Huxlow & Rushden Academy

SIXTH FORM

One Trust, Two Schools, One Vision of Excellence

Transition Task

Chemistry A Level

Prepared by: C Woodward

Retrieval questions

You need to be confident about the definitions of terms that describe measurements and results in A Level Chemistry.

Learn the answers to the questions below then cover the answers column with a piece of paper and write as many answers as you can. Check and repeat.

Practical science key terms

When is a measurement valid?	when it measures what it is supposed to be measuring
When is a result accurate?	when it is close to the true value
What are precise results?	when repeat measurements are consistent/agree closely with each other
What is repeatability?	how precise repeated measurements are when they are taken by the same person, using the same equipment, under the same conditions
What is reproducibility?	how precise repeated measurements are when they are taken by different people, using different equipment
What is the uncertainty of a measurement?	the interval within which the true value is expected to lie
Define measurement error	the difference between a measured value and the true value
What type of error is caused by results varying around the true value in an unpredictable way?	random error
What is a systematic error?	a consistent difference between the measured values and true values
What does zero error mean?	a measuring instrument gives a false reading when the true value should be zero
Which variable is changed or selected by the investigator?	independent variable
What is a dependent variable?	a variable that is measured every time the independent variable is changed
Define a fair test	a test in which only the independent variable is allowed to affect the dependent variable

What are control variables?	variables that should be kept constant to avoid them affecting the dependent variable
What does an atom consist of?	a nucleus containing protons and neutrons, surrounded by electrons
What are the relative masses of a proton, neutron, and electron?	1, 1, and $\frac{1}{1836}$ respectively
What are the relative charges of a proton, neutron, and electron?	+1, 0, and -1 respectively
How do the number of protons and electrons differ in an atom?	they are the same because atoms have neutral charge
How does the number of protons differ between atoms of the same element?	it does not differ – all atoms of the same element have the same number of protons
What force holds an atom nucleus together?	strong nuclear force
What is the proton number / atomic number of an element?	the number of protons in the atom's nucleus of an element
What is the mass number of an element?	number of protons + number of neutrons
What is an isotope?	an atom with the same number of protons but different number of neutrons
What is the equation for relative isotopic mass?	relative isotopic mass = $\frac{\text{mass of an isotope}}{\frac{1}{12}^{th} \text{ mass of 1 atom of } ^{12}C}$
What is the equation for relative atomic mass *(Ar)*?	relative atomic mass = $\frac{\text{weighted mean mass of 1 atom}}{\frac{1}{12}^{th} \text{ mass of 1 atom of } ^{12}C}$
What is the equation for relative molecular mass *(Mr)*?	relative molecular mass = $\frac{\text{average mass of 1 molecule}}{\frac{1}{12}^{th} \text{ mass of 1 atom of } ^{12}C}$
What is an ion?	an atom or group of atoms with a charge (a different number of electrons to protons)
Define the term cation	a positive ion (atom with fewer electrons than protons)
Define the term anion	a negative ion (atom with more electrons than protons)
What is the function of a mass spectrometer?	it accurately determines the mass and abundance of separate atoms or molecules, to help us identify them
What is a mass spectrum?	the output from a mass spectrometer that shows the different isotopes that make up an element
What is a binary compound?	a compound which contains only two elements

Maths skills

1 Core mathematical skills

A practical chemist must be proficient in standard form, significant figures, decimal places, SI units, and unit conversion.

1.1 Standard form

In science, very large and very small numbers are usually written in standard form. Standard form is writing a number in the format A × 10x where A is a number from 1 to 10 and x is the number of places you move the decimal place.

For example, to express a large number such as 50 000 mol dm−3 in standard form, A = 5 and x = 4 as there are four numbers after the initial 5.

Therefore, it would be written as 5×104 mol dm−3.

To give a small number such as 0.000 02 Nm2 in standard form, A = 2 and there are five numbers before it so x = −5.

So it is written as 2×10−5 Nm2.

Practice questions

Change the following values to standard form.

- boiling point of sodium chloride: 1413 °C
- largest nanoparticles: 0.0 001×10−3 m
- number of atoms in 1 mol of water: 1806×1021

Change the following values to ordinary numbers.

- 5.5×10−6
- 2.9×102
- 1.115×104
- 1.412×10−3
- 7.2×101

1.2 Significant figures and decimal places

In chemistry, you are often asked to express numbers to either three or four significant figures. The word significant means to 'have meaning'. A number that is expressed in significant figures will only have digits that are important to the number's precision.

It is important to record your data and your answers to calculations to a reasonable number of significant figures. Too many and your answer is claiming an accuracy that it does not have, too few and you are not showing the precision and care required in scientific analysis.

For example, 6.9301 becomes 6.93 if written to three significant figures.

Likewise, 0.000 434 56 is 0.000 435 to three significant figures.

Notice that the zeros before the figure are not significant – they just show you how large the number is by the position of the decimal point. Here, a 5 follows the last significant digit, so just as with decimals, it must be rounded up.

Any zeros between the other significant figures are significant. For example, 0.003 018 is 0.003 02 to three significant figures.

Sometimes numbers are expressed to a number of decimal places. The decimal point is a place holder and the number of digits afterwards is the number of decimal places.

For example, the mathematical number pi is 3 to zero decimal places, 3.1 to one decimal place, 3.14 to two decimal places, and 3.142 to three decimal places.

Practice questions

Give the following values in the stated number of significant figures (s.f.).

- 36.937 (3 s.f.)
- 258 (2 s.f.)
- 0.043 19 (2 s.f.)
- 7 999 032 (1 s.f.)

Use the equation:

number of molecules = number of moles × 6.02 × 1023 molecules per mole

to calculate the number of molecules in 0.5 moles of oxygen. Write your answer in standard form to 3 s.f.

Give the following values in the stated number of decimal places (d.p.).

- 4.763 (1 d.p.)
- 0.543 (2 d.p.)
- 1.005 (2 d.p.)
- 1.9996 (3 d.p.)

1.3 Converting units

Units are defined so that, for example, every scientist who measures a mass in kilograms uses the same size for the kilogram and gets the same value for the mass. Scientific measurement depends on standard units – most are *Système International* (SI) units.

If you convert between units and round numbers properly it allows quoted measurements to be understood within the scale of the observations.

Multiplication factor	Prefix	Symbol
109	giga	G
106	mega	M
103	kilo	k
10-2	centi	c
10–3	milli	m
10–6	micro	µ
10–9	nano	n

Unit conversions are common. For instance, you could be converting an enthalpy change of 488 889 J mol−1 into kJ mol−1. A kilo is 103 so you need to divide by this number or move the decimal point three places to the left.

488 889 ÷ 103kJ mol−1 = 488.889 kJ mol−1

Converting from mJ mol−1 to kJ mol−1, you need to go from 103 to 10−3, or move the decimal point six places to the left.

333 mJ mol−1 is 0.000 333 kJ mol−1

If you want to convert from 333 mJ mol−1 to nJ mol−1, you would have to go from 10−9 to 10−3, or move the decimal point six places to the right.

333 mJ mol−1 is 333 000 000 nJ mol−1

Practice question

Calculate the following unit conversions.

- 300 µm to m
- 5 MJ to mJ
- 10 GW to kW

2 Balancing chemical equations

2.1 Conservation of mass

When new substances are made during chemical reactions, atoms are not created or destroyed – they just become rearranged in new ways. So, there is always the same number of each type of atom before and after the reaction, and the total mass before the reaction is the same as the total mass after the reaction. This is known as the conservation of mass. You need to be able to use the principle of conservation of mass to write formulae, and balanced chemical equations and half equations.

2.2 Balancing an equation

The equation below shows the correct formulae but it is not balanced.

$H_2 + O_2 \rightarrow H_2O$

While there are two hydrogen atoms on both sides of the equation, there is only one oxygen atom on the right-hand side of the equation against two oxygen atoms on the left-hand side. Therefore, a two must be placed before the H_2O.

$H_2 + O_2 \rightarrow 2H_2O$

Now the oxygen atoms are balanced but the hydrogen atoms are no longer balanced. A two must be placed in front of the H_2.

$2H_2 + O_2 \rightarrow 2H_2O$

The number of hydrogen and oxygen atoms is the same on both sides, so the equation is balanced.

Practice question

Balance the following equations.

- $C + O_2 \rightarrow CO$
- $N_2 + H_2 \rightarrow NH_3$
- $C_2H_4 + O_2 \rightarrow H_2O + CO_2$

2.3 Balancing an equation with fractions

To balance the equation below:

$C_2H_6 + O_2 \rightarrow CO_2 + H_2O$

Place a two before the CO_2 to balance the carbon atoms.

Place a three in front of the H_2O to balance the hydrogen atoms.

$C_2H_6 + O_2 \rightarrow 2CO_2 + 3H_2O$

There are now four oxygen atoms in the carbon dioxide molecules plus three oxygen atoms in the water molecules, giving a total of seven oxygen atoms on the product side.

To balance the equation, place three and a half in front of the O_2.

$C_2H_6 + 3\frac{1}{2}O_2 \rightarrow 2CO_2 + 3H_2O$

Finally, multiply the equation by 2 to get whole numbers.

$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$

Practice question

1 Balance the equations below.

- $C_6H_{14} + O_2 \rightarrow CO_2 + H_2O$

$NH_2CH_2COOH + O_2 \rightarrow CO_2 + H_2O + N_2$

2.4 Balancing an equation with brackets

Ca(OH)2+ HCl → CaCl2 + H2O

Here the brackets around the hydroxide (OH−) group show that the Ca(OH)2 unit contains one calcium atom, two oxygen atoms, and two hydrogen atoms.

To balance the equation, place a two before the HCl and another before the H2O.

Ca(OH)2+ 2HCl → CaCl2 + 2H2O

Practice question

Balance the equations below.

- Mg(OH)2 + HNO3 → Mg(NO3)2 + H2O
- Fe(NO3)2 + Na3PO4 → Fe3(PO4)2 + NaNO3

3 Rearranging equations and calculating concentrations

3.1 Rearranging equations

In chemistry, you sometimes need to rearrange an equation to find the desired values.

For example, you may know the amount of a substance (n) and the mass of it you have (m), and need to find its molar mass (M).

The amount of substance (n) is equal to the mass you have (m) divided by the molar mass (M):

You need to rearrange the equation to make the molar mass (M) the subject.

Multiply both sides by the molar mass (M):

$M \times n = m$

Then divide both sides by the amount of substance (n):

Practice questions

Rearrange the equation to make:

- n the subject of the equation
- V the subject of the equation.

Rearrange the equation PV = nRT to make:

- n the subject of the equation
- T the subject of the equation

3.2 Calculating concentration

The concentration of a solution (a solute dissolved in a solvent) is a way of saying how much solute, in moles, is dissolved in 1 dm3 or 1 litre of solution.

Concentration is usually measured using units of mol dm−3. (It can also be measured in g dm3.)

The concentration of the amount of substance dissolved in a given volume of a solution is given by the equation:

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

where n is the amount of substance in moles, c is the concentration, and V is the volume in dm3.

The equation can be rearranged to calculate:

·the amount of substance n, in moles, from a known volume and concentration of solution

·the volume V of a solution from a known amount of substance, in moles, and the concentration of the solution.

Practice questions

- Calculate the concentration, in mol dm−3, of a solution formed when 0.2 moles of a solute is dissolved in 50 cm3 of solution.
- Calculate the concentration, in mol dm−3, of a solution formed when 0.05 moles of a solute is dissolved in 2.0 dm3 of solution.
- Calculate the number of moles of NaOH in an aqueous solution of 36 cm3 of 0.1 mol dm−3.

4 Molar calculations
4.1 Calculating masses and gas volumes
The balanced equation for a reaction shows how many moles of each reactant and product are involved in a chemical reaction.
If the amount, in moles, of one of the reactants or products is known, the number of moles of any other reactants or products can be calculated.
The number of moles (n), the mass of the substance (m), and the molar mass (M) are linked by:

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

Note: The molar mass of a substance is the mass per mole of the substance. For $CaCO_3$, for example, the atomic mass of calcium is 40.1, carbon is 12, and oxygen is 16. So the molar mass of $CaCO_3$ is:
$40.1 + 12 + (16 \times 3) = 100.1$. The units are g mol^{-1}.

Look at this worked example. A student heated 2.50 g of calcium carbonate, which decomposed as shown in the equation:
$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
The molar mass of calcium carbonate is 100.1 g mol^{-1}.

a Calculate the amount, in moles, of calcium carbonate that decomposes.
= 2.50/100.1 = 0.025 mol

b Calculate the amount, in moles, of carbon dioxide that forms.
From the balanced equation, the number of moles of calcium carbonate =number of moles of carbon dioxide = 0.025 mol

Practice questions
In a reaction, 0.486 g of magnesium was added to oxygen to produce magnesium oxide.
$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$
- Calculate the amount, in moles, of magnesium that reacted.
- Calculate the amount, in moles, of magnesium oxide made.
- Calculate the mass, in grams, of magnesium oxide made.

Oscar heated 4.25 g of sodium nitrate. The equation for the decomposition of sodium nitrate is:
$2NaNO_3(s) \rightarrow 2NaNO_2(s) + O_2(g)$
- Calculate the amount, in moles, of sodium nitrate that reacted.
- Calculate the amount, in moles, of oxygen made.

0.500 kg of magnesium carbonate decomposes on heating to form magnesium oxide and carbon dioxide. Give your answers to 3 significant figures.
$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$
- Calculate the amount, in moles, of magnesium carbonate used.
- Calculate the amount, in moles, of carbon dioxide produced.

5 Percentage yields and percentage errors
5.1 Calculating percentage yield
Chemists often find that an experiment makes a smaller amount of product than expected. They can predict the amount of product made in a reaction by calculating the percentage yield.
The percentage yield links the actual amount of product made, in moles, and the theoretical yield, in moles:

$$\text{percentage yield} = \frac{\text{actual amount (in moles) of product}}{\text{theoretical amount (in moles) of product}} \times 100$$

Look at this worked example. A student added ethanol to propanoic acid to make the ester, ethyl propanoate, and water.
$C_2H_5OH + C_2H_5COOH \rightarrow C_2H_5COOC_2H_5 + H_2O$
The experiment has a theoretical yield of 5.00 g.
The actual yield is 4.50 g.

The molar mass of C2H5COOC2H5 = 102.0 g mol−1
Calculate the percentage yield of the reaction.
Actual amount of ethyl propanoate: = 4.5/102 = 0.0441 mol
Theoretical amount of ethyl propanoate: = 5.0/102 = 0.0490 mol
percentage yield = (0.0441/0.0490) × 100% = 90%

Practice questions

- Calculate the percentage yield of a reaction with a theoretical yield of 4.75 moles of product and an actual yield of 3.19 moles of product. Give your answer to 3 significant figures.
- Calculate the percentage yield of a reaction with a theoretical yield of 12.00 moles of product and an actual yield of 6.25 moles of product. Give your answer to 3 significant figures.

5.3 Calculating percentage error in apparatus

The percentage error of a measurement is calculated from the maximum error for the piece of apparatus being used and the value measured:

$$\text{percentage error} = \frac{\text{maximum error}}{\text{measuredvalue}} \times 100\%$$

Look at this worked example. In an experiment to measure temperature changes, an excess of zinc powder was added to 50 cm3 of copper(II) sulfate solution to produce zinc sulfate and copper.
Zn(s) + CuSO4(aq) → ZnSO4(aq) + Cu(s)
The measuring cylinder used to measure the copper(II) sulfate solution has a maximum error of ±2 cm3.

a Calculate the percentage error.
percentage error = (2/50) × 100% =4%

b A thermometer has a maximum error of ±0.05 °C.
Calculate the percentage error when the thermometer is used to record a temperature rise of 3.9 °C. Give your answer to 3 significant figures.
percentage error =(2 × 0.05)/3.9 × 100% =2.56%
(Notice that two measurements of temperature are required to calculate the temperature change so the maximum error is doubled.)

Practice questions

A gas syringe has a maximum error of ±0.5 cm3. Calculate the maximum percentage error when recording these values. Give your answers to 3 significant figures.

- 21.0 cm3
- 43.0 cm3

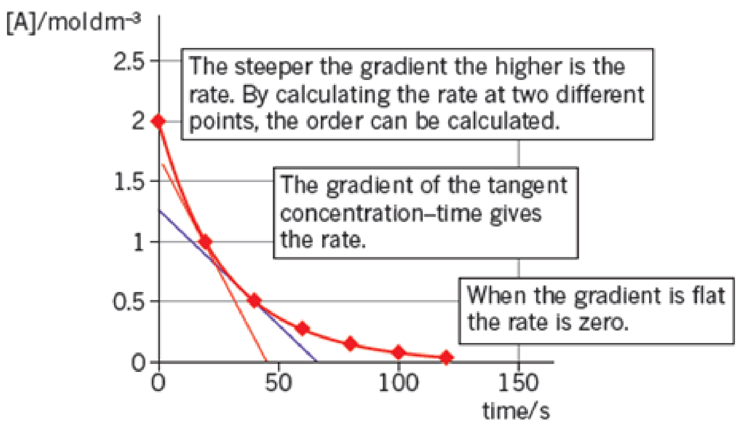
A thermometer has a maximum error of ±0.5 °C. Calculate the maximum percentage error when recording these temperature rises. Give your answers to 3 significant figures.

- 12.0 °C
- 37.6 °C

6 Graphs and tangents

6.1 Deducing reaction rates

To investigate the reaction rate during a reaction, you can measure the volume of the product formed, such as a gas, or the colour change to work out the concentration of a reactant during the experiment. By measuring this concentration at repeated intervals, you can plot a concentration–time graph.



Note: When a chemical is listed in square brackets, it just means 'the concentration of' that chemical. For example, $[O_2]$ is just shorthand for the concentration of oxygen molecules.

By measuring the gradient (slope) of the graph, you can calculate the rate of the reaction. In the graph above, you can see that the gradient changes as the graph is a curve. If you want to know the rate of reaction when the graph is curved, you need to determine the gradient of the curve. So, you need to plot a tangent.

The tangent is the straight line that just touches the curve. The gradient of the tangent is the gradient of the curve at the point where it touches the curve.

Looking at the graph above. When the concentration of A has halved to 1.0 mol dm^{-3}, the tangent intercepts the y-axis at 1.75 and the x-axis at 48.

The gradient is $\frac{-1.75}{48} = -0.0365$ (3 s.f.).

So the rate is 0.0365 mol dm^{-3} s^{-1}.

Practice question

Using the graph above, calculate the rate of reaction when the concentration of A halves again to 0.5 mol dm^{-3}.

6.2 Deducing the half-life of a reactant

In chemistry, half-life can also be used to describe the decrease in concentration of a reactant in a reaction. In other words, the half-life of a reactant is the time taken for the concentration of the reactant to fall by half.

Practice question

The table below shows the change in concentration of bromine during the course of a reaction.

Time / s	$[Br_2]$ / mol dm^{-3}
0	0.0100
60	0.0090
120	0.0066
180	0.0053
240	0.0044
360	0.0028

- Plot a concentration–time graph for the data in the table.
- Calculate the rate of decrease of Br_2 concentration by drawing tangents.
- Find the half-life at two points and deduce the order of the reaction.

The development of organic chemistry as a science

Specification references

- 4.1.1
- HSW7 Know that scientific knowledge and understanding develops over time.
- HSW11 Evaluate the role of the scientific community in validating new knowledge and ensuring integrity.

Introduction

As a science, organic chemistry is less than 200 years old. Most historians of science date its origin to the early part of the 19th century, a time in which a mistaken belief was dispelled.
Originally the word ‘organic’ applied to those substances that were produced by living organisms. Berzelius wrote in 1815 that ‘the essential difference between inorganic and organic compounds was that the formation of organic compounds could only be achieved by the influence of a ‘vital force’ which was present in nature’. It was thought that no organic material could be synthesised in the laboratory. Sugar, dyes, starch, oils, alcohol, known since the earliest times, were thought to only be made by nature.

Learning outcomes

After completing the worksheet you should be able to:

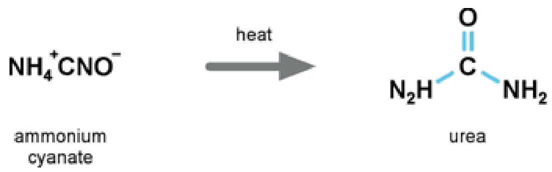
- recognise the contribution some scientists have made to the discipline of organic chemistry
- be aware of how scientists have been influenced by their own beliefs and the effect this had on the way they approached their work.

Background

Vitalism

Scientists first began to distinguish between organic compounds and inorganic compounds during the 1780s. Organic compounds were defined as ‘compounds that could be obtained from living organisms’. Inorganic compounds were those that came from non-living sources. Along with this distinction, a belief called ‘vitalism’ grew. According to this idea, a special influence known as a ‘vital force’ was necessary for the synthesis of an organic compound. Such synthesis, chemists held then, could take place only in living organisms. It could not take place in the test tubes and flasks of a chemistry laboratory.

Between 1828 and 1850 a number of compounds that were clearly ‘organic’ were synthesised from sources that were clearly ‘inorganic’. Friedrich Wöhler accomplished the first of these syntheses in 1828. Wöhler found that evaporating an aqueous solution of the inorganic compound, ammonium cyanate, produced the organic compound, urea (a constituent of urine).



Followers of Berzelius argued that ammonium cyanate was not truly inorganic and, even if it were, the change from ammonium cyanate to urea was merely the result of an alteration of the positions of the atoms within the molecule. The molecule of urea was not, in any real sense, built up of a completely different substance.

If Wöhler's synthesis of urea did not settle the matter of the vital force, Kolbe's synthesis of acetic acid did. In 1845, Adolph Wilhelm Hermann Kolbe (1818–84), a pupil of Wöhler's, succeeded in synthesising acetic acid, unquestionably an organic substance. Furthermore, he synthesised it by a method which showed a clear line of chemical change from the constituent elements, carbon, hydrogen, and oxygen, to the final product, acetic acid.

With a growing number of preparations of other organic substances in the laboratory, it became increasingly evident that organic substances were subject to the same chemical laws as inorganic substances.

Although ‘vitalism’ died slowly and did not disappear completely from scientific circles until 1850, its passing made possible the flowering of the science of organic chemistry that has occurred since 1850.

Organic chemistry today

For the sake of convenience, we still use the term 'organic chemistry' to designate the study of carbon compounds. There are about 100 known elements, and the question naturally arises as to why the element carbon should be assigned such a unique place of honour as a separate branch of chemistry, whereas all the other elements and their compounds are put together to constitute the other branch called 'inorganic chemistry'.

Over two million compounds are known that contain the element carbon, and about 80 000 new carbon compounds are made each year. It is therefore convenient to study the compounds of carbon separately, and this branch of chemistry is known as organic chemistry.

Questions

- Describe why supporters of 'vitalism' believed it was not possible to synthesise organic compounds. (2 marks)
- Suggest where chemists thought the 'vital force' came from. (1 mark)
- Give the molecular formulae for ammonium cyanate and urea. (1 mark)
- What can you deduce from their molecular formulae? (1 mark)
- With Wöhler's synthesis of urea, the original distinction between organic chemistry and inorganic began to fade. Explain why there was initial scepticism to Wöhler's claims. (3 marks)
- Kolbe was the first chemist to truly synthesise an organic substance from its elements. The compound's 'trivial' name is acetic acid. Find out its systematic name. (1 mark)
- Draw the structural formula of ethanoic acid. (1 mark)
- Ethanoic acid cannot be made directly from its elements; but it can be made indirectly from its elements in three steps:
  - produce carbon monoxide;
  - react the carbon monoxide with hydrogen to produce, methanol, CH3OH;
  - finally, react methanol with more carbon monoxide to produce ethanoic acid.

  Suggest an equation for each step in the synthesis of ethanoic acid. Use structural formulae to represent the organic molecules (3 marks)