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Numbers, Units and Measurement

1

OBJECTIVES

This unit:

- Explains standard notation
- Tests you on the use of your calculator
- Describes how to work out the units of a quantity
- Defines accuracy and precision
- Looks at errors and the use of significant figures

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1.1 Very small and very big numbers

Science often involves very large and very small numbers. Such numbers may be cumbersome to write down, and an abbreviated notation (known as 'standard' or 'scientific' notation) is often used. This relies upon the following mathematical notation:

$10^{-6} = 0.000\,001$

$10^{-5} = 0.000\,01$

$10^{-4} = 0.000\,1$

$10^{-3} = 0.001$

$10^{-2} = 0.01$

$10^{-1} = 0.1$

$10^{0} = 1$

$10^{1} = 10$

$10^{2} = 100$

$10^{3} = 1\,000$

$10^{4} = 10\,000$

$10^{5} = 100\,000$

$10^{6} = 1\,000\,000$

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Now let us look at an example of standard notation. Think of the number 100. This is the same as 1 × 100. In standard notation we write this as 1×10^2. Similarly,

2300 becomes 2.3×1000 or 2.3×10^3

$6\,749\,008$ becomes $6.749\,008 \times 10^6$

$0.001\,245\,0$ becomes 1.2450×10^{-3}

EXERCISE **1A**

Standard notation

Express the following in standard notation

(i) 0.000 0345

(ii) 300 000 000

(iii) 0.082 057 5

(iv) 3.5

(v) 602 200 000 000 000 000 000 000

(vi) 17

Sometimes, we are only interested in approximate values. The symbol ≈ means 'approximately' e.g. $6.023 \times 10^{23} \approx 6 \times 10^{23}$.

1.2 Logarithms

Logarithms to the base 10

The logarithm (or 'log') of a number to the base ten is the power to which the number 10 has to be raised in order to equal that number. For example, $100 = 10^2$. Therefore, the log of 100 is 2. Similarly, since $0.0001 = 1 \times 10^{-4}$, the log of 0.0001 is −4.

What is the log of 150? The log of 150 is the value of x in the expression

$$150 = 10^x$$

We carry out this operation on a calculator. In many types of calculator, this is done by entering *log*, the *number*, and then pressing the = button. The log of 150 is 2.176. We write this as

$$\log(150) = 2.176$$

What if you are provided with the log of x and asked to find x? Using the above example, how do we get back to 150 from 2.176? To do this we would need to evaluate $10^{2.176}$. (To carry out this operation on a scientific calculator we use the 10^x key, a common sequence of operations being *shift*, *10*^{x}, *number*, and =.) We then write

$$10^{+2.176} = 150$$

Similarly,

$$10^{-0.9104} = 0.1229$$

Logarithms to the base e (natural logs)

The symbol 'e' is a mathematical constant (like π) where

$$e = 2.718\ldots$$

The logarithm of a number to the base e is the power to which the number e has to be raised in order to equal that number. For example, $e^{3.912} = 50$ so that the natural log of 50 is 3.912. In this book we symbolize natural logs as 'ln':

$$\ln(50) = 3.912$$

It follows from the definition of natural logs that $\ln(e^x) = x$.

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Manipulating logarithms

The following general formulae are useful and apply to logs of any base:

$$\log(ab) = \log a + \log b$$

$$\log\left(\frac{a}{b}\right) = \log a - \log b$$

For example,

$$\log\left(\frac{yZ}{km}\right) = \log y + \log z - \log k - \log m$$

Using your calculator

We are now in a position to summarize the type of calculations you need to be able to do on your calculator in preparation for later units. You will need to be able to

1. enter numbers in standard notation form;
2. add, subtract, divide and multiply numbers;
3. square numbers and find their square roots;
4. use the calculator memory;
5. calculate log x, ln x, e^x and 10^x.

The way you carry out such calculations varies slightly according to the make of your calculator. Refer to the calculator instructions for further information – or ask a knowledgeable friend! Now try Exercise 1B.

EXERCISE **1B**

Quick test on calculator use

Use your calculator to evaluate the following:

(i) $\left(\frac{45.6}{2.34}\right)^2$

(ii) $\sqrt{300.7}$

(iii) log (1.2×10^{-2})

(iv) $10^{-4.56}$

(v) ln (0.178×8.456)

(vi) $e^{-5.20}$

(vii) $e^{-E/RT}$, where
$E = 30\,000$,
$R = 8.3145$ and
$T = 298$

1.3

Units

International system of units

The international system of units (usually known as SI units, from the French *Système International*) consists of several *base units* from which all other units (such as those of volume or energy) are derived. Some of the base units are shown in Table 1.1.

Because the base units are sometimes too large or too small for use, SI prefixes (Table 1.2) are used to produce smaller or bigger units. For example, the milligram (= 0.001 g, and symbolized mg) is used if we are reporting small masses.

The cubic metre (written m^3) is too large for most purposes in chemistry, and the cubic decimetre, dm^3 (or litre) is commonly used. There are 1000 dm^3 in 1m^3. Also, there are 1000 cubic centimetres (cm^3) in a cubic decimetre (see Fig. 1.1). Summarizing,

$$1\,m^3 = 1000\,dm^3 = 1\,000\,000\,cm^3$$

Table 1.1 Base units

Property	Base unit	Symbol for unit
Mass	kilogram	kg
Length	metre	m
Time	second	s
Temperature	kelvin	K
Amount of substance	mole	mol

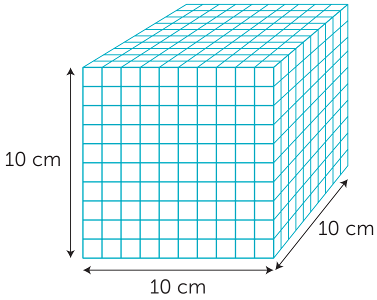

Fig. 1.1 There are 1000 cm^3 in 1 dm^3.

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Table 1.2 SI prefixes

Value	Prefix	Symbol	Value	Prefix	Symbol
10^9, billion	giga-	G	10^{-3}	milli-	m
10^6, million	mega-	M	10^{-6}	micro-	μ
10^3, thousand	kilo-	k	10^{-9}	nano-	n
10^{-1}	deci-	d	10^{-12}	pico-	p
10^{-2}	centi-	c	10^{-15}	femto-	f

You will also meet units raised to negative powers, such as m^{-3} (read as 'per metre cubed'). Remember,

$$m^{-3} = \frac{1}{m^3}$$

Now try Exercise 1C.

EXERCISE **1C**

Working with different units

(i) The radius of the hydrogen atom is approximately 40 pm (picometres). Convert this to nm (nanometres).

(ii) 1.0 cm^3 of a solution contains 0.010 g of salt. What is the mass of salt contained in 1 dm^3 of the same solution?

(iii) The wavelength of orange-yellow light is roughly 600 nm. Express the wavelength in metres.

Amount of substance

One of the most important physical quantities in chemistry is the *amount of substance*, which has the unit of **mole** (symbol mol). The greater the number of particles (atoms, ions or molecules) in a piece of material, the greater is the amount of substance. **Concentration** is another physical quantity which is particularly important to chemists. Concentration is a measure of the packing of particles per unit volume, and is commonly expressed in the units of mol per dm^3, usually written as mol dm^{-3}. Moles and concentration are dealt with in more detail in Units 8 and 9.

Temperature

The hotness or coldness of a material is called its temperature. The units of temperature used in science are degrees **Celsius** (°C) or **kelvin** (K). For more details, see page 156.

Force and energy

The quantities of force and energy occur throughout chemistry and we will need to be familiar with their units.

The SI unit of *force* is the **newton** (N). 1 N is the force needed to give a mass of 1 kg an acceleration of 1 $m\,s^{-2}$. (If the 1 kg mass were stationary before applying the force, this would mean that the mass would have a velocity of 1 $m\,s^{-1}$ after 1 second, 2 $m\,s^{-1}$ after 2 seconds and so on.) The formal definition of the newton is therefore

$$1\ N = 1\ kg\ m\,s^{-2}$$

The SI unit of *energy* is the **joule** (J). 1 joule is the energy used up in pushing against a force of 1 newton over a distance of 1 metre. This means that we can write

$$1\ J = 1\ Nm$$

Units of physical quantities

A physical quantity consists of a number and a unit. For example, suppose we measure the volume of a block and find it to be 4.5 cm^3:

$$\underset{\text{the number}}{\underset{\uparrow}{4.5}} \qquad \underset{\text{the unit}}{\underset{\uparrow}{\text{cm}^3}}$$

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Mathematically, the physical quantity consists of a number multiplied by a unit:

$$\text{physical quantity} = \text{number} \times \text{unit} \tag{1.1}$$

In our example,

$$\text{physical quantity (i.e. volume)} = 4.5 \times \text{cm}^3$$

For convenience, the physical quantity is usually written without the multiplication sign – here as 4.5 cm^3. This may be compared with the algebraic expression $4.5y$ (i.e. $4.5 \times y$).

A laboratory balance. This balance measures masses as low as 0.0001 g.

Labelling axes on graphs

Suppose that we are plotting the volume of a gas (V, in dm^3) against the temperature of the gas (T, in kelvin, K). First, consider the y-axis. We might be tempted to label this axis as ‘V (dm^3)’. However, we are not actually plotting the quantity *volume* but simply the *number* part of the quantity. Rearrangement of equation (1.1) shows that

$$\text{number} = \frac{\text{physical quantity}}{\text{unit}}$$

This shows that in plotting the number part we are really plotting $\dfrac{\textit{physical quantity}}{\text{unit}}$

Hence the y-axis is labelled Volume/dm^3. This is usually written as V/dm^3. Similar reasoning leads to the x-axis being labelled Temperature/K or T/K.

Deriving the units of a quantity

To illustrate the derivation of the units of a quantity, consider the following question: with mass in kg and the volume in m^3, what are the units of density?

We start with the definition of **density**:

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

We find the units of density by substituting the units of mass and volume into the equation defining density:

$$\text{units of density} = \frac{\text{units of mass}}{\text{units of volume}} = \frac{\text{kg}}{\text{m}^3} = \text{kg m}^{-3}$$

The *units of density* are therefore kilograms per cubic metre (see Example 1.1).

EXAMPLE **1.1**

The ratio $\dfrac{E}{RT}$

occurs frequently in chemistry. E is an energy per mole of substance (in units of joules per mole, symbolized J mol^{-1}), T is the temperature (in K) and R is a universal constant with the units of J mol^{-1} K^{-1} (read as joule per mole per kelvin). What are the units of E/RT?

Answer

$$\text{units of } \frac{E}{RT} = \frac{\text{J mol}^{-1}}{\text{J mol}^{-1}\ \text{K}^{-1}\ \text{K}}$$

EXERCISE **1D**

Deriving units

(i) The energy (q joules) needed to raise the temperature of a material (of mass m grams) by ΔT kelvin may be calculated by the equation

$$q = m \times C \times \Delta T$$

where C is the specific heat capacity of the material. What are the units of C?

(ii) The mass m (in grams) of an amount of substance n (in moles) is related to the molar mass M of that substance by the expression

$$m = n \times M$$

What are the units of M?

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EXAMPLE 1.1 *(continued)*

Cancelling like units and remembering that K is really K^{+1}

$$\frac{\cancel{J}\ \cancel{mol^{-1}}}{\cancel{J}\ \cancel{mol^{-1}}\ \cancel{K^{-1}}\ \cancel{K}}$$

so that E/RT is unitless, i.e. it is simply a number.

Comment

The most commonly encountered form of this expression is $e^{-E/RT}$ (e = 2.718). Note that as E increases, $e^{-E/RT}$ *decreases*. Now try Exercise 1D.

1.4 Errors in experiments

Types of experimental error

Suppose we measure the temperature of a liquid. The difference between a *single* measurement of temperature and the *true* temperature is the *absolute error* of the measurement. The **absolute error** can have a positive or a negative value. Generalizing:

Absolute error = experimental value − true value

It is always good practice to repeat measurements. For example, if we are measuring the concentration of pesticide in a lake, we might fill a large bottle with lake water and later (back in the laboratory) withdraw 50 cm^3 portions of the lake water from the bottle and analyse each portion for pesticide using the same analytical technique. The *mean* pesticide concentration is then obtained by averaging the concentrations found in each 50 cm^3 portion. In this case, the absolute error in the pesticide measurements is the difference between the *mean* measurement and the true concentration.

There are two main kinds of error that we need to consider in experimental measurements.

1 The first type are called **random errors**, random because they cause repeat measurements on the same sample to go up and down.

Random errors will cause successive measurements to be scattered, although averaging a large number of such measurements will produce a mean measurement which will not be greatly affected by random error. However, repeating a measurement many times may not be practical (for example, there may be too little sample available, or the measurements might be too time consuming or expensive to carry out). It is for this reason that precise measurements are highly desirable.

2 The second type are called **systematic errors**. A systematic error affects all measurements, and makes all measurements either higher or lower than the 'true value'. Systematic errors do not average out, no matter how many repeat measurements are made.

Examples of random errors

Random errors are introduced whenever there is a *subjective* part to the experiment (such as estimating when a solution has reached the mark in a pipette, or recognizing the onset of a colour change during a titration), *or* where the experimental measurements are influenced by rapidly fluctuating conditions (e.g. air draughts).

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Examples of systematic errors

A simple example of a systematic error is provided by a balance. Balances are often set to read a mass of zero before being used to weigh a sample. Suppose a speck of dust falls upon the pan of a balance *after* zeroing. This will cause the indicated mass of any object to be greater than the true mass. For example, if the speck has a mass of 0.0001 g, all objects will have an apparent mass which is 0.0001 g too high.

Another example of systematic error involves the analysis of chromium in blood. If the blood samples are stored in stainless steel vessels prior to analysis, then some chromium may dissolve out of the steel into the sample. This introduces a systematic error which causes the measurement (here the chromium concentration) to be overestimated.

Systematic errors are often difficult to recognize, particularly in measurements of the concentration of substances (**quantitative analysis**) in which the concentrations of materials are being found in the presence of substances which *interfere* with the measurement (see Box 1.1).

BOX **1.1**

Example of systematic errors – the analysis of aluminium ions (Al^{3+}) in tea

One way of determining the concentration of Al^{3+} in tea is to add a **complexing agent** (usually a complicated organic compound) to the tea (Fig. 1.2). This combines with the Al^{3+} ion to produce a red coloured substance (a coloured **complex**):

$$Al^{3+} + \text{complexing agent} \rightarrow \underset{\text{(red colour)}}{Al^{3+}\ \text{complex}}$$

The stronger the concentration of Al^{3+} in the tea, the stronger will be the intensity of the red colour.

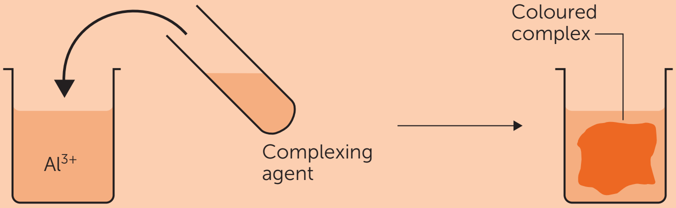

Fig. 1.2 Determination of Al^{3+} ion by coloured complex formation.

However, if the tea contains traces of heavy metal ions (such as copper), it happens that these ions will also form red coloured compounds with the complexing agent. If we presumed that all the red colour was due to the Al^{3+} complex, this would produce a systematic error in which the Al^{3+} concentration in the tea is *overestimated*.

If the tea contains traces of fluoride ions (F^-), these ions react directly with the Al^{3+} producing stable aluminium fluoride complexes and so preventing the Al^{3+} ions from reacting with the complexing agent. This leads to a systematic error in which the Al^{3+} concentration in the tea is *underestimated*.

In an ideal measurement, we would separately measure the concentrations of ions (F^-, Cu^{2+} etc.) which interfere with the measurements of the Al^{3+} ion, and correct the measured Al^{3+} concentration accordingly.

EXERCISE **1E**

Systematic errors

In order to compare the alcohol content of several wines, a student poured samples of each wine into open test tubes and the next day analysed each for alcohol using a standard analytical technique. Comparison of the student's results with those obtained by other laboratories showed that her alcohol concentrations were consistently low. Suggest one reason for the systematic error. (**Hint:** What happens to wine when it is left open to the air?)

Accuracy and precision

Repeat measurements on the same sample which are close together are said to be *precise*:

Precise measurements have a small random error.

A measurement which is close to the true value is said to be *accurate*:

Accurate measurements have a small systematic error.

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The 'rifle shooting analogy' helps us to distinguish between **accuracy** and **precision** (Fig. 1.3). In a rifle competition, the aim is to hit the bullseye. Competitor A is a precise shot (the shots are close together) but inaccurate (no bullseye); B is a precise and accurate shot (three bullseyes); C is neither precise nor accurate.

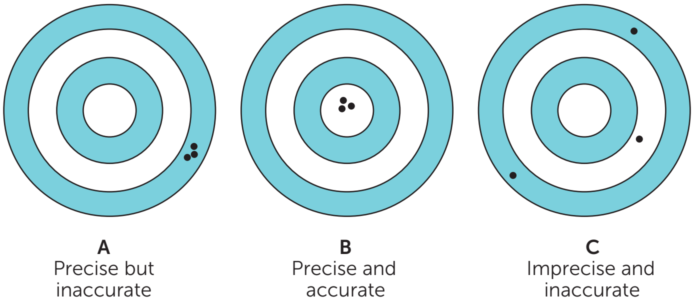

Fig. 1.3 Accuracy and precision – the rifle shooting analogy.

We have already noted that measurements are usually repeated several times and that the random errors will be nearly completely cancelled out in the mean measurement provided that enough repeat measures are made. This is why the precision of measurements is important: *the greater the precision, the fewer the number of repeat measurements that need to be made in order for the random errors to be eliminated.* The fewer the repeat measurements that are required, the quicker and cheaper are the measurements.

See **Extension 1** on the website, www.macmillanihe.com/Lewis-Chemistry-5

One way of showing the precision of a set of repeat measurements is to quote the standard deviation of those results. The lower the standard deviation of a set of results, the better is the precision of those results.

What do we mean by the true value of a measurement?

If a measurement is accurate, it must give a true value. But how do we know the true value? If we are using a new analytical instrument which detects lead, we might test the accuracy of the instrument by analysing solutions whose lead concentrations [Pb] are known (i.e. **standard solutions**). If the analysis is accurate, the difference

$$[\mathrm{Pb}]^{\text{in}}_{\text{standard}} - [\mathrm{Pb}]^{\text{found by}}_{\text{instrument}}$$

should be close to zero. If we are analysing mixtures where the true concentrations are unknown, the absence of systematic errors is essential if we are to have faith in the final result. In such cases, accuracy may be estimated by comparing the results of *different* analytical methods for the *same* sample.

Professional organizations (such as the British Standards Institute) often publish the most reliable analytical methods in the form of **standard methods**, in which the likely sources of experimental error are highlighted.

Quality of analytical measurements in laboratories

Chemical analysis is a multi-billion dollar business in all industrialized countries, involving thousands of highly skilled scientists and technicians. Poor quality analysis is a barrier to international trade, technological advancement and the development of government policies such as health and safety. For example, unless the concentrations of pollutants can be reliably and accurately measured, regulations based on the highest permissible concentration of an airborne pollutant are unenforceable.

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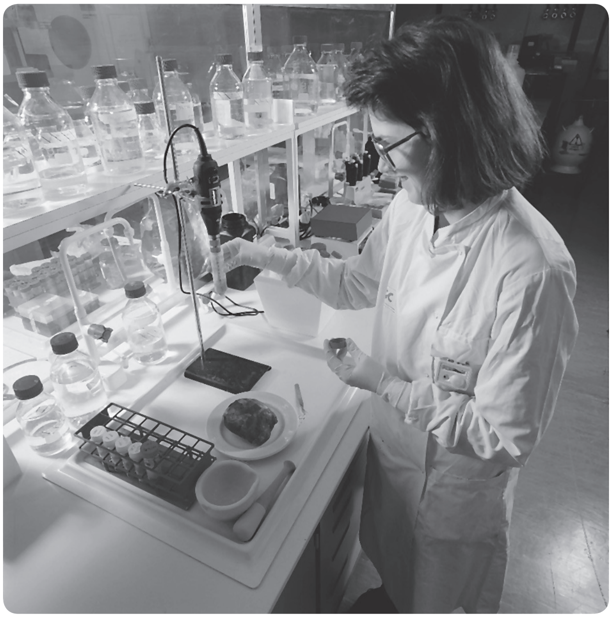
Analytical chemists distinguish between an *analytical technique* and an *analytical method*. The analytical technique used to take the measurement itself (such as gas chromatography or ultra-violet spectroscopy) is one part of a 'recipe' of operations that make up an analytical method. The analytical method may include details of sample preparation and how interference from other substances can be minimized. If the method is recognized by a national or international professional body it is called a **standard analytical method**.



Quality control is very important in analytical laboratories, so that customers know that they can rely upon the results of any analysis.

1.5 Reporting measurements

Significant figures and measurement uncertainty

If we asked someone to measure the length of a piece of wire with a standard ruler and they reported its length as 19.843 cm, we would have every right to be sceptical: 19.843 contains *five significant figures*, a number of figures that cannot be justified when we are using a ruler.

We might estimate the **uncertainty** in the length measurement as ±0.2 cm. This means that the measurement is at worst 0.2 cm too high or 0.2 cm too low. It follows that we are justified only in including the first decimal place of the measurement and we then report the length as 19.8 ± 0.2 cm. Alternatively, we might report the measurement as simply 19.8 cm, a number which contains *three* significant figures. Neglecting the ±0.2 cm is less informative, but because of an agreement between scientists about the meaning of significant figures, even simply writing 19.8 cm carries with it some information about the minimum uncertainties involved in the measurement.

To explain this further, suppose that you report the length of the wire to a friend as 19.8 cm but provide no further information. What could your friend say about the likely uncertainties in the measurement? By general agreement, it is assumed that the uncertainty in the measurement is *equal to at least one digit in the last significant figure*. In our example, reporting the length as 19.8 cm implies that the total uncertainty in the measurement is equal to at least one digit in the first decimal place. In other

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words, the *minimum* uncertainty is ±0.1 cm. As we have seen, the actual uncertainty is estimated to be greater, as ±0.2 cm.

In order to report the correct number of significant figures in a measurement, an estimate of the uncertainties is obviously required. Sometimes this will be nothing more than an informed guess of the likely effect of random errors. In more sophisticated measurements, further experiments may need to be carried out in order to assess the importance of both random and systematic errors.

What is the connection between uncertainty and absolute error? In the above examples, the uncertainty (the ± part) is an estimate of the *range of the absolute error* that is likely to apply to the reported measurement. For example, suppose the uncertainty is ± 0.2 cm. This is equivalent to saying that the absolute error in the reported measurement is between 0.2 cm *above* the measurement and 0.2 cm *below* the measurement.

EXAMPLE **1.2**

The mass of a coin is displayed on a balance as 10.0078 g. The uncertainty of this measurement is estimated as ±0.002 g. How many significant figures are we justified in using when reporting the mass of the coin?

Answer

The uncertainty shows that three decimal places can be justified in the measurement. This means that we are justified in reporting the mass of the coin to five significant figures, i.e. as 10.008 g. This implies that the *minimum* uncertainty in the measurement is ±0.001 g.

BOX **1.2**

Recognizing the number of significant figures

The easiest way to recognize the number of significant figures in a number is to express the number in standard notation and count the number of digits (including zeros) in the number that multiplies the 10^x part. For example, 0.00233 becomes 2.33×10^{-3} in standard notation. Since there are three digits in 2.33 the number of significant figures is three. Other examples are as follows:

Number	Standard notation	Number of significant figures
0.002 330	2.330×10^{-3}	4
235.5	2.355×10^{2}	4
0.000 056 767 6	5.67676×10^{-5}	6
14	1.4×10^{1}	2
1302	1.302×10^{3}	4
150	1.50×10^{2} or 1.5×10^{2}	3 or 2

The number of significant figures in the number 150 is ambiguous. If we mean 1.50×10^{2}, then there are three significant figures. If we mean 1.5×10^{2} then there are only two significant figures.

Number of significant figures in a quantity calculated by multiplication or division

Suppose we carry out an experiment to find the density of a lump of metal. We require two measurements, namely: (i) the mass of the metal and (ii) its volume. Suppose the

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mass of the metal was reported as 10.0078 g whereas the volume of the metal (which is more difficult to determine accurately) was reported as 2.8 cm³. The density is now calculated as

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{10.0078}{2.8} = 3.574\,214\,29\ \text{g cm}^{-3}$$

where the number 3.574 214 29 is the one that might be displayed on a calculator. It is absurd to report the density as 3.574 214 29 g cm^{-3}, since this would suggest an uncertainty of about ±0.000 000 01 g cm^{-3}! Again, there is a rule to guide us: the number of *significant figures* in the final calculated figure is *set equal to the number of significant figures in the most uncertain contributing measurement*. The density calculation depends upon measurements of volume and mass, but the volume measurement was only reported to two significant figures and is therefore the most uncertain of the two measurements. Accordingly, the density should also be reported to two significant figures:

$$\text{density} = 3.\mathbf{6}\ \text{g cm}^{-3}$$

(where the **5** of the 3.**5**74 214 29 is rounded up to **6** as explained in Box 1.3).

Number of significant figures in a quantity calculated by addition or subtraction

The rule here is that the number of *decimal places* in the final calculated figure is *set equal to the smallest number of decimal places in the contributing measurements*.

As an example, consider two different samples of water whose volumes were determined by two different methods as 41.66 cm³ and 2.1 cm³, respectively. Following the rule, and rounding up, the total volume is reported as 43.8 cm³.

BOX **1.3**

Rounding up

Suppose the mass of a coin is incorrectly reported as 5.6489 g (i.e. five significant figures), and that the uncertainties involved only justify the use of four significant figures. This means that we must *round up* to the fourth significant figure.

The rules we use are:

1 In considering the rounding up of the *n*th significant figure, we consider the next (i.e. the (*n* + 1)th) significant figure *only*.
2 The *n*th significant figure is only rounded up if the (*n* + 1)th figure is *equal to or greater than* 5.

So, rounding 5.6489 g (five significant figures) to four significant figures gives 5.649 g. However, if we wanted to report the mass of 5.6489 g to three significant figures our mass becomes 5.65 g, with the fourth significant figure (5.64**8**9) causing the 4 to round up to 5 and the fifth significant figure having no part to play. If we wanted to report the original mass to two significant figures, we go from 5.6489 to 5.6 because 4 is not equal to or greater than 5 and the other original figures (the 8 and 9 in 5.6489) are irrelevant. Finally, the original mass becomes 6 g when expressed to one significant figure with the 6 in 5.6489 causing rounding up. In summary:

Coin mass	Number of significant figures
5.6489	5
5.649	4
5.65	3
5.6	2
6	1

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EXERCISE **1F**

Significant figures and rounding up

(i) How many significant figures are present in the following numbers: (a) 0.02, (b) 20.02, (c) 890, (d) 0.007 65?

(ii) The atomic mass of the oxygen-16 atom is 15.9949 atomic mass units, but a student uses an approximate value of 16.10. Is this justified?

(iii) Round up 0.034 67 to (a) three significant figures, (b) two significant figures and (c) one significant figure.

Number of significant figures in a logarithmic quantity

The log of 1.97×10^3 is 3.294 466, but how many decimal places should the answer contain? The rule here is that the number of *decimal places* in the answer is equal to the number of *significant figures* in the initial number: 1.97×10^3 contains three significant figures, so log $(\mathbf{1.97} \times 10^3) = 3.\mathbf{294}$.

The reverse also applies. If we wish to find the number whose log is 0.8234, we calculate $10^{0.8234}$, which is 6.658 86. The number of significant figures in the answer equals the number of decimal places in the initial number, so that $10^{0.8234} = 6.659$.

Quantities are often expressed in logarithmic form so as to compress them and make very big or very small numbers more manageable. An example is found in the quantity known as **pH**, which is calculated from the hydrogen ion concentration (symbolized $[H^+(aq)]$) in a solution by the equation:

$$\text{pH} = -\log\,[H^+(aq)]$$

This equation is read as 'pH equals the negative of the log of the hydrogen ion concentration'. The reverse of this equation is

$$[H^+(aq)] = 1 \times 10^{-\text{pH}}$$

read as 'the hydrogen ion concentration equals 10 to the power of the negative of the pH value'.

Applying the rules of significant figures, if the hydrogen ion concentration is 4.403×10^{-3} mol dm^{-3} then

$$\text{pH} = -\log\,(\mathbf{4.403} \times 10^{-3}) = -(-2.356\,25) = 2.356\,25 = 2.\mathbf{3563}$$

i.e. four decimal places and rounded up.

If the pH of a solution was expressed as 6.81 then

$$[H^+(aq)]\ (\text{in mol dm}^{-3}) = 1 \times 10^{-\text{pH}} = 1 \times 10^{-6.81} = 1.5488 \times 10^{-7} = \mathbf{1.5} \times 10^{-7}$$

i.e. two significant figures and rounded down.

Further examples of pH calculations are found in Unit 9 (page 152) and in Unit 17 (page 298).

EXERCISE **1G**

Examples of the use of significant figures

(i) Significant figures in multiplication

The concentration of hydrogen ions in a solution was calculated using the equation

$$[H^+(aq)] = \sqrt{2.04 \times 10^{-8} \times c} \quad \text{mol dm}^{-3}$$

Experiments show that c has a mean value of 0.0108. Only two significant figures are justified in c. Report $[H^+(aq)]$ to the correct number of significant figures.

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(ii) Significant figures in addition

The mass of metals in a sample of waste water was determined by analysis to be as follows: Cu^{2+} 0.132 mg, Pb^{2+} 0.3 mg, Zn^{2+} 10.00 mg. What is the total mass of metal present?

(iii) Significant figures in logarithmic quantities

The pK_a of an acid is calculated by the equation:

$$pK_a = -\log K_a$$

where K_a is the *acidity* constant of the acid. At 25°C, the K_a value of ethanoic acid is 8.4×10^{-4} mol dm^{-3}. Calculate the pK_a of ethanoic acid.

REVISION QUESTIONS

Questions 1.1 to 1.4 are multiple choice questions. Select the correct answer.

1.1 The temperature of a liquid is reported as 25.0 ± 0.5°C. The 0.5 represents the estimated:

(a) Precision of the measurement
(b) Accuracy of the measurement
(c) Type of error in the measurement
(d) Uncertainty in the measurement

1.2 The concentration of calcium in water may be determined by atomic absorption spectrometry (AAS), in which the aqueous solution is sprayed into a flame which breaks down the calcium compounds into atoms. However, if phosphate ions (PO_4^{3-}) are present in the sample, they bind the calcium as calcium phosphate, which resists atomization in the flame. Therefore, the presence of phosphate in the sample causes the measured calcium concentration to be:

(a) Underestimated, causing a random error in the measurement
(b) Underestimated, causing a systematic error in the measurement
(c) Overestimated, causing a systematic error in the measurement
(d) Overestimated, causing a random error in the measurement

1.3 Measurements which are precise will always have a:

(a) Large systematic error
(b) Small systematic error
(c) Large random error
(d) Small random error

1.4 A solution containing exactly 10.0 ppm (parts per million) of manganese ion (Mn^{2+}) is sent to a laboratory to test the quality of its measurements. The lab analysed three samples of the solution, obtaining the following results:

Sample number	[Mn^{2+}]/ppm
1	10.0
2	10.5
3	11.0

The absolute error in the mean measurement is:

(a) 0.0 ppm
(b) 0.5 ppm
(c) 1.0 ppm
(d) 21.5 ppm

1.5 The radii of several atoms and ions (in different units) are as follows: Cr^{3+} 0.069 nm, F^- 1.36×10^{-6} cm, O 1.40×10^{-5} mm. Express the radii in metres and arrange the particles in order of increasing size.

1.6 The temperature of a water bath was reported as 27.1°C. What does the number of significant figures tell you about the minimum uncertainty involved in this measurement?

1.7 Report the following measurements to four significant figures:

(i) 0.123 47 V,
(ii) 12.45 m,
(iii) 0.003 557 57 cm,
(iv) 1200.5 K.

1.8 Calculate $e^{-E/RT}$ (with E = 20 000 J mol^{-1} and R = 8.3145 J mol^{-1} K^{-1}) at

(i) T = 300 K and
(ii) T = 3000 K.

In each case, express the result to three significant figures.

1.9 A pain killing tablet contains 154 mg caffeine, 101 mg aspirin and 0.23 g filler. Express the total mass of the tablet (in grams) to the correct number of significant figures.

1.10 The ionic product of an aqueous solution (a solution with water as the solvent), symbolized K_w is defined by the equation

$$K_w = [H^+(aq)] \times [OH^-]$$

where the brackets denote the concentrations (in mol dm^{-3}) of H^+(aq) and OH^- ions at that temperature. What are the units of K_w? Calculate K_w for a solution at 25°C if $[H^+(aq)] = 5 \times 10^{-10}$ mol dm^{-3} and $[OH^-] = 2.0 \times 10^{-5}$ mol dm^{-3}. Express your answer to the correct number of significant figures.

1.11 In spectroscopy, the **absorbance** (A) of a solution of a substance is defined by the expression

$$A = \varepsilon \times c \times b$$

where ε (pronounced epsilon) is the molar absorption coefficient of the substance, c the concentration of the substance (in moles per metre cubed) and b the thickness of the sample (in metres); A is unitless. What are the units of ε?

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1.12 Label the following as random or systematic errors:

- **(i)** The variation of the mass recorded by a balance because of air draughts in the laboratory.
- **(ii)** A thermostat in a water bath that registers 30°C when the real temperature of the water is 25°C.
- **(iii)** The leaking of a gas cell which is being used to hold a sample of gas so that its pressure can be measured.

1.13

- **(i)** The hydrogen ion concentration of a solution is 8.987×10^{-6} mol dm^{-3}. Calculate the pH of the solution.
- **(ii)** The pH of a solution was 11.344. Calculate the hydrogen ion concentration of the solution.

DID YOU KNOW?

Celsius and his temperature scale

The *Centigrade* temperature scale, as its name suggests, consists of 100 degrees between the two 'fixed points'. It was first proposed by the Swedish scientist Anders Celsius in 1742. Celsius suggested that 0°C should be the temperature of boiling water and 100°C should be the freezing point of water. The scale was reversed after his death to the current version with the boiling point of water being set at 100°C and the freezing point of water at 0°C. To avoid confusion with the SI prefix centi- and in honour of the discoverer, the temperature was renamed the *Celsius scale* in 1948.

Additional material to support this unit is available on our website at www.macmillanihe.com/Lewis-Chemistry-5. See **Extension 1**: Numbers, units and measurement. This includes more about the relative sizes of quantities, the use of standard deviation in interpreting the results of analysis and examples (in preparation for other units) of the use of the equation for a straight line.

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