) dipole-dipole
(i) dispersion (j) dispersion

10. *(a) H_2O – hydrogen bonds ; HCl – dipole-dipole forces ; F_2 – dispersion forces.*
(b) A highly electronegative atom (F, N or O) is bonded to a hydrogen atom.
(c) H_2O because it has a much smaller mass than the other molecules, bent shape.

11. *(a) H_2Te has the higher molar mass – stronger dispersion forces.*
(b) H_2S exhibits dipole-dipole forces, Ar has weaker dispersion forces / similar masses.
(c) H_2O exhibits hydrogen bonding, H_2S exhibits weaker dipole-dipole forces.
(d) H_2O – stronger hydrogen bonds than CH_4 with dispersion forces.

12. *When HCl is placed in water it ionises forming H^+ and Cl^-. These are highly soluble. Cl_2 is a non-polar molecule and is not very soluble in the polar solvent (water).*

13. *(a) HF exhibits strong hydrogen bonding while the others exhibit weaker dipole-dipole forces.*
(b) Because it has a greater molar mass which leads to stronger dispersion forces.

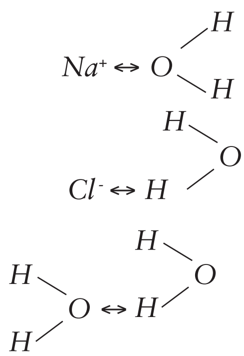
14. *For molecules with very large molar masses the strength of dispersion forces becomes quite significant as indicated by high boiling point of I_2.*

15.

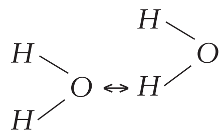
NON-POLAR	POLAR
petrol	ammonia
kerosene	ethanol
turpentine	methylated spirits

16. *Ionic: copper (II) sulfate, ammonium chloride, sodium hydroxide, silver chloride.*
Polar: sugar, ammonia, hydrogen chloride, hydrogen sulfide.
Non-polar: carbon, methane, sulfur.

17.
(a) *Beaker A:* $Na^+ \leftrightarrow Cl^-$



Beaker B:



sugar ↔ water

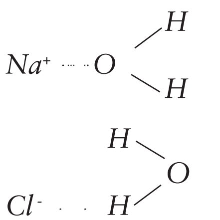
sugar ↔ sugar

Beaker C: *iodine ↔ iodine*

water ↔ water

(b) *Beaker C (I_2 does not dissolve)*

(c)



18. *The relative height that the molecules travelled up the stationary phase gives is inversely related to the strength of the interactions between the molecule and the stationary phase. That would indicate that A would have the greatest dipole-dipole interaction with the stationary phase and is likely to the be the most polar molecule and C the least polar.*

For the Experts

19.

(a) Cholesterol non-polar and so would have a low solubility in water and blood.

(b) The protein region of the lipoprotein is polar and responsible for the lipoprotein being soluble in water and consequently blood. The lipid region of lipoprotein is non-polar and responsible for attractions between the lipoprotein and the cholesterol. The combination of the two attractions allows for the cholesterol to be carried through the blood by the lipoproteins.

(c) The protein region is the polar region of the lipoprotein. The greater the percentage of protein in the lipoprotein, the greater the polarity of the lipoprotein and the greater its density. This means that HDLs are more polar and more dense than LDLs.
As the stationary phase is highly polar, the lipoprotein that travelled more slowly through the stationary phase would be the more polar, this would be the HDL. The peak at 6.5 would be the HDL and the peak at 9.5 would be the LDL.

(d) The height of the LDL peak appears to be about twice that of the HDL which would suggest that the amount of LDL in the blood was about twice the amount of HDL.

CHP 8: THEORY AND GASES

Chapter Questions

8.1

	SOLID	LIQUID	GAS
Shape	fixed	that of container to level surface	takes up complete volume of container
Volume	fixed	fixed	volume of container
Density	high	high	low
Ease of compression	incompressible	almost incompressible	easily compressed
Ease of diffusion	negligible	slow	readily diffuses
Ease of flow	negligible	flows	flows rapidly

8.2

(a) There is a lot of space between gas particles.

(b) Particles in solids are very close together.

(c) Particles in solids are held in fixed positions and cannot flow, unlike the particles in gases and liquids.

8.3

(a) Gases (b) Solids / liquids (c) Gases

8.4 In the gas phase molecules are widely spaced and it requires very little force to reduce this distance of separation, i.e. gases are easy to compress.

8.5 No, gases tend to behave with very similarly so long as they are not at temperatures or pressures likely to cause a phase change.

8.6. The gas particles are small enough to escape through the walls of the balloon.

8.7 The over ripe banana emits a gas which can be smelt. According to the kinetic theory these gas particles will be in a constant state of motion and will gradually spread throughout a room.

8.8

DESCRIPTION	TEMP (°C)	TEMP (K)
Boiling point of helium	-269	4
Boiling point of oxygen	-183	90
Melting point of mercury	-39	234
Melting point of ice	0	273
Normal laboratory temperature	25	298
Normal body temperature	37	310
Boiling point of water	100	373
Melting point of aluminium	660	933
Melting point of NaCl	751	1024
Melting point of tungsten	3422	3695

8.9

(a)

(i) He atom 2m , 500 ms^{-1}
H_2 molecule m , 500 ms^{-1}
Since KE = ½mv^2, He atom has greatest KE (twice that of H_2)

(ii) The effect of halving the speed of the He atom reduces its KE to ¼ of what it was. H_2 unchanged, hence it has the greatest KE (twice that of He).

(b) No. The heavier molecule (O_2) will be travelling a little slower since the average KE (KE = ½mv^2) must be the same for both.

8.10

(a) The area represents the total number of particles.

(b) T_1.

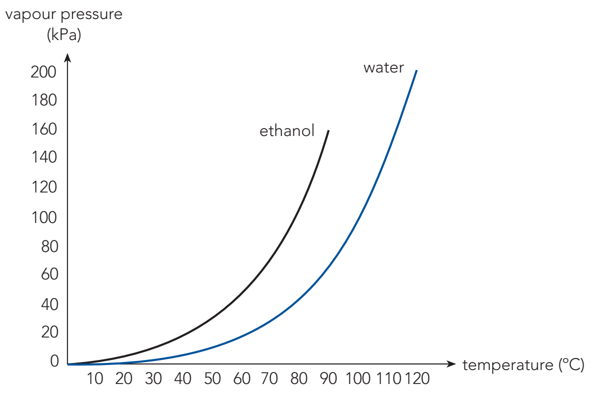
8.11

(a) There is a range of particle energies (velocities). Those with sufficient energy are able to escape the attractive forces of other particles.

(b) The average KE drops (temperature drops) since the particles escaping take the high KE with them.

(c) Temperature of the remaining liquid drops since average KE is lower.

8.12



8.13 (a) 78°C (b) 87°C

8.14

The intermolecular forces for ethanol are weaker as particles are more easily vaporised.

8.15

A pressure cooker is a closed system and the pressure above the liquid can be much greater than 100 kPa (standard atmospheric pressure). Hence water will not boil until the vapour pressure is high enough.

8.16

$n(Cl_2) = V/22.71 = 32.5/22.71 = 1.43$

8.17

$n(N_2) = m/M = 695/28.02 = 24.8$

$V(N_2) = n \times 22.71 = 563$ L

8.18

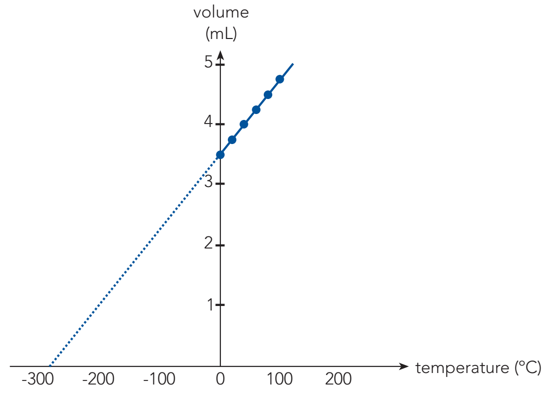
$n(H_2) = m/M = 18\ 000\ 000/2.016 = 8.93 \times 10^6$

$V(H_2) = n \times 22.71 = 2.03 \times 10^8$ L (203 000 cubic metres)

8.19

As temperature increases molecular particles gain more kinetic energy, move faster and hence collide more often with container walls and with greater force. The volume expands since external pressure is constant.

8.20



8.21

(a) ≈ –273° C (b) absolute zero

Review Questions

1. (a) solid (b) gas (c) liquid or gas
   (d) solid (e) gas

2.

(a) Particles in a solid are held in a fixed position.

(b) Gas particles are in a constant state of motion and exhibit no forces of attraction for other particles hence they disperse or spread to all regions of their container.

(c) When heated particles gain more kinetic energy and move more rapidly. With this increased rate of motion they will tend to push each other further apart, ie the substance expands.

3.

(a) Gases behave as ideal gases when they exhibit no attraction for other particles, ie at temperatures and pressures well away from those needed to cause a phase change.

(b) At lower temperatures and higher pressures, gas particles are close enough together and moving slow enough for intermolecular forces of attraction to effect their motion.

4. No, at the same temperature the two gases would have the same average kinetic energy. As the hydrogen molecules are lighter than the nitrogen molecules they will be moving with a higher speed.

5. (a) 123 K, 253 K, 273 K, 373 K
   (b) -173°C, -73°C, 0°C, 227°C

6. B, as the absolute temperature is directly proportional to average kinetic energy. In case B the average temperature has doubled and so the average kinetic energy will also have doubled.

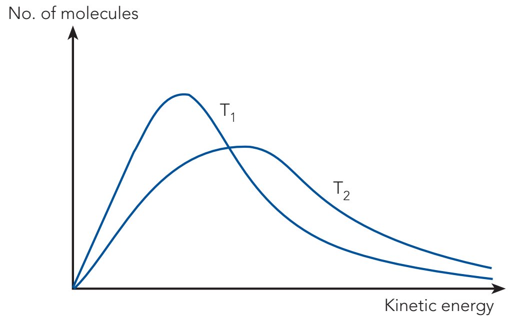
7.

(a) The most common kinetic energy of the molecules.

(b) The average kinetic energy would be the same for both N_2 and O_2 (same temperature).

(c) N_2 molecules would move faster on average because they are lighter than O_2 molecules. Because $KE = \frac{1}{2}mv^2$, smaller masses have a greater velocity for the same KE.

(d)



More particles would have a higher kinetic energy.

8.

(a) Methanol as it vaporises more easily.

(b) Methanol as it vaporises more easily.

(c) Ethanol as it requires a higher temperature for molecules to escape to vapour phase.

(d) Methanol molecules have weaker intermolecular forces.

9.

(a) Gas molecules are very far apart. Their volume is negligible compared to the total volume taken up by the gas.

(b) Molecules of air are constantly colliding with each other and the walls of their container. More air in the football means collisions (pressure) on the internal walls of the football.

(c) Molecules of air in the tyre move faster in warmer temperatures. This means that they hit the inside walls of the tyre more often and with greater force.

10.

(a) The He container as He particles have the lowest mass, they will be moving the fastest and hitting the container walls most often.

(b) The He container for the same reason as (a).

(c) As the containers have the same number of moles of gas at the same volume and temperature, they will have the same pressure.

11.

(a) Salty water would boil at the highest temperature, a little over 100°C. Salty solutions have a reduced vapour pressure due to the effect of ions in solution. The attraction forces between the ions and water molecules reduces the number of water molecules in the vapour phase.

(b) Water boils at a lower temperature at high altitudes due to the lower atmospheric pressure. As an example at an altitude of 3000 m the atmospheric pressure is 70.1 kPa and hence water would boil at 90°C in an open container.

12.

(a) The balloon will deflate or shrink.

(b) The air molecules trapped in the balloon move about more slowly at lower temperatures. Hence the collision between the water molecules and their surroundings become less frequent or forceful. This reduces the pressure in the balloon and it shrinks.

13.

(a) An ideal gas behaves completely according to the kinetic theory of gases.

(b) Real gas particles do attract each other very slightly. At low temperatures and high pressures this allows liquids to form.

(c) Yes – the trend appears to be that the gases made up of larger molecules (higher molar mass) have the smaller molar volume.

14. (a) $n = V/22.71 = 19.8$

(b) $n = V/22.71 = 890$

(c) $n = V/22.71 = 1.41 \times 10^{-3}$

(d) $n = V/22.71 = 6.61$

15. (a) $V = n \text{ X } 22.71 = 52.7$ L

(b) $V = n \times 22.71 = 4.18 \times 10^6$ L

(c) $n = m/M = 72.5/70.9 = 1.02$

$V = n \times 22.71 = 23.2$ L

16. Tyre would require 7.5 L of CO_2 if compared to standard conditions.

$n(CO_2)$ required $= 7.5/22.71 = 0.33$

$m(CO_2)$ required $= n \times M = 0.33 \times 44.01 = 14.5$ g

Yes, a 16.0 g canister contains enough CO_2 to inflate the tyre.

For the Experts

17.

(a) Diving time $= \dfrac{(210 - 50) \times 18}{22 \times 2.5} =$ 52.4 minutes.

Yes, the diver has sufficient gas in this cylinder for a 45 minute dive.

(b) If the gas was stored at STP, $n(\text{gas}) = V/22.71 = 18/22.71 = 0.793$

As the gas is 21% oxygen, $n(O_2)$ = 0.793 × 0.21 = 0.166
Gas is stored at 210 times standard pressure and so $n(O_2)$ would be 210 × 0.166 = 35.0

$M(O_2) = n \times M$ = 35.0 × 32.0 = 1.12 × 103 g

CHP 9: AQUEOUS SOLUTIONS

Chapter Questions

9.1 *Saturated solution, unsaturated solution, supersaturated solution and unsaturated solution.*

9.2

SOLUTION	SOLUTE	SOLVENT	SOLUTION TYPE
sea water	salt	water	solid in liquid
carbonated drink	carbon dioxide	water	gas in liquid
air	gases	gases	gas in gas
ammonia solution	ammonia	water	gas in liquid
brass alloy	zinc	copper	metal in metal
deep sea diver's gas	oxygen nitrogen	helium	gas in gas
vinegar	ethanoic acid	water	liquid in liquid

9.3
(a) *Yes, it's saturated.*
(b) *Excess solute is present.*
(c) *Heat the solution and/or add water.*

9.4
(a) *CO_2 gas*
(b) *Gas is dissolved under pressure, reduction of pressure causes gas bubbles to appear.*
(c) *Gas solubility is low at higher temperatures.*
(d) *(i) the higher the temperature, the lower the solubility.*
(ii) the higher the pressure, the higher the solubility.

9.5
(a) $KNO_{3(s)} \rightarrow K^+(aq) + NO_3^-(aq)$
(b) $CaCl_{2(s)} \rightarrow Ca^{2+}(aq) + 2\ Cl^-(aq)$
(c) $Al_2(SO_4)_{3(s)} \rightarrow 2\ Al^{3+}(aq) + 3\ SO_4^{2-}(aq)$
(d) $C_{12}H_{22}O_{11(s)} \rightarrow C_{12}H_{22}O_{11(aq)}$
(e) $HNO_3 \rightarrow H^+(aq) + NO_3^-(aq)$

9.6
(i) *H_2O, Na^+ and Cl^-*
(ii) *H_2O, CH_3COOH, CH_3COO^- and H^+*
(iii) *CH_3CH_2OH and H_2O*

9.7
(i) *many ions in solution – good conductivity*
(ii) *few ions in solution – poor conductivity*
(iii) *no ions in solution – zero conductivity*

9.8
(a) $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$
(b) $(NH_4)_2SO_{4\,(s)} \rightarrow 2NH_4^+(aq) + SO_4^{2-}(aq)$
(c) $BaSO_{4(s)} \rightarrow Ba^{2+}(aq) + SO_4^{2-}(aq)$

9.9
(a) $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$
(b) $H_2SO_{4\,(aq)} \rightarrow H^+(aq) + HSO_4^-(aq)$
and $HSO_4^-\,(aq) \rightleftharpoons H^+(aq) + SO_4^{2-}(aq)$
(c) $HNO_{3\,(aq)} \rightarrow H^+(aq) + NO_3^-(aq)$
(d) $CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$
(e) $NH_{3\,(aq)} + H_2O(l) \rightleftharpoons NH_4^+\,(aq) + OH^-(aq)$

9.10
- ***Dissociation** is the separation of ions from an ionic solid into solution.*
- ***Ionisation** occurs when atoms from a molecular compound separate by becoming ions in solution.*

9.11

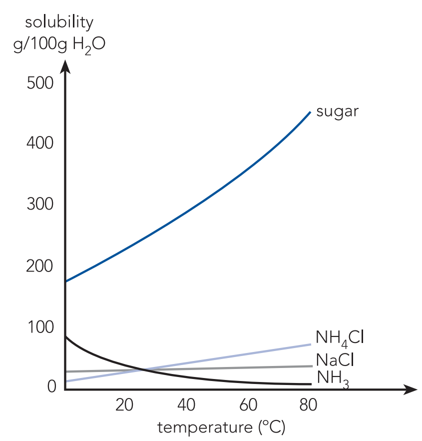
ELECTROLYTE STRENGTH	BRIEF DESCRIPTION	EXAMPLES
Strong	completely ionise or dissociate in solution	NaCl , *AgCl , K_2CO_3 , *$BaSO_4$, $NaCH_3COO$
Weak	partially ionise or dissociate in solution	CH_3COOH , NH_3 , H_3PO_4 , H_2O , H_2CO_3
Non	do not ionise or dissociate in solution	CH_3CH_2OH , CH_4 , $C_{12}H_{22}O_{11}$

9.12
(i) *good conductors – $HCl_{(aq)}$, KCH_3COO , $CaCl_2$*
(ii) *poor conductors – $NH_{3(aq)}$, H_2CO_3 , *$CaSO_4$, *AgCl*
(iii) *non conductors – CH_3OH , $C_6H_{12}O_6$, C_8H_{18}*

**Note:* AgCl, $BaSO_4$ are insoluble salts but to the extent that they dissolve they are completely dissociated.*

9.13
The reaction would produce the salt KCH_3COO which would dissociate to from K^+ ions and CH_3COO^- ions. The resulting solution would be classified as a strong electrolyte.

9.14



9.15

(a)

(i) An increase in temperature increases solubility of solids.

(ii) An increase in temperature decreases solubility of gases.

(b) Yes, solubility of sugar at 50°C = 257 g/100g. $\therefore$ 514 g would dissolve in 200 mL of water (we only put in 500 g).

(c)

(i) Crystals will appear when solution is saturated.

100 g/250 mL H_2O = 40 g/100 g H_2O

$\therefore$ occurs at ≈ 27°C.

(ii) saturated solution.

(iii) supersaturated solution.

9.16

(a) (i) Freezing point is lowered.

(ii) Boiling point is raised.

(b) Directly related to concentration of solute particles (mol L^{-1}), e.g. 200 g of MgF_2 has more effect than 200 g of NaCl because of the greater mol L^{-1} concentration and hence more ions in solution.

(c) Salt lowers the F.P. of water and hence the ice melts.

(d) Ethylene glycol lowers the F.P. of water and hence prevents the water in radiators from freezing in cold climates. In warm climates it allows water to reach a higher temperature before it boils.

(e) As water evaporates the remaining solution becomes more concentrated (salt does not evaporate) and hence boiling point is raised.

(f) A solute decreases the vapour pressure of a liquid, e.g. adding salt to boiling water reduces the vapour pressure of the water and hence boiling stops. At a higher temperature the vapour pressure of the water again reaches atmospheric pressure and boiling occurs.

9.17

(a) $n = cV = (1.75)(5.0) = 8.75$ mol

(b) for 250 mL of NaOH solution

$n = cV = (1.75)(0.25) = 0.4375$ mol

$m(NaOH) = nM = (0.4375)(40.0)$

$= 17.5$ g

9.18

(a)

(i) % solids (by mass)

$$= \frac{86.5}{2500} \times 100 = 3.46\%$$

(ii) $\text{ppm} = \dfrac{86500 \text{ mg}}{2.50 \text{ kg}} = 34600 \text{ ppm}$

(b)

(i) $\text{ppm (NaCl)} = \dfrac{73400 \text{ mg}}{2.50 \text{ kg}} = 29360 \text{ ppm}$

(ii) $n\text{ (NaCl)} = \dfrac{m}{M} = \dfrac{73.4}{58.44} = 1.26 \text{ mol}$

$\text{Volume } H_2O = \dfrac{m}{\rho} = \dfrac{2.50}{1.03} = 2.427 \text{ L}$

$\therefore c\text{ (NaCl)} = \dfrac{n}{V} = \dfrac{1.25}{2.427} = 0.515 \text{ mol L}^{-1}$

9.19

(a) $M(KNO_3) = 101.1 \text{ g mol}^{-1}$

$c = \dfrac{20 \text{ g}}{0.250 \text{ L}} = 80.0 \text{ g L}^{-1}$

or $c = (80.0) \div (101.1) = 0.791 \text{ mol L}^{-1}$

or $c = \dfrac{n}{V} = \dfrac{20.0/101.1}{0.250} = 0.791 \text{ mol L}^{-1}$

(b) $c = \dfrac{87.7 \text{ g}}{0.750 \text{ L}} = 117 \text{ g L}^{-1}$

or $c = (117) \div (46.068) = 2.54 \text{ mol L}^{-1}$

or $c = \dfrac{n}{V} = \dfrac{87.7/46.068}{0.750} = 2.54 \text{ mol L}^{-1}$

(c) $c = \dfrac{75.6 \text{ g}}{2.0 \text{ L}} = 37.8 \text{ g L}^{-1}$

or $c = (37.8) \div (60.05) = 0.629 \text{ mol L}^{-1}$

or $c = \dfrac{n}{V} = \dfrac{75.6/60.05}{2.0} = 0.629 \text{ mol L}^{-1}$

(d) $c = \dfrac{44.5 \text{ g}}{0.375 \text{ L}} = 119 \text{ g L}^{-1}$

or $c = (119) \div (342.3) = 0.348 \text{ mol L}^{-1}$

or $c = \dfrac{n}{V} = \dfrac{44.5/342.3}{0.375} = 0.348 \text{ mol L}^{-1}$

9.20

(i) soluble – $Zn(NO_3)_2$, $AlCl_3$, $(NH_4)_3PO_4$ $MgSO_4$, $BaCl_2$, Na_2CO_3 , $NaNO_3$, NaCl

(ii) slightly soluble - $PbCl_2$, $Ca(OH)_2$, $CaSO_4$, Ag_2SO_4 , $PbBr_2$

(iii) insoluble - AgI , $BaSO_4$, AgCl , $PbSO_4$, $Mg(OH)_2$, AgBr

9.21

(a) No precipitate.

(b) $Zn(NO_3)_{2(aq)} + 2NaOH_{(aq)} \rightarrow Zn(OH_2)_{(s)} + 2NaNO_{3(aq)}$

$Zn^{2+}{}_{(aq)} + 2OH^-{}_{(aq)} \rightarrow Zn(OH)_{2(s)}$

(c) $CaCl_{2(aq)} + 2AgNO_{3(aq)} \rightarrow 2AgCl_{(s)} + Ca(NO_3)_{2\,(aq)}$

$Ag^+{}_{(aq)} + Cl^-{}_{(aq)} \rightarrow AgCl_{(s)}$

(d) $BaCl_{2(aq)} + (NH_4)_2SO_{4\,(aq)} \rightarrow BaSO_{4\,(s)} + 2NH_4Cl_{\,(aq)}$

$Ba^{2+}{}_{(aq)} + SO_4{}^{2-}{}_{(aq)} \rightarrow BaSO_{4\,(s)}$

(e) $Pb(NO_3)_{2(aq)} + 2NaI_{(aq)} \rightarrow PbI_{2(s)} + 2NaNO_{3(aq)}$

$Pb^{2+}{}_{(aq)} + 2I^-{}_{(aq)} \rightarrow PbI_{2(s)}$

9.22

	K^+	Pb^{2+}	NH_4^+	Ba^{2+}	Ag^+
Cl^-	-	$PbCl_{2(s)}$ white	-	-	AgCl(s) white
I^-	-	$PbI_{2(s)}$ yellow	-	-	AgI(s) yellow
SO_4^{2-}	-	$PbSO_{4(s)}$ white	-	$BaSO_{4(s)}$ white	$Ag_2SO_{4(s)}$ white
CO_3^{2-}	-	$PbCO_{3(s)}$ white	-	$BaCO_{3(s)}$ white	$Ag_2CO_{3(s)}$ white

9.23

$CaO_{(s)} + H_2O_{(l)} \rightarrow Ca^{2+}_{(aq)} + 2\,OH^-_{(aq)}$ Cadmium ions in the water will form the insoluble $Cd(OH)_2$ precipitate when lime is added.

$$Cd^{2+}_{(aq)} + 2\,OH^-_{(aq)} \rightarrow Cd(OH)_{2(s)}$$

9.24

(a) *Strontium, $Sr^{2+}_{(aq)} + SO_4^{2-}_{(aq)} \rightarrow SrSO_{4(aq)}$*

(b) *Calcium may cause a similar set of results. $Ca^{2+}_{(aq)}$ is sparingly soluble with sulfate ions and as only a small amount of dilute sodium sulfate was added the calcium ions may have remained soluble.*

Review Questions

1. *Any undissolved solid indicates a saturated solution. If otherwise, it may be supersaturated (any addition of NaCl would cause crystallisation) or unsaturated (any additional NaCl would dissolve).*

2.

	SUBSTANCE TYPE			
	Covalent Network	Covalent Molecular	Metallic	Ionic
Soluble	–	HCl, NH_3, CH_3COOH, $C_{12}H_{22}O_{11}$, ethanol	–	NaCl, Na_2CO_3, $Cu(NO_3)_2$
Insoluble	SiO_2, diamond	olive oil, kerosene, petrol	Zn, Fe, Au	–

3.

(a) *A conductivity test.*

(b) *If no conduction – water.*
If minor conduction – ethanoic acid.
If good conduction – sulfuric acid.

4.

(a) $HCl_{(g)} \rightarrow HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$

(b) $MgSO_{4(s)} \rightarrow Mg^{2+}_{(aq)} + SO_4^{2-}_{(aq)}$

(c) $CH_3COOH_{(aq)} \rightleftharpoons CH_3COO^-_{(aq)} + H^+_{(aq)}$

(d) $H_3PO_{4(aq)} \rightleftharpoons H_2PO_4^-_{(aq)} + H^+_{(aq)}$

(e) $NaOH_{(s)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$

(f) $NH_{3(g)} + H_2O_{(l)} \rightleftharpoons NH_4^+_{(aq)} + OH^-_{(aq)}$

(g) $KI_{(s)} \rightarrow K^+_{(aq)} + I^-_{(aq)}$

(h) $Al_2(SO_4)_{3(s)} \rightarrow 2Al^{3+}_{(aq)} + 3SO_4^{2-}_{(aq)}$

5. *Both ammonia and ethanoic acid solutions are weak conductors of electricity because they only slightly ionise in aqueous solution. Hence there are few ions to transfer the current. When they are combined, however, ammonium ethanoate is produced. This salt is highly soluble and a good electrolyte.*

$NH_{3(aq)} + CH_3COOH_{(aq)} \rightarrow NH_4CH_3COO_{(aq)}$ and
$NH_4CH_3COO_{(aq)} \rightarrow NH_4^+_{(aq)} + CH_3COO^-_{(aq)}$

6. $ppm = \dfrac{\text{*mg of solute}}{\text{kg of solution}}$

**mg = milligrams*
= 1.0 × 10^-3^ grams
= 1.0 × 10^-6^ kg

$= \dfrac{(5.50)(1{,}000{,}000)}{50{,}000}$

$= 110$ ppm

Just under the accepted range.

7. $c(N) = 1.0$ mg L^{-1}

$= \dfrac{1.0 \times 10^{-3} \text{mol } L^{-1}}{14.0}$

$= 7.14 \times 10^{-5}$ mol L^{-1}

8. *river water conc (salt) = 20.0 g L^{-1}*
∴ mass in 750 L = (20.0)(750)
= 1.5×10^4 g (15 kg)
lake water conc (salt) = 0.50 mol L^{-1}
∴ n (salt) in 750 L = (0.50)(750) = 375 mol
∴ m (salt) in 750 L = (375)(58.44) =
2.19×10^4 g (21.9 kg)

9.

(a) $Ag^+_{(aq)} + Br^-_{(aq)} \rightarrow AgBr_{(s)}$

(b) $Fe^{3+}_{(aq)} + 3OH^-_{(aq)} \rightarrow Fe(OH)_{3(s)}$

(c) $3Cu^{2+}_{(aq)} + 2PO_4^{3-}_{(aq)} \rightarrow Cu_3(PO_4)_{2(s)}$

(d) $2Cr^{3+}_{(aq)} + 3CO_3^{2-}_{(aq)} \rightarrow Cr_2(CO_3)_{3(s)}$

(e) $Pb^{2+}_{(aq)} + 2I^-_{(aq)} \rightarrow PbI_{2(s)}$

(f) $Sr^{2+}_{(aq)} + SO_4^{2-}_{(aq)} \rightarrow SrSO_{4(s)}$

10. (a) *Sodium ions*
(b) *Cu^{2+} or copper (II) ions*
(c) *Barium ions*

11.

(a) $n(Na_2CO_3) = m/M$
$= 15.2/(45.98+12.01+48.00) = 0.143$
$c(Na_2CO_3) = n/V = 0.143/0.375 = 0.382$ mol L^{-1}

(b) $n(CaCl_2) = m/M = 2500/(40.08 + 70.90) = 22.53$
$c(CaCl_2) = n/V = 22.53/45000 = 5.01 \times 10^{-4}$ mol L^{-1}

(c) $n(HCl) = m/M = 9.47/(1.008 + 35.45) = 0.260$
$c(HCl) = n/V = 0.260/0.250 = 1.04$ mol L^{-1}

12. *ppm = m(Hg) in mg/m (sludge) in kg = 0.1026/0.150 = 0.684 ppm*

13. *m(NaF) in toothpaste* = $0.750 \times 0.32/100 = 2.40 \times 10^{-3}$ *g*

14.
(i) *Done in text.*
(ii) *Dissolved oxygen is essential to aquatic plants and animals for respiration.*
(iii) *Dissolved compounds of nitrogen (usually nitrates) and phosphorus (phosphates) are essential nutrients to plant growth. Excess nutrients however, can cause eutrophication.*
(iv) *Increased salinity of soil reduces their suitability for farming. Rivers and water supplies become more salty.*
(v) *Dissolved salts of heavy metals (mercury, lead, chromium, cadmium, etc) are toxic if they accumulate in the body to significant amounts. For humans, fish can be such a source.*
(vi) *Excess dissolved nutrients stimulate plant growth such as algal bloom. The eventual bacterial decay of the algae depletes the oxygen in the water. This causes animal life to suffer.*

For the Experts

15.

(a) $n = \dfrac{m}{M} = \dfrac{15.0}{30.97} \times \dfrac{4.5}{100} = 0.0218 \text{ mol}$

$\therefore c = \dfrac{n}{V} = \dfrac{0.0218}{9.0 \text{ L}} = 2.42 \times 10^{-3} \text{ mol L}^{-1}$

(b)
(i) $Ca(H_2PO_4)_2\,(aq) \rightarrow Ca^{2+}(aq) + 2H_2PO_4^{-}(aq)$
$CaSO_4 \cdot 2H_2O\,(aq) \rightarrow Ca^{2+}(aq) + SO_4^{2-}(aq) + 2H_2O\,(l)$
(ii) *mass of P* $= \dfrac{8.8}{100} \times \dfrac{95}{100} \times 10{,}000 \text{ kg}$

$= 836$ *kg of P*

CHP 10: ACIDS AND BASES

Chapter Questions

10.1
(a)

ACIDIC SOLUTIONS	BASIC SOLUTIONS
$[H^+] > [OH^-]$	$[H^+] < [OH^-]$
$[H^+] > 1.00 \times 10^{-7}$ mol L^{-1}	$[H^+] < 1.00 \times 10^{-7}$ mol L^{-1}
$[OH^-] < 1.00 \times 10^{-7}$ mol L^{-1}	$[OH^-] > 1.00 \times 10^{-7}$ mol L^{-1}

10.2
(a) $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$, *acidic*
(b) $HClO(aq) \rightarrow H^+(aq) + ClO^-(aq)$, *acidic*
(c) $KOH(s) \rightarrow K^+(aq) + OH^-(aq)$, *basic*
(d) $Ba(OH)_2(s) \rightarrow Ba^{2+}(aq) + 2OH^-(aq)$, *basic*

10.3

	Acid	Base
(a)	CH_3COOH	H_2O
(b)	HSO_4^-	CO_3^{2-}
(c)	H_3PO_4	$NaOH$
(d)	HCl	NH_3
(e)	H_2O	PO_4^{3-}

10.4
Acid: $H_2O(l) + HCO_3^-(aq) \rightleftharpoons H_2CO_3\,(aq) + OH^-(aq)$
Base: $H_2O(l) + H_2PO_4^-(aq) \rightleftharpoons H_3O^+\,(aq) + HPO_4^{2-}(aq)$

10.5
$Ca(OH)_2\,(s) \rightarrow Ca^{2+}\,(aq) + 2OH^-(aq)$
The production of OH^- *ions will make the soil basic.*

10.6
$SO_2\,(g) + H_2O(l)\ (\rightarrow H_2SO_3\,(aq)) \rightleftharpoons H^+(aq) + HSO_3^-(aq)$
$SO_3\,(g) + H_2O(l)\ (\rightarrow H_2SO_4(aq)) \rightarrow H^+(aq) + HSO_4^-(aq)$

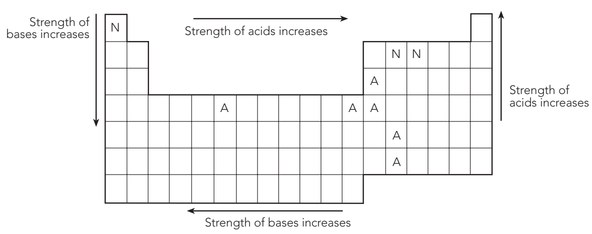
10.7
(a) *Ethanoic acid is a weak acid, only a small % of* CH_3COOH *molecules will ionise to form* H^+ *ions. Therefore the* $[H^+]$ *will be considerably lower than the* $[CH_3COOH]$.
(b) $pH = -\log[H^+] = -\log(4.16 \times 10^{-4}) = 3.38$

10.8
(a) $[H^+] = 1.00 \times 10^{-9}$ mol L^{-1}
(b) $[H^+] < 1.00 \times 10^{-7}$ mol L^{-1}, *solution is basic*

10.9
(a) $HBr(aq) \rightarrow H^+(aq) + Br^-(aq)$
(b) $H_3PO_4 \rightleftharpoons H^+ + H_2PO_4^- \rightleftharpoons H^+ + HPO_4^{2-} \rightleftharpoons H^+ + PO_4^{3-}$
(c) $SO_2 + H_2O \rightarrow H_2SO_3 \rightleftharpoons H^+ + HSO_3^-$
(d) $SO_3 + H_2O \rightarrow H_2SO_4 \rightarrow H^+ + HSO_4^- \rightleftharpoons H^+ + SO_4^{2-}$

10.10
(a) $Na_2O(s) + H_2O(l) \rightarrow 2Na^+(aq) + 2OH^-(aq)$
(b) $NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$
(c) $PO_4^{3-}(aq) + H_2O(l) \rightleftharpoons HPO_4^{2-}(aq) + OH^-(aq)$
(d) $KHCO_3(s) + H_2O(l) \rightarrow K^+(aq) + HCO_3^-(aq) + H_2O \rightleftharpoons K^+(aq) + H_2CO_3(aq) + OH^-(aq)$

10.11



A = amphoteric
N = neutral (H_2O, CO, NO *and* N_2O)
10.12

(a) $2H^+(aq) + Mg(OH)_2(s) \rightarrow Mg^{2+}(aq) + 2H_2O(l)$
(b) $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$
(c) $SO_3(g) + H_2O(l) \rightarrow H^+(aq) + HSO_4^-(aq)$
(d) $2Al(s) + 2OH^-(aq) + 6H_2O(l) \rightarrow 2[Al(OH)_4]^-(aq) + 3H_2(g)$
(e) $SO_3(g) + 2OH^-(aq) \rightarrow SO_4^{2-}(aq) + H_2O(l)$
(f) $Al(OH)_3(s) + OH^-(aq) \rightarrow [Al(OH)_4)]^-(aq)$
(g) $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
(h) $3ZnCO_3(s) + 2H_3PO_4(aq) \rightarrow Zn_3(PO_4)_2(s) + 3CO_2(g) + 3H_2O(l)$
(i) $H_3PO_4(aq) + Al(OH)_3(s) \rightarrow AlPO_4(s) + 3H_2O(l)$
(j) $CO_3^{2-}(aq) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$
(k) $NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$
(l) $NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$
(m) $Fe(OH)_3(s) + 3H^+(aq) \rightarrow Fe^{3+}(aq) + 3H_2O(l)$
(n) $SO_3(g) + 2OH^-(aq) \rightarrow SO_4^{2-}(aq) + H_2O(l)$
(o) $BaO(s) + 2H^+(aq) \rightarrow Ba^{2+}(aq) + H_2O(l)$
(p) $Ba^{2+}(aq) + 2OH^-(aq) + 2H^+(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s) + H_2O(l)$

10.13
Amphoteric substances can act as acids or bases.

10.14
(a) $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$
(b) $2Al(s) + 2OH^-(aq) + 6H_2O(l) \rightarrow 2[Al(OH)_4]^-(aq) + 3H_2(g)$
(c) $Cr(OH)_3(s) + OH^-(aq) \rightarrow [Cr(OH)_4]^-(aq)$
(d) $Cr(OH)_3(s) + 3H^+(aq) \rightarrow Cr^{3+} + 3H_2O(l)$
(e) $Cr_2O_3(s) + 6H^+(aq) \rightarrow 2Cr^{3+} + 3H_2O(l)$
(f) $Cr_2O_3(s) + 2OH^-(aq) + 3H_2O(l) \rightarrow 2[Cr(OH)_4]^-(aq)$

10.15
(i) $CO_2 + H_2O \rightarrow H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$
(ii) $SO_2 + H_2O \rightarrow H_2SO_3 \rightleftharpoons H^+ + HSO_4^-$

10.16
Because both oxides produce H^+ ions they will decrease the pH of rainwater.

10.17
Acid rain will gradually cause the structures to deteriorate as the acid dissolves the limestone/marble.

10.18 *Decrease the pH.*

10.19
$2NO_2(g) + H_2O(l) \rightarrow 2H^+(aq) + NO_3^-(aq) + NO_2^-(aq)$

10.20
$6NO_2(g) + 8NH_3(g) \rightarrow 7N_2(g) + 12H_2O(g)$

Review Questions

1.
(a) *ammonium ion*
(b) *Brønsted Lowry theory suggests that the NH_4^+ is acidic as it donates a proton to the H_2O.*

2. (a) *acid* - HS^-, *base* - CO_3^{2-}
(b) *acid* - H_2O, *base* - HPO_4^{2-}
(c) *acid* - CH_3COOH, *base* - CH_3NH_2
(d) *not an acid base reaction.*

3.
(a) $Zn(s) + 2OH^-(aq) + 2H_2O(l) \rightarrow [Zn(OH^-)_4]^{2-}(aq) + H_2(g)$
(b) $Al(OH)_3(s) + 3H^+(aq) \rightarrow Al^{3+}(aq) + 3H_2O(l)$
(c) *Add sodium hydroxide solution to both, the $Mg(OH)_2$ will not dissolve but the white $Al(OH)_3(s)$ will dissolve to form a colourless solution.*

4.
(a) $HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$
(b) $NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$
(c) $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
(d) $2CH_3COOH(aq) + Na_2CO_3(s) \rightarrow 2Na^+(aq) + 2CH_3COO^-(aq) + CO_2(g) + H_2O(l)$
(e) $H_2CO_3(aq) + 2OH^-(aq) \rightarrow CO_3^{2-}(aq) + 2H_2O(l)$
(f) $MgCO_3(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + CO_2(g) + H_2O(l)$

5.
(a) $2H^+(aq) + MgO(s) \rightarrow H_2O(l) + Mg^{2+}(aq)$
(b) $H^+(aq) + CH_3COO^-(aq) \rightarrow CH_3COOH(aq)$
(c) $H_3PO_4(aq) + Fe(OH)_3(s) \rightarrow FePO_4(s) + 3H_2O(l)$
(d) $SO_2(g) + 2OH^-(aq) \rightarrow SO_3^{2-}(aq) + H_2O(l)$
(e) $CO_2(g) + Ca^{2+}(aq) + 2OH^-(aq) \rightarrow CaCO_3(s) + H_2O(l)$
(f) $NH_3(aq) + CH_3COOH(aq) \rightarrow NH_4^+(aq) + CH_3COO^-(aq)$

6.
(a) 3
(b) 5
(c) $[OH^-] = 1 \times 10^{-3}$ mol L^{-1},
$$[H^+] = \frac{1 \times 10^{-14}}{1 \times 10^{-3}}$$
$= 1 \times 10^{-11}$, pH = 11
(d) 1.52
(e) $[OH^-] = 2 \times 7.5 \times 10^{-6} = 1.5 \times 10^{-5}$ mol L^{-1}
$$[H^+] = \frac{1 \times 10^{-14}}{1.5 \times 10^{-5}} = 6.67 \times 10^{-10}, \text{pH} = 9.18$$

7.

MATERIAL	pH	[H+]	[OH–]
vinegar	3.00	1.00×10^{-3}	1.00×10^{-11}
toothpaste	6.80	1.58×10^{-7}	6.33×10^{-8}
oven cleaner	13.5	3.16×10^{-14}	0.316
window cleaner	9.75	1.79×10^{-10}	5.62×10^{-5}

8. *HCl and HNO_3 are strong acids while CH_3COOH is a weak acid. The $[H^+]$ in CH_3COOH will be less than in solution of HCl and HNO_3 of the same concentration*

and so will produce a different colour when indicator is added.

9.
(a) *As a base it will neutralise the spilt acid. Solid mops up solution and contains it. It is a weak base and so the addition of too much Na_2CO_3 is not a major safety concern. The Na_2CO_3 will react with the acid to produce CO_2 which will cause the mixture to fizz. Na_2CO_3 is added when the fizzing stops, indicating that the acid has been neutralised.*
(b) *Add enough dry acid such as sodium hydrogensulfate to neutralise the $NaOH_{(aq)}$ or enough to soak up the liquid.*

10.
(a) *The concentration of acid is a measure of the number of moles of acid that has been dissolved to form each litre of solution. The strength of an acid is a measure of the extent to which the molecules ionise to form H^+ ions when dissolved in water.*
(b) *(i) 0.1 mol L^{-1} NaOH*
(ii) 0.1 mol L^{-1} CH_3COOH
(iii) 6 mol L^{-1} HCl

11.
(a) *Add H_2SO_4 (aq) to both; the $Ba(OH)_{2(aq)}$ solution will produce a white precipitate while with the $NH_{3(aq)}$ no visible change will occur but the strong ammonia smell will disappear.*
(b) *Dissolve both in water and add universal indicator. The $NaCH_3COO_{(aq)}$ will turn blue/purple while the $NaCl_{(aq)}$ will turn green.*
(c) *Add $Ba(OH)_{2(aq)}$ to both, a white precipitate will form in the H_2SO_4 (aq) while the $HCl_{(aq)}$/$Ba(OH)_{2(aq)}$ mixture will remain a colourless solution.*

12. $2\ H_3PO_4 + 3\ K_2CO_3 \rightarrow 2\ K_3PO_4 + 3\ CO_2 + 3\ H_2O$

$n(K_2CO_3) = c.V = 0.250 \times 0.0624 = 0.0156$
$n(CO_2) = n(K_2CO_3) = 0.0156$
$m(CO_2) = n.M = 0.0156 \times (12.01 + 32.00) = 0.687\ g$

13. $2\ NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2\ H_2O$

$n(NaOH) = c.V = 2.77 \times 7.99 = 22.1$
$n(H_2SO_4) = ½\ n(NaOH) = 11.1$
$V(H_2SO_4) = n/c = 11.1/5.75 = 1.92\ L$

14. $Ca(OH)_2 + 2\ HCl \rightarrow CaCl_2 + 2\ H_2O$
$n(Ca(OH)_2) = c.V = 0.330 \times 0.00824 = 2.72 \times 10^{-3}$
$n(CaCl_2) = n(Ca(OH)_2) = 2.72 \times 10^{-3}$
$m(CaCl_2) = n.M = 2.72 \times 10^{-3} \times (40.08 + 70.90) = 0.302\ g$

15. $Mg + 2\ HNO_3 \rightarrow Mg(NO_3)_2 + H_2$
(a) $n(HNO_3) = c.V = 1.75 \times 0.0500 = 0.0875$
$n(Mg)$ lost $= ½\ n(HCl) = 0.04375$
$m(Mg)$ lost $= n.M = 0.04375 \times 24.31 = 1.06\ g$
(b) $V(H_2) = n \times 22.71 = 0.994\ L$

16.
(a) *acid + base → salt + water*
(b) *acid + reactive metal → salt + hydrogen gas*
(c) *acid + metal oxide → salt + water*
(d) *acid + carbonate → salt + water + carbon dioxide*
(e) *acid + hydrogen carbonate → salt + water + carbon dioxide*
(f) *(strong) acid + metal sulfite → salt + water + sulfur dioxide*
(g) *(strong) acid + metal sulfide → salt + hydrogen sulfide*
(h) *non-metal oxide + base → salt + water*

For the Experts

17. $CO_{2(g)} + H_2O_{(l)} \rightarrow H_2CO_{3\ (aq)} \rightleftharpoons H^+_{(aq)} + HCO_3^-{}_{(aq)}$
$SO_{2(g)} + H_2O_{(l)} \rightarrow H_2SO_{3\ (aq)} \rightleftharpoons H^+_{(aq)} + HSO_3^-{}_{(aq)}$

18. $CO_{2(g)} + 2OH^-{}_{(aq)} \rightarrow CO_3^{2-}{}_{(aq)} + H_2O_{(l)}$

19. $CaCO_{3(s)} + SO_{2(g)} \rightarrow Ca^{2-}{}_{(aq)} + SO_3^{2-}{}_{(aq)} + CO_{2(g)}$
$Na_2CO_{3(aq)} + CO_3^{2-}{}_{(aq)} + SO_{2(aq)} \rightarrow SO_3^{2-}{}_{(aq)} + CO_{2(g)}$

CHP 11: REACTION RATES

Chapter Questions

11.1

VERY FAST	SLOW	VERY SLOW
burning precipitate neutralisation explosion	rusting fermentation setting digestion	photosynthesis weathering

11.2
(a) *Vigorous reaction, bubbles of colourless gas, heat released.*
(b) $CaCO_{3(s)} + 2H^+_{(aq)} \rightarrow Ca^{2+}{}_{(aq)} + H_2O(l) + CO_{2(g)}$
(c) *The reaction would gradually slow down, fewer gas bubbles.*
(d) • *mass of $CaCO_3$ decreases*
• *H^+ concentration decreases*
• *Ca^{2+} concentration increases*
• *volume of CO_2 increases*
(e) *Concentration of H^+ ions decreases. Available surface area of zinc also decreases.*
(f) *1st minute ≈ 400 mL $CO_{2(g)}$ produced*
3rd minute ≈ 125 mL $CO_{2(g)}$ produced
Reaction rate ≈ 400/125 ≈ 3.2 times faster during 1st minute.

(g)

- *concentration of the acid*
- *temperature*
- *how finely divided the marble chips were.*

11.3
(a) 3, H–H , H–H , O–O
(b) 4, O–H (four of)

11.4
Where a bond is broken or formed - energy greater than minimum required and orientation is appropriate.

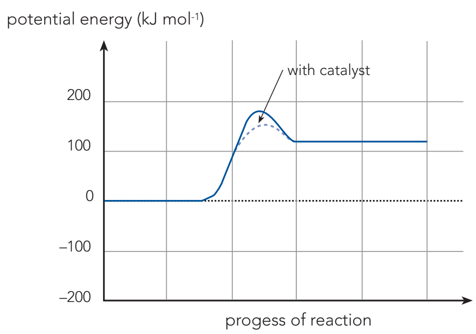
11.5
E.g. 3 is slowest since many bonds have to break.

11.6
At first the concentration of the acid was greatest. As acid is consumed in the reaction, concentration decreases and so does reaction rate.

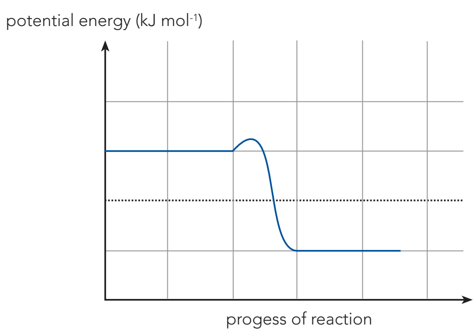
11.7
(a) S.A. = (6) (2 × 2) = 24 cm^2
(b) S.A. = (8) (6) (1 × 1) = 48 cm^2
(c) Rate would be twice as fast (48/24) since surface area is twice as great.

11.8
This increases surface area and therefore the effectiveness of the catalyst.

11.9 (a)(b)



11.10
(a)



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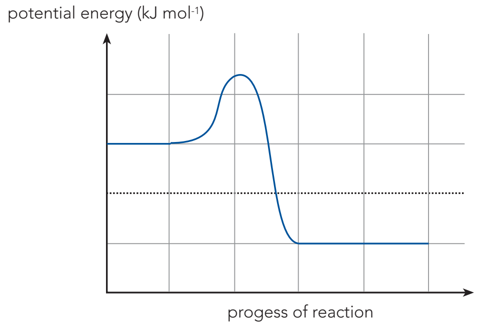

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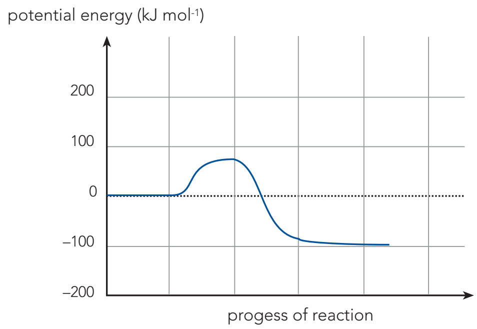
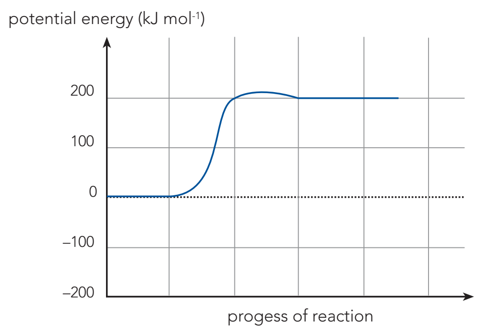


(d)



11.11
(a) Total number of particles available for reaction.
(b) The number of particles with sufficient energy to undergo a successful collision (i.e. activation energy).
(c) A greater proportion of particles have sufficient energy to undergo a reaction.

Review Questions

1.
(a) No change. No change in concentration or surface area available for reaction.
(b) Increase in reaction rate as surface area has increased.
(c) Decrease in reaction rate. Lower temperature means fewer collisions and less forceful collisions.
(d) Increase in reaction rate since concentration has increased.

2.
(a) (iii) precipitation of $AgCl_{(s)}$ *fastest – no bonds to break.*
(b) (ii) oxidation of glucose slowest. The greatest number of bonds to be broken.

3. *Concentration of reactant particles increases reaction rate as the number of collisions will increase. This will also increase the number of successful collisions.*

4.
(a) $Zn_{(s)} + 2H^{+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + H_{2(g)}$.
(b) *The concentration of the* H^+ *ions is greatest at first then gradually decreases as the reaction proceeds and the available* H^+ *is consumed.*
(c) *With 6.0 mol* L^{-1} *acid the time for 400 mL was 1.0 minute. With 2.0 mol* L^{-1} *it would be 3.0 minutes.*
(d)
- H^+ *concentration*
- *temperature*
- *how finely divided the zinc was*

5. *Sugar cubes have much less surface area than sugar granules. Hence slower to dissolve.*

6.
(a) *flour dust is combustible since the surface area exposed to oxygen gas in the air is very great*
(b) *a large activation energy is required for this reaction*
(c) *the lower temperature greatly reduces the chemical processes which cause food to spoil*
(d) *as a fire burns, it consumes much of the oxygen around it. A breeze helps to increase oxygen availability and also removes the carbon dioxide produced which would inhibit burning.*

7. *A temperature increases the velocity of reacting particles is greater. This means that there will be more collisions and importantly more forceful collisions. These two factors will cause a large increase in the number of successful collisions (reaction rate).*

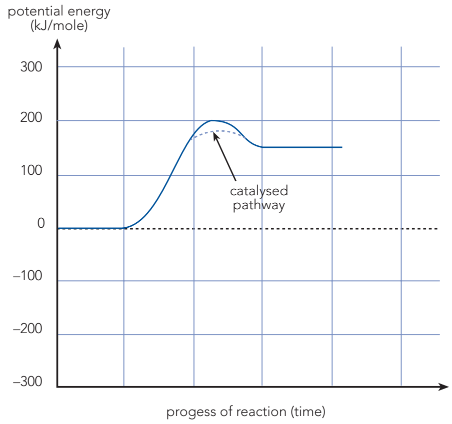
8.
(a) *manganese (IV) oxide,* MnO_2.
(b) *platinum.*
(c) *enzymes.*

9.
(a) *exothermic.*
(b) $E_a \approx 150$ *kJ.*
(c) $\Delta H \approx 220$ *kJ (reverse reaction).*
(d) (i) *none.* (ii) *will reduce it.*

10.
(a) E_a – *this is the energy that colliding particles must have in order to form an activated complex.*
(b) *Reactant particles which collide with sufficient energy form an activated complex and are in a transition state. Bonds are breaking and forming.*
(c) *Heat of reaction (ΔH) is the difference in potential energy (enthalpy) between the reactants and products.*

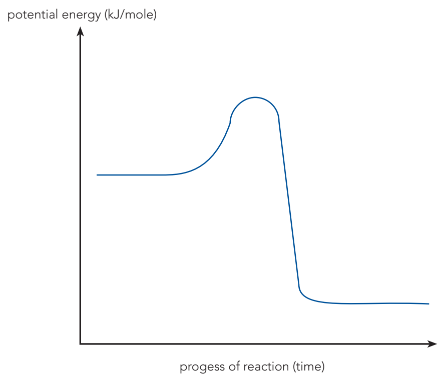
$\Delta H = H\ (products) - H\ (reactants)$.

11.
(a)(b)

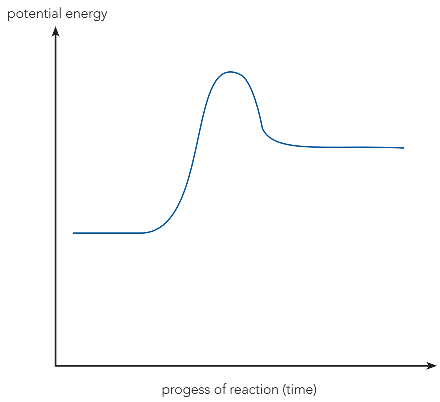


(c) (i) $\Delta H = 150$ *kJ* (ii) *Ea = <200 kJ, >150 kJ*

12.
(a)



(b)



For the Experts

13.
(a) $2\ CO_{(g)} + O_{2(g)} \rightarrow 2\ CO_{2(g)}$
(b) $2\ CO_{(g)} + 2\ NO_{(g)} \rightarrow 2\ CO_{2(g)} + N_{2(g)}$
(c) *The Pd and Rh are catalysts, they speed up the conversion of toxic pollutants to less harmful materials by lowering the activation energy of the reaction. This allows a faster reaction*

rate which is important as the exhaust gases only spend a small amount of time travelling through the catalytic converter.

(d) *Chemical reactions require reactants to collide. For Pd and Rh to act as catalysts they must be involved in collisions with the reactants and this will occur on the surface of the ceramic support structure. The support structure is of a honeycombed shape to provide a much larger surface area. The greater the surface area the greater the likelihood the reactants colliding with the catalyst.*

(e) *From a chemical reaction perspective the amount of catalyst in the structure should remain constant and not diminish with time as a catalyst is not permanently consumed in a chemical reaction.*

ANSWERS TO TRIAL TESTS

TT1: ATOMIC STRUCTURE

Section 1

1. *a*	6. *d*
2. *c*	7. *b*
3. *c*	8. *b*
4. *d*	9. *c*
5. *a*	10. *a*

[20]

Section 2

11.

(a) (i) 37 (ii) 17 (iii) 20 (iv) 17

(b) (i) 2,5 (ii) 2,8 (iii) 2,8,8 (iv) 2,8,8,1

[8]

12.

(a) Valence electrons are the outer shell electrons of an atom.

(b) (i) 2 (ii) 18 (iii) 11 (iv) 13

[6]

13.

(a) They are all in the third energy level of the atom.

(b) They all have one valence electron.

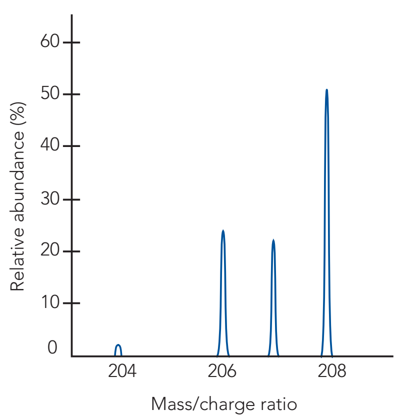
(c) The chemical properties of an element depend mainly on the number of valence electrons it has. Elements of any particular group all have the same number of valence electrons and hence their chemical properties are similar.

(d) The inert gases all have a full outer shell of electrons. Hence they do not react as they already have a very stable electron structure.

(e) Inert gases all have a very stable electron structure of 8 valence electrons (except He which has 2). Other elements always react in a way which tries to achieve this preferred outer-level configuration of 8 valence electrons (the octet rule).

[10]

14. (a)



(b) A_r *(Pb) =* $\sum$ *(isotopic mass × abundance %)/100*

= $\sum$ *((203.97 × 1.4) + (205.97 × 24.1) + (206.98 × 22.1) + (207.98 × 52.4))/100*

= 2.86 + 49.64 + 45.74 + 108.98

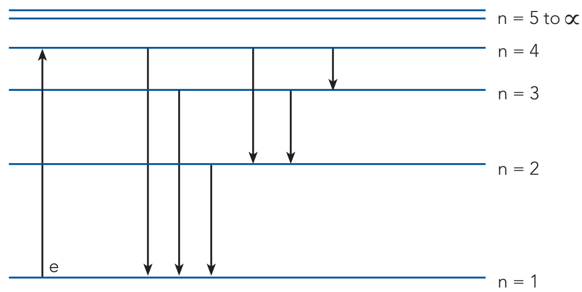
= 207.2

[12]

15.

(a) Energy has been absorbed.

(b) There are six unique downward transitions as shown on the diagram below.



(c) The n = 4 to n = 3 involves the smallest change in energy hence lowest frequency radiation and longest wavelength.

(d) Green light has the greater frequency of the two colours red and green. Hence since the n = 4 to n = 2 involves the greater change in energy it will result in a higher frequency radiation and green light.

[12]

16.

(a) The intense flame will atomise the sample, that is, form gaseous atoms.

(b) Any vaporised target metal atoms in the flame will absorb some of the specific wavelength light from the cathode lamp.

(c) Light from the cathode lamp is pulsed and easily distinguished from other light.

(d) The amount of absorbance, if any, of pulsed light from the cathode lamp indicates the presence and concentration of the metal being analysed for.

[12]

Total = 80 marks

TT 2: MATERIALS AND BONDING

Section 1

1. *a*
2. *d*
3. *b*
4. *d*
5. *b*
6. *d*
7. *d*
8. *c*
9. *b*
10. *c*

[20]

Section 2

11.

(a) ***elements:*** *oxygen, potassium*
compounds: *sugar, steam, ethanol, copper (II) oxide*
mixtures: *air, vinegar, limestone*

[3]

(b)
(i) *oxygen, potassium, sugar, steam, ethanol, copper(II) oxide*
(ii) *limestone*
(iii) *oxygen, potassium*

[3]

12.

filtration/evaporation: *solubility/boiling point*
distillation: *boiling point*
crystallisation: *relative solubility*
fractional distillation: *different boiling points*
chromatography: *solubility,selective adsorption*

[10]

13.

(a) *Zinc oxide and/or Titanium dioxide nanoparticles. They are good absorbers of UV radiation and form a transparent layer.*
(b) *Silver nanoparticles. They have antibacterial properties.*

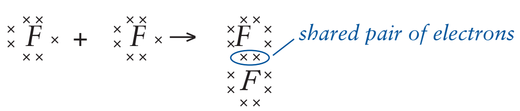
[6]

14. (a)
(i) Initial surface area
$(30 \times 20 \times 2) + (30 \times 10 \times 2) + (20 \times 10 \times 2)$
$= 2200\ cm^2$
(ii) Final total surface area
$10 \times [(30 \times 5 \times 2) + (30 \times 4 \times 2) + (5 \times 4 \times 2)]$
$= 5800\ cm^2$
Surface area change = 5800/2200 = 2.64
The greater surface area allows more contact with air (and oxygen). This allows more rapid burning to take place.

[10]

15.

(a) *The atoms bond to get a full outer energy level. They do this by sharing a pair of valence electrons.*



(b) *The bond is a result of both F nuclei attracting the same pair of electrons. This electrostatic attraction creates a very strong covalent bond.*
(c) *Covalent molecular solids are poor conductors because the covalent bond involves localised electrons that are not free to move. The covalent bond results in a sharing of electrons and not the formation of charged ions.*
(d) *An ionic bond would form between fluorine and sodium atoms. The valence electron of a sodium atom is loosely held and easily attracted by the electronegative fluorine atoms. This transfer of an electron creates a positively charged sodium ion and a negatively charged fluoride ion. An ionic lattice results.*

[12]

16.

METALLIC	IONIC	COVALENT MOLECULAR	COVALENT NETWORK
potassium	calcium oxide lithium bromide	ammonia sulfur dioxide water	diamond silicon dioxide

[4]

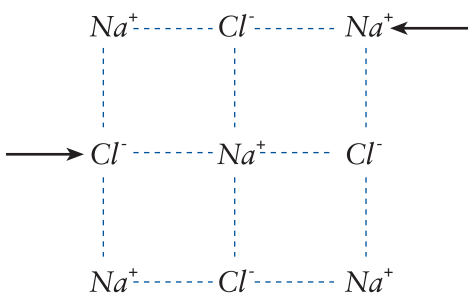
17.

(a) *An ionic bond forms when metal atoms with nearly empty valence shells donate their valence electrons to non-metal atoms with nearly full valence shells.*

e.g.

This electron transfer causes the formation of oppositely charged ions (i.e. Na becomes Na^+ *; Cl becomes* Cl^-*). An ionic bond is the attraction between these oppositely charged ions.*

(b) *Ionic solids are hard because the ionic lattice is held together by strong bonds making it difficult to displace the ions at the surface (i.e. making it difficult to scratch). Ionic solids are brittle because any distortion of its lattice will cause like charged ions to repel and shatter the lattice.*



If one layer is moved, similar charged ions are brought near to each other and the lattice *shatters due to great repulsive forces between like charged ions.*

(c) *The ions in an ionic lattice are unable to move freely as they are surrounded by oppositely charged ions in all directions. All valence electrons are strongly held and localised near each ion. Hence there is no way that current (charge) can be transferred.*

[12]

Total = 80 marks

TT 3: CARBON CHEMISTRY

Section 1

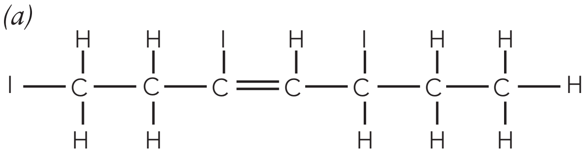
1. c
2. a
3. d
4. c
5. b
6. d
7. c
8. d
9. d
10. b

[20]

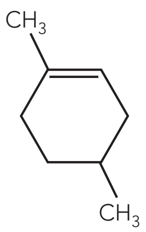
Section 2

11.
(a) *hexane*
(b) *1,2,4–trichloropentane*
(c) *1-chloro-3,3-dimethylbut-1-yne*
(d) *chloro-3-methylbut-2-ene*
(e) *3,6-dichlorocyclohexene*
(f) *1,2,4,-trifluorobenzene* [12]

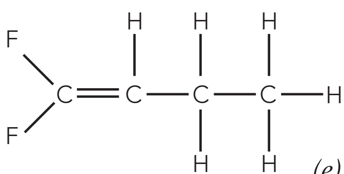
12.
(a)



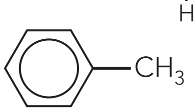
(b)



(c)



(d)



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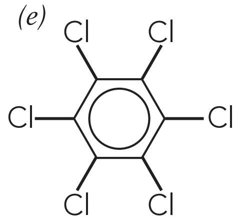

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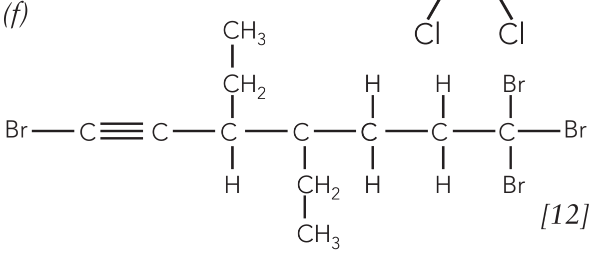


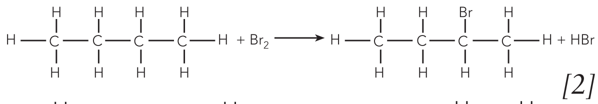
[12]

13.
(a) $C_6H_{12} + Br_2 \rightarrow C_6H_{11}Br + HBr$
yellow/brown solution turns colourless [3]

(b) $C_6H_{10} + I_2 \rightarrow C_6H_{10}I_2$
red/brown solution turns colourless [3]

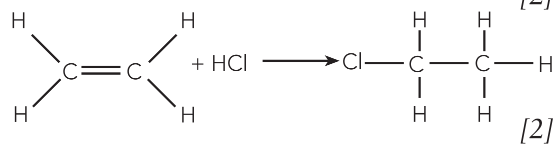
(c) $C_7H_{16} + 11O_2 \rightarrow 7CO_2 + 8H_2O$
liquid burns with yellow flame [3]

14.
(a)



[2]

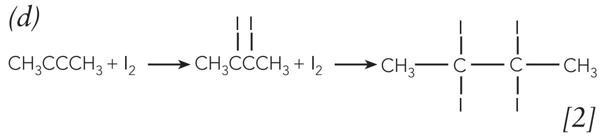
(b)



[2]

(c) $C_6H_6 + F_2 \rightarrow C_6H_5F + HF$ [2]

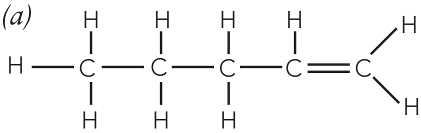
(d)



[2]

(e) $2C_3H_6 + 9O_2 \rightarrow 6CO_2 + 6H_2O$ [2]

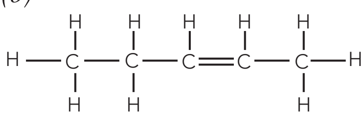
15.
(a)



pent-1-ene

[2]

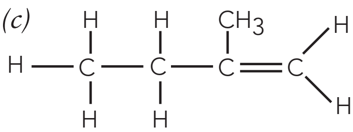
(b)



cis pent-2-ene
(also trans pent-2-ene)

[2]

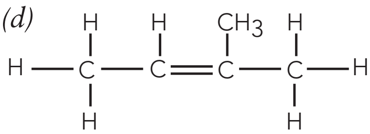
(c)



2-methylbut-1-ene

[2]

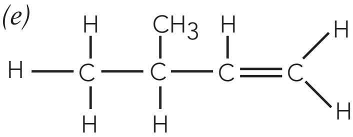
(d)



2-methylbut-2-ene

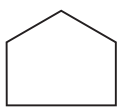
[2]

(e)



3-methylbut-1-ene

[2]



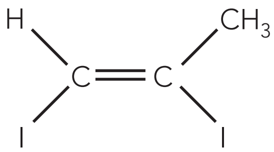
cyclopentane is a 6th possibility, and other cycloalkanes.

16.

(a) (i) *cis 1,2-dichloroethene* (ii) *trans but-2-ene* [1]

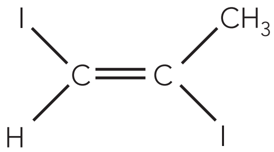
(b)

(i)



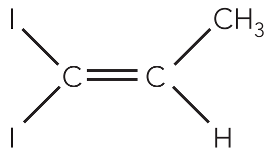
[2]

(ii)



[2]

(c)



[2]

Total = 80 marks

TT 4: CHEMICAL REACTIONS AND ENERGY CHANGE

Section 1

1. b	6. a
2. c	7. b
3. b	8. d
4. c	9. a
5. c	10. d

[20]

Section 2

11.

(a) *exothermic*

(b) *exothermic*

(c) *endothermic*

(d) *endothermic*

[4]

12.

(a) *bonds broken H-H (2), O=O (1)* [2]

(b) *bonds formed O-H (4)* [2]

(c) *bond forming since energy released* [3]

(d) *product bonds are stronger as more energy is released when they form than is absorbed when reactant bonds break.* [3]

13.

(a) *A chemical change. New products* CO_2 *and* H_2O *are formed.* [2]

(b) *propane + oxygen* $\rightarrow$ *carbon dioxide + water vapour* [2]

(c) $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$ [3]

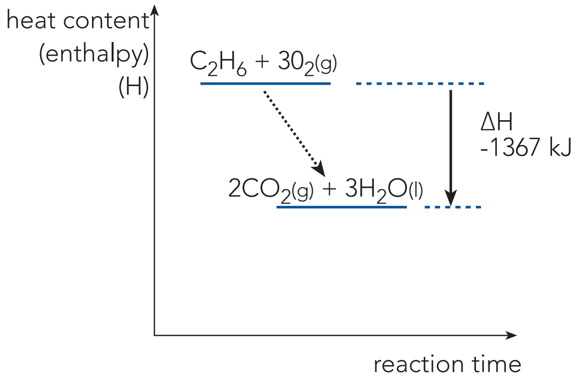
(d) $2C_8H_{18}(g) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$ [3]

14.

(a) $C_2H_6O(g) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l) + 1367\ kJ$

[2]

(b)



[8]

(c) $M(C_2H_6O) = (2)(12.01) + (6)(1.008) + (16.00) = 46.07\ g\ mol^{-1}$

Moles of C_2H_6O *in 1.0 kg* $= 1000/46.07 = 21.71\ mol$

Energy produced per kg $= -1367 \times 21.71 = -29.7\ MJ$

[4]

(d) $M(CO_2) = (12.01) + (2)(16) = 44.01\ g\ mol^{-1}$

$n(CO_2)/n(C_2H_6O) = 2$ *from equation*

$n(CO_2)$ *produced* $= 2 \times 21.71 = 43.42$

mass of CO_2 *produced* $= 43.42 \times 44.01 = 1.91\ kg$

[4]

15.

(a)

Methane

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

$\Delta H = -891\ kJ\ mol^{-1}$

Butane

$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$

$\Delta H = -2219\ kJ\ mol^{-1}$

Octane

$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(l)$

$\Delta H = -5471\ kJ\ mol^{-1}$

[6]

(b) $M(CH_4) = 12.01 + (4)(1.008) = 16.04\ g\ mol^{-1}$

Moles of CH_4 *in 1.0 kg*

$= 1000/16.04$

$= 62.34\ mol$

Energy produced per kg

$= -891 \times 62.34$

$= -55549\ kJ$

Energy per kilogram of methane $= 55.5\ MJ$

Similarly, for propane 50.3 MJ and for octane 45 MJ

Mole ratios (see equations and table).

Also to find n(CO_2)/MJ of energy

e.g. for Octane = $8 \times 1 \times 10^6/5471 \times 10^3$
= 1.46

FUEL	ΔH kJmol^{-1} 25 °C , 1 atm.	MJ of energy/ kg of fuel	n(CO_2)/ n(fuel)	n(CO_2)/ MJ of energy
Methane	-891	55.5	1	1.12
Propane	-2219	50.3	3	1.39
Octane	-5471	47.9	8	1.46

[9]

(c) *Natural gas can be readily piped from its source to the user with little treatment. It burns cleanly, has a high energy content and relatively low CO_2 emission.*

[2]

(d) *Natural gas can only be liquefied at very high pressures posing extra costs and risks. Propane is more easily liquefied and bottled as LPG.*

[1]

Total = 80 marks

TT 5: CHEMICAL CALCULATIONS

Section 1

1. b
2. c
3. c
4. c
5. c
6. d
7. d
8. a
9. b
10. b

[20]

Section 2

11.
(a) *These are relative atomic masses compared to a carbon-12 atom taken as 12.0. Specifically we can say that an oxygen atom (16.0) is 4 times the mass of a helium atom (4.0).*

[2]

(b) $M_r(H_2SO_4) = (2)(1.008) + 32.07 + (4)(16.0)$
$= 98.086$
$M_r(HCl) = 1.008 + 35.45 = 36.46$

[3]

(c) *% by mass for* H_2SO_4

$= \dfrac{2.016}{98.086} \times 100 = 2.055\%$

% mass for HCl

$= \dfrac{1.008}{36.46} \times 100 = 2.76\%$

i.e. HCl has the greater %H by mass.

[3]

12.
(a) $n(Au) = \dfrac{6.022 \times 10^{25}}{6.022 \times 10^{23}} = 100$ mol

(b) *number of Pb atoms* $= (0.25)(6.022 \times 10^{23})$
$= 1.51 \times 10^{23}$ *atoms*

(c) *molecules of* $H_2O = (2.5)(6.022 \times 10^{23})$
$= 1.51 \times 10^{24}$

(d) $n(SO_2) = \dfrac{64.0}{64.07} \approx 1.00$ mol

$\therefore$ *molecules* $(SO_2) = 6.02 \times 10^{23}$

(e) $n(NaCl) = cV = (0.0200)(0.250)$
$= 5.00 \times 10^{-3}$ mol
$\therefore n(Cl^-) = 5.00 \times 10^{-3}$ mol

[10]

13. $2NH_{3(g)} + 3Cl_{2(g)} \rightarrow N_{2(g)} + 6HCl_{(g)}$
23.5 L ?
from equation
2 mol NH_3 *gives 6 mol HCl*
$\therefore$ *2 Vol* NH_3 *gives 6 Vol HCl*
Hence volume of the HCl produced
$= 23.5 \times \dfrac{6}{2} = 70.5$ L

[8]

14.
(a) $H_2SO_4 + 3H_2S \rightarrow 4S + 4H_2O$

(b) *m(S) required* $= 37.6 \times \dfrac{100}{97.6} = 38.5$ g

n(S) required $= \dfrac{m}{M} \times \dfrac{38.5}{32.07} = 1.20$

$n(H_2S) = \frac{3}{4}n(S) = 0.75 \times 1.20 = 0.901$
$M(H_2S) = 2.016 + 32.07$
$= 34.086$ g mol^{-1}
$m(H_2S)$ *required* $= nM$
$= 0.901 \times 34.086$
$= 30.7$ g

[8]

15.
(a) $Na_2SO_{4(aq)} + BaCl_{2(aq)} \rightarrow BaSO_{4(s)} + 2NaCl_{(aq)}$
150 mL
2.0 mol L^{-1} ?

[2]

(b) $n(BaCl_2) = cV = (2.0)(0.150) = 0.30$ mol
$\therefore n(BaSO_4) = 0.30$ *mol (from equation)*
$\therefore m(BaSO_4) = (0.30)(137.3 + 32.07 + 64)$
$= 70.0$ g

[4]

(c) $c(Cl^-) = \dfrac{n \text{ (of } Cl^- \text{ ions in solution)}}{V \text{ (total vol of solution)}}$

Since $BaCl_{2(aq)} \rightarrow Ba^{+2}{}_{(aq)} + 2Cl^-{}_{(aq)}$

$n(Cl^-) = (2.0)(0.150)(2) = 0.60$ mol

$\therefore\ c(Cl^-) = \dfrac{0.60}{0.400} = 1.5 \text{ mol L}^{-1}$

[4]

16.

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

$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

(b) for CH_4:

$\% \text{ C} = \dfrac{12.01}{16.04} \times 100 = 74.9\%$

for C_3H_8:

$\% \text{ C} = \dfrac{(3)(12.01)}{44.09} \times 100 = 81.7\%$

Propane has the highest carbon content.

[6]

(c) Firstly find moles of each fuel

$n(CH_4) = \dfrac{1.00}{16.04} = 6.23 \times 10^{-2}$ mol

$n(C_3H_8) = \dfrac{1.00}{44.09} = 2.27 \times 10^{-2}$ mol

Next look at mole ratios (and hence volume ratios) for each reaction.

for CH_4: 1 Vol (CH_4) $\rightarrow$ 3 Vol (products)

for C_3H_8: 1 Vol (C_3H_8) $\rightarrow$ 7 Vol (products)

Hence from 1.0 g of each fuel

for CH_4:

$V(\text{products}) = (6.23 \times 10^{-2})(22.71)(3) = 4.25$ L

for C_3H_8:

$V(\text{products}) = (2.27 \times 10^{-2})(22.71)(7) = 3.61$ L

Methane produces the greatest volume of gaseous products.

[6]

Total = 80 marks

TT 6: INTERMOLECULAR FORCES

Section 1

1. d	6. c
2. d	7. a
3. c	8. a
4. b	9. b
5. d	10. d

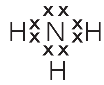
[20]

Section 2

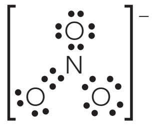
11. (a) ammonia (b) fluorine

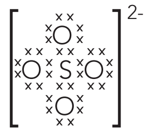
(c) any noble gas [6]

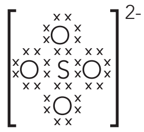
12.

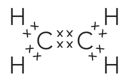
(a) 

(b) 

(c) 

(d) 

(e) 

(f) 

[12]

13.

NAME	SHAPE	POLAR OR NON-POLAR
chloroform	tetrahedral	polar
hydrogen sulfide	bent	polar
phosphorous trichloride	pyramidal	polar
dichlorine monoxide	bent	polar

[8]

14. H_2O has a higher BP than H_2S because H_2O exhibits hydrogen bonding while H_2S exhibits the weaker dipole-dipole forces. H_2S has a lower BP than H_2Se because it has a much smaller molar mass – dispersion forces are much weaker.

[6]

15. As the molar mass increases so does the strength of the dispersion forces. This is because dispersion forces increase with an increase in the number of electrons. Hence increased boiling points.

[6]

16. A substance will usually dissolve in water if the adhesive forces between the water molecules and the solute are similar in strength to the forces holding the solute molecules together. This is likely to be true for polar solutes as the negative region of the polar molecule will be attracted to the positive region around the hydrogen atoms of the water molecule (and the positive region to the negative region around the oxygen atom). Non-polar solutes tend not to dissolve in water as the hydrogen bonding holding water molecules to each other is much stronger than molecules to the water molecules, i.e. the water molecules do not attract the non-polar molecules and so the solute is not likely to break up and dissolve.

[6]

17.

(a) *Because of its shape (two atom linear) and the difference in electronegativity between C and O, CO has an uneven charge distribution and is polar. CO_2 is non-polar because even though the bonds are polar, the 3 atom linear shape gives an even or symmetrical charge distribution.*

[4]

(b) *CO_2 is a larger molecule than CO and so will have exhibit greater strength dispersion forces than CO. The strength of the dispersion forces in CO_2 are greater than the dipole-dipole forces of CO as shown by these properties.*

[2]

18.

(a) *The pencil line is a mixture of substances that will not dissolve in the mobile phase whereas the ink may dissolve.*

[2]

(b) *The lipsticks used were probably not very soluble in water indicating that they are non-polar. The ethanol/acetone mixture would be more successful at dissolving the lipsticks.*

[2]

(c) *The height to which the compounds travel up the chromatography paper is related to the strength of intermolecular attractions between the compounds and the molecules in the paper. The greater the height travelled the weaker the attractions for the molecules in the stationary phase.*

[2]

(d) *None of the lipstick brands appears to match the unknown as the heights the colours have travelled to do not match any of the samples.*

[2]

(e) *The boiling points of the coloured compounds would be too high to easily convert to a gas for the use of GC and so HPLC would be used.*

[2]

Total = 80 marks

TT 7: KINETIC THEORY AND GASES

Section 1

1. *c*	6. *a*
2. *d*	7. *a*
3. *b*	8. *a*
4. *b*	9. *d*
5. *a*	10. *c*

[20]

Section 2

11.

(a) *There exists a great deal of space between gas particles.*

(b) *Gas particles move in random straight line motion. Hence easily diffuse.*

(c) *Greater number of particles means that there will be more collisions per second with the walls of the balloon.*

(d) *Air particles in the tyre move about more rapidly hence collisions with the walls of the tyre are more frequent and more forceful.*

[12]

12.

(a) *Gas has been transferred from the small cylinder to the balloon. The balloon now contains more particles, this will cause more molecular collisions with the walls of the balloon. The increased number of collisions creates more pressure which will cause the volume of the balloon to increase.*

(b) *As there is no change in volume, the increased number of gas gas molecules inside the tyre increases the number of collisions with the walls of the tyre. This will cause the pressure inside the tyre to increase.*

[8]

13.

(a)

Equipment:

(i) *glass capillary tube with one end sealed and the other end having a mercury bead that traps a small but measureable volume of air.*

(ii) *thermometer.*

(iii) *beaker of water at room temperature.*

(iv) *bunsen burner.*

Method:

(i) *Place the capillary tube and thermometer in the water.*

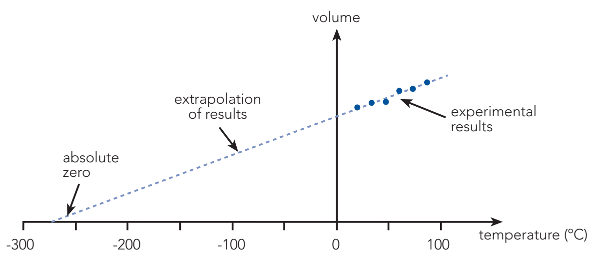
(ii) *Record the temperature of the water and the water and the volume of trap gas.*

(iii) *Heat the water and measure the volume of the trapped mass of gas for temperature 10 °C increments from 20° C to 80° C.*

Prediction: Draw a graph of volume of trapped air (vertical axis) versus temperature and extrapolate the graph to the predict the temperature at which the volume will be zero.

[6]

(b)



[6]

(c) *Absolute zero is the temperature at which molecules stop moving.*

[4]

14.

(a) *The balloon will expand in size. This is because there is an imbalance between the internal pressure in the balloon (unchanged) and the external pressure which is now lower. The expansion of the balloon will cause its internal pressure to be reduced. This occurs because the air particles have more space to travel in and hence there are fewer collisions with the internal wall of the balloon.*

[3]

(b) *Since the air is cooler the original pressure of the balloon will be reduced as it rises. This occurs because the air particles will move less quickly, there will be fewer collisions and less forceful collisions. The effect on our prediction that the balloon will expand as it rises will be that it will not expand as much. It will depend on how significant the effect of the temperature is. This can only be calculated using the gas laws (not in this course).*

[3]

(c)

(i) *Boiling occurs when the vapour pressure reaches atmospheric pressure. At 2000 m the atmospheric pressure is only 79.5 kPa and hence water boils at a significantly lower temperature. Hence the eggs just get very hot but don't cook.*

[3]

(ii) *Ideally if the climbers had a small pressure cooker it would allow water to boil at a higher temperature. Another possible solution would be to add large amounts of salt to the water as this would raise the boiling point.*

[3]

15.

(a) (i) -158° C (ii) 566°C (iii) 219°C

(b) $n(CH_4) = m/M = 2.95/16.042 = 0.184$

$V(CH_4) = n \times 22.71 = 4.18$ L

(c) $V = l \times w \times h = 10.0 \times 7.00 \times 2.80 = 196\ m^3$

$= 1.96 \times 10^5$ L

$n(O_2) = 0.21 \times 1.96 \times 10^5/22.71 = 1.81 \times 10^3$

$m(O_2) = n.M = 1.81 \times 10^3 \times 32.00 = 5.80 \times 10^4$ g

[12]

Total = 80 marks

TT 8: AQUEOUS SOLUTIONS

Section 1

1. b	6. b
2. b	7. d
3. a	8. a
4. b	9. d
5. c	10. c

[20]

Section 2

11.

(a) (i) *Dissociation is the separation of ions from an ionic solid into a solution.*

(ii) *Ionisation occurs when atoms from a molecular compound separate by becoming ions in solution.*

(b) (i) $KBr_{(s)} \rightarrow K^+_{(aq)} + Br^-_{(aq)}$

(ii) $HCl_{(g)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$

[8]

12.

(a) $Ba^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4\,(s)}$

A white precipitate forms.

(b) $Ca(OH)_{2(s)} \rightleftarrows Ca^{2+}_{(aq)} + 2OH^-_{(aq)}$

(c) $Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}$

A white precipitate forms.

[9]

13.

(a) *Electrolytes are substances which form ions in aqueous solutions.*

(b) (i) K_2CO_3 , $NaCH_3COO$, $MgSO_4$

(ii) H_3PO_4 , NH_3 , H_2CO_3 , H_2O

(iii) $C_{12}H_{22}O_{11}$, CH_3CH_2OH

[5]

14.

(a) *A saturated solution is a solution that contains as much dissolved solute as is possible to dissolve at that temperature.*

(b) *Salt lowers the freezing point of water and hence it is more difficult to freeze than pure water. The presence of the salt particles (Na^+, Cl^- ions) between the water molecules reduces the effectiveness of the attractive forces between water molecules.*

(c) *A salt solution contains a lot of free and mobile ions (Na^+, Cl^- ions) which allow current to flow. Sugar molecules in solution remain as neutral molecules and cannot conduct electricity.*

(d) $MgCl_{2(aq)} \rightarrow Mg^{2+}_{(aq)} + 2Cl^-_{(aq)}$

$NH_4Cl_{(aq)} \rightarrow NH_4^+{}_{(aq)} + Cl^-_{(aq)}$

The $MgCl_2$ releases more ions in solution and hence is a better conductor.

(e) *The attractive forces between salt particles (Na^+, Cl^- ions) and water molecules reduces somewhat the ability of water molecules to escape from the liquid. This reduces vapour pressure. Since boiling can only occur when vapour pressure is equal to atmospheric pressure a salt solution will raise boiling point.*

[10]

15.

(a) At 20°C solubility of NaCl ≈ 360 g/L

∴ mass (NaCl) in 5.0 L

≈ 360 × 5 ≈ 1800 g or 1.8 kg

(b) 100 g/0.250 L = 400 g/L

∴ crystals of NH_4Cl appear ≈ 27°C

(c) At 10°C solubility of KNO_3
≈ 200 g/L i.e. ≈ 40 g/200 mL
$\therefore$ mass (KNO_3) that will ppt
$\approx 100 - 40 \approx 60$ g

[12]

16.

(a)

(i) $c(BaCl_2) = \frac{2.55\ g}{0.250\ L} = 10.2\ gL^{-1}$

(ii) $n(BaCl_2) = \frac{m}{M} = \frac{2.55}{208.2}$

$= 1.224 \times 10^{-2} mol$

$c = \frac{n}{V} = \frac{1.224 \times 10^{-2}}{0.250} = 0.0490\ mol\ L^{-1}$

[6]

(b) 5.0% = 5.0 g/100 g of beer
= 5.0 g/100 mL (density = 1.0 g mL^{-1})

$\therefore$ mass of beer $= 5.0 \times \frac{330}{100}$

= 16.5 g

[4]

(c)

(i) % by mass $= \frac{155.5\ g}{5000\ g} \times 100 = 3.11\%$

(ii) ppm $= \frac{155500\ mg}{5.0\ kg} = 3.11 \times 10^4$ ppm

[6]

Total = 80 marks

TT 9: ACIDS AND BASES

Section 1

1. b
2. a
3. d
4. c
5. d
6. a
7. d
8. a
9. d
10. b

[20]

Section 2

11.

(a) Base, NaOH $\rightarrow Na^+_{(aq)} + OH^-_{(aq)}$
(b) Acid, $CH_3COOH \rightarrow CH_3COO^-_{(aq)} + H^+_{(aq)}$
(c) Base, $Ba(OH)_2 \rightarrow Ba^{2+}_{(aq)} + 2\ OH^-_{(aq)}$
(d) Acid, $H_2S \rightarrow HS^-_{(aq)}) + H^+_{(aq)}$

[8]

12.

(a) Beaker B: a big reduction in the number of bubbles produced compared to Beaker A.
(b) CH_3COOH is a weak acid and so a 1 mol L^{-1} solution of CH_3COOH will contain fewer H^+ ions than 1 mol L^{-1} HCl. The lower concentration of H^+ leads to a slower reaction rate.

[1]

13.

(a) (i) Acid + Metal Hydroxide → Salt + Water
(ii) Acid + Metal Carbonate → Salt + Water + Carbon Dioxide
(iii) Acid + Metal Oxide → Salt + Water

[3]

(b) (i) $CaCO_{3(s)} + 2H^+_{(aq)} \rightarrow Ca^{2+}_{(aq)} + CO_{2\ (g)} + H_2O_{\ (l)}$
(ii) $Mg(OH)_{2(s)} + 2H^+_{(aq)} \rightarrow Mg^{2+}_{(aq)} + 2H_2O_{\ (l)}$
(iii) $3HCO_3^-{}_{(aq)} + H_3PO_{4(aq)} \rightarrow PO_4^{3-}{}_{(aq)} + 3CO_{2(g)} + 3H_2O_{\ (l)}$
(iv) $SO_{2\ (g)} + H_2O_{(l)} \rightarrow H_2SO_{3\ (aq)} \rightleftharpoons H^+_{(aq)} + HSO_3^-{}_{(aq)}$
(v) $NH_4^+{}_{(aq)} + OH^-{}_{(aq)} \rightarrow NH_{3\ (aq)} + H_2O_{(l)}$
(vi) $2H^+_{(aq)} + Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + H_{2\ (g)}$
(vii) $Na_2O_{(s)} + 2H^+_{(aq)} \rightarrow 2Na^+_{(aq)} + H_2O_{(l)}$
(viii) $NH_{3\ (aq)} + H^+_{(g)} \rightarrow NH_4^+{}_{\ (aq)}$

[16]

14.

TURNS RED LITMUS BLUE	TURNS GREEN IN UNIVERSAL INDICATOR	HAS A PH GREATER THAN 7	HAS A PH LESS THAN 7
$Ca(OH)_2$	NaCl	$Ca(OH)_2$	HBr
Na_2O	$Ca(NO_3)_2$	Na_2O	SO_2

[8]

15.

(a)
1. 1 mol L^{-1} HNO_3
2. 1 mol L^{-1} CH_3COOH
3. 1×10^{-6} mol L^{-1} HCl
4. 1 mol L^{-1} NaCl
5. 1 mol L^{-1} Na_2CO_3
6. 1 mol L^{-1} KOH

[6]

(b) To neutalise the acidic soil a weak base such as Na_2CO_3 or $CaCO_3$ should be added.

[2]

16.

(a) A polyprotic acid has more than one H^+ ion that can be donated to a base.
e.g. $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$
the H_2SO_4 has 2 protons (H^+) that can be donated to the NaOH.

[3]

(b) The strength of an acid can be tested by testing its electrical conductivity. The greater the strength of the acid the better the solution will conduct electricity, at constant concentration.

[4]

(c) 1. $Al(OH)_3 + 3H^+ \rightarrow Al^{3+}(aq) + 3H_2O(l)$
2. $Al(OH)_3 + OH^- \rightarrow [Al(OH)_4]^-(aq)$
An amphoteric substance can react with acids and bases. Equation 1 slows $Al(OH)_3$ reacting with an acid while reaction 2 shows $Al(OH)_3$ reacting with a base.

[3]

(d) *Water is a weak electrolyte because only a very small fraction of molecules break up to form ions.*
in $H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$ *[3]*
1 litre of H_2O *contains* 1.00×10^{-7} *mol* L^{-1} H^+ *and* OH^- *ions at 25°C.*

Total = 80 marks

TT 10: REACTION RATES

Section 1

1. *d*	6. *b*
2. *a*	7. *d*
3. *d*	8. *b*
4. *a*	9. *a*
5. *d*	10. *d*

[20]

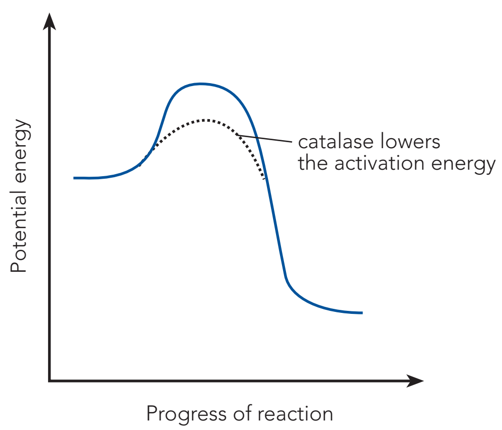
Section 2

11.
(a) $C_7H_{16} + 11\ O_2 \rightarrow 7\ CO_2 + 8\ H_2O$ *[4]*

(b) *By making into a fine mist the surface area of the petrol is greatly increased and this increases the likelihood of a successful collision occurring between the petrol and oxygen.* *[4]*

(c) *Increased gas pressure means reactant particles are much closer to each (increased concentration) and are far more likely to collide. Which increases the reaction rate.* *[4]*

12. *The catalase provides an alternate reaction pathway with a lower activation energy meaning that more collisions involving* H_2O_2 *molecules will result in the dispropotionation of the* H_2O_2 *to form* O_2 *and* H_2O.



[8]

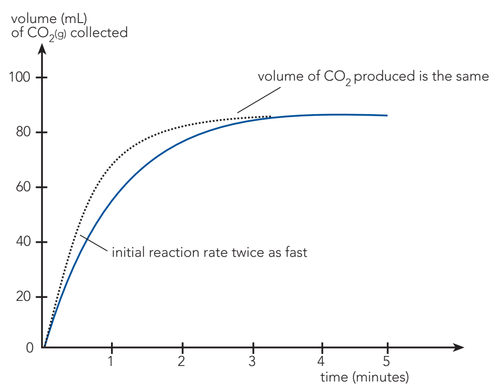
13. *Increased temperature increases the average kinetic energy of the reactant particles. This leads to more collisions occurring with an energy greater than activation energy and consequently an increased reaction rate.* *[6]*

14.
(a) *Granulated sugar has a much greater surface area per mass of sugar than does cubed sugar. The increased surface area allows greater contact with solvent and a greater rate of collisions at a particle level.*
(b) *The activation energy for the reaction is quite high. The flame allows the reaction to begin and as it is an exothermic reaction it becomes self supporting.*
(c) *Food spoilage is a chemical reaction. At lower temperatures the proportion of particles able to undergo a reaction is greatly reduced. Hence reaction rate is slowed and food keeps longer.*
(d) *Magnesium ribbon burns by combining with oxygen to produce magnesium oxide, light and heat. Air has only about 21% oxygen. In pure oxygen the rate of collisions between magnesium atoms and oxygen molecules is greatly increased and hence reaction rate is increased.*

[16]

15.
(a) $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$ *or*
$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$ *[2]*
(b) *Reaction rate is fastest at the beginning (steeper slope of graph).* *[2]*
(c) *As reactants are consumed their concentration or available surface area is reduced and hence reaction rate is reduced.* *[3]*
(d) *Reaction stops. One or both of the reactants has been totally consumed.* *[3]*
(e) *Easiest to measure pH. But could also monitor:*
- *mass of marble chips remaining,*
- Ca^{2+} *ion concentration,*
- H^+ *ion concentration or pH.* *[2]*

(f)
(i) *The reaction rate would be faster.*
(ii) *Concentration of* H^+ *ions greater,* ∴ *greater number of collisions/sec.*
(iii)



[6]

Total = 80 marks

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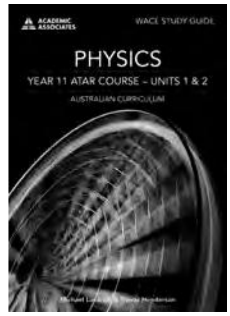

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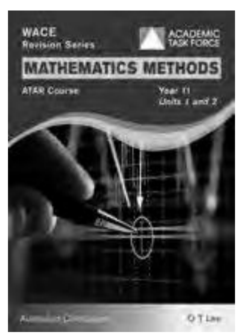

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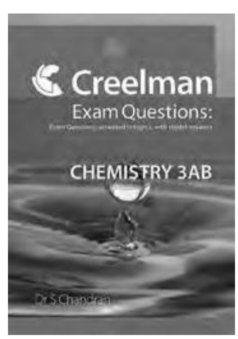

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