IB Chemistry
Errors and Uncertainties

Random Uncertainties

These occur due to the limitations of the apparatus used or they are predictions due to the accuracy of the readings taken.

Uncertainties therefore are given a numerical (±) value and are included in the **data collection** (DC) and then processed in the **data processing** (DPP).

Due to the (±) nature, uncertainties are usually called **random** as they can lead to the experimental answer being either above or below the literature/actual value.

For some apparatus the uncertainty is a standard value.

E.g.   Burette reading ± 0.05 cm^3
         Pipette ± 0.05 cm^3
         Volumetric flask ± 0.6 cm^3
         Electronic balance to 2 dp ± 0.005g
         Thermometer (0-110^{0}C) ± 1^{0}C

Sometimes the uncertainty must be predicted at a reasonable level depending of the circumstances.

E.g. Using a pipette to measure the volume of a carbonated liquid the uncertainty may increase to ± 0.2 cm^3

Reading off the burette twice to give the volume in a titration makes the accuracy ± 0.1 cm^3 but this may also depend on obtaining concordant results. If the nearest 2 titrations are 22.40 and 22.80 cm^3 then the uncertainty may well be predicted to be ± 0.2cm^3.

A temperature change involves 2 thermometer readings and so should be at least be ± 2^{0}C accurate, however, measuring a melting point depends on the accuracy of the thermometer used, how quickly the substance was heated and the nature of the apparatus and so could quite reasonably rise to ± 3 or 4^{0}C.

Measuring cylinders depends on the size used. For example a 10cm^3 measuring cylinder may have an uncertainty of ± 0.1 cm^3 whereas a 50cm^3 measuring cylinder this value could rise to ± 0.5 cm^3.

Uncertainties should clearly accompany the readings taken. They may be included in the table for the dependant and independent variables.
E.g.

		Rough	1	2	3	Mean
Burette Reading: Volume of HCl / cm^3 $\Delta vol. = \pm 0.05cm^3$	Initial	0.00	20.00	0.00	0.00	
	Final	20.00	41.05	21.10	20.85	
Titre: Actual Volume of HCl used / cm^3 $\Delta vol. = \pm 0.1cm^3$		20.00	21.05	21.10	20.85	21.08

They should accompany any control variables.

E.g. Volume of NaOH used = 25. 00 cm^3 ± 0.05 cm^3 (from pipette)
Volume of phenolphthalein indicator used = 4 drops ± 1 drop.

Systematic Errors

Errors are used to explain why the experimental value is either above or below the literature value. These are normally described in the evaluation of the experiment. Since they result in a consistently high or low reading then they are called **systematic**. Although the total size of the error can be estimated (i.e. 2% below the literature value) the errors are normally explained qualitatively (e.g. misinterpretation of colour change etc)

Propagation of Uncertainties in Data Processing.

This is a statement made by the IB when discussing the propagation of uncertainties through a calculation.

Standard level candidates are *not* *expected to process uncertainties in calculations. However, they can make statements about the minimum uncertainty, based on the least significant figure in a measurement, and can also make statements about the manufacturer's claim of accuracy. They can estimate uncertainties in compound measurements, and can make educated guesses about uncertainties in the method of measurement. If uncertainties are small enough to be ignored, the candidate should note this fact.*

Higher level candidates *should be able to express uncertainties as fractions,* $\triangle x/x$*, and as percentages,* $(\triangle x/x) \cdot 100$*. They should also be able to propagate uncertainties through a calculation.*

When calculating total uncertainties during a calculation make sure that only the uncertainties that will actually affect the calculation are used. E.g. In a titration, the uncertainty in the amount of indicator used will not feature in the calculation.

Express each uncertainty as a %.

E.g. Uncertainty for volume of HCl (titre): 0.1cm^3 = 0.1/21.08 = ±0.5% (1 s.f.)
Uncertainty for volume of NaOH (pipette): 0.05cm^3 = 0.05/25.00 = ±0.2% (1 s.f.)

Then add up to give the total uncertainty.

Total uncertainty = 0.7% (1 s.f.) approx = 1%

Note that it is a good idea to round up/down a total uncertainty to 1 sig fig as it is unreasonable to give predicted uncertainty to such a high degree of accuracy.

This value can now be given with the final answer.

E.g. 0.164g/100 cm^3 ± 1%

This could then be converted to an actual value   *E.g 0.164g/100 cm^3 ±0.0016g*

Now you can see how the uncertainty affects the last significant figure and how the final answer could be given as a range

E.g. 0.01656g-0.01624g/100cm^3 so that students should be able to appreciate the accuracy of the third significant figure and also appreciate the number of significant figures that the answer should reasonably be given to.

E.g 0. 0166-0.0162g/100cm^3 or 0.0164g ±0.002g/100cm^3

For both SL and HL students, the number of significant figures given in the final answer displays an important appreciation of the uncertainties encountered in the practical.
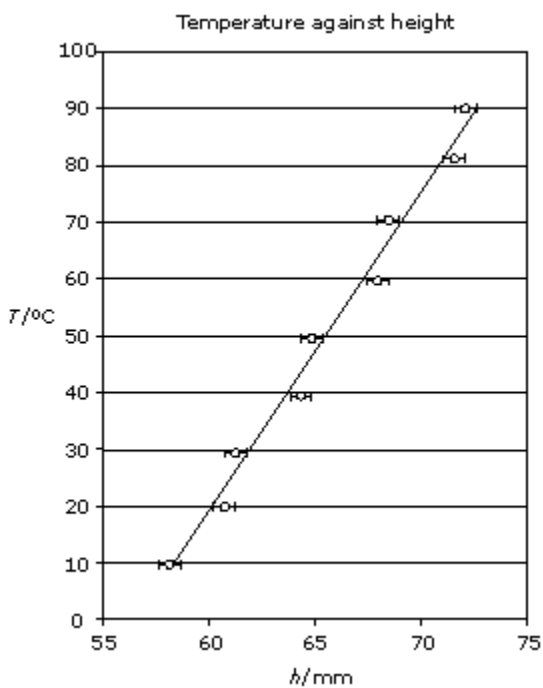
Graphical uncertainties

This is what the IB states about propagating uncertainties in graphs.

*Note: Standard level and higher level candidates are **not** expected to construct uncertainty bars on their graphs.*

Students should show appreciation of uncertainties in their graphs by drawing a line of best fit.

If a gradient is required then some appreciation of how this would be affected by drawing other possible lines of best fit is required.

Although not required by the IB, we encourage HL students to appreciate uncertainty in the gradient by drawing error bars or boxes if appropriate.

Differences Between SL and HL Students

Both SL and HL students must be able to predict appropriate levels of uncertainty in data collection.

HL students only must be able to propagate these uncertainties through a calculation to give % uncertainty and then convert this into a value that corresponds to the appropriate allocation of significant figures to their answer during data processing.

Although SL students do not need to do this to the same extent, they should be encouraged to give a simple sum of the total % uncertainty to show that they understand the number of significant figures that will be appropriate in their answer.

As stated, neither SL or HL students need to draw error bars on a graph. However both need to be able to draw a line of best fit. They can then show some understanding of how there is an uncertainty in the gradient of this line. HL students could give a maximum and minimum possible gradient.

Uncertainties in the Conclusion

When comparing the experimental result to the literature value, it is important to determine the % by which it is either above or below the expected answer.

If this % difference is greater than the % difference estimated from the random uncertainties propagated in the data processing, then this leads to a conclusion that there were other **systematic errors** that resulted in a consistently high or low answer. This should be stated clearly as part of the conclusion.

Explanation of Errors in the Evaluation

The students can then predict and explain the nature and importance of these **systematic errors** in their evaluation. Systematic errors lead to the experimental answer being consistently higher or lower than the literature value. These can often be predicted by looking at the qualitative observations during a practical. (E.g. A colour change could be consistently over estimated). Also look at some of the control variables. (E.g. A standard $0.05mol\ dm^{-3}$ solution could actually be $0.054mol\ dm^{-3}$ leading to consistently lower than expected results, or too few drops of indicator were added leading to difficulty in identifying the end point and an overestimation of the given answer.

During the evaluation try to discuss which systematic errors specifically lead to the experimental answer being either above or below the given answer. Give these in order of importance.

The IB are very keen on students displaying that they understand the difference between random and systematic errors.

*E.g. **Conclusion**:*

Melting point (or freezing point) is the temperature at which the solid and the liquid are in equilibrium with each other:

Melting points can be used as an accurate means of identification of a substance and its purity. Melting points vary due to the molar mass of the compound and the strength of the intermolecular bonding. Each compound therefore has its own unique melting point. The purity is displayed by the compound melting at a precise temperature rather than over a temperature range. By comparing the melting point to a list of possibilities the identification of the compound can be predicted.

Literature value of melting point of para-dichlorobenzene = 53.1°C (Handbook of Chemistry and Physics).

Experimental melting point = 58⁰ ± 2%

Difference in literature value to experimental value = 58 – 53.1 = 4.9⁰C

% difference above literature value = 4.9 / 53.1 x 100 = 9.2%

The fact that % difference is greater than % uncertainty predicted means random uncertainties alone cannot explain the difference and some systematic error(s) must be present.

Evaluation of procedure and modifications:

i. Duplicate readings were not taken. Other groups of students had % uncertainty > % difference, ie in their case random errors could explain the % difference, so repeating the investigation is important.
ii. The thermometer—how accurate was it? Was it consistently reading a temperature higher than the actual?
iii. The melting point apparatus heated the solid up too quickly, so that when the solid had melted the temperature had actually already risen above this point on the thermometer.
iv. The capillary tube used for heating was too thick, insulating the solid from the heat, meaning that it appeared to melt at a higher temperature.
v. It was difficult to see when the solid had completely melted due to the small size of the sample used. This also meant that the temperature rises were much faster.

Improvements

The sample in the test tube was not as large as in other groups. Thus the temperature rises were much faster than for other groups. Also it was harder to see the solid melt. A greater quantity of solid, plus use of a more accurate thermometer would have provided more accurate results.

The thermometer should have been calibrated in order to eliminate any systematic errors due to the use of a particular thermometer. Calibration against the boiling point of water (at 1 atmosphere) or better still against a solid of known melting point (close to the sample) should be done.

Heat the solid more slowly to see the melting point more clearly.

Carry out a trial with a solid of known melting point to find the effect of the capillary tube on insulating the heat.