Significant figures

Data should be written in tables to the same number of significant figures. This number should be determined by the resolution of the device being used to measure the data or the uncertainty in measurement. For example, a sample labelled as “1 mol dm^{-3} acid” should not be recorded in a table of results as 1.0 mol dm^{-3}.

There is sometimes confusion over the number of significant figures when readings cross multiples of 10. Changing the number of decimal places across a power of ten retains the number of significant figures **but changes the accuracy.** The same number of decimal places should therefore generally be used, as illustrated below.

0.97
0.98
0.99
1.00
1.10

99.7
99.8
99.9
100.0
101.0

It is good practice to write down all digits showing on a digital meter.

Calculated quantities should be shown to the number of significant figures of the data with the least number of significant figures.

Example:

Calculate the concentration, in mol dm^{-3}, of a solution of sodium hydroxide that contains 0.28 mol of NaOH in 465 cm^{3} of water.

$$\text{Concentration} = \frac{0.28}{475} \times 1000 = 0.59$$

Note that the concentration can only be quoted to two significant figures as the number of moles is only quoted to two significant figures.

Uncertainties

Students should know that every measurement has some inherent uncertainty.

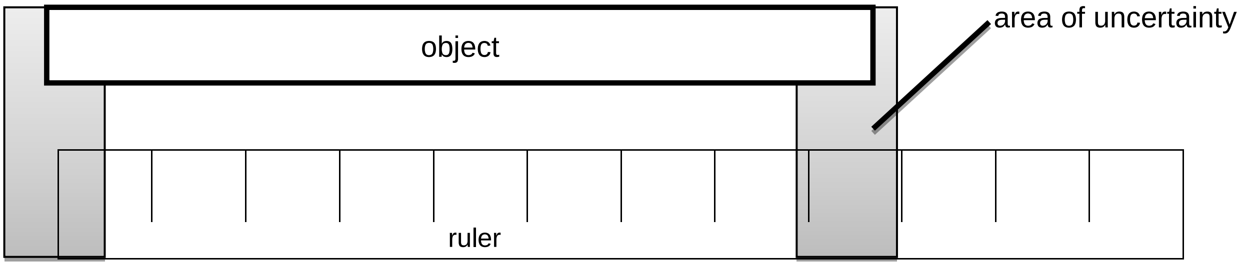
The uncertainty in a measurement using a particular instrument is no smaller than plus or minus half of the smallest division or greater. For example, a temperature measured with a thermometer is likely to have an uncertainty of ±0.5 °C if the graduations are 1 °C apart.

Students should be aware that measurements are often written with the uncertainty. An example of this would be to write a voltage was (2.40 ± 0.005) V.

Measuring length

When measuring length, **two** uncertainties must be included: the uncertainty of the placement of the zero of the ruler and the uncertainty of the point the measurement is taken from.

As both ends of the ruler have a ±0.5 scale division uncertainty, the measurement will have an uncertainty of ±1 division.



For most rulers, this will mean that the uncertainty in a measurement of length will be ±1 mm.

Other factors

There are some occasions where the resolution of the instrument is not the limiting factor in the uncertainty in a measurement.

Best practice is to write down the full reading and then to write to a fewer significant figures when the uncertainty has been estimated.

Examples:

A stop watch has a resolution of hundredths of a second, but the uncertainty in the measurement is more likely to be due to the reaction time of the experimenter. Here, the student should write the full reading on the stop watch (eg 12.20 s) and reduce this to 12 s later.

If a student measures the length of a piece of wire, it is very difficult to hold the wire completely straight against the ruler. The uncertainty in the measurement is likely to be higher than the ±1 mm uncertainty of the ruler. Depending on the number of “kinks” in the wire, the uncertainty could be reasonably judged to be nearer ± 2 or 3 mm.

Repeated measurements

If measurements are repeated, the uncertainty can be calculated by finding half the range of the measured values.

For example:

Repeat	1	2	3	4
Time (s)	12.3	13.2	12.7	12.2

13.2 – 12.2 = 01.0 so
Mean time: (12.6 ± 00.5) s

Percentage uncertainties

The percentage uncertainty in a measurement can be calculated using:

$$percentage\ uncertainty = \frac{uncertainty}{value}\ x\ 100\%$$

The percentage uncertainty in a repeated measurement can be calculated using:

$$percentage\ uncertainty = \frac{uncertainty}{mean\ value}\ x\ 100\%$$

Titration

Titration is a special case where a number of factors are involved in the uncertainties in the measurement.

Students should carry out a rough titration to determine the amount of titrant needed. This is to speed up the process of carrying out multiple samples. The value of this titre should be ignored in subsequent calculations.

In titrations one single titre is never sufficient. The experiment is usually done until there are at least two titres that are concordant ie within a certain allowable range, often 0.10 cm^3. These values are then averaged.

For example:

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

Here, titres 1 and 3 are within the allowable range of 0.10 cm^3 so are averaged to 23.23 cm^3.

Unlike in some Biology experiments (where anomalous results are always included unless there is good reason not to), in Chemistry it is assumed that repeats in a titration should be concordant. If they are not then there is likely to have been some experimental error. For example the wrong volume of solution added from the burette, the wrong amount of solution measuring the pipette or the end point might have been misjudged.

The total error in a titre is caused by three factors:

Error	**Uncertainty**
Reading the burette at the start of the titration	Half a division = ±0.05 cm^3
Reading the burette at the end of the titration	Half a division = ±0.05 cm^3
Judging the end point to within one drop	Volume of a drop = ± 0.05 cm^3
Total	± **0.15 cm^3**

This will, of course, depend on the glassware used, as some burettes are calibrated to a higher accuracy than others.

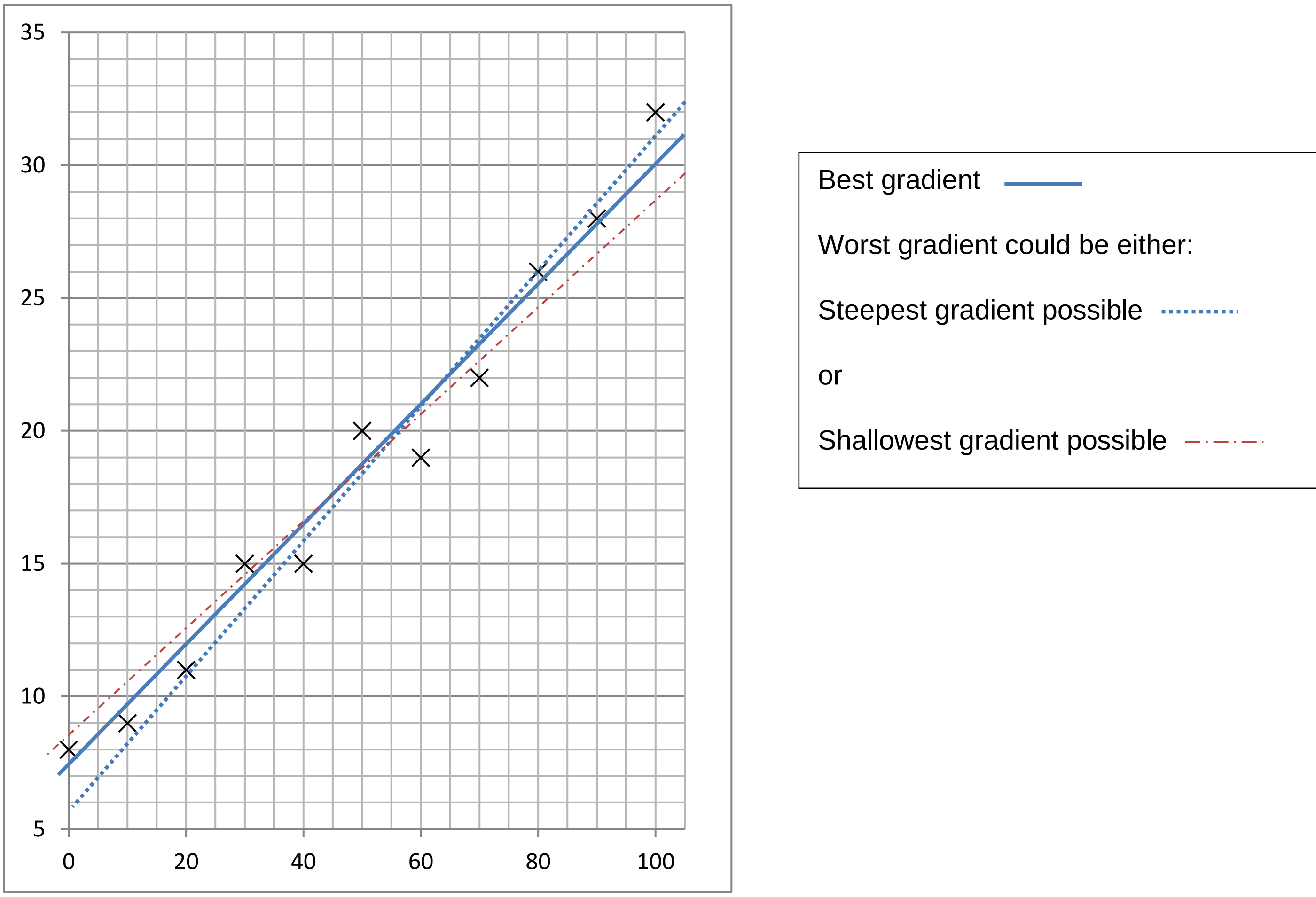
Uncertainties from gradients

To find the uncertainty in a gradient, two lines should be drawn on the graph. One should be the “best” line of best fit. The second line should be the steepest or shallowest gradient line of best fit possible from the data. The gradient of each line should then be found.

The uncertainty in the gradient is found by:

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

Note the modulus bars meaning that this percentage will always be positive.



In the same way, the percentage uncertainty in the y-intercept can be found:

$$\text{percentage uncertainty} = \frac{|\text{best } y\ intercept - \text{worst } y\ intercept|}{\text{best } y\ intercept} \times 100\%$$

Combining uncertainties

Percentage uncertainties should be combined using the following rules:

Combination	Operation	Example
Adding or subtracting values $\boldsymbol{a = b + c}$	Add the absolute uncertainties Δa = Δb + Δc	Initial volume in burette = 3.40 ± 0.05 cm^3 Final volume in burette = 28.50 ± 0.05 cm^3 Titre = 25.10 ± 0.10 cm^3
Multiplying values $\boldsymbol{a = b \times c}$	Add the percentage uncertainties εa = εb + εc	Mass = 50.0 ± 0.1 g Temperature rise (T) = 10.9 ± 0.1 ^oC 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 = ± 26 J (Note – the uncertainty in specific heat is taken to be zero)
Dividing values $\boldsymbol{a = \frac{b}{c}}$	Add the percentage uncertainties εa = εb + εc	Mass of salt in solution= 100 ± 0.1 g Volume of solution = 250 ± 0.5 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 = ± 0.0012 g cm^{-3}
Power rules $\boldsymbol{a = b^c}$	Multiply the percentage uncertainty by the power εa = c × εb	Concentration of H^+ ions = 0.150 ± 0.001 mol dm^{-3} rate of reaction = $k[H^+]^2$ = 0.207 mol dm^{-3} s^{-1} (Note – the uncertainty in *k* is taken as zero and its value in this reaction is 0.920 dm^6 mol^{-2} s^{-1}) Percentage uncertainty in concentration = 0.67 % Percentage uncertainty in rate = 1.33 % Absolute uncertainty in rate = ± 0.003 mol dm^{-3} s^{-1}

Note: Absolute uncertainties (denoted by Δ) have the same units as the quantity.
Percentage uncertainties (denoted by ε) have no units.
Uncertainties in trigonometric and logarithmic functions will not be tested in A-level exams.