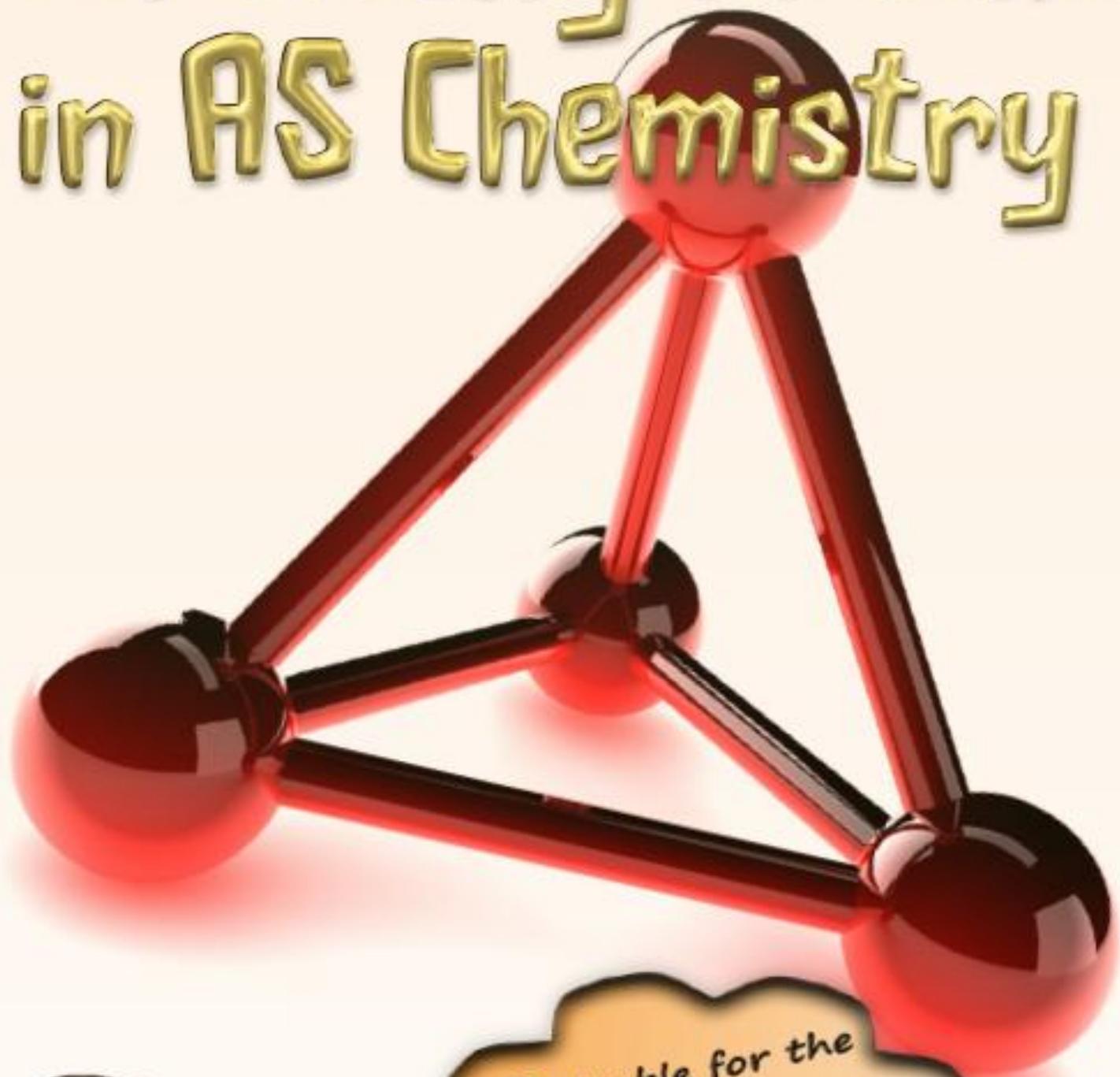


Surviving Maths in AS Chemistry



Suitable for the
new specs from
AQA, OCR, WJEC
and Edexcel!

Dr. Robert Mitchell

CT Publications

Surviving Maths in AS Chemistry

by

Dr Robert Mitchell



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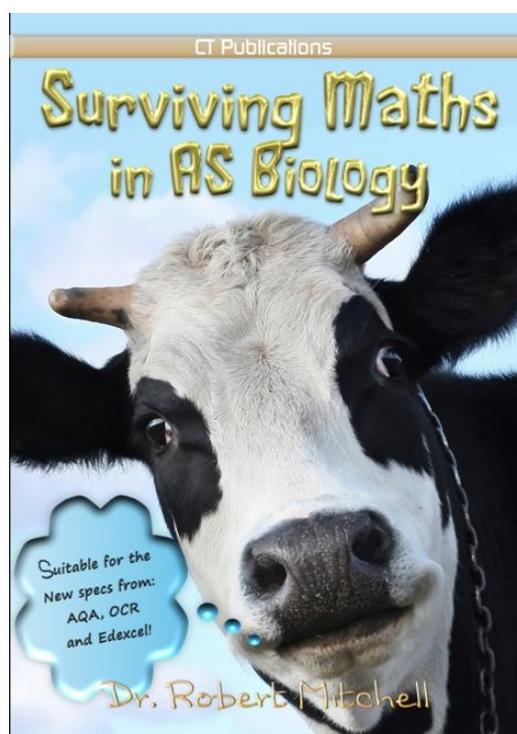
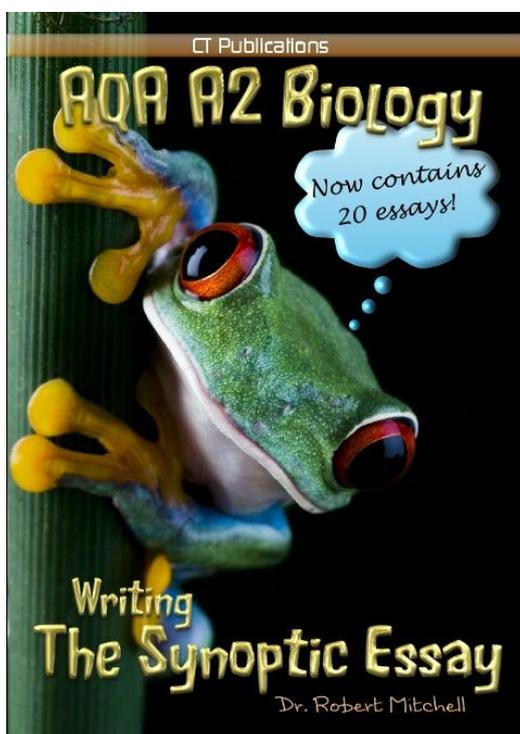
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About the author

Rob is a private tutor in chemistry and biology in Bolton. He's formerly worked in medical research as technician, research assistant and post-doctoral researcher and has contributed to the publication of over 40 research papers. During a varied career in science, he's been a project leader in industry, a lecturer and examiner and blogs daily as *Chemicalguy*. He likes dogs, and pies, going to the movies and walking!

Other books by the author

AQA A2 Biology; Writing the Synoptic Essay	May 2010
Surviving Maths in AS Biology	September 2010
Ultimate Exam Preparation; AQA Chemistry Unit 1	October 2010 (in press)
Ultimate Exam Preparation; AQA Biology Unit 1	November 2010 (in press)
Uprooting the Tree of Life (Popular Science)	December 2010 (in press)
Biofuelishness (Popular Science)	March 2011 (in preparation)



Preface

Love it or hate it, you can't escape from doing maths in chemistry A-level. Since the introduction of the new-style specifications in 2008, the exam papers have included up to 20% of calculation and mathematical transformations. In the ISA or EMPA parts of the specifications this can increase to almost 50%. This means that those of you wanting to secure the grade A, or A* qualifications will be unlikely to achieve it without mastering the mathematical principles.

All exam boards publish the same set of *Mathematical Requirements*. While these are a basic set of criteria, I have found over the years that students often struggle with these concepts to the point that they can impact severely on the outcome in their exams. Part of the reason for this is that students often do not link and carry forwards some of the material from the GCSE. Even when they do, being able to calculate proportions in a GCSE maths class, for example is not necessarily an indicator of them being able to apply proportional changes in a titration calculation in A-level chemistry.

This book aims to put this right! It is split into four main sections. *Section 1* covers the basic mathematical requirements outlined in the specifications using examples from the AS chemistry syllabus. *Section 2* then systematically covers all the AS mathematical content of the chemistry courses giving you simple and robust techniques for getting these calculations right every time! *Section 3* will then give you many examples of exam-style questions using the styling from AQA, Edexcel, OCR and WJEC. These of course come with mark schemes and a breakdown of the points so you can see how they are awarded in *Section 4*.

As with all the books from CTPublications the emphasis is on showing you how to do the content and get the exam points rather than helping you understand why you are doing it. I wish you the very best of luck to you all in your exams and future careers.

Dr Robert Mitchell
Summer 2010

Student resources

I publish two regular blogs covering various aspects of studying A-level chemistry and biology. All updates on new products and services are posted on these blogs **before** any other announcements. They are found at:

www.chemicalguy.wordpress.com [chemistry]

www.howscienceworks.wordpress.com [biology]

Pop along to this book's dedicated website for some more exam questions and worked examples.

www.SurvivingMathsInASChemistry.co.uk

I also recommend using www.thestudentroom.co.uk for free help and support.

How to use this book



Also consider: 1. Lose the book 2. Buy it again 3. Write an excellent review on Amazon.co.uk or CTPublications.co.uk, and 4. Recommend the book to all your classmates, teachers, head of science and college library book-buying officer. ☺

So many books show you *what you need to know* but miss out the obvious! How to do it! You should assume you know nothing and read through **the entire** book. You'll find an *End of Section Test* at ... errm the end of each section which you should complete before attempting the exam questions. Always, always, always monitor your performance and be critical of your own answers when marking your efforts. If you always work on fixing the weaker areas you will gain the most improvement in the least time!

You will also notice the 🍷 placed strategically throughout. This symbol infers that those points are common mistakes you **must** be aware of and avoid.

The terms *sig fig* and *dp* refer to the number of significant figures and decimal places respectively that a number is rounded to.

Section 1: Mathematical Requirements

All exam boards publish the same set of mathematical criteria. These are a series of statements that identify what you should be able to understand, or do, on entry into the AS level. Most of it might be the stuff you'd prefer to forget, or prefer not to remember that you never knew how to do it in the first place! Work through the following material even if you don't like it, you'll be glad you did later.

Reference	Relevant Topics
Arithmetical and numerical computation	
<ul style="list-style-type: none"> Recognise and use expressions in decimal and standard form. Use ratios, fractions and percentages. Make estimates of the results of calculations (without using calculators). Use calculators to find and use power, exponential and logarithmic functions. 	<p>All quantitative calculations involving moles, masses, volumes and concentrations (AS & A2). Also applicable to rate, equilibrium, pH, <i>Ecell</i> calculations in A2. Estimation of answers will be exploited in relative atomic mass calculations (AS). Logarithms used in the context of ionisation energy (AS) and pH calculations (A2).</p>
Handling data	
<ul style="list-style-type: none"> Use an appropriate number of significant figures. Find arithmetic means. 	<p>A core component of all quantitative work at AS and A2 and is particularly relevant to the practical aspects of ISA or EMPA.</p>
Algebra	
<ul style="list-style-type: none"> Understand and use the symbols =, <, <<, >, >>, α, \sim. Change the subject of an equation. Substitute numerical values into algebraic equations using appropriate units for physical quantities. Solve simple algebraic equations. Use logarithms in relation to quantities which range over several orders of magnitude. 	<p>Symbols used appropriately in calculations or comparisons. Rearranging equations, substituting in values, solving and changing units are all core skills that will be tested in the context of all quantitative calculations involving moles, titrations, rates, equilibrium etc.</p> <p>Logarithms will be mainly used in the context of <i>pH</i>, <i>pKa</i>, <i>pKw</i> in A2.</p>
Graphs	
<ul style="list-style-type: none"> Translate information between graphical, numerical and algebraic forms. Plot two variables from experimental or other data. Understand that $y = mx + c$ represents a linear relationship. Determine the slope and intercept of a linear graph. Calculate the rate of change from a graph showing a linear relationship. Draw and use the slope of a tangent to a curve as a measure of rate of change. 	<p>The graphical component will mainly be relevant for the practical work, particularly in the context of experiments involving energy changes (AS), estimations of reaction rates (AS) and acid-base (AS & A2) or redox titrations (A2).</p>
Geometry and trigonometry	
<ul style="list-style-type: none"> Appreciate angles and shapes in regular 2D and 3D structures. Visualise and represent 2D and 3D forms including two-dimensional representations of three-dimensional objects. Understand the symmetry of 2D and 3D shapes. 	<p>Usually assessed in the context of molecular shapes and estimation of bond angles (AS).</p> <p>The representation of displayed formulae of covalent molecules (AS & A2).</p> <p>The impact of symmetry on physical properties.</p>

Arithmetical and Numerical Computation

Selecting the right calculator

Given the importance of the mathematical component of AS chemistry to your overall A-level success, it is imperative that you buy the right calculator early on and become *very* familiar with how to use it. You cannot just assume that the one you used at GCSE will make do. The range and complexity of the functions your calculator will need to have increases exponentially (pun very much intended!) in A-level. Here are just a few other things to consider before you part with your money on the shiny new abacus:

- **Always** carry a spare calculator battery with you, particularly around exam time ... *Sod's Law* states that it will go just as the exam is about to start!
- **Never** throw away the instruction leaflet ... at this point you don't know for sure what you will be using it for.
- **Before** you buy the calculator, **check** the requirements for the other subjects you do, particularly if you do mathematics, physics or biology ... a multitude of other functions such as statistical analysis or graphical functions may be required.

When selecting an appropriate calculator, ensure that it:

- Is described as a *scientific calculator* and can calculate numbers in the range at least 1×10^{-14} to 1×10^{24}
- Is able to do the following *logarithmic* functions; \log \ln
- Can express numbers in *standard form* using either $\times 10^x$ or \exp
- Can easily do squares, powers and square roots using x^2 x^n \sqrt{n}

Fractions, ratios, percentages and Stuff

The same numerical value can be expressed in different ways. For example, the decimal number 0.005 is the same as the fraction $\frac{1}{200}$ and can be expressed as 5×10^{-3} in standard form. It can also be expressed as a ratio of *one in two hundred* or as 0.5%. In science, we use these different expressions of the same numbers in different contexts. How some of these different forms of a number are used is outlined below.

Fractions

- If the value is less than *one*, a fraction $\frac{x}{y}$ can be used to express it.
- While useful, fractions are not easy to compare. If I asked you which was smaller, $\frac{11}{17}$ or $\frac{12}{18}$ it is not easy to give a definitive answer.
- In such cases the number at the bottom, the *denominator*, can be made the same. When it is, it is called a *common denominator*.

Decimals

- The transformation of fractions into decimals leads to a result that can be compared instantly. In such cases, the common denominator effectively becomes 1.
- On a calculator this is done by dividing the top number of the fraction by the bottom number. In the above example, $\frac{11}{17}$ becomes 0.647 and $\frac{12}{18}$ becomes 0.667.
- This simple transformation now shows that $\frac{11}{17}$ is smaller than $\frac{12}{18}$.

Standard form

- Decimals have limited use when a number becomes very large or very small. In such cases, standard form is used to provide a consistent way of presenting and handling the number.
- Numbers in standard form are usually seen as $x.yz \times 10^n$ where x lies in the range between 1 and 9 and n can be a *negative* or *positive* number.
- The fraction $\frac{1}{120}$ has a decimal value of 0.00833. In standard form this becomes 8.33×10^{-3} (3 sig fig). The large number 19878956 would become 1.99×10^7 (3 sig fig).
- This powerful form of scientific numbering is used throughout chemistry, especially when small amounts of materials are used.
- To put the number into standard form follow the simple steps below:

		Example 1	Example 2
(i)	Write out the number.	0.00833	19878956
(ii)	Place a decimal point between the first two non-zero parts of the number.	0.008•33	1•9878956
(iii)	Move toward the original decimal point noting the direction you move and <i>number</i> of jumps.	3 places left	7 places to the right
(iv)	Construct the number – if moving left, the power is <i>negative</i> , if moving right the power is <i>positive</i> .	8.33×10^{-3}	1.9878956×10^7
(v)	Round to three significant figures	8.33×10^{-3}	1.99×10^7

Percentages

- Percent means *out of 100*, and fractions or decimals are converted to percentages by simply multiplying them by 100.
- It's effective as our brains most easily use numbers between 1 and 100.
- Conversion of fractions or decimals into percentages allows an instant *comparison* which carries *meaning*.
- For example, would you choose to smoke if I said that 90% of all smokers died of lung cancer before age 60? What if I said that 1% died instead? Your brain can easily process and use these figures to make reasoned judgements.

Ratios

- Fractions can also be considered to be ratios.
- A ratio maintains a constant relationship between the top and bottom numbers of the fraction.
- A fraction of $\frac{1}{10}$ can translate as *one in ten*. So if *one in ten* students get a grade A, then by scaling the numbers by equal amounts on the top and bottom can give an appropriate expectation of other combinations of numbers. For example of $\frac{1}{10}$ is the same ratio as $\frac{2}{20}$ or $\frac{8}{80}$. So I could reasonably expect 6 grade A results out of a group of 60 students.

Powers and exponents

- When a number is raised to a power, for example 2^4 it means that the 2 is multiplied by *itself*, four times ($2 \times 2 \times 2 \times 2 = 16$).
- On a calculator this is achieved using the x^n or equivalent button.
- This calculation can be done by pressing $2 \ x^n \ 4 \ =$
- For standard form, scientific calculators have a function which multiplies the value by the 10^x . The $x10^x$ or exp buttons will automatically put that part of the standard form in place. So pressing $1.99 \ x10^x \ 7$ will enter 1.99×10^7 into the calculator.

Logarithms

- Taking a logarithm, or *log* of a number, is a mathematical transformation that scientists use to compress data that are spread over a wide range, again to make the numbers more manageable and comparable.
- On a calculator simply type \log followed by the number you wish to log.
- So the log of 2.99 is found by typing $\log \ 2.99 \ =$ which gives 0.476.
- A logarithm can be converted back to the original number by using the *antilog* function, or $\text{shift} \ \log \ 0.476 \ =$ which gives 2.99.

Handling Data

Rounding

- When a decimal number is rounded, it loses some of its precision and can therefore introduce some inaccuracies in calculations if not handled appropriately.
- To round the number *up* or *down* follow the simple steps below:

		1 dp	2dp	3 dp
(i)	Write out the original number and decide on how many decimal places you need. e.g. 2.849 to 2dp.	2.8494	2.8494	2.8494
(ii)	Decide whether the figures after number of decimal places you want is less than or greater than 5xxx	2.8 494 Where 494 is <i>less than</i> 500	2.8 494 Where 94 is <i>greater than</i> 50	2.849 4 Where 4 is <i>less than</i> 5
(iii)	If it is <i>less than</i> ; then remove the extra decimal places,	2.8		2.849
(iv)	If it is <i>greater than</i> ; then increase the last digit remaining by 1		2.85	

- In practice rounding is a lot easier to *do* than it is to *describe*!
- It is advisable only to round your answer *up* or *down* to the appropriate number of decimal places **at the end** of a calculation to minimise the risk of errors creeping in along the way.
- In general, either **one** or **two decimal places** are usually specified in *relative atomic mass* calculations. **Two decimal places** are used for *pH calculations* in A2 whereas most other calculations (in particular those involving moles and masses etc.), use **three significant figures** as the appropriate level of precision (see below).

Significant figures

- In chemistry the use of significant figures tends to be preferable to decimal places. This is because different numbers have different *magnitudes*, or sizes. A good example of this is money. If you had £1,000,000 pounds in your bank account, it would barely seem relevant whether it was £1,000,000.01 or £1,000,000.99.

Like rounding, showing you how a number can be expressed to a certain number of significant figures is easier to *do* than to describe. There are a few rules to assigning significant figures:

- (i) All non-zero numbers are significant (e.g. 9700 has 2 significant numbers, 9 and 7).
- (ii) Zeros appearing between two other digits are significant (e.g. 101 has three significant figures, 1, 0 and 1).
- (iii) Leading zeroes are not significant (e.g. 0.0052 has 2 significant figures 5 and 2).
- (iv) Trailing zeros of a number before a decimal point are significant (e.g. 91200.1 has 6 significant figures 9, 1, 2, 0, 0 and 1).
- (v) If appropriate, a number can be rounded up or down (e.g. 512 to two significant figures is 510 whereas 587 to two significant figures is 590).

Examples:

907.459	0.000528715	have 6 sig fig
907.46	0.00052872	have 5 sig fig
907.5	0.0005287	have 4 sig fig
907	0.000529	have 3 sig fig
910	0.00053	have 2 sig fig
900	0.0005	have 1 sig fig

Arithmetic means

A mean is one of three kinds of mathematical average (mode and median are others). The mean is found using the equation:

$$\text{Mean} = \frac{\Sigma \text{ values being averaged}}{\text{number of values to be averaged}}$$

Most people use the principle, “*add them all up and divide by how many there are.*” For example to find the average of 35, 38, 42 and 56 the equation:

$$\text{Mean} = \frac{35 + 38 + 42 + 56}{4} = 42.8 \text{ (3 sig fig)}$$

One common error is to **fail to check** that the average value calculated lies **within** the range of the numbers being averaged. So in the example above 42.8 lies between 35 and 56.

Algebra

Symbols

You are required to know the meaning and context of the following symbols. A *translation* is included to show how the mathematics can be interpreted. These symbols may appear in algebraic expressions or be used in text or summaries in chemistry. Some examples are given below.

Symbol	Translation	Example of context
=	Is equal to	The pH of 0.1 mol dm^{-3} HCl = the pH of 0.1 mol dm^{-3} HBr
<	Is less (or lower than) than	The oxidising strength of iodine < chlorine
<<	Is much less than	The acidic strength of ethanoic acid << hydrochloric acid
>	Is greater than	The reducing strength of iodide ions > chloride ions
>>	Is much greater than	The atomic radius of francium >> hydrogen, or $A_r(\text{Fr}) \gg A_r(\text{H})$
α	Is proportional to	The rate of a first order reaction α the concentration, or rate α [A]
\sim	Is approximately equal to	The electronegativity of carbon \sim hydrogen
Δ	Change in	Standard enthalpy change of formation of Cu(s), $\Delta H_f(\text{Cu})$
Σ	Sum of	Sum of the enthalpy changes of formation of the products, $\Sigma \Delta H_f(\text{products})$
%	Percent (or <i>out of 100</i>)	10% of 5 is 0.5, is the same as writing $\frac{10}{100} \times 5 = 0.5$

Rearranging equations

Incorrect equation rearrangements, or changing the subject of an equation as it is known, accounts for a large percentage of mathematical errors in chemistry exams. It's a skill that many science teachers expect their students to have already and so they do not show them simple ways of doing it.

There are several rearrangement methods, all of which result in the same

rearrangements if done properly. Two have been included here for you to choose from. I suggest you practice both and use the one that gives to the correct answer most of the time.

For Equations involving multiplication and division:

Method 1: Rearrangement

In chemistry the rearrangement of equations such as $moles = \frac{mass}{molar\ mass}$ is a common expectation in exams. You can rearrange these types quickly by applying the following steps.

For example; **to make mass the subject of the equation:**

- (i) Focus on the *subject*, the variable you want to find, e.g. *mass*.
- (ii) Whatever another variable is doing **to** the *subject*, then **do the opposite** to it (divide if it's multiplying, multiply if it's dividing), on **both sides** of the equation. This means:

$$molar\ mass \times moles = molar\ mass \times \frac{mass}{molar\ mass}$$

- (iii) As anything divided by itself = 1, then the above cancels down to:

$$molar\ mass \times moles = \cancel{molar\ mass} \times \frac{mass}{\cancel{molar\ mass}}$$

$$molar\ mass \times moles = mass$$

For example; **to make molar mass the subject of the equation:**

- (i) If the *variable* you want to make the *subject* of the equation is currently underneath another one, then **first of all** multiply both sides by it. This means:

$$molar\ mass \times moles = \frac{molar\ mass \times mass}{molar\ mass}$$

(ii) Cancel this down to get a linear equation:

$$\mathbf{molar\ mass} \times \mathbf{moles} = \frac{\mathbf{molar\ mass} \times \mathbf{mass}}{\mathbf{molar\ mass}}$$

$$\mathbf{molar\ mass} \times \mathbf{moles} = \mathbf{mass}$$

(iii) Repeat step (ii) but this time divide both sides by moles, so:

$$\frac{\mathbf{molar\ mass} \times \mathbf{moles}}{\mathbf{moles}} = \frac{\mathbf{mass}}{\mathbf{moles}}$$

(iv) Cancelling this down gives the rearranged equation:

$$\frac{\mathbf{molar\ mass} \times \cancel{\mathbf{moles}}}{\cancel{\mathbf{moles}}} = \frac{\mathbf{mass}}{\mathbf{moles}}$$

$$\mathbf{molar\ mass} = \frac{\mathbf{mass}}{\mathbf{moles}}$$

Method 2: Linearised method

This method is simpler, but it requires you to **learn** the *linearised* versions of the equations first. All you do is **learn a** linear version and then just *put everything except the subject under the other side of the equation*. It's that simple! Using the above example, a linear version of the equation would be:

$$\mathbf{mass} = \mathbf{moles} \times \mathbf{molar\ mass}$$

Common *linearised* equations you will need for AS chemistry include:

- ✓ *Mass = moles x molar mass*
- ✓ *Moles (x 1000) = concentration x volume*
- ✓ *Gas volume = moles x 24(000)*
- ✓ *Pressure x volume = moles x gas constant x temperature*
- ✓ *Energy change = mass of solution heated x specific heat capacity x temperature change*
- ✓ *Energy change = enthalpy change x moles*

For equations involving addition and subtraction:

The principles involved in rearranging equations with addition or subtraction is the same as for multiplication in *Method 1* above. If the variable you want to make the subject is *positive* then do the *opposite* function (addition or subtraction) for *all* other variables, *on both sides*. For example, to make ΔH_1 the subject from the equation $\Delta H_r = \Delta H_1 + \Delta H_2$:

- (i) $\Delta H_r - \Delta H_2 = \Delta H_1 + \Delta H_2 - \Delta H_2$
- (ii) Which simplifies to $\Delta H_r - \Delta H_2 = \Delta H_1$
- (iii) Repeat the process with any other variables present

If the variable you want to make the subject is *negative* then start by adding that variable to both sides and then repeating the process outlined above. For example, to make ΔH_1 the subject from the equation $\Delta H_r = \Delta H_2 - \Delta H_1$:

- (i) $\Delta H_r + \Delta H_1 = \Delta H_2 - \Delta H_1 + \Delta H_1$
- (ii) Which simplifies to $\Delta H_r + \Delta H_1 = \Delta H_2$
- (iii) $\Delta H_r - \Delta H_r + \Delta H_1 = \Delta H_2 - \Delta H_r$
- (iv) Hence, $\Delta H_1 = \Delta H_2 - \Delta H_r$
- (v) Repeat the process with any other variables present

You are most likely to meet these rearrangements under the *enthalpy changes* section of your specification.

- ✓ $\Delta H_{\text{reaction}} = \sum \Delta H_f (\text{products}) - \sum \Delta H_f (\text{reactants})$
- ✓ $\Delta H_{\text{reaction}} = \sum \Delta H_C (\text{reactants}) - \sum \Delta H_C (\text{products})$
- ✓ $\Delta H_{\text{reaction}} = \sum \Delta H (\text{bonds broken}) - \sum \Delta H (\text{bonds made})$

Conversion of Units

SI Units

A physical unit gives a sense of scale to a value and allows comparison of different numbers. As long as, e.g., I know the *gram* is smaller than a *kilogram*, then I know that 20 g of NaCl represents a *much lower* mass than 20 kg. This may be an obvious point but it never ceases to amaze me how many students completely disregard units and their inter-conversions.

As a general rule of thumb, if the examiners do not ask for a unit then it is best *not* to stipulate one. This is because in mark schemes a correct unit is often

seen as neutral whereas an incorrect one can be penalised. This said, many calculations in chemistry *do* require you to interconvert between units and do clearly specify that you should include a suitable unit in your answer.

As there are so many different units of physical parameters such as temperature, mass, amount, length, volume, pressure etc., the scientific community uses an International Standard (*SI*) unit. Smaller or larger subunits extend out from this, usually in factors of 1000. On the table below, the units in **bold** type are the commonest ones for that parameter that you will use at AS level.

1×10^{-6}	1×10^{-3}	SI unit	1×10^3	1×10^6
Milligram (mg)	gram (g)	Kilogram (kg)	Tonne	
Micromole (μmol)	Millimole (mmol)	Mole (mol)		
Centimetre cubed (cm^3)	Decimetre cubed (dm^3)	Meters cubed (m^3)		
		Pascals (Pa)	Kilopascals (kPa)	Megapascals (MPa)

If you are converting units from **left** \rightarrow **right**, then multiply by the number by 1×10^{-3} (**which is the same thing** as saying *divide* by 1000) for each box covered. So, e.g., 20 mg = 20×10^{-3} g = 20×10^{-6} kg etc.

If you are converting units from **right** \rightarrow **left**, then multiply by 1×10^3 (**or times** by 1000) for each box covered. So e.g., 3 tonnes = 3000 kg = 3,000,000 g.

For *temperature*, the SI unit is the Kelvin, *K*. The *Celsius* scale is 273° lower than this so:

- ✓ To convert Celsius to Kelvin, *add* 273 to the $^{\circ}\text{C}$; 25°C is 298 K
- ✓ To convert Kelvin to Celsius, *subtract* 273 from the K; $0\text{K} = -273^{\circ}\text{C}$.

Absolutes and Concentrations

It is important to distinguish between *absolutes* and *concentrations*. An *absolute* is a physical amount of material, be it 10 g of NaOH or 25 dm^3 of carbon dioxide. On the other hand, a concentration represents a certain *amount* of a substance *in a stated volume*. The *molar concentration* is expressed in mol dm^{-3} , the same as saying *moles per decimetre cubed*, or *molarity*, *M*. So a 1 mol dm^{-3} solution of sodium chloride contains one mole of sodium chloride dissolved in every decimetre cubed of solution.

In many of the mole questions you will face, especially those involving *titrations*, you will have to apply the process of *proportions* to convert a concentration to an absolute or *vice versa*. For example, if there is 1 mole of NaCl per decimetre cubed (1000 cm^3) then there is 0.5 mol in 500 cm^3 or 0.1 mol in 100 cm^3 . How you do this effectively will be included in the section on titration questions.

Graphs

A graph is a diagrammatic means of representing the relationship between two variables. There are many types of graph, such as line, scatter, pie, histogram some of which you should probably be familiar with.

In chemistry, line graphs are one of the most important types you will meet and much of your practical work will look at how a *dependent variable* (shown on the *y-axis*), changes as the *independent variable* (shown on the *x-axis*) changes.

When a graph is plotted there are a number of rules you need to follow:

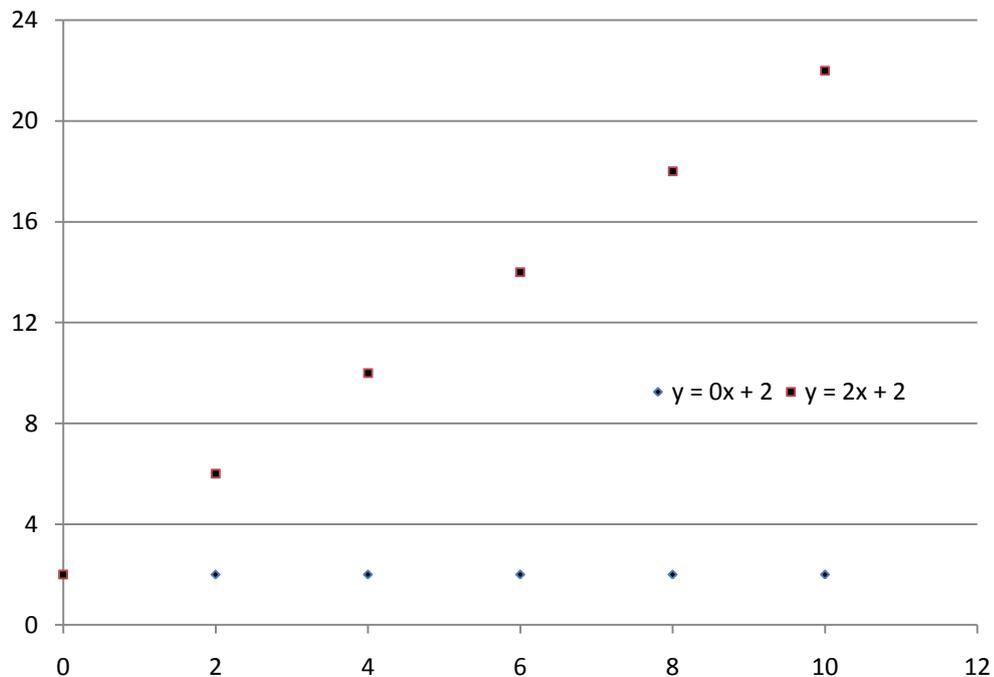
- ✓ Use as much of the graph paper as possible.
- ✓ The dependent variable (e.g. colour, rate, volume, mass etc.) is placed on the vertical (*y*-) axis. If in doubt this is usually the *thing* you are *measuring*.
- ✓ The independent variable (e.g. time, concentration) is placed on the horizontal axis (*x*-) axis. If in doubt this is usually the *thing* you are *changing*.
- ✓ You must always label the axis stating clearly what the variable is and give its unit e.g. time (s) or mass (g).

Linear relationships

In maths, a linear relationship can be described by the equation $y = mx + c$ where *m* is the *gradient*, the slope of the line and *c* is the *intercept*, the point at which the line crosses the *y-axis* when $x = 0$. Consider the table of data over the page:

x	$y = 0x + 2$	$y = 2x + 2$
0	2	2
2	2	6
4	2	10
6	2	14
8	2	18
10	2	22

When the data are plotted the following two graphs are obtained. Fitting a straight *line of best fit* through the data would show that when $x = 0$, $y = 2$ in both cases, hence the intercept, $c = 2$.

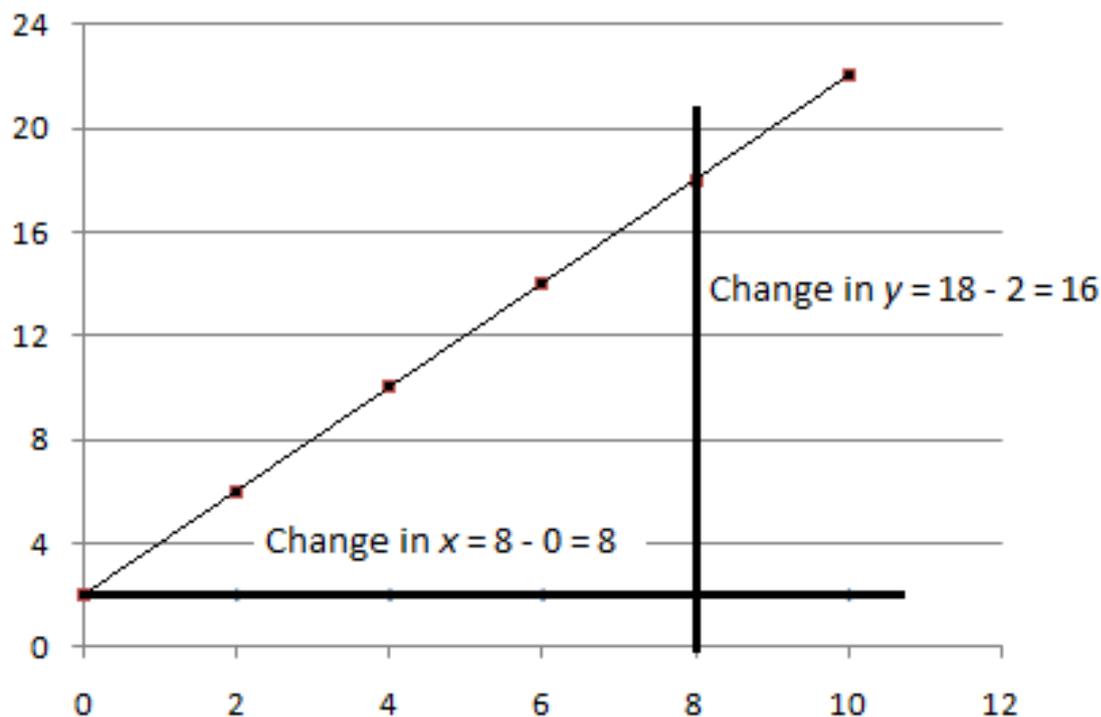


The *gradient* of the line, the value of m , can be found by using the equation:

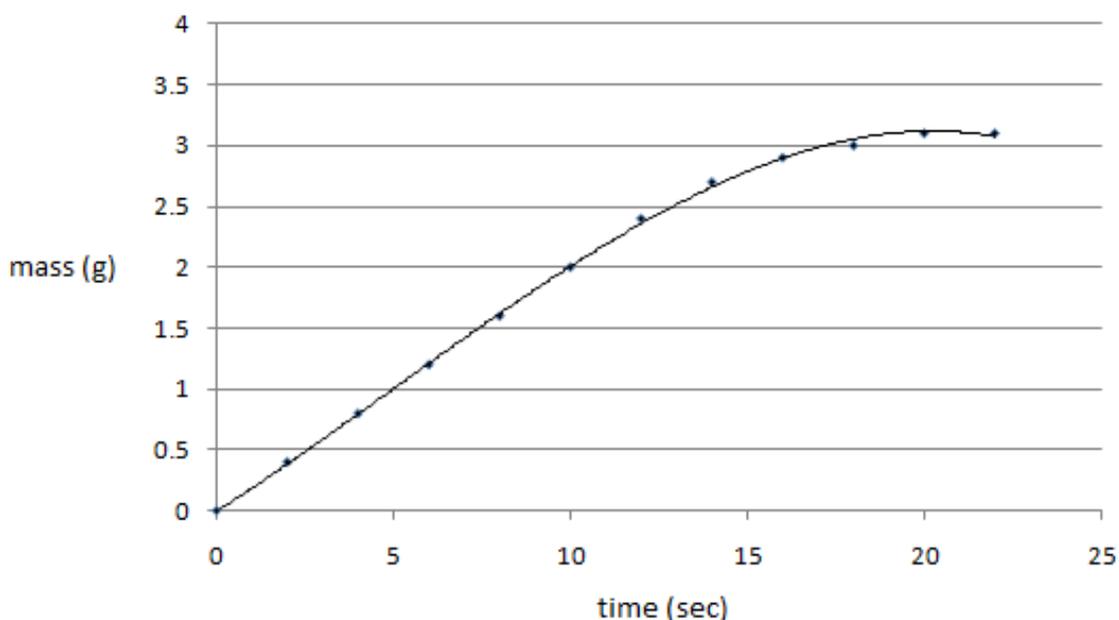
$$\text{Gradient} = \frac{\text{change in } y}{\text{change in } x}$$

For linear relationships this can be found from the graph as shown below:

- From low down on the y -axis where the *line of best fit* crosses, draw a *horizontal line* across the page.
- Draw a *vertical line* up from the x -axis to cross your horizontal line and the *line of best fit*.
- It is always best to use the lines across that use up more than half the graph paper, in order to get the most precise estimate of the gradient.
- Calculate the difference in the points at which the lines cross the *line of best fit*, both vertically (the change in y) and horizontally (the change in x). This is shown on the graph below.
- The gradient is then found using the equation above.

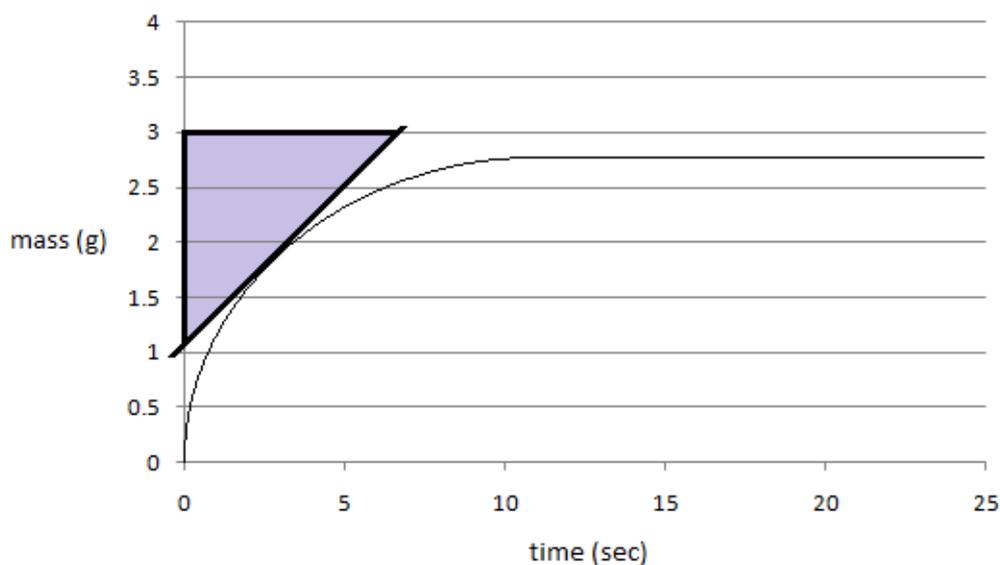


In chemical reactions, the *rate of reaction* can be determined experimentally using this method. When a graph of, for example, mass of product formed is plotted against time, the gradient of the relationship at any given time is the reaction rate (expressed in *g per second*, or g sec^{-1} in this case).



Notice that at first there is a straight line relationship between mass and time, so the mass of product formed is proportional to time. As the reaction proceeds towards completion the graph starts to tail off and become flat.

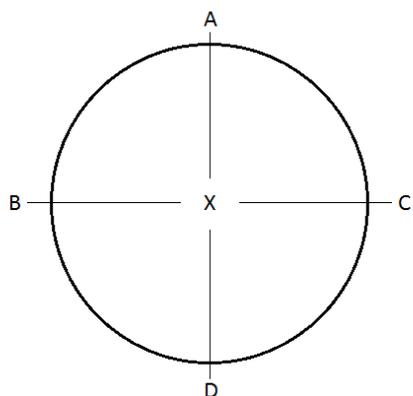
If the rate of reaction is needed at a given point on this *shoulder* region, it can be found by drawing a tangent to the curve and determining *its* gradient. A tangent is a straight line or plane that touches a curve or curved surface at a point but does not intersect it at that point. Calculation of the gradient of the tangent by using the method outlined above will give the rate of reaction.



Geometry

When studying bond shapes, the bond angles around any three atoms are important. These atoms may be arranged two-dimensionally (planar or flat) or three-dimensionally.

Imagine a circle split in two by a cross with a centre point (X) as shown in the diagram below. The points are also shown where the intersecting line crosses the circle at A, B, C and D.

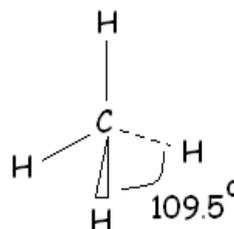


You need to be aware that the angle, between X and any two *opposite* points on the diagram (A-X-D or B-X-C), is **180°**. Between any two *adjacent* points in the diagram (A-X-B or C-X-D etc), the angle is **90°**. You may also no doubt recall that when a circle is split into three then the internal angle is **120°**. A familiarity with these angles will expedite your understanding of molecular shapes later in the AS modules.

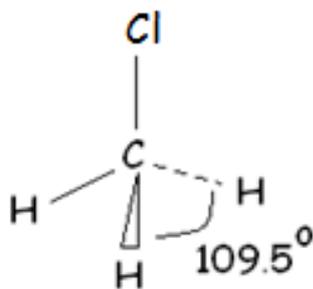
Molecules can be either *two dimensional* or *three dimensional* structures in which the atoms are arranged in certain positions relative to one another. As you progress through your studies, you will learn the angle that these adopt in space reflects the number and arrangement of the electrons around the central atom.

But paper is a two-dimensional medium and it is therefore difficult to represent three-dimensional structures on paper. A convention has been developed which attempts to show this.

If a bond is considered to be in the plane of the paper then it is shown as a *solid* line, if it goes back away from the paper then it is a *dashed* line, and if it comes forward *out* from the plane of the paper then it is shown as an *enlarging* triangle, as shown for methane below. In such cases as in methane when the number of bond is four, the shape forms a regular tetrahedron (four-sided figure) with a bond angle of 109.5° .



Another key concept for molecular shapes is symmetry. Symmetry is an attribute of a shape; if a shape is symmetrical it forms an exact reflection on opposite sides of a dividing line or plane. The circle shown on the previous page is symmetrical and if it was folded along the



line A-X-D, the two halves would superimpose exactly. Visualising symmetry in three-dimensions is harder. Methane, above, is *symmetrical* in three-dimensions because the molecule is the same when viewed from different directions. A similar molecule, chloromethane, CH_3Cl is *not symmetrical* as the presence of the chlorine atom means the molecule is no longer the same when viewed in different positions.

End of Section 1 Test

Fractions, decimals, ratios and percentages

1. Express the following fractions as (i) percentages, (ii) decimals and (iii) standard form. Give all your answers to *three* significant figures:

a. $\frac{5}{81}$ (i) % (ii) (iii)

b. $\frac{4}{19}$ (i) % (ii) (iii)

c. $\frac{1}{125}$ (i) % (ii) (iii)

2. In a recent college chemistry test it was shown that 18 of out every 57 students gained a grade B or above.

- a. Express this information as (i) a percentage, (ii) a decimal and (iii) in standard form. Give all your answers to three significant figures.

(i) % (ii) (iii)

- b. Another teacher is setting the same test. How many students would he expect to achieve a grade B or above from a class of 36?

- c. Using the same test a third teacher found that four students gained grade B or above. How many students were in her class?

Score _____ / 14 [_____ %]

Powers and logs

3. Calculate the following:

a. 4^5 _____ b. 2^3 _____ c. 3^7 _____ d. 4^2 _____

4. Determine the *logarithms* of the following *numbers*:

a. 59 _____ (1dp)

b. 7892 _____ (2 sig fig)

c. 4.58×10^{-6} _____ (3 sig fig)

d. 4^2 _____ (4 sig fig)

e. 1.111 _____ (2 sig fig)

5. Determine the *numbers* of the following *logarithms*. Give your answers to the appropriate precision.

a. 59 _____ (2 sig fig)

b. -0.7892 _____ (4 sig fig)

c. 4.58×10^{-6} _____ (1 sig fig)

d. $4^{0.3}$ _____ (1 sig fig)

e. 10.8 _____ (standard form)

Score _____ / 14 [_____ %]

Rearranging equations

6. By rearranging the equation $A \times B \times C = D \times E \times F$, make each of the following the subject of the equation:

(i) $A =$

(ii) $C =$

(iii) $D =$

(iv) $E =$

7. Rearrange the following equation

$$A = \frac{B \times C}{D \times E}$$

(i) $B =$

(ii) $C =$

(iii) $D =$

(iv) $E =$

8. By rearranging the equation $A + B - C = D + E - F$, make each of the following the subject of the equation:

(i) $A =$

(ii) $B =$

(iii) $C =$

(iv) $E =$

(v) $F =$

9. Rearrange the following equation

$$A = \frac{B + C}{1000}$$

- (i) $B =$
 (ii) $C =$
 (iii) $1000 =$

Score _____ / 16 [_____ %]

Unit conversions

10. Express the following *masses* as the specified units:

- a. 10 g = _____ kg c. 145 kg = _____ g
 b. 802 mg = _____ kg d. 14 mg = _____ g

11. Express the following *volumes* as the units specified in the questions:

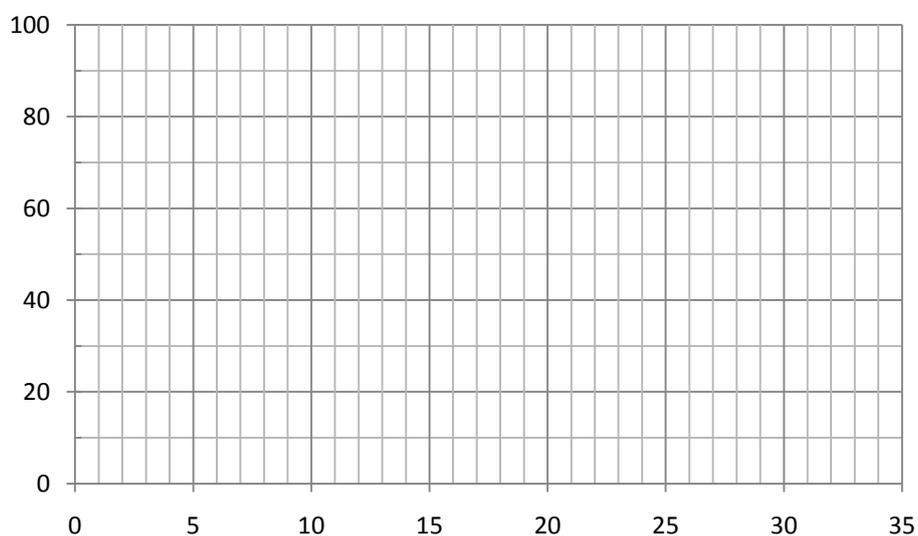
- a. 100 cm³ = _____ dm³ c. 24 dm³ = _____ cm³
 b. 802 m³ = _____ cm³ d. 0.01 dm³ = _____ m³

Score _____ / 8 [_____ %]

Gradient, intercept and rate

12. Plot a graph of the following data and use it to deduce the missing volumes, and calculate the gradient and intercept of the line [6 marks].

Time (seconds)	Volume of CO ₂ released (cm ³)
5	18
10	33
15	48
20	?
25	?
30	93



Score _____ / 8 [_____ %]

Grand total of _____ / 60 = _____ %

Section 2: Calculations in AS Chemistry Specifications

Very few concepts in chemistry strike fear into the hearts of students like moles and mole questions. A close second is energetic calculations, and the AS chemistry specs have both! If the teacher doesn't put it in the right way at the start then your confidence in these relatively straight forward concepts gets dented, and it often doesn't recover even by A2. This section puts this right, and shows you just how to go about moles and energetic content without the fear...

Mole Calculations

Moles and numbers

What is a mole and why use it?

The *mole* is just simply an *amount* or *number*, usually referred to as a certain *amount* of substance. Granted it's a strange number (one mole of any substance contains 6×10^{23} particles). This number is often referred to as *Avogadro's constant* (L). The mole is a fundamental concept in chemistry because when reactions take place it is the *numbers* of particles reacting and **not** their *masses* or *volumes* that is important. Here's a simple analogy:

Suppose you were having three friends around for tea, you might want to make them banana split for pudding. You'd need a total of four bananas for you and your mates. If I sent you out to the shop to get 1 kg of bananas, do you have enough? The answer is simply ... I don't know! It really depends on the mass of one banana. If one banana weighed 200 g then you'd have five bananas, but if each weighed 333 g then you'd only have three. So to put that mathematically...

$$\text{Number (or moles) of bananas} = \frac{\text{mass of bananas bought}}{\text{mass of one banana}}$$

Converting numbers to moles and vice versa

Avogadro's constant (L) tells us that one mole of any particle contains 6×10^{23} particles. It follows that:

$$\text{Number of particles} = \text{moles} \times 6.0 \times 10^{23}$$

$$\text{moles} = \frac{\text{number of particles}}{L}$$

You must be careful *not* to get confused with this seemingly straight-forward concept, but it can be tricky when considering subtle differences between atoms and molecules, or ions. When it comes to atoms, it is an easy conversion:

One mole of argon, Ar contains $1 \times 6 \times 10^{23} = 6 \times 10^{23}$ argon **atoms**.

Two moles of argon, Ar contains $2 \times 6 \times 10^{23} = 12.04 \times 10^{23}$ argon **atoms**.

One mole of chlorine, Cl₂ contains $1 \times 6 \times 10^{23} = 6 \times 10^{23}$ chlorine **molecules**. Because each chlorine **molecule** is made of two chlorine **atoms**, then **one mole** of chlorine contains $2 \times 6 \times 10^{23} = 12 \times 10^{23}$ **atoms**.

One mole of methane, CH₄ contains $1 \times 6 \times 10^{23} = 6 \times 10^{23}$ methane **molecules**. Because each methane **molecule** is made of one carbon **atom** and four hydrogen **atoms**, then **one mole** of methane contains $5 \times 6 \times 10^{23} = 3 \times 10^{24}$ **atoms**. But there would be $1 \times 6 \times 10^{23} = 6 \times 10^{23}$ carbon **atoms** and $4 \times 6 \times 10^{23}$ hydrogen **atoms**.

One mole of sodium chloride, NaCl contains $1 \times 6 \times 10^{23} = 6 \times 10^{23}$ sodium chloride **particles**. Because each sodium chloride **particle** is made of one sodium **ion** and one chloride **ion**, then **one mole** of sodium chloride contains $2 \times 6 \times 10^{23}$ **ions** = 12×10^{23} **ions** in total. But there would be $1 \times 6 \times 10^{23} = 6 \times 10^{23}$ sodium **ions** and $1 \times 6 \times 10^{23} = 6 \times 10^{23}$ chloride **ions**.

Moles and mass

Relative atomic mass

All *substances* have a *molar mass*, the mass of one mole of the substance. For atoms of an element the concept of the *relative atomic mass* (RAM or A_r) is used. This is effectively the weighted average mass of one mole of atoms of an element compared to 1/12 of the mass of one mole of ¹²C atoms. While this is a bit of a mouthful, it is straight forward to use. The values of RAM are found on the Periodic Table, as the larger of the two numbers found top and bottom of each element. For example, the RAM of chlorine, Cl is 35.5, while that of sodium, Na is 23. The unit of RAM is g mol⁻¹, the same as saying *grams per mole*. In other words in 23 g of sodium atoms there would be *one mole* of sodium atoms, or 6×10^{23} atoms.

- Relative isotopic masses are whole numbers as they are the *number* of protons + neutrons.
- Relative atomic masses are *not always* whole numbers. This is because they are composed of different *abundances* of isotopes with different mass numbers.

RAM and mass spectrometry

Relative atomic masses are calculated from the data generated by mass spectrometers. In your AS studies you will learn mass spectrometers generate two numbers, the *abundance* and the *mass to charge ratio*, or m/z of isotopes of elements. These can be used to determine the RAM of the element using the equation:

$$RAM = \frac{\Sigma \left(\text{abundance} \times \frac{m}{z} \right)}{\text{total abundance}}$$

Regardless of whether the data in the question is in the form of text, a table or a graph the calculation is always the same.

- Be careful to **always** write out the substituted equation.
- Check whether the abundance data is given as a percentage, % or numerical values.
- Make sure the value calculated is within the range of the m/z ratios of the isotopes.
- Check how many decimal places or significant figure the question demands.
- Remember a correct answer *does not always* guarantee all the method marks.
- The correct unit for RAM is g mol^{-1} but **only** volunteer this if they ask you to.

? Worked Example 1: Using mass spectrometry a sample of copper was found to be composed of 33.4% ^{63}Cu and 66.6% ^{65}Cu . Use the data to determine the relative atomic mass of the copper in the sample. Give your answer to **one** decimal place.

$$RAM = \frac{(33.4 \times 63) + (66.6 \times 65)}{100} = 64.33$$

$$RAM = \underline{64.3} \text{ (1 dp)}$$

The data may also be presented in the form of a table or graph. Try to look beyond the *way* in which the data is presented. It is *not* trying to confuse you. The *How Science Works* component of the A-level expects you to be able to process and use information in different formats.

In the worked examples below you will notice that although the information looks different, the way in which the question is **always** answered is the same.

? **Worked Example 2:** The data in the table below were obtained from the mass spectrum of a sample of chromium extracted from a meteorite. Give your answer to 3 significant figures.

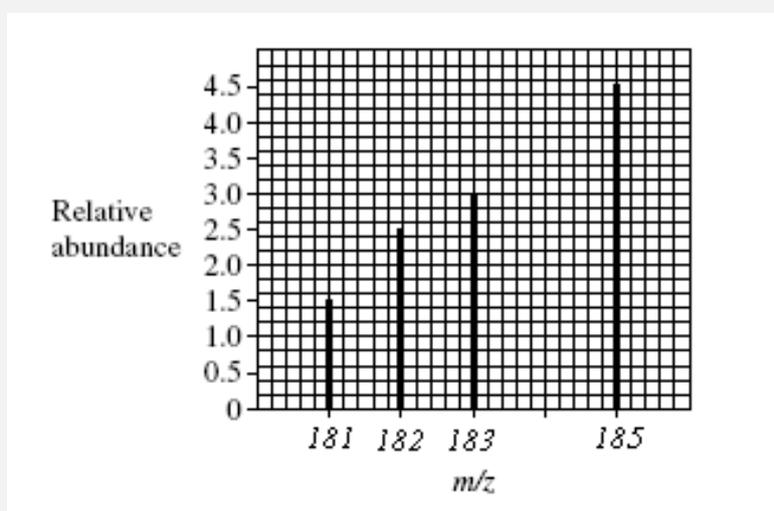
m/z	50	52	53	54
Relative abundance (%)	4.3	83.8	9.5	2.4

Calculate the relative atomic mass of chromium in this sample giving your answer to three significant figures.

$$RAM = \frac{(4.3 \times 50) + (83.8 \times 52) + (9.5 \times 53) + (2.4 \times 54)}{100} = 52.06$$

$$RAM = \underline{52.1} \text{ (3 sig fig)}$$

? **Worked Example 3:** The mass spectrum of element R is shown below. Use this spectrum to calculate the relative atomic mass of R, giving your answer to one decimal place. Identify element R.

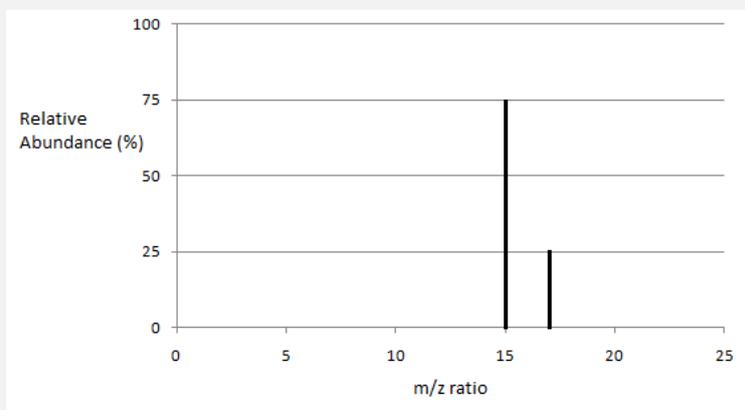


$$RAM = \frac{(1.5 \times 181) + (2.5 \times 182) + (3 \times 183) + (4.5 \times 185)}{(1.5 + 2.5 + 3 + 4.5)} = 183.30$$

$$RAM = \underline{183.3} \text{ (1 dp)}$$

Identity of element R = Tungsten, W

? Worked Example 4: Z is an element extracted from a piece of moon rock. The mass spectrum of element Z is shown below. Use this spectrum to calculate the relative atomic mass of Z, giving your answer to one decimal place. Identify element Z.



$$RAM = \frac{(75 \times 15) + (25 \times 17)}{100} = 15.50$$

$$RAM = \underline{15.5} \text{ (1 dp)}$$

Identity of element Z = Oxygen, O

Relative molecular (and formula) mass

For substances that are composed of molecules such as chlorine Cl₂, the concept of *relative molecular mass* (RMM or M_r) is used. This is effectively the sum of the RAMs of all the atoms which make up the molecule. So in the case of chlorine, the RMM would be $35.5 + 35.5 = 71 \text{ g mol}^{-1}$. In other words in 71 g of chlorine gas there would be one mole of chlorine molecules (6×10^{23} molecules) but as *each* molecule is composed of *two* atoms there would be 12×10^{23} chlorine atoms.

For ionic compounds the concept of *relative formula mass* (RFM) is used. Like RMM, this is effectively the sum of the RAMs of all the *ions* which make up the compound. For sodium chloride, the RFM would be $23 + 35.5 = 58.5 \text{ g mol}^{-1}$.

You will often see the term *molar mass* referred to. This is simply the mass of one mole of a substance. This is a generic term, and so you must be careful when using it in the different contexts of atoms, molecules or ions. Calculation of relative molecular (or formula) masses is a standard expectation in chemistry exams. While you are unlikely to be asked this as a

stand-alone question, a calculation of RMM is often part of bigger mole questions. Molar mass, RMM or RFM are given by the formulae:

$$\text{Molar mass} = \Sigma (\text{RAM of all atomic particles})$$

$$\text{RMM} = \Sigma (\text{RAM of all atoms})$$

$$\text{RFM} = \Sigma (\text{RAM of all ions})$$

Translated, these equations simply mean that to find the RMM or RFM you simply *add up* the RAM of all the atoms involved.

? Worked Example 5: Calculate the relative molecular mass, M_r of butan-2-ol. Give a suitable unit.

Butan-2-ol has the structural formula $\text{CH}_3\text{CH}(\text{OH})\text{CH}_2\text{CH}_3$ and so has a molecular formula $\text{C}_4\text{H}_{10}\text{O}$.

Hence the RMM = $(4 \times 12) + (10 \times 1) + (1 \times 16) = \underline{74 \text{ g mol}^{-1}}$

? Worked Example 6: Calculate the relative formula mass of calcium carbonate giving your answer to 3 significant figures.

Calcium carbonate has the formula CaCO_3 .

Hence the RFM = $(40.1 \times 1) + (12 \times 1) + (3 \times 16) = 100.1$
= 100 (3 sig fig)

% Compositions

One commonly included component of multi-part questions at A-level is the calculation of percentage by mass of a given element in a compound. Given that the different elements in a compound have different relative atomic masses, each one contributes a different mass to the overall RMM or RFM.

The % by mass of an element in a compound is calculated by dividing its contribution to the mass by the total RMM (RFM) and multiplying by 100. Put mathematically this is given by the equation:

$$\% \text{ by mass of } X = \frac{\text{number of atoms } (X) \times \text{RAM } (X)}{\text{RMM or RFM of compound}} \times 100$$

$$\% \text{ by mass of Na in NaCl} = \frac{1 \times 35.5}{58.5} \times 100 = \mathbf{60.7\%}$$

? Worked Example 7: Calculate the percentage by mass of iron in iron(III) oxide. Give your answer to three significant figures.

Iron(III) oxide has the formula Fe_2O_3 .

The % by mass of Fe =

$$\frac{55.8 + 55.8}{55.8 + 55.8 + 16 + 16 + 16} \times 100 = 69.92$$

$$= \mathbf{69.9\% (3 \text{ sig fig})}$$

Mass

The interconversion of mass and moles is one of the commonest you will perform in questions. The relationships are given by:

$$\text{Moles} = \frac{\text{mass}}{\text{molar mass}}$$

$$\text{Molar mass} = \frac{\text{mass}}{\text{moles}}$$

$$\text{Mass} = \text{moles} \times \text{molar mass}$$

Empirical and molecular formulae

The **empirical formula** of a compound is usually defined as the *simplest whole number ratio* of atoms of each element in a *compound*. The **molecular**

formula on the other hand is the *actual number* of atoms of each element in a *molecule*. The concept has particular relevance to organic chemistry where there are a wide range of compounds containing different numbers of carbon, hydrogen and oxygen atoms in the molecules.

For example the *molecular formulae* of ethene and propene are C_2H_4 and C_3H_6 respectively. In both cases the simplest whole number ratio of carbon to hydrogen is 1:2, and so the *empirical formula* of **both** would be CH_2 .

The process of finding the empirical formula from % by mass data usually follows the same steps:

1. In a table, write out the % by mass for each element. The % of any *missing* element can be found by *subtracting the sum of the other elements* from 100.
2. Find mole **ratio** by dividing each % by mass by the *relative atomic mass* of each element in turn. Be careful to use **3 sig fig** so as not to *over round*.
3. Divide *all* the elements by the *smallest* mole ratio calculated in step 2.
4. If the ratios obtained are whole numbers then this is now the empirical formula of the compound.

Note: If you get non-whole number ratios such as 0.333, 0.250, 0.500 then multiply **all** of the element's ratios by 3, 4 and 2 respectively.

To deduce the molecular formula from this information you will need the relative molecular mass (RMM) and empirical mass (EM) of the compound. The number of times the empirical formula *goes into* the molecular formula is give by:

$$\text{number of units} = \frac{RMM}{EM}$$

The multiply the empirical formula by the number of units to get the molecular formula.

- ☛ Be careful not to over round the empirical formula calculations. This will lead to spectacularly wrong answers!
- ☛ Always use 3 sig fig.
- ☛ One of the commonest errors is to invert the first division. You **must** check that you *divide the % by RAM* and *not* the other way around.

? Worked Example 7: Determine the empirical formula of a hydrocarbon that is 82.8 % carbon by mass. Given the relative molecular mass of the hydrocarbon is 58, deduce the molecular formula.

Element	Carbon (%)	Hydrogen (%)
% by mass	82.8	100 - 82.8 = 17.2
Mole ratio	$\frac{82.8}{12} = 6.90$	$\frac{17.2}{1} = 17.2$
Simplest ratio	$\frac{6.90}{6.90} = 1$	$\frac{17.2}{6.90} = 2.49$
Adjustment for a "0.5 ratio" = x 2	2	5

Hence the empirical formula = C₂H₅

Empirical mass = (2 x 12) + (5 x 1) = 29

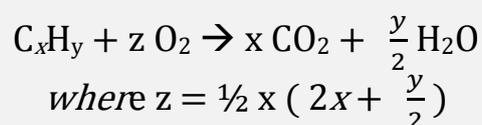
Given the relative molecular mass = 58

Number of empirical units = 58/29 = 2 and so **2 empirical** units are in the **molecular** formula.

Molecular formula = C₄H₁₀

Combustion analysis

When hydrocarbons are burned in a plentiful supply of oxygen they form CO₂ and H₂O. You may not have seen this kind of equation before but *trust me* ☺ when I say you can *simplify* it to....



The volumes of the gases, or masses of the water formed can therefore be used to deduce the empirical formula of the compound. If the molar mass of the original compound is known, then the molecular formula can also be deduced using the following steps:

1. Deduce the moles of carbon dioxide and the moles of water formed.
2. As *one mole* of C in the compound forms *one mole* of CO₂ it follows that the moles of C in original compound = moles of CO₂ formed.
3. As *two moles* of H in the compound form *one mole* of H₂O it follows that the moles of H in the original compound = moles of H₂O formed x 2.
4. Determine the empirical ratio of C to H, and hence the empirical formula by dividing the *larger moles* by the *smaller moles* as above.
5. Determine the molecular formula using the empirical mass and RMM (as above).

? Worked Example 8: A hydrocarbon W ($M_r = 70 \text{ g mol}^{-1}$) was burned in an unlimited supply of oxygen. On combustion, 4.9 g of W formed 8.4 dm³ of carbon dioxide and 6.3 g of water. Use this information to deduce (i) the empirical formula, (ii) the molecular formula and (iii) deduce the identify of W. Under the conditions used, the molar volume was 24 dm³.

Substance formed	Carbon dioxide (8.4 dm ³)	Water (6.3 g)
Moles	$\frac{8.4}{24} = 0.350$	$\frac{6.3}{18} = 0.350$
Elements	As $n\text{C}$ in W = $n\text{CO}_2$ $n\text{C}$ in W = 0.350	As $n\text{H}$ in W = $2 \times n\text{H}_2\text{O}$ $n\text{H}$ in W = 0.700
Empirical ratio	$\frac{0.350}{0.350} = 1$	$\frac{0.700}{0.350} = 2$

The empirical formula = CH₂, so empirical mass = (1x12)+(2x1) = 14

Number of empirical units in W = $\frac{70}{14} = 5$

Therefore the molecular formula is C₅H₁₀. W is propene.

Moles and solution volumes and concentrations

When a substance gets dissolved in a certain volume of solution, it is not only important to consider the *amount* (in moles) that is absolutely present, but also its *concentration* (the amount per unit volume). In chemistry the *molar concentration*, in mol dm^{-3} (often spoken as *moles per decimetre cubed*) is a universally accepted unit, though its other name, *molarity* given by the symbol, *M* is sometimes used. GCSE boards often used the unit, mol/dm^3 . These are all the same thing, but you must standardise the way you present your concentrations according to how the board uses it in their papers.

The only trickiness to consider when converting moles into concentrations and *vice versa* is the habit of examiners to sometimes change the unit of volume in the question. Volumes in AS chemistry practicals are often in the form of cm^3 , whereas the volume part of the molar concentration unit is dm^3 .

This **does not** require a conversion **if** the volumes are in **dm³**:

$$\text{Moles} = \text{concentration (mol dm}^{-3}\text{)} \times \text{volume (dm}^3\text{)}$$

$$\text{or ... "Moles} = \text{conc} \times \text{vol"}$$

This **does** require a conversion **if** the volumes are in **cm³**:

$$\text{Moles} = \frac{\text{concentration (mol dm}^{-3}\text{)} \times \text{volume (cm}^3\text{)}}{1000}$$

$$\text{or ... "Moles} = \frac{\text{conc} \times \text{vol}}{1000}$$

It's therefore **very important** that you pay close attention to the units of *volume* that are stated in the questions so you **do not fall** into a clearly marked *trap*! Given that we can now express moles in terms of concentration and volume, we will be expected to rearrange the equation to express concentrations and volumes in terms of moles.

☛ You should practice these conversions very often by writing them out **over and over** again. They can cost a lot of marks if you mess them up in multi-part questions!

Since in AS almost 100% of exam questions use volumes in cm^3 , I shall perform the rearrangements on that basis. It follows that:

$$\text{moles} = \frac{\text{concentration} \times \text{volume}}{1000}$$

$$\text{concentration} = \frac{\text{moles} \times 1000}{\text{volume}}$$

$$\text{volume} = \frac{\text{moles} \times 1000}{\text{concentration}}$$

? Worked Example 9: A student dissolved 0.1 mol of NaCl into 250 cm^3 of solution. Use this information to calculate (a) the concentration of NaCl in mol dm^{-3} , (b) the volume of solution that contains 0.055 mol (stating an appropriate unit), and (c) the number of moles of NaCl in 89 cm^3 of the solution. Give your answers to 2 sig fig.

$$(a) \text{ Concentration} = \frac{0.1 \times 1000}{250} = 0.40 \text{ mol dm}^{-3}$$

$$(b) \text{ Volume} = \frac{0.055 \times 1000}{0.4} = 137.5 \text{ cm}^3 \text{ *ECF}$$

$$(c) \text{ Moles} = \frac{0.4 \times 89}{1000} = 0.0356 \text{ mol} = 0.036 \text{ (2 sig fig) *ECF}$$

**ECF means error carried forward. If you made a mistake in calculating the original concentration then it is deemed unfair to continue to penalise you for other questions that relied on that result. So long as you use the mistaken answer correctly afterwards, you will be awarded these marks.*

? Worked Example 10: A teacher spilled 450 cm^3 of a 1 mol dm^{-3} solution of sodium hydroxide in the laboratory. Calculate the number of moles of NaOH spilled giving your answer to 2 decimal places.

$$\text{Moles} = \frac{1 \times 450}{1000} = 0.45 \text{ mol dm}^{-3}$$

Moles and gas volumes

You will need to check which version of the moles and gas volumes you need to be using for your specific board. Essentially there are two concepts (i) the *molar volume*, and (ii) the *ideal gas equation*.

The molar volume

The molar volume is the volume occupied by 1 mole of any gas at room temperature and pressure. It is assumed to be 24 dm³ (or 24000cm³). This is inaccurate as the terms *room temperature* and *room pressure* are not standardised.

For volumes expressed in **dm³**, it follows that:

$$\text{moles} = \frac{\text{volume}}{24}$$

$$\text{Volume} = \text{moles} \times 24$$

For volumes expressed in **cm³**, it follows that:

$$\text{moles} = \frac{\text{volume}}{24000}$$

$$\text{Volume} = \text{moles} \times 24000$$

? Worked Example 11: When a student reacted some of calcium carbonate with an excess hydrochloric acid she found that 75 cm³ of CO₂ was released. (a) Calculate the number of moles of gas produced. (b) Calculate the volume of gas, in dm³ that would have been produced if she had used enough calcium carbonate to yield 0.5 moles of the gas.

$$(a) \text{ moles of gas} = \frac{75}{24000} = 3.13 \times 10^{-3} \text{ mol}$$

$$(b) \text{ volume of gas} = 0.5 \times 24 = 12 \text{ dm}^3$$

The idea gas equation, $PV = nRT$

The ideal gas equation represents a more accurate way of estimating volumes, pressures and temperatures of gases, but even this is flawed to a certain extent. It assumes that a gas shows *ideal* behaviour which is rarely the case. That said it is a straight forward equation to use and rearrange. If you are in any doubt use the *linearised method* that was described in Section 1.

$$PV = nRT$$

$$V = \frac{nRT}{P}$$

$$P = \frac{nRT}{V}$$

$$n = \frac{PV}{RT}$$

$$T = \frac{PV}{nR}$$

Where: P is Pressure (Pa), V is Volume (m^3), T is Temperature (K), n is moles (mol) and R is the gas constant ($J K^{-1} mol^{-1}$)

☛ The units for these equations are fixed. You must pay special attention to the units given in the question and convert them accordingly.

? Worked Example 12: Calculate the pressure, in kPa (3 sig fig) exerted by 0.15 mol of methane in a vessel of 234 cm^3 at a temperature of 16°C . The gas constant, R has a value of $8.31 \text{ J K}^{-1} \text{ mol}^{-1}$.

First of all do the conversions:

$$n = 0.15 \quad T = 16 + 273 = 289\text{K} \quad V = 234 \times 1 \times 10^{-6} = 2.34 \times 10^{-4} \text{ m}^3$$

$$R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$$

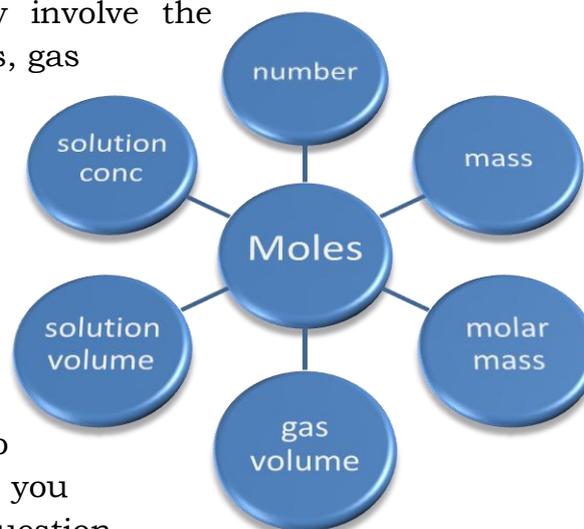
$$\text{as } P = \frac{nRT}{V} \text{ it follows that: } P = \frac{0.15 \times 8.31 \times 289}{0.000234} = 1539480.8 \text{ Pa}$$

$$\text{Pressure} = 1539480.7 \times 1 \times 10^{-3} = \underline{1540 \text{ kPa}} \text{ (3 sig fig)}$$

Putting it all together

Mole calculations at A-level usually involve the interconversions of moles with masses, gas and solution volumes and concentrations. The last few pages have described how each one can be linked to moles and back individually. But using the diagram on the right it is possible to see how several could be linked together in a question.

In such cases the strategy should be to use the information provided to calculate moles. Once you have these, you can then link to the final part of the question using the equations you have been familiarising yourself with earlier in this section. Many of the more advanced titration questions, described later will use this approach.

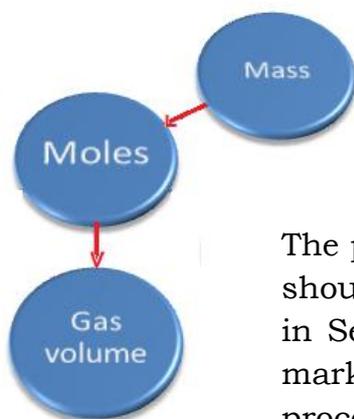


? **Worked Example 13:** What volume, in cm^3 is occupied by 88 g of CO_2 .

$$\text{moles of carbon dioxide} = \frac{\text{mass}}{\text{RMM}} = \frac{88}{12+16+16} = 2 \text{ mol}$$

At room temperature and pressure 1 mole occupies 24 dm^3

$$\text{Volume} = \text{moles} \times 24 = 2 \times 24 = 48 \text{ dm}^3$$



Notice how *Worked Example 13* used the fragment of the diagram shown on the left. Mass was converted into moles and then moles on to volume.

The permutations are too numerous to cover here but you should hopefully get the general idea. The exam questions in Section 3 will provide you with more practice and the mark schemes have worked answers so you can follow the processes involved.

Atom economy

Atom economy is a new concept to AS chemistry, introduced with these new A-levels since 2008/2009. It is usually defined as:

$$\% \text{ atom economy} = \frac{\text{mass of desired product} \times 100}{\text{total mass of products}}$$

Note that because an equation is balanced on either side it can be defined as:

$$\% \text{ atom economy} = \frac{\text{mass of desired product} \times 100}{\text{total mass of reactants}}$$

In practice the % atom economy is calculated using the *Mr* of the compounds in the equations and so is more accurately given by:

$$\% \text{ atom economy} = \frac{\text{molar mass of desired product} \times 100}{\text{total molar masses of products}}$$

Atom economy is a measure of the efficiency with which reactants are converted to products in a given chemical equation. A low % atom economy means a large amount of waste is formed, which, if the reaction was used in manufacturing would mean a large amount of waste would have to be removed and processed. This would impact severely on the profitability of a given process.

The manufacturer cannot change the atom economy, it is fixed for a given reaction, but he can alter the *actual* reaction used to decrease the amount of waste. For example, consider the production of ethylamine which can be made via two different routes.

Either reaction 1 $\text{CH}_3\text{CN} + 2\text{H}_2 \rightarrow \text{CH}_3\text{CH}_2\text{NH}_2$

Or reaction 2 $\text{CH}_3\text{CH}_2\text{Br} + 2\text{NH}_3 \rightarrow \text{CH}_3\text{CH}_2\text{NH}_2 + 2\text{NH}_4\text{Br}$

In the production of $\text{CH}_3\text{CH}_2\text{NH}_2$ from CH_3CN and H_2 , reaction 1 has an atom economy of 100% as there is *only one product*. However, when producing the same compound from $\text{CH}_3\text{CH}_2\text{Br}$ and NH_3 there is an atom economy of 18.7%:

$$\% \text{ atom economy of reaction 2} = \frac{45 \times 100}{241} = 18.7\%$$

It is clear therefore that a manufacturer would choose the first method as there is no waste and so a much higher profitability.

? Worked Example 14: Calculate the % atom economy for the production of iron from iron(III) oxide.



$$\begin{aligned} \% \text{ atom economy} &= \frac{\text{molar mass of desired product} \times 100}{\text{total molar masses of products}} = \frac{11160}{229.6} \\ &= \underline{\underline{48.6\%}} \end{aligned}$$

Percentage yield

When I ask students the meaning of percentage yield, I usually get told that it's ... " what you got over what you should have got ". I often have to prod them to add ... " x 100!"

Perhaps a little more accurately it is:

$$\% \text{ yield} = \frac{\text{actual amount of product formed} \times 100}{\text{theoretical amount product}}$$

Percentage yield is a measure of efficiency of the conversion of products, and unlike atom economy, it can be manipulated and altered by changing the reaction conditions. A manufacturer might then choose a reaction with a high atom economy but one that gives a low yield. The Haber process is like this. When nitrogen and hydrogen react together, they only form ammonia. This means the atom economy is 100%. But the equilibrium that is established means that the yield obtained is low, even when the conditions are optimised.

Again, the sheer number of possible permutations in a question makes it impossible to cover them all in fine detail, but included below is an example of

the kind of question a student might expect. Percentage yield is best calculated from moles (see the example below) and a generic strategy that's suitable for most types of question is outlined below:

1. Using the equation, identify a *reactant* that you can calculate moles for, and use the information given in the question to calculate it (let's call this ***moles of reactant***).
2. Using the equation, identify a *product* that you can calculate moles for, and use the information to calculate it (let's call this ***moles of product***).
3. Using the equation, determine the mole ratio of the *product to reactant* (let's call it the ***PR ratio***).
4. It follows that:

$$\text{The } \mathbf{\textit{theoretical moles of product}} = \mathbf{\textit{moles of reactant}} \times \mathbf{\textit{PR ratio}}$$

5. It follows that:

$$\% \text{ yield} = \frac{\mathbf{\textit{moles of product}} \text{ formed} \times 100}{\mathbf{\textit{theoretical moles of product}}}$$

I should stress that there are many different strategies for these kinds of questions, but this is a robust step-by-step process that works in most cases.

? Worked Example 15: When 75 cm³ of a 0.1 mol dm⁻³ solution of silver nitrate was added to an excess of sodium chloride solution, 0.9 g of the white precipitate silver chloride, AgCl was formed. Calculate the % yield of AgCl.



$$\text{Moles of reactant, AgNO}_3 = \frac{75 \times 0.1}{1000} = 0.0075 \text{ mol}$$

$$\text{Moles of product, AgCl} = \frac{0.9}{108 + 35.5} = 0.00627 \text{ mol}$$

Product to reactant ration = 1:1 therefore

Theoretical moles of product = 0.0075 x 1 = 0.0075 mol.

$$\% \text{ yield} = \frac{0.00627 \times 100}{0.0075} = 83.6\%$$

Titration

Titration questions, (or associated mole questions) seem to cause more anxiety to students than any others, yet they often follow a standardised process similar to the % yield calculations above. Again, the sheer numbers of possible permutations in a question make it impossible to cover them all in fine detail, but included below is an example of the kind of question a student can expect. Titration questions are best calculated using the suitable generic strategy that is outlined below:

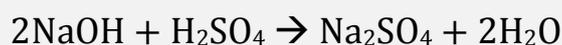
1. Using the equation, identify a *reagent* that you can calculate moles for, and use the information in the question to calculate it (call this **got**).
2. Using the equation, identify a reagent that the question is asking of you, and let's call this the **wanted**.
3. Using the equation, determine the "**want to got ratio**" (the *WG ratio*).
4. It follows that you can deduce the mole you want using:

$$\text{moles wanted} = \text{moles got} \times \text{WG ratio.}$$
5. Shoot for goal! Once you have the moles of the desired reagent from part 4, you can convert this to mass, concentration, volume, or anything else the question demands of you.

Before the worked example it is worth pointing out that this method looks silly, but works brilliantly. You must however be wary of the following:

- ☛ If you get stuck in a multi-part question, put in any value. You will still get errors carried forwards and gain some credit.
- ☛ Remember, scoring 3 or 4 out of 5 is better than 0 out of 5! If you don't think you can go all the way in a mole question, do what you can!
- ☛ Always work out moles and use the ratios, that way you always get at least 2 marks.
- ☛ Using this **want to got** method may seem a little odd at first, but it gets over those concerns of whether to multiply or divide by 2s or 3s when the mole ratios are odd.

? **Worked Example 16:** Sodium hydroxide reacts with sulphuric acid according to the following equation.



A 25 cm³ of a 0.15 mol dm⁻³ solution of sodium hydroxide exactly neutralised 58 cm³ of a dilute sulphuric acid solution. Calculate the concentration of the sulphuric acid.

Mole of *got* (NaOH) =

$$\frac{\text{concentration} \times \text{volume}}{1000} = \frac{25 \times 0.15}{1000}$$

= 0.00375 mol

Reagent you *want* is H₂SO₄

From the equation the WG ratio = 1 H₂SO₄ (*want*) : 2 NaOH (*got*)

Therefore, moles of H₂SO₄ = moles got (NaOH) $\times \frac{1}{2}$

= 0.00375 $\times \frac{1}{2}$ = 0.00188 mol

Question asks for a concentration and:

$$\text{concentration} = \frac{\text{moles} \times 1000}{\text{volume}}$$

$$\text{Concentration of sulphuric acid} = \frac{0.00188 \times 1000}{58} = \underline{\underline{0.0324}} \text{ mol dm}^{-3}$$

Water of Crystallisation

Some ionic compounds are *anhydrous* and do not have any associated water in the crystal structure. Others contain *water of crystallisation* and as such they are termed *hydrated* compounds. You can tell whether a compound is hydrated as it will have a $\cdot x\text{H}_2\text{O}$ in its structure, e.g. $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.

The questions in AS usually involve finding a value for the x in $\cdot x\text{H}_2\text{O}$. It is easiest to imagine the x as a *mole ratio* to the compound. In the above example, the mole ratio of CuSO_4 to H_2O is 1 to 5 for hydrated copper(ii) sulphate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.

The strategy for these questions usually involves:

1. Identifying the difference in the mass of water between the hydrated and anhydrous forms.
2. Calculating the moles of the anhydrous material.

3. Calculating the moles of water.
4. Deducing the mole ratio between the anhydrous and water to calculate a value for x .

? Worked Example 17: When 6.40 g of hydrated magnesium sulphate, $\text{MgSO}_4 \cdot x\text{H}_2\text{O}$ are heated, 3.14 g of anhydrous salt were formed. Deduce the value of x in the formula.

$$\begin{aligned} \text{Mass of water} &= \text{mass of hydrated salt} - \text{mass of anhydrous salt} \\ &= 6.40 - 3.14 \text{ g} = 3.26 \text{ g} \end{aligned}$$

$$\begin{aligned} \text{The moles of anhydrous salt} &= \frac{\text{mass}}{\text{RFM}} \\ &= \frac{3.14}{24.3 + 32.1 + 16 + 16 + 16 + 16} = 0.0261 \text{ mol} \end{aligned}$$

$$\text{The moles of water} = \frac{3.27}{18} = 0.181 \text{ mol}$$

Therefore the mole ratio of $\text{H}_2\text{O} : \text{MgSO}_4 = 0.181 : 0.0261 = 6.93$

$X = 7$, therefore the formula is $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$

Before we finish with moles, let us just summarise some very general key points that we have come across.

- Always do calculations to 3 sig fig unless otherwise stated.
- Never round too soon in a calculation
- Remember that there is usually more than one way to do a calculation.
- **Don't panic!** If you cannot do one step, put in a value so you can get *errors carried forward* marks.

Energetics

Energetics calculations use three main equations. You *must learn* and be *completely familiar* with:

$$\Delta H_{\text{reaction}} = \Sigma \Delta H_{(\text{bonds broken})} - \Sigma \Delta H_{(\text{bonds made})}$$

$$\Delta H_{\text{reaction}} = \Sigma \Delta H_f(\text{products}) - \Sigma \Delta H_f(\text{reactants})$$

$$\Delta H_{\text{reaction}} = \Sigma \Delta H_C(\text{reactants}) - \Sigma \Delta H_C(\text{products})$$

$$\Sigma = \text{sum of, } \Delta H = \text{enthalpy change}$$

Bond Enthalpy Calculations

Definition of bond enthalpy

Bond enthalpies are often defined slightly differently in different boards, and you must always use and learn the definition for the board you are doing! Essentially though, it is defined as ... “The heat **energy change** when **one mole of covalent bonds is broken** in the gas phase.”

Bond enthalpies are *always endothermic*, i.e. heat energy must be taken *in*, in order to break the bonds. This means bond enthalpies *always* have a positive, +ve value when using them in calculations.

ΔH from bond enthalpies

The enthalpy change of a reaction, $\Delta H_{\text{reaction}}$ is considered to be the difference between the amount of energy it takes to break bonds and that released when the new bonds are formed. This is described by the equation:

$$\Delta H_{\text{reaction}} = \Sigma \Delta H_{(\text{bonds broken})} - \Sigma \Delta H_{(\text{bonds made})}$$

The bond enthalpies usually given at AS are *mean* bond enthalpies and as such represent the average values from a range of compounds. The data will be presented in one of two ways to enable you to either calculate $\Delta H_{\text{reaction}}$ from bond enthalpies or *vice versa*. You were shown how to rearrange equations in the **Algebra** section, so if you are unsure go back and try the

examples. The bonds broken are those in the *reactants* and the bonds made are those in the *products*.

☛ Be careful to make sure you use data for **all** the bonds!

ΔH from bond enthalpies

? **Worked Example 18:** Use the following mean bond enthalpies to calculate the ΔH of the reaction between ethene and hydrogen to form ethane.



Bond	C - C	C = C	C - H	H - H
Bond enthalpy, kJ mol^{-1}	348	612	413	436

$$\begin{aligned} \Sigma \Delta H_{(\text{bonds broken})} \\ &= 612 (1 \times \text{C} = \text{C}) + 1652 (4 \times \text{C} - \text{H}) + 436 (1 \times \text{H} - \text{H}) \\ &= 2700 \text{ kJ mol}^{-1} \end{aligned}$$

$$\begin{aligned} \Sigma \Delta H_{(\text{bonds made})} \\ &= 348 (1 \times \text{C} - \text{C}) + 2478 (6 \times \text{C} - \text{H}) \\ &= 2826 \text{ kJ mol}^{-1} \end{aligned}$$

$$\begin{aligned} \Delta H_{\text{reaction}} &= \Sigma \Delta H_{(\text{bonds broken})} - \Sigma \Delta H_{(\text{bonds made})} \\ &= 2700 - 2826 \\ &= \underline{-126} \text{ kJ mol}^{-1} \end{aligned}$$

Bond enthalpies from ΔH

? **Worked Example 19:** Use the following mean bond enthalpy data and the ΔH of the reaction for the formation of ammonia to calculate a value for the mean bond enthalpy of the N-H bond. Give your answer to 3 sig fig.



Bond	$\text{N} \equiv \text{N}$	H - H
Bond enthalpy, kJ mol^{-1}	944	436

$$\begin{aligned} \Sigma \Delta H_{(\text{bonds broken})} \\ &= 944 (1 \times N \equiv N) + 1308 (3 \times H - H) \\ &= 2252 \text{ kJ mol}^{-1} \end{aligned}$$

$$\begin{aligned} \Sigma \Delta H_{(\text{bonds made})} \\ &= 6x (3 \times N - H) \\ &= ? \text{ kJ mol}^{-1} \end{aligned}$$

$$\begin{aligned} \Delta H_{\text{reaction}} &= \Sigma \Delta H_{(\text{bonds broken})} - \Sigma \Delta H_{(\text{bonds made})} \\ -92 &= 2252 - 6x \end{aligned}$$

$$\begin{aligned} \text{Rearranging,} & \quad 6x - 92 = 2252 \\ \text{Hence} & \quad 6x = 2252 + 92 \end{aligned}$$

$$\text{Calculating,} \quad 6x = 2344$$

$$\text{Hence} \quad x = \frac{2344}{6} \quad \text{so } x = 391$$

Mean bond enthalpy of N - H = + 391 kJ mol⁻¹

Hess' Law Calculations

The definition of Hess's law may vary from board to board, but the meaning is the same. It states that ... "the enthalpy change for a reaction is independent of the route taken from reactants to products."

There are two main approaches to using Hess's law to calculate the ΔH of a reaction; either by (i) linking ΔH of reaction to the standard enthalpies of formation, ΔH_f^\ominus or standard enthalpies of combustion, ΔH_c^\ominus using an equation (the term **equation approach** will be used here), or (ii) with an energy cycle (the term **cycle approach** will be used here).

ΔH of a reaction using enthalpies of formation

Enthalpies of formation, ΔH_f^\ominus are often defined slightly differently in different boards, and you must always use and learn the definition for the board you are doing! Essentially though, it is defined as ... "The **enthalpy change** when **one mole** of a substance is formed from its **constituent elements under standard conditions**, all reactants and products in their **standard states**."

Enthalpies of formation can be *negative* or *positive*, and in general the more negative it is, the more thermodynamically stable the compound is. It is important to remember that the ΔH_f^\ominus of all elements is **zero** as there is no net release of energy when one mole of element is formed from one mole of element!

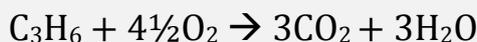
The equation approach:

$$\Delta H^\ominus_{\text{reaction}} = \Sigma \Delta H_f^\ominus (\text{products}) - \Sigma \Delta H_f^\ominus (\text{reactants})$$

This is the *simpler* of the two approaches; the tricky part is to make sure you get the products and reactants part the right way around, and to be careful with any negatives.

- ☛ Remember that the unit of energy is kJ mol^{-1} so when there is more than one mole of a substance in the equation, its ΔH_f^\ominus must be **multiplied by the number of moles**.
- ☛ Remember that the ΔH_f^\ominus of all elements is zero so they may not appear in the data.
- ☛ They will try to confuse you with phrases like ...“use the enthalpies of formation to calculate the enthalpy of combustion, “ or so on. Don't panic! Always **use** the information they give you and *not* what they are asking you to do.

? Worked Example 20: Use the following standard enthalpies of formation to calculate the enthalpy of combustion of propene. All reactants and products are gaseous.



Compound	C_3H_6	CO_2	H_2O
Standard enthalpy of formation, $\Delta H_f^\ominus \text{ kJ mol}^{-1}$	+20	-394	-242

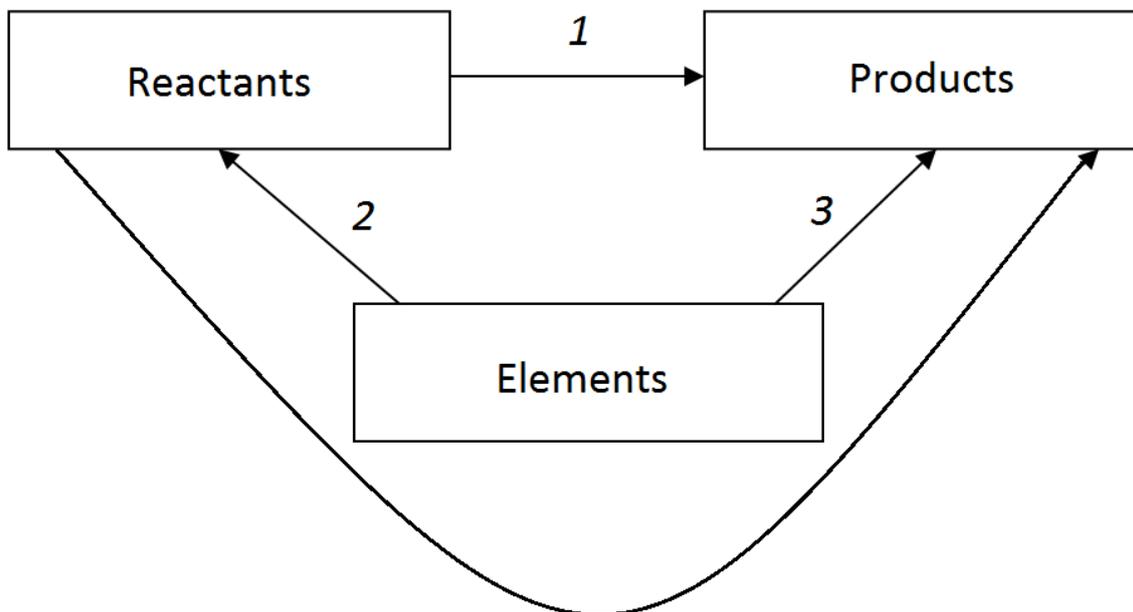
$$\Delta H_{\text{reaction}} = \Sigma \Delta H_f (\text{products}) - \Sigma \Delta H_f (\text{reactants})$$

$$\Delta H_{\text{reaction}} = [(3 \times -394) + (3 \times -242)] - [+20]$$

$$= -1182 + (-726) - 20 = -1928 \text{ kJ mol}^{-1}$$

The cycle approach:

This approach is a visual representation of the equation approach. In the scheme below, reaction 1 is the $\Delta H_{\text{reaction}}$ whereas reactions 2 and 3 represent the $\Delta H_{\text{f}}(\text{reactants})$ and the $\Delta H_{\text{f}}(\text{products})$ respectively.



- (i) You are calculating ΔH for the reaction that moves from the reactants box to the products box (reaction 1).
- (ii) The enthalpies of formation are the reactions represented by arrows 2 and 3 as they are the energy changes **from** the elements **to** the reactants and products.
- (iii) Now imagine following the curved arrow from the reactants to the elements and then on to the products. This is like train journey that you cannot move directly from Manchester to London, but have to change at Birmingham.
- (iv) If you move **against** the arrow then **change** the sign of the ΔH_{f} , but if you're travelling **with** the direction as the arrow, then **keep** the sign.
- (v) So, in the above case $\textcircled{1} = -\textcircled{2} + \textcircled{3}$. This is exactly the same as saying $\textcircled{1} = \textcircled{3} - \textcircled{2}$ or $\Delta H_{\text{reaction}} = \sum \Delta H_{\text{f}}(\text{products}) - \sum \Delta H_{\text{f}}(\text{reactants})$

This approach gets you exactly the same answer for the worked example above.

ΔH of a reaction using enthalpies of combustion

Enthalpies of combustion, $\Delta H_{\text{C}}^{\ominus}$ are often defined slightly differently in different boards, and you must always use and learn the definition for the board you are doing! Essentially though, it is defined as ... “The **enthalpy change** when **one mole** of a substance is **completely burned in excess oxygen under standard conditions**, all reactants and products in their **standard states**.”

Enthalpies of combustion are always *negative*, and in general the more negative it is, the more thermodynamically **unstable** the compound is.

The equation approach:

$$\Delta H_{\text{reaction}} = \sum_{\text{c r}} \Delta H_{\text{C}} (\text{reactants}) - \sum_{\Delta \text{ p}} \Delta H_{\text{C}} (\text{products})^*$$

**Some students refer to this as the cr Δ p approach to help themselves remember which way around it is. Check out what the letters spell ☺.*

This is the *simpler* of the two approaches; the tricky part is to make sure you get the products and reactants the right way around, and to be careful with any negatives.

- ☛ Remember that the unit of energy is kJ mol^{-1} so when there is more than one mole of a substance present, its $\Delta H_{\text{C}}^{\ominus}$ must be **multiplied by the number of moles**.
- ☛ Remember that the $\Delta H_{\text{C}}^{\ominus}$ of all elements is **never** zero; they **will** appear in the table.
- ☛ They will try to confuse you with phrases like ...“use the enthalpies of combustion to calculate the enthalpy of formation, “ or so on. **Don't panic!** Always use the information they give you and *not* what they are asking you to do.

? Worked Example 21: Use the following enthalpies of combustion to calculate the enthalpy of formation of propene. Note that all reactants and products are gaseous.

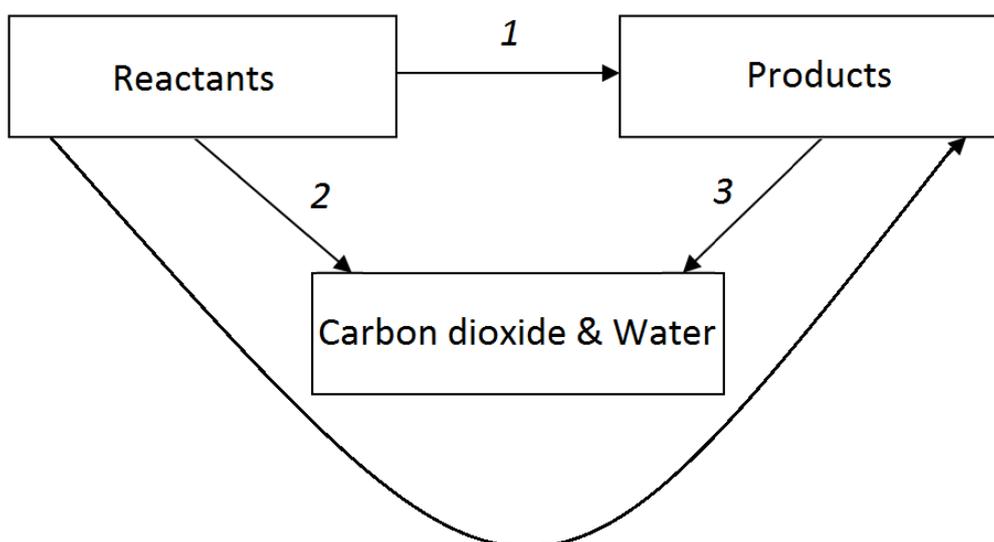


Compound	C_3H_6	C	H_2
Standard enthalpy of combustion, $\Delta H_{\text{C}}^{\ominus} \text{ kJ mol}^{-1}$	- 1928	-394	-242

$$\begin{aligned}\Delta H_{\text{reaction}} &= \Sigma \Delta H_{\text{C}} (\text{reactants}) - \Sigma \Delta H_{\text{C}} (\text{products}) \\ \Delta H_{\text{reaction}} &= [(3 \times -394) + (3 \times -242)] - [-1928] \\ &= -1182 + (-726) + 1928 \\ &= +20 \text{ kJ mol}^{-1}\end{aligned}$$

The cycle approach:

This approach is a visual representation of the equation approach. In the scheme below, reaction 1 is the $\Delta H_{\text{reaction}}$ whereas reactions 2 and 3 represent the $\Delta H_{\text{C}} (\text{reactants}) - \Delta H_{\text{C}} (\text{products})$ respectively.



- (i) The reaction that you are calculating ΔH for, moves from the reactants to the products boxes (reaction 1).
- (ii) The enthalpies of combustion are the reactions represented by arrows 2 and 3 as they are the energy changes **from** the reactants **to** carbon dioxide & water.
- (iii) Now imagine following the curved arrow from the reactants to the carbon dioxide & water and then on to the products. If you move **against** the arrow then **change** the sign of the ΔH_{C} , but if you're travelling **with** the direction as the arrow, then **keep** the sign.
- (iv) So, in the above case ① = ② – ③. This is exactly the same as saying $\Delta H_{\text{reaction}} = \Sigma \Delta H_{\text{C}} (\text{reactant}) - \Sigma \Delta H_{\text{C}} (\text{products})$

This approach gets you exactly the same answer for the *Worked Example 21*.

Calorimetric Calculations

The enthalpy changes of some reactions can be measured directly using a modification of a bomb calorimeter. The reaction under test is placed inside a polystyrene cup in order to minimise any heat losses to the surroundings which would render the results inaccurate. The polystyrene cup in turn is placed inside a lagged beaker to add further insulation. A thermometer can be used to measure the temperature change of the reaction and then an estimate of ΔH can be made using the formula below.

$$q = m c \Delta T$$

q = heat energy

m = mass of water heated

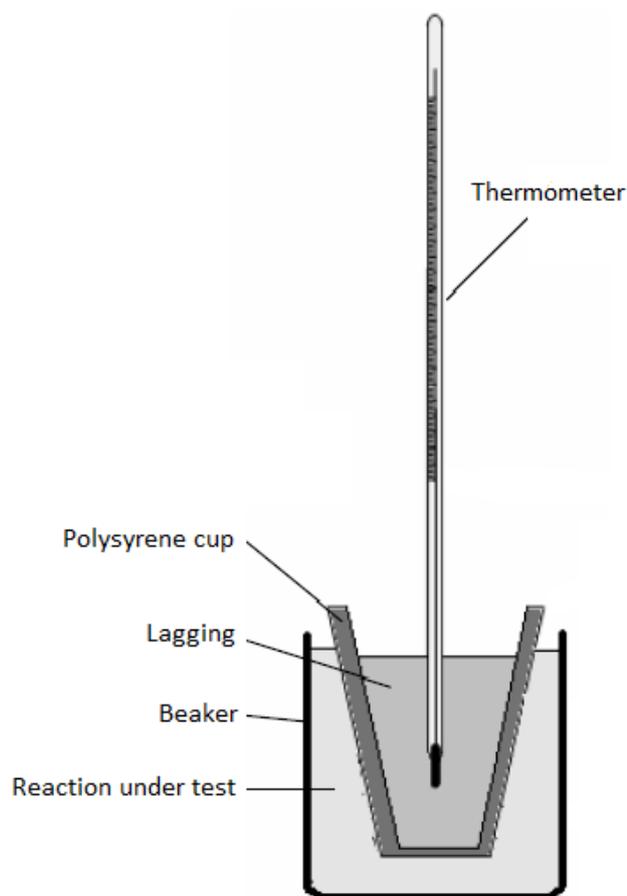
c = specific heat capacity

ΔT = change in temperature

There are three fundamentally different reactions that are investigated in this way;

- (i) displacement reactions,
- (ii) neutralisation reactions, and
- (iii) combustion reactions.

The heat energy change, Q can then be converted to a ΔH by *dividing* by the number of *moles* which caused the energy change.

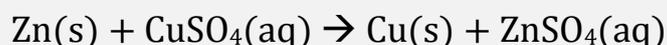


Each type of reaction, while ostensibly different, demands the same strategy:

1. Calculate q using the $q = mc\Delta T$ equation.
2. Convert q (in J) in Q (in kJ) by dividing by 1000.
3. Add a sign to Q . If the temperature goes down (endothermic) use a + **ve** sign or if the temperature goes up (exothermic) use a - **ve** sign.
4. Calculate the number of moles, n that reacted.
5. Calculate ΔH using $\Delta H = \frac{Q}{n}$

ΔH from displacement reactions

? Worked Example 22: Zinc and copper(ii) sulphate solution react according to the equation:



When 0.48 g of powdered zinc was added to 25 cm³ of 0.25 mol dm⁻³ copper sulphate solution the temperature rose from 21.0°C to 31.9°C. Calculate the ΔH of the reaction. State a suitable unit in your answer. Give your answer to 3 sig fig.

The specific heat capacity of water, $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$

$$q = mc\Delta T = 25 \times 4.18 \times (31.9 - 21) = 1139 \text{ J}$$

$$\text{Since the temperature increased, } Q = \frac{q}{1000} = -1.139 \text{ kJ}$$

Moles of Zn =

$$\frac{\text{mass}}{\text{RAM}} = \frac{0.48}{65.4} = 0.00734 \text{ mol}$$

Moles of CuSO₄ =

$$\frac{\text{conc} \times \text{vol}}{1000} = \frac{0.25 \times 25}{1000} = 0.00625 \text{ mol}$$

Since the mole ratio of Zn : CuSO₄ is 1:1, it follows that:

0.00625 mol of Zn reacted* to increase the temperature,

therefore $n = 0.00625 \text{ mol}$.

using $\Delta H = \frac{Q}{n}$, the enthalpy change of the reaction

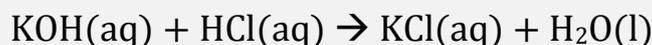
$$= -1.139 / 0.00625$$

$$= -182 \text{ kJ mol}^{-1} \text{ (3 sig fig)}$$

* The smaller of the two mole values is used as that is the limiting factor. The other reagent is present in a slight excess.

ΔH from neutralisation reactions

? **Worked Example 23:** Potassium hydroxide solution reacts with a solution of hydrochloric acid according to the equation:



25 cm³ of 1 mol dm⁻³ KOH and 25 cm³ of 1 mol dm⁻³ HCl were mixed in a polystyrene beaker, the temperature rose from 19°C to 25.5°C. Calculate the ΔH of the reaction. State a suitable unit in your answer which should be given to 2 significant figures.

The specific heat capacity of water, $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$

$$q = mc\Delta T = (25 + 25) \times 4.18 \times (25.5 - 19) = 1400.3 \text{ J}$$

$$\text{Since the temperature increased, } Q = \frac{q}{1000} = -1.400 \text{ kJ}$$

$$\begin{aligned} \text{Moles of KOH} &= \\ &= \frac{\text{conc} \times \text{vol}}{1000} = \frac{25 \times 1}{1000} = 0.025 \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{Moles of HCl} &= \\ &= \frac{\text{conc} \times \text{vol}}{1000} = \frac{25 \times 1}{1000} = 0.025 \text{ mol} \end{aligned}$$

Since the mole ratio of KOH : HCl 1:1, it follows that:

only 0.025 mol of KOH reacted to increase the temperature.

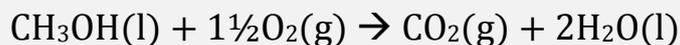
Therefore $n = 0.025 \text{ mol}$.

using $\Delta H = \frac{Q}{n}$, the enthalpy change of the reaction

$$= -1.400 / 0.025 = -56 \text{ kJ mol}^{-1} \text{ (2 sig fig)}$$

ΔH from combustion reactions

? Worked Example 24: When 0.7 g of methanol, CH_3OH was burned in a plentiful supply of oxygen, the heat released was used to increase the temperature of 220 cm^3 of water by 5.3°C :



Use this data to calculate the $\Delta H_{\text{c}}^\ominus$ of methanol.

1 cm^3 has a mass of 1 g.

The specific heat capacity of water, $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$

$$q = mc\Delta T = 220 * 4.18 * 5.3 = 4853 \text{ J}$$

$$\text{Since the temperature increased, } Q = \frac{q}{1000} = -4.853 \text{ kJ}$$

Moles of $\text{CH}_3\text{OH} =$

$$\frac{\text{mass}}{\text{RAM}} = \frac{0.7}{32} = 0.0219 \text{ mol}$$

Using:

$$\Delta H = \frac{Q}{n}$$

$$\Delta H = \frac{-4.853}{0.0219}$$

$$= -221 \text{ kJ mol}^{-1} \text{ (3 sig fig)}$$

* The mass, in g of water is the same as its volume, in cm^3 as its density = 1

- ☛ Students often mix up and then use the wrong masses in combustion calculations. You must use the **mass of water heated** up in the $q = mc\Delta T$ part.
- ☛ The **mass of fuel** burned is used to calculate **moles**.

Half-life (WJEC only)

Only STUDENTS DOING WJEC BOARD SHOULD READ THIS SECTION

Radioactive decay follows a defined mathematical pattern, regardless of which isotope it is. The key concept is the *half-life*, the time taken for half of the original number of radioactive atoms to decay. Even different isotopes of the same elements can have very different half lives, ranging from nanoseconds to millions of years! With each half-life that passes, the number of radioactive atoms remaining decays by half:

Half-life	Fraction of original particles
0	1
1	$\frac{1}{2}$
2	$\frac{1}{4}$
3	$\frac{1}{8}$
4	$\frac{1}{16}$

You are likely to be asked either to (i) determine the number of radioactive particles remaining after a specified time using a given half-life, (ii) calculate the time taken for a given fraction of the original radioactive particles to be formed, or (iii) when given the original and final numbers of particles and the time taken, you may have to determine the half-life. While there are more complex mathematical expressions that can be used to calculate these I personally favour the simple and very visual approaches...

Number of particles

? Worked Example 25: The half-life of an isotope of Caesium is 4.3 days. What proportion of radioactivity is remaining after 17.2 days?

*Construct the following table, automatically filling in the **half-life** and **fraction** columns.*

Then fill in the third column, time, using multiples of half-life (in this example 4.3, 8.6, 12.9 and so on.

When you reach the time specified in the question simply read the answer from the other column, so ...

Half-life	Fraction of original particles	Time (days)
0	1	0
1	$\frac{1}{2}$	4.3
2	$\frac{1}{4}$	8.6
3	$\frac{1}{8}$	12.9
4	$\frac{1}{16}$ ✓	17.2 ✓

Therefore after 17.2 days, $\frac{1}{16}$ of the original numbers would remain.

Time taken

? Worked Example 26: The half-life of an isotope of Iodine is 31.5 days. How long does it take for the amount of radioactivity in the sample to decay to $\frac{1}{8}$ of the original value.

*Construct the following table, automatically filling in the **half-life** and **time** columns, using multiples of the half-life given in the question.*

*Then fill in the **fraction** column starting from 1, $\frac{1}{2}$, $\frac{1}{4}$ and so on.*

When you reach the fraction specified in the question simply read the answer from the other column, so ...

Half-life	Fraction or original particles	Time (days)
0	1	0
1	$\frac{1}{2}$	31.5
2	$\frac{1}{4}$	63.0
3	$\frac{1}{8}$ ✓	94.5 ✓

Therefore it would take **94.5 days** to reach $\frac{1}{8}$ of the original radioactivity.

Half-life determination

? **Worked Example 27:** It took 40 years for 5.3 g of radioactive isotope ^{112}Rg to decrease to 1.325 g. Use this information to calculate the half-life of ^{112}Rg .

Construct a table like those in the previous examples.

*Fill in the **half-life** column again, but use **mass** instead of fraction in the second column.*

Keep halving the mass until you reach the mass that was specified in the question.

By reading across at the half-life column you can now see how many half-lives pass.

Half-lives	Mass (g)
0	5.3
1	2.65
2 <input checked="" type="checkbox"/>	1.325 <input checked="" type="checkbox"/>

$$\text{Half - life} = \frac{\text{total time}}{\text{number of half - lives}}$$

Therefore it would take **2** half-lives to reach 1.325 g.

$$\begin{aligned} \text{The half - life} &= \frac{40}{2} \\ &= \underline{\underline{20 \text{ years}}} \end{aligned}$$

End of Section 2 Test

Mole questions

Moles and number

Deduce the number of ...

1. hydrogen atoms in 3 mol of hydrogen gas
2. chlorine molecules in 0.45 mol of chlorine gas
3. chloride ions in 0.26 mol of calcium chloride, CaCl_2
4. atoms in 7.6 mol of methane, CH_4
5. hydrogen atoms in 0.5 mol of ammonia, NH_3
6. phosphorous atoms in 0.78 mol P_4O_{10}

Deduce the number of moles ...

7. of 4.78×10^{21} sodium atoms
8. of 5.12×10^{22} chlorine molecules
9. of 6.78×10^{22} chlorine atoms

Moles and mass

How many moles of sodium atoms are there in the following compounds?

10. sodium hydroxide, 5 g
11. sodium chloride, 10 g
12. sodium carbonate, 3.75 g
13. sodium phosphide, Na_3P , 0.154 g

What is the mass of ...?

14. chlorine gas, 20 mol
15. sulphuric acid, H_2SO_4 , 0.1 mol
16. hydrogen chloride, 0.15 mol
17. oxygen atoms, 2 mol
18. hydrogen peroxide, 2 mol
19. 4.2×10^{23} magnesium atoms

What is the mass of the following numbers of particles?

20. atoms of iron, 2.34×10^{23}
21. molecules of chlorine, 6.78×10^{21}
22. atoms of chromium, 4.19×10^{22}
23. molecules of carbon dioxide, 5
24. oxygen atoms in 5 carbon dioxide molecules

Moles and solution concentration and volume

Deduce the number of moles of acid in;

25. 75 cm³ of 0.030 mol dm⁻³ HCl solution
26. 25 cm³ of 0.025 mol dm⁻³ HI solution
27. 55 dm³ of 0.100 mol dm⁻³ H₂SO₄ solution
28. 1000 cm³ of 0.030 mol dm⁻³ H₃PO₄ solution
29. 880 dm³ of 0.100 mol dm⁻³ HCl solution

Assuming that all the hydrogen ions are released into solution, deduce the number of moles of hydrogen ions in;

30. 75 cm³ of 0.030 mol dm⁻³ HCl solution
31. 25 cm³ of 0.025 mol dm⁻³ HI solution
32. 55 dm³ of 0.100 mol dm⁻³ H₂SO₄ solution
33. 1000 cm³ of 0.03 mol dm⁻³ H₃PO₄ solution
34. 16 cm³ of 0.200 mol dm⁻³ H₂SO₄ solution
35. 16 dm³ of 0.200 mol dm⁻³ H₂SO₄ solution

What is the concentration of a solution containing ...?

36. 0.15 mol in 5 cm³ of solution
37. 0.015 mol in 5 dm³ of solution
38. 0.15 mol in 50 cm³ of solution
39. 0.15 mol in 500 cm³ of solution
40. 0.015 mol in 5 cm³ of solution
41. 0.0015 mol in 50 cm³ of solution

What volume of the following solutions contains 0.0027 mol of hydrogen ions?

42. 0.025 mol dm⁻³ HCl
43. 0.200 mol dm⁻³ H₂SO₄ solution
44. 0.150 mol dm⁻³ H₃PO₄ solution
45. 1.000 mol dm⁻³ H₂SO₄ solution
46. 1.000 mol dm⁻³ H₃PO₄ solution
47. 1.000 mol dm⁻³ HCl solution

Moles and the molar gas volume

Given that one mole of any gas occupies 24 dm³ at room temperature, deduce ...

48. The volume of 3.5 mol HCl(g) in cm³
49. The volume of 0.5 mol HCl(g) in dm³
50. The mass of 2000 cm³ methane
51. The mass of 2 dm³ methane

52. The number of moles in 98 cm³ methane
53. The number of moles in 98 dm³ methane
54. The volume of 3.56 × 10²⁴ hydrogen molecules

Moles and the ideal gas equation

Given the ideal gas equation, $PV = nRT$ and that $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ deduce;

55. The pressure, in kPa of 2 mol of hydrogen at 25°C in a vessel of 5 dm³.
56. The pressure, in Pa of 2 mol of hydrogen at 25°C in a vessel of 5 cm³.
57. The pressure, in MPa of 20 mol of hydrogen at 255°C in a vessel of 5 m³.
58. The pressure, in kPa of 20 mol of hydrogen at 25K in a vessel of 5 dm³.
59. The pressure, in Pa of 2 mol of hydrogen at 25°C in a vessel of 5 cm³.
60. The volume, in m³ of 2 mol of hydrogen at 25°C at a pressure of 5 kPa.
61. The volume, in dm³ of 0.2 mol of hydrogen at 225°C at a pressure of 5 kPa.
62. The volume, in dm³ of 0.2 mol of hydrogen at 25K at a pressure of 50 kPa.
63. The volume, in dm³ of 20 mol of hydrogen at 25°C at a pressure of 50 kPa.
64. The temperature, in °C when 0.5 mol of hydrogen are in a vessel of 3 dm³ at a pressure of 15 Pa.
65. The temperature, in K when 0.5 mol of hydrogen are in a vessel of 3 dm³ at a pressure of 15 MPa.
66. The temperature, in °C when 5 mol of hydrogen are in a vessel of 30 cm³ at a pressure of 15 Pa.
67. The temperature, in K when 0.5 mol of hydrogen are in a vessel of 0.3 dm³ at a pressure of 15 kPa.
68. The number of moles of hydrogen in a vessel of 10 dm³ at a pressure of Pa at 298K.
69. The number of moles of hydrogen in a vessel of 10 cm³ at a pressure of 10kPa at 278K.
70. The number of moles of hydrogen in a vessel of 100 cm³ at a pressure of 1,000Pa at 298°C.
71. The number of moles of hydrogen in a vessel of 10 m³ at a pressure of 10,000Pa at 298K.

Relative atomic mass and mass spectroscopy

72. A sample of a metal, M was composed of two isotopes, ^{58}M and ^{60}M with an abundance of 72% and 28% respectively.
- Use the data to calculate the relative atomic mass of M to one decimal place. Show your working.
 - Suggest an identity of M .
73. A sample of a non-metal extracted from a meteorite as found to have the following abundances:

m/z ratio	125	127	128
abundance	1.2	3.8	2.9

Use the data to calculate the relative atomic mass of the non-metal giving your answer to 4 significant figures. Show your working.

74. A sample of chromium with a relative atomic mass of 52.0 was found to contain 2 isotopes. In this sample, 75.0% of the atoms have a relative mass of 51.2. Calculate the relative mass of the other isotope.

Relative molecular and formula mass

75. Calculate the relative formula mass of the following salts and calculate the percentage of the Group 1 metal in each.
- $K_2S_2O_8$ $g\ mol^{-1}$ %
 - $Rb_2(NO_3)_2$ $g\ mol^{-1}$ %
76. A sample iron ore was found to contain a compound of iron that was 31.7% Fe and 27.4% S by mass, with the rest assumed to be oxygen. Use the information to determine the empirical formula of the compound.

77. A hydrocarbon was distilled from a batch of crude oil was found to contain 82.7% C by mass and have an M_r of 58 g mol^{-1} . Determine (i) the empirical **and** (ii) the molecular formulae of the compound.

Empirical formula =

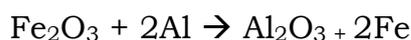
Molecular formula =

% atom economy and % yield

78. The second step of the extraction of titanium from its ore involves the reduction of titanium(iv) chloride with magnesium. The equation for the reaction is given below.

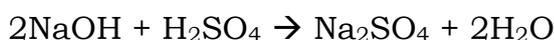


- (i) Calculate the % atom economy for the reaction.
 (ii) In an experimental batch extraction of titanium, 280 g of TiCl_4 produced 38 g of Ti. Calculate the % yield of this extraction.
79. In an experimental extraction of iron from iron(ii) oxide using Al, it was found that 2400 g of Fe_2O_3 formed a certain mass of iron. (i) Use the equation below to calculate the atom economy of the reaction. (ii) Given the % yield of the process was 80%, calculate the mass of iron produced.

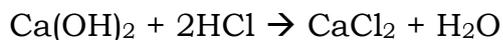


Titration and water of crystallisation

80. In a titration, a student found that 34.5 cm^3 of a 0.15 mol dm^{-3} sodium hydroxide solution were required to exactly neutralise 25 cm^3 of a sulphuric acid solution of unknown concentration. Calculate the concentration of the sulphuric acid solution. Show all your working



81. What volume of a 0.25 mol dm^{-3} solution of calcium hydroxide is required to exactly neutralise 25 cm^3 of a 0.1 mol dm^{-3} hydrochloric solution. Show all your working

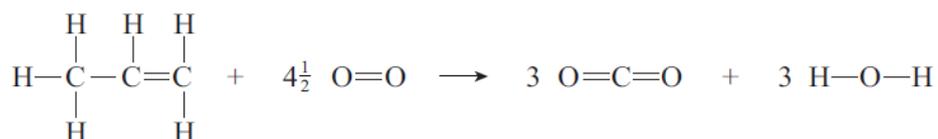


82. A 0.350 g sample of hydrated calcium chloride, $\text{CaCl}_2 \cdot x\text{H}_2\text{O}$ was heated until constant mass. Upon cooling in a desiccator, the mass of the anhydrous compound was 0.264 g . Use the data to calculate (i) a value for x and (ii) hence deduce the relative formula mass of the hydrated compound.

Energetics Questions

83. Use the mean bond enthalpy data from the table and the equation given below to calculate a value for the standard enthalpy of combustion of propene. All substances are in the gaseous state.

Bond	C=C	C-C	C-H	O=O	O=C	O-H
Mean bond enthalpy kJ mol^{-1}	612	348	412	496	743	463

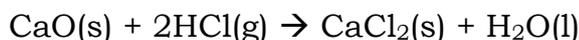


84. The mean bond enthalpy of $\text{H}-\text{H} = +436 \text{ kJ mol}^{-1}$. Use the following equation and data from the table in Q8 above to calculate a value for the $\text{C}-\text{H}$ bond enthalpy in ethane.

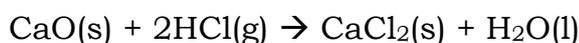


85. Use Hess's Law together with the data given in the table below, to calculate the standard enthalpy change for the following reaction.

	CaO(s)	HCl(g)	CaCl ₂ (s)	H ₂ O(l)
$\Delta H_f / \text{kJ mol}^{-1}$	-602	-92	-642	-286

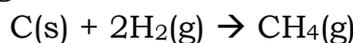


86. In an experiment, an excess of solid calcium oxide was added to 40 cm³ of 2.2 mol dm⁻³ hydrochloric acid. The initial temperature of the solution was 21°C. After reaction, the temperature had risen to 53°C. (The specific heat capacity of water is 4.2 J K⁻¹ g⁻¹)



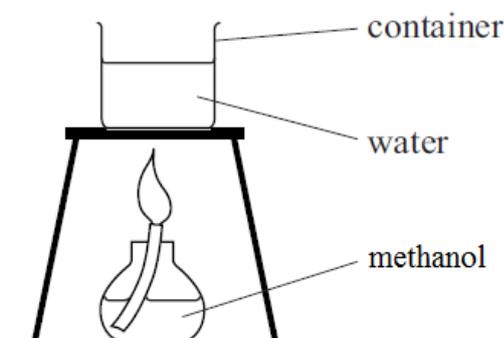
Use this information to calculate the enthalpy change for the reaction of one mole of calcium oxide with hydrochloric acid. For your calculation you should assume that all the heat from the reaction is used to raise the temperature of 40 g of water.

87. Using the enthalpy of combustion data given below, calculate the standard enthalpy change for the formation of methane.



$$\Delta H_C \text{ (kJ mol}^{-1}\text{)} \quad \text{C(s)} = -394 \quad \text{H}_2\text{(g)} = -286 \quad \text{CH}_4\text{(g)} = -890$$

88. An experiment was carried out to determine a value for the enthalpy of combustion of liquid methanol using the apparatus shown in the diagram.



Burning 3.5 g of methanol caused the temperature of 150 g of water to

rise by 50 °C. Use this information to calculate a value for the enthalpy of combustion of methanol, CH₃OH

(The specific heat capacity of water is 4.18 JK⁻¹ g⁻¹. Ignore the heat capacity of the container.)

Half-life questions

89. 5 g of a radioactive isotope of iodine, ¹³¹I has a half-life of 8 days. Calculate what mass of is present after 32 days.

90. Caesium-131 has a half-life of 4 years. What fraction of the original amount present is left after 16 years.

Section 3: Examination-style Questions

Ok so now you know how well you did on the End of Section tests it's time to do some real exam-style questions. The questions included here are a mix of questions that follow the styles asked by AQA, OCR, Edexcel and WJEC in the AS modules. As each board covers the same topics in a different order, you need to check which types of question are relevant for the paper you're sitting. The mark schemes include worked answers. I recommend you study them *after* you sit the questions to work out how best to improve.

Questions involving Moles

- Diamond is a rare allotrope of the element carbon. Diamond is one of the hardest materials known to man and is used in some kinds of industrial drilling equipment. A student decided to investigate the number of atoms in a diamond drill bit. He found that the mass of the drill bit was 0.124g.
 - Calculate the amount, in moles, of carbon atoms in the drill bit assuming that it is made of pure diamond. (1 mark)
 - Using the Avogadro constant, N_A , calculate the number of carbon atoms in the student's drill bit. (1 mark)
- The following questions are about chlorine, a gaseous element in Group 7 of the Periodic Table. It can dissolve in water to form a compound, chloric(i) acid which has a wide variety of uses in the bleaching industry.
 - Chloric(i) acid has the following percentage composition by mass:
Hydrogen, 1.90%; Chlorine, 67.62% and the remainder being oxygen.

Using this information, calculate the empirical formula of chloric(i) acid. Show all of your working. (2 marks)
 - Another compound of chlorine is chloric(v) acid, HClO_3 . Calculate the percentage by mass of chlorine in chloric(v) acid. (2 marks)
 - Chlorine also forms compound X with carbon and hydrogen which has the empirical formula, CHCl_2 . The relative molecular mass of X is 168 g mol^{-1} .

Use this information to determine the molecular formula of compound X. (1 mark)
- The Group 2 element magnesium, Mg is obtained as a silver coloured ribbon that reacts slowly with air to form magnesium oxide.
 - Magnesium has three stable isotopes, which are ^{24}Mg , ^{25}Mg and ^{26}Mg .
 - A sample of magnesium ribbon was composed of ^{24}Mg : 78.6%; ^{25}Mg : 10.1%; ^{26}Mg : 11.3%. Calculate the relative atomic mass of this sample of magnesium. Give your answer to **four** significant figures. (2 marks)

- (ii) Deduce the number of neutrons in the isotope ^{26}Mg . (1 mark)
- (b) Magnesium forms a hydrated sulphate, $\text{MgSO}_4 \cdot x\text{H}_2\text{O}$ that is sold as Epsom salts.

A sample of Epsom salts was heated to remove the water. The mass of anhydrous MgSO_4 remaining was 3.02 g and 3.14 g of water was driven off.

- (i) Calculate the amount, in moles, of anhydrous MgSO_4 formed. (2 marks)
- (ii) Calculate the amount, in moles, of water removed. (1 mark)
- (iii) Calculate a value for x in the formula $\text{MgSO}_4 \cdot x\text{H}_2\text{O}$. (1 mark)
4. Strontium carbonate, SrCO_3 reacts with dilute hydrochloric acid as shown in the equation below.

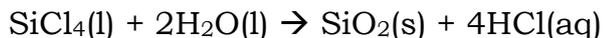


- (a) 6.25×10^{-3} moles of SrCO_3 reacts with $0.200 \text{ mol dm}^{-3}$ HCl.
- (i) Calculate the volume, in cm^3 , of $0.200 \text{ mol dm}^{-3}$ HCl required to react exactly with 6.25×10^{-3} moles of SrCO_3 . (2 marks)
- (ii) Calculate the volume, in cm^3 , of CO_2 formed at room temperature and pressure. (1 mark)
- (b) Calculate the % atom economy of the reaction assuming strontium chloride to be the desired product. (1 mark)
5. Barium is a Group 2 element that reacts vigorously with water according to the following equation:



- (a) A 0.33 g sample of pure barium was added to 250 cm^3 of water. Calculate the number of moles of barium added. (1 mark)
- (b) Calculate the volume, in dm^3 of hydrogen gas produced at room temperature and pressure. (1 mark)
- (c) Calculate the concentration, in mol dm^{-3} , of the barium hydroxide solution formed. (1 mark)

6. Silicon(IV) oxide (SiO_2 $M_r = 60.1$) is used as an abrasive in industry. The abrasive can be made as shown in the following equation:



- (a) Calculate the percentage atom economy for the formation of SiO_2 . (1 mark)
- (b) In an experiment 3.25 g of SiCl_4 were added to 500 cm^3 of water.
- (i) Calculate the amount, in moles, of SiCl_4 in 3.25 g. (2 marks)
- (ii) Calculate the maximum amount, in moles, of SiO_2 which could be formed in this experiment assuming that 500 cm^3 of water is an excess. (1 mark)
- (iii) Calculate the maximum mass of SiO_2 formed in this experiment. (1 mark)
- (iv) In this experiment only 0.49 g of SiO_2 were produced. Calculate the percentage yield of SiO_2 . (1 mark)
7. Magnesium can react with dilute nitric acids under suitable conditions. The equation for the reaction is given below:



A 0.0947 g sample of magnesium was added to 45.3 cm^3 of 0.275 mol dm^{-3} nitric acid solution. The acid was present in excess.

- (a) (i) Calculate the amount, in moles of magnesium in the 0.0947 g sample. (1 mark)
- (ii) Deduce the amount, in moles, of nitric acid needed to react completely with this sample of magnesium. (1 mark)
- (iii) Calculate the amount, in moles of the nitric acid that was originally added to this sample of magnesium. (1 mark)
- (iv) Deduce the amount, in moles of nitric acid that remained unreacted. (1 mark)
- (b) In a different experiment, 0.589 mol of hydrogen gas was produced when a different sample of magnesium was added to an excess of nitric acid. Calculate the volume that this gas would occupy at 245K and 120 kPa. Clearly state a unit in you answer. (The gas constant, $R = 8.31\text{JK}^{-1}\text{mol}^{-1}$) (3 marks)

- (c) When concentrated nitric acid reacts with magnesium, an oxide of nitrogen is produced which contains 30.4% by mass of nitrogen. Calculate the empirical formula of the oxide and show your working. (3 marks)

8. A carbonate of a Group 2 metal, MCO_3 reacts with hydrochloric acid according to the following equation:



A 0.387 g sample of MCO_3 underwent complete reaction with 10.49 cm^3 of a 0.500 mol dm^{-3} solution of hydrochloric acid.

- (a) (i) Calculate the amount, in moles of HCl which reacted with the 0.387 g MCO_3 . (1 mark)
- (ii) Deduce the amount, in moles, of MCO_3 in 0.387 g. (1 mark)
- (iii) Calculate the *relative formula mass* of MCO_3 . (1 mark)
- (b) Use your answer from part (a)(iii) to deduce the relative atomic mass of metal M and hence suggest its identity. (if you have been unable to calculate a value for the relative formula mass of MCO_3 you can assume it to be 100.1 but this is not the correct answer.) (2 marks)
9. When in a large excess, chlorine can undergo reaction with methane to form tetrachloromethane and hydrogen chloride. The equation for the reaction is given below.



- (a) Calculate the % atom economy for the formation of tetrachloromethane in this reaction. Give your answer to 3 significant figures. (1 mark)
- (b) In a different reaction the CCl_4 produced was found to be contaminated with another compound containing chlorine. This compound was found to contain 83.5% by mass of chlorine and 14.1% by mass of carbon. The remainder was assumed to be hydrogen.

Calculate the empirical formula of the contaminant. (3 marks)

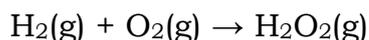
10. Cyclohexene, C_6H_{10} can be prepared by the dehydration of cyclohexanol. A student reacted 8.75 g of cyclohexanol, $C_6H_{12}O$, and obtained 0.0348 mol of cyclohexene.
- (a) What is the relative molecular mass of cyclohexene? (1 mark)
- (b) What is the percentage by mass of carbon in cyclohexene. (1 mark)
- (c) Calculate the percentage yield of cyclohexene. (1 mark)
11. What is the empirical formula of the liquid that contains 38.4 % carbon, 4.80 % hydrogen and 56.8 % chlorine by mass? (1 mark)
- A CH_3Cl
B C_2H_5Cl
C C_2H_3Cl
D $C_3H_5Cl_3$
12. Calcium oxide reacts with dilute hydrochloric acid as shown in the following equation: $CaO(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l)$
How many moles of Calcium oxide, CaO , are required to neutralize 40 cm^3 of 0.250 mol dm^{-3} hydrochloric acid, HCl ? (1 mark)
- A 0.020
B 0.0010
C 0.010
D 0.0050
13. Which of the following contains the greatest number of hydrogen atoms? (1 mark)
- A 1.5 moles of ammonia, C_2H_2
B 0.5 moles of methane, CH_4
C 1 mole of hydrogen gas, H_2
D 2 moles of water, H_2O
14. The following data were obtained from the mass spectrum of a sample of chromium.
- | Mass/charge ratio | % abundance |
|-------------------|-------------|
| 50.0 | 6.3 |
| 52.0 | 82.2 |
| 53.0 | 9.1 |
| 54.0 | 2.4 |
- Calculate the relative atomic mass of chromium in this sample, giving your answer to three significant figures. (1 mark)

15. Hydrogen peroxide decomposes on heating to form water and oxygen:



What mass of hydrogen peroxide is required to give 8 g of oxygen gas? (1 mark)

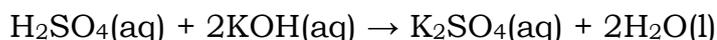
16. In the presence of a chromium catalyst, hydrogen and oxygen react to form hydrogen peroxide.



If all volumes are measured at 270°C and one atmosphere pressure, calculate the volume of gaseous hydrogen peroxide formed on reaction of 75 cm³ of hydrogen react completely with 55 cm³ of oxygen. (1 mark)

17. Calculate the mass of calcium hydroxide, Ca(OH)₂, present in 100 cm³ of a 0.300 mol dm⁻³ solution assuming the molar mass of Ca(OH)₂ is 74.0 g mol⁻¹. (1 mark)

18. Sulphuric acid, H₂SO₄, reacts with sodium hydroxide, KOH, as shown below.

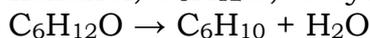


0.0100 mol of sulphuric acid (34 cm³) is neutralized with dilute potassium hydroxide, concentration 0.400 mol dm⁻³. Calculate the volume of potassium hydroxide required. (2 mark)

19. Deduce the empirical formula of a compound that was analysed and found to contain 1.45 g carbon 0.482 g hydrogen 1.69 g nitrogen. (2 marks)

[Relative atomic masses: C = 12; H = 1; N = 14]

20. The dehydration of cyclohexanol, C₆H₁₂O, to cyclohexene, C₆H₁₀ is:

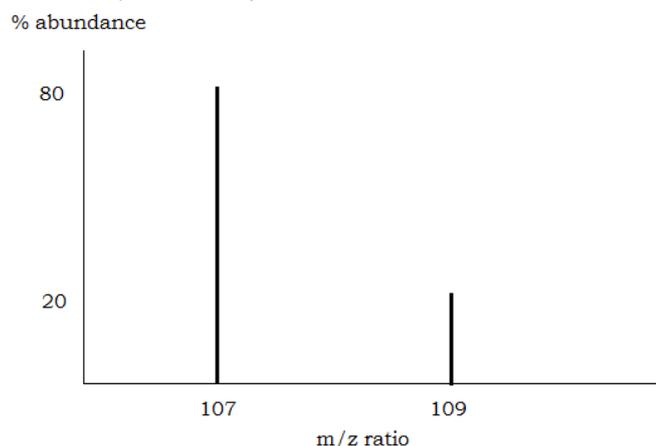


86.0 g of cyclohexanol produced 42.8 g of cyclohexene.

[Molar masses / g mol⁻¹ : cyclohexanol = 100; cyclohexene = 82]

- (a) Calculate the percentage yield of cyclohexene. (1 mark)
- (b) Calculate the percentage atom economy for the formation of cyclohexene from cyclohexanol. (1 mark)

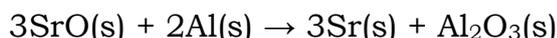
21. Using the mass spectrum below, calculate the relative atomic mass of this sample of silver. (2 marks)



22. Which of the following gas samples occupies the smallest volume at the same temperature and pressure?
[Relative atomic masses: H = 1; C = 12; O = 16; F = 19; Ne = 20]
A 1 gram of methane, CH₄
B 1 gram of ozone, O₃
C 1 gram of hydrogen fluoride, HF
D 1 gram of neon
23. State which one of the following gases contains the greatest number of molecules. (1 mark)
A 8 g of hydrogen
B 8.8 g of carbon dioxide
C 16 g of oxygen
D 28 g of nitrogen
24. When calcium carbonate is heated strongly in air, calcium oxide is made.
 $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$
Calculate the maximum mass of calcium oxide formed when 3.400 mole of pure calcium carbonate is heated. (2 marks)
25. What mass of boron contains the same number of atoms as there are molecules in 16 g nitrogen, N₂. (1 mark)
26. Calculate the empirical formula of lead carbonate which has the following composition, Pb 77.5%; C 4.50%; O 18.0% by mass. (3 marks)
27. Calculate the molar mass of Sn₃O₄. (1 mark)
28. Calculate the percentage by mass of nitrogen in ammonium sulfate, (NH₄)₂SO₄. (2 marks)

Questions involving Energy

1. Strontium is a reactive Group 2 metal that can be extracted from strontium oxide (SrO) using aluminium. Powdered strontium oxide and powdered aluminium is heated strongly. The equation for this extraction process is shown below.



Some standard enthalpies of formation are given in the table below.

Substance	SrO(s)	Al ₂ O ₃ (s)
$\Delta H / \text{kJ mol}^{-1}$	-538	-1669

Use the information provided to calculate the ΔH of the extraction reaction. (3 marks)

2. Some students devised an experiment which they believed would enable them to investigate the strength of the intermolecular forces between butyl ethanoate molecules (CH₃COOCH₂CH₂CH₂CH₃) and trichloromethane molecules (CHCl₃). They mixed 0.10 mol of each of the two liquids in a calorimeter and recorded the following results. The starting temperature of both liquids was the same.

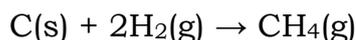
Mass of 0.10 mol of butyl ethanoate	11.6 g
Mass of 0.10 mol of trichloromethane	11.95 g
Increase in temperature (ΔT) on mixing	9.5 K

- (a) Write an expression for the heat change (q) which relates mass (m), specific heat capacity (c) and change in temperature (ΔT). (1 mark)
- (b) Calculate the amount of heat required to increase the temperature of 11.60 g of butyl ethanoate by 9.5 K during the mixing process. (You should assume that c for ethyl ethanoate = $1.92 \text{ J g}^{-1}\text{K}^{-1}$) (1 mark)
- (c) Calculate the amount of heat required to increase the temperature of 11.95 g of trichloromethane by 9.5 K during the mixing process. (You should assume that c for trichloromethane = $0.96 \text{ J g}^{-1}\text{K}^{-1}$) (1 mark)
- (d) Using the values from parts (b) and (c), calculate the molar enthalpy change in kJ mol^{-1} for the mixing process. (2 marks)

3. A student carried out an experiment to determine the enthalpy change of combustion of hexan-1-ol, $\text{CH}_3(\text{CH}_2)_5\text{OH}$.
In the experiment, 1.67 g of hexan-1-ol was burnt. The energy was used to heat 200 cm^3 of water from $23.0 \text{ }^\circ\text{C}$ to $88.0 \text{ }^\circ\text{C}$.
- (a) Calculate the energy released, in kJ, during combustion of 1.67 g hexan-1-ol. (1 mark)
The specific heat capacity of water = $4.18 \text{ J g}^{-1} \text{ K}^{-1}$.
Density of water = 1.00 g cm^{-3} .
- (b) Calculate the amount, in moles, of hexan-1-ol that was burnt. (2 marks)
- (c) Calculate the enthalpy change of combustion of hexan-1-ol. Give your answer to three significant figures. (3 marks)
4. The standard enthalpy change of formation of pentane can be determined indirectly.
Calculate the standard enthalpy change of formation of pentane using the standard enthalpy changes of combustion below. (3 marks)

Substance	$\Delta H_c / \text{kJ mol}^{-1}$
C	-394
H_2	-286
C_5H_{12}	-3578

5. Carbon and hydrogen can react according to the equation below to form methane.



The enthalpy changes of combustion of carbon, hydrogen and methane are shown below.

Substance	Enthalpy change of combustion / kJ mol^{-1}
carbon, C(s)	-394
hydrogen, $\text{H}_2(\text{g})$	-286
methane, $\text{CH}_4(\text{g})$	-891

Which one of the following expressions gives the correct value for the enthalpy change of formation of methane in kJ mol^{-1} ?

- A $394 + 286 - 891$
 B $-394 - (2 \times 286) + 891$
 C $-394 - 286 + 891$
 D $394 + (2 \times 286) - 891$

6. Given the following data:

$$\Delta H_f [\text{MnO}(\text{s})] = -240 \text{ kJ mol}^{-1}$$

$$\Delta H_f [\text{Mn}_2\text{O}_3(\text{s})] = -930 \text{ kJ mol}^{-1}$$

select the expression which gives the enthalpy change, in kJ mol^{-1} , for the reaction: $2\text{MnO}(\text{s}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{Mn}_2\text{O}_3(\text{s})$

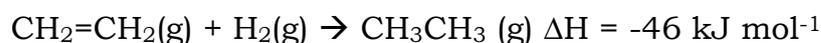
A $+930 - (240 \times 2)$

B $-930 + (240 \times 2)$

C $(-930 \times \frac{1}{2}) + 240$

D $(+930 \times \frac{1}{2}) - 240$

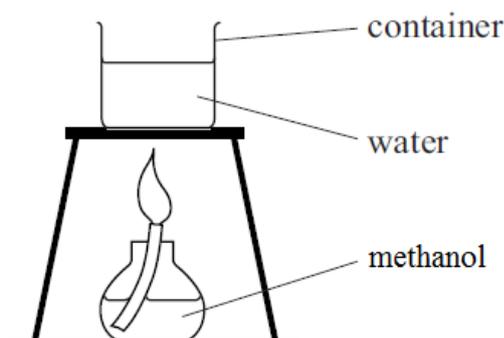
7. Ethane may be manufactured from ethene and hydrogen by the direct hydrogenation reaction:



Use the standard enthalpy change for the reaction and the standard enthalpy change of formation value, ΔH_f below to calculate a value for the enthalpy change of formation of gaseous ethane. (3 marks)

Compound	$\Delta H_f / \text{kJ mol}^{-1}$
$\text{CH}_2=\text{CH}_2(\text{g})$	+52.3

8. In order to find the enthalpy of combustion of ethanol, David set up the apparatus as shown below.



Here are his results:

Mass of ethanol before the experiment / g	15.9
Mass of ethanol after the experiment / g	12.9
Mass of water / g	250
Temperature of water before the experiment / $^{\circ}\text{C}$	19.2
Temperature of water after the experiment / $^{\circ}\text{C}$	60.4

To find the enthalpy of combustion he used the following equation he found in a text book:

$$\Delta H_c = \frac{-m c \Delta T}{n}$$

Where m is the mass of water heated, c is the specific heat capacity of water ($4.18 \text{ J g}^{-1} \text{ mol}^{-1}$), ΔT is the temperature rise and n is the number of moles of ethanol burned. He calculated the M_r of ethanol to be 46 g mol^{-1} .

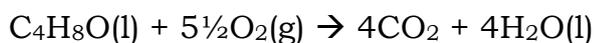
- State why the mass of ethanol is measured both before and after the experiment. (1 mark)
 - Calculate the number of moles of ethanol burned in the experiment. (1 mark)
 - Calculate a value for the enthalpy change of combustion (ΔH_c) of ethanol, in kJ mol^{-1} . (2 marks)
9. The mean bond enthalpy values of some elements are given in the table below.

Bond	Mean bond enthalpy / kJ mol^{-1}
H – H	+ 436
I – I	+ 151
H - I	+ 299

What is the enthalpy change for the reaction shown below in kJ mol^{-1} ?



10. Butanone, $\text{C}_4\text{H}_8\text{O}$ undergoes complete combustion forming carbon dioxide and water.



In an experiment to calculate the enthalpy change of combustion for butanone, 3.2 g of butanone was burned completely in oxygen. The heat energy from this combustion raised the temperature of 300 g of water from $20.0 \text{ }^\circ\text{C}$ to $81.4 \text{ }^\circ\text{C}$. The specific heat capacity of water is $4.18 \text{ J g}^{-1} \text{ mol}^{-1}$

- Calculate the number of moles of butanone present in the 3.20 g.
- Using the expression energy transferred (J) = mass x specific heat capacity x temperature change, calculate the heat energy transferred to raise the temperature of 300 g of water from $20.0 \text{ }^\circ\text{C}$ to $81.4 \text{ }^\circ\text{C}$. (1 mark)

(c) Use your answers to (a) and (b) to calculate a value for the enthalpy change of combustion of butanone. Give your answer to three significant figures, include a sign and unit. (3 marks)

11. Use the standard enthalpy changes of combustion, ΔH_c of given in the table below to find the standard enthalpy change of formation for propanoic acid, $\text{CH}_3\text{CH}_2\text{COOH}$, in kJ mol^{-1} . Give your answer to 3 sig fig.

Substance	$\Delta H_c / \text{kJ mol}^{-1}$
C(s)	- 394
H ₂ (g)	- 286
CH ₃ CH ₂ COOH(l)	- 870

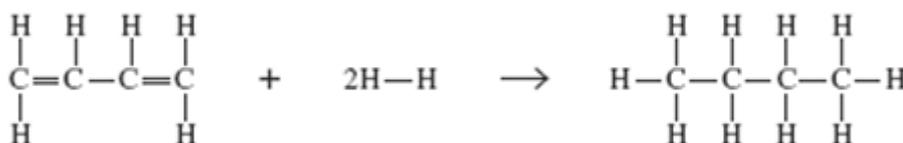


12. Use the equation below, and the mean bond enthalpy data given in the table to calculate a value for the enthalpy change of the following reaction. (3 marks)



Bond	Mean bond enthalpy / kJ mol^{-1}
C - H	413
C - Cl	346
Cl - Cl	243
H - Cl	432

13. Buta-1,3-diene can be hydrogenated to form butane.



The standard enthalpy change for the reaction is -240 kJ mol^{-1} . Use this and the data below to calculate a value for the C = C bond enthalpy. (3 marks)

Bond	Mean bond enthalpy / kJ mol^{-1}
C - H	412
C - C	348
H - H	436

Questions involving half-life

1. The half life of cobalt, ^{60}Co , used in radiotherapy, is 5.1 years. Calculate how long it will take for the activity of the isotope to decay to $\frac{1}{4}$ of its original activity. (1 mark)
2. An isotope of magnesium, ^{27}Mg , is used to detect leaks in water pipes. It decays by β -emission with a half life of 9.3 minutes. Calculate how long it will take for the activity of the isotope to decay to one eighth of its original activity. (1 mark)
3. A radioactive isotope of cobalt has a half-life of 74 days. Starting with 64 g, calculate the mass of this isotope remaining after 222 days. (1 mark)

Section 4: Answers to the Questions

Marking your answers is not just about finding whether you got the questions right or wrong! Thoroughness and careful scrutiny of the mark schemes, going back to work out what went wrong are the hallmarks of the A-graders. If you wannabe an A-grader then you'd better act like it. Mark the questions in turn and go back and see what went wrong. Identify which questions you tend to make common errors on, and know what those common errors are. That way, you'll be aware of them and are less likely to make them!

End of Section 1 Test Answers

Fractions, decimals, ratios and percentages

1. Express the following fractions as (i) percentages, (ii) decimals and (iii) standard form. Give all your answers to *three* significant figures:

a. $\frac{5}{81}$ (i) **6.17 %** (ii) **0.0617** (iii) **6.17×10^{-2}**

b. $\frac{4}{19}$ (i) **21.1 %** (ii) **0.211** (iii) **2.11×10^{-1}**

c. $\frac{1}{125}$ (i) **0.800 %** (ii) **0.008** (iii) **8.00×10^{-3}**

2. In a recent college chemistry test it was shown that 18 of out every 57 students gained a grade B or above.

- a. Express this information as (i) a percentage, (ii) a decimal and (iii) in standard form. Give all your answers to three significant figures.

(i) **31.6 %** (ii) **0.316** (iii) **3.16×10^{-1}**

- b. Another teacher is setting the same test. How many students would he expect to achieve a grade B or above from a class of 36?

11

- c. Using the same test a third teacher found that four students gained grade B or above. How many students were in her class?

13

Score _____ / 14 [_____ %]

Powers and logs

3. Calculate the following:

a. $4^5 = \mathbf{1024}$ b. $2^3 = \mathbf{8}$ c. $3^7 = \mathbf{2187}$ d. $4^2 = \mathbf{16}$

4. Determine the *logarithms* of the following *numbers*:

a. 59 **1.8** (1dp)
 b. 7892 **3.9** (2 sig fig)
 c. 4.58×10^{-6} - **5.34** (3 sig fig)
 d. 4^2 **1.204** (4 sig fig)
 e. 1.111 **0.046** (2 sig fig)

5. Determine the *numbers* of the following *logarithms*. Give your answers to the appropriate precision.

a. 59 **1.0×10^{59}** (2 sig fig)
 b. -0.7892 **0.1625** (4 sig fig)
 c. 4.58×10^{-6} **1** (1 sig fig)
 d. $4^{0.3}$ **30** (1 sig fig)
 e. 10.8 **6.31×10^{10}** (standard form)

Score _____ / **14** [_____ %]

Rearranging equations

6. By rearranging the equation $A \times B \times C = D \times E \times F$, make each of the following the subject of the equation:

$$(i) \quad A = \frac{D \times E \times F}{B \times C}$$

$$(ii) \quad C = \frac{D \times E \times F}{A \times B}$$

$$(iii) \quad D = \frac{A \times B \times C}{E \times F}$$

$$(iv) \quad E = \frac{A \times B \times C}{D \times F}$$

7. Rearrange the following equation

$$A = \frac{B \times C}{D \times E}$$

$$(i) \quad B = \frac{A \times D \times E}{C}$$

$$(ii) \quad C = \frac{A \times D \times E}{B}$$

$$(iii) \quad D = \frac{B \times C}{A \times E}$$

$$(iv) \quad E = \frac{B \times C}{A \times D}$$

8. By rearranging the equation $A + B - C = D + E - F$, make each of the following the subject of the equation:

$$(i) \quad A = D + E - F - B + C$$

$$(ii) \quad B = D + E - F - A + C$$

$$(iii) \quad C = A + B - D - E + F$$

$$(iv) \quad E = A + B - C - D + F$$

$$(v) \quad F = D + E - A - B + C$$

9. Rearrange the following equation

$$A = \frac{B + C}{1000}$$

(i) $B = (1000 A) - C$

(ii) $C = (1000 A) - B$

(iii) $1000 = \frac{B+C}{A}$

Score _____ / 16 [_____ %]

Unit conversions

10. Express the following *masses* as the specified units:

a. 10 g = **0.01 kg** c. 145 kg = **145 000 g**
 b. 802 mg = **8.02x10⁻⁴ kg** d. 14 mg = **0.014 g**

11. Express the following *volumes* as the units specified in the questions:

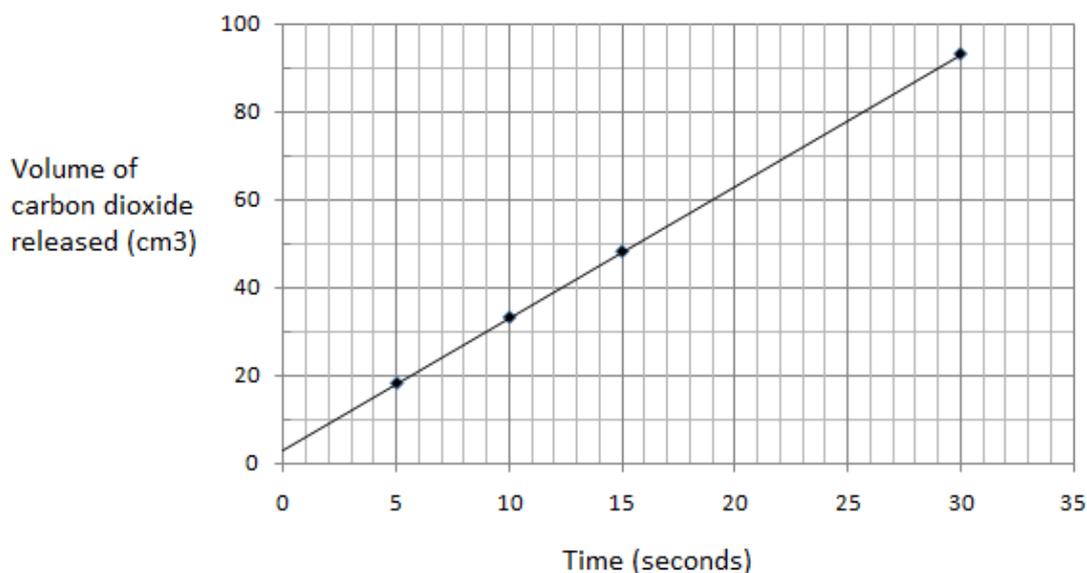
a. 100 cm³ = **0.1 dm³** c. 24 dm³ = **24 000 cm³**
 b. 802 m³ = **8.02x10⁶ cm³** d. 0.01 dm³ = **1x10⁻⁵ m³**

Score _____ / 8 [_____ %]

Gradient, intercept and rate

12. Plot a graph of the following data and use it to deduce the missing volumes, and calculate the gradient and intercept of the line [6 marks].

Time (seconds)	Volume of CO ₂ released (cm ³)
5	12
10	22
15	32
20	42 [1 mark]
25	52 [1mark]
30	62



- [1 mark] correctly labelled x-axis [1 mark] correctly labelled y-axis
 [1 mark] correctly plotted points [1 mark] line placed through data
 [1 mark] gradient = 2 [1 mark] intercept = 2

Score _____ / 8 [_____ %]

Grand total of _____ / 60 = _____ %

End of Section 2 Test Answers

Mole questions

Moles and number

Deduce the number of ...

- | | |
|---|---|
| 1. hydrogen atoms in 3 mol of hydrogen gas | 3.60×10^{24} |
| 2. chlorine molecules in 0.45 mol of chlorine gas | 2.70×10^{23} |
| 3. chloride ions in 0.26 mol of calcium chloride, CaCl_2 | 3.12×10^{23} |
| 4. atoms in 7.6 mol of methane, CH_4 | 2.28×10^{25} |
| 5. hydrogen atoms in 0.5 mol of ammonia, NH_3 | 9.00×10^{23} |
| 6. Phosphorous atoms in 0.78 mol P_4O_{10} | 1.87×10^{24} |

Deduce the number of moles ...

- | | |
|--|---|
| 7. of 4.78×10^{21} sodium atoms | 7.97×10^{-3} |
| 8. of 5.12×10^{22} chlorine molecules | 0.085 |
| 9. of 6.78×10^{22} chlorine atoms | 0.113 |

Moles and mass

How many moles of sodium atoms are there in the following compounds?

- | | |
|---|---|
| 10. sodium hydroxide, 5 g | 0.125 |
| 11. sodium chloride, 10 g | 0.171 |
| 12. sodium carbonate, 3.75 g | 0.0708 |
| 13. sodium phosphide, Na_3P , 0.154 g | 4.62×10^{-3} |

What is the mass of ...?

- | | |
|---|---------------|
| 14. chlorine gas, 20 mol | 1420 g |
| 15. sulphuric acid, H_2SO_4 , 0.1 mol | 9.81 g |
| 16. hydrogen chloride, 0.15 mol | 5.48 g |
| 17. oxygen atoms, 2 mol | 32.0 g |
| 18. hydrogen peroxide, 2 mol | 68.0 g |
| 19. 4.2×10^{23} magnesium atoms | 17.0 g |

What is the mass of the following numbers of particles?

- | | |
|--|--|
| 20. atoms of iron, 2.34×10^{23} | 21.8 g |
| 21. molecules of chlorine, 6.78×10^{21} | 0.802 g |
| 22. atoms of chromium, 4.19×10^{22} | 3.63 g |
| 23. molecules of carbon dioxide, 5 | 3.67×10^{-22} g |
| 24. oxygen atoms in 5 carbon dioxide molecules | 2.67×10^{-22} g |

Moles and solution concentration and volume

Deduce the number of moles of acid in;

- | | | |
|-----|--|-----------------------------|
| 25. | 75 cm ³ of 0.030 mol dm ⁻³ HCl solution | 2.25x10⁻³ |
| 26. | 25 cm ³ of 0.025 mol dm ⁻³ HI solution | 6.25x10⁻⁴ |
| 27. | 55 dm ³ of 0.100 mol dm ⁻³ H ₂ SO ₄ solution | 5.5 |
| 28. | 1000 cm ³ of 0.030 mol dm ⁻³ H ₃ PO ₄ solution | 0.03 |
| 29. | 880 dm ³ of 0.100 mol dm ⁻³ HCl solution | 88 |

Assuming that all the hydrogen ions are released into solution, deduce the number of moles of hydrogen ions in;

- | | | |
|-----|---|-----------------------------|
| 30. | 75 cm ³ of 0.030 mol dm ⁻³ HCl solution | 2.25x10⁻³ |
| 31. | 25 cm ³ of 0.025 mol dm ⁻³ HI solution | 6.25x10⁻⁴ |
| 32. | 55 dm ³ of 0.100 mol dm ⁻³ H ₂ SO ₄ solution | 11 |
| 33. | 1000 cm ³ of 0.03 mol dm ⁻³ H ₃ PO ₄ solution | 0.09 |
| 34. | 16 cm ³ of 0.200 mol dm ⁻³ H ₂ SO ₄ solution | 6.40x10⁻³ |
| 35. | 16 dm ³ of 0.200 mol dm ⁻³ H ₂ SO ₄ solution | 6.4 |

What is the concentration of a solution containing ...?

- | | | |
|-----|--|--------------------------|
| 36. | 0.15 mol in 5 cm ³ of solution | 30 |
| 37. | 0.015 mol in 5 dm ³ of solution | 3x10⁻³ |
| 38. | 0.15 mol in 50 cm ³ of solution | 3 |
| 39. | 0.15 mol in 500 cm ³ of solution | 0.3 |
| 40. | 0.015 mol in 5 cm ³ of solution | 3 |
| 41. | 0.0015 mol in 50 cm ³ of solution | 0.03 |

What volume of the following solutions contains 0.0027 mol of hydrogen ions?

- | | | |
|-----|--|----------------------------|
| 42. | 0.025 mol dm ⁻³ HCl | 108 cm³ |
| 43. | 0.200 mol dm ⁻³ H ₂ SO ₄ solution | 6.75 cm³ |
| 44. | 0.150 mol dm ⁻³ H ₃ PO ₄ solution | 6 cm³ |
| 45. | 1.000 mol dm ⁻³ H ₂ SO ₄ solution | 1.35 cm³ |
| 46. | 1.000 mol dm ⁻³ H ₃ PO ₄ solution | 0.90 cm³ |
| 47. | 1.000 mol dm ⁻³ HCl solution | 2.70 cm³ |

Moles and the molar gas volume

Given that one mole of any gas occupies 24 dm³ at room temperature, deduce ...

- | | | |
|-----|---|------------------------------|
| 48. | The volume of 3.5 mol HCl(g) in cm ³ | 84 000 cm³ |
| 49. | The volume of 0.5 mol HCl(g) in dm ³ | 12 dm³ |
| 50. | The mass of 2000 cm ³ methane | 1.33 g |
| 51. | The mass of 2 dm ³ methane | 1.33 g |

- | | |
|--|-----------------------------|
| 52. The number of moles in 98 cm ³ methane | 4.08x10⁻³ |
| 53. The number of moles in 98 dm ³ methane | 4.08 |
| 54. The volume of 3.56 x 10 ²⁴ hydrogen molecules | 142.4 dm³ |

Moles and the ideal gas equation

Given the ideal gas equation, $PV = nRT$ and that $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ deduce;

- | | |
|--|-------------------------------|
| 55. The pressure, in kPa of 2 mol of hydrogen at 25°C in a vessel of 5 dm ³ . | 991 kPa |
| 56. The pressure, in Pa of 2 mol of hydrogen at 25°C in a vessel of 5 cm ³ . | 9.91x10⁸Pa |
| 57. The pressure, in MPa of 20 mol of hydrogen at 255°C in a vessel of 5 m ³ . | 0.0175 MPa |
| 58. The pressure, in kPa of 20 mol of hydrogen at 25K in a vessel of 5 dm ³ . | 831 kPa |
| 59. The pressure, in Pa of 2 mol of hydrogen at 25°C in a vessel of 5 cm ³ . | 9.91x10⁸ Pa |
| 60. The volume, in m ³ of 2 mol of hydrogen at 25°C at a pressure of 5 kPa. | 0.991 m³ |
| 61. The volume, in dm ³ of 0.2 mol of hydrogen at 225°C at a pressure of 5 kPa. | 166 dm³ |
| 62. The volume, in dm ³ of 0.2 mol of hydrogen at 25K at a pressure of 50 kPa. | 0.831 dm³ |
| 63. The volume, in dm ³ of 20 mol of hydrogen at 25°C at a pressure of 50 kPa. | 991 dm³ |
| 64. The temperature, in °C when 0.5 mol of hydrogen are in a vessel of 3 dm ³ at a pressure of 15 Pa. | 273°C |
| 65. The temperature, in K when 0.5 mol of hydrogen are in a vessel of 3 dm ³ at a pressure of 15 MPa. | 10.8 K |
| 66. The temperature, in °C when 5 mol of hydrogen are in a vessel of 30 cm ³ at a pressure of 15 Pa. | 273°C |
| 67. The temperature, in K when 0.5 mol of hydrogen are in a vessel of 0.3 dm ³ at a pressure of 15 kPa. | 1.083 K |
| 68. The number of moles of hydrogen in a vessel of 10 dm ³ at a pressure of 10 Pa at 298K. | 4.04x10⁻⁵ |
| 69. The number of moles of hydrogen in a vessel of 10 cm ³ at a pressure of 10kPa at 278K. | 4.32x10⁻⁵ |
| 70. The number of moles of hydrogen in a vessel of 100 cm ³ at a pressure of 1,000Pa at 298°C. | 2.11x10⁻⁵ |
| 71. The number of moles of hydrogen in a vessel of 10 m ³ at a pressure of 10,000Pa at 298K. | 40.4 |

Relative atomic mass and mass spectroscopy

72. A sample of a metal, M was composed of two isotopes, ^{58}M and ^{60}M with an abundance of 72% and 28% respectively.
- a. Use the data to calculate the relative atomic mass of M to one decimal place. Show your working.

$$\text{RAM} = \frac{(72 \times 58) + (28 \times 60)}{100} = \mathbf{58.6 \text{ (1dp)}}$$

- b. Suggest an identity of M .

Nickel

73. A sample of a non-metal extracted from a meteorite as found to have the following abundances:

m/z ratio	125	127	128
abundance	1.2	3.8	2.9

Use the data to calculate the relative atomic mass of the non-metal giving your answer to 4 significant figures. Show your working.

$$\text{RAM} = \frac{(125 \times 1.2) + (127 \times 3.8) + (128 \times 2.9)}{1.2 + 3.8 + 2.9} = \mathbf{127.1 \text{ (4 sig fig)}}$$

74. A sample of chromium with a relative atomic mass of 52.0 was found to contain 2 isotopes. In this sample, 75.0% of the atoms have a relative mass of 51.2. Calculate the relative mass of the other isotope.

$$\mathbf{52} = \frac{(51.2 \times 75) + (\text{relative mass} \times 25)}{100} = \mathbf{54.4 \text{ (3 sig fig)}}$$

Relative molecular and formula mass

75. Calculate the relative formula mass of the following salts and calculate the percentage of the Group 1 metal in each.

a.	$K_2S_2O_8$	270.2	$g \text{ mol}^{-1}$	28.9	%
b.	$Rb_2(NO_3)_2$	295	$g \text{ mol}^{-1}$	58.0	%

76. A sample iron ore was found to contain a compound of iron that was 31.7% Fe and 27.4% S by mass, with the rest assumed to be oxygen. Use the information to determine the empirical formula of the compound.

Fe 31.7	S 27.4	O 40.9
0.568	0.854	2.55
1	1.5	4.5
2	3	9



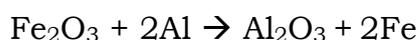
77. A hydrocarbon was distilled from a batch of crude oil was found to contain 82.7% C by mass and have an M_r of 58 g mol^{-1} . Determine (i) the empirical **and** (ii) the molecular formulae of the compound.

Empirical formula = C₂H₅

Molecular formula = C₄H₁₀

% atom economy and % yield

79. In an experimental extraction of iron from iron(III) oxide using Al, it was found that 2400 g of Fe_2O_3 formed a certain mass of iron. (i) Use the equation below to calculate the atom economy of the reaction. (ii) Given the % yield of the process was 80%, calculate the mass of iron produced.



$$\% \text{ atom economy} = \frac{55.8 + 55.8}{55.8 + 55.8 + 27 + 27 + 16 + 16 + 16} \times 100 = \mathbf{52.2\%}$$

$$\text{Moles of } \text{Fe}_2\text{O}_3 \text{ used} = 2400 / (55.8 + 55.8 + 16 + 16 + 16) = 15.04$$

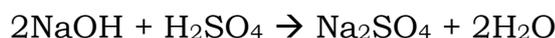
$$\text{Theoretical moles of iron} = 2 \times n\text{Fe}_2\text{O}_3 = 30.08$$

$$\text{Theoretical mass of iron} = 30.08 \times 55.8 = 1677 \text{ g}$$

$$80\% \text{ yield, mass (g)} = 0.8 \times 1677 = \mathbf{1341 \text{ g}}$$

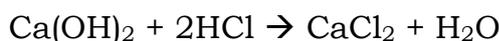
Titration and water of crystallisation

80. In a titration, a student found that 34.5 cm³ of a 0.15 mol dm⁻³ sodium hydroxide solution were required to exactly neutralise 25 cm³ of a sulphuric acid solution of unknown concentration. Calculate the concentration of the sulphuric acid solution. Show all your working



$$\begin{aligned} n_{\text{NaOH}} &= \frac{34.5 \times 0.15}{1000} = \mathbf{5.18 \times 10^{-3}} \\ n_{\text{H}_2\text{SO}_4} &= \frac{1}{2} \times n_{\text{NaOH}} = \mathbf{2.59 \times 10^{-3}} \\ \text{concentration H}_2\text{SO}_4 &= \frac{0.00259 \times 1000}{25} = \mathbf{0.104 \text{ mol dm}^{-3}} \end{aligned}$$

81. What volume of a 0.25 mol dm⁻³ solution of calcium hydroxide is required to exactly neutralise 25 cm³ of a 0.1 mol dm⁻³ hydrochloric solution. Show all your working



$$\begin{aligned} n_{\text{HCl}} &= \frac{25 \times 0.1}{1000} = \mathbf{2.50 \times 10^{-3}} \\ n_{\text{Ca}(\text{OH})_2} &= \frac{1}{2} \times n_{\text{HCl}} = \mathbf{1.25 \times 10^{-3}} \\ \text{volume Ca}(\text{OH})_2 &= \frac{0.00125 \times 1000}{0.25} = \mathbf{5.00 \text{ cm}^3} \end{aligned}$$

82. A 0.350 g sample of hydrated calcium chloride, CaCl₂·xH₂O was heated until constant mass. Upon cooling in a desiccator, the mass of the anhydrous compound was 0.264 g. Use the data to calculate (i) a value for x and (ii) hence deduce the relative formula mass of the hydrated compound.

$$\text{Mass of water lost} = 0.350 - 0.264 = \mathbf{0.086 \text{ g}}$$

$$n_{\text{CaCl}_2} = 0.264 / 111.1 = \mathbf{2.38 \times 10^{-3}}$$

$$n_{\text{H}_2\text{O}} = 0.086 / 18 = \mathbf{4.78 \times 10^{-3}}$$

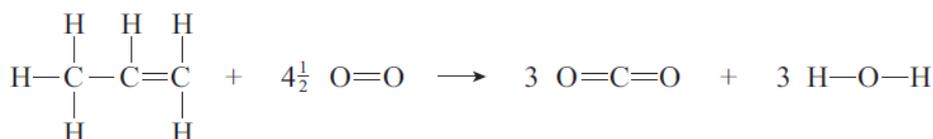
$$\text{CaCl}_2 : \text{H}_2\text{O} = 1 : 2 \text{ so ...}$$

$$\mathbf{x = 2 \text{ therefore RFM} = 40.1 + 35.5 + 35.5 + 18 + 18 = 147.1}$$

Energetics Questions

83. Use the mean bond enthalpy data from the table and the equation given below to calculate a value for the standard enthalpy of combustion of propene. All substances are in the gaseous state.

Bond	C=C	C - C	C -H	O=O	O=C	O-H
Mean bond enthalpy kJ mol^{-1}	612	348	412	496	743	463



$$\begin{aligned}
 \Sigma \Delta H_{(\text{bonds broken})} &= \mathbf{612} (1 \times \text{C}=\text{C}) + \mathbf{2472} (6 \times \text{C}-\text{H}) + \mathbf{348} (1 \times \text{C}-\text{C}) + \mathbf{2232} (4\frac{1}{2} \times \text{O}=\text{O}) \\
 &= 5664 \text{ kJ mol}^{-1} \\
 \Sigma \Delta H_{(\text{bonds made})} &= \mathbf{4458} (1 \times \text{C}=\text{O}) + \mathbf{2778} (6 \times \text{O}-\text{H}) \\
 &= 7236 \text{ kJ mol}^{-1} \\
 \Delta H_{\text{reaction}} &= \Sigma \Delta H_{(\text{bonds broken})} - \Sigma \Delta H_{(\text{bonds made})} = 5664 - 7236 = \underline{\mathbf{-1572 \text{ kJ mol}^{-1}}}
 \end{aligned}$$

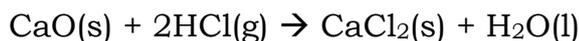
84. The mean bond enthalpy of H-H = +436 kJ mol^{-1} . Use the following equation and data from the table in Q83 above to calculate a value for the C-H bond enthalpy in ethane.



$$\begin{aligned}
 \Sigma \Delta H_{(\text{bonds broken})} &= \mathbf{612} + \mathbf{6x} + \mathbf{348} + \mathbf{436} \\
 \Sigma \Delta H_{(\text{bonds made})} &= \mathbf{696} + \mathbf{8x} \\
 \Delta H_{\text{reaction}} &= \Sigma \Delta H_{(\text{bonds broken})} - \Sigma \Delta H_{(\text{bonds made})} \\
 -140 &= (612 + 6x + 348 + 436) - (696 + 8x) \\
 -140 &= -2x + 700 \text{ so } x (\text{C-H}) = \mathbf{+420 \text{ kJ mol}^{-1}}
 \end{aligned}$$

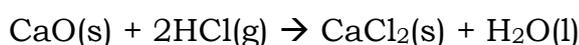
85. Use Hess's Law together with the data given in the table below, to calculate the standard enthalpy change for the following reaction.

	CaO(s)	HCl(g)	CaCl ₂ (s)	H ₂ O(l)
$\Delta H_f / \text{kJ mol}^{-1}$	-602	-92	-642	-286



$$\begin{aligned}\Delta H_{\text{reaction}} &= \Sigma \Delta H_{\text{f}}(\text{products}) - \Sigma \Delta H_{\text{f}}(\text{reactants}) \\ \Delta H_{\text{reaction}} &= [-642 - 286] - [-602 - 92 - 92] \\ \Delta H_{\text{reaction}} &= -142 \text{ kJ mol}^{-1}\end{aligned}$$

86. In an experiment, an excess of solid calcium oxide was added to 40 cm³ of 2.2 mol dm⁻³ hydrochloric acid. The initial temperature of the solution was 21°C. After reaction, the temperature had risen to 53°C. (The specific heat capacity of water is 4.2 J K⁻¹ g⁻¹)



Use this information to calculate the enthalpy change for the reaction of one mole of calcium oxide with hydrochloric acid. For your calculation you should assume that all the heat from the reaction is used to raise the temperature of 40 g of water.

$$q = mc\Delta T = 40 \times 4.2 \times (53) = 8904 \text{ J}$$

$$\text{Since the temperature increased, } Q = \frac{q}{1000} = -8.904 \text{ kJ}$$

Moles of HCl =

$$\frac{\text{conc} \times \text{vol}}{1000} = \frac{2.2 \times 40}{1000} = 0.0880 \text{ mol}$$

Since the mole ratio of CaO : HCl is 1:2, it follows that:
0.044 mol of CaO reacted to increase the temperature,

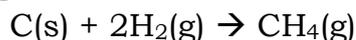
therefore $n = 0.044 \text{ mol}$.

using $\Delta H = \frac{Q}{n}$, the enthalpy change of the reaction

$$= -8.904 / 0.044$$

$$= -202 \text{ kJ mol}^{-1} \text{ (3 sig fig)}$$

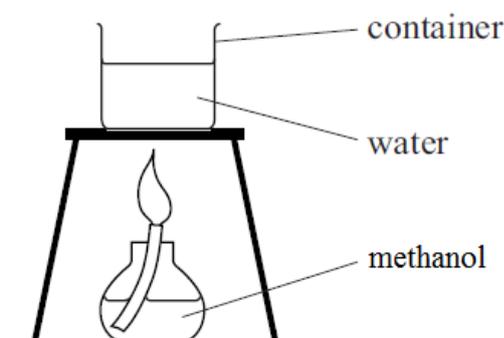
87. Using the enthalpy of combustion data given below, calculate the standard enthalpy change for the formation of methane.



$$\Delta H_{\text{C}} (\text{kJ mol}^{-1}) \quad \text{C(s)} = -394 \quad \text{H}_2(\text{g}) = -286 \quad \text{CH}_4(\text{g}) = -896$$

$$\begin{aligned} \Delta H_{\text{reaction}} &= \Sigma \Delta H_{\text{C}} (\text{reactants}) - \Sigma \Delta H_{\text{C}} (\text{products}) \\ \Delta H_{\text{reaction}} &= [(1 \times -394) + (2 \times -286)] - [(-896)] \\ &= -70 \text{ kJ mol}^{-1} \end{aligned}$$

88. An experiment was carried out to determine a value for the enthalpy of combustion of liquid methanol using the apparatus shown in the diagram.



Burning 3.5 g of methanol caused the temperature of 150 g of water to rise by 50 °C. Use this information to calculate a value for the enthalpy of combustion of methanol, CH₃OH.

$$q = mc\Delta T = 150 \times 4.18 \times 50 = 31350 \text{ J}$$

$$\text{Since the temperature increased, } Q = \frac{q}{1000} = -31.35 \text{ kJ}$$

Moles of CH₃OH =

$$\frac{\text{mass}}{\text{RAM}} = \frac{3.5}{32} = 0.101 \text{ mol}$$

$$\Delta H = \frac{Q}{n}$$

$$\Delta H = \frac{-31.5}{0.101}$$

$$= -310 \text{ kJ mol}^{-1} \text{ (3 sig fig)}$$

Half-life questions

89. 5 g of a radioactive isotope of iodine, ^{131}I has a half-life of 8 days. Calculate what mass of is present after 32 days.

0.313 g

90. Caesium-131 has a half-life of 4 years. What fraction of the original amount present is left after 16 years?

1/16

Grand total of _____ / 90 = _____ %

Answers to examination-style questions

Mole Questions

1. (a) $n_{\text{C}} = 0.124/12 = 0.0103 \text{ mol}$
- (b) $\text{Atoms} = 0.103 \times 6 \times 10^{23} = 6.2 \times 10^{21}$
2. (a)
- | | | |
|-----------|----------------|-----------------------|
| H | Cl | O |
| 1.9 | 67.62 | $100 - (67.62 + 1.9)$ |
| $1.9 / 1$ | $67.62 / 35.5$ | $30.48 / 16$ |
| 1.9 | 1.905 | 1.905 |
| 1 | 1 | 1 |
- Empirical formula = HClO
- (b) $\% \text{ Cl} = (35.5 \times 100) / (1 + 35.5 + 16 + 16 + 16) = 42.0\%$
- (c) Empirical mass = $12 + 1 + 35.5 + 35.5 = 84 \text{ g mol}^{-1}$
 Units = $168 / 84 = 2$ so molecular formula = $\text{C}_2\text{H}_2\text{Cl}_4$
3. (a)(i) $\text{RAM} = [(78.6 \times 24) + (10.1 \times 25) + (11.3 \times 26)] / 100 = 24.33$
- (a)(ii) Neutrons = mass number – atomic number = $26 - 12 = 14$
- (b)(i) Moles = $3.02 / (24.3 + 32.1 + 16 + 16 + 16 + 16) = 0.0251$
- (b)(ii) Moles = $3.14 / 18 = 0.174$
- (b)(iii) $X = 0.174 / 0.0251 = 6.93$, or 7
4. (a)(i) $n_{\text{HCl}} = 2 \times n_{\text{SrCO}_3} = 2 \times 6.25 \times 10^{-3} = 0.0125 \text{ mol}$
 volume = $(0.0125 \times 10000) / 0.200 = 62.5 \text{ cm}^3$
- (a)(ii) $n_{\text{CO}_2} = n_{\text{SrCO}_3} = 6.25 \times 10^{-3} \text{ mol}$
 volume = $6.25 \times 10^{-3} \times 24000 = 150 \text{ cm}^3$
- (b) $\% \text{ atom economy} = (87.6 + 71) / (87.6 + 71 + 18 + 44) \times 100 = 71.9\%$
5. (a) $n_{\text{Ba}} = 0.33/137.3 = 2.40 \times 10^{-3}$
- (b) $n_{\text{H}_2} = n_{\text{Ba}} = 2.40 \times 10^{-3}$
 volume = $2.40 \times 10^{-3} \times 24 = 0.0577 \text{ dm}^3$
- (c) Concentration = $(2.40 \times 10^{-3} \times 1000) / 250 = 9.6 \times 10^{-3} \text{ mol dm}^{-3}$

6. (a) % atom economy = $(28.1 + 32) / [(28.1 + 32) + (4 \times 36.5)]$
 $= 0.292 \times 100 = 29.2\%$
- (b)(i) $n \text{SiCl}_4 = 3.25 / (28.1 + 35.5 + 35.5 + 35.5 + 35.5) = 0.0191$
- (b)(ii) $n \text{SiO}_2 = n \text{SiCl}_4 = 0.0191$
- (b)(iii) Mass = $0.0191 \times (60.1) = 1.148 \text{ g}$
- (b)(iv) % yield = $(0.49 \times 100) / 1.148 = 42.68\%$

7. (a)(i) $n\text{Mg} = 0.0947 / 24.3 = 3.89 \times 10^{-3} \text{ mol}$
- (a)(ii) $n\text{HNO}_3 = 2 \times n\text{Mg} = 7.79 \times 10^{-3} \text{ mol}$
- (a)(iii) $n\text{HNO}_3 \text{ added} = (45.3 \times 0.275) / 1000 = 0.0125$
- (a)(iv) $n\text{HNO}_3 \text{ unreacted} = 0.0125 - 7.79 \times 10^{-3} = 4.71 \times 10^{-3} \text{ mol}$
- (b) $Volume = \frac{nRT}{P} = \frac{0.589 \times 8.31 \times 245}{120000} = 0.010 \text{ m}^3$

(c)

N	O
30.4	69.6
30.4/14	69.6/16
2.17	4.35
1	2

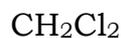


8. (a)(i) $n\text{HCl} = (10.49 \times 0.500) / 1000 = 5.25 \times 10^{-3} \text{ mol}$
- (a)(ii) $n\text{MCO}_3 = \frac{1}{2} \times n\text{HCl} = 2.63 \times 10^{-3} \text{ mol}$
- (a)(iii) $\text{RFM} = \text{mass} / \text{mol} = 0.387 / 2.63 \times 10^{-3} = 147.1 \text{ g mol}^{-1}$
- (b) $\text{RAM M} = 147.1 - (12 + 16 + 16 + 16) = 87.1$, Strontium
 OR $M = 100.1 - (12 + 16 + 16 + 16) = 40.1$, Calcium

9. (a) % atom economy = $(12 + 142) / (154 + 146) = 51.3\%$

(b)

C	H	Cl
14.1	$100 - 97.6 = 2.4$	83.5
$14.1/12 = 1.175$	$2.4/1 = 2.4$	$83.5/35.5 = 2.35$
1	2.04	2



10. (a) $\text{RMM} = (6 \times 12) + (10 \times 1) = 82 \text{ g mol}^{-1}$
 (b) $\% = (6 \times 12 \times 100) / 82 = 87.8\%$
 (c) $n \text{ cyclohexanol} = 8.75 / 100 = 0.0875$
 $\text{theoretical } n \text{ cyclohexene} = 0.0875$
 $\% \text{ yield} = (0.0348 \times 100) / 0.0875 = 39.8\%$

11.

C	H	Cl
38.4	4.80	56.8
$38.4/12 = 3.2$	$4.8/1 = 4.8$	$56.8/35.5 = 1.6$
2	3	1

 $\text{C}_2\text{H}_3\text{Cl}$, option C

12. $n \text{ HCl} = 40 \times 0.25 / 1000 = 0.01 \text{ mol}$
 $n \text{ CaO} = \frac{1}{2} n \text{ HCl}$, so 0.005, option D
13. A $1.5 \times 2 = 3$
 B $0.5 \times 4 = 2$
 C $1 \times 2 = 2$
 D $2 \times 2 = 4$, option D
14. $\text{RAM} = (50 \times 6.3) + (52 \times 82.2) + (53 \times 9.1) + (54 \times 2.4)$
 $= 5201 / 100 = 52.0$ (3sf)
15. $8 \text{ g O}_2 = 8 / 32 = 0.25 \text{ mol}$
 $\text{Mol H}_2\text{O}_2 = 2 \times \text{mol O}_2 = 0.50$
 $\text{Mass H}_2\text{O}_2 = 34 \times 0.50 = 17 \text{ g}$
16. Mole ratio 1:1 so 55 cm^3 formed
17. $n\text{Ca(OH)}_2 = (100 \times 0.3) / 1000 = 0.03 \text{ mol}$
 $\text{mass} = 0.03 \times 74 = 2.22 \text{ g}$
18. $n\text{H}_2\text{SO}_4 = (0.01 \times 34) / 1000 = 3.4 \times 10^{-4} \text{ mol}$
 $n\text{KOH} = 2 \times n\text{H}_2\text{SO}_4 = 6.8 \times 10^{-4} \text{ mol}$
 $\text{volume} = (6.8 \times 10^{-4} \times 1000) / 0.4 = 1.7 \text{ cm}^3$

19.

C	H	N
1.45	0.482	1.69
$1.45 / 12 = 0.120$	$0.482 / 1 = 0.482$	$1.69 / 14 = 0.121$
1	4	1

 CH_4N

20. (a) n cyclohexanol = $86 / 100 = 0.86$
 theoretical n cyclohexene = 0.86
 n cyclohexene = $42.8 / 82 = 0.521$
 $\% \text{ yield} = (0.521 \times 100) / 0.86 = 60.6\%$
- (b) $\% \text{ atom economy} = (82 \times 100) / 100 = 82\%$
21. $\text{RAM} = (80 \times 107) + (20 \times 109) / 100 = 107.4$
22. A $1 / 16 \text{ mol} = 0.0625 \text{ mol}$
 B $1 / 48 \text{ mol} = 0.021 \text{ mol}$
 C $1 / 20 \text{ mol} = 0.05 \text{ mol}$
 D $1 / 20 \text{ mol} = 0.05 \text{ mol}$
 B
23. A = $8.0 / 2 = 4 \text{ mol}$
 B = $8.8 / 44 = 0.2 \text{ mol}$
 C = $16 / 32 = 0.5 \text{ mol}$
 D = $28 / 28 = 1 \text{ mol}$
 A
24. $n\text{CaO} = n\text{CaCO}_3 = 3.400$
 $\text{mass} = 3.4 \times (40.1 + 16) = 190.7 \text{ g}$
25. $16 \text{g N}_2 = 0.571 \text{ mol}$, $= 2 \times 0.571 \text{ mol of atoms} = 1.142$
 Moles boron atoms = moles of N atoms = 1.142
 Mass of boron = $1.142 \times 10.8 = 12.3 \text{ g}$
- 26.
- | | | |
|----------------|------------|-----------|
| Pb | C | O |
| 77.5 | 4.5 | 18 |
| $77.5 / 207.2$ | $4.5 / 12$ | $18 / 16$ |
| 0.374 | 0.375 | 1.125 |
| 1 | 1 | 3 |
- PbCO₃
27. $\text{RFM} = 118.7 + 118.7 + 118.7 + 16 + 16 + 16 + 16 = 420.1$
28. $\% = (14 + 14) / (14 + 14 + 8 + 32.1 + 64) = 0.211 \times 100 = 21.1\%$

Energetics Questions

1.

$$\Delta H^{\ominus}_{\text{reaction}} = \sum \Delta H_{\text{f}}^{\ominus} (\text{products}) - \sum \Delta H_{\text{f}}^{\ominus} (\text{reactants})$$

$$\Delta H^{\ominus}_{\text{reaction}} = (-1669) - (3 \times -538)$$

$$\Delta H^{\ominus}_{\text{reaction}} = -55 \text{ kJ mol}^{-1}$$
2.
 - (a) $q = m c \Delta T$
 - (b) Heat = $11.6 \times 1.92 \times 9.5 = 212 \text{ J}$ (3 sig fig)
 - (c) Heat = $11.95 \times 0.96 \times 9.5 = 109 \text{ J}$ (3 sig fig)
 - (d) But as it is 0.1 mol, the values are multiplied by 10 so
 $\Delta H = 2120 + 1090 = 3210 \text{ J mol}^{-1} = 3.21 \text{ kJ mol}^{-1}$
3.
 - (a) $q = m c \Delta T = 200 \times 4.18 \times (88 - 23) = 54340 \text{ J} = 54.34 \text{ kJ}$
 - (b) $M_r \text{ hexan-1-ol} = (6 \times 12) + (14 \times 1) + (1 \times 16) = 102 \text{ g mol}^{-1}$
 $n_{\text{hexan-1-ol}} = 1.67 / 102 = 0.0164 \text{ mol}$
 - (c) $\Delta H = -q / n = -54.34 / 0.0164 = -3310 \text{ kJ mol}^{-1}$ (3 sig fig)
4.

$$5\text{C} + 6\text{H}_2 \rightarrow \text{C}_5\text{H}_{12}$$

$$\Delta H^{\ominus}_{\text{reaction}} = \sum \Delta H_{\text{C}}^{\ominus} (\text{reactants}) - \sum \Delta H_{\text{C}}^{\ominus} (\text{products})$$

$$\Delta H^{\ominus}_{\text{reaction}} = [(5 \times -394) + (6 \times -286)] - [-3578]$$

$$\Delta H^{\ominus}_{\text{reaction}} = -108 \text{ kJ mol}^{-1}$$
5.

$$\Delta H^{\ominus}_{\text{reaction}} = \sum \Delta H_{\text{C}}^{\ominus} (\text{reactants}) - \sum \Delta H_{\text{C}}^{\ominus} (\text{products})$$

$$\Delta H^{\ominus}_{\text{reaction}} = -394 + (2 \times -286) - -891$$
 Option B
6.

$$\Delta H^{\ominus}_{\text{reaction}} = \sum \Delta H_{\text{f}}^{\ominus} (\text{products}) - \sum \Delta H_{\text{f}}^{\ominus} (\text{reactants})$$

$$\Delta H^{\ominus}_{\text{reaction}} = -930 - (-240 \times 2)$$
 Option B
7.

$$\Delta H^{\ominus}_{\text{reaction}} = \sum \Delta H_{\text{f}}^{\ominus} (\text{products}) - \sum \Delta H_{\text{f}}^{\ominus} (\text{reactants})$$

$$-46 = \Delta H_{\text{f}}^{\ominus} (\text{ethane}) - 52.3 + 0$$

$$\Delta H_{\text{f}}^{\ominus} (\text{ethane}) = -46 + 52.3 = +6.3 \text{ kJ mol}^{-1}$$
8.
 - (a) In order to calculate moles of ethanol used as ΔH is kJ mol^{-1}
 - (b) Mass = before – after = $15.9 - 12.9 = 3 \text{ g}$
 $n_{\text{ethanol}} = 3 / 46 = 0.0652 \text{ mol}$
 - (c) $\Delta H_{\text{C}}^{\ominus} = - [(250 \times 4.18 \times (60.4 - 19.2))] / 0.0652$
 Divide by 1000 to convert to kJ
 $\Delta H_{\text{C}}^{\ominus} = -660 \text{ kJ mol}^{-1}$

9. $\Delta H_{\text{reaction}} = \Sigma \Delta H_{\text{(bonds broken)}} - \Sigma \Delta H_{\text{(bonds made)}}$
 $\Delta H_{\text{reaction}} = (436 + 151) - (299 + 299)$
 $\Delta H_{\text{reaction}} = -11 \text{ kJ mol}^{-1}$
10. (a) $M_r \text{ butanone} = (4 \times 12) + (8 \times 1) + (1 \times 16) = 72 \text{ g mol}^{-1}$
 $n_{\text{butanone}} = 3.20 / 72 = 0.0444 \text{ mol}$
- (b) Heat transferred = $300 \times 4.18 \times (81.4 - 20) = 76996 \text{ J} = 77 \text{ kJ}$
- (c) $\Delta H_{\text{C}}^{\ominus} = -77 / 0.0444 = -1730 \text{ kJ mol}^{-1}$ (3 sig fig)
11. $\Delta H^{\ominus}_{\text{reaction}} = \Sigma \Delta H_{\text{C}}^{\ominus}(\text{reactants}) - \Sigma \Delta H_{\text{C}}^{\ominus}(\text{products})$
 $\Delta H^{\ominus}_{\text{reaction}} = (3 \times -394) + (3 \times -286) - -870$
 $\Delta H^{\ominus}_{\text{reaction}} = -1170 \text{ kJ mol}^{-1}$ (3 sig fig)
12. $\Delta H_{\text{reaction}} = \Sigma \Delta H_{\text{(bonds broken)}} - \Sigma \Delta H_{\text{(bonds made)}}$
 $\Delta H_{\text{reaction}} = [(4 \times 413) + 243] - [(3 \times 413) + 346 + 432]$
 $\Delta H_{\text{reaction}} = 1895 - 2017 = -122 \text{ kJ mol}^{-1}$ (3 sig fig)
13. $\Delta H_{\text{reaction}} = \Sigma \Delta H_{\text{(bonds broken)}} - \Sigma \Delta H_{\text{(bonds made)}}$
 $-240 = [2x + 348 + (6 \times 412) + (2 \times 436)] - [(3 \times 348) + (10 \times 412)]$
 $-240 = 2x + 3692 - 5164$
 $-240 = 2x - 1472$
 $1232 = 2x$
 $\text{C}=\text{C} = 1232 / 2 = +616 \text{ kJ mol}^{-1}$

Half life

1.

Proportion	Time (years)
1	0
$\frac{1}{2}$	5.1
$\frac{1}{4}$	10.2

10.2 years

2.

Proportion	Time (minutes)
1	0
$\frac{1}{2}$	9.3
$\frac{1}{4}$	18.6
$\frac{1}{8}$	27.9

27.9 minutes

3.

Mass (g)	Time (minutes)
64	0
32	74
16	148
8	222

8 g

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