CHAPTER 2

Introduction to the chemistry of life

ANIMATIONS

Because living tissues are composed of chemical building blocks, the study of anatomy and physiology depends upon some understanding of biochemistry – the chemistry of life. This chapter introduces core concepts in chemistry that will underpin the remaining chapters in this book.

Atoms, molecules and compounds

Learning outcomes

After studying this section, you should be able to:

- define the following terms: atomic number, atomic weight, isotope, molecular weight, ion, electrolyte, pH, acid and alkali
- describe the structure of an atom
- discuss the types of bond that hold molecules together
- outline the concept of molar concentration
- explain the importance of buffers in the regulation of pH.

All matter in our universe is built of particles called *atoms*. An *element* contains only one type of atom, e.g. carbon, sulphur or hydrogen. Substances containing two or more types of atom combined are called *compounds*. For instance, water is a compound containing both hydrogen and oxygen atoms.

There are 92 naturally occurring elements, but the wide variety of compounds making up living tissues are composed almost entirely of only four: carbon, hydrogen, oxygen and nitrogen. Small amounts (about 4% of body weight) of others are present, including sodium, potassium, calcium and phosphorus.

Atomic structure

Atoms are mainly empty space, with a tiny central nucleus containing *protons* and *neutrons* surrounded by clouds of tiny orbiting *electrons* (Fig. 2.1). Neutrons carry no electrical charge, but protons are positively charged, and electrons are negatively charged. Because atoms contain equal numbers of protons and electrons, they carry no net charge.

These subatomic particles differ also in terms of their mass. Electrons are so small that their mass is negligible, but the bigger neutrons and protons carry one atomic mass unit each. The physical characteristics of electrons, protons and neutrons are summarised in Table 2.1.

Atomic number and atomic weight

What makes one element different from another is the number of protons in the nuclei of its atoms (Fig. 2.2). This

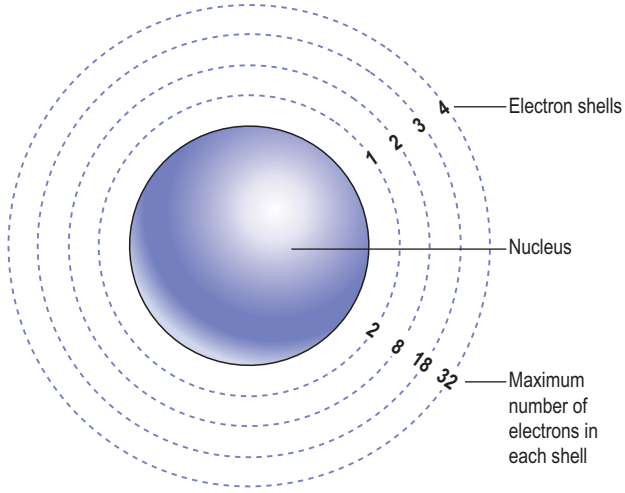

Figure 2.1 The atom, showing the nucleus and four electron shells.

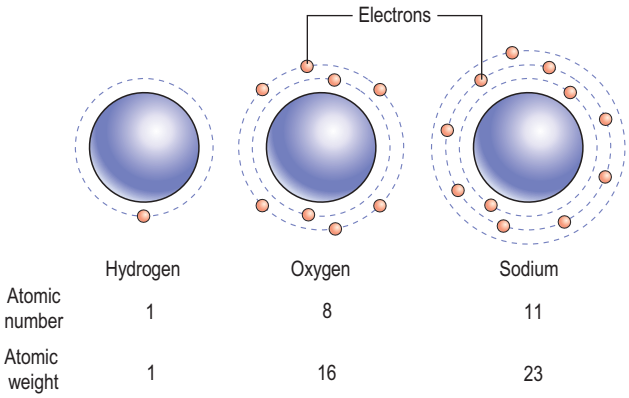

Figure 2.2 The atomic structures of the elements hydrogen, oxygen and sodium.

Table 2.1 Characteristics of subatomic particles

Particle	Mass	Electric charge
Proton	1 unit	1 positive
Neutron	1 unit	Neutral
Electron	Negligible	1 negative

is called the *atomic number* and each element has its own atomic number, unique to its atoms. For instance, hydrogen has only one proton per nucleus, oxygen has eight and sodium has 11. The atomic numbers of hydrogen, oxygen and sodium are therefore 1, 8 and 11, respectively. The *atomic weight* of an element is the sum of the protons and neutrons in the atomic nucleus.

The electrons are shown in Figure 2.1 as though they orbit in concentric rings round the nucleus. These shells

represent the different energy levels of the atom's electrons, not their physical positions. The first energy level can hold only two electrons and is filled first. The second energy level can hold only eight electrons and is filled next. The third and subsequent energy levels hold increasing numbers of electrons, each containing more than the preceding level.

When the atom's outer electron shell does not contain a stable number of electrons, the atom is *reactive* and can donate, receive or share electrons with one or more other atoms to achieve stability. The great number of possible combinations of different types of atom yields the wide range of substances of which the world is built and on which biology is based. This is described more fully in the section discussing molecules and compounds.

Isotopes. These are atoms of an element in which there is a *different number of neutrons in the nucleus*. This does not affect the electrical activity of these atoms because neutrons carry no electrical charge, but it does affect their atomic weight. For example, there are three forms of the hydrogen atom. The most common form has one proton in the nucleus and one orbiting electron. Another form (*deuterium*) has one proton and one neutron in the nucleus. A third form (*tritium*) has one proton and two neutrons in the nucleus and one orbiting electron. Each is an *isotope* of hydrogen (Fig. 2.3).

Because the atomic weight of an element is actually an average atomic weight calculated using all its atoms, the true atomic weight of hydrogen is 1.008, although for most practical purposes it can be taken as 1.

Chlorine has an atomic weight of 35.5, because it contains two isotopes, one with an atomic weight of 35 (with 18 neutrons in the nucleus) and the other 37 (with 20 neutrons in the nucleus). Because the proportion of these two forms is not equal, the average atomic weight is 35.5.

Molecules and compounds

As mentioned earlier, the atoms of each element have a specific number of electrons around the nucleus. When the number of electrons in the outer shell of an element is either the maximum number (Fig. 2.1), or a stable proportion of this fraction, the element is described as *inert* or chemically unreactive, and it will not easily combine with other atoms. These elements are the inert gases – helium, neon, argon, krypton, xenon and radon.

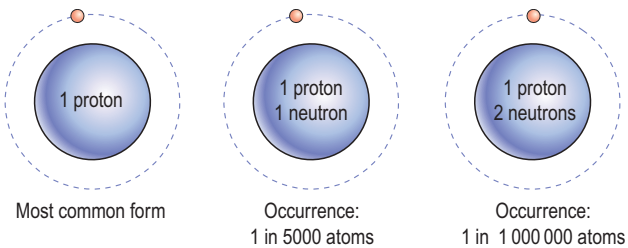

Figure 2.3 The isotopes of hydrogen.

Molecules consist of two or more atoms that are chemically combined. The atoms may be of the same element, e.g. a molecule of atmospheric oxygen (O_2) contains two oxygen atoms. Most substances, however, are compounds and contain two or more different elements, e.g. a water molecule (H_2O) contains two hydrogen atoms and an oxygen atom. **2.1**

Compounds containing carbon and hydrogen are classified as *organic*, and all others as *inorganic*. Living tissues are based on organic compounds, but the body requires inorganic compounds too.

Covalent and ionic bonds. The vast array of chemical processes on which life is based is completely dependent upon the way atoms come together, bind and break apart. For example, the humble water molecule is a crucial foundation of all life on Earth. If water was a less stable compound, and the atoms came apart easily, human biology could never have evolved. On the other hand, the body is dependent upon the breaking down of various molecules (e.g. sugars, fats) to release energy for cellular activities. When atoms are joined together, they form a chemical bond that is generally one of two types: *covalent* or *ionic*.

Covalent bonds are formed when atoms share their electrons with each other. Most molecules are held together with this type of bond; it forms a strong and stable link between its constituent atoms. A water molecule is built using covalent bonds. Hydrogen has one electron in its outer shell, but the optimum number for this shell is two. Oxygen has six electrons in its outer shell, but the optimum number for this shell is eight. Therefore, if one oxygen atom and two hydrogen atoms combine, each hydrogen atom will share its electron with the oxygen atom, giving the oxygen atom a total of eight outer electrons, making it stable. The oxygen atom shares one of its electrons with each of the two hydrogen atoms, so that each hydrogen atom has two electrons in its outer shell, and they too are stable (Fig. 2.4).

Ionic bonds are weaker than covalent bonds and are formed when electrons are transferred from one atom to another. For example, when sodium (Na) combines with chlorine (Cl) to form sodium chloride (NaCl), the only

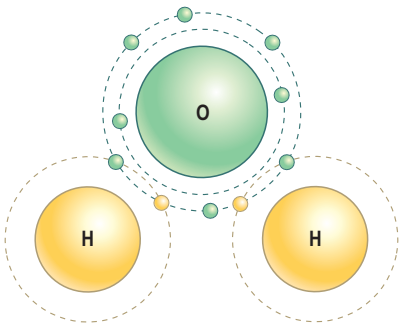

Figure 2.4 A water molecule, showing the covalent bonds between hydrogen (yellow) and oxygen (green).

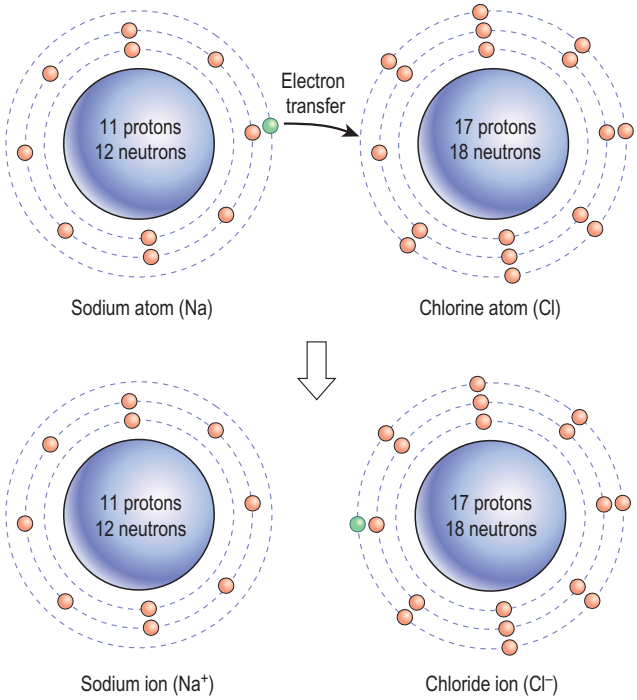

Figure 2.5 Formation of the ionic compound, sodium chloride.

electron in the outer shell of the sodium atom is transferred to the outer shell of the chlorine atom (Fig. 2.5).

This leaves the sodium atom with eight electrons in its outer (second) shell, and therefore stable. The chlorine atom also has eight electrons in its outer shell, which, although not filling the shell, is a stable number. The sodium atom is now positively charged because it has given away a negatively charged electron, and the chloride ion is now negatively charged because it has accepted sodium's extra electron. The two atoms, therefore, stick together because they are carrying opposite, mutually attractive, charges.

When sodium chloride is dissolved in water the ionic bond breaks and the two atoms separate. The atoms are charged, because they have traded electrons, so are no longer called atoms, but *ions*. Sodium, with the positive charge, is a *cation*, written Na^+, and chloride, being negatively charged, is an *anion*, written Cl^-. By convention the number of electrical charges carried by an ion is indicated by the superscript plus or minus signs. **2.2**

Electrolytes

An ionic compound, e.g. sodium chloride, dissolved in water is called an *electrolyte* because it conducts electricity. Electrolytes are important body constituents because they:

- conduct electricity, essential for muscle and nerve function
- exert osmotic pressure, keeping body fluids in their own compartments
- act as buffers (p. 24) to resist pH changes in body fluids.

Many biological compounds, e.g. carbohydrates, are not ionic, and therefore have no electrical properties when dissolved in water. Important electrolytes other than sodium and chloride include potassium (K^+), calcium (Ca^{2+}), bicarbonate (HCO_3^-) and phosphate (PO_4^{3-}).

Table 2.2 Examples of normal plasma levels

Substance	Molar concentrations	Equivalent concentration in other units
Chloride	97–106 mmol/L	97–106 mEq/L
Sodium	135–143 mmol/L	135–143 mEqL
Glucose	3.5–5.5 mmol/L	60–100 mg/100 mL
Iron	14–35 mmol/L	90–196 mg/100 mL

Measurement of substances in body fluids

There is no single way of measuring and expressing the concentration of different substances in body fluids. Sometimes the unit used is based on weight in grams or fractions of a gram (see also pp. 479–80), e.g. milligrams, micrograms or nanograms. If the molecular weight of the substance is known, the concentration can be expressed as moles, millimoles or nanomoles per litre. A related measure is the *milliequivalent (mEq)* per litre.

Sometimes it is most convenient to measure the quantity of a substance in terms of its activity; insulin, for instance, is measured in *international units* (IU).

Table 2.2 gives examples of the normal plasma levels of some important substances, given in molar concentrations and alternative units.

Acids, bases and pH

pH is the measuring system used to express the concentration of hydrogen ions ($[H^+]$) in a fluid, which is an indicator of its acidity or alkalinity. Living cells are very sensitive to changes in $[H^+]$, and since the biochemical processes of life continually produce or consume hydrogen ions, sophisticated homeostatic mechanisms in the body constantly monitor and regulate pH.

An acid substance releases hydrogen ions when in solution. On the other hand, a basic (alkaline) substance accepts hydrogen ions, often with the release of hydroxyl (OH^-) ions. A salt releases other anions and cations when dissolved; sodium chloride is therefore a salt because in solution it releases sodium and chloride ions.

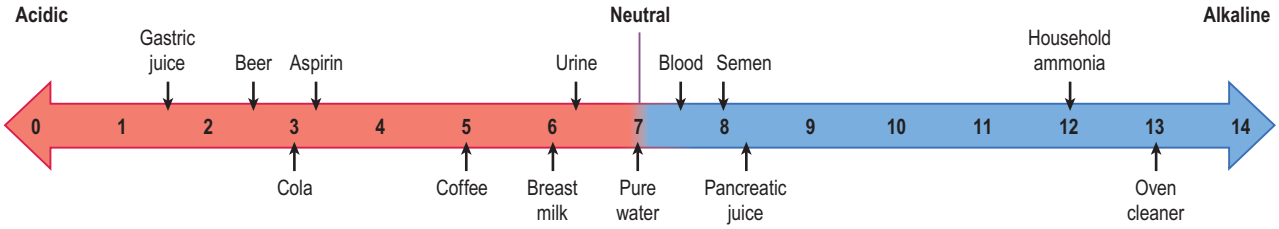

Figure 2.6 The pH scale.

The pH scale

The standard scale for measurement of hydrogen ion concentration in solution is the pH scale. The scale measures from 0 to 14, with 7, the midpoint, as neutral; this is the pH of pure water. Water is a neutral molecule, neither acid nor basic (alkaline), because when the molecule breaks up into its constituent ions, it releases one H^+ and one OH^-, which balance one another. With the notable exception of gastric juice, most body fluids are close to neutral, because they contain *buffers*, themselves weak acids and bases, to keep their pH within narrow ranges.

A pH reading below 7 indicates an *acid solution*, while readings above 7 indicate basic (alkaline) solutions. Figure 2.6 shows the pH of some common fluids (see also, p. 479). A change of one whole number on the pH scale indicates a 10-fold change in $[H^+]$. Therefore, a solution of pH 5 contains 10 times as many hydrogen ions as a solution of pH 6.

Not all acids ionise completely when dissolved in water. The hydrogen ion concentration is, therefore, a measure of the amount of *dissociated acid* (ionised acid) rather than of the total amount of acid present. Strong acids dissociate more extensively than weak acids, e.g. hydrochloric acid dissociates extensively into H^+ and Cl^-, while carbonic acid dissociates much less freely into H^+ and HCO_3^-.

Likewise, not all bases dissociate completely. Strong bases dissociate more fully, i.e. they release more OH^- than weaker ones.

pH values of body fluids

The pH of body fluids are generally maintained within relatively narrow limits.

The highly acid pH of the gastric juice is maintained by hydrochloric acid secreted by the parietal cells in the walls of the gastric glands. The low pH of the stomach fluids destroys microbes and toxins swallowed in food or drink. Saliva has a pH of between 5.4 and 7.5, which is the optimum value for the action of salivary amylase, the enzyme present in saliva which initiates the digestion of carbohydrates. Amylase is destroyed by gastric acid when it reaches the stomach.

Blood pH is kept between 7.35 and 7.45, and outwith this narrow range there is severe disruption of normal physiological and biochemical processes. Normal metabolic activity of body cells constantly produces acids and bases, which would tend to alter the pH of the tissue fluid and blood. Chemical *buffers*, which can reversibly bind hydrogen ions, are responsible for keeping body pH stable.

Buffers

Despite the constant cellular production of acids and bases, body pH is kept stable by systems of buffering chemicals in body fluids and tissues. These buffering mechanisms temporarily neutralise fluctuations in pH, but can function effectively only if there is some means by which excess acid or bases can be excreted from the body. The organs most active in this way are the *lungs* and the *kidneys*. The lungs are important regulators of blood pH because they excrete carbon dioxide (CO_2). CO_2 increases $[H^+]$ in body fluids because it combines with water to form carbonic acid, which then dissociates into a bicarbonate ion and a hydrogen ion.

$$\underset{\text{carbon dioxide}}{CO_2} + \underset{\text{water}}{H_2O} \leftrightarrow \underset{\text{carbonic acid}}{H_2CO_3} \leftrightarrow \underset{\text{hydrogen ion}}{H^+} + \underset{\text{bicarbonate ion}}{HCO_3^-}$$

The lungs, therefore, help to control blood pH by regulating levels of excreted CO_2. The brain detects rising $[H^+]$ in the blood and stimulates breathing, causing increased CO_2 loss and a fall in $[H^+]$. Conversely, if blood pH becomes too basic, the brain can reduce the respiration rate to increase CO_2 levels and increase $[H^+]$, decreasing pH towards normal (see Ch. 10).

The kidneys regulate blood pH by adjusting the excretion of hydrogen and bicarbonate ions as required. If pH falls, hydrogen ion excretion is increased and bicarbonate conserved; the reverse happens if pH rises. In addition, the kidneys generate bicarbonate ions as a by-product of amino acid breakdown in the renal tubules; this process also generates ammonium ions, which are rapidly excreted.

Other buffer systems include body proteins, which absorb excess H^+, and phosphate, which is particularly important in controlling intracellular pH. The buffer and excretory systems of the body together maintain the *acid–base balance* so that the pH range of body fluids remains within normal, but narrow, limits.

Acidosis and alkalosis

The buffer systems described above compensate for most pH fluctuations, but these reserves are limited and, in extreme cases, can become exhausted. When the pH falls below 7.35, and all the reserves of alkaline buffers are used up, the condition of *acidosis* exists. In the reverse situation, when the pH rises above 7.45, the increased alkali uses up all the acid reserve and the state of *alkalosis* exists.

Acidosis and alkalosis are both dangerous, particularly to the central nervous system and the cardiovascular system. In practice, acidotic conditions are commoner than alkalotic ones, because the body tends to produce more acid than alkali. Acidosis may follow respiratory problems, if the lungs are not excreting CO_2 as efficiently as normal, or if the body is producing excess acids (e.g. diabetic ketoacidosis, p. 237) or in kidney disease, if renal H^+ excretion is reduced. Alkalosis may be caused by loss of acidic substances through vomiting, diarrhoea, endocrine disorders or diuretic therapy, which stimulates increased renal excretion. Rarely, it may follow increased respiratory effort, such as in an acute anxiety attack where excessive amounts of CO_2 are lost through overbreathing (hyperventilation).

Important biological molecules

Learning outcomes

After studying this section, you should be able to:

- describe in simple terms the chemical nature of sugars, proteins, lipids, nucleotides and enzymes
- discuss the biological importance of each of these important groups of molecules.

Carbohydrates

Carbohydrates (sugars and starches) are composed of carbon, oxygen and hydrogen. The carbon atoms are normally arranged in a ring, with the oxygen and hydrogen atoms linked to them. The structures of glucose, fructose and sucrose are shown in Figure 2.7. When two sugar molecules combine to form a bigger sugar molecule, a water molecule is expelled and the bond formed is called a *glycosidic linkage*.

Glucose, the cells' preferred fuel molecule, is a *monosaccharide* (mono = one; saccharide = sugar). Monosaccharides can be linked together to form bigger sugars, ranging in size from two sugar units (*disaccharides*), e.g. sucrose (table sugar) (Fig. 2.7), to long chains containing many thousands of monosaccharides, such as starch. Such complex carbohydrates are called *polysaccharides*.

Glucose can be broken down in either the presence (*aerobically*) or the absence (*anaerobically*) of oxygen, but the process is much more efficient when O_2 is used. During this process, energy, water and carbon dioxide are released (pp. 315–6). To ensure a constant supply of glucose for cellular metabolism, blood glucose levels are tightly controlled. Functions of sugars include:

- providing a ready source of energy to fuel cell metabolism (p. 313)
- providing a form of energy storage, e.g. glycogen (p. 310)
- forming an integral part of the structure of DNA and RNA (pp. 438, 441)
- acting as receptors on the cell surface, allowing the cell to recognise other molecules and cells.

Amino acids and proteins

Amino acids always contain carbon, hydrogen, oxygen and nitrogen, and many in addition carry sulphur. In human biochemistry, 20 amino acids are used as the principal building blocks of protein, although there are others; for instance, there are some amino acids used only in certain proteins, and some are seen only in microbial products. The amino acids used in human protein synthesis have a basic common structure, including an amino group (NH_2), a carboxyl group (COOH) and a hydrogen atom. What makes one amino acid different from the next is a variable side chain. The basic structure and three common amino acids are shown in Figure 2.8. As in the formation of glycosidic linkages, when two amino acids join up the reaction expels a molecule of water and the resulting bond is called a *peptide bond*.

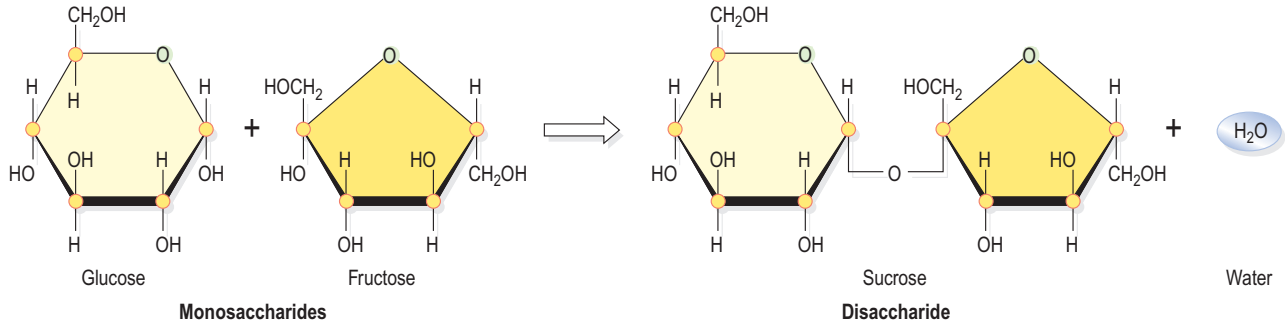

Figure 2.7 The combination of glucose and fructose to make sucrose.

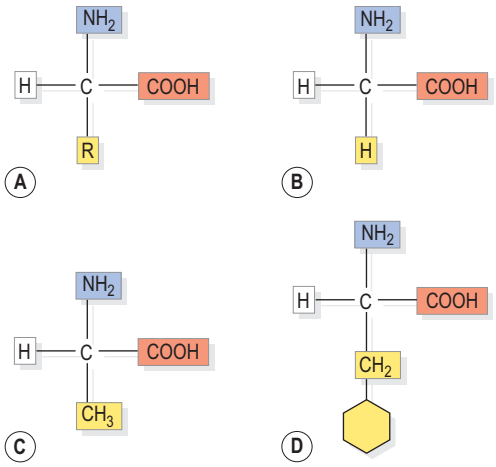

Figure 2.8 Amino acid structures. A. Common structure, R = variable side chain. **B.** Glycine, the simplest amino acid. **C.** Alanine. **D.** Phenylalanine.

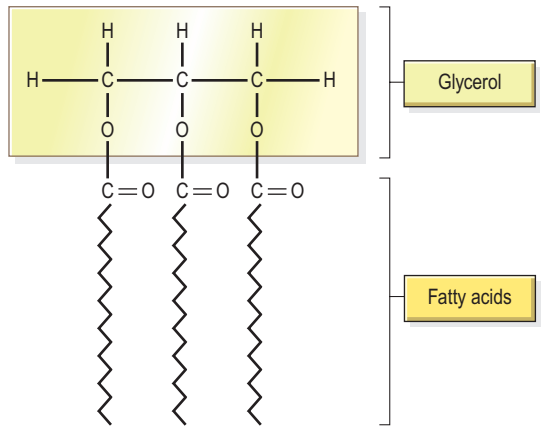

Figure 2.9 Structure of a fat (triglyceride) molecule.

Proteins are made from amino acids joined together, and are the main family of molecules from which the human body is built. Protein chains can vary in size from a few amino acids long to many thousands. They may exist as simple, single strands of protein, for instance some hormones, but more commonly are twisted and folded into complex and intricate three-dimensional structures that may contain more than one kind of protein, or incorporate other types of molecule, e.g. haemoglobin (Fig. 4.6). Such complex structures are stabilised by internal bonds between constituent amino acids, and the function of the protein will depend upon the three-dimensional shape it has been twisted into. One reason why changes in pH are so damaging to living tissues is that hydrogen ions disrupt these internal stabilising forces and change the shape of the protein (denaturing it), leaving it unable to function. Many important groups of biologically active substances are proteins, e.g.:

- carrier molecules, e.g. haemoglobin (p. 65)
- enzymes (p. 28)
- many hormones, e.g. insulin (p. 227)
- antibodies (pp. 381–2).

Proteins can also be used as an alternative energy source, usually in starvation. The main source of body protein is muscle tissue, so muscle wasting is a feature of starvation.

Lipids

The lipids are a diverse group of substances whose common property is an inability to mix with water (i.e. they are *hydrophobic*). They are made up mainly of carbon, hydrogen and oxygen atoms, and some contain additional elements, like nitrogen or phosphorus. The most important groups of lipids include:

- *phospholipids*, integral to cell membrane structure. They form a double layer, providing a water-repellant barrier separating the cell contents from its environment (p. 32)
- certain vitamins (p. 278). The fat-soluble vitamins are A, D, E and K
- *fats (triglycerides)*, stored in adipose tissue (p. 41) as an energy source. Fat also insulates the body and protects internal organs. A molecule of fat contains three fatty acids attached to a molecule of glycerol (Fig. 2.9). When fat is broken down under optimal conditions, more energy is released than when glucose is fully broken down.

Fats are classified as *saturated* or *unsaturated*, depending on the chemical nature of the fatty acids present. Saturated fat tends to be solid, whereas unsaturated fats are fluid.

- *prostaglandins* are important chemicals derived from fatty acids and are involved in inflammation (p. 377) and other processes.
- *steroids*, including important hormones produced by the gonads (the ovaries and testes, p. 455 and p. 459) and adrenal glands (p. 244). *Cholesterol* is a steroid that stabilises cell membranes and is the precursor of the hormones mentioned above, as well as being used to make bile salts for digestion.

Nucleotides

Nucleic acids

These are the largest molecules in the body and are built from nucleotides. They include deoxyribonucleic acid (DNA, p. 438) and ribonucleic acid (RNA, p. 441).

Adenosine triphosphate (ATP)

ATP is a nucleotide built from ribose (the sugar unit), adenine (the base) and three phosphate groups attached to the ribose (Fig. 2.10A). It is sometimes called the energy

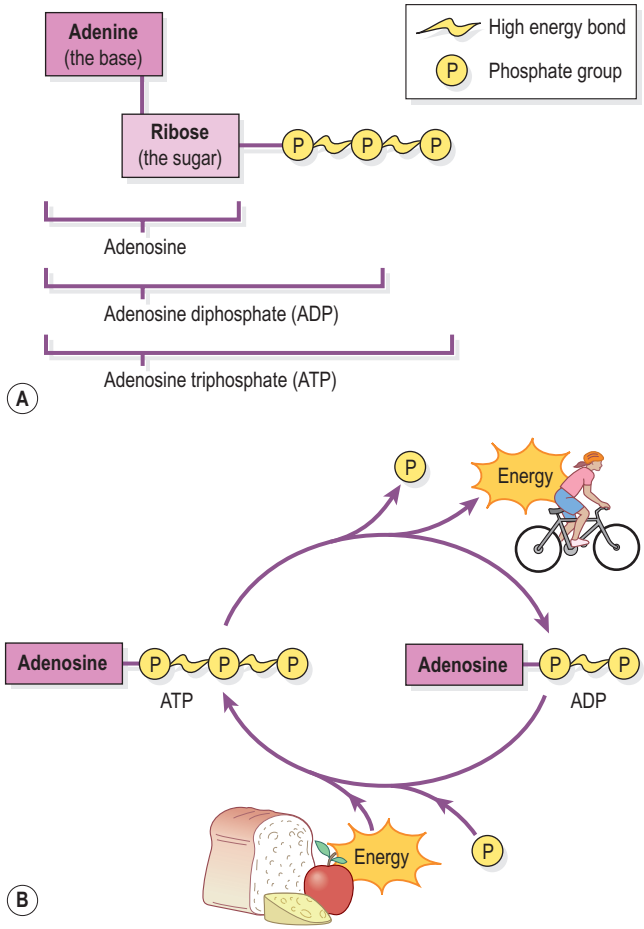

Figure 2.10 ATP and ADP. A. Structures. **B.** Conversion cycle.

currency of the body, which implies that the body has to 'earn' (synthesise) it before it can 'spend' it. Many of the body's huge number of reactions release energy, e.g. the breakdown of sugars in the presence of O_2. The body captures the energy released by these reactions, using it to make ATP from adenosine diphosphate (ADP). When the cells need chemical energy to fuel metabolic activities, ATP is broken down again into ADP, releasing water, a phosphate group, and energy from the splitting of the high-energy phosphate bond (Fig. 2.10B). Energy generated from ATP breakdown fuels muscle contraction, motility of spermatozoa, anabolic reactions and the transport of materials across membranes.

Enzymes

Many of the body's chemical reactions can be reproduced in a test-tube. Surprisingly, the rate at which the reactions then occur usually plummets to the extent that, for all practical purposes, chemical activity ceases. The cells of the body have developed a solution to this apparent problem – they are equipped with a huge array of enzymes. Enzymes are proteins that act as *catalysts* for

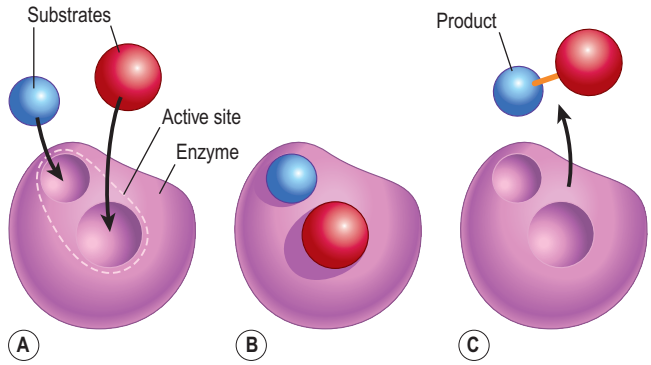

Figure 2.11 Action of an enzyme. A. Enzyme and substrates. **B.** Enzyme–substrate complex. **C.** Enzyme and product.

biochemical reactions – that is, they speed the reaction up but are not themselves changed by it, and therefore can be used over and over again. Enzymes are very selective and will usually catalyse only one specific reaction. The molecule(s) entering the reaction is called the *substrate* and it binds to a very specific site on the enzyme, called the *active site*. Whilst the substrate(s) is bound to the active site the reaction proceeds, and once it is complete the product(s) of the reaction breaks away from the enzyme and the active site is ready for use again (Fig. 2.11).

Enzyme action is reduced or stopped altogether if conditions are unsuitable. Increased or decreased temperature is likely to reduce activity, as is any change in pH. Some enzymes require the presence of a *cofactor*, an ion or small molecule that allows the enzyme to bind its substrate(s). Some vitamins act as cofactors.

Enzymes can catalyse both synthetic and breakdown reactions, and their names (almost always!) end in ~ase. When an enzyme catalyses the combination of two or more substrates into a larger product, this is called an *anabolic reaction*. *Catabolic reactions* involve the breakdown of the substrate into smaller products, as occurs during the digestion of foods.

Movement of substances within body fluids

Learning outcomes

After studying this section, you should be able to:

- compare and contrast the processes of osmosis and diffusion
- using these concepts, describe how molecules move within and between body compartments.

Movement of substances within and between body fluids, sometimes across a barrier such as the cell membrane, is essential in normal physiology.

In liquids or gases, molecules distribute from an area of high concentration to one of low concentration, assuming that there is no barrier in the way. Between two such areas, there exists a *concentration gradient* and movement of substances occurs *down* the concentration gradient, or downhill, until the molecules are evenly spread throughout, i.e. equilibrium is reached. No energy is required for such movement, so this process is described as *passive*.

There are many examples in the body of substances moving *uphill*, i.e. against the concentration gradient; in this case, energy is required, usually from the breakdown of ATP. These processes are described as *active*. Movement of substances across cell membranes by active transport is described on page 37.

Passive movement of substances in the body proceeds usually in one of two main ways – *diffusion* or *osmosis*. **2.3**

Diffusion

Diffusion refers to the movement of molecules from an area of high concentration to an area of low concentration, and occurs mainly in gases, liquids and solutions. Sugar molecules heaped at the bottom of a cup of coffee that has not been stirred will, in time, become evenly distributed throughout the liquid by diffusion (Fig. 2.12). The process of diffusion is speeded up if the temperature rises and/or the concentration of the diffusing substance is increased.

Diffusion can also occur across a semipermeable membrane, such as the plasma membrane or the capillary wall. Only molecules small or soluble enough to cross the membrane can diffuse through. For example, oxygen diffuses freely through the walls of the alveoli (airsacs in the lungs), where oxygen concentrations are high, into the bloodstream, where oxygen concentrations are low. However, blood cells and large protein molecules in the plasma are too large to cross and so remain in the blood.

Osmosis

While diffusion of molecules across a semipermeable membrane results in equal concentrations on both sides of the membrane, *osmosis* refers specifically to diffusion of water down its concentration gradient. This is usually because any other molecules present are too large to pass through the pores in the membrane. The force with which this occurs is called the *osmotic pressure*. Imagine two solutions of sugar separated by a semipermeable membrane whose pores are too small to let the sugar molecules through. On one side, the sugar solution is twice as concentrated as on the other. After a period of time, the concentration of sugar molecules will have equalised on both sides of the membrane, not because sugar molecules have diffused across the membrane, but because osmotic pressure across the membrane 'pulls' water from the dilute solution into the concentrated solution, i.e. water has moved down its concentration gradient. Osmosis proceeds until equilibrium is reached, at which point the solutions on each side of the membrane are of the same concentration and are said to be *isotonic*. The importance of careful control of solute concentrations in the body fluids can be illustrated by looking at what happens to a cell (e.g. a red blood cell) when it is exposed to solutions that differ from normal physiological conditions.

Plasma osmolarity is maintained within a very narrow range because if the plasma water concentration rises, i.e. the plasma becomes more dilute than the intracellular fluid within the red blood cells, then water will move down its concentration gradient across their membranes and into the red blood cells. This may cause the red blood cells to swell and burst. In this situation, the plasma is said to be *hypotonic*. Conversely, if the plasma water concentration falls so that the plasma becomes more concentrated than the intracellular fluid within the red blood cells (the plasma becomes *hypertonic*), water passively moves by osmosis from the blood cells into the plasma and the blood cells shrink (Fig. 2.13).

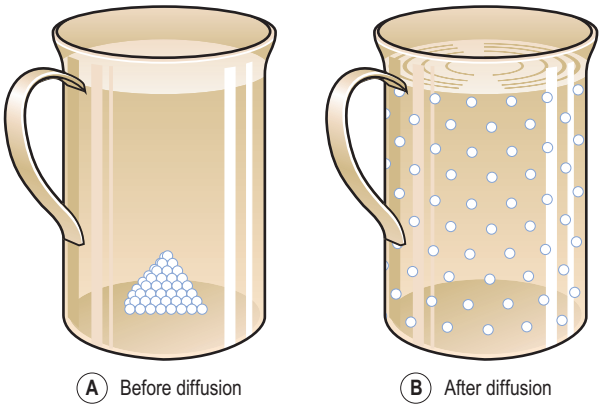

Figure 2.12 The process of diffusion: a spoonful of sugar in a cup of coffee.

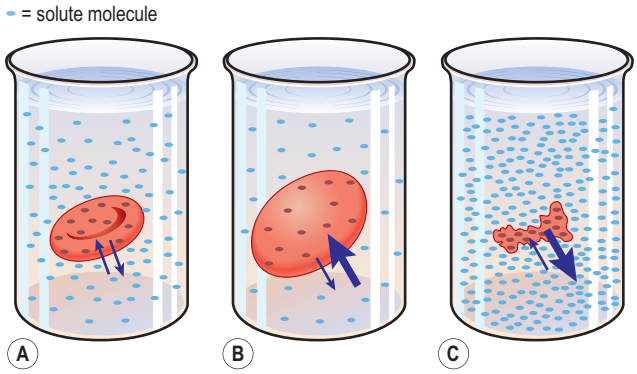

Figure 2.13 The process of osmosis. Net water movement when a red blood cell is suspended in solutions of varying concentrations (tonicity): **A.** Isotonic solution. **B.** Hypotonic solution. **C.** Hypertonic solution.

Body fluids

Learning outcomes

After studying this section, you should be able to:

- define the terms intra- and extracellular fluid
- using examples, explain why homeostatic control of the composition of these fluids is vital to body function.

The total body water in adults of average build is about 40L, around 60% of body weight. This proportion is higher in babies and young people and in adults below average weight. It is lower in the elderly and in obesity in all age groups. About 22% of body weight is extracellular water and about 38% is intracellular water. It is also lower in females than males, because females have proportionately more adipose than muscle tissue than males, and adipose tissue is only 10% water compared to 75% of muscle tissue.

Most of our total body water is found inside cells (about 70%, or 28L of the average 40L). The remaining 30% (12L) is extracellular, mostly in the interstitial fluid bathing the tissues, with nearly all the remainder found in plasma (Fig. 2.14).

Extracellular fluid

The extracellular fluid (ECF) consists mainly of blood, plasma, lymph, cerebrospinal fluid and fluid in the interstitial spaces of the body. Other extracellular fluids are present in very small amounts; their role is mainly in lubrication, and they include joint (synovial) fluid, pericardial fluid (around the heart) and pleural fluid (around the lungs).

Interstitial or intercellular fluid (tissue fluid) bathes all the cells of the body except the outer layers of skin. It is the medium through which substances diffuse from blood to body cells, and from cells to blood. Every body cell in contact with ECF is directly dependent upon the composition of that fluid for its well-being. Even slight changes can cause permanent damage, therefore, ECF composition is closely regulated. For example, a fall in plasma potassium levels may cause muscle weakness and cardiac arrhythmia, because of increased excitability of muscle and nervous tissue. Rising blood potassium also interferes with cardiac function, and can even cause the heart to stop beating. Potassium levels in the blood are only one of the many parameters under constant, careful adjustment by the homeostatic mechanisms of the body.

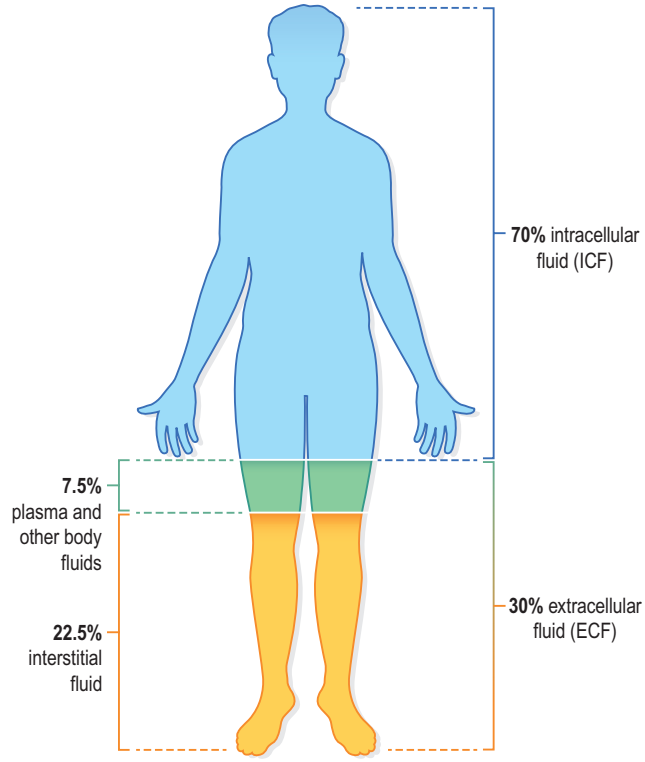

Figure 2.14 Distribution of body water in a 70 kg person.

Intracellular fluid

The composition of intracellular fluid (ICF) is largely controlled by the cell itself, because there are selective uptake and discharge mechanisms present in the cell membrane. In some respects, the composition of ICF is very different from ECF. For example, sodium levels are nearly 10 times higher in the ECF than in the ICF. This concentration difference occurs because, although sodium diffuses into the cell down its concentration gradient, there is a pump in the membrane that selectively pumps it back out again. This concentration gradient is essential for the function of excitable cells (mainly nerve and muscle). Conversely, many substances are found inside the cell in significantly higher amounts than outside, e.g. ATP, protein and potassium. Water, however, passes freely in both directions across the cell membrane, and changes in water concentration of the ECF therefore have immediate consequences for intracellular water levels (Fig. 2.13).

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