

AQA Chemistry A-level

Practical skills: Uncertainties



Sources of uncertainties

Every measurement that is made will have some **inherent uncertainty**. It is important to consider if the true value lies in the range that is outlined by the quoted uncertainty. Experiments should be designed with an **attempt to reduce the uncertainty** in each value and therefore the outcome of the experiment. Procedures should be used so that individual uncertainties are minimised in order to provide a realistic value and uncertainty for the outcome.

When assessing uncertainty, a number of issues have to be considered. These include:

- Resolution of instruments
- Manufacturer's tolerance on these instruments
- Judgments made by the person carrying out the experiment
- Adopted procedures such as repeated readings
- The size of increments available such as the size of drops from a pipette.

Numerical questions require a number of these factors to be considered. **Resolution** is usually the key factor that has to be considered when assessing a numerical uncertainty.

Longer questions often require an **evaluation** of arrangements and procedures to assess how uncertainty could be reduced. This can include **discussions** on the use of different apparatus design or procedure. The effect of these changes on experimental results can be different depending on the specific experimental situation meaning only the experimenter can assess their impact and usefulness.

Absolute and percentage uncertainties

Absolute uncertainty: the **true numerical value** (often with units), indicating the range in which the true value lies.

Percentage uncertainty: this expresses the absolute uncertainty as a **proportion of the measurement**. It is sometimes necessary to calculate percentage uncertainty so that the total uncertainty (in a value with multiple variables) can be found. (see combining uncertainties).



Readings and measurements

When discussing uncertainties, measurements should be separated into two forms:

- Readings: the values found from a **single judgement** when using a piece of equipment.
- Measurements: the values taken as the **difference between the judgements** of two values.

Readings

Uncertainty in a reading for a specific instrument is said to be **plus or minus half of the smallest division**.

Example: Thermometer

Only one judgement has to be made (the height of the liquid). Therefore it is a reading and it can be assumed that the zero value has been correctly set.

A temperature measured using a thermometer with graduations of 1°C is said to have an uncertainty of $\pm 0.5^{\circ}\text{C}$.

It is common to write down values with the uncertainty, for example, $2.40 \pm 0.01 \text{ V}$. It is normally quoted to the **same number of decimal places** as the value.

Measurements

Both of the uncertainties in a measurement must be taken into account; the uncertainty of the **zero marker** and the uncertainty of the **measurement point**.

Example: Burettes and rulers

Both the starting point and the end point of the measurement must be judged. This leads to two uncertainties overall, making it a measurement.

Both ends of the ruler have a ± 0.5 scale division uncertainty, therefore the measurement will have an overall uncertainty of ± 1 division.

For most rulers, the uncertainty in a measurement of length will be $\pm 1 \text{ mm}$.

Summary

- The uncertainty of a **reading** (one judgement) is at least **± 0.5 of the smallest scale reading**.
- The uncertainty of a **measurement** (two judgements) is at least **± 1 of the smallest scale reading**.





Experimental method

The **method** by which measurements are taken can also affect the uncertainty.

In some situations, **resolution** of an instrument is not the limiting factor in the uncertainty of a measurement.

It is considered best practice to **write down the full reading** and then to write to **fewer significant figures** when the uncertainty has been estimated.

Examples:

1. A stopwatch has a resolution of hundredths of a second, but the measurement is likely to be more affected by the **reaction time** of the experimenter.
Therefore the full reading on the stopwatch (eg 12.20 s) should be recorded, carrying all the significant figures for all repeats. The number of significant figures can then be reduced to a **more appropriate value** during an averaging process later on.
2. If a digital measuring instrument is being used, the **uncertainty cannot be taken from the last figure in the readout** as it is not always the full value. Manufacturers quote the percentage uncertainties for the different ranges of that specific instrument.

Uncertainties in given values

In exams, it can be assumed that the uncertainty in a given value is **± 1 the last significant figure**.

Example: Boiling points

A boiling point quoted as being 78 °C can be assumed to have an uncertainty of ± 1 °C. It could be lower than this but without knowing experimental details, there is no evidence to assume otherwise.

Repeated measurements

Repeating a measurement is a good method for **reducing the uncertainty**.

In some experiments, it allows **anomalies** (values that lie far away from the general trend of the other measurements) to be identified more easily. These outlying results can then be **removed from the data set** before calculating a **mean**.

On other occasions, **outliers can be important to include** and so uncertainty can be found using a different method. If measurements are repeated, the uncertainty can be calculated by finding **half the range of the measured values**.

Example: Distance

Repeat	1	2	3	4
Distance/m	1.23	1.32	1.27	1.22

$$1.32 - 1.22 = 0.10 \text{ therefore}$$

$$\text{Mean distance: } (1.26 \pm 0.05) \text{ m}$$





From this, **percentage uncertainty** can be found by **dividing the uncertainty by the mean distance** and **multiplying it by one hundred**.

The method required to be used in an exam question will be made clear in the wording of the question.

Titration:

Titration requires a **number of factors** to be considered to find the uncertainty in the measurement. Measuring one single titre is **not sufficient**. It should be repeated until at least two **concordant titre volumes** have been recorded (within 0.10 cm^3). These values can then be averaged to find the mean titre volume.

Example: Titre

Titration	Rough	1	2	3
Final reading	24.20	47.40	24.10	47.35
Initial reading	0.35	24.20	0.65	24.10
Titre/ cm^3	23.85	23.20	23.45	23.25

Titres 1 and 3 are concordant so can be averaged to find a **mean titre volume** of 23.23 cm^3 . The total error in a titre is a result of **three factors**:

Error	Uncertainty
Reading the burette at the start of the titration	Half a division = $\pm 0.05 \text{ cm}^3$
Reading the burette at the end of the titration	Half a division = $\pm 0.05 \text{ cm}^3$
Judging the end point to within one drop	Volume of a drop = $\pm 0.05 \text{ cm}^3$
Total	$\pm 0.15 \text{ cm}^3$

This depends on the glassware used, as some burettes are **calibrated to a higher accuracy** than others.





Uncertainties in exams

In exam questions, it will be made clear whether the question is asking for the uncertainty of a reading, a measurement, or given data to be calculated.

Mark schemes will allow **sensible alternative answers** and will **credit clear working methods**.

It can be helpful to read the **examiner report** to see the common mistakes that are made to help **improve answers** to practical questions.

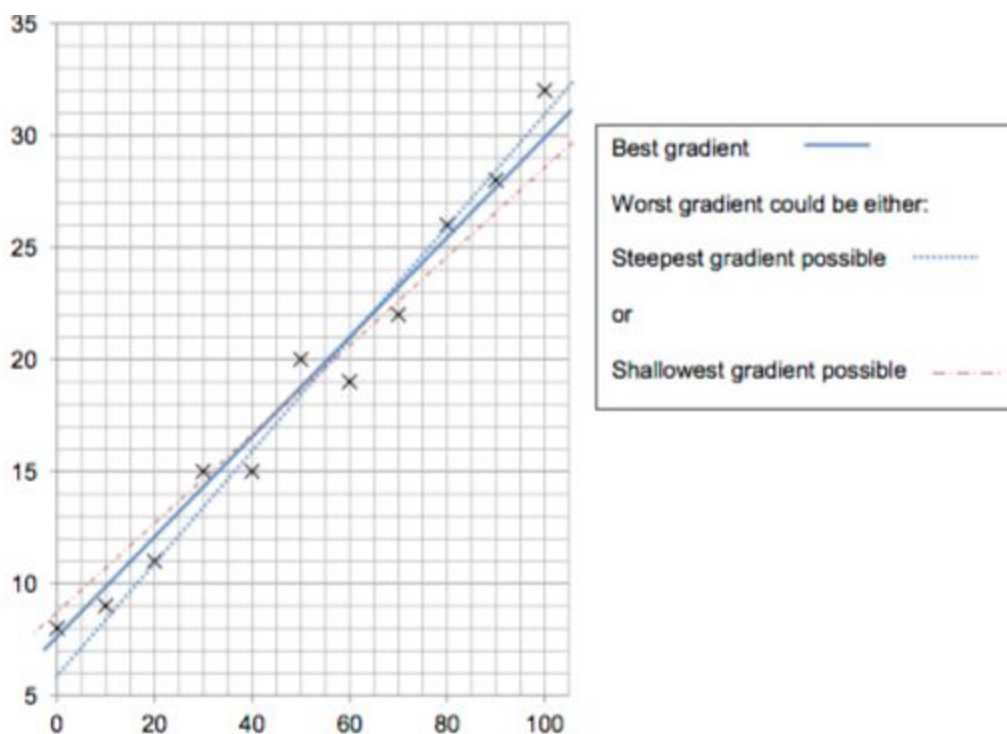
Uncertainties from gradients

Uncertainty in a gradient can be found using **two lines drawn on the graph**:

- The **“best”** line of best fit.
- Either the **steepest or shallowest gradient** line of best fit possible from the data.

The gradient of each line should then be found, and the uncertainty calculated using:

$$\text{percentage uncertainty} = |\text{best gradient} - \text{worst gradient}| \times 100\%$$



The **modulus bars** in this equation mean that this value will always be **positive**.

Percentage uncertainty in the **y-intercept** can also be found using the same method with the varying intercept values.





Combining uncertainties

When **adding or subtracting** values, the individual **absolute uncertainty** in each value is **added** to find the total uncertainty.

When **multiplying or dividing** values, the individual **percentage uncertainty** in each value is **added** to find the total uncertainty.

When a value is **raised to a power**, the **percentage uncertainty** is **multiplied** by the **same value as the power**.

Summary

(taken from the AQA Chemistry Practical Handbook)

Combination	Operation	Example
Adding or subtracting values $a = b + c$	Add the absolute uncertainties $\Delta a = \Delta b + \Delta c$	Initial volume in burette = $3.40 \pm 0.05\text{cm}^3$ Final volume in burette = $28.50 \pm 0.05\text{cm}^3$ Titre = $25.10 \pm 0.10\text{cm}^3$
Multiplying values $a = b \times c$	Add the percentage uncertainties $\epsilon a = \epsilon b + \epsilon c$	Mass = $50.0 \pm 0.1\text{ g}$ Temperature rise (T) = $10.9 \pm 0.1^\circ\text{C}$ Percentage uncertainty in mass = 0.20% Percentage uncertainty in T = 0.92% Heat change = 2278 J Percentage uncertainty in heat change = 1.12% Absolute uncertainty in heat change = $\pm 26\text{ J}$ (Note – the uncertainty in specific heat is taken to be zero)
Dividing values $a = \frac{b}{c}$	Add the percentage uncertainties $\epsilon a = \epsilon b + \epsilon c$	Mass of salt in solution = $100 \pm 0.1\text{ g}$ Volume of solution = $250 \pm 0.5\text{cm}^3$ Percentage uncertainty in mass = 0.1% Percentage uncertainty in volume = 0.2% Concentration of solution = 0.400 g cm^{-3} Percentage uncertainty of concentration = 0.3% Absolute uncertainty of concentration = $\pm 0.0012\text{ g cm}^{-3}$
Power rules $a = b^c$	Multiply the percentage uncertainty by the power $\epsilon a = c \times \epsilon b$	Concentration of H^+ ions = $0.150 \pm 0.001\text{ mol dm}^{-3}$ rate of reaction = $k[\text{H}^+]^2 = 0.207\text{ mol dm}^{-3}\text{ s}^{-1}$ (Note – the uncertainty in k is taken as zero and its value in this reaction is $0.920\text{ dm}^6\text{ mol}^{-2}\text{ s}^{-1}$) Percentage uncertainty in concentration = 0.67% Percentage uncertainty in rate = 1.33% Absolute uncertainty in rate = $\pm 0.003\text{ mol dm}^{-3}\text{ s}^{-1}$

