Guidelines for Writing a Lab Report

General points.

- Ascertain before you undertake the investigation which criteria (if any) are being assessed.
- Check that you are clear about all the aspects to be assessed.
- Record all your work as you proceed in your log book or laboratory notebook.
- Record the title of the experiment (or piece of work), the date and the name(s) of any partner(s) you worked with.
- Record precise details of all equipment used, e.g. a balance weighing to + or - 0.001 g, a thermometer measuring from -10 to + 110 ^{o}C to an accuracy of + or - 0.1^{o}C, a 25.00 cm^3 pipette measuring to + or -0.04 cm^3 etc.
- Record precise details of any chemicals used, e.g. copper(II) sulfate pentahydrate $CuSO_4.5H_2O(s)$.
- Record all measurements accurately to the correct number of significant figures and include all units.
- Record all observations. Include colour changes, solubility changes, whether heat was evolved or taken in etc.
- Draw up a checklist to cover each criterion being assessed. As you write the laboratory account check that each aspect is addressed fully. (Some students give each aspect a sub-heading. Although this is not strictly necessary it does help to draw the aspect to the attention of the teacher).
- Your work may be hand-written (in ink) or word-processed. Ensure that it is neat, correct and legible.
- Write clearly and make a spell check.
- Hand your work in on time. Teachers are within their rights to refuse to mark work handed in late as you may benefit from using other students' marked assignments.
- Learn from your mistakes. In the early part of the course do not expect to get everything correct the first time you do it. Find out why you lost marks and improve your next presentation.

Error and uncertainty in practical work

The **error** is the difference between the result obtained and the generally accepted 'correct' result found in the data book or other literature. If the 'correct' result is available it should be recorded and the percentage error calculated and commented upon in your conclusion. Without the 'correct ' value no useful comment on the error can be made.

Uncertainty occurs due to the limitations of the apparatus itself and the taking of readings from scientific apparatus. For example during a titration there are generally four separate pieces of apparatus, each of which contributes to the uncertainty.
e.g. when using a balance that weighs to $\pm$ 0.001 g the uncertainty in weighing 2.500 g will equal

$$\frac{0.001}{2.500} \times 100 = 0.04\%$$

Similarly a pipette measures 25.00 cm^3 $\pm$ 0.05 cm^3.
The uncertainty due to the pipette is thus $\frac{0.05}{25.00} \times 100 = 0.2\%$

Assuming the uncertainty due to the burette and the volumetric flask is 0.50% and 0.10% respectively the overall uncertainty is obtained by summing all the individual uncertainties:
Overall uncertainty = 0.04 + 0.2 + 0.50 + 0.10 = 0.84% $\simeq$ 0.8% (1 significant figure)
Hence if the answer is 1.87 mol dm^{-3} the uncertainty is 0.8% or 0.01496 mol dm^{-3}
The answer should be given as 1.87 $\pm$ 0.01 mol dm^{-3}. (1 significant figure)

Sometimes it is not possible to give precise percentage uncertainties. For example in a titration the end-point taken could vary according to the person carrying out the titration. In such cases you should state the colour change taken (e.g. until a faint permanent pink colour was obtained). Any assumptions made which can add to the uncertainty (e.g. the specific heat capacity of the solution was taken to be the same as that for pure water) should be stated in the evaluation.

Significant figures

Whenever a measurement of a physical quantity is taken there will be uncertainty in the reading. The measurement quoted should include the first figure that is uncertain. This should include zero if necessary. Thus a reading of 25.30^{o}C indicates that the temperature was taken with a thermometer that is accurate to + or - 0.05^{o}C. If a thermometer accurate to only + or - 0.5^{o}C was used the temperature should be recorded as 25.3^{o}C.

Zero can cause problems when determining the number of significant figures. Essentially zero only becomes significant when it comes *after* a non-zero digit (1,2,3,4,5,6,7,8,9).

000123.4 0.0001234	1.0234 1.2340
zero not a significant figure	zero is a significant figure
values quoted to 4 sig. figs.	values quoted to 5 sig. figs.

Zeros after a non-zero digit but before the decimal point may or may not be significant depending on how the measurement was made. For example 123 000 might mean exactly one hundred and twenty three thousand or one hundred and twenty three thousand to the nearest thousand. This problem can be neatly overcome by using scientific notation.

1.23000×10^{6} quoted to six significant figures
1.23×10^{6} quoted to three significant figures.

Calculations.
1. When adding or subtracting it is the number of decimal places that is important.

e.g. 7.10 g	+	3.10 g	=	10.20 g
3 sig. figs.		3 sig. figs.		4 sig. figs.

This answer can be quoted to four significant figures since the balance used in both cases was accurate to + or - .01g.

2. When multiplying or dividing it is the number of significant figures that is important. The number with the least number of significant figures used in the calculation determines how many significant figures should be used when quoting the answer.
e.g. When the temperature of 0.125 kg of water is increased by 7.2^{o}C the heat required =
$0.125 \text{ kg} \times 7.2^{o}\text{C} \times 4.18 \text{ kJ kg}^{-1} \, {}^{o}\text{C}^{-1} = 3.762 \text{ kJ}$.
Since the temperature was only recorded to two significant figures the answer should strictly be given as 3.8 kJ.
In practice the IB does not tend to penalise in exams if the number of significant figures in an answer differs by one from the correct number (unless the question specifically asks for them) but will penalise if they are grossly wrong.