PRINCIPLES OF SCIENCE FOR NURSES

Joyce James, Colin Baker & Helen Swain

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PREFACE

The nursing profession has steadily developed from an untrained workforce to one that provides the highly skilled and educated nurses required in today's health service. Ever since Florence Nightingale started the first nursing school at St Thomas's Hospital the need for educated nurses has been recognised. The continued development of nursing education has seen nurses change from dependent to independent decision makers, a need which has become more important as specialist fields of expertise have evolved. A basic understanding of science and technology is vital to working effectively in this environment.

This book explains basic scientific concepts that are related to nursing practice. This will encourage and enable you to apply these concepts to your everyday working experience. Both the simplest and the most advanced practices and procedures have a fundamental underlying scientific principle. The ability to recognise this will provide a greater understanding of why things are done the way they are and should give you more confidence in contributing to the total care of your patients as well as enhancing your personal satisfaction and achievement.

A simple activity like giving a saline drip to restore fluid balance for example, could trigger a series of questions:

- Why does the fluid balance need restoring?
- Why is the bottle placed above the patient's head?
- What is normal saline?
- Why is it important to keep the lines sterile?

We hope to stimulate this spirit of enquiry in you.

There are a number of ways in which you may wish to use this book but as we recognise that you are not likely to study more than one chapter at a time, each chapter is designed to 'stand alone' as far as is possible. As you work through a chapter, activities and exercises are provided to encourage enquiry and test your understanding as you go. Simple and clear diagrams are provided where appropriate to support the explanations in the text. At the end of the chapter there are further questions to try, with answers provided.

There are several other important features. A glossary is provided which may be helpful for quick reference if there are terms in the text that you are uncertain about. At the start of this book important aspects of expressing measurements and the use of units have been summarised. These are essential for accurate reporting of the readings you have to record each day on the ward. Lastly, there are suggestions for further reading and some internet sites you can visit.

We wish you good luck with your studies!

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INTRODUCTION

Homeostasis

Human beings are made up of millions of cells which are organised into tissues, organs and systems. The conditions inside the body are constantly changing as we break down food and synthesise molecules in response to particular situations. However, for each cell to function properly the fluid that surrounds the cell (extracellular fluid) and the fluid inside the cell (intracellular fluid) have to stay roughly the same. All the parts of the body work together to keep the extracellular and intracellular fluids as close as possible to the best or optimum conditions for the cell. The preservation of this internal environment of the body is called homeostasis.

There are only about 15 litres (L) of extracellular fluid (including blood) in the body, which has to cope with eliminating waste products and transporting nutrients to all of the cells within the body. Moreover, the concentrations of ions such as potassium and hydrogen within a cell have also to be maintained by transport into and out of the same fluids. An enormous task! It is easy to understand that the levels of all these factors will be constantly changing but must be quickly corrected to keep the extracellular fluid at an optimum for cell function. An uncorrected effect in an individual cell can quickly affect surrounding cells and ultimately loss of whole organ function and death.

Feedback mechanisms

Homeostasis can be thought of as a constantly self-adjusting or dynamic system which keeps the extracellular fluid within a fairly narrow range of parameters. This is made possible by a number of feedback mechanisms. Each feedback mechanism has three parts: a set point, a sensory receptor and an effector.

The set point is used as a reference to compare the fluid passing through it to the 'norm' for that fluid. Many set points are found in the hypothalamus which controls the release of hormones from the pituitary gland. The hypothalamus is connected to the pituitary gland and is located in the forebrain beneath the cerebral cortex. Sometimes the set point is referred to as an integrator because it receives information and selects different information, for the effectors, to bring about change (Fig. 0.1). One example is the control of body temperature; sensory receptors in the skin will respond to the environmental temperature and send a signal to the set point (integrator) in the brain The integrator responds by sending out signals to the muscles and glands which cause the correct response to maintain homeostasis.

This type of control is called negative feedback and is used by the body in hundreds of situations. Its effect is to cancel out the change (either less or more of a component) so that the levels are returned to the normal range.

The same kind of system is used to control the temperature of central heating, ovens and irons, when one part doesn't work the whole system is affected. Generally the speed at which homeostatic negative feedback occurs is not very fast, for example changes in blood glucose levels rely on the circulation of the blood to register that the glucose levels are not normal. They are also not corrected very accurately and overshooting the mark is common, which will have to be corrected again by the effectors. These reactions are automatic and self-adjusting: we cannot control them voluntarily.

Nurses need to realise that 'normal' levels will be within an acceptable range and should

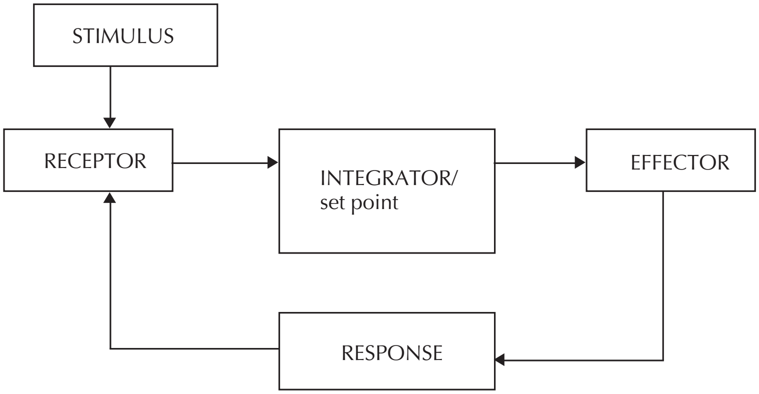


Fig. 0.1 Negative feedback mechanism.

be able to judge when changes occur. Taking the temperature of the patient is just one activity, which can indicate a homeostatic mechanism which is in or out of control.

Negative feedback control is used to control many different body functions which include respiration, breathing, and body temperature, but it is not the only type of control system, there are a few positive mechanisms as well.

Positive feedback increases an activity for a limited time to bring about a desired effect. A good example is the series of uterine contractions which expel the baby from the uterus in childbirth. This can be thought of as an unstable homeostatic state that lasts for a relatively short time. The fetus (stimulus) presses on the walls of the uterus and causes the pituitary gland to stimulate the production and secretion of oxytocin, which makes the muscles in the uterus wall contract, increasing the pressure on the fetus until it is expelled.

Many positive feedback mechanisms are harmful, e.g. cardiogenic shock (reduced blood flow to organs due to heart attack). During a heart attack less blood is circulated to the tissues, which become starved of oxygen. The tissues respond by changing their metabolism so that less oxygen is required. This leads to an increase in the volume of fluid bathing the cells (interstitial fluid) and a decrease in the volume of circulating blood. The positive feedback loop continues and makes the condition worse until it becomes irreversible and death results (Fig. 0.2).

Unstable homeostatic states

When homeostatic mechanisms can no longer function, an unstable state or a stressed state predominates. Stress can be temporary (acute), e.g. injury or shock, or continual (chronic) as in long term disease. As we get older we become increasingly unable to respond to stress because our homeostatic mechanisms do not seem to function as well.

Many of the activities of a nurse are about restoring a homeostatic state to the patient who has become ill because they can no longer control their internal environment. Helping the patient to restore normal body temperature, fluid volume, chemical composition and electrolyte concentration, nutrients or the correct pH is a vital role for the nurse. Techniques and practices are constantly changing to accommodate new knowledge and better patient care but all have the same objective – to achieve the set point for the patient.

Measurement and units

Blood pressure, pulse, temperature, height and weight are a few of the measurements

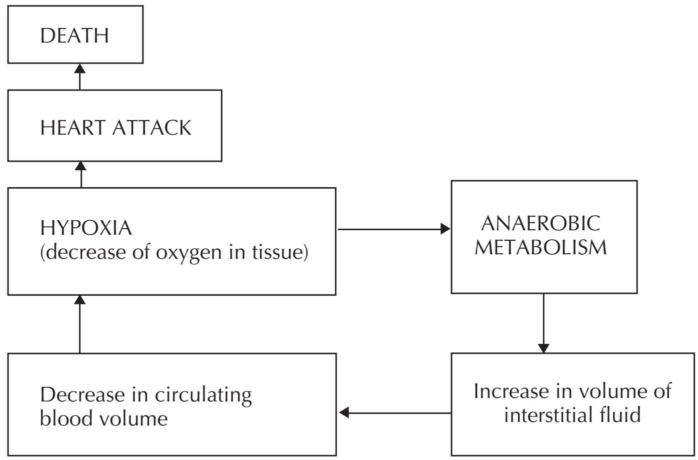


Fig. 0.2 Positive feedback – cardiogenic shock.

and recordings done by a nurse every day. There is always a need for accurate measurement and accurate recording. Records need to be interpreted by other staff so what is recorded must be clear and unambiguous. Continuous monitoring of a patient enables a good patient history to be built up and allows any changes in patient condition to be picked up early. It is equally important to understand measurements correctly in providing doses of medication.

Qualitative and quantitative measurement

There are three different types of measurement:

- **Qualitative** measurements are based on visual signs or the absence of signs but do not involve numbers. They rely on the observer's judgement so they are subjective.
- **Semi-qualitative** measurements record information on a relative scale. They are numerical estimates of an observation, for example a score of 5.3 in competitive ice-skating.
- **Quantitative** measurement – numbers are used to state the quantity which has been measured. This will normally require the use of an instrument to make the measurement. Examples are the measurement of height, weight, temperature and blood pressure. Most measurements of this type also require a unit to define the measurement. For example, it is not adequate to express body weight as 65 – do we mean pounds or kilograms? If it is kilograms then we must express the body weight as 65 kg to avoid any ambiguity! This type of measurement can be considered to be objective as it is based on a universally accepted scale.

An example of qualitative signs being converted to relative numerical values is the Apgar score used to indicate the condition of a new infant (Table 0.1).

A score is given for each of the five signs 1 and 5 minutes after the birth. A score of seven or above is considered satisfactory. Scores lower than seven indicate that some resuscitation may be required. In such cases, scores at subsequent 5-minute intervals would indicate the extent of recovery. In practice, the midwife will probably simply estimate the score without going through the whole process. Experience allows this to be done with accuracy.

Table 0.1 Apgar score

Sign	Score		
	0	1	2
Heart rate (beats min^{-1})	Absent	Slow (below 100)	Over 100
Respiratory effort	Absent	Slow, irregular	Good, crying
Muscle tone	Flaccid	Some flexion of extremities	Active motion
Reflex irritability	No response	Grimace	Cough or sneeze
Colour	Blue, pale	Body pink, extremities blue	Completely pink

Recording numerical values

A few important points need to be borne in mind.

- When expressing values of less than one, the decimal point should always be preceded by a zero; hence 0.5 and not .5 should be written. This is simply a practical point to ensure clarity (in this case to avoid the number being read hurriedly as 5 which could be disastrous if it led to an error in drug dosage).
- You may be required to quote a value to 'so many' decimal places. This refers to the number of digits after the decimal place. So, 156.17 is quoted to two decimal places, 37.9 to one decimal place.
- The rule for rounding off to one less decimal place is that if the last number is 5 or more then add one to the previous decimal place, so 0.825 becomes 0.83 but 0.824 becomes 0.82.
- How the results of a measurement or calculation are expressed must reflect the accuracy of the least accurate individual measurement that has been taken. The concept of significant figures is used to indicate the numbers that are accurate. All non-zero numbers are significant but zeros at the end of a number are not unless they are followed by a decimal point. Look at the following simple calculations:

$200 \times 1.62 = 324$ but as the value 200 is only given to one significant figure, the answer should be given as 300!

$200.0 \times 1.62 = 324$ again and both 200.0 and 1.62 are given to three significant figures so this time the answer can be given as 324!

Units of measurement

As we saw above, when measuring (for example, mass, length or time) we need to indicate which system of units has been used for the measurement. The two common systems of units used in the United Kingdom are SI (Système International) and the more traditional imperial units. The simpler and more logical SI system is increasingly used in the hospital environment but there is a need to be aware of both types of units. For example, blood pressure is still quoted in mmHg. Conversion from one system to another can be achieved using conversion tables or simple conversion formulae. Some commonly used conversion factors are given in Table 0.2:

So for example to convert 1 pound (16 ounces) into grams you simply multiply 16 by the conversion factor (28.35) to give 453.6 g.

Table 0.2 Conversion to SI units

Non-SI unit	Equivalent SI unit	Conversion factor
pound (lb)	kilogram (kg);	0.454
yard	metre (m);	0.914
ounce (oz)	gram (g)	28.359
mm Hg (pressure)	pascal (Pa)	133.30
pounds per square inch (psi)	pascal (Pa)	6895.156
calorie	joule (J)	4.184

Table 0.3 Base units

Quantity	SI unit
length	metre (m)
mass	kilogram (kg)
time	second (s)
current	ampere (A)
temperature	kelvin (K)
amount of substance	mole (mol)

Table 0.4 Common derived units

Physical Quantity	SI unit
force	newton (N)
energy	joule (J)
pressure	pascal (Pa)
potential difference	volt (V)
frequency	hertz (Hz)
volume	cubic metres (m^3)

Exercise 0.1

Convert:
a) 75 pounds to kilograms
b) 110 yards to metres
c) 1500 calories to joules
Express each answer to two significant figures.

A distinction is sometimes made between base (fundamental) and derived units, the latter are units of quantity that require the measurement of more than one of the base units. Examples are given in Tables 0.3 and 0.4.

Expressing derived units

There are a number of different ways we can write some derived units. You should be aware of their equivalence. For example, speed (calculated as distance divided by time) can be written in each of the following ways: metres per second, m/s or $m\,s^{-1}$. The reciprocal index ($^{-1}$) is used to indicate that time (s) is the divisor. Similarly, pressure can be expressed as newtons per square metre or N/m^2 or $N\,m^{-2}$.

Dealing with large and small numbers

In science we often have to deal with very large or very small numbers. For example the number of particles in a mole is approximately 600 000 000 000 000 000 000 000, X-rays have a wavelength of about 0.000 000 000 01 metres! These are very cumbersome numbers to write down and it would be difficult to ensure you wrote down exactly the right number of noughts.

Expressing these sorts of numbers is far easier in the standard form based on the use of indices. Consider multiples of 10:

100	$= 10 \times 10$
	$= 10^2$
1000	$= 10 \times 10 \times 10$
	$= 10^3$
10 000	$= 10 \times 10 \times 10 \times 10$
	$= 10^4$
100 000	$= 10 \times 10 \times 10 \times 10 \times 10$
	$= 10^5$
1 000 000	$= 10 \times 10 \times 10 \times 10 \times 10 \times 10$
	$= 10^6$

and so on. Note that the figure above the ten ('ten to the power of' or the index) is equal to the number of zeros in the number written out longhand and the number of times 10 is multiplied by itself. Hence:

2500	$= 2.5 \times 1000$	$= 2.5 \times 10^3$
697 000	$= 6.97 \times 100\,000$	$= 6.97 \times 10^5$

and we have a convenient and compact way of dealing with these types of figures.

Small numbers (less than 1)

Numbers with a value of less than one can be expressed as follows, using the reciprocal index:

0.1	1/10	$= 1.0 \times 10^{-1}$
0.01	1/100	$= 1.0 \times 10^{-2}$
0.001	1/1000	$= 1.0 \times 10^{-3}$
0.0037	37/1000	$= 3.7 \times 10^{-3}$
0.000 007 3	73/1 000 000	$= 7.3 \times 10^{-6}$

The negative sign in front of the index value indicates that the power of ten is dividing and not multiplying the given digit. As a rule of thumb, the index value indicates the number of places the decimal place has been moved to the left.

Exercise 0.2

Numbers written in the form 1.7×10^3 and 9.26×10^{-5} are said to be written in the standard form. Express the following numbers in the standard form:

a) 0.000 000 000 5
b) 301 000 000
c) 78 100 000
d) 128 400

Prefixes for SI units

When dealing with multiples and submultiples of the basic SI units, the prefixes in Table 0.5 are used. Since a lot of measurements associated with the body and medication are very small, a correct understanding and use of these prefixes is very important.

Simple conversions in SI units only require the use of the prefixes for multiples and submultiples.

$$\text{kilogram} \underset{\text{Divide by } 1000\,(10^3)}{\overset{\text{Multiply by } 1000\,(10^3)}{\rightleftarrows}} \text{gram} \underset{\text{Divide by } 1000\,(10^3)}{\overset{\text{Multiply by } 1000\,(10^3)}{\rightleftarrows}} \text{milligram}$$

Table 0.5 Prefixes for SI units

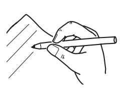
Prefix	Symbol	Multiple	Standard notation
mega	M	one million	10^{6}
kilo	k	one thousand	10^{3}
deci	d	one tenth	10^{-1}
centi	c	one hundredth	10^{-2}
milli	m	one thousandth	10^{-3}
micro	mc	one millionth	10^{-6}
nano	n	one thousand millionth	10^{-9}

So, 1 **kilo**gram = 1000 grams
1 **milli**gram = 10^{-3} gram (0.001 gram)

To convert from milligrams to micrograms, the table above indicates that there are 10^3 micrograms in 1 milligram (why?), so 1000 or 10^3 is the factor for this conversion. One point to note is that in scientific notation a microgram is abbreviated to µg. However for greater clarity in nursing work the abbreviation mcg or the full name should be used.

Measuring volumes

Although the cubic metre is the correct SI unit of volume the most commonly used metric measurement of volume is the litre (L) or millilitre (mL). One litre is equal to 1000 cm^3 so you can see that 1 cm^3 is also equivalent to 1 mL. It may also be useful to know that a litre is approximately equal to 0.11 gallons and 2.2 pints.



Exercise 0.3

Express:
a) 1×10^{-6} g in micrograms (mcg)
b) 474 joules as kilojoules (kJ)
c) 700 millilitres (mL) as litres (L)

Percentages, fractions and ratios

These can be used to express the same information in a variety of forms. The term ‘per cent’ means ‘in each hundred’ so a percentage value expresses the number of parts in a hundred parts of whatever is being measured. Hence if we state that 4% of the population have the AB blood group we are saying that four in every hundred people have that type of blood. The same statistic can be expressed as a fraction – 4/100 that we can simplify to 1/25 by dividing both numbers in the fraction by four. This fraction can be converted back to a percentage by multiplying the fraction by 100. The general equation for calculating a percentage is:

Percentage value

$$= \frac{\text{value of item being measured}}{\text{total number of items in system}} \times 100$$

Ratios can be also used to express quantities in a similar way: 4% can also be expressed as the ratio 4 : 96. In this case we are saying that there are 4 people with the AB blood group to every 96 people with other blood groups in a population of 100. Ratios can be simplified in the same way as fractions by dividing each value by the same factor. So, 1 : 25 states the same ratio as 4 : 96 or 4%.

Percentages and ratios are often used in describing the concentration of solutions. Hence a 1% solution of sodium citrate contains one part of sodium citrate (the solute) in a hundred parts of solution. The value is independent of the units of measurement. So, 5 g citric acid in 250 g of solution would be described by the following percentage:

$$\frac{5}{250} \times 100 = 2\%$$

2 kg (2000 g) of this 2% solution would contain 2000 × 2/100 = 40 g sodium citrate.

You may also find very low concentrations described in parts per million (ppm). A 1 ppm solution contains one part of solute dissolved in one million parts of solution.

Exercise 0.4

a) Change these fractions to percentages: 3/5, 17/20, 9/25
b) Change these ratios to the form 1 in 'x' and then write as percentages: 1 : 24, 1 : 9, 1 : 40

Many simple drug calculations can be worked out by using the formula below which uses the fraction of the dose required provided by each measure (pill, spoonful or solution of known concentration) to identify the number of measures to be given.

Number of measures to be given

$$= \frac{\text{dose prescribed}}{\text{dose per measure}}$$

For example, a patient is ordered 0.25 mg of digioxin, orally. The digioxin available is in tablets containing 125 micrograms (mcg). First we have to recognise that we *must* be consistent with our units – the 0.25 mg dose prescribed is equivalent to 250 mcg. Then, substituting the values in the formula:

Number of measures to be given

$$= \frac{\text{dose prescribed}}{\text{dose per measure}} = \frac{250}{125} = 2 \text{ tablets}$$

Exercise 0.5

An injection of morphine 8 mg is required. Ampoules available contain 10 mg in 2 ml. What volume is drawn up for injection? Hint: use the formula above, the fraction produced will be the fraction of the ampoule required.

Averages

Averages allow you to express a series of values or readings as a single, meaningful value. The most common type of average value is the mean. It is worked out as follows:

$$\text{Mean} = \frac{\text{sum of all the readings}}{\text{number of readings taken}}$$

For example, the seven nurses on a ward have ages of 22, 26, 30, 32, 33, 39 and 51 respectively. The average (mean) age is:

$$\frac{22 + 26 + 30 + 32 + 33 + 39 + 51}{7} = 33$$

We might also quote the median, the middle value in the range of figures. This is 32 in this case, not necessarily the same as the mean value.

Exercise 0.6

Calculate the average weekly temperature (°C) of a patient from the following data. To how many decimal places should you express your answer?

Mon	Tues	Wed	Thurs	Fri	Sat	Sun
36.4	36.8	37.0	37.1	37.5	37.9	36.9

Graphs and conversion charts

Numerical information is often conveniently presented in graphs or conversion charts. In Fig. 0.3 the data from the exercise above have been plotted graphically; time (days of the week) on the x-axis and temperature on the y-axis. This type of presentation of data enables us to see trends at a glance as well as take readings directly from the graph. For example, the value plotted for Wednesday's temperature is simply read off on the y-axis as 37.0 °C. It can immediately be seen that the patient's temperature rose steadily until Saturday and then dropped back down again.

Summary

- Quantitative measurements are based on numerical figures, qualitative measurement on other visual signs

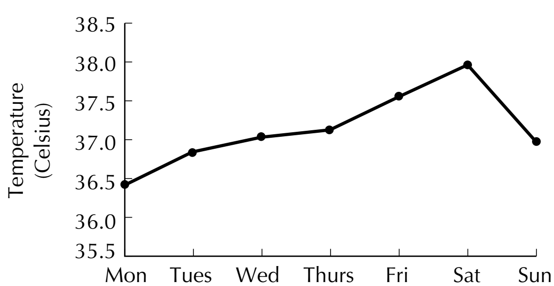


Fig. 0.3 Temperature plotted against time.

- The SI system of units has now largely replaced the 'older' imperial measures. If required, interconversion between these units can be achieved using simple conversion factors
- The 'standard form' is important in expressing large and small quantities in a convenient way. The use of prefixes to SI units can be used to achieve the same aim
- Percentages, fractions and ratios can be readily and simply interconverted
- Averages can provide a single meaningful figure to represent a series of measurements
- The graphical representation of data is a useful and convenient way of presenting information.

Review questions

1. Steve Jones cycles to work every morning. His resting pulse rate is $72\,\mathrm{min}^{-1}$ but when he arrives at work this has increased to $81\,\mathrm{min}^{-1}$. What is the percentage increase in his pulse rate?

2. Blood makes up 7% of the body weight. If each of the following people donated one pint of blood, what percentage of their blood has each person given:
   (a) Polly 50 kg
   (b) Joyce 10 stone
   1 pint of blood weighs 600 g; 1 stone = 6.35 kg

3. A patient requires 1.250 g of an antibiotic that is supplied in tablet form, each containing 250 mg of the antibiotic. How many tablets are needed for the correct dose?

4. Fig. 0.4 illustrates the effect of drinking 1 L of water on the rate of urine production. From the graph, state the time in minutes that the maximum amount of urine is produced. After 120 min, what is the volume of urine being produced?

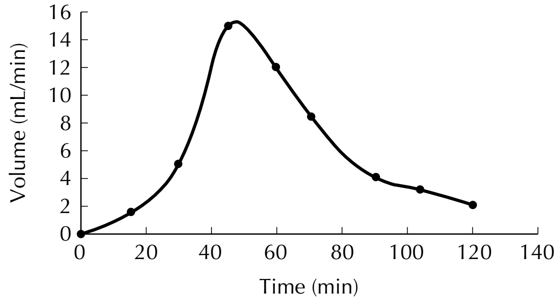


Fig. 0.4 Urine production plotted against time.

Chapter One
ATOMS

Learning objectives

By the end of the chapter the reader will be able to:
- Identify the states of matter
- Describe the structure of the atom in terms of electrons, protons and neutrons
- Define the terms element, molecule, ion, radical, compound, atomic number, mass number and isotope
- List the main elements contained in the body
- Explain how ionic and covalent bonds are formed
- Distinguish between polar and non-polar molecules
- Explain the concept of the mole.

Introduction

Atoms, the compounds they form and the chemical reactions which they undergo form the basis of life. All living creatures (and that includes you!) are essentially a collection of atoms – a very complex and organised collection of atoms, but still atoms. Any understanding of human physiology must therefore go hand in hand with an understanding of the amazing atom.

Matter

Everything in the universe is made up of matter. Matter is anything that occupies space and has mass. All matter is made up of particles. There are three states of matter: solid, liquid and gas (Fig. 1.1).

Changes of state

When a solid is heated it changes state and forms a liquid. If you continue to heat the liquid, then it will change state again, to form a gas. Some substances can go from the solid state to the gaseous state directly, for example iodine, this is known as sublimation.

All particles possess kinetic energy (energy of movement). Solid particles have a small amount of kinetic energy, liquid particles more and gas particles the most. Changes of state always involve the addition or removal of energy, normally in the form of heat (Figs 1.2 & 1.3).

The core body temperature, in humans, is 37 °C. Body temperature is altered by illness, exercise or climatic changes. One of the mechanisms for maintaining body temperature is sweating. Up to 12 L of water a day can be lost through sweating. Sweat evaporates from the surface of the skin cooling the body. Heat is removed from the body by changing the state of sweat from a liquid to a gas.

State	**Arrangement of Particles**
SOLID	Solids have a definite shape and volume. Very strong forces of attraction between particles. Particles only vibrate about a point. Solids are rigid.
LIQUID	Liquids flow. They have a definite volume, but their shape depends on the container. Weak forces of attraction between the particles. Particles move and slide over each other.
GAS	Gases flow. Gases have no definite volume or shape and they fill all the space available. Almost no forces of attraction between particles. Particles move in a random manner colliding with other particles and the walls of the container.

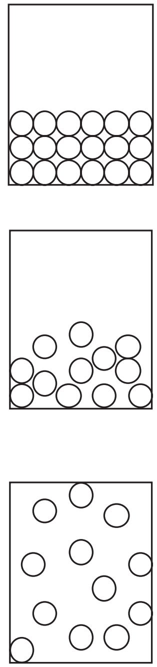


Fig. 1.1 States of matter.

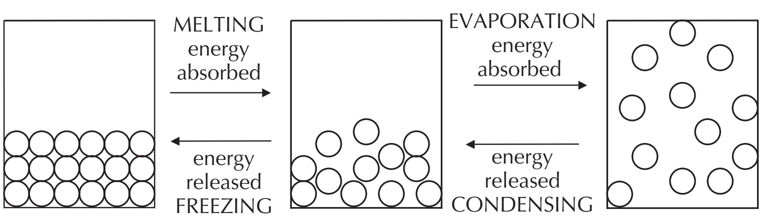


Fig. 1.2 Changes of state.

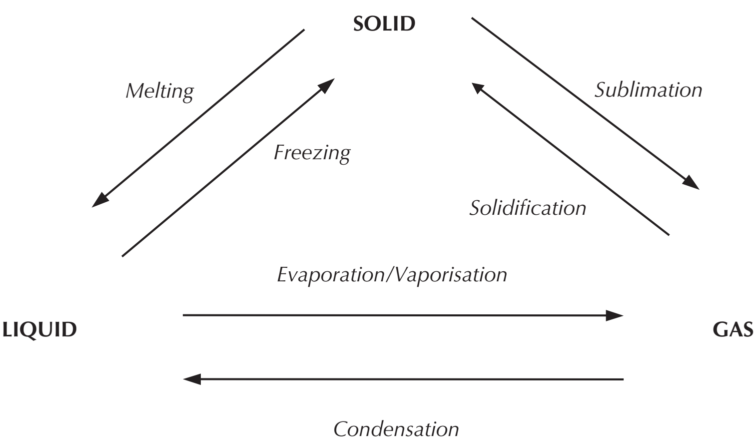


Fig. 1.3 Summary of states of matter.

Atoms

About 2500 years ago a Greek philosopher called Democritus suggested that all matter in the universe is made up of tiny, indivisible particles. In 1807 John Dalton called these tiny particles ‘atoms’, from the Greek word *atomos* meaning indivisible. We now know that atoms are not indivisible, the nuclear or atomic bomb being the result of splitting the atom. However the atom still remains the smallest particle that can exist independently and the smallest unit that can enter a chemical reaction.

Atoms are approximately 1/100 000 000 of a centimetre across; 100 million atoms side by side would therefore measure 1 cm (Fig. 1.4).

Atomic structure

Atoms are made up of protons, neutrons and electrons – these are called sub-atomic particles (Fig. 1.5).

The central nucleus contains protons and neutrons. The mass of the atom is concentrated in the central nucleus and is very small. Moving around the nucleus are electrons. The electrons are arranged in orbitals or shells or energy levels at different distances from the nucleus. The mass of the electron is so small it can be ignored when working out the mass

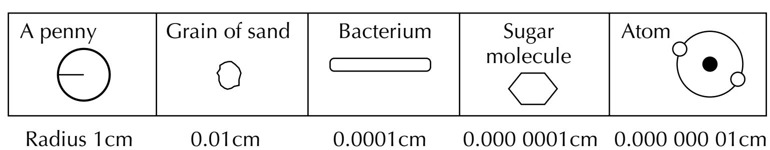


Fig. 1.4 Relative size of an atom.

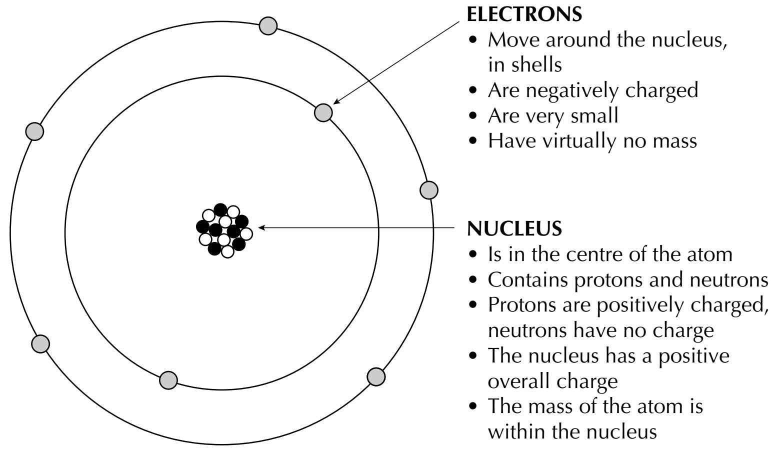


Fig. 1.5 Atomic structure.

of the atom. Chemical reactions involve only electrons (Table 1.1).

Elements

An element is a substance in which all the atoms have the same chemical properties. Over 100 elements exist. A Russian scientist called Dmitri Mendeléev conveniently arranged these into a table called the periodic table (Fig. 1.6 & p. 227). He arranged the elements in horizontal rows – periods and vertical columns – groups. It is called the periodic table since the structure and chemical reactivity of the elements arranged in the rows, or periods, show repeating or periodic characteristics.

Each element in the periodic table is represented by a symbol. Most elements have two letters in their symbol; the first letter is a capital, and the second letter is *always* small. If there is only one letter this is always a capital. These symbols come from either the English name (C for carbon) or the Latin (Na for natrium, English sodium; K for kalium, English potassium). Medical terminology often uses Latin terms so hypernatraemia and hyperkalaemia are used to indicate high levels of sodium and potassium in the body.

Table 1.1 Sub-atomic particles

Particle	Position	Mass (relative to that of a proton)	Charge (relative to that on a proton)
Proton	Nucleus	1	+1
Neutron	Nucleus	1	0
Electron	Shells	$^{1}/_{1840}$	–1

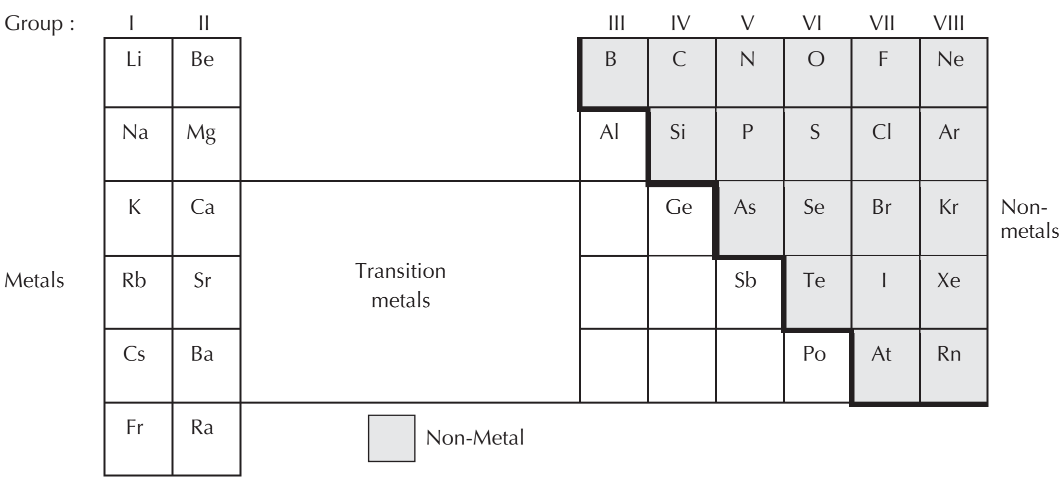


Fig. 1.6 Abridged periodic table, showing groups, metals and non-metals.

Activity

Using the periodic table provided on p. 227 find the symbols that represent the following elements:

Element	Symbol
Mercury	
Iron	
Gold	
Lead	

Classification of the elements

Elements can be conveniently divided into two broad categories; metals and non-metals. This can be done on the basis of their properties (Table 1.2).

Metals are found on the left-hand side of the periodic table, e.g. sodium and calcium and non-metals are found on the right-hand side, e.g. oxygen and chlorine.

Elements in the body

The number of elements found in the body is small. In fact 95% of a living organism is made up of only four elements – carbon, oxygen, hydrogen and nitrogen (Fig. 1.7).

Trace elements for example, iron (Fe), iodine (I), copper (Cu), chromium (Cr) and zinc (Zn) are

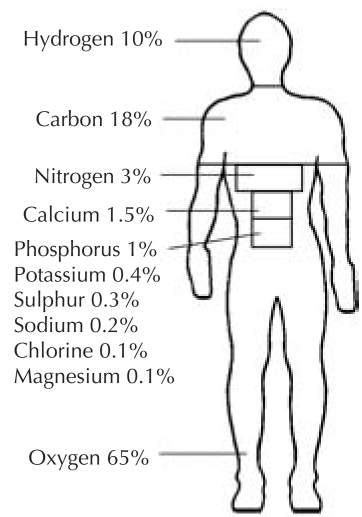


Fig. 1.7 Elements in the body.

Table 1.2 Classification of the elements

Property	Metals	Non-metals
physical state	solid (except mercury)	solid, liquid or gas
lustre	shiny	dull
boiling point/melting point	high	low
conductors of heat/electricity	good	poor, good insulators
affinity for electrons	donates electrons	accepts electrons

needed in very small quantities by the body. Many are found as parts of enzymes or are needed for enzyme activity. In excess they can be toxic.

Toxic elements e.g. mercury (Hg), lead (Pb), arsenic (As) and thallium (Tl) can all cause death, their toxic effect is cumulative.

> *Activity*
>
> **Mercury is a toxic element and can lead to death. In hospitals it is commonly used in thermometers and sphygmomanometers. Look up the symptoms of mercury poisoning and the type of treatment given to such a patient.**

Atomic number and mass number

The atomic number is the number of protons that the atom contains (Fig. 1.8). In a neutral atom, this is also the same as the number of electrons. The elements are arranged in order of their increasing atomic number in the periodic table.

The mass number is the number of protons, plus the number of neutrons (Fig. 1.8). The mass number is always larger, roughly twice the size of the atomic number.

Isotopes

Isotopes are atoms of the same element that have the same atomic number but different mass numbers. Isotopes have a different numbers of neutrons in their nucleus. For example iodine:

	$^{123}_{54}I$	$^{125}_{54}I$	$^{127}_{54}I$	$^{131}_{54}I$
protons	*54*	*54*	*54*	*54*
electrons	*54*	*54*	*54*	*54*
neutrons	*69*	*71*	*73*	*77*

Isotopes which are radioactive are called radioisotopes. These are covered in more detail in Chapter 11. Isotopes can be written as either ^{131}I or iodine-131.

Isotopes have the same chemical properties, since this is dependent on the number of electrons in an atom, but they will have different physical properties (e.g. boiling and melting points) as their masses are different.

The element iodine is essential for the production of the hormone thyroxine by the thyroid gland. Radioactive iodine-131 is used as a diagnostic test for hyperthyroidism (overactive thyroid), e.g. Graves disease, or hypothyroidism (myxoedema – underactive thyroid) and thyroid cancer. The patient is given a solution of sodium iodide containing a small amount (trace) of iodine-131. The quantity of iodine-131 absorbed by the thyroid gland or excreted in the urine is measured using a gamma detector.

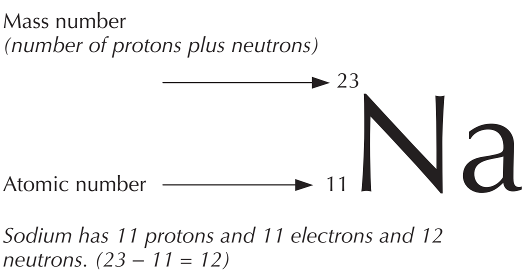


Fig. 1.8 Atomic and mass numbers.

Relative atomic mass

An atom of the carbon-12 isotope was given a relative atomic mass of 12.0000u (u = atomic mass unit). The relative masses of all other atoms were then obtained by comparison to this standard atom. The relative atomic mass is also called 'atomic mass', 'average atomic mass' and in the past 'atomic weight'.

Most elements contain a mixture of isotopes. The relative atomic mass of an element is the average mass of one atom, taking account of all its isotopes and their relative proportions compared with an atom of carbon-12.

For example chlorine consists of two isotopes with mass numbers of 35 and 37. These isotopes are written as chlorine-35 and chlorine-37 respectively. Naturally occurring chlorine contains 75% of chlorine-35 and 25% of chlorine-37. The relative atomic mass of chlorine is:

$$\frac{(75 \times 35) + (25 \times 37)}{100} = 35.5\ \text{u}$$

Most relative atomic masses are not whole numbers because most elements have many isotopes (Table 1.3).

Electron arrangement

Electrons are arranged in shells around the nucleus. The number of electrons in a neutral atom is given by the atomic number. The rules for filling the shells are:

Table 1.3 The relative atomic masses of some elements

Element	Symbol	Relative atomic mass
Carbon	C	12.011
Hydrogen	H	1.008
Magnesium	Mg	24.312
Sodium	Na	22.989
Sulphur	S	32.064

1. The shell nearest the nucleus fills first
2. Only a given number of electrons are allowed in each shell

Examples

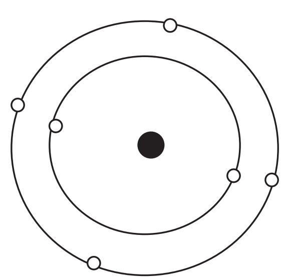
1st shell 2   2nd shell 8   3rd shell 18

The diagrams like those in Figs 1.9 and 1.10 are quite cumbersome to draw each time, so a shorthand form is often used to show how electrons are arranged – this is called the electronic configuration.

- The electronic configuration for carbon is written as C 2, 4
- The electronic configuration for sodium is written as Na 2, 8, 1

Exercise 1.1

The atomic number of lithium is 3. Draw a diagram to show the electronic arrangement and write the electronic configuration for this atom.



CARBON – atomic number 6
It has 6 electrons,
2 occupy the first shell (making it full),
leaving 4 in the second shell.

Fig. 1.9 Electronic arrangement of carbon.

SODIUM – atomic number 11
It has 11 electrons,
2 occupy the first shell (making it full),
8 occupy the second shell (making it full),
leaving 1 in the third shell.

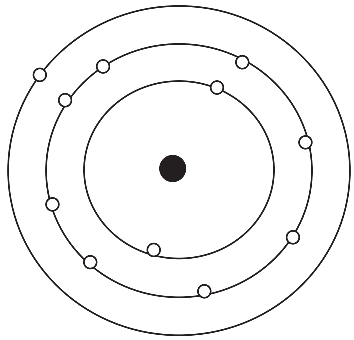


Fig. 1.10 Electronic arrangement of sodium.

Compounds

Jilly was a chemist, but Jilly is no more,
For, what she thought was H_2O, was H_2SO_4.

A compound is a substance that contains two or more elements joined together by a chemical bond (Table 1.4). The property of the compound will usually be very different from those of the individual elements in it,

e.g.	sodium *reacts violently with water*	+	chlorine *corrosive gas*	→	sodium chloride (table salt) *nice with chips, yummy!*

Table 1.4 Simple compounds

Name of compound	Elements found and their symbols	How the atoms join up (structural formula)	Formula of compound
Water	Hydrogen (H) and oxygen (O)	H–O–H	H_2O
Carbon monoxide	Carbon (C) and oxygen (O)	C≡O	CO
Ethanol (alcohol)	Carbon (C) and oxygen (O) and hydrogen (H)	H–CH_2–CH_2–OH	C_2H_5OH
Ammonia	Nitrogen (N) and hydrogen (H)	H–N–H │ H	NH_3

Compounds are generally named after the elements that reacted to form them. A shorthand way of describing a compound is by writing its name as a chemical formula. The formula of a compound indicates the number and type of each atom present, e.g. carbon dioxide contains one carbon atom and two oxygen atoms (indicated by the prefix 'di') and its formula is CO_2 (note the number 2 is written as a subscript).

Chemical equations

Chemical equations are a way of representing chemical changes which occur when elements