Appendix 1

Background chemistry for biologists

THE ISSUES

- Elements, atoms, molecules and compounds
- How atoms form molecules
- Ions
- Why carbon is important to life
- Chemical equations
- Acids and bases
- pH and buffers
- Isotopes
- Oxidation and reduction
- Hydrolysis and condensation
- Molarity and concentration

Elements, atoms, molecules and compounds

Elements are the basic units of pure substances that make up our entire world. The Earth is composed of about 92 stable elements in varying quantities, including substances like carbon, hydrogen, calcium and sodium. The names of the elements are usually written down as a **chemical symbol** – a convenient shorthand for the full name. For example, the symbol for carbon is C, for hydrogen it is H, for calcium it is Ca, and for sodium it is Na (derived from its Latin name of natrium).

Most elements are divided into two groups – **metals** and **non-metals** – based on their appearances, states and properties (Table A1.1). Living things are built from both metal and non-metallic elements, but it is non-metallic elements that make up the bulk. The range of elements that compose living things is quite different from the composition of the environment around them. The four most common elements of the Earth's crust are oxygen (O), silicon (Si), aluminium (Al) and iron (Fe), whereas the bulk of living things (about 99%) is formed from hydrogen, oxygen, carbon and nitrogen (N) (Table A1.2).

Table A1.1 Metallic and non-metallic elements

Non-metallic elements e.g. hydrogen, oxygen, carbon, silicon	Metallic elements e.g. calcium, potassium, magnesium, iron
usually solids or gases (at room temperature)	usually solids (at room temperature)
solids appear dull	have a shiny appearance
low density, with a low melting point	dense, with a high melting point
usually do not conduct electricity	conduct electricity

Table A1.2 The relative percentage abundance of the elements of living things

Element	Symbol	Organisms	Earth's crust
hydrogen	H	49.0	0.22
carbon	C	25.0	0.19
oxygen	O	25.0	47.0
nitrogen	N	0.27	<0.1
calcium	Ca	0.073	3.5
potassium	K	0.046	2.5
silicon	Si	0.033	28.0
magnesium	Mg	0.031	2.2
phosphorus	P	0.030	<0.1
sodium	Na	0.015	2.5
others		traces	13.7

Trace elements include iron (Fe), manganese (Mn), copper (Cu), cobalt (Co) and zinc (Zn)
Iron (Fe) makes up almost 5% of the Earth's crust.

Table adapted from: Wood, E. J. and Pickering, W. R. (1982), *Introducing Biochemistry*, John Murray (page 7).

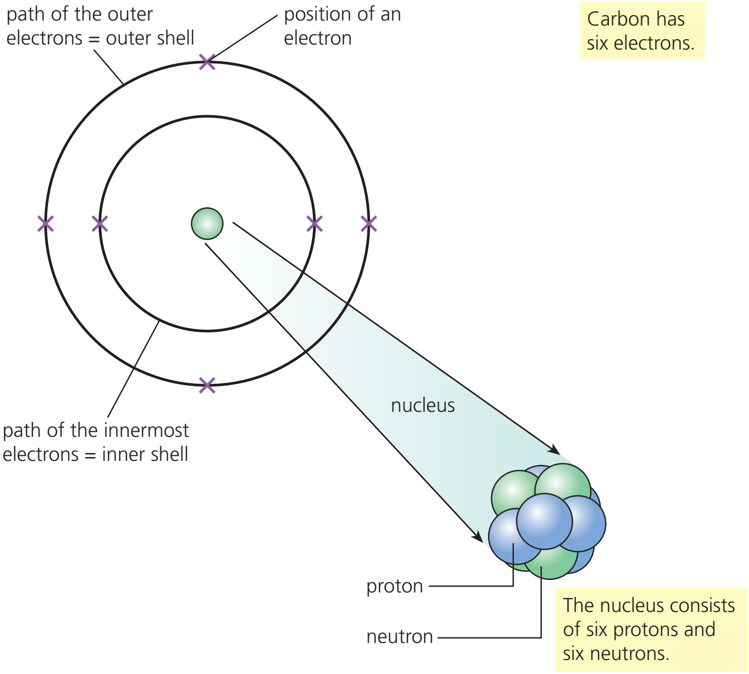
An **atom** is the smallest part of an element that can take part in a chemical change. The atoms of different elements are of different sizes, but all atoms are incredibly small. Because of the small size of atoms we cannot refer to their mass by a standard unit, like the gram, for example. Instead, we compare the mass of an atom relative to an agreed standard. For this purpose the reference atom is that of carbon. The carbon atom is given a relative atomic mass* of 12 ($A_r = 12$). By comparison, atoms of hydrogen ($A_r = 1$) are much lighter than carbon, but atoms of nitrogen ($A_r = 14$) are slightly heavier, and atoms of potassium ($A_r = 39$) substantially heavier. Table A1.3 gives the relative atomic mass of the atoms of 13 common elements of living things.

Table A1.3 The relative atomic mass (A_r) of the atoms of 13 elements of living things

Element	Symbol	A_r
hydrogen	H	1
oxygen	O	16
carbon	C	12
nitrogen	N	14
phosphorus	P	31
sulfur	S	32
chlorine	Cl	35.5
sodium	Na	23
potassium	K	39
calcium	Ca	40
magnesium	Mg	24
iron	Fe	56
copper	Cu	63.5

The structure of an atom is shown in Figure A1.1. This is the carbon atom. You can see that this atom is made of three kinds of smaller particles.

Figure A1.1 The structure of a carbon atom



At the centre of the atom is the nucleus. This consists of protons (positively charged particles) and, usually, neutrons, which are uncharged particles. (The single exception is the nucleus of the hydrogen atom which contains no neutron, only a single proton.) Protons and neutrons have virtually the same mass. (Note that the nucleus of an atom must not be compared or confused with the cell nucleus.)

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*Note that the term 'relative atomic mass' replaces the former term 'atomic weight'.

Around the nucleus occur incredibly tiny particles called electrons (negatively charged particles), moving in orbits or 'shells'. Electrons have almost no mass at all. The final point to note is that the whole atom is electrically neutral. This is because the number of protons (positively charged) in the nucleus is equal to the number of electrons (negatively charged) in the surrounding orbits.

A **molecule** is the smallest part of most elements or compounds that can exist alone under normal conditions. For example, oxygen and nitrogen are not found as single atoms, but as pairs. Each atom naturally combines with another atom of the same type to form a molecule. So, molecular oxygen has the symbol O_2 and molecular nitrogen N_2, to indicate that there are two atoms in each case. Some molecules that form naturally contain more than one element, and so we call these compounds.

A **compound** is a molecule which contains two or more elements chemically combined together. For example, carbon dioxide (CO_2) contains atoms of carbon and oxygen, molecules of water (H_2O) contain atoms of hydrogen and oxygen, and molecules of common salt, sodium chloride (NaCl), contain atoms of sodium and chlorine.

How atoms form molecules

Atoms naturally combine together (they 'bond') to form molecules in ways that have a stable arrangement of electrons in the outer shells of each atom. Atoms are most stable when their outer shell of electrons is complete. The electron shell closest to the nucleus can hold up to two electrons and then it is full. Similarly, the second shell can hold up to a maximum of eight electrons, the third shell can hold up to 18 electrons, and the fourth shell can hold up to 32 electrons. (There are further shells in the largest atoms, but they do not concern us here because the elements that make up living things have atoms that are among the smallest and lightest in the Earth's crust.)

How are these bonds formed that achieve complete outer shells? This may happen in two different ways. Atoms *either* may share electrons (covalent bonding), *or* they may give and take electrons (ionic bonding).

In **covalent bonding**, electrons are shared between atoms. Covalent bonds are the strongest bonds occurring in biological molecules. This means that the greatest amount of energy has to be used to break them. So, covalent bonds provide great stability to biological molecules, many of which are very large and elongated. Bonding of this kind is common in non-metallic elements such as hydrogen, nitrogen, carbon and oxygen.

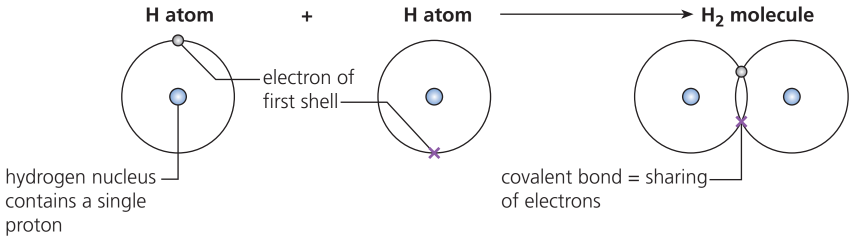
When one pair of electrons is shared, as in the hydrogen molecule, a single covalent bond is formed, known as a **single bond**. If two pairs of electrons are shared, then a **double bond** is formed. Double bonds are quite common between two carbon atoms (Figure A1.2), or between carbon and oxygen.

In **ionic bonding**, atoms gain or lose electrons to form positive or negative ions that have a stable arrangement of outer electrons (Figure A1.3). For example, sodium ions and chloride ions are much more stable than sodium atoms and chlorine atoms, because the sodium loses the single electron in the outer shell to form a complete shell, and the chlorine gains the electron to complete a third-level shell.

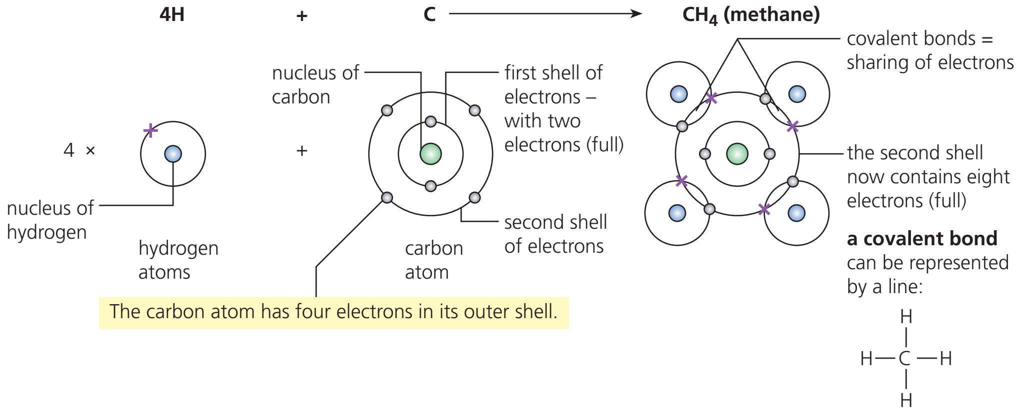
Ionic bonding is the electrostatic attraction between oppositely charged particles. The electrostatic attraction acts in all directions around the ion so, for example, the positive sodium ion is surrounded by negative chloride ions, while each chloride ion attracts and is surrounded by sodium ions. Because of this, in solid sodium chloride the ionic bonds hold the ions together in a regular arrangement (known as a **crystal lattice**), as seen in crystals of salt. In solution, however, the presence of water molecules, surrounding the sodium and chloride ions, causes them to be separated and dispersed.

Figure A1.2 Sharing electrons in covalent bonds

1 In the hydrogen molecule a covalent bond is formed.



2 In **methane**, four single bonds are formed with hydrogen atoms to make the outer shell of carbon up to eight electrons.



3 In **ethene (C_2H_4)**, the carbon atoms make up their outer shells to eight electrons by forming a double bond (with the other carbon atom) and two single bonds (with hydrogen atoms).

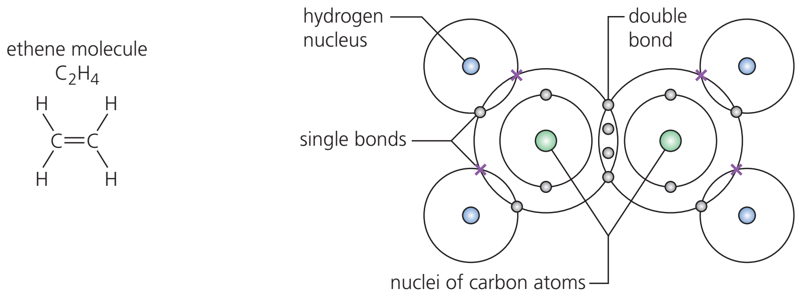
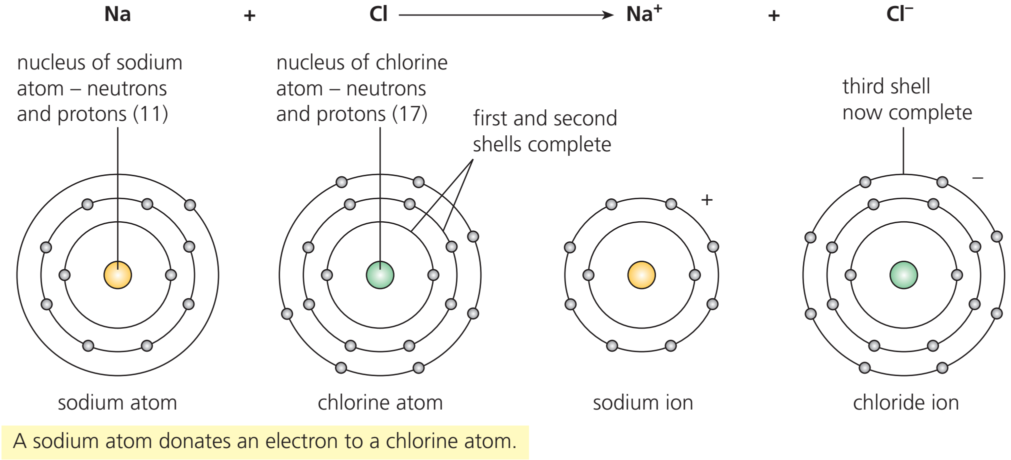


Figure A1.3 Transfer of electrons in ionic bonds



Ions

Ions are the charged particles formed in ionic bonding. They are very stable because their outermost electron shell is complete. Positively charged ions are called **cations** and negatively charged ions are called **anions**. Six ions are of particular biological importance, as these ions occur widely in living cells. Table A1.4 details some of their specialized roles in functioning cells.

Table A1.4 Some ions of importance in biological systems

Positively charged ions (cations)	
Na^+ (sodium ion)	involved in the setting up of the action potential of a nerve fibre and the flow of the action potential (impulse)
K^+ (potassium ion)	involved in the setting up of the action potential of a nerve fibre and the flow of the action potential (impulse)
Ca^{2+} (calcium ion)	involved in the contraction of the muscle myofibrils by combining with blocking molecules so that the myosin head of the cross bridge can attach to the actin
Negatively charged ions (anions)	
NO_3^- (nitrate ion)	plants reduce nitrate to ammonia and combine it with an organic acid forming an amino acid; from this come all 20 amino acids that are used to manufacture proteins
PO_4^{3-} (phosphate ion)	the phosphate ion is combined with ADP to form ATP; ATP is the primary energy currency of cells, involved in energy-requiring reactions and processes, like protein synthesis and muscle contraction
HCO_3^- (hydrogencarbonate ion)	this is the form in which carbon dioxide is transported in the blood (plasma and red cells); it is formed when CO_2 reacts with water, catalysed by carbonic anhydrase enzyme

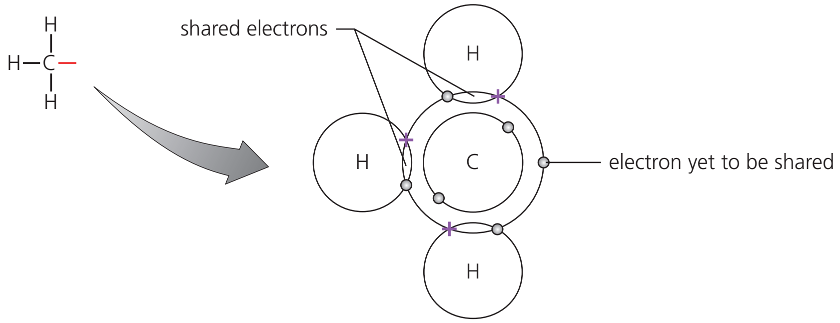
Why carbon is important to life

The importance of carbon lies in its unique collection of properties. These are so remarkable, in fact, that we can say that they make life possible.

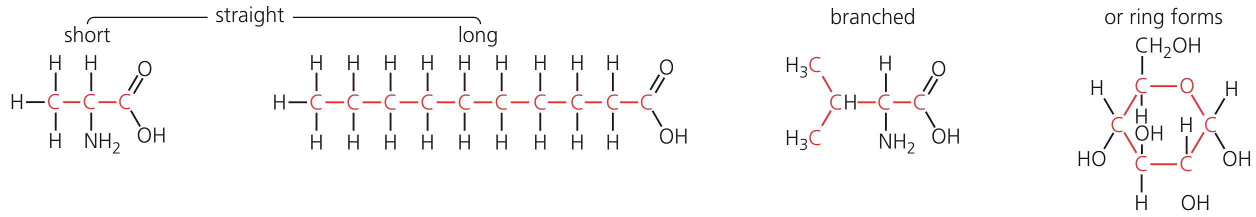
- Carbon is a relatively small atom and is able to form four strong, stable bonds. The bonds that carbon atoms form are covalent bonds. A great deal of energy has to be put into carbon molecules to break them.
- Carbon atoms are able to react with each other to form extended chains. These 'carbon skeletons' may be straight chains, branched chains or rings (Figure A1.4). So, a vast number of extremely stable carbon compounds exist – more than the combined total of known compounds of all the other elements, in fact. Biologists believe that this great wealth of carbon compounds has made possible the diversity of life as we know it.

Figure A1.4 Covalent bonds and carbon 'skeletons'

Covalent bonds are formed by sharing of electrons, one from the carbon atom and one from the neighbouring atom it reacts with.



Carbon atoms bond with other carbon atoms to form carbon 'skeletons'.



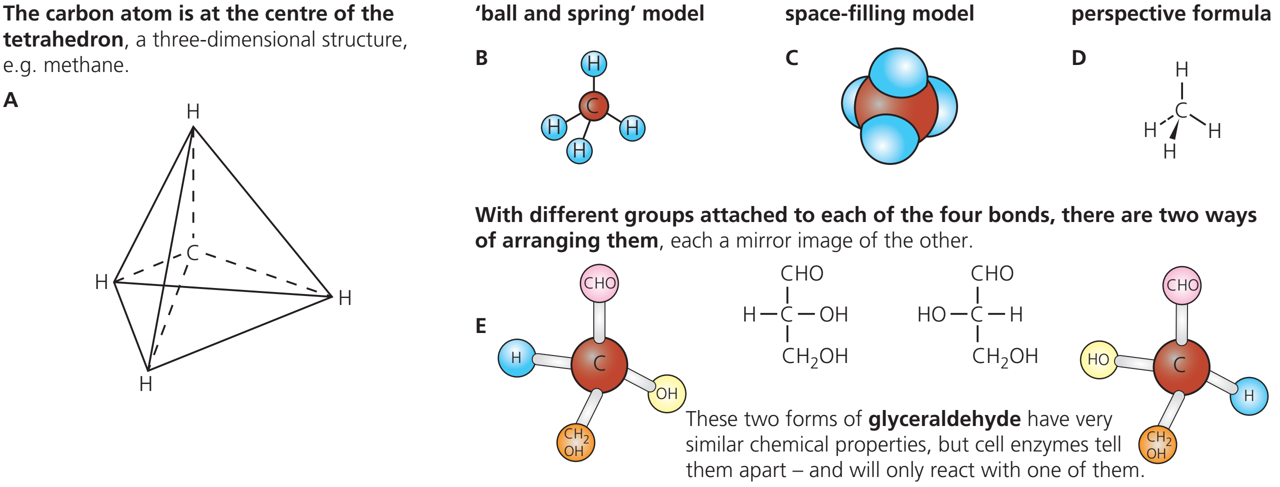
short chain (in the amino acid alanine)

long chain (in a fatty acid)

branched chain (in the amino acid valine)

the ring form (of α-glucose)

- The four covalent bonds of carbon atoms point to the corners of a regular **tetrahedron** (a pyramid with a triangular base) (Figure A1.5). This is because the four pairs of electrons repel each other, and so position themselves as far away from each other as is possible. If there are different groups attached to each of the four bonds around a carbon atom then there are two different ways of arranging the groups. This leads to forms of molecules that are 'mirror images' of each other. Carbon atoms with four different atoms or groups attached are said to be asymmetric. This is another cause of variety among organic molecules.



The carbon atom is at the centre of the tetrahedron, a three-dimensional structure, e.g. methane.

'ball and spring' model

space-filling model

perspective formula

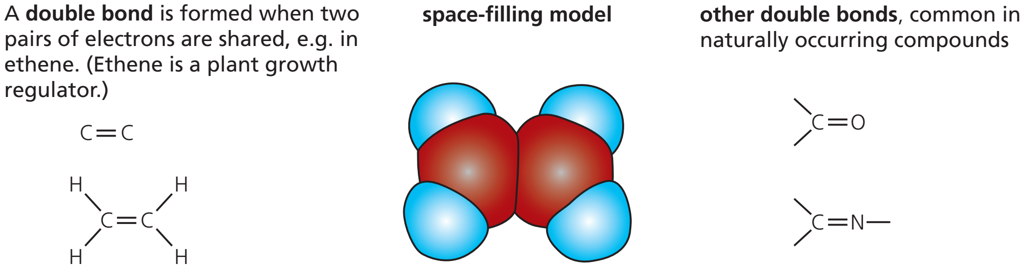
With different groups attached to each of the four bonds, there are two ways of arranging them, each a mirror image of the other.

These two forms of **glyceraldehyde** have very similar chemical properties, but cell enzymes tell them apart – and will only react with one of them.

Figure A1.5 The tetrahedral carbon atom

- Carbon atoms form covalent bonds with other atoms, such as oxygen, hydrogen, nitrogen and sulfur, forming different groups of organic molecules with distinctive properties.
- Carbon compounds, although numbering 2–3 million in total, fit into a relatively small number of 'families' of compounds. The families are identified by a part of their molecule called the **functional group**; this gives them their characteristic chemical properties. Chemical families we often come across in biology include the alcohols, organic acids, ketones and aldehydes. The remainder of the organic molecule, apart from the functional group, has little or no effect on the chemical properties of the functional group and is referred to as the **R group**.
- Carbon atoms can form more than one bond between them. For example, carbon may share two pairs of electrons to form a double bond (Figure A1.6). Carbon compounds that contain double carbon–carbon bonds are known to chemists as '**unsaturated**'; carbon compounds not containing double bonds are **saturated** compounds.

Figure A1.6 Double bonds or carbon═carbon bonds



A **double bond** is formed when two pairs of electrons are shared, e.g. in ethene. (Ethene is a plant growth regulator.)

space-filling model

other double bonds, common in naturally occurring compounds

In fact carbon, nitrogen and oxygen all form double bonds. This is another way in which variety is introduced to the range of carbon compounds that make up cells.

Because of these features, molecules containing a carbon skeleton exist in vast numbers.

Chemical equations

In a chemical reaction, the reacting molecules (**reactants**) are converted into new molecules (**products**). The change involved in the reaction is summarized in an equation. The process of constructing a chemical equation has three steps.

1 A **word equation** is written. Examples are:

hydrogen + oxygen → water

calcium + chlorine → calcium chloride

2 Below the word equation, the **formulae** for the substances involved are written using the symbols for the elements. Molecules must be shown containing the correct number of atoms. For this you need to know how many bonds an atom can form. This is known as the combining power or **valency** (Table A1.5).

hydrogen + oxygen → water
$H_2 + O_2 \rightarrow H_2O$
calcium + chlorine → calcium chloride
$Ca + Cl \rightarrow CaCl_2$

The valency of an atom, ion or group of atoms is needed for writing chemical formulae. Valency is the number of electrons used by an atom in bonding:

- the number of electrons lost by a cation in formation
- the number of electrons gained by an anion in formation
- the number of electrons shared by an atom in forming a covalent bond.

Table A1.5 Combining power or valencies

Valencies of covalent atoms					
1	**2**	**3**	**4**		
H	O	N	C		
	S	P			
Charges on some ions					
–3	–2	–1	+1	+2	+3
Single element ions					
		Cl^-	H^+	Ca^{2+}	
			Na^+	Mg^{2+}	
			K^+		
			Cu^+ (copper(I))	Cu^{2+} (copper(II))	
				Fe^{2+} (iron(II))	Fe^{3+} (iron(III))
Ions containing more than one element					
PO_4^{3-} (phosphate)	SO_4^{2-} (sulfate)	NO_3^- (nitrate)	NH_4^+ (ammonium)		
	CO_3^{2-} (carbonate)	OH^- (hydroxyl)			
		HCO_3^- (hydrogencarbonate)			

3 The equation is then **balanced**, so that the numbers of atoms on each side of the equation are equal, since matter is neither created nor destroyed in chemical reactions.

hydrogen + oxygen → water
$2H_2 + O_2 \rightarrow 2H_2O$
calcium + chlorine → sodium chloride
$Ca + 2Cl \rightarrow CaCl_2$

Acids and bases

An **acid** is a compound that produces hydrogen ions in water (Table A1.6). We are familiar with the sharp taste that acids such as lemon juice or vinegar give to the tongue. These are relatively weak acids – weak enough to use on foods. The stronger the acid, the more dangerous and corrosive it is, and the more hydrogen ions it produces. An example of a **strong acid** is hydrochloric acid. In water, this acid dissociates completely. The word 'dissociate' means 'separates into its constituent ions'.

$$HCl \rightleftharpoons H^+ + Cl^-$$

With organic acids such as citric acid (present in lemon juice) and ethanoic acid (of vinegar), which we recognize as **weak acids**, relatively few molecules are dissociated, and few hydrogen ions are present.

$$C_2H_5COOH \rightleftharpoons C_2H_5COO^- + H^+$$

A **base** is a substance that can accept a hydrogen ion and so neutralize an acid. From this reaction a salt and water only are formed. Many bases are insoluble in water. Those that are soluble in water are called **alkalis**. Examples of strong bases (that are also alkalis) are sodium hydroxide and potassium hydroxide. Strong alkalis, like strong acids, are completely dissociated in water.

$$NaOH \rightleftharpoons Na^+ + OH^-$$

Some substances, called **indicators**, change colour when mixed with acid or alkalis. Examples of indicators are litmus and universal indicator, which turn red in acid (see Figure A1.7).

Table A1.6 Acids and bases in biological systems

Strong acids		
HCl hydrochloric acid	$H^+ + Cl^-$	produced in the gastric glands of the stomach
H_2SO_4 sulfuric acid	$2H^+ + SO_4^{2-}$	produced when sulfur trioxide (atmospheric pollutant) dissolves in rain water
HNO_3 nitric acid	$H^+ + NO_3^-$	produced when oxides of nitrogen (NO_x) (atmospheric pollutant) dissolve in rain water
Weak acids		
C_2H_5COOH ethanoic acid	$H^+ + C_2H_5COO^-$	formed by bacterial action on ethanol
H_2CO_3 carbonic acid	$H^+ + HCO_3^-$	formed when carbon dioxide dissolves in water
Strong alkalis		
NaOH sodium hydroxide	$Na^+ + OH^-$	commonly used laboratory alkali, very corrosive when concentrated
KOH potassium hydroxide	$K^+ + OH^-$	commonly used laboratory alkali, very corrosive when concentrated
Other bases		
$CaCO_3$ calcium carbonate $MgCO_3$ magnesium carbonate		both these bases are found in soils and tend to neutralize acid rain

pH and buffers

pH is a measure of the acidity or alkalinity of a solution. Strictly, pH is a measure of the hydrogen ion concentration. The pH scale runs from 0 to 14. pH 7 is neutral; this is the pH of pure water, where the concentrations of hydrogen ions and hydroxyl ions are low and equal in number.

A solution with a pH of less than 7 is an acidic solution; strong acids have a pH of 0–2. Solutions with a pH of more than 7 are alkaline. Because the pH scale is logarithmic, a change on the scale of one unit (for example, from pH 4 to pH 5) represents a ten-fold change in the amount of hydrogen ions present. This means that pH values cannot be added up and an average (mean) value found.

pH is measured experimentally, using either an indicator solution or a pH meter. For example, universal pH indicator is a mixture of several different indicators, and changes colour with pH, as shown in Figure A1.7.

pH is very important in living organisms, largely because it affects the shape of enzymes, almost all of which are proteins. In the mammal's body there are mechanisms to maintain the pH at a constant level, at a value just slightly above pH 7.0. If the pH varies very much from this value this is quickly fatal. For plants that obtain essential mineral ions from the soil solution, the pH of the soil affects the availability of ions for absorption.

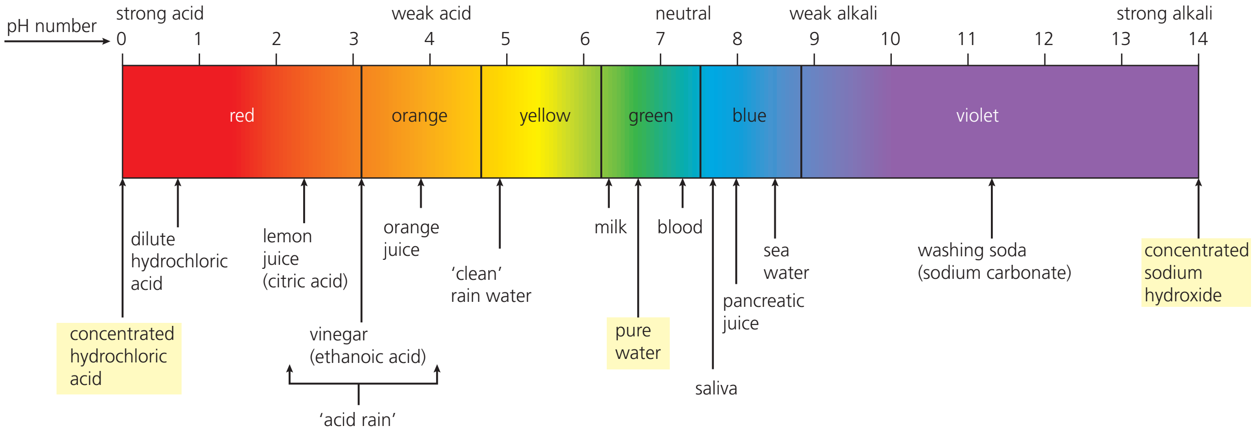


Figure A1.7 The pH scale of universal indicator solution

A **buffer solution** acts to resist pH change when diluted, or if a little acid or alkali is added. Many buffers used in laboratory experiments contain a weak acid (such as ethanoic acid) and its soluble salt (for example, sodium ethanoate). In this case, if acid is added, the excess hydrogen ions are immediately removed by being combined with ethanoate ions to form undissociated ethanoic acid. Alternatively, if alkali is added, the excess hydroxyl ions immediately combine with hydrogen ions, forming water. At the same time, more of the ethanoic acid dissociates, adding more hydrogen ions to the solution. The pH does not change in either case.

In the body of the mammal, the blood is very powerfully buffered by the presence of a mixture of phosphate ions, hydrogencarbonate ions and blood proteins. The pH of blood is held between 7.35 and 7.45.

Isotopes

Atoms of certain elements exist in more than one form, and these are known as **isotopes** of the element. All the isotopes of an element have the same number of orbiting electrons and the same number of protons in their nuclei, so they have the same **chemical properties**. Consequently, they react identically. However, the isotopes of a particular element have different numbers of neutrons in the nucleus, and so they have different masses.

Isotopes are classified as stable or unstable. **Stable isotopes** persist in nature because they do not undergo radioactive decay. Oxygen-16 and oxygen-18 are stable isotopes. **Unstable isotopes** are radioactive. This means they break down steadily, often emitting α or β particles. This process is known as radioactive decay. The product of the decay process is a stable isotope. For example, carbon-14 (radioactive) decays to nitrogen-14, emitting β particles in the process.

The radiation emitted by radioactive isotopes fogs photographic film placed near the radioactive source in the dark, a technique known as **autoradiography**. The presence of radioactive isotopes can also be detected by instruments, such as a Geiger–Müller tube (**Geiger counter**). Non-radioactive isotopes are detected by their difference in mass, using a **mass spectrometer**. Examples of stable and radioactive isotopes and the uses they may be put to in biology are given in Table A1.7.

Table A1.7 Some isotopes of importance in biology

Stable isotopes	
oxygen-16 and oxygen-18	these isotopes have been used in experiments showing that oxygen given off in photosynthesis comes from the ‘splitting’ of water
potassium-40 and argon-40	these isotopes are used in the dating of volcanic lava deposits in the process of determining the approximate age of fossils originally trapped in the cooling lava
Stable/radioactive(*) isotopes	
carbon-12/carbon-14(*)	these isotopes have been used to investigate the ‘fixation’ of carbon from carbon dioxide into sugar during photosynthesis
phosphorus-31/phosphorus-32(*) and sulfur-32/sulfur-35(*)	these isotopes have been used to show that DNA, and not protein, carries the genetic message

Oxidation and reduction

When the fuel methane (CH_4) is burned in air, carbon dioxide and water are formed:

$$2CH_4 + 4O_2 \rightarrow 2CO_2 + 4H_2O$$

Some oxygen has been added to the carbon, so the carbon has been **oxidized**. At the same time, some of the oxygen has been **reduced** to water. Oxidation and reduction always occur together in this way. When one substance in a reaction is oxidized another is automatically reduced; that is, a reduction/oxidation reaction has occurred. The shorthand name for **red**uction/**ox**idation reactions is **redox** reactions.

In redox reactions in biology, hydrogen atoms are often involved. We have seen that a hydrogen atom is an electron and a proton. Gaining one or more hydrogen atoms (a case of reduction) involves gaining one or more electrons. In fact, the best definition of oxidation is the loss of electrons and of reduction is the gain of electrons. Remembering this definition has given countless people problems. A mnemonic has been devised, namely:

OILRIG = **O**xidation **I**s **L**oss of electrons; **R**eduction **I**s **G**ain of electrons

Redox reactions mostly take place in biological systems because of the presence of a compound with a strong tendency to take electrons from another compound (an **oxidizing agent**) or the presence of a compound with a strong tendency to donate electrons to another compound (a **reducing agent**). For example, in respiration (in which glucose is oxidized to carbon dioxide and water) a common oxidizing agent is a large molecule known as NAD, and the corresponding oxidizing agent is NADH. In photosynthesis (in which carbon dioxide is reduced to glucose), a common reducing agent is a large molecule known as NADPH, and the corresponding oxidizing agent is NADP.

Another feature of oxidation and reduction is **energy change**. When reduction occurs, energy is absorbed (an **endergonic reaction**). When oxidation occurs, energy is released (an **exergonic reaction**). An example of energy release in oxidation is the burning of a fuel in air. Here, energy is given out as heat. In fact, the amount of energy in a molecule depends on its degree of oxidation. An oxidized substance has less stored energy than a reduced substance. This is illustrated by the fuel molecule methane (CH_4), which has more stored chemical energy than carbon dioxide (CO_2).

Hydrolysis and condensation

When two molecules join together to form a larger molecule with the loss of a molecule of water, a **condensation reaction** has occurred. This happens when sugar molecules are condensed together to make disaccharides and polysaccharides. Other examples are when glycerol reacts with fatty acids to form triglycerols and when an organic acid reacts with an alcohol to form an ester.

$$\mathrm{R{-}COOH} + \mathrm{HO{-}R} \xrightleftharpoons[\text{hydrolysis reaction}]{\text{condensation reaction}} \mathrm{R{-}COOR} + H_2O$$

Hydrolysis is the reverse of condensation. It occurs when water splits a molecule into two or more smaller molecules.

Molarity and concentration

The **mole** is the scientific unit for the amount of substance. We use the mole in situations where we need a unit that contains the same number of atoms (or molecules or ions), whatever the substance being referred to. All atoms are incredibly small, but atoms of different elements vary in mass. Some atoms are very heavy, others are lighter, but a mole of each will contain exactly the same number of atoms. So the mole is a particularly useful way of describing concentration, and when you use the mole, you count the atoms (or molecules or ions) by weighing them!

A mole (symbol mol) of a substance contains the relative atomic mass (A_r) of the substance expressed in grams. We have seen that the relative atomic mass of carbon-12 is taken as 12, and that hydrogen has a relative atomic mass of 1. Similarly oxygen has a relative atomic mass of 16. So, 1 mol of carbon has a mass of 12 g, 1 mol of hydrogen 1 g, and 1 mol of oxygen 16 g.

Actually, hydrogen and oxygen both exist as molecules, H_2 and O_2. Consequently 1 mol of hydrogen molecules (H_2) has a mass of 2 g and 1 mol of oxygen molecules (O_2) has a mass of 32 g. Moles all contain the same number of units of substance.

The **concentration** of a solution can be expressed in moles, too. A **molar solution** contains 1 mol in a litre. For example, the relative molecular mass of glucose ($C_6H_{12}O_6$) is 180 ($C = 12 \times 6$; $H = 1 \times 12$; $O = 16 \times 6$). So, we make a molar solution of glucose by taking the relative molecular mass in grams (180 g) and dissolving it in 1 litre of distilled water.

In the same way, the relative molecular mass of sodium chloride is 58.5 (relative atomic mass of sodium = 23, chlorine = 35.5 – see Table A1.3). So a molar solution of sodium chloride contains 58.5 g NaCl per litre. However, unlike glucose, sodium chloride ionizes ($NaCl \rightarrow Na^+ + Cl^-$). This means a molar solution of sodium chloride contains twice as many particles (ions) as does a molar solution of glucose (molecules). This brings us back to the definition – that is, a mole always contains the same number of units of substance, whether that substance exists as atoms, or molecules or ions.