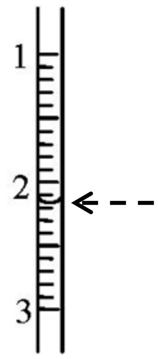
Significant Figures

1. Significant figures (abbreviated as s.f.)

When recording raw data in the laboratory, always write down both the number (with correct significant figures) and the unit of your measurement. Taking the buret on the right as an example, the smallest division is 0.1 mL, and the reading of the buret is "2.18 mL". The first two numbers "2" and "1" are certain digits that can be accurately read. But "8" is an estimated uncertain number, that varies with different people. Hence, the raw data "2.18 mL" has three significant digits which comprises two certain digits and one uncertain digit. Keep in mind that **the significant figures include all certain digits plus one uncertain digit**.



For weighing in the laboratory, one item may be measured as "1.10 g" or "1.0725 g" by different electronic balances (Fig. A-1). This is because the precisions of the two balances are ± 0.01 g and ± 0.0001 g, respectively. If the data is recorded as "1.1 g", the omission of the final zero will mislead the precision to be ± 0.1 g and thus is incorrect. **When recording a measurement from instrument, write down all the digits on display; the last digit shown is the uncertain digit.**

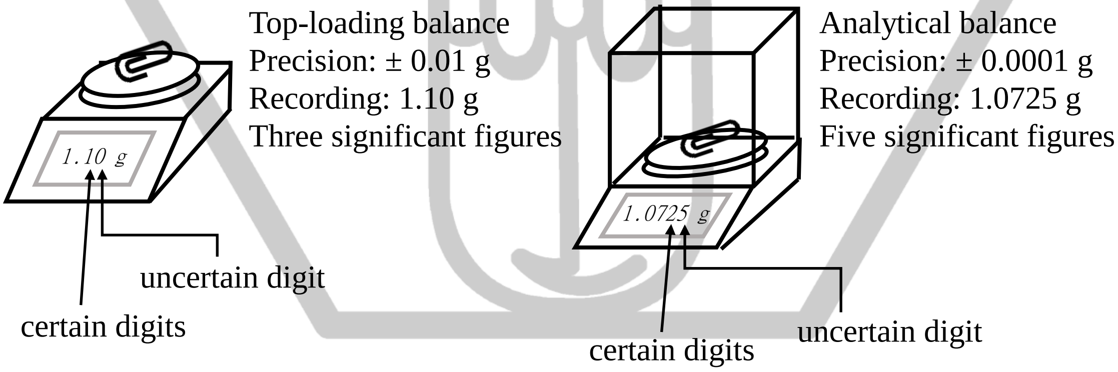


Figure A-1 Electronic balances and precisions

A zero may or may not be significant, depending on its location in a number.

(a) Zeros between non-zero integers are always significant. For example, 1.0725 has five significant figures.

(b) Zeros that come before non-zero integers are never significant. For example, 0.011 has two significant figures.

(c) If the zeros come after non-zero integers and come after the decimal point, they are significant. For example, 1.10 has three significant figures.

However, the situation would still become confusing because trailing zeros may or may not be significant. For example, "1500 mL" may have two, three, or four significant figures, depending on the uncertainty of the measuring instrument. To avoid this ambiguity, change the number to scientific notation. It is easy to determine the significant figures. Give examples below.

1.**5** × 10^3 mL (two significant figures, uncertain by ± 100 mL)

1.5**0** × 10^3 mL (three significant figures, uncertain by ± 10 mL)

1.50**0** × 10^3 mL (four significant figures, uncertain by ± 1 mL)

Furthermore, exact numbers obtained from definitions or by counting numbers of objects can be considered to have an **infinite** number of significant figures.

1 atm = 101325 Pa = 760 torr ≒ 760 mmHg ≒ 76 cmHg

0°C = 273.15 K

0.2786 g × 8 = 2.229 g

2. Significant figures in calculations

A second set of rules specifies how to handle significant figures in calculations. Remember to do the rounding after all calculation steps are completed.

The rounding-off procedure is when the first discarded digit equals or exceeds five, increase the last significant digit by 1; when the first discarded digit is less than five, retain the last significant digit as is.

(1) In addition and subtraction, retain one uncertain digit in the sum or difference.

$$\begin{array}{r} 1.7\mathbf{5} \\ +\quad 9.\mathbf{1} \\ \hline 10.\mathbf{85} \end{array}$$

last significant digit first discarded digit

retain one uncertain digit → three s.f.

The first discarded digit is 5, increase the last significant digit, 8, by 1.

Ans: 10.**9** (rounded)

$$\begin{array}{r} 172.6\mathbf{2} \\ -\quad 1.\mathbf{3} \\ \hline 171.\mathbf{32} \end{array}$$

last significant digit first discarded digit

retain one uncertain digit → four s.f.

The first discarded digit is 2, retain the last significant digit as 3.

Ans: 171.**3** (rounded)

(2) In multiplication and division, the answer should have the same number of significant figures as the number with the least number of significant

figures.

$$\underset{4\text{ s.f.}}{200.\mathbf{5}} \times \underset{3\text{ s.f.}}{3.2\mathbf{1}} = 64\mathbf{3.605} = 64\mathbf{4}\ (\text{rounded}) \rightarrow 3\text{ s.f.}$$

$$\underset{3\text{ s.f.}}{22\mathbf{2}} \div \underset{2\text{ s.f.}}{1\mathbf{1}} = 2\mathbf{0.1818} = 2\mathbf{0}\ (\text{rounded}) \rightarrow 2\text{ s.f.}$$

$$\frac{(\overset{3\text{ s.f.}}{452.3\mathbf{7}} - 21\mathbf{1}) \times \overset{5\text{ s.f.}}{0.2651\mathbf{4}} \times \overset{5\text{ s.f.}}{273.0\mathbf{5}}}{\underset{2\text{ s.f.}}{2.\mathbf{6} \times 10^2} \times \underset{3\text{ s.f.}}{0.010\mathbf{3}} \times \underset{3\text{ s.f.}}{29\mathbf{8}}} = 2\mathbf{1.8964589} = 2\mathbf{2}\ (\text{rounded}) \rightarrow 2\text{ s.f.}$$

(3) When taking a logarithm of a number, keep as many digits to the right of the decimal point as there are significant figures in the original number:

Example: $\underline{3.77} \times 10^4$ has **three** significant figures.

Taking logarithm,

$\log(\underline{3.77} \times 10^4)$

$= \log(10^4) + \log(\underline{3.77})$

$= 4 + 0.\underline{576}341$

$= 4.\underline{576}341$

$= 4.\underline{576}$ (retaining **three** digits to the right of the decimal point)

Another example is converting $[H^+]$ to pH value:

Example: $[H^+] = \underline{2.5} \times 10^{-12}\ M$ has **two** significant figures.

$pH = -\log[H^+]$

$= -\log(\underline{2.5} \times 10^{-12})$

$= 12 - \log(\underline{2.5})$

$= 12 - 0.\underline{39}794$

$= 11.\underline{60}206$

$= 11.\underline{60}$ (retaining **two** digits to the right of the decimal point)

(4) When taking an antilogarithm of a number, keep as many digits as there are digits to the right of the decimal point in the original number:

Example: pH = 8.$\underline{74}$ is a number having **two** significant digits on the right of the decimal point.

Taking antilogarithm,

$$
\begin{aligned}
[H^+] &= \text{antilog}\,(-8.\underline{74})\ M \\
&= 10^{-8.\underline{74}}\ M \\
&= \underline{1.8}197 \times 10^{-9}\ M \\
&= \underline{1.8} \times 10^{-9}\ M \text{ (retaining } \textbf{two} \text{ significant figures)}
\end{aligned}
$$