

Transition from GCSE to A Level

The Priory Academy LSST, Cross O' Cliff Hill, Lincoln, Lincolnshire, LN5 8PW

Moving from GCSE Science to A Level can be a daunting leap – and it's a big one! You'll be expected to remember a lot more facts, equations, and definitions, and you will need to learn new maths skills and develop confidence in applying what you already know to unfamiliar situations.

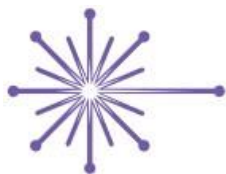
The plan:

- to pre-learn some useful knowledge from the first chapters of your A Level course – you'd be surprised how often we get questions about precision and validity from core practical assessments.
- understand and practice of some of the maths skills you'll need.

Atomic structure

Some more bits here that you'll just need to know.

What are the relative masses of a proton, neutron, and electron?	1, 1, and $\frac{1}{1840}$ 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
What is the atomic number of an element?	the number of protons in the nucleus of a single atom of an element
What is the mass number of an element?	number of protons + number of neutrons
What is an ion?	an atom, or group of atoms, with a charge
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 the total number of electrons that each electron shell (main energy level) can contain?	$2n^2$ electrons, where n is the number of the shell
How many electrons can the first three electron shells hold each?	2 electrons (first shell), 8 electrons (second shell), 18 electrons (third shell)
What are the first four electron sub-shells (orbitals) called?	s, p, d, and f (in order)
How many electrons can each orbital hold?	a maximum of 2 electrons
Define the term ionisation energy, and give its unit	the energy it takes to remove a mole of electrons from a mole of atoms in the gaseous state, unit = kJ mol^{-1}
What is the equation for relative atomic mass (A_r)?	relative atomic mass = $\frac{\text{average mass of 1 atom}}{\frac{1}{12} \text{ mass of 1 atom of } ^{12}\text{C}}$
What is the equation for relative molecular mass (M_r)?	relative molecular mass = $\frac{\text{average mass of 1 molecule}}{\frac{1}{12} \text{ mass of 1 atom of } ^{12}\text{C}}$



Maths skills

1 Core mathematical skills

A practical chemist must be proficient in standard form – FACT. significant figures, decimal places, SI units, and unit conversion. Most of the time we do our calculations (and in exams too) in standard form. The markschemes don't mind but sort of default to standard form as its easier. And be warned – the exam board WILL expect your answers to the appropriate number of significant figures without reminding you.

1.1 Guidance

In science, very large and very small numbers are usually written in standard form. Standard form is writing a number in the format $A \times 10^x$ 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 5×10^4 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: $0.0\ 001 \times 10^{-3}$ m
 - number of atoms in 1 mol of water: 1806×10^{21}
- Change the following values to ordinary numbers.
 - 5.5×10^{-6}
 - 2.9×10^2
 - 1.115×10^4
 - 1.412×10^{-3}
 - 7.2×10^1

1.2 Significant figures and decimal places

You will need to express answers to calculations to the smallest number of significant figures displayed in the question. Just a habit you need to develop – annoyingly it even gets us every now and then.

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

- 3 Give the following values in the stated number of significant figures (s.f.).
a 36.937 (3 s.f.) b 258 (2 s.f.) c 0.043 19 (2 s.f.) d 7 999 032 (1 s.f.)
- 4 Use the equation:
number of molecules = number of moles \times 6.02×10^{23} 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.
- 5 Give the following values in the stated number of decimal places (d.p.).
a 4.763 (1 d.p.) b 0.543 (2 d.p.) c 1.005 (2 d.p.) d 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
10^9	giga	G
10^6	mega	M
10^3	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 $488\,889 \text{ J mol}^{-1}$ into kJ mol^{-1} . A kilo is 10^3 so you need to divide by this number or move the decimal point three places to the left.

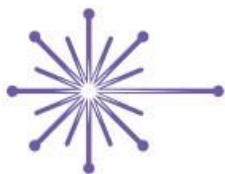
$$488\,889 \div 10^3 \text{ kJ mol}^{-1} = 488.889 \text{ kJ mol}^{-1}$$

Converting from mJ mol^{-1} to kJ mol^{-1} , you need to go from 10^3 to 10^{-3} , or move the decimal point six places to the left.

$$333 \text{ mJ mol}^{-1} \text{ is } 0.000\,333 \text{ kJ mol}^{-1}$$

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.

$$333 \text{ mJ mol}^{-1} \text{ is } 333\,000\,000 \text{ nJ mol}^{-1}$$



Practice question

- 6 Calculate the following unit conversions.
- a 300 μm to m
 - b 5 MJ to mJ
 - c 10 GW to kW

2 Balancing chemical equations

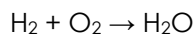
2.1 Conservation of mass

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.

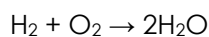
You need to be able to use the principle of conservation of mass to write formulae, and balanced chemical equations and half equations.

2.2 Balancing an equation

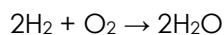
The equation below shows the correct formulae but it is not balanced.



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 .



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 .



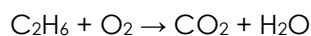
The number of hydrogen and oxygen atoms is the same on both sides, so the equation is balanced.

Practice question

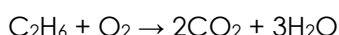
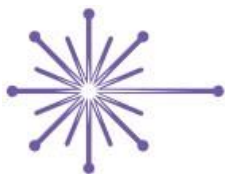
- 1 Balance the following equations.
- a $\text{C} + \text{O}_2 \rightarrow \text{CO}$
 - b $\text{N}_2 + \text{H}_2 \rightarrow \text{NH}_3$
 - c $\text{C}_2\text{H}_4 + \text{O}_2 \rightarrow \text{H}_2\text{O} + \text{CO}_2$

2.3 Balancing an equation with fractions

To balance the equation below:

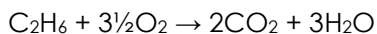


- Place a two before the CO_2 to balance the carbon atoms.
- Place a three in front of the H_2O to balance the hydrogen atoms.

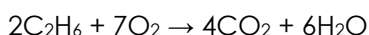


There are now four oxygen atoms in the carbon dioxide molecules plus three oxygen atoms in the water molecules, giving a total of seven oxygen atoms on the product side.

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

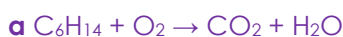


- Finally, multiply the equation by 2 to get whole numbers.

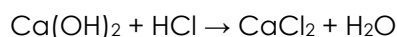


Practice question

- 2 Balance the equations below.

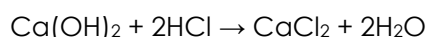


2.4 Balancing an equation with brackets



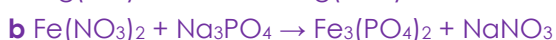
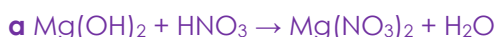
Here the brackets around the hydroxide (OH^-) group show that the $\text{Ca}(\text{OH})_2$ unit contains one calcium atom, two oxygen atoms, and two hydrogen atoms.

To balance the equation, place a two before the HCl and another before the H_2O .



Practice question

- 3 Balance the equations below.



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):

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

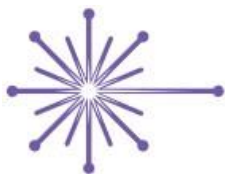
You need to rearrange the equation to make the molar mass (M) the subject.

Multiply both sides by the molar mass (M):

$$M \times n = m$$

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

$$m = \frac{m}{N}$$



Practice questions

- 1 Rearrange the equation $c = \frac{n}{V}$ to make:
 - a n the subject of the equation
 - b V the subject of the equation.
- 2 Rearrange the equation $PV = nRT$ to make:
 - a n the subject of the equation
 - b 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

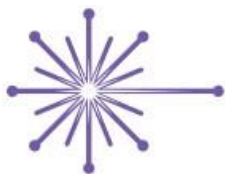
- 3 Calculate the concentration, in mol dm⁻³, of a solution formed when 0.2 moles of a solute is dissolved in 50 cm³ of solution.
- 4 Calculate the concentration, in mol dm⁻³, of a solution formed when 0.05 moles of a solute is dissolved in 2.0 dm³ of solution.
- 5 Calculate the number of moles of NaOH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



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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:



The molar mass of calcium carbonate is 100.1 g mol^{-1} .

a Calculate the amount, in moles, of calcium carbonate that decomposes.

$$n = \frac{m}{M} = 2.50/100.1 = 0.025 \text{ mol}$$

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

1 In a reaction, 0.486 g of magnesium was added to oxygen to produce magnesium oxide.



a Calculate the amount, in moles, of magnesium that reacted.

b Calculate the amount, in moles, of magnesium oxide made.

c Calculate the mass, in grams, of magnesium oxide made.

2 Oscar heated 4.25 g of sodium nitrate. The equation for the decomposition of sodium nitrate is:



a Calculate the amount, in moles, of sodium nitrate that reacted.

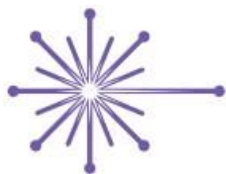
b Calculate the amount, in moles, of oxygen made.

3 0.500 kg of magnesium carbonate decomposes on heating to form magnesium oxide and carbon dioxide. Give your answers to 3 significant figures.



a Calculate the amount, in moles, of magnesium carbonate used.

b Calculate the amount, in moles, of carbon dioxide produced.



5 Percentage yields and percentage errors

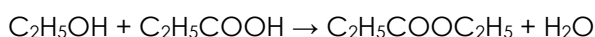
5.1 Calculating percentage yield

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.



The experiment has a theoretical yield of 5.00 g.

The actual yield is 4.50 g.

The molar mass of $\text{C}_2\text{H}_5\text{COOC}_2\text{H}_5 = 102.0 \text{ g mol}^{-1}$

Calculate the percentage yield of the reaction.

$$\text{Actual amount of ethyl propanoate: } n = \frac{m}{M} = 4.5/102 = 0.0441 \text{ mol}$$

$$\text{Theoretical amount of ethyl propanoate: } n = \frac{m}{M} = 5.0/102 = 0.0490 \text{ mol}$$

$$\text{percentage yield} = (0.0441/0.0490) \times 100\% = 90\%$$

Practice questions

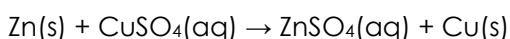
- 1 Calculate the percentage yield of a reaction with a theoretical yield of 4.75 moles of product and an actual yield of 3.19 moles of product. Give your answer to 3 significant figures.
- 2 Calculate the percentage yield of a reaction with a theoretical yield of 12.00 moles of product and an actual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

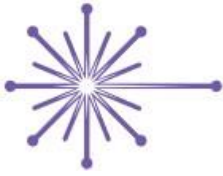
5.3 Calculating percentage error in apparatus

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

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

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.





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The measuring cylinder used to measure the copper(II) sulfate solution has a maximum error of $\pm 2 \text{ cm}^3$.

- a** Calculate the percentage error.

$$\text{percentage error} = (2/50) \times 100\% = 4\%$$

- b** A thermometer has a maximum error of $\pm 0.05 \text{ }^\circ\text{C}$.

Calculate the percentage error when the thermometer is used to record a temperature rise of $3.9 \text{ }^\circ\text{C}$. Give your answer to 3 significant figures.

$$\text{percentage error} = (2 \times 0.05)/3.9 \times 100\% = 2.56\%$$

(Notice that two measurements of temperature are required to calculate the temperature change so the maximum error is doubled.)

Practice questions

- 3** A gas syringe has a maximum error of $\pm 0.5 \text{ cm}^3$. Calculate the maximum percentage error when recording these values. Give your answers to 3 significant figures.
a 21.0 cm^3 **b** 43.0 cm^3
- 4** A thermometer has a maximum error of $\pm 0.5 \text{ }^\circ\text{C}$. Calculate the maximum percentage error when recording these temperature rises. Give your answers to 3 significant figures.
a $12.0 \text{ }^\circ\text{C}$ **b** $37.6 \text{ }^\circ\text{C}$

ANSWERS TO QUESTIONS

1. Maths

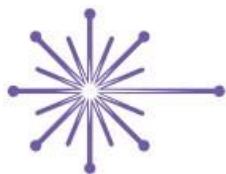
Practice questions

Changing to Standard Form:

- a.** Boiling point of sodium chloride: $1413 \text{ }^\circ\text{C}$ Standard form: $1.413 \times 10^3 \text{ }^\circ\text{C}$
- b.** Largest nanoparticles: $0.000001 \times 10^{-3} \text{ m}$ Standard form: $1 \times 10^{-9} \text{ m}$
- c.** Number of atoms in 1 mol of water: 1806×10^{21} Standard form: 1.806×10^{24}

Changing to Ordinary Numbers:

- a.** 5.5×10^{-6} Ordinary number: 0.0000055
- b.** 2.9×10^2 Ordinary number: 290
- c.** 1.115×10^4 Ordinary number: 11150
- d.** 1.412×10^{-3} Ordinary number: 0.001412
- e.** 7.2×10^1 Ordinary number: 72



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- 1 a 1.413×10^3 °C b 1.0×10^{-7} m
c 1.806×10^{21} atoms
- 2 a 0.000 0055 b 290
c 11150 d 0.001 412
e 72
- 3 a 36.9 b 260
c 0.043 d 8 000 000
- 4 Number of molecules = $0.5 \text{ moles} \times 6.022 \times 10^{23} = 3.011 \times 10^{23} = 3.01 \times 10^{23}$
- 5 a 4.8 b 0.54
c 1.01 d 2.000
- 6 a 0.0003 m b 5×10^9 mJ
c 1×10^7 kW

2 Balancing chemical equations

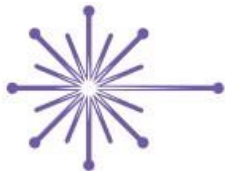
Practice questions

- 1 a $2\text{C} + \text{O}_2 \rightarrow 2\text{CO}$ b $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$
c $\text{C}_2\text{H}_4 + 3\text{O}_2 \rightarrow 2\text{H}_2\text{O} + 2\text{CO}_2$
- 2 a $\text{C}_6\text{H}_{14} + 9\frac{1}{2}\text{O}_2 \rightarrow 6\text{CO}_2 + 7\text{H}_2\text{O}$ or $2\text{C}_6\text{H}_{14} + 19\text{O}_2 \rightarrow 12\text{CO}_2 + 14\text{H}_2\text{O}$
b $2\text{NH}_2\text{CH}_2\text{COOH} + 4\frac{1}{2}\text{O}_2 \rightarrow 4\text{CO}_2 + 5\text{H}_2\text{O} + \text{N}_2$
or $4\text{NH}_2\text{CH}_2\text{COOH} + 9\text{O}_2 \rightarrow 8\text{CO}_2 + 10\text{H}_2\text{O} + 2\text{N}_2$
- 3 a $\text{Mg}(\text{OH})_2 + 2\text{HNO}_3 \rightarrow \text{Mg}(\text{NO}_3)_2 + 2\text{H}_2\text{O}$
b $3\text{Fe}(\text{NO}_3)_2 + 2\text{Na}_3\text{PO}_4 \rightarrow \text{Fe}_3(\text{PO}_4)_2 + 6\text{NaNO}_3$

3 Rearranging equations and calculating concentrations

Practice questions

- 1 a $n = cv$ b $v = \frac{n}{c}$
- 2 a $n = \frac{PV}{RT}$ b $T = \frac{PV}{nR}$
- 3 $\frac{0.2}{0.050} = 4.0 \text{ mol dm}^{-3}$
- 4 $\frac{0.05}{2} = 0.025 \text{ mol dm}^{-3}$
- 5 $\frac{36}{1000} \times 0.1 = 3.6 \times 10^{-3} \text{ mol}$



4 Molar calculations

Practice questions

1 a $\frac{0.486}{24.3} = 0.02 \text{ mol}$ b 0.02 mol

c $0.02 \times 40.3 = 0.806 \text{ g}$

2 a $\frac{4.25}{85} = 0.05 \text{ mol}$ b $\frac{0.05}{2} = 0.025 \text{ mol}$

3 a $\frac{500}{84.3} = 5.93 \text{ mol}$ b 5.93 mol

5 Percentage yields and percentage errors

Practice questions

2 $3.19/4.75 \times 100 = 67.2\%$

2 $6.25/12.00 \times 100 = 52.1\%$

3 a $0.5/21 \times 100 = 2.38\%$ b $0.5/43 \times 100 = 1.16\%$

4 a $0.5 \times (2/12) \times 100 = 8.33\%$ b $0.5 \times (2/37.6) \times 100 = 2.66\%$