

Unit I: Measurement and Data Processing

Vocabulary: precision, accuracy, repeatable, reproducible, systematic error, random error, experimental error, total percentage error, tolerance, significant figures, uncertainty, interpolation, extrapolation, independent variable, dependent variable, control variable, line of best fit, gradient, slope, line of minimum and maximum fit, error bars

1. Measurement

**ACT - thickness of aluminum foil*

When we measure, we are comparing to a known standard. All measurements have some degree of uncertainty. e.g. if students used different thermometers to measure the temperature of water, we might get several different results, of which we would take the average. Raw data should always be recorded in a table, but is often also shown in graphical form. It is easier to spot trends, and also to identify data points that don't fit the trend (outliers).

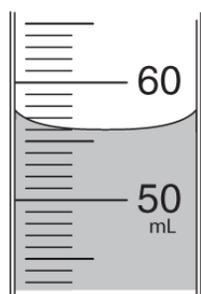
Uncertainty – the range of possible values for a measurement. e.g. $25.0 \pm 0.5\text{mL}$ means that the actual value could be as high as 25.5mL or as low as 24.5mL .

In analogue instruments, the uncertainty range or tolerance is often written on the instrument.



e.g. a 1000mL volumetric flask with uncertainty range of $\pm 0.80\text{mL}$.

This is useful if you are filling to the line. If you are measuring an unknown volume of liquid, take it from the bottom of the meniscus.



First, record the digits that you are certain of. Then record the last digit which is an estimate. E.g. This graduated cylinder would be 56.0mL . The uncertainty is conventionally half of the smallest division. Smallest division is 1mL , so uncertainty is $\pm 0.5\text{mL}$.

$56.0 \pm 0.5\text{mL}$

In digital instruments, the uncertainty will be \pm the smallest scale division.

There are other uncertainties during a lab which should be recorded even if they cannot be quantified. E.g. judging a colour change during a titration, or the reaction time of an experimenter using a stopwatch, etc.

2. Significant Figures/Digits

The digits in measurement up to and including the first uncertain digit e.g. 27.33g has four significant digits, 27 has two significant digits.

If you see a measurement of 10.00g , this would have 4 significant digits. (the zeroes are significant because they tell you the instrument used has an uncertainty range of $\pm 0.01\text{g}$.) Conversely, a measurement of 10g would only have 1 significant digit. Some rules:

- a) leading zeroes are not significant e.g. 0.0034 cm³ has 2 SD
- b) trailing zeroes are not significant unless a decimal is present e.g. 5400 has 2 SD, 5400.0 has 5 SD
- c) zeroes between non-zeroes are significant e.g. 2002 has 4 SD

If you are converting units for a measurement, the number of SD should stay the same:

e.g. 1000.0mL (5 SD) is equal to 1.0000L (also 5 SD)

How can you convert 1.00L into mL and keep 3 SD?

We have to use scientific notation: 1.00×10^3 mL

3. Calculations with significant digits

Evaluate this calculation: 1.7L of a liquid weighs 4g. A student calculates the density to be 2.35294118 g/L. Is this a logical answer?

a) Multiplication and Division

The answer cannot be more precise than the least-precise value in the calculation i.e. round so that the number of SD in the answer is the same as the least number of SD in the calculation.

NOTE: SD only applies to measurements!

e.g. the density of the above example should be rounded to 2 g/L

b) Addition and subtraction

The answer should be rounded to the place value of the least significant digit used in the calculation. It may be helpful to write the numbers vertically e.g. $1.01\text{L} + 25\text{L} = 26\text{L}$. (the answer can be no more precise than the “ones” column)

4. Experimental Error

Experimental Error is the difference between the recorded value (i.e. what you calculate in the lab) and the literature value. This is usually recorded as a percentage and called the total percentage error. There are two types of experimental error

a) Random errors a.k.a. random uncertainties

Cannot be avoided and results in an equal probability of a reading that is too high or too low, caused by: insufficient data, readability of the instrument, changes in surroundings e.g. temperature or pressure changes. NOTE: called an “error” but not the same as a mistake!

e.g. reading the volume of solution in a burette (there is uncertainty with any measurement)

e.g. the mass recorded from an electronic balance

e.g. the response time of a person timing a reaction with a stopwatch.

The effect of random uncertainties can be minimized by repeated trials. If the results are duplicated by the individual, we say the results are repeatable. If the results are duplicated by others, we say the results are reproducible.

b) Systematic Error

These are the result of poor experimental design or procedure. They cannot be reduced by repeated trials. They are also hard to evaluate unless you know the accepted or true value.

e.g. heat lost to the environment when you are measuring the temperature of an exothermic rxn.

How would this affect your results?

e.g. evaporation of water from a solution used in a reaction

How would this affect your results?

NOTE: on an IB lab report, in your "Evaluation", you should have separate paragraphs to discuss random and systematic error. When comparing your result to the literature value, consider if random error alone can account for the difference. In other words, is the literature value within your calculated range of uncertainty? If not, then suggest how systematic error may have affected your results, and suggest some changes in procedure to reduce this.

5. Precision and Accuracy

Precision refers to how reproducible a measurement is. For a measurement to be reproducible, there must be little random error.

Accuracy refers to how close a measurement is to the true or accepted value. For a measurement to be accurate, there must be little systematic error.

6. Uncertainties in calculated results a.k.a. Propagation of Error

When uncertainties in the raw data are used in a calculation, the answer will also have uncertainty.

a) Addition and Subtraction

Consider two burette readings:

Initial reading: $15.05 \pm 0.05 \text{ cm}^3$

Final reading: $37.20 \pm 0.05 \text{ cm}^3$.

Given the uncertainties, what is the maximum possible volume and the minimum possible volume for the amount titrated?

b) Multiplication and Division

Change the absolute uncertainties to percentage uncertainties and add them together. You can then change back to absolute uncertainty for your answer.

e.g. Calculate the density of a liquid that has a mass of $24.0 \pm 0.5\text{g}$ and a volume of $2.0 \pm 0.1\text{ cm}^3$.

i) convert the absolute uncertainties to percentage uncertainties

ii) add the percentage uncertainties together (usually rounded to one SD)

iii) convert the percentage uncertainty back to absolute uncertainty (don't use rounded value)

7. Percentage Uncertainty vs. Percentage Error (2 different things!)

Percentage uncertainty is what you did in #6 above. It is the absolute uncertainty/measured value * 100%. Simply, it is expressing the uncertainty as a percentage.

Percentage error is a measure of how close the experimental value is to the literature or accepted value.

Percentage error = (accepted value – experimental value)/accepted value * 100%

Note: In a lab, one uncertainty may be much larger and therefore have a major effect on results. This should be stated. As well, can the difference between the experimental result and the literature be explained by the uncertainties? Suggest procedural improvements.

8. Graphing

- Give a title
- Label axes with both quantities and units. Use as much space as possible
- Independent variable (the “cause”) goes on the x-axis. This is the variable that is being manipulated. This is also where “time” usually goes. The dependent variable (the “effect”) goes on the y-axis. This is what is being measured.
- Plot data points clearly. Error bars are not necessary for IB Chem but useful for future!

- If a linear relationship, use a ruler to draw the line of best fit. This helps to eliminate the effect of random uncertainty. Identify data points that do not fit the trend.
- Identify the units for the slope (rise/run)
- The origin (0,0) may or may not be relevant, so think carefully about whether it should be included.

Interpolation – determining data points between those which were measured.

Extrapolation - finding values outside the range of what was measured. (p. 320)

9. Unit Conversions

See handout “Unit Conversions”. This is a review of Math 10 but you will do unit conversions throughout Chem 11.

10. Metric unit conversions

See handout. You are responsible for knowing the commonly used metric prefixes: milli, centi, kilo, etc.

11. The Mole – a very special unit!

The Mole is a unit used often in chemistry as it describes the number of particles in a substance.

1 mol is equal to 6.02×10^{23} particles. 6.02×10^{23} is called Avagadro’s number.

Conversion factors:

e.g. How many molecules in 3.5 mol of ammonia?

e.g. How many oxygen atoms in 0.8 mol of carbon dioxide?

12. Derived units

Derived units are units that are made by combining two or more units e.g. density

e.g. The heat change equation is $\Delta H = c \times m \times \Delta T$, where H is measured in Joules, mass in grams, and temp. in °C.

Derive a unit for “c”:

e.g. Derive a unit for the Universal Gas Constant (R) for the Ideal Gas Law: $PV = nRT$, where

P (atm)

V (L)

n (number of particles, measured in moles)

T (K)