2015

Practical and mathematical skills

A comprehensive

guide to A level

Chemistry Practical

Endorsement

Contents

Practical Endorsement

The 2015 OCR A level specification includes a new Practical Endorsement. This means that there is no longer coursework but it has been replaced with regular practical opportunities throughout the year. This is great for Chemists! ☺

You will be expected to keep a neat record of all assessed practical work to provide evidence for you Practical Endorsement. It is also important to keep your lab work safe and in a logical order as you will need it to revise the skills you have learnt throughout the course for your written examinations.

This booklet aims to give you hints and tips for practical work, as well as give you worked examples of mathematical skills and questions for you to have a go!

As you will be given a 'skills test' at the start of the A level Chemistry course, it is important that you work through this booklet to ensure that your maths skills are up to scratch! Don't forget to keep it safe for the duration of the course. This will allow you to go back and remind yourself of the key skills required to meet the demand of A level Chemistry.

Miss Davis and Mr Spencer ☺

sdavis@wadebridge.cornwall.sch.uk

tspencer@wadebridge.cornwall.sch.uk

Equipment lists

You will need to be able to use the following sixth form equipment confidently and safely.

- Pipettes (10 cm^3 and 25 cm^3)
- Pipette fillers
- One burette (50 cm^3)
- One volumetric flask (250 cm^3, 150 cm^3 or 100 cm^3)
- Two conical flasks (250 cm^3)
- One wash bottle
- Four measuring cylinders (250 cm^3, 100 cm^3, 50 cm^3 and 10 cm^3)
- Two filter funnels
- Evaporating basin (at least 30 cm^3)
- Beakers (400 cm^3, 250 cm^3 and 100 cm^3)
- Quickfit apparatus to carry out an organic preparation including distillation
- Plastic cup for use as a calorimeter
- Thermometer (–10 to +110 °C) or equivalent
- Stop clocks/watches reading to 1 s or better.
- Pipeclay triangle
- Porcelain crucible + lid
- Test-tubes and boiling tubes
- Test-tube holders
- Dropping pipettes
- Two stands and clamps
- Bunsen burner
- Balance reading to at least two decimal places
- Glass rods
- Heat proof mat
- Tripod and gauze
- Melting point apparatus or oil bath/thiele tubes
- Apparatus for filtration under reduced pressure
- Melting point tubes

Hazard symbols

The following is a sample of commonly used hazard symbols. It is recommended that you know these as a minimum. A comprehensive guide to hazardous chemicals can be found at www.cleapss.org.uk
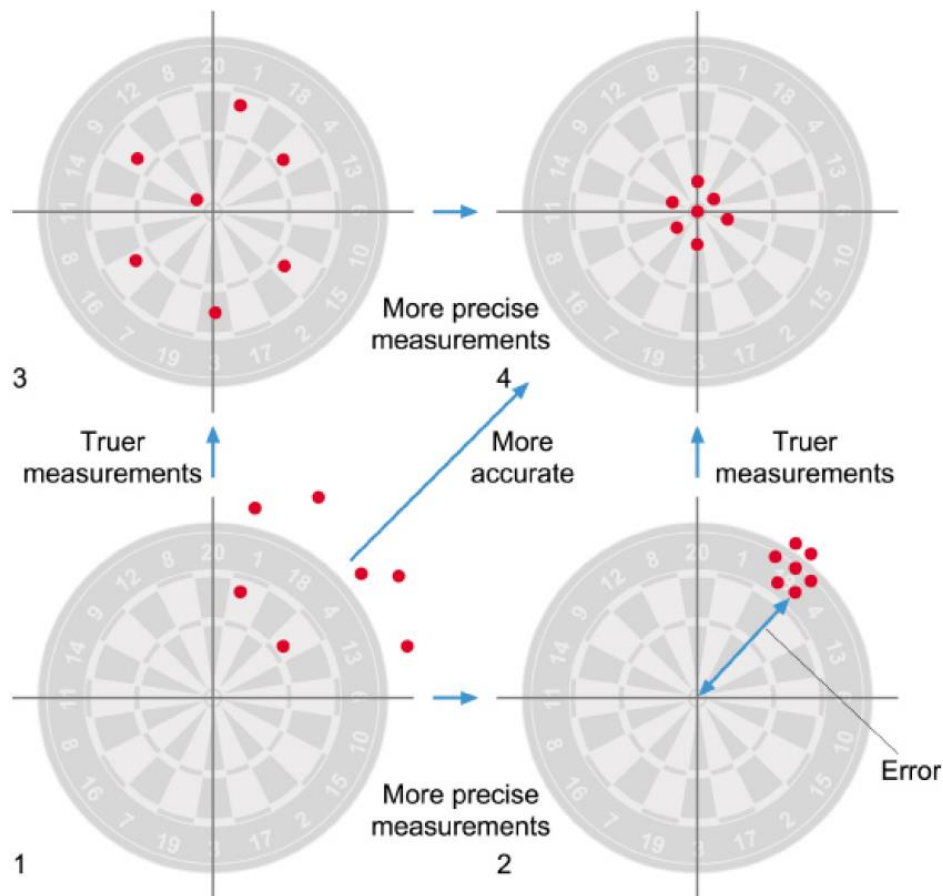
Oxidising

Harmful

Highly Flammable

Corrosive

Toxic

Irritant

What is the difference between accuracy and precision?

In this analogy, the dartboard represents our experiment, the position f the darts represents our results and the bull represents our true value.

In situation 1, the darts are relatively spread out and it is clear that the average position is some way from the bull (or true value). The results are neither **precise** (closely grouped) nor **true** (centred on the bull).

In 2, the darts are grouped more closely together, but the average position is the same as in 1. In this situation, the **precision** is greater but the **trueness** is the same as situation 1.

In 3, the darts are relatively spread out but the average position is close to the bull. Here the **precision** is much lower than in situation 2, but the **trueness** is better.

In 4, the darts are grouped closely together and their average position is close to the bull. We could say that they are the most precise and accurate.

In case 2 the magnitude of the error for one of the results is shown.

## How can I record accurate measurements?

Digital measuring devices

When using a digital measuring device (such as a modern top pan balance or ammeter), record *all* the digits shown.
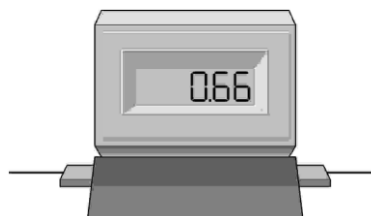
What would you record here? _____

Non-digital measuring devices

When using a non-digital device (such as a ruler or a burette), record all the figures that are known for certain plus one that is estimated.
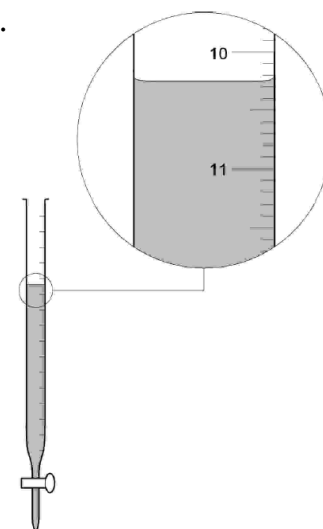
What would you record here? _____

As a general rule, the uncertainty of a piece of equipment is often taken to be half a division on either side of the smallest unit on the scale you are using. However, the accuracy of measurements does also depend on the quality of the apparatus used, such as a balance, thermometer or glassware.

For example, a 100 cm^3 measuring cylinder is graduated in divisions every 1 cm^3.

• A Class A measuring cylinder has a maximum error of half a division or 0.5 cm^3

• A Class B measuring cylinder has a maximum error of a whole division or 1 cm^3.

Because of this variability, assessed tasks will state the **maximum error** in any measurement that is being made.

Calculating maximum error

When glassware is manufactured there will always be a maximum error. This is usually marked on the glassware. We can express the maximum error as a percentage.

Volumetric or standard flask (Class B)

A 250 cm3 volumetric flask has a maximum error of 0.2 cm^3

0.2 ÷ 250 = 0.0008

0.0008 x 100 = **0.08%.**

Pipette (Class B)

A 25 cm^3 pipette has a maximum error of 0.06 cm^3

0.06 ÷ 25 = 0.0024

0.0024 x 100 = **0.24%.**

How badly do maximum errors affect our results?

The significance of the maximum error in a measurement depends upon how large a quantity is being measured. We can calculate the percentage error using the following formula.

$$\text{Percentage error} = \frac{\text{maximum error}}{\text{quantity measured}} \times 100\%$$

1. For example, a two-decimal place balance had a maximum error of 0.005 g.
    For a mass measurement of 2.56 g, what would be the maximum percentage
    error?

2. Calculate the percentage error for a mass measurement of 0.12 g, weighed
    out on the same balance.

3. Which measurement would be most badly affected by the maximum error of
    the balance and how would it affect your experiment?

4. Why do you think that is?

Multiple measurements in the same experiment

When carrying out an experiment that relies on obtaining a resultant mass by difference, the maximum percentage error of the equipment has a larger effect on our results. This is due to a cumulative effect of the maximum errors.

Worked example

Say we did an experiment in which we weighed two different samples and heated them in an open container. We then calculated the mass lost. If we used a digital balance with a maximum error of 0.005 g we could calculate the overall error to the final measurement, using the following formula:

$$\text{Percentage error} = \frac{2 \times \text{maximum error in each measurment}}{\text{quantity measured}} \times 100\%$$

Mass of crucible + crystals before heat = 23.45 g maximum error = 0.005 g

Mass of crucible + crystals after heat = 23.21 g maximum error = 0.005 g

Mass lost = 0.23 g maximum overall error = 2 x 0.005 g

Reading burettes

A burette is graduated in divisions every 0.1 cm^3. A burette is a non-digital device, so we record all figures that are known for certain plus one that is estimated.

Because the divisions are so small, the most we can practically estimate, is half of a division (0.05 cm^3). We therefore record burette measurements to two decimal places with the last figure either '0' or '5'. We could say that the maximum error in each measurement = 0.05 cm^3.

The overall maximum error in any volume measured always comes from two measurements; Measurement 1 is the reading we take when we fill it to zero. Measurement 2 is the reading we take when we have let some out. Therefore the overall maximum error = 2 x 0.05 cm^3 = 0.1 cm^3.

In a titration, a burette will typically deliver about 25.00 cm^3 so the percentage error is small.

Percentage error = (2 × 0.05) ÷ 25.00 = 0.004

× 100 = **0.4%**

For small volumes, the percentage error becomes more significant

For delivery of 2.50 cm^3, the percentage error would be:

$$\text{Percentage error} = (2 \times 0.05) \div 2.50 = 0.04$$

$$\times 100 = \underline{\textbf{4\%}}$$

Example for you to try

Calculate the percentage error for a burette delivery of 13.55 cm^3.

Explain why you think the percentage error increases as the volume measured decreases.

Recording volumes during titrations

During a titration, you should record both the initial and final burette readings you can then use the difference between the two measurements to calculate the mean titre in cm^3. It is usual practice to record titration results in a table like the one shown below.

	trial	1	2	3
final burette reading / cm^3				
initial burette reading / cm^3				
titre / cm^3				
titres used to calculate mean (tick)				
mean titre / cm^3				

When you record the titre, always record it to two decimal places.

How do I calculate mean titres

When recording a mean titre, you should take an average of the concordant titres, *i.e.* those that agree to within 0.10 cm^3. Where this is not possible, the two titres that have the closest agreement should be used.

Worked example

A student carried out an acid-base titration and recorded three titres.

1. 25.80 cm^3
2. 25.30 cm^3
3. 25.20 cm^3.

You would use titre 2 and 3 to calculate the mean titre because they agree to within 0.1 cm^3.

The mean titre is……..

$$(25.30 + 25.20) \div 2 \text{ cm}^3 = \underline{\textbf{25.25 cm}^3}$$

The overall maximum error is………..

$$2 \times 0.05 = \underline{\textbf{0.1 cm}^3}.$$

There is a case for arguing that the accumulated errors indicate that one decimal place is more appropriate but this should **not** be used. We do not use one decimal place because the maximum error is the worst-case scenario and it is likely that the actual titre will in reality be more accurate than one decimal place.

Watch out for averages!

If you were to get concordant titres within 0.05 cm^3 of one another, you may feel pretty pleased with yourself right?! Yes you would have been accurate in your measurements, but you might encounter a problem when calculating your mean titre.

A student obtained three recorded titres.

1. 25.80 cm^3
2. 25.25 cm^3
3. 25.20 cm^3

The mean titre would be

$$(25.25 + 25.20) ÷ 2 \text{ cm}^3 = \underline{\textbf{25.225 cm}^3}$$

The problem is that the mean titre has a value that is more accurate than the burette can measure. The value of 25.225 cm^3 should more correctly be 'rounded' to 25.25 cm^3.

It would seem very unfair not to credit a mean titre of 25.225 cm^3 in this case, especially as this student has carried out the titration better than the first student.

How to round off

If you find yourself in the situation outline above you will have to round off. The rules are very simple. If the last figure is between 5 and 9 inclusive round up; if it is between 0 and 4 inclusive round down.

When do I round off?

It is important to be careful when rounding off in a calculation with two or more steps. If you follow the golden rule (right) you can't go far wrong

If you round off after each step, and then use the rounded figure as the starting figure for the next step, it could make a big difference to the final answer. This introduces a **rounding error**.

Students often introduce rounding errors in multi-step calculations and as a result they lose lots of marks!

> Always leave rounding off till the end of a calculation

How many significant figures should I use?

Ever heard of the expression "your chain is only as strong as the weakest link"? The same is true of a calculation in chemistry. The result of a calculation that involves measured quantities cannot be more certain than the *least* certain of the information that is used. Therefore the result should contain the same number of significant figures as the measurement that has the *smallest* number of significant figures, because that is all you can be sure of.

A common mistake by students is to simply copy down the final answer from the display of a calculator. This often has far more significant figures than the measurements justify and you will lose marks for this.

The amount we round to, depends on the number of significant figures allowed by the accuracy the initial measurements. When a number contains zeros, some confusion can arise. The rules for working out the number of significant figures are given below: (I strongly urge you to learn them)!

1. **Zeros between digits are significant.**
   - For example 2004 has four significant figures.

2. **Zeros to the left of the first non-zero digit are not significant (even when there is a decimal point in the number)**
   - For example 0.002 has one significant figure.

3. **When a number with a decimal point ends in zeros to the right of the decimal point these zeros are significant.**
   - For example 2.0040 has five significant figures.

4. **When a number with no decimal point ends in several zeros, these zeros may or may not be significant. The number of significant figures should then be stated.**

   - For example: 20 000 (to 3 sf) means that the number has been measured to the nearest 100 while 20 000 (to 4 sf) means that the number has been measured to the nearest 10.

<u>The following examples of the number of significant figures should help you work out the number of significant figures in your data.</u>

1. The number 14.33 has four significant figures - all non-zero digits are significant.

2. The number 2000. We cannot tell the number of significant figures. Since there is no decimal point, the zeros may or may not be significant. With numbers with zeroes at the end we must state the number of significant figures.

3. The number 2000.0 has five significant figures - the decimal point implies that we have measured to the nearest 0.1.

4. The number 1.045 has four significant figures - zeros between digits are significant.

5. The number 0.025 has two significant figures - zeros to the left of the decimal point only fix the position of the decimal point. They are not significant.

6. The rules are the same when dealing with numbers expressed in standard form so 6×10^{23} has one significant figure, 6.022×10^{23} has four significant figures.

Rounding to specific significant figures – a worked example

The number 350.99 rounded to:

4 sig fig is 351.0

3 sig fig is 351

2 sig fig is 350

1 sig fig is 400

Notice that when rounding you only look at the one figure beyond the number of figures to which you are rounding, *i.e.* to round to three sig fig you only look at the fourth figure.

How do we know the number of significant figures?

In the example above, 351 has been rounded to the 2 sig fig value of 350.

However, if we came across this on its own, it would be impossible to know whether the final zero in 350 is significant (and the value to 3 sig figs) or insignificant (and the value to 2 sig figs).

In such cases, standard form should be used and is unambiguous:

• 3.5×10^2 is to 2 sig figs

• 3.50×10^2 is to 3 sig figs

Worked example

When 6.074 g of a carbonate is reacted with 50.0 cm^3 of 2.0 mol dm^{-3} HCl (aq) (which is an excess), a temperature rise of 5.5 °C is obtained. The specific heat capacity of the solution is 4.18 J g^{-1} K^{-1},

The heat produced = 50.0 × 4.18 x 5.5 for which a calculator gives 1149.5 J

We ÷ J by 1000 to get KJ = 1.1495 kJ

Therefore, the heat produced = **1.1 kJ**

This is because the least certain measurement (the temperature rise) is only to 2 significant figures. Therefore the answer should also be quoted to 2 significant figures.

> **BE CAREFUL OF EARLY ROUNDING!**
>
> **Don't fall into the trap of rounding your answer too early. If you were going to use this figure to do a further calculation e.g. calculating the enthalpy change per mole; then you do not round until you get the final answer**

Say the carbonate had a molar mass of 84.3 g mol^{-1}, the enthalpy change per mole of carbonate can be calculated from the value above. If we do the right thing and use the calculator value of 1.1495 kJ for the heat produced, the enthalpy per mole would = 15.95371255 kJ mol^{-1}. Rounded to 2 sig figs would give us 16 kJ mol–1.

However if we do the wrong thing and use the rounded value of 1.1 kJ for the heat produced, the enthalpy per mole would = 15.26671057 kJ mol$^{-1.}$. This figure rounded to 2 sig figs would give us 15 kJ mol^{-1} . This would be marked as wrong because of the rounding error!

Errors in procedure

Don't forget that the accuracy of a final result also depends on the procedure used. For example, in an enthalpy experiment, the measurement of a temperature change may be precise but there may be large heat losses to the surroundings which affect the accuracy of overall result.

You will remember from KS4 that random errors are unpredictable. They can be caused by operator error or erratic changes in the way that the test was carried out. Random errors effect particular data points and can cause anomalous results. Systematic errors are more predictable. They are usually caused by faults within the system, e.g. faulty equipment or procedures. They usually affect all data points equally.

Anomalous readings

Where an experiment uses repeated measurements of the same quantity, such as repeated titration readings, anomalous readings should be identified. If a titre is clearly outside the range of all other readings, it can be judged as being anomalous and should be ignored when the mean titre is calculated.

Similarly, if a plotted graph reveals that a value is anomalous, then it should be ignored.

Decimal places

Sometimes, rather than rounding off to the nearest whole number, you might need to be a little more accurate. You might need to include some of the digits after the decimal point. (very likely in Chemistry!)

In these cases, we can round off the number up to a certain number of decimal places.

Do not confuse this with rounding off using significant figures, as this is slightly different!

Remember

The same rules for rounding up apply here:

- **5 or more**, we '**round up**'.
- **4 or less**, it stays as it is

Worked examples

1. Write 2.6470588 g to 2 decimal places (dp).

Answer

*Did you get the answer **2.65 g**? You needed to round up. We want 2 decimal places.*

Look at the 2nd decimal digit.

The 2nd decimal digit is 4. So the number lies between 2.64 and 2.65

Look at the next digit.

The next digit is 7, so we have to round up. So the answer is **2.65 g (2 dp)**.

2. Round off the number 3.9762645 mol dm^{-3} to 1 dp.

Answer

The number lies between 3.9 and 4.0. The 7 after the 9 means you have to round up. So the answer is **4.0 mol cm^{-3}**

3. 0.067295 mol to 3 dp

Answer

0.067 mol

Converting units

1cm^3	**=**	**0.001dm^3**
1000cm^3	**=**	**1dm^3**
25cm^3	=	0.025dm^3
300cm^3	=	0.3dm^3
1g	**=**	**0.001kg**
1000g	**=**	**1kg**
500g	=	0.5kg
23g	=	0.023kg
1J	**=**	**0.001kJ**
1000J	**=**	**1kJ**
68J	=	0.068kJ
7543J	=	7.543kJ

Example for you to try

1. Convert 63 cm^3 to dm^3

2. Convert 3456 g into kg to 2dp

3. Round 1.25 mol to 1dp

4. 25.701 cm^3 to dm^3 to 2 dp

5. 845299 J to kJ to 3 dp

6. 0.011 dm^3 to cm^3

7. 53.273 kg to g

8. 9.283854772106 kJ to J to 3 dp

9. 3.45981 mol dm^3 to 2 dp

10. 15.17625 g to 1 dp

Using Standard form

Mathematicians, scientists and engineers (and your calculator) prefer to write and work with very large and very small numbers in standard form.

A number is in standard form when it is written like this:

This part is made up of a number from 1 up to (but not including) 10.

6.7×10^6

This part is written as a power of 10, and the power is a whole number.

You could think of 1000 as being 1 x 10 x 10 x 10 and write it as 1×10^3.

You could think of 10 000 as being 1 x 10 x 10 x 10 x 10 and write it as 1×10^4.

Complete the table below, the first three rows have been completed for you.

Number	Number in standard form
1 000 000	1×10^6
100 000	1×10^5
10 000	1×10^4
1 000	
100	
10	
1	
0.1	
0.01	
0.001	

The power of ten is the place value of the first significant figure.

As the number is less than one the power of ten is negative.

200 could be written as 2×10^2.

300 could be written as 3×10^2.

250 could be written as 2.5×10^2.

Complete the table on the left, the first three rows have been completed for you.

Number	Number in standard form
200	
230	
300	
399	
400	
415	
500	
550	
9870	

Complete the table below; the first three rows have been completed for you.

Number	Number in standard form
1 000 000 000	1×10^9
12 000 000	1.2×10^7 (not 12×10^6)
15 000	1.5×10^4 (not 15×10^3)
99 000	
12 000 000 000 000	
10 000 000 000 000 000	
99 000 000 000 000 000	
155 000 000	
1 380 000 000	
22 000 000	

Write as ordinary numbers:

Number in standard form	Ordinary number
2.34×10^3	
6.25×10^4	
9.03×10^5	
8×10^6	
3.56×10^7	
1.6×10^6	
4.4×10^{10}	
8.01×10^3	
1.11×10^2	
9.9×10^8	

Write in standard form:

Number	Number in standard form
0.008	8×10^{-3}
0.07	7×10^{-2}
0.55	5.5×10^{-1}
0.000052	
0.048	
0.0086	
0.00086	
0.000086	
0.0000000001	
0.000455	

Balancing equations

NaOH + HCl → NaCl + H₂O
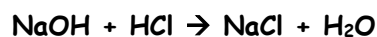
Reactants	Products
Na = 1	Na = 1
O = 1	O = 1
H = 1 + 1 = 2	H = 2
Cl = 1	Cl = 1

The equation is balanced! ☺

Mg(OH)₂ + H₂SO₄ → MgSO₄ + H₂O
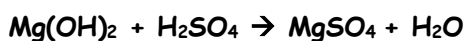
Reactants	Products
Mg = 1	Mg = 1
O = 1 x 2 = 2 + 4 = 6	O = 4 + 1 = 5
H = 1 x 2 = 2 + 2 = 4	H = 2
S = 1	S = 1

The equation is not balanced! ☹

Answer:

Mg(OH)₂ + H₂SO₄ → MgSO₄ + **2H₂O**
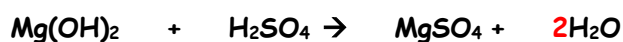
Reactants	Products
Mg = 1	Mg = 1
O = 1 x 2 = 2 + 4 = 6	O = 4 + **2 = 6**
H = 1 x 2 = 2 + 2 = 4	H = 2 **x 2 = 4**
S = 1	S = 1

Try another one!

CH₄ + O₂ → CO₂ + H₂O
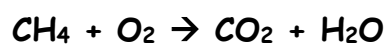
Reactants	Products
C =	C =
H =	H =
O =	O =

Answer:

CH₄ + 2O₂ → CO₂ + 2H₂O

21
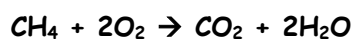
Practice makes perfect! ☺

BALANCING EQUATIONS – 1

Balance the equations shown below. Remember that you may only write numbers in front of any of the formulae but you may not change any formulae. Some of the equations may already be balanced.

1	H_2	+	Cl_2	→	HCl							
2	Zn	+	O_2	→	ZnO							
3	Cl_2	+	Al	→	$AlCl_3$							
4	Na	+	O_2	→	Na_2O							
5	Mg	+	O_2	→	MgO							
6	Mg	+	HCl	→	$MgCl_2$	+	H_2					
7	Fe_2O_3	+	Al	→	Fe	+	Al_2O_3					
8	$CaCl_2$	+	KOH	→	$Ca(OH)_2$	+	KCl					
9	HCl	+	Na_2CO_3	→	$NaCl$	+	H_2O	+	CO_2			
10	HNO_3	+	$NaOH$	→	$NaNO_3$	+	H_2O					
11	HNO_3	+	$Ca(OH)_2$	→	$Ca(NO_3)_2$	+	H_2O					
12	H_2SO_4	+	KOH	→	K_2SO_4	+	H_2O					
13			$NaNO_3$	→	$NaNO_2$	+	O_2					
14	KI	+	$Pb(NO_3)_2$	→	KNO_3	+	PbI_2					
15	$CaCl_2$	+	Na_2SO_4	→	$CaSO_4$	+	$NaCl$					
16	HCl	+	K_2SO_3	→	KCl	+	H_2O	+	SO_2			
17	KOH	+	$MgSO_4$	→	$Mg(OH)_2$	+	K_2SO_4					
18	K	+	H_2O	→	KOH	+	H_2					
19	$NaOH$	+	H_3PO_4	→	Na_3PO_4	+	H_2O					
20			$Pb(NO_3)_2$	→	PbO	+	NO_2	+	O_2			

Percentage yield

Yield describes how much product has been collected. The **maximum** mass of products that can be made from the reactants used is called the **predicted yield.**

To calculate predicted yield you must use the following equation:

Moles = Mass/ Mr

Worked example:

Thermal decomposition of Calcium carbonate (limestone)

$CaCO_3 \rightarrow CaO + CO_2$

Jo heats 2.5g of Calcium carbonate to a constant mass.

1. Calculate the predicted yield of carbon dioxide.

Mr of $CaCO_3$ is 40 + 12 + (16 x 3) = 100 g/ mol

Mass of $CaCO_3$ is 2.5g

Moles = Mass/ Mr   Moles = 2.5 / 100   Moles = 0.025 moles

$CaCO_3$: CO_2 is a 1:1 ratio, therefore moles of CO_2 is also 0.025 moles

Mr of CO_2 is 12 + (16 x 2) = 44 g/ mol

Mass = Mr x Moles   Mass = 44 x 0.025   Mass = 1.1g

Therefore the <u>theoretical yield of carbon dioxide in this reaction is 1.1g.</u>

The mass that you **make in a reaction** is called the **actual yield.** We can compare the predicted and actual yield by using the percentage yield, which is calculated from the formula below:

| Percentage yield | = | <u>Actual yield</u> | x | 100 |
| | | Theoretical yield | | |

Jo only collect 0.4g of carbon dioxide. <u>Therefore her actual yield is 0.4g.</u>

Percentage yield = 0.4 / 1.1 x 100 = 36.36 %

1. Use the formula above to help you calculate the percentage yield in these reactions:

Actual Yield (g)	Predicted Yield (g)	Percentage Yield (g)
50	100	
25	50	
8	40	
80	120	
33	100	
2	5	

Use the method of calculating mass of reactants and products to help you answer the question below: SHOW YOUR WORKINGS CLEARLY!

2. In an experiment, 11.5g of sodium is reacted with an excess of chlorine. The mass of sodium chloride made is 23.4g.

a) What is the maximum amount of product that could be made.

(answer: 29.25 g)

b) What is the percentage yield?

(answer: 80%)

Practical skills

Investigating rate of reaction using marble chips and acid.

Sam wanted to find out how surface area affects the rate of reaction. He measured 3g of powdered calcium carbonate and added 30cm^3 of hydrochloric acid. Every 10 seconds he measured the amount of carbon dioxide produced using a gas syringe. He repeated the experiment using medium chips and large chips.

1. Identify the variables in this investigation:

Independent variable is…

Dependent variable is…

Control variables are…

2. Make a list of all the equipment Sam needs.

3. Which measuring cylinder should Sam used?

        10cm^3 25cm^3 50cm^3 100cm^3

4. Write a risk assessment for this investigation.

5. Write a step by step method for this investigation.

6. Other than using a gas syringe, what other method could Sam have used?

7. Create an appropriate results table for Sam's investigation.

Graphs

1. Construct a graph to show the following data about small marble chips and hydrochloric acid. (Think carefully about what makes a good graph!)

Time (s)	0	10	20	30	40	50	60	70	80	90	100
Volumes of CO_2 (cm^3)	0	10	20	28	34	38	40	41	41	48	41
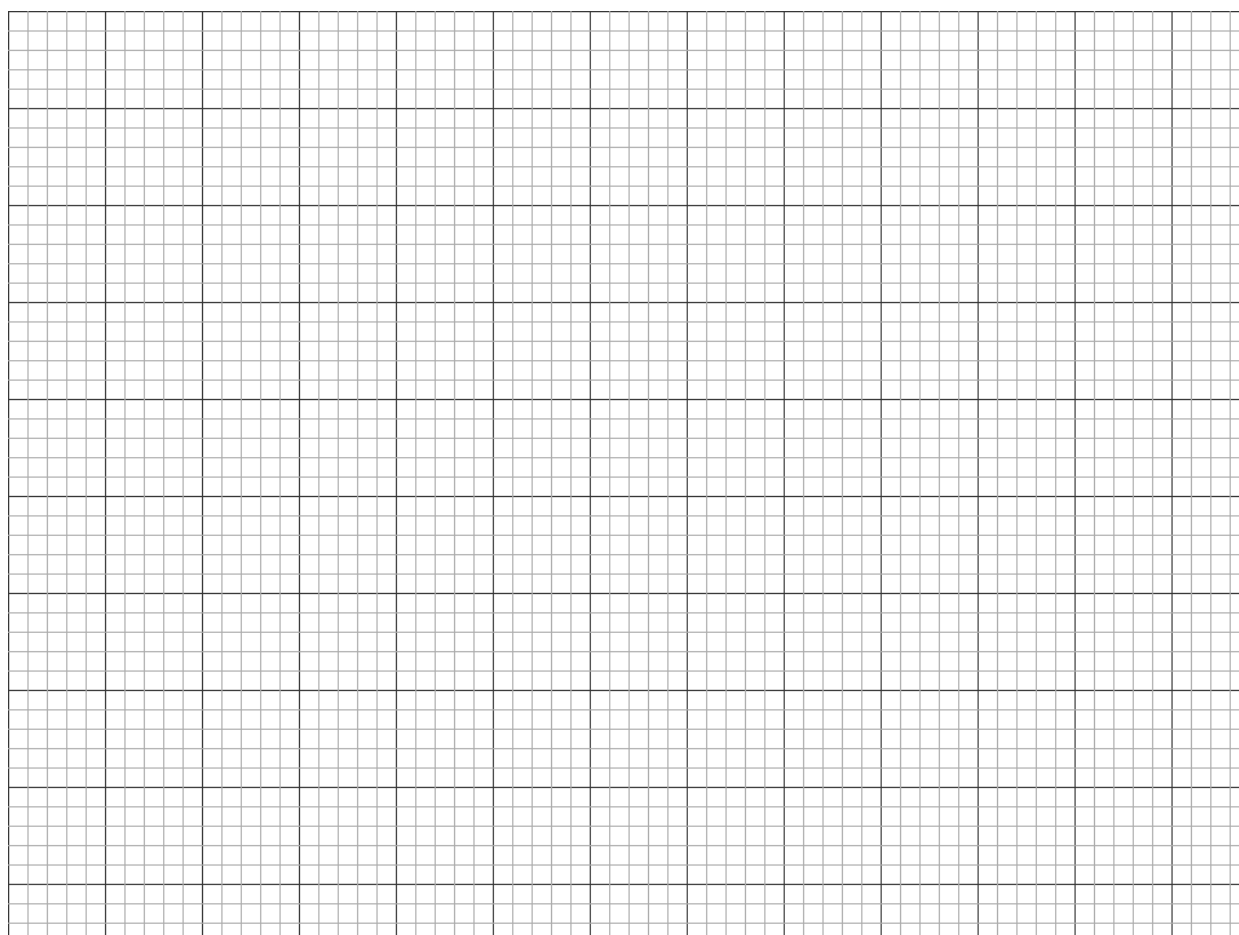
2. Circle the anomalous result.

3. How did you know this was the anomalous result?

Moles Calculations

mol dm^{-3} can also be shown as M

1cm^3 to equal to 1 ml

1 dm^3 is equal to 1 litre

Concentration = Moles / Volume

As volumes are usually in cm^3 you need to divide by 1000 to get the volume into dm^3

Therefore the equation is usually:

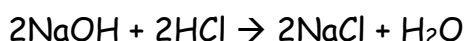Concentration = Moles / Volume/ 1000

1)   If it takes 54 cm^3 of 0.1 M NaOH to neutralize 125 cm^3 of an HCl
     solution, what is the concentration of the HCl?

Fill in what we know from the question:

	NaOH	HCl
Volume (dm^3)	54/1000 = 0.054	125/1000 = 0.125
Concentration (mol dm^{-3})	0.1	?
Moles (mol)	0.0054	0.0054

We can calculate moles of NaOH:
Moles = conc x volume   Moles = 0.1 x 0.054   Moles = 0.0054 or 5.4 x10^{-3} mol

$$2NaOH + 2HCl \rightarrow 2NaCl + H_2O$$

NaOH : HCl is a 2:2 ratio or 1:1 ratio, therefore there is also 0.0054 moles of HCl

Concentration of HCl = moles/ volume   Conc = 0.0054 / 0.125

Conc = **0.0432 mol dm^{-3}**   or   0.0432M   or   4.32 x10^{-2} M

2)   If it takes 25 ml of 0.05 M HCl to neutralize 345 ml of NaOH
      solution, what is the concentration of the NaOH solution?

3)   If it takes 50 cm^3 of 0.5 M KOH solution to completely neutralize 125
      cm^3 of sulfuric acid solution (H_2SO_4), what is the concentration of the
      H_2SO_4 solution? *(hint: you need a balanced equation first!)*