S215 Chemistry: essential concepts

Are you ready for S215?

Contents

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2.1

1 Introduction

If you are intending to study S215, you should make sure that you have the necessary background knowledge and skills to give yourself the best possible chance of completing the module successfully.

Read through these preparatory notes carefully, and work through the questions. It will be assumed that students taking S215 will be familiar with the science content of S104 and the maths content of S141 (or S111 and S112). Working through these notes will serve as a reminder of some of the facts, skills and conceptual knowledge that are required. **If you are coming to S215 without having studied these modules then it is essential that you establish whether or not your knowledge and experience give you a sound basis on which to tackle this Level 2 Course**.

All S215 material is delivered via the module website which is optimised for use on Internet Explorer (9 and 10), Firefox, Safari and Chrome on Windows 7 to 10.

Versions of the module for iPAD, Android tablet and Kindle are available but note that the module contains interactive material which is best studied online, although the majority is available on the iPAD version. .pdf and Word optimised for screen reader versions are available to download.

2 Study skills

You will find it more efficient if you have acquired the following skills, many of which are not covered in the quiz.

Basic study skills: organising your study time; pacing your study; effective reading to identify and extract relevant information from irrelevant or redundant material; retrieving data from scientific texts and accounts.

Writing skills: writing coherently; structuring and presenting arguments in a logical sequence often based on a range of sources, including published papers, websites, tables and diagrams; writing a scientific account with appropriate diagrams.

Cognitive skills: recognising trends and patterns in data; using evidence to support or refute theories and arguments; assessing the adequacy/limitations of explanations.

Problem-solving skills: solving problems using given evidence (including negative evidence), and using more than one source of information.

Computer skills: word processing and spreadsheet skills; familiarity with Windows operations and web browsers; manipulating diagrams for inclusion in TMAs; familiarity with online forums; searching online for data.

3 Suggested prior study

S104 Book 4. Book 3 Chapters 4, 6 - 11, Book 7 Chapter 3, Maths Skills **and**

S141 Maths for Science Chapters 1, 2, 3, 4, 5, 6, 7, 9, 10.

S111 Topics 1, 3, 4, 5, 8, 9, 10.

If you have not previously studied with the OU, the relevant background may be found in:

Any general ‘A’ level Chemistry textbook, such as:

A-Level Chemistry - Core Text Fourth Edition Eileen Ramsden, Nelson Thornes.(2000) ISBN 978-0-7487-5299-7.

Or a general chemistry text such as

Chemistry Rob Lewis and Wynne Evans, Palgrave Macmillan Ltd. (2011) ISBN 9780230291829 (print) 978023044921 (e-book)

These are by no means the only options; you should be able to start S215 if you have studied chemistry to one of the following levels: A level, Certificate of Sixth Year Studies or Advanced Higher (Scotland), or HNC. In addition you should have studied mathematics to at least GCSE level or equivalent.

4 Key science concepts for S215

This Section outlines the key concepts that you will need to understand before embarking on S215. Important terms that are central to S215 have also been highlighted in **bold type**.

As a general rule, if you are able to answer the questions correctly you have a sound basis on which to successfully complete S215. If after working through the questions you are still unsure about whether or not S215 is the right course for you, we advise you to seek further help and advice from the STEM Student Support Team.

4.1 Elements and compounds

An **element** is a substance made up of only one type of basic building block. Elements can be classified as metals, non-metals or metalloids.

Around 80% of elements are classified as metals. In most metals the atoms pack together as closely as they can. A feature of metals is that the some electrons do not remain in the proximity of a specific nucleus. In bulk metals, these electrons, rather than being associated with any particular metal atom, can be thought to be part of a shared ‘sea’ of electrons that move freely. The attraction between these electrons and the positively charged metal ions constitutes metallic bonding.

Elements combine with each other to form **chemical compounds**, in which the proportions of atoms of the different elements are in a simple ratio. This ratio may be established by analysis of the resulting compound. Elements can be represented by chemical symbols e.g H for hydrogen and O for oxygen. For compounds, the ratio of the different elements are represented by subscript numbers after the element symbol, e.g. H_2O for water.

The relative combining power of an element in a compound is governed by its **valency**. This can be defined as the combining power of the element. In H_2O, H has a valency of 1 and O has a valency of 2.

The **empirical formula** of a compound is a formula with the lowest whole number values of the subscripts consistent with the ratio of the elements in the compound. For example the empirical formula of hydrazine, N_2H_2 is NH.

Question 1

(a) Give the name of the element and the number of protons for Mg^{2+}.

(b) Is this element a metal, non-metal or semimetal?

(c) Is Mg^{2+} a neutral atom, an anion or a cation?

Question 2

(a) In its oxides, sulphur exhibits valencies of +4 and +6. What are the empirical formulae of these oxides (assuming a valency 2 for oxygen)?

(b) What is the valency of phosphorus in the compound P_2O_5?

4.2 Atomic structure and chemical bonding

You should be familiar with the idea of 'matter' being composed of **atoms**. The core of an atom is its **nucleus**, containing **protons** and **neutrons**. These two types of particle have approximately the same mass. The sum of the masses of the protons and neutrons in the atom constitutes the **atomic mass**. The nucleus is surrounded by negatively charged particles of comparatively negligible mass, known as **electrons**. Any chemical element is characterized by its **atomic number**, ***Z***, where *Z* is the number of protons in the nucleus.

Atoms of the same element that differ only in the number of neutrons they contain are known as **isotopes** of that element, and have different atomic masses to other isotopes of the same element. The mass of any atom is expressed by the **relative atomic mass, ram**, which compares the mass of that atom to the atomic mass of the isotope of carbon with six neutrons (^{12}C). On this scale, the relative atomic mass of ^{12}C is defined as 12.000.

The electrons surrounding a nucleus occupy **energy levels** (or shells), described by their **principal quantum number**, ***n***.

For all elements, electrons fill the energy levels in a well-defined order from the lowest energy level up.

Chemical bonding involves interaction of the outer-shell (or 'valence') electrons of atoms. There are two main types of chemical bonding. Atoms can bond with other atoms by, either *transferring* outer-shell electrons to form positively and negatively charged ions (**ionic bonding**), or *sharing pairs* of electrons (**covalent bonding**).

The sharing of an electron pair is not equal if there are significant differences in **electronegativity** between the two atoms involved. In such cases, one atom has a greater 'share' of the electron pair in a covalent bond, thereby developing a slight negative charge (electrons are negatively charged, remember); the corresponding 'electron-poor' atom, which has a small deficiency in negative charge, carries a slight positive charge. The molecule so formed is **polar**, such as water.

In ionic bonding there is a transfer of electrons from one element to another. Thus, the ionic compound $A^+ B^-$ contains **positive ions**, A^+, and negative ions, B^-.

Positive ions are called **cations**, and negatively charged ions are **anions**.

The formation of ions from the atoms can be notionally represented as follows:

$$A \longrightarrow A^+ + e^-$$

$$B + e^- \longrightarrow B^-$$

Ionic bonding occurs when the atoms concerned have very different electronegativities: **metals** from the left-hand side of the Periodic Table combine with **non-metals** from the right-hand side of the Table. In the above example, if A is Na and B Cl, then the ions A^+ and B^- each have a **noble gas** electron configuration. Solid ionic salts are made up of regular arrays of positive and negative ions.

Question 3

Predict whether the following will be ionic compounds or covalently bonded molecules: KBr, PH_3, CO_2 and $Ca(OH)_2$.

The energy level (shell) corresponding to a given value of n contains sub-levels, or sub-shells, which are described by the **orbital quantum number *l***. For a given value of n the orbital quantum number, l, may take any whole number from 0 up to $n – 1$. Sub-levels, with $l = 0, 1, 2, 3$ are labelled s, p, d and f.

For a given value of the orbital quantum number, l, there are $(2l + 1)$ different values of the **magnetic quantum number, m_l** , ranging from $–l$ to $+l$. Thus there is one s level, three p levels five d levels and seven f levels.

There are two possible values of the **spin quantum number, m_s**.+½ and −½.

So s, p, d, f levels are therefore capable of holding 2, 6, 10 and 14 electrons, respectively.

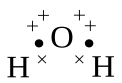
Each electron in the atom must have a unique set of these four quantum numbers.

If the atomic number of a particular element is known, it is possible to establish the occupation of these energy levels in an atom of that element – its **electron configuration**. This configuration uniquely defines the position of that element in the **Periodic Table of elements**.

Question 4

What is the largest number of electrons that can be accommodated by filling the 1s, 2s, and 2p subshells?

Non-metals can also bond covalently by sharing electrons between atoms to form electron-pair bonds. This sharing of electrons can be depicted in **Lewis structures**, like that shown for water below. In this structure, crosses are used to represent the outer-shell electrons of the oxygen atom, and black spots are used to represent the electrons from the hydrogen atoms.



The common (that is, most important) **valency** of a non-metal is equal to the number of covalent bonds it can form in order to obtain the electron configuration of the next higher noble gas. For example, fluorine ($Z = 9$) has an electron configuration $1s^22s^22p^5$. It requires one electron to acquire the noble gas electron configuration of neon, $1s^22s^22p^6$. To do this, it can form one covalent bond, and it therefore has a valency of one.

Fluorine has *seven* electrons in its outer principal shell ($n = 2$), and it is therefore located in the same column of the Periodic Table as chlorine. Thus, valency, like electron configuration is related to an element’s position in the Periodic Table.

Question 5

(a) Draw the Lewis structures for the following molecules (the atomic numbers of the elements concerned are F = 9; O = 8; C = 6):

(i) OF_2

(ii) CF_4.

(b) Show how the participating atoms attain stable noble gas electron configurations in the following ionic compounds (the atomic numbers of the elements concerned are Mg = 12; O = 8; Ca = 20; F = 9):

(i) MgO

(ii) CaF_2.

4.3 Periodic classification

The chemical elements can be arranged in an ordered form, known as the **Periodic Table**, based on their respective **electron configurations**; this order is found to be compatible with the similarity of the chemical properties of elements in columns of the Periodic Table. For example, sodium (whose atomic number, *Z*, equals 11) has the electron configuration $1s^22s^22p^63s^1$. Lithium (electron configuration $1s^22s^1$) and potassium (electron configuration $1s^22s^22p^63s^23p^64s^1$) both have similar chemical properties to sodium, because each of these elements contains one electron in its outer (‘s’) shell. Because of their similar electron configurations, the three elements are each placed in the same column of the Periodic Table.

In the long form of the Periodic Table, elements are distinguished between **main Group elements**, **transition elements**, lanthanoids and actinoids.

Question 6

(a) A chemical element has the atomic number 13. What is its:

(i) electron configuration (in s, p, d notation)

(ii) most common valency?

(b) Repeat the exercise for the element whose atomic number is 20.

Question 7

Give the full electronic configuration of Mg^{2+} and S^{2-}

Question 8

Two isotopes of boron (boron-10 and boron-11) are found naturally in the environment in the proportions of 19.9% and 80.1%. Calculate the average relative atomic mass and give your answer to one decimal place.

4.4 The mole

The **mole** is used to quantify the number of atoms or formula units involved in a reaction or contained in a volume of gas or solution. The relative atomic mass in grams of every element contains exactly the same number of atoms. Consequently, the mole is defined as the amount of substance containing the same number of particles as there are atoms in exactly 12 g of ^{12}C. This number is called Avogadro’s constant, and it has a value 6.022×10^{23} mol^{-1}.

The mass of a mole of a substance is obtained by adding together the relative atomic masses of all of the atoms in the formula unit, and following this by the unit of mass, g (grams).

Concentrations of solutions are commonly expressed as the number of moles of solute dissolved in 1 dm^3 of solution (mol dm^{-3}).

Question 9

9.8 g of sulfuric acid (H_2SO_4) is dissolved in water so that the total volume is 250 cm^3. Express the concentration of the resulting solution in mol dm^{-3} (the relative atomic masses of the elements concerned are H = 1, S = 32; O = 16).

4.5 Chemical equations

Chemical equations are written showing the **reactants** on the left-hand side and the **products** on the right-hand side. These are connected by an arrow indicating the direction of the reaction, as in Equation 1. If a chemical equation is **balanced**, in other words, if the numbers of each type of atom, and the overall electric charge, are exactly the same on both sides of the equation, then the arrow is replaced by an equals sign (as in Equation 2).

$$Na(s) + Br_2(l) \rightarrow NaBr(s) \qquad (1)$$

$$2Na(s) + Br_2(l) = 2NaBr(s) \qquad (2)$$

On each side of the balanced Equation 2, there are two sodium atoms and two bromine atoms, and an overall electric charge of zero. The physical **states** of reactants and products are denoted by suffixes: (s) for solid, (l) for liquid, (g) for gas and (aq) for aqueous solution.

Question 10

Balance the following equation for the combustion of heptane, C_7H_{16}.

$$C_7H_{16}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g) \qquad (3)$$

4.6 Dissociation in solution

An **acid** may be defined as a substance that dissociates to give hydrogen ions in water:

$$HX(aq) = H^+(aq) + X^-(aq) \qquad (4)$$

The equilibrium constant for this dissociation is represented as

$$K_a = \frac{[H^+(aq)][X^-(aq)]}{[HX(aq)]} \qquad (5)$$

where the brackets [ ] indicate ‘concentration of’. Acids with very high values of the acid dissociation constant, K_a, are known as **strong acids**, and are virtually completely dissociated. Examples are nitric acid, HNO_3, and hydrochloric acid, HCl. Acids with low values of K_a are only slightly dissociated, and are known as **weak acids**. An example of a weak acid is ethanoic (or acetic) acid, CH_3COOH, for which $K_a = 1.8 \times 10^{-5}$ mol dm^{-3}. Such a low value of K_a means that the equilibrium (5) lies well to the left in the case of ethanoic acid.

Dissociation of water

Water is only partially dissociated into aqueous hydrogen ions and aqueous hydroxide ions:

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq) \qquad (6)$$

The vast majority of water molecules remain undissociated. In other words, the equilibrium in Reaction 6 lies well to the left, which means that the concentration of water molecules remains effectively constant. The product of the concentration of hydrogen ions and the concentration of hydroxide ions present in water is also a constant, denoted by $\boldsymbol{K_w}$**,** where

$$K_w = [H^+(aq)][OH^-(aq)] = 1.0 \times 10^{-14}\ mol^2\ dm^{-6}\ (at\ 25\ °C)$$

As equal numbers of $H^+(aq)$ and $OH^-(aq)$ ions are present in pure water, the concentration of hydrogen (and hydroxide) ions in pure water is 1×10^{-7} mol litre^{-1}. This is the basis of the **pH scale**, in which a hydrogen ion concentration of 1×10^{-n} mol litre^{-1}, is expressed as a pH of $+n$. The pH of pure water is therefore 7. Acid solutions have a pH less than 7, and alkaline solutions a pH greater than 7.

Question 11

0.05 mol of NaOH is dissolved in pure water at 25 °C to give a total volume of 0.5dm^3. Calculate the pH of the resulting solution.

Question 12

Are the following solutions acidic, neutral or basic? Give your reasons.

(i) NaOH of concentration 1.0×10^{-2} mol dm^{-3}.

(ii) HCl of concentration 1.0×10^{-5} mol dm^{-3}.

(iii) NaCl of concentration 1.0×10^{-3} mol dm^{-3}.

(iv) CH_3COOH (ethanoic, or acetic, acid) of concentration 1.0 mol dm^{-3}.

4.7 Oxidation and reduction

Oxidation is said to occur when the proportion of oxygen in a compound increases. Conversely, **reduction** occurs when the proportion of oxygen decreases or hydrogen is added. The term oxidation is also used more generally to include any reaction in which an atom loses electrons. Conversely, reduction involves a gain of electrons. For example, on going from left to right in Reaction 7, a magnesium ion is reduced to magnesium metal by gaining two electrons. At the same time, a molecule of hydrogen gas loses two electrons, and by so doing is oxidized to hydrogen ions.

$$Mg^{2+}(aq) + H_2(g) = Mg(s) + 2H^+(aq) \qquad (7)$$

In a reaction, the oxidation of one species must be accompanied by a corresponding reduction of another species involved in the reaction.

Question 13

For the following reactions, state which species are being oxidized and which are being reduced:

(i) $Sn^{2+}(aq) + 2Fe^{3+}(aq) = Sn^{4+}(aq) + 2Fe^{2+}(aq)$

(ii) $Mg(s) + 2Ag^+(aq) = Mg^{2+}(aq) + 2Ag(s)$

(iii) $Zn(s) + 2H^+(aq) = Zn^{2+}(aq) + H_2(g)$

4.8 Electromagnetic radiation

Spectroscopy is widely used in chemistry and in S215 you will learn about several types of spectra using radiation from several region of the electromagnetic spectrum..

The **wavelength** of a wave is the distance between two neighbouring, equivalent points on the wave profile and has the symbol λ. The **frequency** of a wave is the number of cycles of the wave that pass a given point in one second and has the symbol f. It has the unit Hz $\equiv$ s^{-1}.The speed of any wave, v , is related to its frequency and wavelength by the equation $v = f\lambda$. The wavelength and frequency of an **electromagnetic wave** are related by the equation $c = f\lambda$ where c is the speed of light and has a value of 3.00×10^8 m s^{-1} to three significant figures.

Waves may be **diffracted** by apertures whose size is similar to the wavelength of the wave.

Different regions of the **electromagnetic spectrum** are distinguished by the different wavelengths and frequencies of the radiation. **Radio waves** have the longest wavelength and the lowest frequency. Moving towards shorter wavelengths and higher frequencies, we have **microwaves, infrared radiation**, light, **ultraviolet radiation**, and **X-rays**. **Gamma rays** have the shortest wavelength and highest frequency.

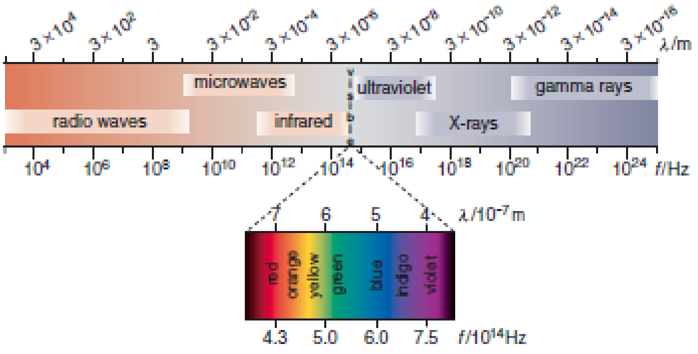


Figure 1 The electromagnetic spectrum

Question 14

Five complete cycles of a water wave travel past a fixed point in two seconds.

What is the frequency of the wave?

Question 15

A red laser beam consists of light with a longer wavelength than that of a green laser beam. What does this tell you about the relative frequencies of red light and green light?

Question 16

(a) A microwave oven operates at a frequency of 2.45 GHz. What is the wavelength of these microwaves?

(b) Technetium-99m, commonly used as a radioactive tracer, produces gamma rays of wavelength 8.8×10^{-12} m. What is their frequency?

(You should assume that the speed of electromagnetic radiation, c , is 3.00×10^8 m s^{-1})

4.9 Atomic spectroscopy

In S215 it is assumed that you are familiar with atomic spectroscopy and you will build on this to study other types of spectroscopy.

The energy of electrons can only have certain specific values in atoms. The electron energies are said to be **quantized**, and the electrons are said to be in particular **energy levels**. If an electron moves from one level to another, there is a change of energy, ΔE, and an associated **absorption** or **emission** of **electromagnetic radiation** occurs such that $\Delta E = h\nu$, where h is the **Planck constant** (whose value is 6.626×10^{-34} Js), and ν is the frequency of the radiation absorbed or emitted.

Question 17

In the hydrogen atom, the frequency of light emitted when an electron falls from level $n = 2$ to $n = 1$ is 2.47×10^{15} Hz, and from $n = 3$ to $n = 1$ the frequency is 2.92×10^{15} Hz. What is the energy of the photon emitted when an electron falls from level $n = 3$ to $n = 2$?

4.10 Schrödinger model of the hydrogen atom

A simple and appealing picture of an atom is of electrons orbiting around the nucleus like planets around the Sun. If electrons really did orbit the nucleus, then they would be constantly accelerating. Any charged object undergoing an acceleration will continuously emit electromagnetic radiation and lose energy. As a result, an electron orbiting in this way would rapidly spiral into the nucleus as its electrical energy reduced. This does not happen, so electrons cannot really be orbiting the nuclei of atoms.

The **Schrödinger model** of the hydrogen atom says that we cannot determine the exact position of the electron at a given time. It is useful to represent the possible positions of electrons in atoms by ‘fuzzy clouds’ surrounding the nucleus. Different clouds will correspond to each quantum state. The model provides definite predictions for the energy levels, but only the probability of the electron’s position and velocity .

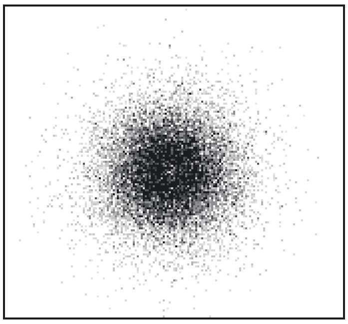


Figure 2 The Schrodinger model of the hydrogen atom says that the position of the electron is indeterminate, so a hydrogen atom in the 1s energy level is represented here by a ‘fuzzy cloud’ surrounding the nucleus, indicating the range of possible positions that the electron may have. The cloud in this case is spherical, with the nucleus at its centre.

Question 18

How, if at all, will the view of the hydrogen atom in the 1s state vary with the angle at which it is viewed?

4.11 Energy changes and chemical reactions

All **chemical reactions** involve changes in **internal energy**.

Consider the reaction

A + B = C

If heat is evolved to the surroundings, then the reaction is said to be **exothermic**; by convention the **enthalpy of reaction**, ΔH, for such a reaction is negative. If heat is absorbed, the reaction is **endothermic** and the value of ΔH for the reaction is positive.

Every chemical reaction has an **energy barrier** that must be overcome before the reactant molecules can form products. This is depicted in Figure 1. In order to react, the reactant molecules must have sufficient internal energy to surmount this barrier. Heating the reaction, to increase the internal energy of the molecules, or using a **catalyst** (which provides an alternative reaction route via an intervening stage of lower energy), will increase the **rate of reaction**. Using a catalyst will not, however, affect the position of the equilibrium.

Question 19

(a) Use the table of bond enthalpies below, Table 1, to calculate the overall molar enthalpy change for the reaction

$CH_4(g) + 2Cl_2(g) = CCl_4(g) + 2H_2(g)$

Table 1 Bond enthalpies.

Bond	**Molar bond enthalpy/kJ mol^{-1}**
C–H	413
Cl–Cl	242
C–Cl	346
H–H	436

(b) Is this reaction endothermic or exothermic?

Every chemical reaction has an **energy barrier** that must be overcome before the reactant molecules can form products. This is depicted in Figure 3. In order to react, the reactant molecules must have sufficient internal energy to surmount this barrier. Heating the reaction, to increase the internal energy of the molecules, or using a **catalyst** (which provides an alternative reaction route via an intervening stage of lower energy), will increase the **rate of reaction**. Using a catalyst will not, however, affect the position of the equilibrium.

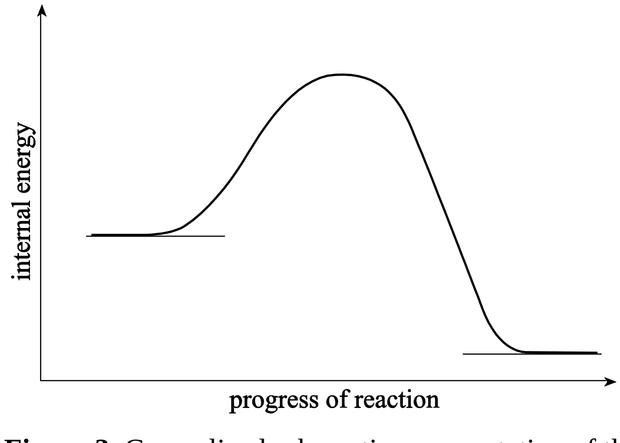


Figure 3 Generalized schematic representation of the course of a reaction.

4.12 Chemical equilibrium and Le Chatelier's Principle

Chemical equilibrium represents a state of dynamic balance. In general, for a chemical reaction that proceeds to equilibrium, we can write

$$\text{Reactants} \rightleftharpoons \text{products}$$

The $\rightleftharpoons$ symbol indicates that the reaction takes place in both directions; in other words, the reaction is reversible. At equilibrium the forward reaction (by convention, the one going from left to right as written) and the reverse reaction (the one going from right to left) occur at the same rate, and consequently, there is *no further change* in the relative amounts of products and reactants present in the reaction mixture. The proportion of a reaction product in an equilibrium mixture is known as the **equilibrium yield** of that substance.

Le Chatelier's principle refers to the effect of external constraints (such as a change of pressure or temperature, or an increase in the concentration of one of the species involved in the reaction) on the equilibrium position of a reaction. It states that:

> when a system in dynamic equilibrium is subject to an external constraint, the system responds in such a way as to minimize that constraint.

Question 21

The reaction

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad (8)$$

is exothermic in the forward direction. How will the equilibrium yield of ammonia be affected by:

(i) increasing the pressure of the reaction

(ii) increasing the temperature of the reaction

(iii) using a catalyst, and

(iv) removing the ammonia product as it is formed?

Question 22

Consider the reaction

$$I_2(g) + Cl_2(g) \rightleftharpoons 2ICl(g)$$

for which ΔH is negative.

Which of the changes listed below leads to an increase in the equilibrium yield of ICl(g):

(i) lowering the temperature of the reaction

(ii) raising the pressure of the reaction

(iii) lowering the pressure of the reaction?

4.13 The chemistry of carbon

Carbon-containing compounds – other than CO, CO_2, carbonates, cyanides and carbides – are referred to as **organic compounds**. Carbon is in Group IV of the Periodic Table and it therefore has a valency of 4. In uncharged organic compounds it always forms four covalent bonds. A significant feature of the chemistry of carbon is its ability to form chains of carbon atoms, producing families of compounds, the members of which have similar chemical properties.

Crude oil is the major source of organic compounds, in particular **hydrocarbons**, which contain only hydrogen and carbon atoms. The **alkane** family of hydrocarbons are fully **saturated** and contain only single C−H and C−C bonds.

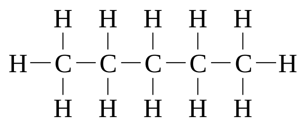
Alkenes are **unsaturated** compounds containing one or more C=C **double bonds**. Both alkanes and alkenes may be **linear**, **branched-chain** or **cyclic**.

The composition and structures of organic molecules can be indicated in a number of ways, depending on the information to be portrayed: these include **molecular formulae** and **structural formulae** (showing *all* of the bonds between atoms).

Compounds with the same molecular formula but different structural formulae are called **isomers**.

Question 23

The hydrocarbon C_5H_{12} has three structural isomers. The straight-chain isomer is



which may be abbreviated to

$CH_3CH_2CH_2CH_2CH_3$

Write out **abbreviated structural formulae** for the two branched-chain isomers of C_5H_{12}.

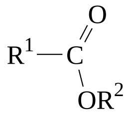
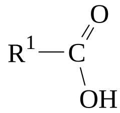
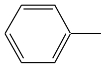
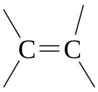
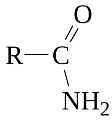
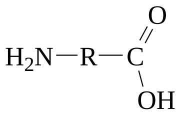
Question 24

In methane, CH_4, there are four CH bonds. How are these bonds arranged in three dimensions?

4.14 Functional groups

Many reactions of organic molecules can be predicted from a knowledge of the **functional groups** they contain. You should be able to recognize the functional groups in Table 2, as well as some of their characteristic reactions. For example, **condensation reactions** between carboxylic acids and alcohols or amines give esters or amides, respectively. The oxidation reactions (defined here as a reaction increasing the proportion of oxygen, or decreasing the proportion of hydrogen, in the product) of alcohols should also be familiar, as should the **reduction** (increasing the proportion of hydrogen, or decreasing the proportion of oxygen in the product) of carboxylic acids, the **hydrolysis** of esters, amides and polyamides (for example, **proteins**), and the **addition reactions** of alkenes.

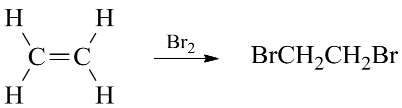
Table 2 Some of the functional groups you should be familiar with..

alcohol	$-OH$	ester	
carboxylic acid		aromatic	
alkene		amide	
haloalkane (X = F, Cl, Br or I)	$R-X$	amino acid	
amine	$-NH_2$		

In Table 2, the symbols R, R^1 and R^2 are used to represent different hydrocarbon groups, or the rest of the molecule.

Question 25

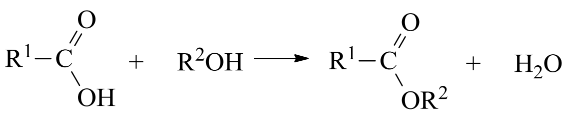
Classify each of the following reactions as either a condensation, oxidation, reduction, hydrolysis or addition reaction.



(9)

$$RCH_2OH \xrightarrow[H_2SO_4]{K_2Cr_2O_7} RCOOH \qquad (10)$$

$$CH_3COOH \xrightarrow{H_2/\text{catalyst}} CH_3CH_2OH \qquad (11)$$



(12)

5 Answers to questions

Question 1

(a) Mg is the symbol for the element magnesium. Mg has an atomic number of 12. It therefore has 12 protons. The ion Mg^{2+} will also have 12 protons. It differs from Mg by having two less electrons.

(b) Magnesium is a metal.

(c) Mg^{2+} is a cation.

S104 Book 4 Chapter 3

S111 Topic 5 Taking it further

Question 2

(a) SO_2, SO_3.

(b) The valency of phosphorus in $P_2O_5 = 5$.

S104 Book 4 Chapter 5 Section 5.4.

S111 Topic 5 Taking it further

Question 3

KBr is made up from a metal, potassium, and a non-metal, bromine. The bonding will be ionic; in other words, it can be formulated as
$K^+ Br^-$.

PH_3 is made up from two non-metals. The molecule involves covalent bonding between hydrogen with a valency of 1, and phosphorus with a valency of 3. The bonding will be covalent.

CO_2 is made up from two non-metals: oxygen has a valency of two, and carbon is in Group IV with a valency of four. The bonding will be covalent.

$Ca(OH)_2$ is made up from a metal, calcium, and OH^-, a polyatomic ion. The bonding between the O and H will be covalent but the bonding between the Ca^{2+} and OH^- will be ionic. The compound can be represented as $Ca^{2+}(OH^-)_2$.

S104 Book 4 Chapter 3

S111 Topic 1 Parts 1 and 3.

Question 4

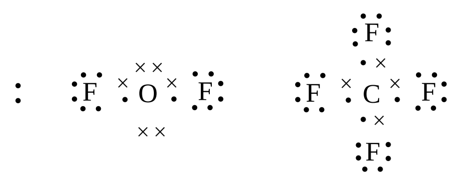
s subshells can take two electrons ($m_s = +½$ and $m_s = –½$) , p subshells can take 6 electrons. The total number for 1s, 2s and 2p is thus 10 electrons.

S104 Book 4 Chapter 5 Section 5.2, Book 7 Chapter 3 Section 3.6

S111 Topic 5 Taking it further

Question 5

(a)



(ii) (iii)

(b) (i)

$$\begin{array}{rcl} & \mathrm{Mg}^{2+} + \mathrm{O}^{2-} & \\ \mathrm{Mg} & \longrightarrow & \mathrm{Mg}^{2+} + 2\mathrm{e} \\ 1s^2 2s^2 2p^6 3s^2 & \longrightarrow & 1s^2 2s^2 2p^6 \quad \text{(Ne configuration)} \\ \mathrm{O} + 2\mathrm{e} & \longrightarrow & \mathrm{O}^{2-} \\ 1s^2 2s^2 2p^4 & \longrightarrow & 1s^2 2s^2 2p^6 \quad \text{(Ne configuration)} \end{array}$$

(ii)

$$Ca^{+2}(F^-)_2$$

$$Ca \longrightarrow Ca^{2+} + 2e$$

$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 \longrightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 \quad \text{(Ar configuration)}$$

$$F + e \longrightarrow F^-$$

$$1s^2 2s^2 2p^5 \longrightarrow 1s^2 2s^2 2p^6 \quad \text{(Ne configuration)}$$

S104 Book 4 Chapter 5 Section 5.5.

S111 Topic 5 Taking it further

Question 6

(a) (i) The electron configuration of the element is $1s^2 2s^2 2p^6 3s^2 3p^1$ (it is in fact aluminium, Al).

(ii) The element has three electrons in its outer shell ($n = 3$). It is in Group 13 of the Periodic Table. Its most common valency is 3.

(b) (i) The electron configuration of the element here is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ (it is in fact calcium,Ca).

(ii) The element has two electrons in its outer shell ($n = 4$), so it is in Group 2 of the Periodic Table. Its most common valency is 2.

S104 Book 4 Chapter 5 Section 5.2

S111 Topic 5 part 2 and Taking it further

Question 7

(a) $1s^2 2s^2 2p^6$

(b) $1s^2 2s^2 2p^6 3s^2 3p^6$

Note that these ions achieve the electron configuration of a noble gas by losing or gaining electrons.

S104 Book 4, Chapter 3 Section 3.4, Chapter 5 Section 5.2

S111 Topic 5 Taking it further

Question 8

The average atomic mass is given by 19.9% of 10 and 80.1% of 11.

$(19.9/100) \times 10 + (80.1/100) \times 11$

$= 1.99 + 8.811 = 10.8$ to one decimal place.

S104 Book 4 Chapter 3 Section 3.2.

S111 Topic 1 Part 1 Section 1.5.3, Book 9 Section 4.4.1.

Question 9

The molar mass of H_2SO_4 corresponds to the sum of the relative atomic masses of two hydrogen atoms, one sulfur atom and four oxygen atoms; expressed in grams this is $(1 \times 2) + 32 + (16 \times 4) = 98$ g. Therefore, 9.8 g corresponds to 0.1 mol. This amount is dissolved in 250 ml, which is 250/1 000 litre = 0.25 litres of solution. Since the concentration is the amount (in moles) dissolved in 1 litre of solution, the concentration of this solution is $4 \times 0.1 = 0.4$ mil litre^{-1}.

S104 Book 4 Chapter 4 Section 4.8.2.,

S111 Topic 4 Part 3 Section 3.7.1.

Question 10

$$C_7H_{16}(l) + 11O_2(g) = 7CO_2(g) + 8H_2O(g)$$

There are seven carbon atoms, sixteen hydrogen atoms and twenty-two oxygen atoms on each side of the equation. No charges are involved in this case.

S104 Book 4 Chapter 6 Section 6.2

S111 Topic 5 Part 2 Section 2.2.2

Question 11

0.05 mol of NaOH contain 0.05 mol of OH^- ions. If present in 0.5 dm^3 of solution, the concentration of OH^- ions will be equivalent to 0.1 mol in 1dm^3 of solution. In other words, the concentration is 0.1 mol dm^{-3}.

The ionic product of water, $K_w = [H^+]\,[OH^-] = 10^{-14}$ mol^2 dm^{-6}. Since in this instance $[OH^-] = 10^{-1}$ mol ldm^{-3}, $[H^+] = K_w/[OH^-] = 10^{-14}/10^{-1}$, or 10^{-13} mol dm^{-3}. The pH of the solution is therefore 13.

S104 Book 4 Chapter 11 Section 11.3

S111 Topic 5 Part 3 Section 3.4.2

Question 12

(a) (i) Sodium hydroxide dissolved in water is completely dissociated into aqueous sodium and hydroxide ions, $Na^+(aq)$ and $OH^-(aq)$, respectively. Hence, in a NaOH solution of concentration 1.0×10^{-2} mol dm^{-3}, $[OH^-(aq)] = 1.0 \times 10^{-2}$ mol dm^{-3}. At 25 °C, the corresponding hydrogen ion concentration will be $[H^+(aq)] = 1.0 \times 10^{-12}$ mol dm^{-3}. The solution will be basic because the hydrogen ion concentration is less than 1.0×10^{-7} mol dm^{31}.

(ii) Hydrochloric acid is a strong acid, which means that it is completely dissociated into its constituent ions, $H^+(aq)$ and Cl^- (aq). Thus, a 1.0×10^{-5} mol dm^{-3} solution of HCl will have $[H^+(aq)] = 1.0 \times 10^{-5}$ mol dm^{-3}. The solution is acidic because the hydrogen ion concentration is greater than 1.0×10^{-7} mol dm^{-3}.

(iii) NaCl dissolves to give only $Na^+(aq)$ and $Cl^-(aq)$ ions. Thus, the concentration of hydrogen ions will be the same as that in pure water. Hence $[H^+(aq)] = 1.0 \times 10^{-7}$ mol dm^{-3}: the solution will be neutral.

(iv) As ethanoic acid is a weak acid, it is only partially dissociated in water. Thus, a 1.0 mol dm^{-3} solution of CH_3COOH will have a hydrogen ion concentration greater than 1.0×10^{-7} mol dm^{-3} but less than 1.0 mol dm^{-3}. In fact at 25 °C, the hydrogen ion concentration is found to be 4.2×10^{-3} mol dm^{-3}: the solution is therefore acidic.

S104 Book 4 Chapter 11 Section 11.3

S111 Topic 5 Part 3 Section 3.4.2

Question 13

(i) Tin loses electrons and is oxidized from Sn^{2+} to Sn^{4+}; iron gains electrons and is reduced from Fe^{3+} to Fe^{2+}.

(ii) Magnesium loses electrons and is oxidized from Mg to Mg^{2+}; silver gains electrons and is reduced from Ag^{+} to Ag.

(iii) Zinc loses electrons and is oxidized from Zn to Zn^{2+}; hydrogen is reduced from H^{+} to H_2.

S104 Book 4 Chapter 8 Section 8.3.

S111 Topic 5 Part 2 Section 5 Section 2.6

Question 14

Frequency , *f*, is the number of cycles of a wave that pass a given point in one second. If five cycles pass a given point in two seconds then the frequency is 5/2 or 2.5 Hz.

S104 Book 3 Chapter 9.

S111 Topic 8 Part 1 Section 1.4.1

Question 15

A longer wavelength corresponds to a lower frequency so red light has a lower frequency than green light.

S104 Book 3 Chapter 8 Section 8.6, Book 7 Chapter 2 Section 2.3

S111 Topic 8 part 1 Section 1.4.1

Question 16

(a) The wavelength and frequency of electromagnetic radiation are related by $c = f\lambda$, where c is the speed of light, 3.00×10^{8} m s^{-1}. For a frequency, f, of 2.45 GHz, $\lambda = 3.00 \times 10^{8}$ ms^{-1}/2.45×10^{9} Hz. = 1.22 m.

(b) A wavelength of 8.8×10^{-12} m corresponds to a frequency $f = 3.00 \times 10^{8}$ m s^{-1}/8.8×10^{-12} m = 0.34×10^{20} s^{-1} = 3.4×10^{19} s^{-1}.

S104 Book 7 Chapter 2 Section 2.3.

S111 Topic 8 Part 4 4.1

Question 17

The transitions that we need to consider are shown in Figure 4.

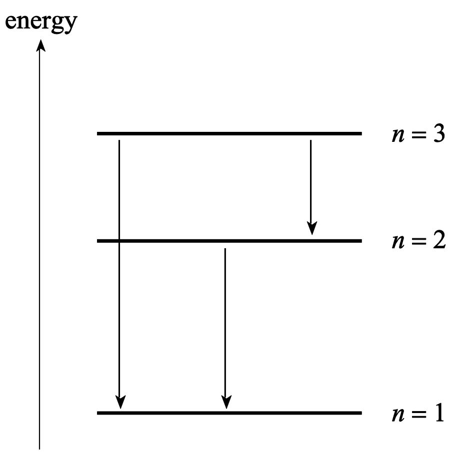


Figure 4 Changes in energy levels of an electron in a hydrogen atom.

The energy of the transition n = 3 to n = 2 is the difference between the energies of the transitions n = 3 to n = 1 and n = 2 to n = 1.

We know that E = hf

$$\Delta E = E_1 - E_2 = hf_1 - hf_2 = h(f_1 - f_2)$$
$$\therefore E = h(2.92 \times 10^{15} - 2.47 \times 10^{15})$$
$$= 6.626 \times 10^{-34}(0.45 \times 10^{15})$$
$$E = 2.98 \times 10^{-19}\ \mathrm{J}$$

S104 Book 7 Chapter 3 Section 3.5.

S111 Topic 8 Part 4 4.1, Topic 10 Taking it further

Question 18

The distribution is spherical. Hence it will appear the same from any angle.

S104 Book 7 Chapter 4 Section 4.2.

Question 19

(a) First you need to consider the energy required to break the bonds in CH_4 and $2Cl_2$. This is 4 × C–H bond energy and 2 × Cl–Cl bond energy. From the table this is

(4 × 413 + 2 × 242) = (1652 + 484) = 2136 kJ mol^{-1}

Now you need to consider the energy released when CCl_4 and $2H_2$ are formed. Since energy is released this is negative and given by −(4 × C–Cl bond energy) and −(2 × HH bond energy).

From the table this is
−(4 × 346) −(2 × 436) = −(1384 + 872) = −2256 kJ mol^{-1}

The enthalpy for the reaction is given by the sum of these two terms (2136 + (−2256)) = −120 kJ mol $^{-1}$.

(b) The enthalpy for the reaction is negative and so the reaction is exothermic.

S104 Book 4 Chapter 9 Sections 9.2 – 9.5

S111 Topic 5 part 2 Section 2.3.

Question 20

(a) See Figure 5.

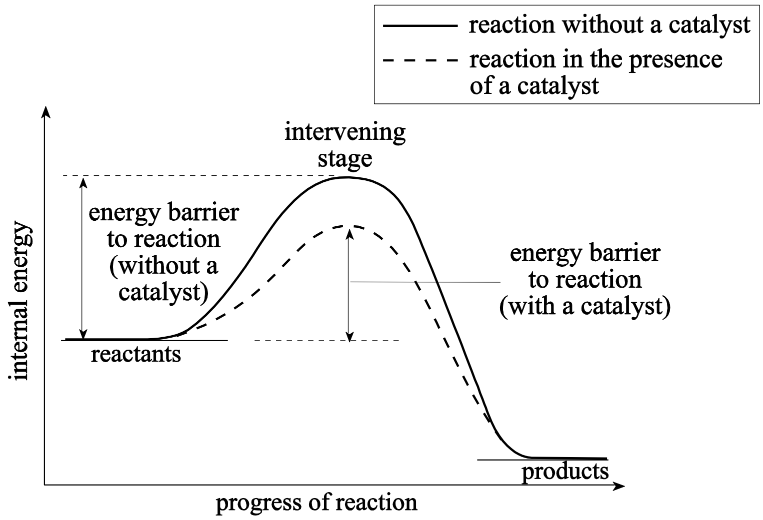


Figure 5 A fully labelled diagram showing the progress of a reaction with and without a catalyst.

(b) Since the products are at lower energy than the reactants, this shows an exothermic reaction.

S104 Book 4 Chapter 10 Section 10.4

Question 21

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad (8)$$

(i) In this case, the external constraint is an increase in pressure. Four moles of gas produce two moles of gas, so the pressure is *reduced* as the products are formed. The formation of product relieves the constraint of increased pressure. Increasing the pressure will therefore *increase* the equilibrium yield of ammonia.

(ii) Here, the external constraint is an increase in temperature. The forward reaction (as written) is said to be exothermic, so the reverse reaction will be endothermic. Temperature is *reduced* by an endothermic reaction, which *absorbs* heat; increasing the temperature promotes the reverse reaction to re-form the reactants. The *equilibrium* yield of ammonia will therefore *decrease* as the temperature is increased.

(iii) A catalyst will have no effect on the position of the equilibrium. It will, however, increase the rate of reaction in both the forward and reverse directions by providing an alternative, lower-energy route for reaction. The equilibrium yield will not be affected.

(iv) Removing the ammonia product as it is formed (for example, by condensation) will prevent equilibrium being reached. The reaction will continue to try to establish an equilibrium by producing more ammonia product (until all the reactants are used up). The equilibrium yield will therefore be unaffected; in fact, the equilibrium is never attained.

S104 Book 4 Chapter 10 Sections 10.5 – 10.8

Question 22

An increased equilibrium yield of ICl(g) will be produced by lowering the temperature of the reaction, as ΔH is negative. A change of pressure will have no influence on the equilibrium yield because there is no change in the number of gaseous molecules in this reaction, and hence no change in volume at equilibrium:

$I_2(g) + Cl_2(g) = 2ICl(g)$

1 vol + 1 vol = 2 vol

S104 Book 4 Chapter 10 Sections 10.5 – 10.8

Question 23

There are three possible isomers of C_5H_{12}. The straight-chain isomer is given in the question. The possible branched-chain isomers of C_5H_{12} are

$$\begin{array}{l}CH_3-\underset{\displaystyle CH_3}{\underset{|}{CH}}-CH_2-CH_3\end{array} \quad \text{and} \quad \begin{array}{l}CH_3-CH_2-\underset{\displaystyle CH_3}{\underset{|}{CH}}-CH_3\end{array}\text{, which are identical,}$$

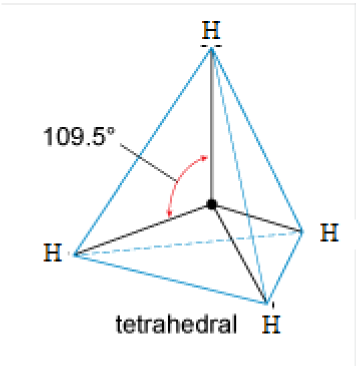
$$\text{and} \quad H_3C-\overset{\displaystyle CH_3}{\overset{|}{\underset{\displaystyle CH_3}{\underset{|}{C}}}}-CH_3$$

S104 Book 4 Chpater 12 Section 12.2

S111 Topic 9 1.1.4

Question 24

The four C-H bonds are arranged tetrahedrally, that is the H atoms lie on the vertices of a tetrahedron.



S104 Chapter 15 Section 15.2.1

S111 Topic 9 part 1 Section 1.1.2.

Question 25

Reaction 8 is an addition reaction; Reaction 9 is an oxidation reaction; Reaction 10 is a reduction reaction; Reaction 11 is a condensation reaction.

S104 Book 4 Chapter 14 Section 14.2

S111 Topic 5 part 3, Topic 9 Part 1 Section 1.2.1, Part 2 Section 2.2.1

6 Mathematical skills for S215

S215 assumes that you have the ability to:

- add, subtract, multiply, and divide
- use a scientific calculator
- manipulate decimal and fractions
- calculate values with the appropriate precision (that is, to the correct number of significant figures).

You will also be expected to be familiar with scientific notation (Section 4.1), handling data in appropriate units (Section 4.2), recording data graphically (Section 4.3), rearranging equations (Section 4.4), trigonometric functions (Section 4.5) and logarithmic and exponential functions (Section 4.6).

6.1 Trigonometric functions

These are functions such as **sine** and **cosine**. They can be defined in terms of a triangle as follows

For all triangles, the internal angles add up to 180°.

A triangle with one internal angle equal to 90°, i.e. a right angle, is known as a **right-angled triangle.**

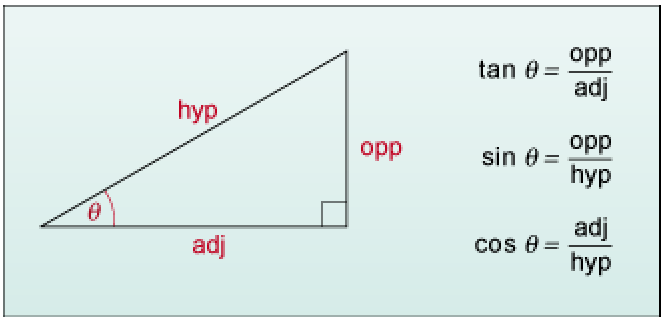


Figure 6 Trigonometric functions defined using a right-angled triangle

The side opposite the right angle in a right-angled triangle is known as

the hypotenuse,

For a right-angled triangle:

cos θ = adjacent /hypotenuse

sin θ = opposite /hypotenuse

tan θ = opposite/adjacent

The crystal structure of lithium iodide consists of lithium and iodide ions (ions are atoms with electric charge due to the loss or gain of electrons), as shown in Figure. Both types of ions can be represented by spheres and, in one model, the spheres can be considered just to touch each other. Trigonometry can then be used to find the radius of the ions.

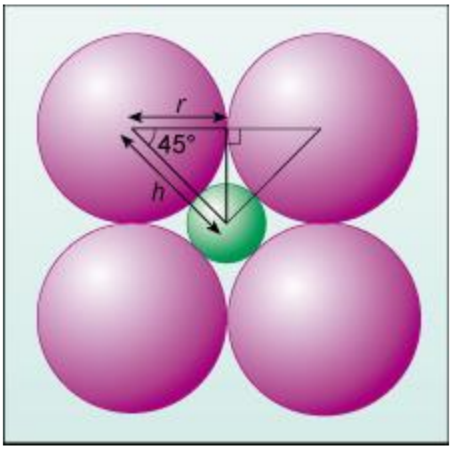


Figure 7 Using trigonometry to find the radius of iodide ions (shown in purple). The small green sphere represents a lithium ion.

If the distance between the centre of a lithium ion and the centre of an iodide ion is known (this is the so-called internuclear distance, and is labelled *h* on Figure 6) then

$$\cos 45^\circ = \frac{\text{adj}}{\text{hyp}} = \frac{r}{h}$$

where *r* is the radius of an iodide ion.

Multiplying both sides by *h* gives

$$r = h \cos 45° \qquad (13)$$

Equation 1 can be used to find the radius of an iodide ion.

Question M1

The internuclear distance, *h*, between the ions shown in Figure 7 is measured to be 302 pm (where 1 pm = 10^{-12} m). Use Equation 1 to find the radius of an iodide ion.

6.2 Fractions and percentages

Consider how you would calculate a certain fraction or percentage of a given number: for example, what is 75%, of 12? Start by thinking what of 12 means. First of all, 12 can be divided into four equal parts or quarters (12 ÷ 4 = 3). Then, since you want three-quarters, which is three times as big, you multiply one of these parts by three (3 × 3 = 9). So 75% of 12 is 9. This calculation can be written as

¾ × 12 = 3/4 × 12/1

12 = 3 × 4 so this equals

(3 × 3 × 4)/(4 × 1)

Cancelling the 4 from top and bottom this equals

3 × 3 =9.

In other words, 'of' a number means multiply by that number. If you see 'of' think 'multiply'.

Finding a percentage of a number can be done in a similar way. For example, 75% of 300 means:

$75/100 \times 300$

$= 75 \times 3$

$= 225$

So 75% of 300 is 225. Note the way in which the cancellation of the zeros has been indicated in this calculation, with each cancellation representing a division by ten.

Question M2

What is 8% of 400?

6.3 Scientific (powers of ten) notation

There is a very wide range of magnitudes of numbers involved in scientific data. For instance, the distance to the nearest star is 40 000 000 000 000 000 metres, whereas the time taken by light to travel 100 metres is 0.000 000 33 seconds. It is clearly not convenient to express very large or very small numbers like these in this conventional way.

A much more manageable form, known as **scientific notation**, uses the fact that large numbers are generated by multiplying several tens together, the number of tens being indicated by a superscript called ‘the **power**’. Thus

$10 = 10^1$ (read as ‘ten to the power one’)

$100 = 10 \times 10 = 10^2$ (‘ten to the power two’)

$1\,000\,000 = 10 \times 10 \times 10 \times 10 \times 10 \times 10 = 10^6$ (‘ten to the power six’).

The distance to the nearest star may now be expressed as

4×10^{16} m (where 16 is the number of zeros involved)

Similarly, the speed of light in a vacuum is approximately 300 000 000 m s^{-1}; in other words,

$3 \times 10 \times 10 \times 10 \times 10 \times 10 \times 10 \times 10 \times 10$ m s^{-1} = 3×10^8 m s^{-1}

Numbers less than 1 can be expressed in a similar way; for example,

$0.1 = 1/10 = 1/10^1$, which is written as 10^{-1}

and

$0.0001 = 1/10\,000 = 1/(10 \times 10 \times 10 \times 10) = 1/10^4$, which is written as 10^{-4}.

Thus, the time for light to travel 100 m can be written as 3.3×10^{-7} s.

Note that when *multiplying* values expressed using powers of 10, the powers must be *added*. For example,

$1\,000 \times 100$ becomes $10^3 \times 10^2 = 10^{(3+2)} = 10^5 = 100\,000$

In order to raise a power by a further power, the powers must be *multiplied*. For example,

$(10^3)^4 = 10^{(3 \times 4)} = 10^{12}$

Question M3

(a) Express the following numbers using scientific (powers of 10) notation:

(i) 8 970

(ii) 1 467 851

(iii) 0.010 01

(iv) 0.0046

(v) 234×10^{2}

(vi) 27×10^{-2}

(b) Evaluate the following using scientific notation:

(i) $10^{2} \times 10^{3}$

(ii) $10^{-2} \times 10^{4}$

(iii) $(10^{4})^{4}$

(iv) $(10^{4})^{1/4}$

6.4 Units

All measured quantities must have units associated with them. The units used in science are known as **SI units**. SI is an abbreviation of 'Système International d'Unités' (international system of units). The advantage of having a standard set of units is that everyone uses the same system, and there is therefore no need to convert laboriously from one system to another in order to compare results from different countries. In the SI system, all units are related to six base units: length is measured in metres (m); time in seconds (s); mass in kilograms (kg); amount in moles (mol), and temperature in kelvin (K, note, not 'degree kelvin', just kelvin), electric current in amperes (A), luminous intensity in candelas (cd)..

An example of the use of SI units is the recording of velocity. If 1 000 metres are travelled in 50 seconds, the velocity is given by:

$$\frac{1000\,\mathrm{m}}{50\,\mathrm{s}} = 20\,\mathrm{m\,s^{-1}}$$

Some units derived from these base units have special names. S215 assumes you have met the following:

Frequency	hertz	Hz (s^{-1})
Energy	joule	J (kg m^{2} s^{-2})
Force	newton	N (kg m s^{-2})
Power	watt	W (kg m^{2} s^{-3} ≡ J s^{-1})
Pressure	pascal	Pa (kg m^{-1} s^{-2} ≡ J m^{-3})
Electric charge	coulomb	C (A s)
Electric potential difference		V (kg m^{2} s^{-3} A^{-1} ≡ J A^{-1} s^{-1})

When carrying out a calculation involving physical quantities, the operation (for example, multiplication or division) is carried out on the units as well as the numbers. For example the value of K_w is determined by multiplying the value of the concentration of OH^{-} ions with that of the concentration of H^{+} ions. If concentration is expressed in units of mol per decimetre cubed, or mol dm^{-3}, the unit of K_w is (mol dm^{-3}) × (mol dm^{-3}) – in other words, (mol $dm^{-3})^2$ or mol^2 dm^{-6}.

A decimetre, dm, is 0.1 m.

6.5 Logarithms and the exponential function

You will find two types of logarithms on your calculator – denoted by **log** and **ln.**

log x is the power to which you have to raise 10 to obtain the number x. So log100 = 2 because 100 = 10^2. The powers do not have to be whole numbers. For example to three sig. figs. 2 = $10^{0.301}$ and therefore log 2 = 0.301.

An important use of log x is in determining pH values. pH tells you how acid a solution is and is defined by pH = – log [H^+] where [H^+] represents the concentration of hydrogen ions in a solution.

It is worth noting that

- It is not possible to obtain the logarithm to base 10 of a negative number, or of zero: if you try this on your calculator it will produce an error message.
- It is possible to obtain logarithms of *pure numbers* only; you cannot obtain the logarithm of a quantity possessing units. Strictly, if a quantity possesses units, then it should be divided by those units before taking the logarithm

ln x is the power to which you have to raise the number e to obtain the number x. Like π and $\sqrt{2}$, e is an irrational number and to four significant figures its value is 2.718

Check that you can use your calculator to raise e to various powers. You are likely to be using a button labelled ‘e^x’ in order to do this; the ‘EXP’ button has a totally different use. There is a need to take particular care over the meaning of ‘e’, ‘EXP’ and ‘exp’ since ‘exp’ is sometimes used to mean ‘e to the power’, so $N = N_0 e^{-\lambda t}$ is sometimes written as $N = N_0 \exp(-\lambda t)$ and $n = n_0 e^{at}$ is sometimes written as $n = n_0 \exp(at)$.

Question M4

(a) What is the value of $e^{0.5}$ to 2 significant figures?

(b) What is the value of ln(0.5) to 2 significant figures?

Question M5

What is the pH of a solution in which the hydrogen ion concentration is 10^{-3} mol dm^{-3}?

Question M6

What is the pH of a solution in which the hydrogen ion concentration is 2×10^{-3} mol dm^{-3}?

6.6 Rearranging equations

The following equation

$$\upsilon_s = \sqrt{\frac{\mu}{p}} \qquad (14)$$

enables you to calculate the S wave speed, υ_s, of seismic waves passing through a rock of density ρ and rigidity modulus μ. But suppose that, instead of knowing ρ and μ and wanting to find υ_s, you know υ_s and ρ and want to find μ. The best way to proceed is to rearrange the equation to make μ the **subject** of the equation, where the word 'subject' is used to mean the term written by itself, usually to the left of the equals sign. That is an equation that starts

$$\mu =$$

There are many different methods taught for rearranging equations, and if you are happy with a method you have learnt previously it is probably best to stick with this method. (However, note that some methods taught for rearranging equations, in particular the so-called 'triangle' method, are not generally applicable and cannot be used to rearrange all the equations you will meet.).

Returning to equation11, to arrange the equation so that μ is the subject, you could proceed as follows.

$$\upsilon_s = \sqrt{\frac{\mu}{p}}$$

You can consider there to be brackets around $\left(\frac{\mu}{p}\right)$ and start by finding an expression for $\left(\frac{\mu}{p}\right)$

The equation can be written as

$$\sqrt{\frac{\mu}{p}} = \upsilon_s$$

which has μ on the left-hand side

Squaring both sides gives

$$\frac{\mu}{p} = \upsilon_s{}^2$$

Now you can multiply both sides by ρ to give

$$\mu = \upsilon_s{}^2 p$$

Question M7

The density of a cubic crystal, ρ, is given by $\rho = m/V$ where m is the mass of atoms or ions in one cell of the crystal of volume V. For a cubic cell of side, r, write the volume in terms of r, and rearrange the equation $\rho = m/V$ to make r the subject.

6.7 Graphical applications

The significance of trends in data is often seen more clearly when those data are presented in the form of a **graph**. For example, a plot of distance travelled against time (as in Figure 8) allows us to calculate **velocity**.

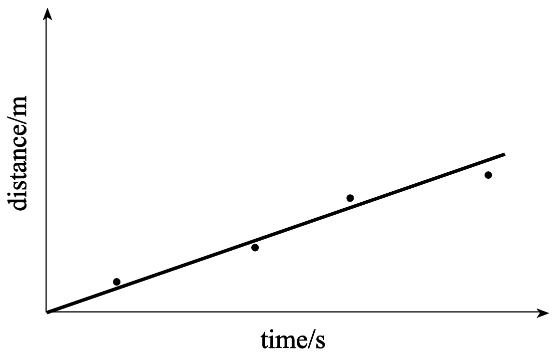


Figure 8 Distance (on the vertical, y, axis) plotted against time (on the horizontal, x, axis).

A linear plot (with y plotted against x) is represented by an equation of the form:

$$y = mx + c \qquad (15)$$

where m, the **proportionality constant**, corresponds to the slope (**gradient**) of the straight line, and c corresponds to the value of the intercept of the line with the y axis; in this example $c = 0$, because the line cuts the vertical axis (and the horizontal axis in this particular case) at the value $y = 0$. The slope (gradient), m is given by $(y_1 - y_2)/(x_1 - x_2)$ where (x_1, y_1) and (x_2, y_2) are the coordinates of any two points on the line.

The unit of the physical quantity corresponding to the slope is obtained as follows (using the example in Figure 8):

$$\frac{\text{distance/m}}{\text{tims/s}} = \text{velocity/ms}^{-1}$$

The fact that it is a linear plot (that is, it's a straight line, or one with a constant slope) means that the velocity is constant.

Not all graphs are straight lines, as Figure 9 shows. The plots in Figure 9 illustrate the progress of the reaction

$$S_2O_8^{2-}(aq) + 2I^-(aq) = 2SO_4^{2-}(aq) + I_2(aq) \qquad (16)$$

for a *fixed* initial concentration of iodide ions and various initial concentrations of peroxodisulfate ions ($S_2O_8^{2-}$). The initial concentration of $S_2O_8^{2-}$ increases as you go from curve (a) to curve (d).

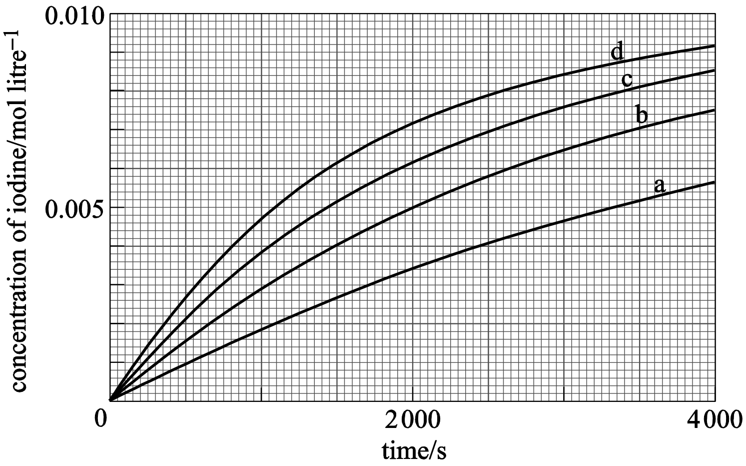


Figure 9 The progress of Reaction 10 as measured by the increase in the concentration of iodine at 25 °C for four different initial concentrations of peroxodisulfate ions, a, b, c and d.

In Figure 9, the early part of each reaction-progress curve approximates to a straight line. For example, curve (a) from 0 to 1 500 s, and curve (d) from 0 to 400 s. For all four lines the value of *c* is 0 because they all cut the *y* axis (showing the concentration of iodine) at the point where the value of *y* is zero.

Experimental data will not normally lie exactly in a straight line, and a 'best fit' line should be drawn, in which the data points are scattered evenly above and below the line. The gradient of such a line can be calculated by selecting two points on the line and dividing the difference in their corresponding values of *y* by the difference in their values of *x*.

Question M8

In an experiment to investigate the relationship between the mass hung from a spring and the extent of stretching of the spring, the following readings were obtained.

Mass added/kg, *x*	1	2	3	4	5
Length of spring/cm, *y*	22.5	26.6	30.6	34.7	38.6

Use graphical means to derive the equation relating mass added (*x*) to spring length (*y*). What is the unextended length of the spring (*l*)?

Question M9

With reference to Figure 9, does the speed (rate) of Reaction 11 increase or decrease as the concentration of peroxodisulfate ions increases?

Question M10

What is the slope of curve (a) in the range of iodine concentration from 0 to 0.002 mol litre^{-1}? Express your value in scientific notation using the appropriate units.

What is the initial rate of Reaction 11

7 Answers to maths questions

Question M1

$r = h \cos 45°$

$h = 302$ pm, $\cos 45° = 0.707$

$r = 302 \times 0.707$ pm = 214 pm to 3 sig. figs

Maths for Science Chapter 9 Section 9.3.

S111 Topic 7 Maths

Question M2

8% of 400 is $(8/100) \times 400$

$= 8 \times 4$

=32

Maths for Science Chapter 1 Section 1.2.

S111 Maths Skills

Question M3

(a) (i) 8.97×10^3

(ii) $1.467\ 851 \times 10^6$

(iii) 1.001×10^{-2}

(iv) 4.6×10^{-3}

(v) 2.34×10^4

(vi) 2.7×10^{-1}

(b) (i) 10^5

(ii) 10^2

(iii) 10^{16}

(iv) 10^1

Maths for Science Chapter 2 Section 2.1

S111 Maths Skills Scientific notation.

Question M4

(a) $e^{0.5} =$ 1.6487212707001281468486507878142 = 1.6 to 2 significant figures.

(b) ln(0.5) = −0.69314718055994530941723212145818 = −0.69 to 2 significant figures.

Maths for Science Chapter 10

Question M5

pH = – log [H^+]. [H^+] = 10^{-3}mol dm^{-3}.

log [H^+] = – 3

pH = 3.

Maths for Science Chapter 10

Question M6

pH = – log [H^+]. [H^+] =2 × 10^{-3}mol dm^{-3}.

log [H^+] = – 3 + 0.301 = –2.699

pH = 2.699.

Maths for Science Chapter 10

Question M7

The volume of a cube of side r is r^3. Hence

$\rho = M/r^3$.

Multiplying both sides by r^3

$\rho \times r^3 = M$

Dividing each side by ρ

$r^3 = M/\rho$

Now r is on the left hand side but it is cubed so you need to take the cube root.

$r = \sqrt[3]{\frac{M}{\rho}}$.

Maths for Science Chapter 6 Section 6.2.

S111 Topic 8 Part 2 Section 2.1.4

Question M8

Plot a graph of the extension of the spring, y, against added mass, x (see Figure 10) This gives a straight-line plot of the form $y = mx + c$, where the slope $m = 16/4 = 4$ cm kg^{-1}. The intercept, $c = 18.6$ cm, represents the length of the unextended spring, with no added masses ($x = 0$).

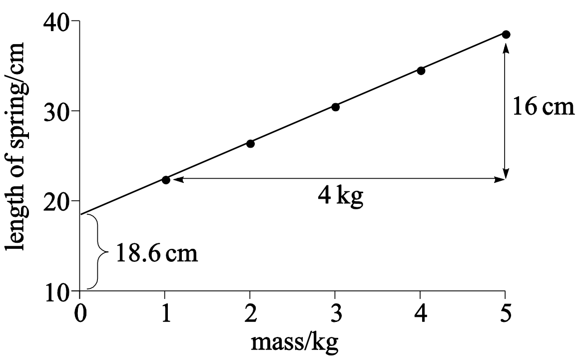


Figure 10 Graph showing the extension of a spring against the mass added to it.

Maths for Science Chapter 7 Section 7.2

S111 Topic 2 Part 3 Section 3.2.

Question M9

The rate of reaction can be defined as the rate at which the concentration of a product increases, or the concentration of a reactant decreases. In Figure5 we can measure the rate of reaction by measuring the rate at which the concentration of one of the products, iodine, increases. As the initial concentration of peroxodisulfate ions increases, the curves are steeper at short times; this means that the rate of the reaction increases.

Maths for Science Chapter 8 Section 8.1

S111 Topic 2 Part 3 Section 3.2.

Question M10

The slope of the linear portion of the curve is

$$0.002 \text{ mol litre}^{-1}/1\,050 \text{ s} \approx 0.000\,002 \text{ mol litre}^{-1} \text{ s}^{-1}$$

$$\approx 2 \times 10^{-6} \text{ mol litre}^{-1} \text{ s}^{-1}$$

The initial rate of reaction is given by

$$\frac{\text{increase in iodine concentration}}{\text{time elapsed}}$$

In other words, the slope of the linear portion of the curve, which, as we have seen, is 2×10^{-6} mol litre^{-1} s^{-1}.

Maths for Science Chapter 7 Section 7.2

S111 Topic 7 Part 4 Section 4.3.1.