Precision and accuracy

An ***accurate*** measurement is close to the true value. Often you don't know the true value. In this case, the *precision* of the measurement gives you some idea of how much you can trust the result, although a measurement can be precise but inaccurate.

Precision is a measure of the agreement among a series of measurements. The precision of a measurement is sometimes expressed (after statistical analysis of data) in the following form:

0.89 ± 0.02 g/mL

12.015 ± 0.005 g

Confidence, a related quantity, is also expressed in this form.

Measurement and precision

•Precision depends upon the equipment used for a measurement and upon the care with which it is made. The way in which a measurement is written down implies something about the expected precision!

•"25.000 g" implies a careful measurement made with good equipment.

- •Repeated measurements would seldom vary from this one by more than ±0.001 g. This measurement is highly repeatable, so it is precise

•"25 g" implies a measurement made with poor equipment (or good equipment and careless technique)

- •Repeated measurements could vary by ±1 g or more, so this one is not very precise

Example (1)

- Cheap balance
  Measurements are trustworthy to the nearest gram

Measurement	25g
Implied precision	±1 g

Example (2)

- Measurements made with a standard lab balance are trustworthy to the nearest milligram (0.001 g)

Measurement	25.000g
Implied precision	±0.001 g

Example (3)

- The analytical balance is very precise (also expensive and delicate!!) Measurements are trustworthy to the nearest 0.1 mg

Measurement	25.0000
Implied precision	±0.0001 g

Significant figures

- Significant figures are the digits in a measurement that are known to be precise (will not vary with repeated measurements).
  - A measurement such as 35.004 g contains 5 significant figures because repeated measurements might give the results shown at right. The 3, 5 and zeros are significant because they do not vary. The 4 is significant because it only changes a little.

35.004
35.003
35.004
35.004
35.004
35.003
35.005
35.004

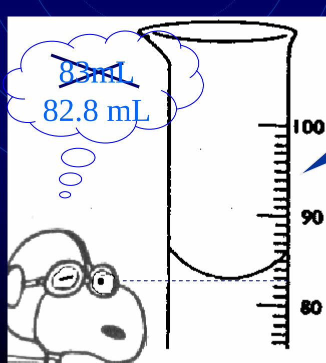
Sig. figs. in lab (1)

All of the digits in the readout of a digital balance or other digital apparatus (**<u>but not the calculator</u>**) are significant and should be written down.

- ➔Don't round off or drop zeros.

Sig. figs. in lab (2)

Ordinarily, you read an analog scale to **<u>one tenth</u> of the smallest graduation** present.



Scale is graduated every 1 mL, so reading should be to the nearest 0.1 mL.

Which digits are significant?

Non-zero digits are significant **96485** (5)

Zeros between non-zero digits are significant **101325** (6)

Trailing zeros are significant if the number contains a decimal point
0.0**500** (3) **1**000 (1?) **1000.** (4) **1.00**x10^3 (3)

Leading zeros are not significant 0.0**821** (3)

" x10^n " in scientific notation is not significant **6.02**x10^{23} (3)

Exact numbers and integers have infinite significant figures
1000 g = 1 kg (both ∞) 6 oxygen atoms (∞)

Logarithms: only digits to the right of the decimal are significant
pH = 4.**85** (2)

Calculations with sig. figs.

- When calculations involving measurements are made, the precision (sig. figs.) of the answer depends on the precision of the measurements themselves.
- The weakest link principle is used, but there are two different rules, one for addition and subtraction and one for multiplication and division.

Addition and subtraction

RULE: When adding and subtracting, the result should have the same number of decimal places as the least precise number added.

Examples

12.345 g + 12.3 g

=24.645 g (incorrect ***s.f.***)
= 24.6 (correct: 1 decimal place like 12.3 g)

12.50 g - 3.50 g

= 9 g (calculator result: incorrect ***s.f.***)
= 9.00 g (correct: 2 decimal places)

Multiplication and division

RULE: When multiplying or dividing, the result should have the same number of significant figures as the number with least significant figures.

Example

Mass = 11.000 g

Volume = 3.0 mL

Density = mass/volume

= 3.66666666 g/mL
(calculator result: too many ***s.f.***)

= 3.7 g/mL (correct: 2 ***s.f.*** like 3.0 mL)

Longer calculations

In a calculation with several steps, the final result must be rounded off to the correct ***s.f.*** Rounding off intermediate results may cause cumulative round-off errors.

Example

$$\frac{5.055}{44} = 0.1148864$$

$$\approx 0.115$$

$$\frac{0.115}{0.0450} = 2.5555556$$

$$= 2.6$$

Intermediate result: <u>one extra</u> sig. fig. retained

Calculator result…

…rounded to 2 sig. figs. because of 44 in 1st step (mult/div rule)