Data Management

Uncertainty in numbers: Precision in measurement

There is some degree of uncertainty in every measurement. This is due to limitations of the measuring device and the experimenter's ability to use the device.
Precision refers to the variation in results obtained when an experiment is repeatedly performed with the same equipment and the same procedure. Thus, for precise work the experimenter must be able to obtain consistently the same results with the same instruments each time he performs a given experiment.
Although each instrument has its own typical uncertainty, we do talk about the typical or average uncertainty value of an instrument. The table below shows some of the typical uncertainty values associated with high school equipment.

Instrument	Typical Uncertainty
triple beam balance (centigram)	±0.01 g
electronic centigram balance (0.01 g)	±0.005 g
electronic decigram balance (0.1g)	±0.05 g
100 cm^3 graduated cylinder	±0.5 cm^3
50 cm^3 graduated cylinder	±0.5 cm^3
10 cm^3 graduated cylinder	±0.05 cm^3
-10 ^{0}C to 110 ^{0}C thermometer	±0.5 ^{0}C
meter stick	±0.05 cm
stop watch	±0.14 s

(Note: uncertainties are usually to 1 significant figure)

The Accumulation Of Errors In Calculated Results
So far we have discussed how to record a single measurement and the uncertainty connected with that measurement. Since each measurement has an uncertainty, the combination of several measurements in a calculation must also have uncertainty.

** When quantities are **added or subtracted**, the maximum uncertainty in the result is the **sum** of the **absolute** uncertainty for each of the component measurements.

example 1. Suppose you weighed a sample of copper and a sample of lead on a triple beam balance,

mass of copper	=	0.97 ± 0.01 g
mass of lead	=	1.66 ± 0.01 g
sum		2.63 ± 0.02 g

example 2. In an experiment, you obtained the mass of copper left over after removing some of the copper sample for testing.

mass of initial copper sample	=	3.06 ± 0.01 g
mass of removed copper	=	1.06 ± 0.01 g
mass of left over copper	=	2.00 ± 0.02 g

** In **multiplication and division**, the derived uncertainty is **not** the simple sum of the uncertainties in the factors. You must add the **relative errors**!! The absolute error is then the fraction of the most probable answer.

example. What is the area of a square that is 2.6 ± 0.5 cm by 2.8 ± 0.5 cm?

First, we determine the product 2.6 cm x 2.8 cm = 7.28 cm^2

Relative error 1 = $\frac{0.5}{2.6}$ = 0.192

Relative error 2 = $\frac{0.5}{2.8}$ = 0.179

Sum of relative errors = 0.371

Absolute error = 0.371 x 7.28 cm^2 2.70 cm^2 (errors are expressed to 1 significant figure)

= 3 cm^2

Therefore the area = **7.3 ± 3 cm^2**

Problems: Try to remember significant figures in your calculations!!

1. Find the area of a rectangle measured at 10.0 ± 0.1 cm by 2.5 ± 0.1 cm.

2. Four groups have done a calorimetric experiment to determine the heat of reaction. The accepted value for the reaction is 38.73 kJ mol^{-1}. The values found by the groups (in kJ mol^{-1}) were:

a) 35.1 ± 0.3 b) 36.5 ± 0.5 c) 33.2 ± 0.1 d) 34.7 ± 0.2

i) Which result is most precise?
ii) Which result is most accurate?

3. What is 24.31 ± 0.3 g - 16.765 ± 0.3 g?

4. Calculate the value and the uncertainty of X, where

X = A(B - C) A = 123 ±0.5 B = 12.7 ± 0.2 C = 4.3 ± 0.1

5. What is the density of a block of pilotium (a substance used to clean the ear wax in ants) measuring 6.2 ± 0.1 cm by 23 ± 0.1 cm by 2.75 ± 0.1 cm and having a mass of 170 ± 0.5 g? (D = M/V)

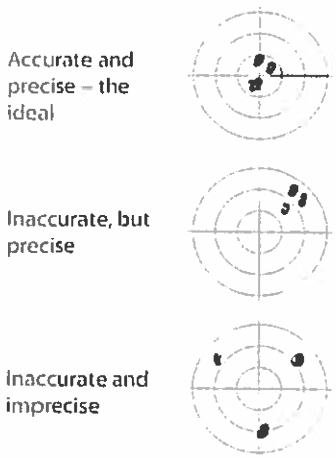
6. Earl Herkemer measured out the following masses on a decigram balance; 2.38 g, 0.42 g, 3.91 g and 5.73 g. What is the total mass including its uncertainty?

Note: You will need to use error calculations in upcoming labs * especially your IA. Please try to learn these simple rules of error analysis.**

*** Just doing a few error calculations and including uncertainties in your data tables on your IA will increase your mark considerably!!!**

Uncertainty and Error in Measurement

Very often experimental results are somewhat faulty as there are always **experimental errors** and **uncertainties** involved. It is vital to differentiate between **accuracy** of a result and the **precision** of a result. Accuracy of a result is a measure of how close the result is to the accepted value. Precision of the result is a measure of the certainty of the value determined.



Example: Two groups of IB chemists are measuring the value of the ideal gas constant (accepted value is 8.314 J mol^{-1} K^{-1}). Group A gets a value of 8.32 ± 0.5 J mol^{-1} K^{-1}. Group B gets a value of 8.02 ± 0.1 J mol^{-1} K^{-1}.
Group A's value is more accurate (i.e. closer to the accepted value) but group B's value is more precise (i.e. has the smaller uncertainty).

Experimental Errors- there are two types we are interested in;

- **Random errors:** random errors make a measurement **less precise** but not in any particular direction. Random errors are mostly due to limitations in the instrument. For example when reading a digital instrument, such as a balance, we should record the uncertainty (or error) as being half of the last digit. A reading of **34.86** g on a digital readout should be recorded as 34.86 ± 0.005 g. Another example, might be that if you weigh a piece of copper wire on a triple beam balance you might record the mass as 2.89 ±0.01 g. The significance of the ±0.01 g is that a repetition of the measurements is expected to yield 2.88, 2.89 or 2.90 g. In many repetitions of the measurement made, the central value is expected to occur

- **Systematic errors:** systematic errors always affect the result in a particular direction and hence the **less accurate**. These errors arise from flaws or defects in the instrument itself or from errors in the way the measurements were taken.

Examples could be

- measuring the mass of water in a cup - but not tarring the cup (your mass would consistently be too high).
- not calibrating a voltmeter properly so it measures 0.05 volts units too low all the time.
- using the same poorly insulated container to measure heat of reaction
- constantly reading the volume in a burrette from the top of the meniscus.
- any others you can think of?