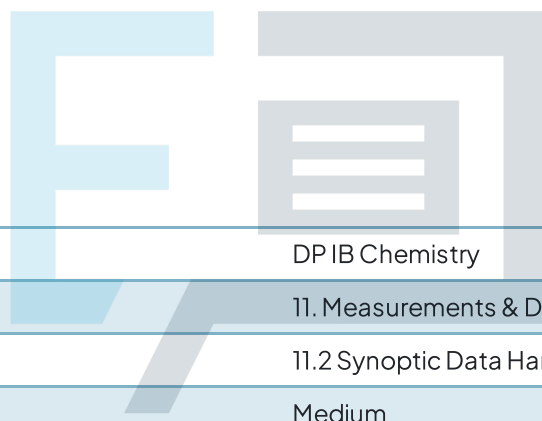




11.2 Synoptic Data Handling & Graphical Skills

Mark Schemes



Course	DP IB Chemistry
Section	11. Measurements & Data Processes
Topic	11.2 Synoptic Data Handling & Graphical Skills
Difficulty	Medium

Exam Papers Practice

To be used by all students preparing for DP IB Chemistry HL
Students of other boards may also find this useful



1a

To show that the copper sulfate is the limiting reagent:

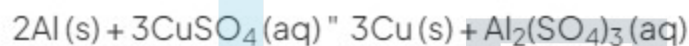
- Calculate the amount of copper sulfate in moles:

$$n(\text{CuSO}_4 \cdot 5\text{H}_2\text{O}) = \frac{2.00 \text{ g}}{249.72 \text{ g mol}^{-1}} = 8.01 \times 10^{-3} \text{ mol}; [1]$$

- Calculate the amount of aluminium in moles

$$n(\text{Al}) = \frac{0.25 \text{ g}}{26.98 \text{ g mol}^{-1}} = 9.27 \times 10^{-3} \text{ mol}; [1]$$

From the equation, the reacting ratio of Al : CuSO₄ is 2 : 3



So, 1 mole of CuSO₄ requires $\frac{2}{3}$ mole of Al for complete reaction

- 0.01×10^{-3} mol of CuSO₄ requires $\frac{2}{3} \times 8.01 \times 10^{-3}$ mol of Al = 5.34×10^{-3} mol; [1]

So, the aluminium is in excess and copper sulfate is the limiting reagent.

The mass of aluminium in excess is:

$$n(\text{Al})_{\text{excess}} = 9.27 \times 10^{-3} - 5.34 \times 10^{-3} = 3.93 \times 10^{-3} \text{ mol}$$

- mass of Al_{excess} = $3.93 \times 10^{-3} \text{ mol} \times 26.98 \text{ g mol}^{-1} = 0.106 / 0.11 \text{ g}; [1]$

Exam Papers Practice

[Total: 4 marks]

- Notice that we are calculating the number of moles of *hydrated* copper sulfate because the number of moles will be the same as the copper sulfate in the equation
- It is really important to show clear steps in your workings so the examiner can see how you have arrived at your answer
- Try to be consistent with significant figures and use the least number given in the data
 - The mass of the aluminium is given to two significant figures
 - However, since the aluminium is an excess reagent, its mass will not influence the mass of the product, so you could express the answer to three significant figures, just like the given mass of hydrated copper sulfate

1b

The actual and percentage yield are:

- The actual yield of copper is $42.61 - 42.18 \text{ g} = 0.43 \text{ g}$; [1]
- The expected yield, by mass, of copper is
Moles x molar mass = $8.01 \times 10^{-3} \text{ mol} \times 63.55 \text{ g mol}^{-1} = 0.509 \text{ g}$; [1]

The percentage yield is therefore:

- $\frac{0.43}{0.509} \times 100 = 84.5\%$; [1]

(Allow 2 or 3 significant figures as long as correctly rounded)

[Total: 3 marks]

- The expected yield of a product is based on the amount of the limiting reagent which is the copper sulfate in this experiment
- The copper sulfate and copper are in a 1:1 ratio so the amount (in moles) of copper is the same as the amount of copper sulfate
 - $2\text{Al}(\text{s}) + 3\text{CuSO}_4(\text{aq}) \rightarrow 3\text{Cu}(\text{s}) + \text{Al}_2(\text{SO}_4)_3(\text{aq})$
- Yield can be calculated by mass or by moles, but the formula to use is the same:
 - $\frac{\text{actual yield}}{\text{expected yield}} \times 100 = \% \text{ yield}$

1c

The percentage uncertainty in the mass of copper is:

- $\frac{\pm 0.02 \text{ g}}{0.43 \text{ g}} \times 100 = 6 \sim 5\%$; [1]

The overall percentage error for the experiment is:

- $\frac{0.509 \text{ g} - 0.43 \text{ g}}{0.509 \text{ g}} \times 100 = 16\%$; [1]

(Allow 15.5%)

[Total: 2 marks]

- To find the percentage uncertainty in a measurement we use the formula
 - $\frac{\text{absolute uncertainty}}{\text{actual measurement}} \times 100 = \% \text{ uncertainty}$
- Notice that we are using $\pm 0.02 \text{ g}$ and not $\pm 0.01 \text{ g}$ as the absolute uncertainty in the mass of copper, because **two** measurements were made to find the mass of copper (the empty beaker and the beaker plus copper) and so we must double the uncertainty when finding the difference between two values
- As a general rule, the calculated % uncertainty should **not** be quoted to:
 - More than one significant figure if it is greater or equal to 2% of the answer
 - More than two significant figures if it is less than 2%
- You should not confuse **percentage error** with **percentage uncertainty** for an experiment
 - Percentage error is how close the experimental value is to an accepted (or literature) value
 - $\frac{\text{accepted value} - \text{experimental value}}{\text{experimental value}} \times 100 = \% \text{ error}$
- Percentage errors are usually quoted to whole numbers
- Sometimes you may be asked to find the **overall** percentage uncertainty for an experiment – this is simply the sum of the individual measurement uncertainties
 - Measurements of excess reagents or solvents would not usually be included in the calculation as they do not affect the accuracy of an experimental value

1d

The impact on the % yield of the following systematic errors:

i) The copper is not fully dry:

- The mass of copper is larger / heavier than it should be so the percentage yield is higher than it should be; [1]

ii) The student adds 1.0 mL of hydrochloric acid:

- The question states that the acid is catalytic (and only provides chloride ions) so it will have no impact on the percentage yield of copper; [1]

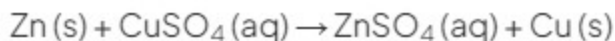
[Total: 2 marks]

- You should be able to discuss systematic errors in all experimental work, their impact on the results and how they can be reduced
- You should also be able to estimate whether a particular source of error is likely to have a major or minor effect on the final result (or none at all)

2a Exam Papers Practice

The limiting reagent is:

- $n(\text{CuSO}_4) = 0.0500 \text{ dm}^3 \times 0.250 \text{ mol dm}^{-3} = 0.0125 \text{ mol}$; [1]
- $n(\text{Zn}) = \frac{3.78 \text{ g}}{65.38 \text{ g mol}^{-1}} = 0.0578 \text{ mol}$; [1]
- From the equation, we can see the ratio of Zn: CuSO₄ is 1:1,



Therefore, the CuSO₄ is the limiting reactant; [1]

[Total: 3 marks]

- You cannot score the final mark if the workings are not correct
- If the ratio is not 1:1 and you are unsure how to determine the limiting reagent, an easy trick is to divide the moles by the coefficient for the substance
- Whichever gives the **LOWEST** number is the **LIMITING** reagent

2b

The enthalpy change for the reaction is:

- Heat evolved, $q = 50 \times 4.18 \times 8.5 = 1800 \text{ J} / 1.8 \times 10^3 \text{ J} / 1.8 \text{ kJ}$; [1]
- Per mol of $\text{CuSO}_4 = \frac{-1.8 \text{ kJ}}{0.0125 \text{ mol}} = -1.4 \times 10^2 \text{ kJ mol}^{-1}$; [1]

[Total: 2 marks]

- The calorimetry equation is given in section 1 of the data booklet
 - $q = mc\Delta T$
 - Make sure you know all the terms and units in the formula
- To determine the enthalpy change for the reaction you need to know the limiting reagent and divide q by the number of moles of the limiting reagent
 - $\Delta H = \frac{q}{n}$
- The specific heat capacity of water is given in section 2 of the data booklet
 - $c = 4.18 \text{ kJ kg}^{-1}\text{K}^{-1}$ or $4.18 \text{ J g}^{-1}\text{K}^{-1}$
 - Most of the time chemists use the g scale rather kg in the lab!
- Don't forget the sign for the enthalpy change - without the negative sign here you would lose the final mark
 - The question tells you the temperature rises, so it is an exothermic reaction which requires a negative sign in front of the value

2c

Two further assumptions made in the calculation of ΔH :

Any **two** from:

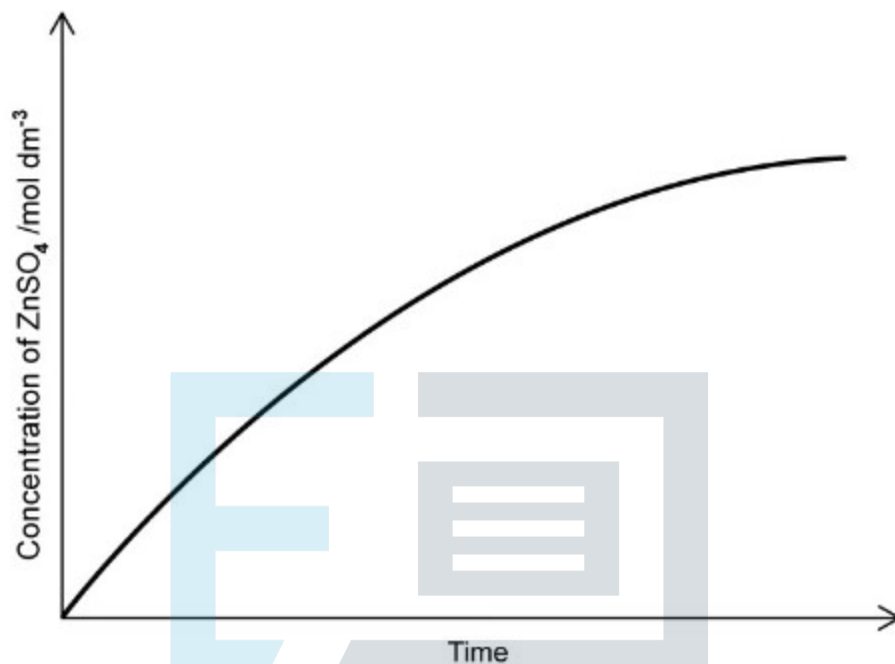
- The density of the solution is the same as water / 1.0 g cm^{-3} ; [1]
- The specific heat capacity of the solution is the same as water / $4.18 \text{ J g}^{-1} \text{ K}^{-1}$; [1]
- There is a complete reaction; [1]
(Allow only if CuSO_4 was identified as the limiting reactant)
- The zinc or copper sulfate does not react with any other substances; [1]

[Total: 2 marks]

- Heat loss is **not** allowed because the question states that you assume all the heat is absorbed by the solution
- Assuming 'a complete reaction' is only a valid answer if CuSO_4 was correctly identified as the limiting reactant, because the calculation in part (b) is based on the moles of copper sulfate
- In the calorimetry equation 'm' is the mass of the substance whose temperature is changing, which is usually water, but frequently students like to add the mass of the solid to the mass of the water to get a total mass
 - Unfortunately, this is not a correct approach, because the specific heat capacity is for water and no other added materials

2d

The sketch graph is as follows:



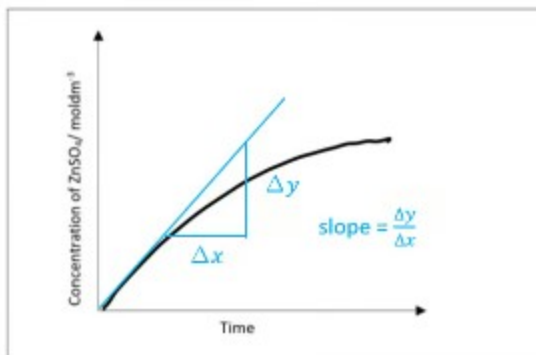
- The graph must show that the initial concentration of ZnSO₄ (aq) is 0
AND
Increases with time; [1]
- The gradient must decrease with time; [1]

To find the initial rate of reaction:

- Draw a tangent to the curve at time t = 0; [1]
- The initial rate of reaction is equal to the slope / gradient of the tangent; [1]

[Total: 4 marks]

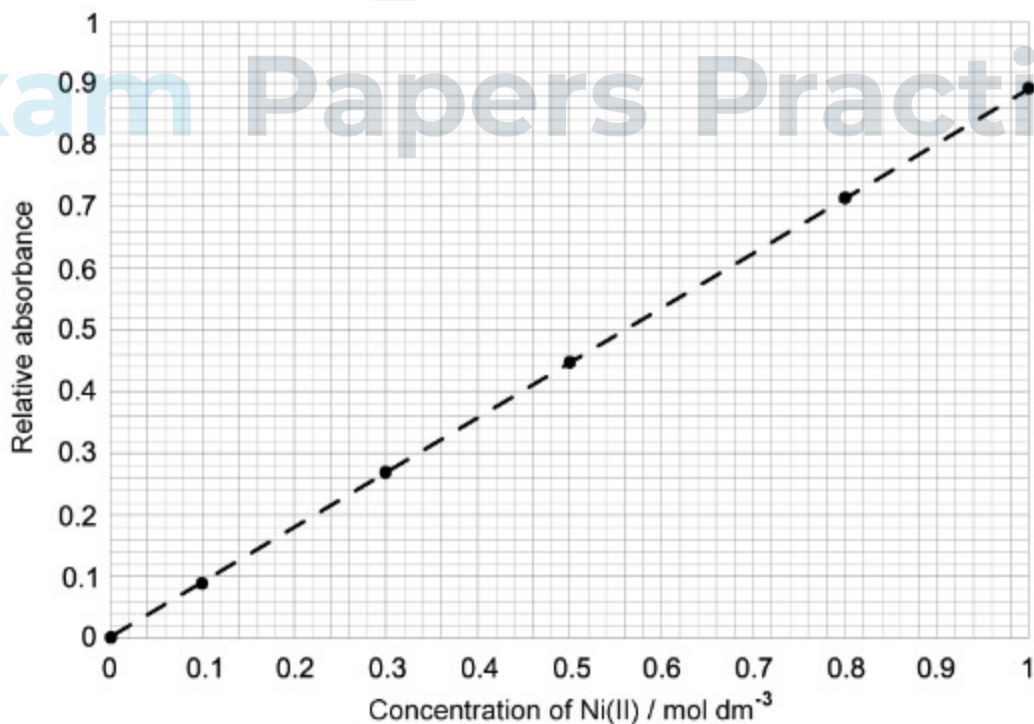
- A suitable diagram instead of a description could be as follows:



- Sketch graphs have labelled but unsealed axes, and are used to show qualitative trends, such as variables that are proportional or inversely proportional

3a

The graph of absorbance against concentration is as follows:





- Correct label, including units, and scale for the x-axis; [1]
- Correct label and scale for the y-axis; [1]
- Accurate plotting of points to within $\pm\frac{1}{2}$ square; [1]
- Correctly drawn trend line; [1]

(Axes reversed would score a maximum of 3 marks)

(Allow graphs that don't start at 0,0)

[Total: 4 marks]

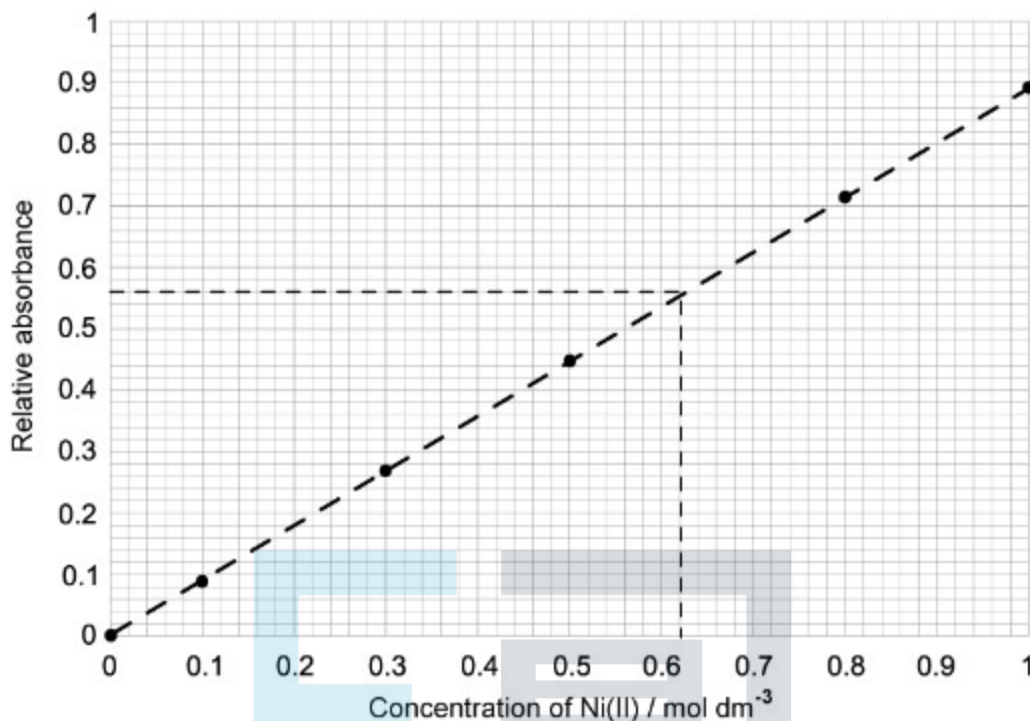
- Graphical techniques are an effective means of communicating the effect of an independent variable on a dependent variable, and can lead to determination of physical quantities
- Drawn graphs have labelled and scaled axes, and are used in quantitative measurements
- You should be able to draw graphs of experimental results including the correct choice of axes and scale
- A rule of thumb is to make sure your scale covers half or more of the graph paper
- Always check if 0,0 is a point as it will help to anchor any trend line

3b

Exam Papers Practice

The variables are:

- Independent: concentration of the Ni(II) ions
AND
Dependent: relative absorbance; [1]
- The concentration is $62 - 0.64 \text{ mol dm}^{-3}$; [1]



- Tie lines shown on the graph; [1]

[Total: 3 marks]

- The **independent variable** is the one that you control or manipulate, and the **dependent variable** is the one that changes as a result of your manipulation
- In this experiment the concentration of the nickel(II) ions is determined by the experimenter so that is the independent variable
- By convention, the independent variable is plotted on the x-axis and the dependent variable on the y-axis
- All other variables that may affect the result are called **controlled variables**
- Use a ruler to draw dotted tie lines to join the axes to the trend line in order to read off accurate values of the unknown data

3c

i) The relationship between the variables is:

- (They are) directly proportional; [1]

ii) Determining the value of m:

- 'm' is the slope (gradient) of the graph

$$\text{Gradient} = \frac{\Delta y}{\Delta x} = \frac{0.890 - 0.000}{1.0 - 0.0} = 0.890 \text{ mol}^{-1}\text{dm}^3; [1]$$

(The correct units must be shown)

(Allow other pairs of values of x and y to find m)

iii) Calculating the concentration of Ni(II) ions:

- $x = \frac{y}{m} = \frac{0.560}{0.890} = 0.63 \text{ mol dm}^{-3}; [1]$

(c = y-intercept = 0)

- The two values are the same; [1]

[Total: 4 marks]

- m is the gradient of the straight line and can be determined graphically or by calculation and you should choose whichever is easiest as long as the question gives you the choice
- You can find the gradient at any position along a straight-line graph, but if the trend line starts at the origin, as this graph does, you can use the maximum value of y and x to calculate the slope
 - The reason for this is that you will minimise the inaccuracy of the graph value by making the largest gradient triangle possible

3d

The relationship is only linear at low concentration because:

- At higher concentrations the solution is darker / more strongly coloured so a higher proportion of light is absorbed
- OR
- At higher concentration very little light passes through the solution
- OR
- Limitations of the colorimeter / spectrometer to detect very low light levels for higher concentrations;
- OR
- Interactions between the ions / molecules of the solvent at high concentrations; [1]

[Total: 1 mark]

- The relationship between concentration and absorbance is known as the Beer-Lambert Law and although you don't need to know the law, you may have to use it in graph-based questions
- It can be particularly useful in colorimetry experiments, but also in determining very low concentrations of heavy metal ions in water or blood using atomic absorption spectroscopy

4a

i) The reacting ratio of $\text{S}_2\text{O}_3^{2-}$ to O_2 is:

- 4:1; [1]

ii) The titre and percentage uncertainty is:

- $26.0 \text{ cm}^3 - 0.20 \text{ cm}^3 = 25.8 \text{ cm}^3$; [1]
- $\frac{\pm 0.2 \text{ cm}^3}{25.8 \text{ cm}^3} \times 100 = 0.8\%$; [1]

**[Total: 3 marks]**

- Working out the reacting ratios from multiple equations can be tricky – just work backwards
 - Imagine you started with 1 mol of $S_2O_3^{2-}$ how much I_2 would that react with?



- The answer is 0.5 mol
- Next take 0.5 mole of I_2 as the product in step 2, how much MnO_2 was needed to make it?



- The answer is 0.5 mol
- Finally take 0.5 mole of MnO_2 as the product in step 1, how much O_2 was needed to make it?



- The answer is 0.25 mol
- So, that means 0.25 mol of O_2 is equivalent to 1 mol of $S_2O_3^{2-}$ or 1 mol of O_2 to 4 mol of $S_2O_3^{2-}$
- Be careful with the absolute uncertainty on burette readings
 - It takes two readings to calculate the titre, so the uncertainty must be doubled
- To find the percentage uncertainty in a measurement we use the formula

$$\frac{\text{absolute uncertainty}}{\text{actual measurement}} \times 100 = \% \text{ uncertainty}$$



4b

The amount of oxygen present in the sample is:

- $n(\text{S}_2\text{O}_3^{2-}) = 0.0258 \text{ dm}^3 \times 0.0120 \text{ mol dm}^{-3} = 3.10 \times 10^{-4} \text{ mol}; [1]$
- $n(\text{O}_2) = \frac{3.10 \times 10^{-4} \text{ mol}}{4} = 7.75 \times 10^{-5} \text{ mol}; [1]$

The concentration in ppm is:

Convert the moles into mg:

$$= 7.75 \times 10^{-5} \text{ mol} \times 32.00 \text{ g mol}^{-1} \times 1000 \text{ mg g}^{-1} = 2.48 \text{ mg}$$

Divide the mass in mg by the volume in dm^3 :

- $\frac{2.48}{0.500 \text{ dm}^3} = 4.96 \text{ ppm}; [1]$

[Total: 3 marks]

- The unit ppm stands for parts per million and is used to express concentration where the concentrations are extremely low, such as in air or water pollution
- It can be expressed in the following ways:
 - One part per 10^6 by mass
 - 1mg of solute is dissolved in 1 dm^3 of solution
 - 1mg of solute is dissolved in 1 kg of solution

4c

Should the student be concerned about the levels of oxygen in the fish tank over a 5-day period:

- Yes / the student should be concerned

AND

The levels have dropped below what is sufficient for aquatic life / the water no longer contains enough oxygen after 5 days / the amount of oxygen after 5 days is too low / insufficient; [1]

The concentration at day 5 in ppm is:

- Convert the moles into mg:

$$= (5.03 \times 10^{-5}) \times 32.00 \times 1000 = 1.61 \text{ mg}$$

Divide the mass in mg by the volume in dm^3 :

$$\frac{1.61 \text{ mg}}{0.500 \text{ dm}^3} = 3.22 \text{ ppm}; [1]$$

[Total: 2 marks]

The comment must be linked to the calculation to justify what has happened to the oxygen levels over the 5-day period

4d

A modification that would make the result more reliable:

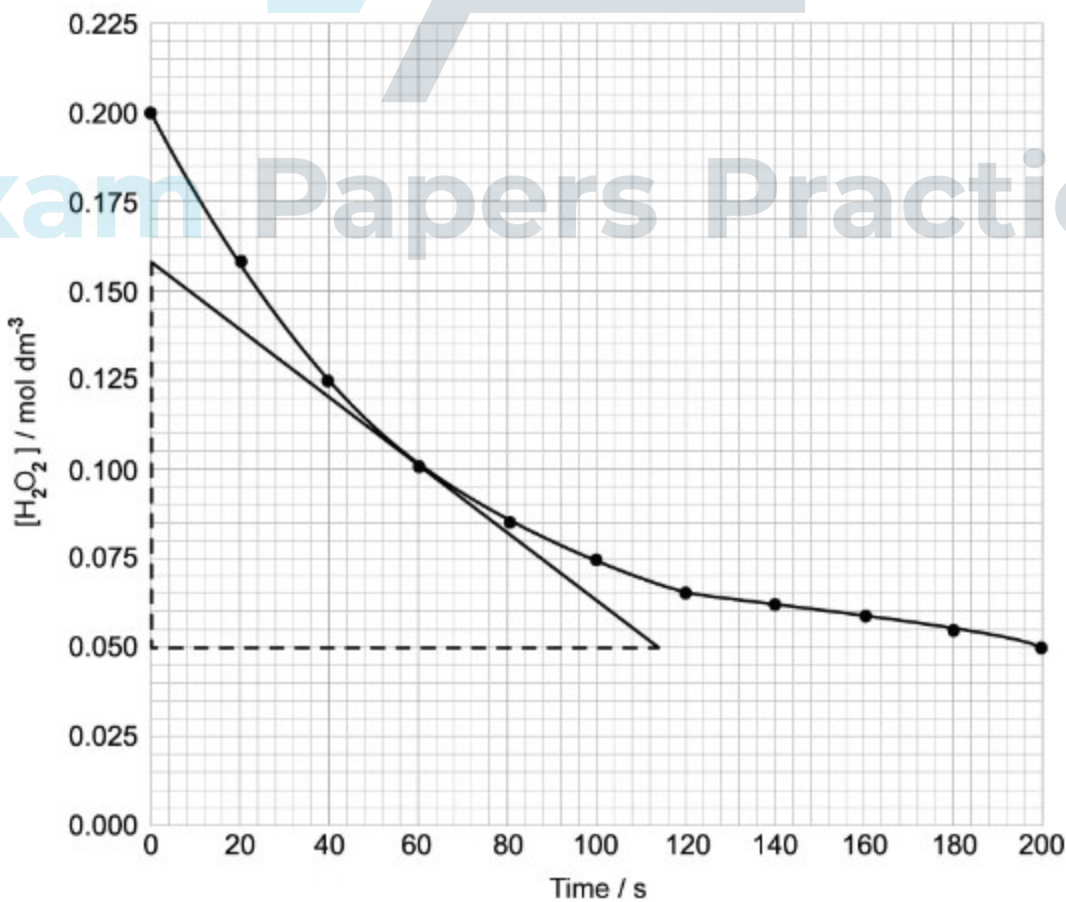
- Several samples should be taken / the procedure should be repeated (and an average taken); [1]

[Total: 1 mark]

- Reliability can only be achieved by repeating a procedure to obtain more results
- Repetition allows you to compensate for the presence of random errors in the data
- Random errors could be caused by things like
 - How easily the measuring instrument can be read
 - An observer misinterpreting the reading
 - Changes in the surrounding, e.g., fluctuations in temperature or air currents
- If you spot a random error in your data, you have two choices:
 - Ignore the data point in your calculations (or graph trend line)
 - Include it and assume its impact will be small when you take averages

5a

The graph of concentration versus time is as follows:



- The points are correctly plotted; [1]
- A smooth best-fit line is drawn; [1]
- A tangent is drawn at 60 s; [1]

Slope (gradient) of the tangent to the line:

- $\text{gradient} = \frac{\Delta y}{\Delta x} = \frac{0.157 - 0.050}{113 - 0} = 9.5 \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1}$; [1]

(Allow a range $(9.3 - 9.5) \times 10^{-4} \text{ mol dm}^{-3} \text{ s}^{-1}$)

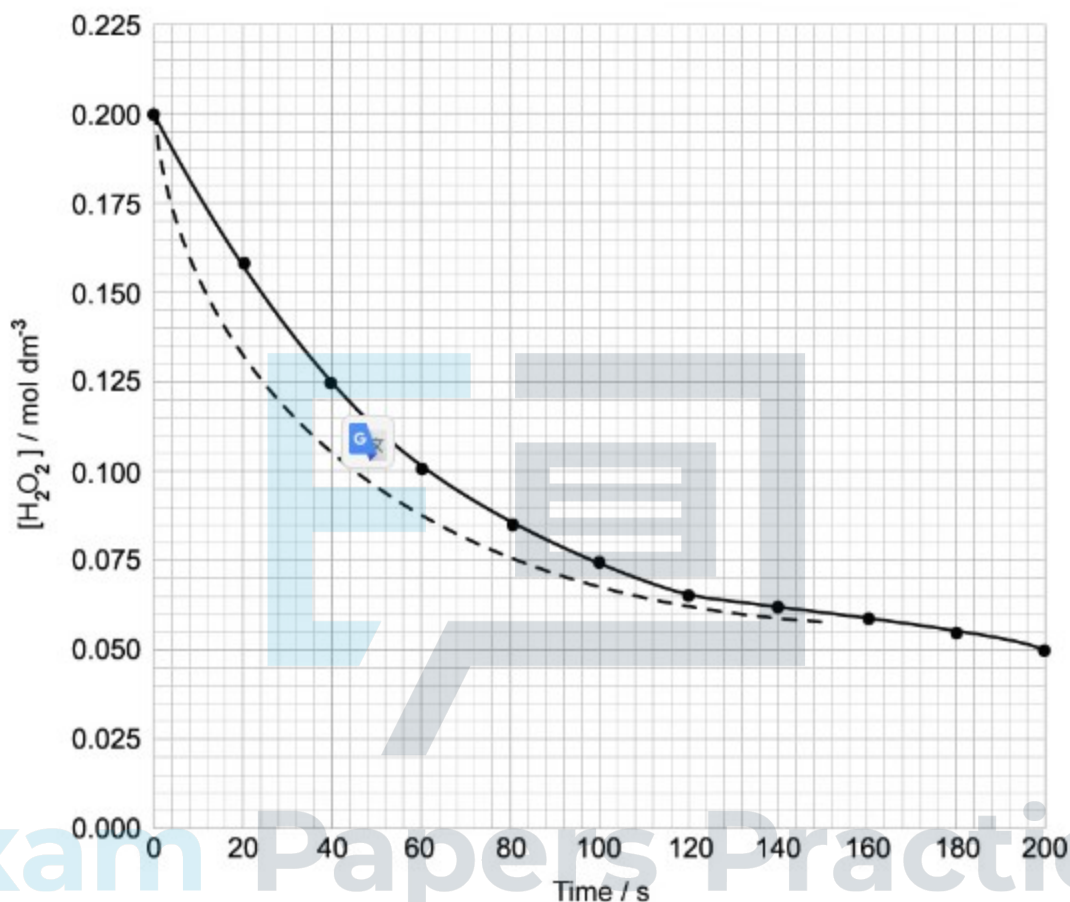
[Total: 4 marks]

- Make sure best-fit lines are single thin lines that aim to go through the majority of the points
 - If you can't go through all the points, try to ensure an even number of points on either side of the line
- Without the correct unit you would lose the final mark

Exam Papers Practice

5b

The shape at 60 °C is:



- Correctly drawn slope that is steeper **AND** levels out at approximately 0.050 mol dm⁻³ H₂O₂; [1]
- The rate of reaction is faster at a higher temperature so the gradient must be steeper; [1]
- The concentration of H₂O₂ is the same/ the quantities have not changed, so the graph must finish at the same level;
OR
 There will be more effective collisions are at higher temperature;
OR
 More particles have energy greater than the activation energy at 60 °C; [1]

[Total: 3 marks]

- Students sometimes lose marks by not getting the final shape of a rate sketch graph correct
- If the only factor you are changing is temperature, surface area or a catalyst then the finishing position on the graph will not change

5c

i) The reason the volume of oxygen can be used to measure the concentration of H_2O_2 :

- The equation shows that 2 mol of H_2O_2 liberates 1 mol of O_2 , so the number of moles of O_2 is proportional to the amount of H_2O_2 remaining
- OR**
- Volumes of gases are proportional to the number of moles of gases (according to Avogadro's hypothesis); [1]

ii) One limitation of the apparatus:

Any **one** from:

- The syringe holds a limited amount of gas / only 100 mL of gas can be collected; [1]
- The gas could leak from the apparatus; [1]
- Some gas could escape before the stopper is connected; [1]
- The barrel of the syringe could stick (giving a false reading); [1]

iii) An alternative method of measuring the reaction rate:

- Measure the change in mass (using a balance) against time;
- OR**
- Collect the gas through water using an inverted measuring cylinder;
- [1]

[Total: 3 marks]

- You should be able to discuss how limitations in experimental design and procedure can lead to systematic errors in measurements and be able to suggest alternative ways of carrying out procedures using normal laboratory apparatus
- Avoid inventing apparatus you have never seen or suggesting unrealistic procedures as you are unlikely to impress the examiners
- Diagrams can be a good way to show how you intend to carry out an experiment
- The more information and clarity you can give the more likely you will hit the target

5d

Which student is correct:

- The second student is correct

AND

Because a mass reading which has three decimal places is more precise than a reading with only two decimal places; [1]

- The accuracy of the result is how close it is to an accepted value;

OR

The balance could be inaccurate / faulty even with three decimal places;

OR

The accuracy is limited by the method / procedure not by the number of decimal places; [1]

[Total: 2 marks]

- Higher precision is achieved by having smaller random uncertainties, which would be the case when you use a three decimal place balance rather than a two decimal place balance
- Accuracy on the other hand is affected by systematic errors, that is, errors in the procedure or method
 - You cannot eliminate systematic errors by repetition