CHEMISTRY

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Chemistry

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Steven Bottle

Steven Bottle is a graduate of the University of Queensland where he completed Honours in Organic Chemistry. After working in various jobs in the pharmaceutical and mining industries, he subsequently undertook a PhD at Griffith University in collaboration with the CSIRO. On completion of his PhD he was awarded an Alexander von Humboldt Fellowship before taking up an academic position at QUT, where he has risen to the rank of full professor and where he currently leads the Molecular Design and Synthesis discipline within the school of Chemistry, Physics and Mechanical Engineering. Steven is a teaching and research academic with an established reputation for excellence in both pure and applied research, matched with demonstrated teaching capabilities and professional expertise. He has a reputation for inventive and pioneering research and has achieved international recognition for his expertise in the chemistry and applications of free radicals, especially in the context of antioxidant drugs and novel materials. Steven's particular interests include the use of stable nitroxide free radicals in synthesis, polymers and other materials as analytical tools and antioxidant drugs. Stable nitroxide free radicals play critical roles as additives (protecting coatings and plastics), as tools to make new materials and even as new antioxidant medicines. Steven has led much of the modern research on discovering new forms of nitroxide free radicals and applying them in a range of contexts, including as medicinally active compounds, detectors of free radical damage in materials and monitors of particulate pollution that impacts on human health.

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Siegbert Schmid obtained his PhD and completed a Habilitation at the University of Tübingen, Germany, and subsequently accepted a position at the School of Chemistry of the University of Sydney. His research interests lie in the synthesis and structural characterisation of aperiodic and other materials with potential technological applications (e.g. electrode materials for rechargeable batteries). In addition, he is active in chemistry education research and has supervised several PhD and Honours students in this area. Siggi's education research aims to improve current teaching practices and learning outcomes for tertiary-level students. He is a Past Chair of the RACI Division of Chemical Education. His contributions to Chemistry Education have been recognised with many awards, including the Vice Chancellor's Award for Outstanding Teaching (The University of Sydney 2012), an Office of Learning and Teaching Citation Award (2012) for Excellence in Teaching, and the Divisional Medal of the Royal Australian Chemical Institute's Division of Chemical Education (2016).

Mauro Mocerino

Professor Mauro Mocerino has enjoyed teaching chemistry at Curtin University for over two decades. During this time he has sought to better understand how students learn chemistry and what can be done to improve their learning. This has developed into a significant component of his research efforts. He also has a strong interest in enhancing the learning in laboratory classes and led the development of a professional development program for those who teach in laboratories. Mauro's other research interests are in the design and synthesis of molecules for specific intermolecular interactions including drug–protein interactions, host–guest interactions, crystal growth modification and corrosion inhibition. He has received numerous awards for his contributions to learning and teaching, including the inaugural Premier's Prize for Excellence in Science Teaching: Tertiary (2003), the Royal Australian Chemical Institute, Division of Chemical Education Medal (2012) and an Office of Learning and Teaching Australian Award for Programs that Enhance Student Learning (2013).

Uta Wille

Uta Wille is a member of the School of Chemistry at the University of Melbourne. She studied chemistry at the University of Kiel in Germany, where she graduated with a PhD in physical chemistry and afterwards completed a Habilitation in organic chemistry. Uta moved to Australia in 2003 to take up an academic position in the School of Chemistry at Melbourne University. Her research interests lie in the area of physical organic chemistry, environmental free radicals and reaction mechanisms, and she is particularly interested in how environmental radical and non-radical oxidants damage biological and manufactured materials exposed to the atmosphere. Uta teaches chemistry at both undergraduate and graduate levels and enjoys sharing her fascination and passion for chemistry with university students. She is currently Assistant Dean Undergraduate Programs in the Faculty of Science, Melbourne University.

Lightboard contributors

Throughout the VitalSource digital text there are numerous worked solutions by leading chemistry educators. These are presented as lightboard videos and help bring to the fore some of the topics that students can struggle with the most. We thank the following contributors for volunteering time out of their busy teaching and researching schedules to spend days in the studio, bringing these concepts to life.

- Uta Wille
- Christopher Thompson
- Gwen Lawrie
- Sonia Horvat

CHAPTER 1

The atom

LEARNING OBJECTIVES

After studying this chapter, you should be able to:

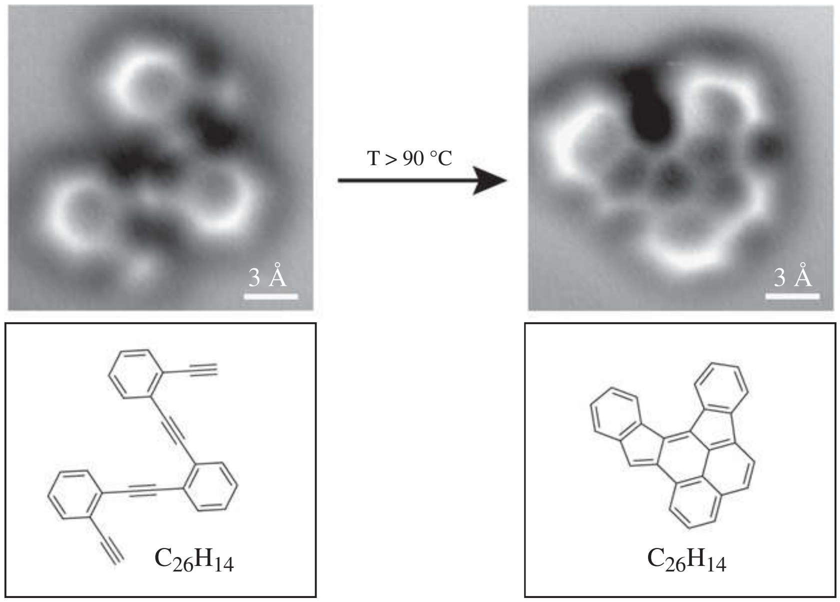
1.1 define atoms, molecules, ions, elements and compounds
1.2 explain how the concept of atoms developed
1.3 describe the structure of the atom
1.4 explain the basis of the periodic table of the elements
1.5 detail the role of electrons in atoms.

What is the universe made of? This question has occupied human thinking for thousands of years. Some ancient civilisations thought that the universe comprised only four elements (earth, air, fire and water) and that everything was made up of a combination of these. Over the past 400 years, the advent of the science called chemistry has allowed us to show that this is not the case. We now know that matter — everything you can see, smell, touch or taste — is made up of atoms, the fundamental building blocks of the universe.

Atoms are incredibly small — far too small to be seen using conventional microscopes. While many experiments over many years have produced results consistent with the existence of atoms, only recently have we been able to 'see' individual atoms and the collections of atoms we call molecules. We now have the technology to observe individual molecules undergoing a chemical reaction, a process in which one chemical substance is converted into another. The images in figure 1.1 show a molecule called an alkyne, which contains three rings made up of carbon atoms, reacting when heated to over 90 °C to give a molecule containing seven rings. The images were obtained using an atomic force microscope (AFM), and the scale on the images (3 Å = 0.000 000 000 3 m) gives an idea of just how tiny atoms truly are.

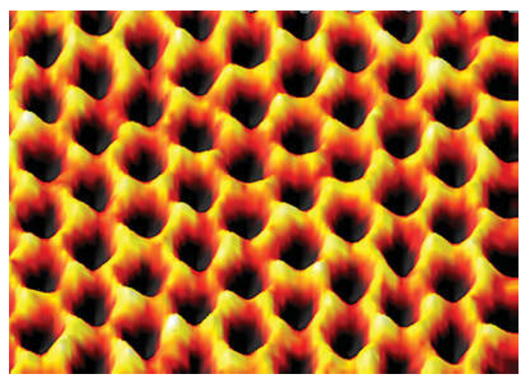
The AFM image in figure 1.2 is of a substance called graphene, which consists of a single layer of hexagonally arranged carbon atoms. Each hexagon has a diameter of approximately 3 Å. The discovery of graphene involved peeling off individual layers of graphite (the 'lead' in a pencil) using sticky tape. This very simple experiment earned graphene's discoverers a Nobel Prize.

FIGURE 1.1 AFM images of a three-ring alkyne reacting when heated to form a seven-ring molecule



Our current knowledge of the structure of the atom, and the way in which atoms pack together in three-dimensional space, owes much to experiments carried out by two Australasian-born scientists, Ernest Rutherford (1871–1937; Nobel Prize in chemistry, 1908) and William Lawrence Bragg (1890–1971; Nobel Prize in physics, 1915), both of whom would doubtless have been astonished by these AFM images. The New Zealand-born Rutherford was the first to show that the atom consists of a positively charged nucleus surrounded by tiny negatively charged electrons. William Lawrence Bragg (born in Australia), together with his British-born father William Henry Bragg, developed the technique of X-ray crystallography, in which X-rays are used to determine the three-dimensional structure of solid matter on the atomic scale. The contribution of the Braggs will be outlined further in the chapter that looks at condensed phases. This chapter is primarily concerned with the atom. It will examine the contribution of Rutherford and others to the determination of the structure of the atom, and will show how a particular structural feature of the atom forms the basis of the periodic table of the elements.

FIGURE 1.2 An AFM image of graphene — a single layer of hexagonally arranged carbon atoms



1.1 The essential concepts in brief

LEARNING OBJECTIVE 1.1 Define atoms, molecules, ions, elements and compounds.

Before we can begin our discussion of chemistry, you need to be familiar with various concepts. We will introduce these briefly here and discuss them in greater detail later in the text.

Chemistry is the study of **matter**, which is anything that has mass and occupies space. Chemists view matter as being composed of various chemical entities. **Atoms** are discrete chemical species comprising a central positively charged nucleus surrounded by one or more negatively charged **electrons**. Atoms are always electrically neutral, meaning that the number of electrons is equal to the number of protons in the nucleus. Chemists regard the atom as the fundamental building block of all matter, so it may surprise you to learn that individual atoms are rarely of chemical interest; free atoms (with the exception of the elements helium, neon, argon, krypton, xenon and radon) are usually unstable. Of much greater interest to chemists are **molecules**, which are collections of atoms with a definite structure held together by chemical bonds. The smallest molecules contain just two atoms, while the largest can consist of literally millions. Most gases and liquids consist of molecules, and most solids based on carbon (organic solids) are also molecular. Like atoms, molecules are electrically neutral and are, therefore, uncharged. Molecules are held together by **covalent bonds**, which involve the sharing of electrons between neighbouring atoms.

Ions are chemical species that have either a positive or negative electric charge. Those with a positive charge are called **cations**; those with a negative charge are called **anions** (respectively designated by a $^+$ or $^-$). Ions can be formally derived from either atoms or molecules by the addition or removal of one or more electrons. For example, removing an electron (e^-) from a sodium, Na, atom gives the Na^+ cation.

$$Na \rightarrow Na^+ + e^-$$

Adding an electron to an oxygen molecule, which consists of two oxygen atoms bonded together and is designated O_2, gives the O_2^- (superoxide) anion.

$$O_2 + e^- \rightarrow O_2^-$$

Elements are collections of one type of atom only. At the time of writing, 118 elements are known. **Compounds** are substances containing two or more elements in a definite and unchanging proportion. Compounds may be composed of molecules, ions or covalently bonded networks of atoms. The **chemical formula** shows the relative number of each type of atom present in a chemical substance. For example, an oxygen molecule contains two oxygen atoms, and therefore has the chemical formula O_2, while a molecule of methane, which contains one carbon atom and four hydrogen atoms, has the chemical formula CH_4. Note that we do not have individual 'molecules' of an ionic compound such as sodium chloride. The chemical formula of sodium chloride, NaCl, simply represents the smallest repeating unit in an enormous three-dimensional array of Na^+ ions and Cl^- ions. The same applies to certain covalently bonded structures. For example, quartz, which is composed of an 'infinite' three-dimensional network of covalently bound Si and O atoms, has the chemical formula SiO_2, which refers not to individual SiO_2 'molecules' but to the smallest repeating unit in the network.

All of the above chemical entities (atoms, molecules, ions, elements and compounds) may be involved as **reactants** in **chemical reactions**, processes in which they undergo transformations generally involving the making and/or breaking of chemical bonds, and which usually result in the formation of different chemical species called **products**.

1.2 The atomic theory

LEARNING OBJECTIVE 1.2 Explain how the concept of atoms developed.

Today, we take the existence of atoms for granted. We can explain many aspects of the structure of the atom and, in fact, current technology allows us to 'see' and even manipulate individual atoms, as we saw in the introduction to this chapter and as further described in the chemical connections feature on imaging atoms. However, scientific evidence for the existence of atoms is relatively recent, and chemistry did not progress very far until that evidence was found.

CHEMICAL CONNECTIONS

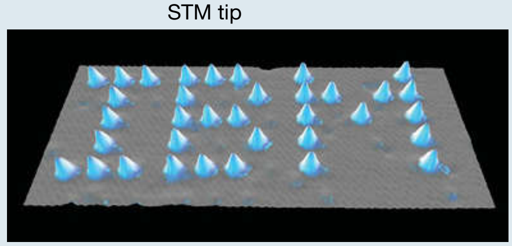
Imaging atoms

We cannot use optical microscopes to see atoms. This is because the dimensions of atoms are smaller than the wavelength of visible light. If we use shorter wavelength radiation, such as a beam of electrons, we can obtain images like those shown at the start of this chapter. However, the apparatus required to obtain such images is expensive and the samples require a significant degree of preparation and careful handling.

In the late twentieth century, two inventions — the scanning tunnelling microscope (STM) and the atomic force microscope (AFM) — revolutionised the imaging of objects having dimensions of the order of nanometres, and have allowed us to 'see' and, more remarkably, manipulate individual atoms. The STM and AFM operate using the same principle — moving the tip of an extremely fine stylus across a surface at a distance of atomic dimensions.

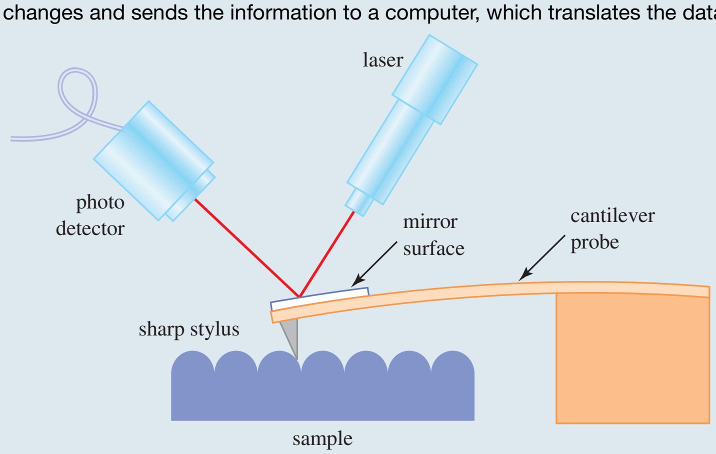
In the case of the STM, the surface must be electrically conducting, and this causes a current to flow between the surface and the tip. The magnitude of this current depends on the distance between the tip and the surface, so as the tip is moved across the surface, computer control of the current at a constant value will cause the tip to move up and down, thereby giving a map of the surface. Because of its tiny size, the tip can also be used to move individual atoms. This was first demonstrated in 1989 when Don Eigler, a scientist at IBM, manipulated 35 atoms of xenon on a nickel surface using an STM to spell the name of his employer (figure 1.3).

FIGURE 1.3 Individual Xe atoms (blue dots) on a nickel surface manipulated by an STM tip



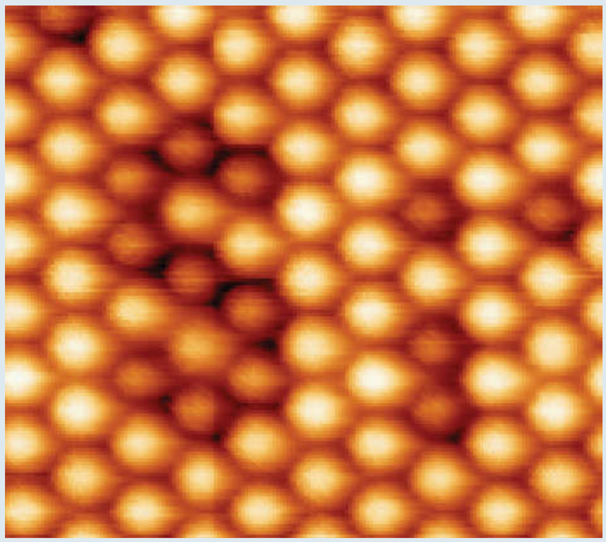
An AFM (illustrated in figure 1.4) is used to study nonconducting samples. The stylus is moved across the surface of the sample under study. Forces between the tip of the probe and the surface cause the probe to flex as it follows the ups and downs of the bumps that are the individual molecules and atoms. A mirrored surface attached to the probe reflects a laser beam at angles proportional to the amount of deflection of the probe. A sensor picks up the signal from the laser and translates it into data that can be analysed by a computer to give three-dimensional images of the sample's surface.

FIGURE 1.4 In an AFM, a sharp stylus attached to the end of a cantilever probe rides up and down over the surface features of the sample. A laser beam, reflected off a mirrored surface at the end of the probe, changes angle as the probe moves up and down. A photodetector reads these changes and sends the information to a computer, which translates the data into an image.



A typical AFM image is shown in figure 1.5. It involves manipulating single atoms to give what is probably the smallest known writing.

FIGURE 1.5 The world's smallest writing? The element symbol for silicon, Si, is spelled out with individual silicon atoms (dark) among tin atoms (light). The silicon atoms were manipulated with the tip of an AFM.



The concept of atoms began nearly 2500 years ago when the Greek philosopher Leucippus and his student Democritus expressed the belief that matter is ultimately composed of tiny indivisible particles; the word 'atom' is derived from the Greek word *atomos*, meaning 'not cut'. The philosophers' conclusions, however, were not supported by any scientific evidence; they were derived simply from philosophical reasoning. The concept of atoms remained a philosophical belief, having limited scientific usefulness, until the discovery of two laws of chemical combination in the late eighteenth century — the law of conservation of mass and the law of definite proportions. These may be stated as follows.

- The **law of conservation of mass**: No detectable gain or loss of mass occurs in chemical reactions. Mass is conserved.
- The **law of definite proportions**: In a given chemical compound, the elements are always combined in the same proportions by mass.

The French chemist Antoine Lavoisier (1743–1794) proposed the law of conservation of mass as a result of his experiments involving the individual reactions of the elements phosphorus, sulfur, tin and lead with oxygen. He used a large lens to focus the sun's rays on a sample of each element contained in a closed jar, and the heat caused a chemical reaction to take place. He weighed the closed jar and its contents before and after the chemical reaction and found no difference in mass, leading him to propose the law. (Lavoisier was beheaded following the French Revolution, the judge at his trial reputedly saying 'the Republic has no need of scientists'.) The law of conservation of mass can be alternatively stated as 'mass is neither created nor destroyed in chemical reactions'.

Another French chemist, Joseph Louis Proust (1754–1826), was responsible for the law of definite proportions, following experiments that showed that copper carbonate prepared in the laboratory was identical in composition to copper carbonate that occurs in nature as the mineral malachite. He also showed that the two oxides of tin, SnO and SnO_2, and the two sulfides of iron, FeS and FeS_2, always contain fixed relative masses of their constituent elements. The law states that chemical elements always combine in a definite fixed proportion by mass to form chemical compounds. Thus, if we analyse *any* sample of water (a compound), we *always* find that the ratio of oxygen to hydrogen (the elements that make up water) is 8 to 1 by mass. Similarly, if we form water from oxygen and hydrogen, the mass of oxygen consumed will always be 8 times the mass of hydrogen that reacts. This is true even if there is a large excess of one of them. For instance, if 100 g of oxygen is mixed with 1 g of hydrogen and the reaction to form water is initiated, all the hydrogen would react but only 8 g of oxygen would be consumed; there would be 92 g of oxygen left over. No matter how we try, we cannot alter the chemical composition of the water formed in the reaction.

WORKED EXAMPLE 1.1

Applying the law of definite proportions

The element vanadium, V, can combine with oxygen, O, to form a compound called vanadium pentoxide. The primary use of this compound is as a catalyst in the production of sulfuric acid, the most produced chemical in the world. A sample of vanadium pentoxide contains 1.274 g of V for each 1.000 g of O. If a different sample of the compound contains 2.250 g of O, what mass of V does it contain?

Analysis

The law of definite proportions states that the proportions of V and O by mass must be the same in both samples. To solve the problem, we will set up the mass ratios for the two samples. In the ratio for the second sample the mass of vanadium will be an unknown quantity. We will use the two ratios to determine the unknown quantity.

Solution

The first sample has a V to O mass ratio of:

$$\frac{1.274\ \text{g V}}{1.000\ \text{g O}}$$

We know the mass of O in the second sample, but not the mass of V. We do know, however, that the V to O mass ratio is the same as that in the first sample. We set up the ratio for the second sample using *x* for the unknown mass of V. Therefore, from the law of definite proportions, we can write the following.

$$\frac{1.274\ \text{g V}}{1.000\ \text{g O}} = \frac{x\ \text{g V}}{2.250\ \text{g O}}$$

We can solve for *x* by multiplying both sides of the equation by 2.250 g O, to give:

$$x\ \text{g V} = \frac{1.274\ \text{g V} \times 2.250\ \text{g O}}{1.000\ \text{g O}} = 2.867\ \text{g V}$$

Is our answer reasonable?

To avoid errors, it is always wise to do a rough check of the answer. Usually, some simple reasoning is all we need to see if the answer makes sense. This is how we might do such a check here: notice that the mass of oxygen in the second sample is more than twice the mass in the first sample. Therefore, we should expect the mass of V in the second sample to be somewhat more than twice what it is in the first. The answer we obtained, 2.867 g V, is more than twice 1.274 g V, so our answer seems to be reasonable.

PRACTICE EXERCISE 1.1

Titanium dioxide, TiO_2, is a compound that is used as a brilliant white pigment in artists' oil colours, as well as in coatings and plastics. A sample of this compound was found to be composed of 1.00 g of titanium and 0.668 g of oxygen. If a second sample of the same compound contains 2.50 g of oxygen, what mass of titanium does it contain?

The laws of conservation of mass and definite proportions served as the experimental foundation for the atomic theory. At the beginning of the nineteenth century, John Dalton (1766–1844), an English scientist, used the Greek concept of atoms to make sense of the laws of conservation of mass and definite proportions. Dalton reasoned that, if atoms really exist, they must have certain properties to account for these laws. He described such properties, and the following list constitutes what we now call **Dalton's atomic theory**.

1. Matter consists of tiny particles called atoms.
2. Atoms are indestructible. In chemical reactions, the atoms rearrange but they do not themselves break apart.
3. In any sample of a pure element, all the atoms are identical in mass and other properties.
4. The atoms of different elements differ in mass and other properties.
5. When atoms of different elements combine to form a given compound, the constituent atoms in the compound are always present in the same fixed numerical ratio.

Dalton's theory easily explained the law of conservation of mass. According to the theory, a chemical reaction is simply a reordering of atoms from one combination to another. If no atoms are gained or lost, and if the masses of the atoms can't change, the mass after the reaction must be the same as the mass before. This explanation of the law of conservation of mass allows us to use a notation system of **chemical equations** to describe chemical reactions. A chemical equation contains the reactants on the left-hand side and the products on the right-hand side, separated by a forward arrow, as demonstrated in the following chemical equation for the formation of liquid water from its gaseous elements.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

The law of conservation of mass requires us to have the same number of each type of atom on each side of the arrow; this being the case, the chemical equation above is described as balanced. We will discuss this concept in detail in the chapter on stoichiometry. Note that this chemical equation also specifies the physical states of the reactants and product. Gases, liquids and solids are abbreviated as (g), (l) and (s), respectively, after each reactant and product.

The law of definite proportions can also be explained by Dalton's theory. According to the theory, a given compound is always composed of atoms of the same elements in the same numerical ratio. Suppose, for example, that elements X and Y combine to form a compound in which the number of atoms of X equals the number of atoms of Y (i.e. the atom ratio is 1 to 1). If the mass of a Y atom is twice that of an X atom, then every time we encounter a sample of this compound, the mass ratio (X to Y) would be 1 to 2. This same mass ratio would exist regardless of the size of the sample so, in samples of this compound, elements X and Y are always present in the same proportion by both number and mass.

Strong support for Dalton's theory came when Dalton and other scientists studied elements that can combine to give at least two compounds. For example, sulfur and oxygen can combine to form both sulfur dioxide, SO_2, and sulfur trioxide, SO_3. The former contains one atom of sulfur and two atoms of oxygen, while the latter contains one atom of sulfur and three atoms of oxygen. Although they have similar chemical formulae, they are different chemically; for example, at room temperature, SO_2 is a colourless gas while SO_3, which melts at 16.8 °C, is a solid or liquid, depending on the temperature of the room. If we analyse samples of SO_2 and SO_3 in which the masses of sulfur are the same, we obtain the results shown in table 1.1.

TABLE 1.1 Mass composition of sulfur dioxide and sulfur trioxide

Compound	Mass of sulfur	Mass of oxygen
SO_2	1.00 g	1.00 g
SO_3	1.00 g	1.50 g

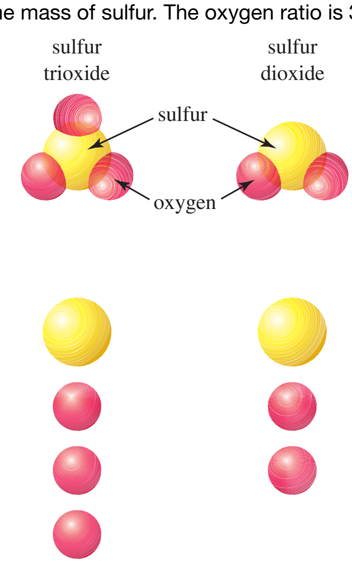
Note that the ratio of the masses of oxygen in the two samples is one of small whole numbers.

$$\frac{\text{mass of oxygen in sulfur trioxide}}{\text{mass of oxygen in sulfur dioxide}} = \frac{1.50\,\text{g}}{1.00\,\text{g}} = \frac{3}{2}$$

Similar observations are made when we study other elements that form more than one compound with each other. These observations form the basis of the **law of multiple proportions**, which states that: Whenever two elements form more than one compound, the different masses of one element that combine with the same mass of the other element are in the ratio of small whole numbers.

Dalton's theory explains the law of multiple proportions in a very simple way. A molecule of sulfur trioxide contains 1 sulfur and 3 oxygen atoms, and a molecule of sulfur dioxide contains 1 sulfur and 2 oxygen atoms (figure 1.6). If we had just one molecule of each, our samples would each contain 1 sulfur atom and, therefore, the same mass of sulfur. Then, comparing the oxygen atoms, we find they are in a numerical ratio of 3 to 2. But because oxygen atoms all have the same mass, the mass ratio must also be 3 to 2. The law of multiple proportions was not known before Dalton presented his theory, and its discovery demonstrates science in action. Experimental data suggested to Dalton the existence of atoms, and the atomic theory suggested the relationships that we now call the law of multiple proportions. When found by experiment, the existence of the law of multiple proportions added great support to the atomic theory. In fact, for many years, it was one of the strongest arguments in favour of the existence of atoms.

FIGURE 1.6 Compounds containing oxygen and sulfur demonstrate the law of multiple proportions. Represented here are molecules of sulfur trioxide, SO_3, and sulfur dioxide, SO_2. Each contains one sulfur atom, and therefore the same mass of sulfur. The oxygen ratio is 3 to 2, both by atoms and by mass.



1.3 The structure of the atom

LEARNING OBJECTIVE 1.3 Describe the structure of the atom.

Even though absolute proof of the existence of atoms was not available around the turn of the twentieth century, scientists were interested in the structure of the atom. While Dalton's theory said that atoms were indestructible and could not be broken apart, experiments around this time showed this was not necessarily true. In particular, the discovery of radiation in the form of X-rays by Wilhelm Röntgen (1845–1923) in 1895 (see figure 1.7) and radioactivity by Antoine Henri Becquerel (1852–1908, pictured in figure 1.8) in

1896 led scientists to believe that the atom was composed of discrete particles, as both forms of radiation involve the release of particles from atoms, thought at that time to be indivisible.

FIGURE 1.7 A reproduction of one of the first ever X-ray images, taken by Wilhelm Röntgen on 22 December 1895. The hand is that of his wife.

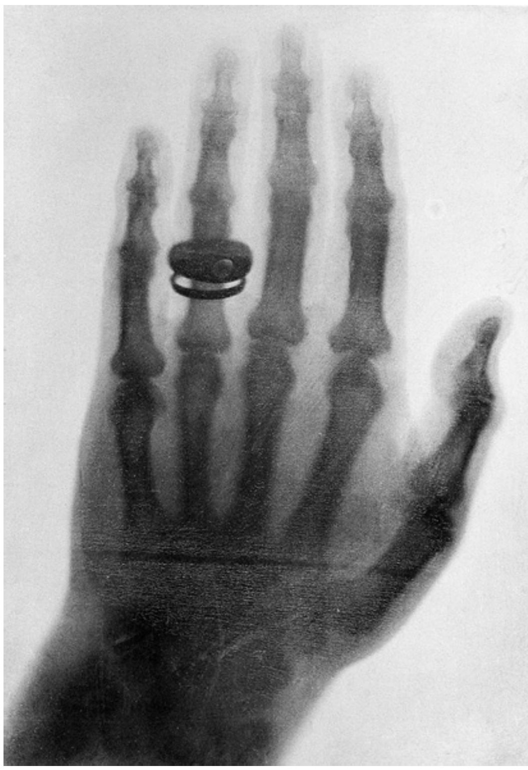
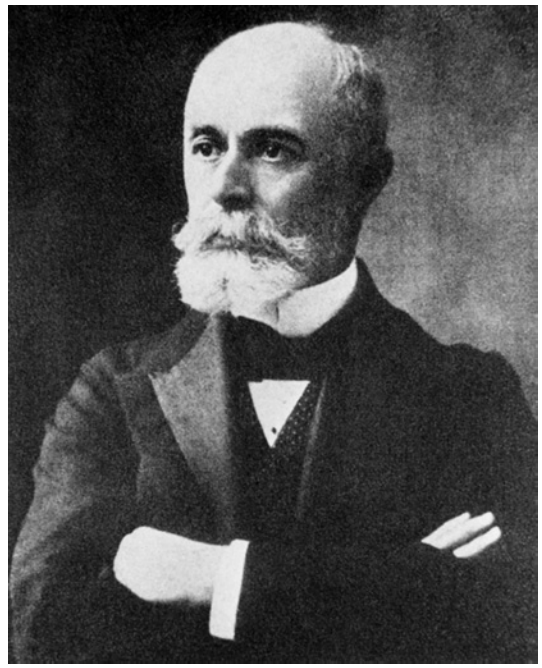
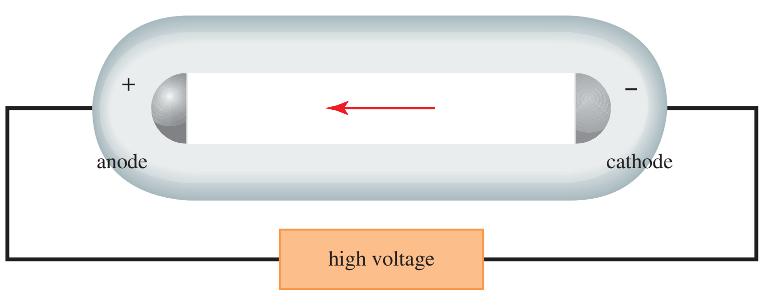


FIGURE 1.8 Antoine Henri Becquerel (1852–1908), a French physicist, discovered radioactivity and was awarded the 1903 Nobel Prize in physics



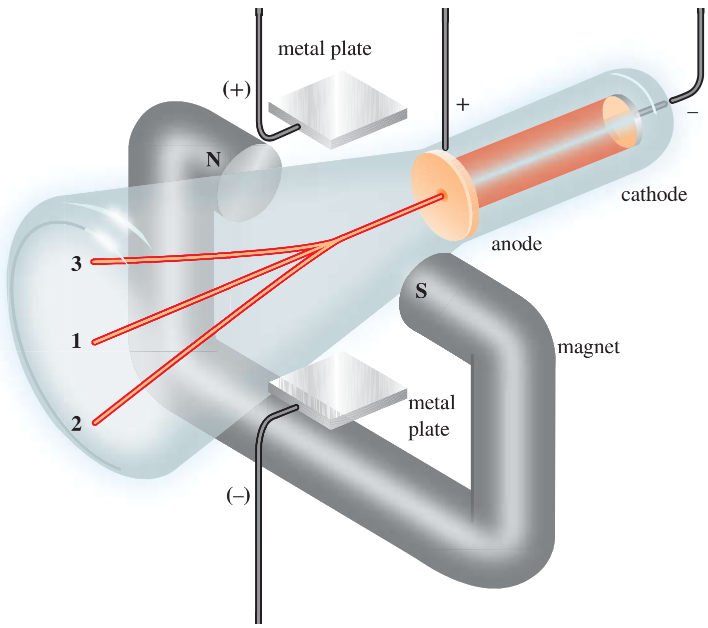
Further evidence for the presence of discrete particles in atoms came from experiments with gas discharge tubes, such as that shown in figure 1.9. When the tube was filled with a low pressure gas and a high voltage was applied between the electrodes, negatively charged particles flowed from the negative electrode (cathode) to the positive electrode (anode).

FIGURE 1.9 Diagram of a gas discharge tube



Because they emanated from the cathode, the particles were called cathode rays. In 1897, the British physicist JJ Thomson passed cathode rays through a magnetic field using the modified discharge tube shown in figure 1.10. The magnetic field caused the path of the cathode rays to bend. Analysis of this effect allowed Thomson to determine the charge to mass ratio of the components of cathode rays, what we now know as electrons.

FIGURE 1.10 Diagram of the apparatus used by JJ Thomson to determine the charge to mass ratio of the electron. The cathode ray beam passes between the poles of a magnet and between a pair of metal electrodes that can be given electric charges. The magnetic field tends to bend the beam in one direction (to point 2) while the charged electrodes bend the beam in the opposite direction (to point 3). By adjusting the charge on the electrodes, the two effects can be made to cancel each other (point 1). The amount of charge on the electrodes required to balance the effect of the magnetic field can be used to calculate the charge to mass ratio.



In 1909, the American chemist Robert Millikan determined the charge on an individual electron by measuring the rates at which charged oil drops fell between electrically charged plates. This, combined with Thomson's results, allowed calculation of the mass of an electron as 9.09×10^{-31} kg. The knowledge that atoms were electrically neutral meant that the electron must have a positively charged counterpart, but its exact nature was not known by the early years of the twentieth century. It was the work of the New Zealand-born scientist Ernest Rutherford (1871–1937) that shed light not only on the positively charged component of the atom, but also on the structure of the atom itself. Around 1909, Rutherford, who had already been awarded the Nobel Prize in chemistry in 1908 for his work on the theory of radioactivity, devised his famous gold foil experiment depicted in figure 1.11. Rutherford took an incredibly thin sheet of gold (only a few atoms thick) and bombarded it with a stream of positively charged particles called **alpha particles**.

Most of the particles went straight through the foil essentially undisturbed, some were deflected through various angles, and about 1 in 8000 was deflected almost straight back towards the source. Of this observation, Rutherford said, 'It was almost as incredible as if you had fired a fifteen-inch [38.1-centimetre] [artillery] shell at a piece of tissue paper and it came back and hit you'. To explain his observations, Rutherford proposed a new model of the atom. He suggested that every atom has a tiny positively charged central core, which he called the **nucleus**, that constitutes most of the mass of the atom.

The positive charge in the nucleus is due to particles, which he called **protons**, and the number of these in the nucleus determines the identity of the atom. The electrical neutrality of the atom requires that there is the same number of electrons in an atom as there are protons in the nucleus, and these surround the central core, as shown schematically in figure 1.12.

FIGURE 1.11 Schematic view of Rutherford's gold foil experiment. When a beam of positively charged alpha particles was 'shot' at a thin gold foil, most of them passed straight through the foil. Some, however, were deflected straight back towards the source.

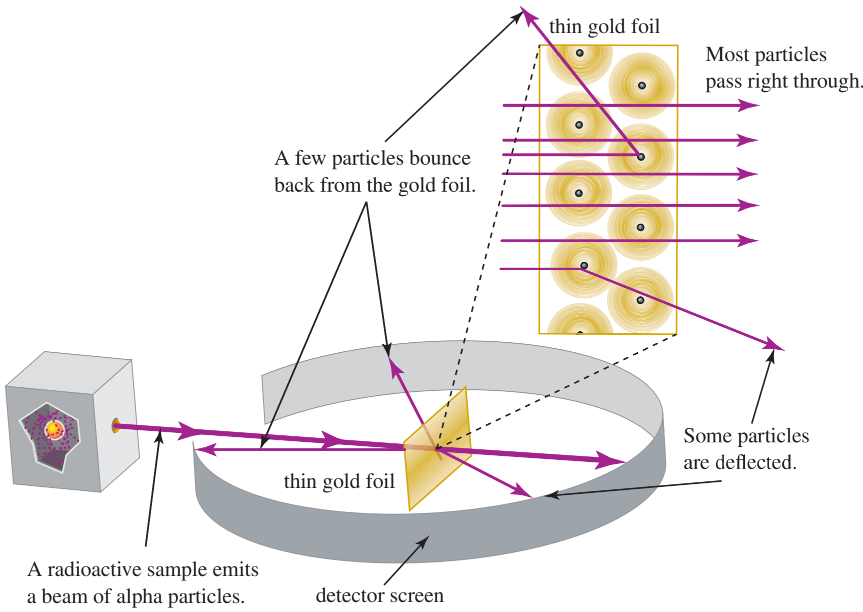
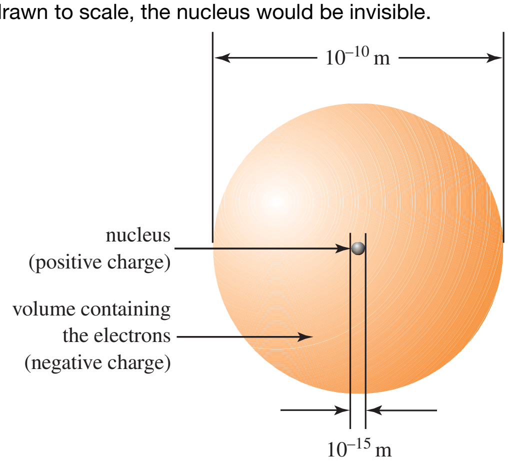


FIGURE 1.12 Diagram (not to scale) showing the nucleus and the volume occupied by the surrounding electrons. If this were drawn to scale, the nucleus would be invisible.



Electrons occupy a volume that is huge compared with the size of the nucleus, but each electron has such a small mass that alpha particles are not deflected by the electrons. Consequently, an alpha particle is deflected only when it passes very near a nucleus, and it bounces straight back only when it collides head-on with a nucleus. Because most of the volume of an atom is essentially empty space, most alpha particles pass through the foil without being affected. From the number of particles deflected and the pattern of deflection, Rutherford calculated that the positive nucleus occupies less than 0.1% of the total atomic volume. We now know that the percentage of the atomic volume is far less than this — to give you some idea of the relative volumes, an atom the size of a rugby stadium would have a nucleus the size of a pea. When Rutherford calculated the nuclear mass based on the number of protons in the nucleus, the value always fell short of the actual mass. In fact, Rutherford found that only about half of the nuclear mass could be accounted for by protons. This led him to suggest that there was another particle in the nucleus that had a mass close to or equal to that of a proton, but with no electric charge. This suggestion initiated a search that finally ended in 1932 with the discovery of the **neutron** by a British physicist named James Chadwick (1891–1974). Because they are found in the nucleus, protons and neutrons are sometimes called **nucleons**. Table 1.2 summarises the **subatomic particles** present in this model of the atom.

Over the intervening years, it has been shown that protons and neutrons are themselves composed of still smaller particles called quarks. The existence of quarks has helped us understand how the atomic nucleus can stay together despite the presence of positively charged protons in close proximity. However, quarks are very unstable outside the confines of the atomic nucleus and are of more interest to physicists than chemists.

TABLE 1.2 **Physical data for the electron, proton and neutron**

Particle	Symbol	Charge (C)	Mass (kg)	Mass (u)
electron	e^-	-1.6022×10^{-19}	9.1094×10^{-31}	5.4858×10^{-4}
proton	p	$+1.6022 \times 10^{-19}$	1.6726×10^{-27}	1.0073
neutron	n	0	1.6749×10^{-27}	1.0087

Note: The charge is measured in coulombs (C). The final column gives the mass in atomic mass units (u); $1\,u = 1.660\,54 \times 10^{-27}$ kg ($\frac{1}{12}$ the mass of the ^{12}C atom).

CHEMICAL CONNECTIONS

How big, or small, is an atom?

The introduction to this chapter stated that atoms are tiny. But just how small is 'tiny'? In order to give you some idea of the atomic scale, we're going to count all the individual atoms in a New Zealand 10-cent piece (the copper-coloured coin in figure 1.13). If we started today, counting at one atom per second, how long do you think this would take? Days? Months? Years? Decades? Have a guess before you read any further.

In order to calculate the number of atoms that we'll be counting, we need to know that a 10-cent piece is predominantly made up of iron (steel), with a thin copper coating, and that it weighs 3.30 g.

We'll assume that a year is 365.25 days and therefore contains 31 557 600 seconds. If you're an average Australasian, you can expect to live for about 80 years, which translates to around 2 500 000 000 (two-and-a-half billion) seconds. So if you started counting at one atom per second on the day you were born, by the time you reached 80, you'd have counted about two-and-a-half billion atoms. But that's not nearly enough.

The oldest known living animal was a clam found in Iceland in 2006. Counting the rings on its shell showed it was 507 years old when it was dredged from the seabed (a process that unfortunately killed it).

FIGURE 1.13 How long would it take to count all the atoms in a New Zealand 10-cent coin?

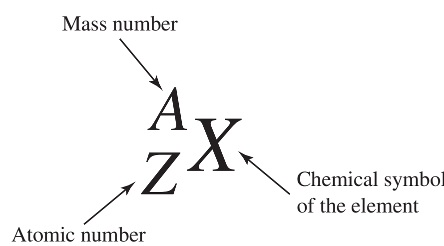
Counting one atom per second for this length of time would get us to about 16 billion atoms. But we're not there yet.

The oldest living thing on Earth is thought to be a patch of seagrass in the Mediterranean that has been dated at around 200 000 years old. An atom a second for that length of time gives us about 6.3 trillion (6 300 000 000 000) atoms. You might think we'd be getting close to counting all the atoms in the 10-cent piece by now, but you'd be wrong. We're not even 1% of the way there. In fact, we're not even anywhere near 1%. We're obviously going to be counting for a while.

The Earth is thought to be about 4.6 billion years old. So let's assume that we've been around since the formation of the Earth, and we've been diligently counting the atoms in the 10-cent piece at one atom per second. After this time, we'd be old, out of breath, and we would have counted around 145 quadrillion (145 000 000 000 000 000) atoms. Surely we must have counted all the atoms by now?

Nope! Still not even close. In fact, the atoms that we've counted by this time would weigh around 13 micrograms, probably just enough for us to see with the naked eye, but nowhere near the number in our 10-cent piece. Even if we counted a thousand atoms per second from the time of the formation of the Earth, we would still be only about 0.5% of the way there.

There are approximately 35 500 000 000 000 000 000 000 atoms in a New Zealand 10-cent piece and, at an atom per second, it would take us about 1 quadrillion years, or roughly 85 000 times the age of the universe, to count these. This is a number of truly incomprehensible magnitude. More incredible still, if we stacked these atoms one on top of another, the resulting chain would go around the Earth about 125 000 times.



To examine how atoms are constructed, we will consider the simplest possible atom, hydrogen, with the **chemical symbol** H. A hydrogen atom consists of a single proton in the nucleus, as well as a single electron. We designate this as $^1_1$H. We use this terminology for any chemical element X as follows.

The **atomic number (*Z*)** is the number of protons in the nucleus. The **mass number (*A*)** is the number of protons in the nucleus plus the number of neutrons (N) in the nucleus.

Note that the atomic number is also equal to the number of electrons in a *neutral* atom (i.e. one in which the number of protons and the number of electrons is the same). A chemical element is defined by its atomic number; all atoms having the same atomic number are atoms of the same element. Therefore, the symbol $^1_1$H tells us that an atom of hydrogen contains 1 proton ($Z = 1$), 1 electron and 0 neutrons ($A = 1$).

If we were to analyse a sample of hydrogen atoms, we would find that roughly 1 atom in every 6600 would have approximately twice the mass of a $^1_1$H atom. These heavier atoms belong to an isotope of hydrogen called deuterium. **Isotopes** are atoms of an element with the same number of protons (i.e. the same value of Z) but different numbers of neutrons (i.e. different values of A). Deuterium atoms are symbolised as $^2_1$H, meaning that there is 1 proton ($Z = 1$) and 1 neutron ($A = 2$) in the nucleus. The $^1_1$H atom is sometimes called protium to distinguish it from deuterium. In chemical terms, deuterium atoms behave essentially identically to hydrogen atoms, but there are some important differences in reactivity when they are bonded to other atoms. Protium and deuterium are examples of **stable isotopes**. This means that the nuclei of these atoms do not undergo any decay processes and are stable indefinitely. In addition, hydrogen also has a third isotope called tritium, $^3_1$H, which has 1 proton and 2 neutrons in the nucleus. It is the least abundant isotope of hydrogen, with only 1 to 10 atoms of tritium in every 10^{18} atoms of hydrogen. Tritium is **radioactive**, meaning that the nucleus is unstable and undergoes spontaneous decay to give an atom of helium, He, a process we will look at in greater detail in the chapter on nuclear chemistry. Helium atoms are characterised by having 2 protons in the nucleus ($Z = 2$). Helium has two stable isotopes, $^3_2$He and $^4_2$He, with 1 and 2 neutrons, respectively, in the nucleus. The element with 3 protons in the nucleus, lithium ($Z = 3$), has the stable isotopes $^{6}_{3}Li$ and $^{7}_{3}Li$ with 3 and 4 neutrons, respectively. Any atom of a specified A and Z is called a **nuclide**. A **radionuclide** is a radioactive nuclide.

WORKED EXAMPLE 1.2

The composition of atoms

The following radioactive isotopes have medical applications. Determine the number of protons, neutrons and electrons in each isotope.

(a) $^{165}_{66}Dy$ (used in the treatment of arthritis)

(b) $^{131}_{53}I$ (used in the treatment of thyroid cancer)

(c) $^{59}_{26}Fe$ (used in studies of iron metabolism)

Analysis

The number of protons is equal to the atomic number (Z), the number of neutrons is found from Z and the mass number (A), and the number of electrons in a neutral atom must equal the number of protons.

Solution

(a) Dy is the chemical symbol for dysprosium. The subscript 66 is Z, which is the number of protons in the nucleus. The superscript 165 is A. We find the number of neutrons by subtracting Z from A: $A - Z = 165 - 66 = 99$ neutrons. Because this is a neutral atom, the number of electrons must equal the number of protons. $^{165}_{66}Dy$ has 66 protons, 99 neutrons and 66 electrons.

(b) I is the symbol for iodine. $Z = 53$ and $A = 131$. Z tells us that the nucleus contains 53 protons. Subtracting Z from A, we find that there are 78 neutrons in this isotope. Finally, the atom is neutral, so there are 53 electrons.

(c) Iron, Fe, has $Z = 26$. A neutral atom of $^{59}_{26}Fe$ has 26 protons, 26 electrons and $59 - 26 = 33$ neutrons.

Is our answer reasonable?

In all cases, the number of protons is equal to the number of electrons, as required for neutral atoms. We have followed the rules for calculating the number of neutrons and have carried out the calculations properly. Our answers should therefore be correct.

PRACTICE EXERCISE 1.2

Determine the number of protons and neutrons in each of the following radioactive isotopes.

(a) $^{177}_{71}Lu$ (used as an imaging and therapeutic agent)

(b) $^{133}_{54}Xe$ (used in studies of the lungs)

(c) $^{192}_{77}Ir$ (used in the form of an internal wire for cancer treatment)

Inclusion of the atomic number in this terminology is almost redundant when the chemical symbol is included, so it is common to see a shorthand version that excludes this. Thus, we often write $^{1}_{1}H$ as simply ^{1}H, as we know that all atoms of hydrogen have $Z = 1$. Using the same shorthand version, deuterium would be written as ^{2}H and tritium as ^{3}H.

Atomic mass

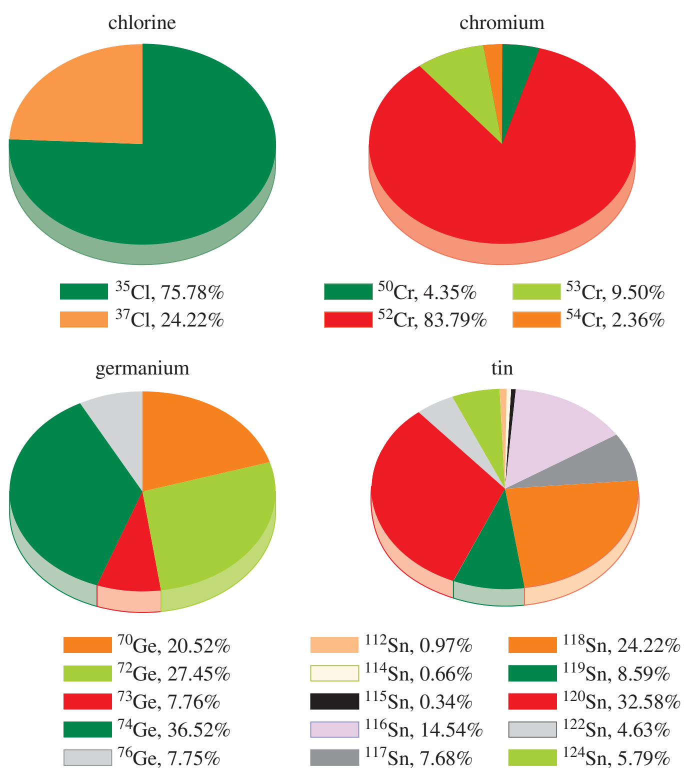
We saw in table 1.2 that the ^{12}C isotope is used as the basis by which **atomic mass** is measured. The **atomic mass unit (u)** is the mass ($1.660\,54 \times 10^{-27}$ kg) equal to $\frac{1}{12}$ the mass of one atom of ^{12}C, and the masses of all atoms are measured relative to this. The atomic mass unit is also known as the Dalton (Da), particularly in biochemistry.

Using this scale, we find that the mass of a single ^{19}F atom is 18.998 403 2 u and that of a single ^{31}P atom is 30.973 762 u. In other words, a single ^{19}F atom weighs $18.998\,403\,2 \times 1.660\,54 \times 10^{-27}$ kg and a single ^{31}P atom weighs $30.973\,762 \times 1.660\,54 \times 10^{-27}$ kg. Because both fluorine and phosphorus have only one naturally occurring isotope, we can be sure that any fluorine atom we chose from a macroscopic sample of fluorine would have a mass of 18.998 403 2 u, while any phosphorus atom chosen from a macroscopic sample of phosphorus would have a mass of 30.973 762 u. We therefore say that the atomic mass of fluorine is 18.998 403 2 u and the atomic mass of phosphorus is 30.973 762 u. However, the majority of elements in the periodic table comprise two or more isotopes, and the mass of a single atom chosen at random from a macroscopic sample of these elements would not be constant — it would depend on which isotope was chosen. We therefore define the atomic mass of these elements as the average mass per atom of a naturally occurring sample of atoms of the element. Consider, for example, the element gallium, Ga. There are two naturally occurring isotopes of Ga, namely $^{69}_{31}$Ga and $^{71}_{31}$Ga, each of which contains 31 protons in the nucleus. Nuclei of the former contain 38 neutrons while those of the latter contain 40. Any naturally occurring sample of gallium will be composed of 60.11% of the $^{69}_{31}$Ga isotope and 39.89% of the $^{71}_{31}$Ga isotope. Given the atomic masses of these isotopes ($^{69}_{31}$Ga = 68.9256 u, $^{71}_{31}$Ga = 70.9247 u) we can calculate the average atomic mass of Ga by taking the sum of the atomic mass of each isotope multiplied by its abundance as follows.

$$\text{average atomic mass of Ga} = (0.6011 \times 68.9256\,\text{u}) + (0.3989 \times 70.9247\,\text{u}) = 69.72\,\text{u}$$

The average atomic mass of Ga, 69.72 u, is just less than the average of the masses of the two isotopes (69.9252 u) because the lighter $^{69}_{31}$Ga isotope is more abundant than the heavier $^{71}_{31}$Ga isotope. Figure 1.14 illustrates the range of isotopic compositions found in four elements, one of which is tin, the element with the largest number of stable isotopes.

FIGURE 1.14 The natural abundances of the isotopes of chlorine (Cl), chromium (Cr), germanium (Ge), and tin (Sn), illustrate the diversity of isotopic distributions. The mass number and relative abundance of each isotope are indicated.



While the distribution of isotopes in samples of most elements is essentially constant, there are 12 elements (H, Li, B, C, N, O, Mg, Si, S, Cl, Br and Tl) which show substantial variation in their isotopic compositions, depending on the source of the element. Consider, for example, hydrogen. As we have seen, this element has three isotopes, ^{1}H, ^{2}H and ^{3}H, the last of which we will neglect in this discussion because of its negligible abundance. If we analysed samples of atmospheric methane and methane from a natural gas well, we would find that the proportion of the ^{2}H isotope in the hydrogen atoms of the former would be greater than that in the latter, and hence the average atomic mass of H in the two samples would be different. Therefore, instead of quoting a single value for the average atomic mass of hydrogen, a range of values of atomic mass is given [1.007 84 u, 1.008 11 u], which corresponds to the lowest and highest values measured in natural samples. Table 1.3 gives the range of atomic mass values for these 12 elements, together with the conventional atomic masses which are used either for routine work or when the source of the sample is unknown.

TABLE 1.3 **Atomic mass ranges and conventional atomic masses for H, Li, B, C, N, O, Mg, Si, S, Cl, Br and Tl**

Element name	Symbol	Atomic number	Atomic mass range (u)	Conventional atomic mass (u)
hydrogen	H	1	[1.007 84, 1.008 11]	1.008
lithium	Li	3	[6.938, 6.997]	6.94
boron	B	5	[10.806, 10.821]	10.81
carbon	C	6	[12.0096, 12.0116]	12.011
nitrogen	N	7	[14.006 43, 14.007 28]	14.007
oxygen	O	8	[15.999 03, 15.999 77]	15.999
magnesium	Mg	12	[24.304, 24.307]	24.305
silicon	Si	14	[28.084, 28.086]	28.085
sulfur	S	16	[32.059, 32.076]	32.06
chlorine	Cl	17	[35.446, 35.457]	35.45
bromine	Br	35	[79.901, 79.907]	79.904
thallium	Tl	81	[204.382, 204.385]	204.38

In this text, we will use the conventional atomic masses listed in table 1.3 in any calculations involving these 12 elements.

WORKED EXAMPLE 1.3

Calculating average atomic masses from isotopic abundances

Naturally occurring titanium, Ti, is a mixture of five isotopes and has the following isotopic composition.

$$^{46}_{22}Ti(8.25\%),\ ^{47}_{22}Ti(7.44\%),\ ^{48}_{22}Ti(73.72\%),\ ^{49}_{22}Ti(5.41\%),\ ^{50}_{22}Ti(5.18\%)$$

The atomic masses of the isotopes are as follows.

$$^{46}_{22}Ti(45.952\,631\,6\ u),\ ^{47}_{22}Ti(46.951\,763\,1\ u),\ ^{48}_{22}Ti(47.947\,946\,3\ u),\ ^{49}_{22}Ti(48.947\,870\,0\ u),$$
$$^{50}_{22}Ti(49.944\,791\,2\ u)$$

Use this information to calculate the average atomic mass of titanium.

Analysis

In a sample containing many atoms of titanium, 8.25% of the total mass is contributed by atoms of ^{46}Ti, 7.44% by atoms of ^{47}Ti, 73.72% by atoms of ^{48}Ti, 5.41% by atoms of ^{49}Ti and 5.18% by atoms of ^{50}Ti. This means that, when we calculate the mass of the hypothetical 'average atom' of Ti, we have to weight it according to both the masses of the isotopes and their relative abundance. (Keep in mind, of course, that such an atom does not really exist. This is just a simple way to see how we can calculate the average atomic mass of this element.)

Solution

We will calculate 8.25% of the mass of an atom of ^{46}Ti, 7.44% of an atom of ^{47}Ti, 73.72% of an atom of ^{48}Ti, 5.41% of an atom of ^{49}Ti and 5.18% of an atom of ^{50}Ti. Adding these contributions gives the total mass of the 'average atom'. Therefore:

$$\begin{aligned}\text{average atomic mass of Ti} &= (0.0825 \times 45.952\,631\,6\,\text{u}) + (0.0744 \times 46.951\,763\,1\,\text{u}) \\ &\quad + (0.7372 \times 47.947\,946\,3\,\text{u}) + (0.0541 \times 48.947\,870\,0\,\text{u}) \\ &\quad + (0.0518 \times 49.944\,791\,2\,\text{u}) \\ &= 47.687\,\text{u}\end{aligned}$$

Is our answer reasonable?

By far the most abundant isotope is ^{48}Ti, so we would expect the average atomic mass to be close to the mass of this isotope. Our calculated average atomic mass, 47.867 u, is indeed just less than the mass of the ^{48}Ti isotope (47.947 946 3 u), because the lighter ^{46}Ti and ^{47}Ti isotopes are slightly more abundant than the heavier ^{49}Ti and ^{50}Ti isotopes. Hence, we can feel confident our answer is correct.

PRACTICE EXERCISE 1.3

Neon has three naturally occurring isotopes. ^{20}Ne has a mass of 19.9924 u and is 90.48% abundant, ^{21}Ne has a mass of 20.9938 u and is 0.27% abundant, and ^{22}Ne has a mass of 21.9914 u and is 9.25% abundant. Using these data, calculate the average atomic mass of neon.

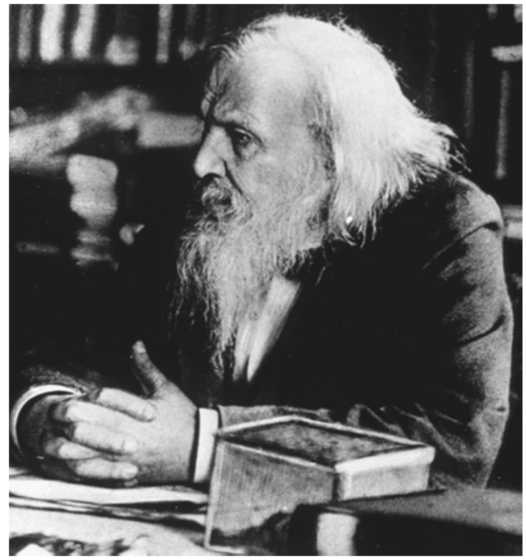
1.4 The periodic table of the elements

LEARNING OBJECTIVE 1.4 Explain the basis of the periodic table of the elements.

We have already seen that the elements H, He and Li can be ordered on the basis of increasing atomic number ($Z = 1, 2$ and 3, respectively). If we continue such an ordering, we obtain the **periodic table of the elements**. The periodic table we use today is based primarily on the efforts of a Russian chemist, Dmitri Ivanovich Mendeleev (1834–1907, pictured in figure 1.15), and a German physicist, Julius Lothar Meyer (1830–1895). Working independently, these scientists developed similar periodic tables only a few months apart in 1869. Mendeleev is usually given the credit, however, because he published his version first.

The extraordinary thing about the work of Mendeleev and Meyer was that they knew nothing of the structure of the atom, so were unaware of the concept of atomic number, which is the basis of the modern periodic table. What they did know, however, were the atomic masses of many of the elements. Bear in mind also that not all of the elements we know today had been discovered at this time. Mendeleev was preparing a chemistry textbook for his students at the University of St Petersburg and, looking for some pattern among the properties of the elements, he found that, when he arranged them in order of increasing atomic mass, similar chemical properties were repeated over and over at regular intervals. For instance, the elements lithium (Li), sodium (Na), potassium (K), rubidium (Rb), and caesium (Cs), are soft metals that react vigorously with water. Similarly, the elements that immediately follow each of these also constitute a set with similar chemical properties. Thus, beryllium (Be), follows lithium; magnesium (Mg), follows sodium; calcium (Ca), follows potassium; strontium (Sr), follows rubidium; and barium (Ba), follows caesium. All of these elements form compounds with oxygen having a 1 : 1 metal to oxygen ratio. Mendeleev used such observations to construct his periodic table, which is illustrated in figure 1.16.

FIGURE 1.15 Dmitri Ivanovich Mendeleev developed the periodic table



At first glance, Mendeleev's original table looks little like the 'modern' table given in figure 1.17. However, a closer look reveals that the rows and columns have been interchanged. The elements in Mendeleev's table are arranged in order of increasing atomic mass. When the sequence is broken at the right places and stacked, the elements fall naturally into columns.

Mendeleev placed elements with similar properties in the same row even when this left occasional gaps in the table. For example, he placed arsenic, As, in the same row as phosphorus because they had similar chemical properties, even though this left gaps in other rows. In a stroke of genius, Mendeleev reasoned, correctly, that the elements that belonged in these gaps had simply not yet been discovered. In fact, on the basis of the location of these gaps, Mendeleev could predict, with astonishing accuracy, the properties of the yet-to-be-found elements, and his predictions helped serve as a guide in the search for them.

The elements tellurium, Te, and iodine, I (note that the German word for 'iodine' is *jod*, which has the abbreviation J in Mendeleev's original table), caused Mendeleev some problems. According to the best estimates at that time, the atomic mass of tellurium was greater than that of iodine. Yet, if these elements were placed in the table according to their atomic masses, they would not fall into the proper rows required by their properties. Therefore, Mendeleev switched their order, believing that the atomic mass of tellurium had been incorrectly measured (it had not), and in so doing violated his ordering sequence based on atomic mass.

The table that Mendeleev developed is the basis of the one we use today, but one of the main differences is that Mendeleev's table lacks the elements helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn). In Mendeleev's time, none of these elements had yet been discovered because they are relatively rare and because they have virtually no tendency to undergo chemical reactions. When these elements were finally discovered, beginning in 1894, another problem arose. Two more elements, argon, Ar, and potassium, K, did not fall into the rows required by their properties if they were placed in the table in the order required by their atomic masses. Another switch was necessary and another exception had been found. It became apparent that atomic mass was not the true basis for the periodic repetition of the properties of the elements. With Rutherford's discovery of the structure of the atom, it became apparent that the elements in the periodic table were arranged in order of increasing atomic *number*, not atomic *mass*, and when this was realised it became obvious that Te and I, and Ar and K, were in fact in the correct positions.

The modern periodic table

The periodic table in use today is shown in figure 1.17. The horizontal rows are called **periods** and are numbered 1 to 7, while the vertical columns are called **groups** and are numbered 1 to 18. The elements are arranged in order of increasing atomic number across each period, and a new period begins after each group 18 element. On the periodic table, the atomic masses are given (generally to four significant figures) below each chemical symbol.

FIGURE 1.16 Mendeleev's original periodic table, taken from the German chemistry journal *Zeitschrift für Chemie*, 1869, 12, 405–6

Ueber die Beziehungen der Eigenschaften zu den Atomgewichten der Elemente. Von D. Mendelejeff. – Ordnet man Elemente nach zunehmenden Atomgewichten in verticale Reihen so, dass die Horizontalreihen analoge Elemente enthalten, wieder nach zunehmendem Atomgewicht geordnet, so erhält man folgende Zusammenstellung, aus der sich einige allgemeinere Folgerungen ableiten lassen.

			Ti = 50	Zr = 90	? = 180
			V = 51	Nb = 94	Ta = 182
			Cr = 52	Mo = 96	W = 186
			Mn = 55	Rh = 104,4	Pt = 197,4
			Fe = 56	Ru = 104,4	Ir = 198
			Ni = Co = 59	Pd = 106,6	Os = 199
H = 1			Cu = 63,4	Ag = 108	Hg = 200
	Be = 9,4	Mg = 24	Zn = 65,2	Cd = 112	
	B = 11	Al = 27,4	? = 68	Ur = 116	Au = 197?
	C = 12	Si = 28	? = 70	Sn = 118	
	N = 14	P = 31	As = 75	Sb = 122	Bi = 210?
	O = 16	S = 32	Se = 79,4	Te = 128?	
	F = 19	Cl = 35,5	Br = 80	J = 127	
Li = 7	Na = 23	K = 39	Rb = 85,4	Cs = 133	Tl = 204
		Ca = 40	Sr = 87,6	Ba = 137	Pb = 207
		? = 45	Ce = 92		
		?Er = 56	La = 94		
		?Yt = 60	Di = 95		
		?In = 75,6	Th = 118?		

1. Die nach der Grösse des Atomgewichts geordneten Elemente zeigen eine stufenweise Abänderung in den Eigenschaften.
2. Chemisch-analoge Elemente haben entweder übereinstimmende Atomgewichte (Pt, Ir, Os), oder letztere nehmen gleichviel zu (K, Rb, Cs).
3. Das Anordnen nach den Atomgewichten entspricht der *Werthigkeit* der Elemente und bis zu einem gewissen Grade der Verschiedenheit im chemischen Verhalten, z. B. Li, Be, B, C, N, O, F.
4. Die in der Natur verbreitetsten Elemente haben *kleine* Atomgewichte

While the atomic mass usually increases with atomic number, you can see the exceptions we mentioned previously (Te and I; Ar and K) as well as Co and Ni. While the isotopic composition and, therefore, the atomic masses of most elements are well established, there are some unstable elements of all the isotopes, which undergo spontaneous radioactive decay. Given that the isotopic composition of such elements cannot be known, it is usual to simply quote the mass number of the longest lived isotope of the element, and these are given in parentheses in the periodic table. Note that there are discontinuities in the periodic table between elements 56 and 72, and between elements 88 and 104, and these two sets of elements are given below the table itself. The elements from 57 to 71 are called the **lanthanoids** (or, less commonly, the **rare earth elements**). Elements 89 to 103 are called the **actinoids**. The lanthanoids and actinoids are generally situated below the rest of the periodic table, simply to save space and to make the table easier to read; note that the lanthanoid and actinoid elements are chemically distinct from the rest of the elements in the periodic table, and do not belong to any of the groups 1 to 18. The lanthanoids and actinoids are sometimes called the ***f*-block elements**, and similar terminology is also used elsewhere in the table; elements in groups 1 and 2 are called the ***s*-block elements**, elements in groups 3 to 12 are called the ***d*-block elements**, and elements in groups 13 to 18 are called the ***p*-block elements**. As we will see, *s*, *p*, *d* and *f* refer to orbitals, particular regions in space in the atom where electrons have a high probability of being found. The *d*-block elements are also called **transition metals**.

FIGURE 1.17 The periodic table of the elements. At the time of writing, 118 elements were known. Atomic masses in parentheses refer to the longest lived isotope of the element. The elements outlined in table 1.3, which have a range of atomic masses, have square brackets around their conventional atomic masses in the periodic table.

Key: atomic number / Symbol / standard atomic mass

metals | nonmetals | metalloids

	alkali metals	alkaline earth metals													pnictogens	chalcogens	halogens	noble gases
	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H [1.008]																	2 He 4.003
2	3 Li [6.94]	4 Be 9.012											5 B [10.81]	6 C [12.01]	7 N [14.01]	8 O [16.00]	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg [24.31]											13 Al 26.98	14 Si [28.09]	15 P 30.97	16 S [32.06]	17 Cl [35.45]	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.97	35 Br [79.90]	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc 97.91	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
6	55 Cs 132.9	56 Ba 137.3	57–71	72 Hf 178.5	73 Ta 180.9	74 W 183.8	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl [204.4]	82 Pb 207.2	83 Bi 209.0	84 Po (209.0)	85 At (210.0)	86 Rn (222.0)
7	87 Fr (223.0)	88 Ra (226.0)	89–103	104 Rf (265.1)	105 Db (268.1)	106 Sg (271.1)	107 Bh (270.1)	108 Hs (277.2)	109 Mt (276.2)	110 Ds (281.2)	111 Rg (280.2)	112 Cn (285.2)	113 Nh (284.2)	114 Fl (289.2)	115 Mc (288.2)	116 Lv (293.2)	117 Ts (294.2)	118 Og (294.2)

lanthanoid series	57 La 138.9	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm 144.9	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
actinoid series	89 Ac 227.0	90 Th 232.0	91 Pa 231.0	92 U 238.0	93 Np 237.0	94 Pu 244.1	95 Am 243.1	96 Cm 247.1	97 Bk 247.1	98 Cf 251.1	99 Es 252.1	100 Fm 257.1	101 Md 258.1	102 No 259.1	103 Lr 262.1

Individual groups within the periodic table are also known by particular names, although this practice is less prevalent than in the past. Group 1 elements are called **alkali metals**, group 2 elements are called **alkaline earth metals**, group 15 elements are called **pnictogens**, group 16 elements are called **chalcogens**, group 17 elements are called **halogens** and group 18 elements are called **noble gases**. Of these, only the terms halogens and noble gases are in common usage.

All elements on the periodic table belong to one of three categories — metals, nonmetals and metalloids — and the groupings are shown by the different colours on the periodic table in figure 1.17. **Metals** are generally good conductors of heat and electricity, are malleable (can be beaten into a thin sheet) and ductile (can be drawn out into a wire), and have the usual metallic lustre. Elements that do not have these characteristics are called **nonmetals**, and the majority of these are gases at room temperature and pressure. The properties of **metalloids** lie somewhere between the metals and nonmetals. The most notable property of these elements is the fact that they tend to be semiconductors, and metalloids such as silicon, Si, and germanium, Ge, have therefore found wide use in silicon chips and transistors. Note that the classification of the recently prepared elements Lv, Ts and Og is somewhat arbitrary, as weighable quantities of these have not yet been obtained.

Naming the elements

All of the elements in the periodic table have one- or two-letter abbreviations of their names. The abbreviations of many elements are simply the first one or two letters of their names (e.g. carbon, C; oxygen, O; lithium, Li) but there are quite a number of elements for which the derivation of the abbreviation is not quite so obvious: for example, potassium, K, tin, Sn, lead, Pb, and iron, Fe. Such apparent anomalies occur because of the way that the elements were historically named. Nowadays, when a new element is discovered, the discoverer usually gets to suggest a name for the element, which is then ratified by IUPAC, the International Union of Pure and Applied Chemistry.

Of all the elements on the periodic table, carbon (C), sulfur (S), iron (Fe), copper (Cu), arsenic (As), silver (Ag), tin (Sn), antimony (Sb), gold (Au), mercury (Hg), lead (Pb) and bismuth (Bi) were known to ancient civilisations so the date of their 'discovery' is not known. Of these, the element symbols Fe, Cu, Ag, Sn, Sb, Au, Hg and Pb were derived from the Latin names *ferrum*, *cuprum*, *argentum*, *stannum*, *stibium*, *aurum*, *hydrargyrum* and *plumbum*. The earliest known discovery of an element was that of phosphorus, P. It was isolated in 1669 by the German alchemist Hennig Brand from the distillation of urine (he was apparently trying to make silver or gold — unsuccessfully, of course!) and was named after the Greek word *phosphoros*, meaning 'bringer of light', as the element glows in the dark (see figure 1.18). Elements have been named after countries (germanium, Ge, francium, Fr, americium, Am, polonium, Po) and even after the places they were first discovered; the Swedish town of Ytterby has the distinction of having four elements (erbium, Er, ytterbium, Yb, yttrium, Y, and terbium, Tb) named after it, as these were first found in mineral deposits close to the town. Surprisingly few elements have been named after people; at present, only 17 people have been immortalised on the periodic table, and they are listed in table 1.4.

FIGURE 1.18 Phosphorus, the first element whose date of discovery is known, was isolated in 1669 by alchemist Hennig Brand as he tried to make silver or gold by distilling urine

TABLE 1.4 People after whom elements have been named

Name	Brief biography	Element named
Vasilii Yefrafovich von Samarski-Bykhovets (1803–1870)	Chief of staff of the Russian Corps of Mining Engineers	samarium, Sm (element 62)
Johan Gadolin (1760–1852)	Finnish chemist; first person to isolate a lanthanoid element	gadolinium, Gd (element 64)

(continued)

TABLE 1.4 *(continued)*

Name	Brief biography	Element named
Pierre Curie (1859–1906) Marie Curie (1867–1934)	Husband and wife scientific team; Pierre (French) and Marie (Polish by birth); jointly awarded the Nobel Prize in physics in 1903	curium, Cm (element 96)
Albert Einstein (1879–1955)	Most famous scientist of the twentieth century, if not all time; German by birth; awarded the Nobel Prize in physics in 1921	einsteinium, Es (element 99)
Enrico Fermi (1901–1954)	Italian physicist; made great advances in the study of nuclear reactions; awarded the Nobel Prize in physics in 1938	fermium, Fm (element 100)
Dmitri Mendeleev (1834–1907)	Russian chemist; renowned for the development of the periodic table	mendelevium, Md (element 101)
Alfred Nobel (1833–1896)	Swedish inventor of dynamite and patron of the Nobel Prizes	nobelium, No (element 102)
Ernest Lawrence (1901–1958)	American inventor of the cyclotron; awarded the Nobel Prize in physics in 1939	lawrencium, Lr (element 103)
Ernest Rutherford (1871–1937)	New Zealand physicist/chemist; made seminal contributions to understanding the structure of the atom; awarded the Nobel Prize in chemistry in 1908	rutherfordium, Rf (element 104)
Glenn Seaborg (1912–1999)	American chemist; first prepared many of the elements beyond uranium in the periodic table; awarded the Nobel Prize in chemistry in 1951	seaborgium, Sg (element 106)
Niels Bohr (1885–1962)	Danish physicist; studied electronic energy levels within atoms, which aided our understanding of the atom; awarded the Nobel Prize in physics in 1922	bohrium, Bh (element 107)
Lise Meitner (1878–1968)	Austrian physicist; made fundamental discoveries concerning nuclear fission; controversially never awarded a Nobel Prize	meitnerium, Mt (element 109)
Wilhelm Röntgen (1845–1923)	German physicist; discoverer of X-rays; awarded the inaugural Nobel Prize in physics in 1901	röntgenium, Rg (element 111)
Nicolaus Copernicus (1473–1543)	Polish astronomer; proposed that the sun, rather than the Earth, was the centre of the solar system	copernicium, Cn (element 112)
Georgii Flerov (1913–1990)	Russian physicist; made significant discoveries in the syntheses of transuranium elements	flerovium Fl (element 114)
Yuri Oganessian (1933–)	Russian physicist; has made important advances in the syntheses of superheavy elements	oganesson, Og (element 118)

1.5 Electrons in atoms

LEARNING OBJECTIVE 1.5 Detail the role of electrons in atoms.

While we have touched briefly on the concept of electrons, we have to this point concentrated primarily on the nucleus of the atom and the way in which the number of protons in the nucleus determines the chemical identity of the atom. However, many of the chemical properties of an atom and, most importantly, its chemical reactivity are determined primarily by the electrons.

One of the most interesting things about electrons is that we cannot really say *exactly* where they are at any particular time, so we usually talk about their *most probable* locations. Electrons occupy regions of space called **orbitals** in atoms. Each orbital has a characteristic electron distribution and energy. For example, the lowest energy situation for a hydrogen atom, the **ground state**, occurs when the single electron occupies an orbital in which its most probable distance from the nucleus is 5.29×10^{-11} m. If we were to take snapshots of the position of the electron in this orbital over time, we would find a spherical distribution. If the ground-state hydrogen atom absorbs a specific amount of energy, the electron can be promoted to a higher energy orbital to form an **excited state** in which the electron lies, on average, further from the nucleus. Such a process is called an **electronic transition**, and the electron distribution in the higher energy orbital is dumbbell shaped. Similarly, the electron in an excited-state hydrogen atom can move to a lower energy orbital through the emission of energy, often in the form of light. Indeed, as we shall see in the chapter on atomic energy levels, such processes are the basis behind both neon and sodium vapour lights. Orbitals have definite energies, so the energy of any electron is dictated by the energy of the orbital it occupies; therefore, an electron in an atom can have only certain well-defined energies. This is a fundamental principle of the science of quantum mechanics called **quantisation**, a phenomenon first proposed by the German physicist Max Planck (1858–1947; Nobel Prize in physics, 1918) in 1900. We will learn more about the quantisation of energy in the chapter on atomic energy levels.

Electrons have a single negative charge, and the overall charge on any chemical species is determined by the number of electrons relative to the number of protons; for example, the oxide ion, O^{2-}, has a 2− charge because there are two more electrons (10) than protons (8) in the ion. Similarly, the Li^+ ion contains three protons and two electrons, so it has a single positive charge. In addition to their negative charge, all electrons have an intrinsic property called **spin**. This can have one of two values, which are commonly called 'spin up' and 'spin down' and are often depicted as follows.

$$\uparrow \text{ (spin up)}$$

$$\downarrow \text{ (spin down)}$$

Each orbital within an atom can contain a maximum of two electrons, one of which must be spin up and the other spin down.

Chemists are interested in electrons because they constitute the chemical bonds that hold atoms together in molecules. Covalent chemical bonds usually consist of one, two or three pairs of electrons shared between atoms, each pair containing electrons of opposite spin. For a molecule to undergo a chemical reaction, usually these bonds must be broken and new ones made; this requires a reorganisation of the electron pairs between the reactant and product molecules, and the ease with which this can be done determines how fast the reaction occurs. Reactions in which one or more electrons are formally transferred between chemical species are also known; such reactions, known as **redox reactions**, are important in a huge number of chemical and biochemical processes; in fact, as you are reading this, iron ions and oxygen molecules are busy exchanging electrons in your blood to transport oxygen around your body.

Because of their importance in both chemical structure and chemical reactivity, electrons occupy a central place in chemistry. In the remaining chapters of this text, we will learn more of the properties of atoms and molecules that are predominantly dictated by electrons.

We have learned much about the atom in the years since Rutherford's seminal experiment. Indeed, so far we have detailed only the very basics of atomic structure; later chapters will outline some of the amazing complexity of the atom. For the moment, it is sufficient for you to appreciate that the atom is composed of a positively charged central nucleus containing protons and neutrons, which is surrounded by negatively charged electrons that can undergo transitions only between well-defined energy levels. And with only 118 different types of these building blocks, we can construct the universe.

SUMMARY

1.1 Define atoms, molecules, ions, elements and compounds.

Atoms are the fundamental building blocks of all matter. Uncharged collections of atoms bonded together in a definite structure are called molecules. These are held together by covalent bonds that share electrons between adjacent atoms. Ions are charged chemical species that may be derived from both atoms and molecules. Cations are positively charged, while anions are negatively charged. Elements comprise only a single type of atom, while compounds are made up of two or more chemical elements. All of these different chemical entities can be involved as reactants in chemical reactions, in which they are transformed to products.

1.2 Explain how the concept of atoms developed.

All matter is composed of atoms. The existence of atoms was proposed on the basis of the following.

- The law of conservation of mass — mass is conserved in chemical reactions
- The law of definite proportions — elements are combined in the same proportions by mass in any particular compound
- The law of multiple proportions — when two elements form more than one compound, the different masses of one element that combine with the same mass of the other element are in the ratio of small whole numbers.

Dalton's atomic theory was the first to propose the existence of atoms on the basis of scientific observations. The basic tenets of his theory are as follows.

1. Matter consists of tiny particles called atoms.
2. Atoms are indestructible. In chemical reactions, the atoms rearrange but they do not themselves break apart.
3. In any sample of a pure element, all the atoms are identical in mass and other properties.
4. The atoms of different elements differ in mass and other properties.
5. When atoms of different elements combine to form a given compound, the constituent atoms in the compound are always present in the same fixed numerical ratio.

Dalton's theory allows us to use chemical equations, in which reactants and products are separated by an arrow, to describe chemical reactions. Such equations are balanced when they contain the same number of each type of atom on each side of the arrow.

Modern apparatus enables us to 'see' individual atoms, and atomic theory is now atomic fact.

1.3 Describe the structure of the atom.

Although Dalton proposed the atom to be indivisible, experiments in the late nineteenth century showed this was not the case. The negatively charged electron was the first subatomic particle to be discovered, while Rutherford's gold foil experiment, in which a thin gold sheet was bombarded with alpha particles, gave evidence for a small, positively charged nucleus. The positive charge is due to subatomic particles called protons, and the number of these in the nucleus determines the identity of the atom in question. The third component of the atom, the neutron, was predicted by Rutherford and found by Chadwick. The atom thus comprises three subatomic particles, the electron, proton and neutron, the latter two collectively being called nucleons. Each type of atom is designated by a chemical symbol, which is determined by its atomic number (Z), the number of protons in the nucleus. The mass number (A) is equal to the number of protons plus the number of neutrons in the nucleus. The terminology used to depict an atom of any element X is ${}^{A}_{Z}X$. All atoms with the same Z are of the same element; however, atoms of the same element can differ in the number of neutrons in the nucleus, and this gives rise to isotopes. Isotopes can be either radioactive (i.e. they decay spontaneously) or stable. A radioactive nucleus is called a radionuclide, while a nuclide is the name given to any atomic nucleus. We can measure atomic mass in atomic mass units (u), where $1\ \text{u} = 1.660\,54 \times 10^{-27}$ kg, and is equal to $\frac{1}{12}$ of the mass of one atom of ^{12}C. The atomic mass of any sample of atoms is the weighted average of the masses of the isotopes present in the sample.

1.4 Explain the basis of the periodic table of the elements.

The periodic table of the elements contains the 118 known elements arranged in order of increasing atomic number, and was developed by both Mendeleev and Meyer. The horizontal rows are called *periods* and the vertical columns *groups*. Elements in the same group tend to have similar chemical properties. The periodic table is divided into sections according to the electron configuration of the elements, namely the *s*-block elements, the *p*-block elements, the *d*-block elements and the *f*-block elements. The *f*-block elements are divided into the lanthanoids (also sometimes called the rare earth elements) and the actinoids, while the *d*-block elements are also called the transition metals. The elements of the periodic table can be classified as metals, nonmetals, or metalloids.

1.5 Detail the role of electrons in atoms.

Electrons occupy regions of space called orbitals. The lowest energy arrangement of electrons in the orbitals of an atom is called the ground state. Electrons can be promoted to higher energy orbitals by absorption of energy to give excited states; conversely, electrons in higher energy orbitals can move to lower energy orbitals with the emission of energy, often as light. Such processes are called electronic transitions. The energies of electrons in atoms are determined by the energies of the orbitals within the atom, so electrons in atoms can have only certain well-defined energies. This is called quantisation, a fundamental principle of quantum mechanics. Electrons have a single negative charge, and one of two possible spins. An orbital in an atom can hold a maximum of two electrons, which must be of opposite spin. Covalent bonds comprise one, two or three pairs of electrons. Chemical reactions often involve reorganising these electrons in bond-making and bond-breaking processes. Redox reactions involve the transfer of one or more electrons between chemical species.

KEY CONCEPTS AND EQUATIONS

Concept	Section	Description/equation
The law of conservation of mass	1.2	The total mass of reactants present before a reaction starts equals the total mass of products after the reaction is finished. We can use this law to check whether we have accounted for all the substances formed in a reaction.
The law of definite proportions	1.2	If we know the mass ratio of the elements in one sample of a compound, the ratio will be the same in a different sample of the same compound.
The law of multiple proportions	1.2	In different compounds containing the same two elements, the different masses of one element that combine with the same mass of the other element are in a ratio of small whole numbers.
Atomic mass	1.3	This is used to determine the mass of any atom relative to $\frac{1}{12}$ that of the ^{12}C isotope.
Periodic table of the elements	1.4	This is a table of the chemical elements arranged in order of increasing atomic number. We can use the periodic table to figure out whether a particular element is a metal, nonmetal or metalloid, predict its chemical reactivity, calculate its number of protons and electrons, obtain its atomic mass and so on. In fact, all of chemistry is contained within the periodic table.

KEY TERMS

actinoids Elements 89 to 103 of the periodic table.
alkali metals The elements in group 1 (except hydrogen) of the periodic table.
alkaline earth metals The elements in group 2 of the periodic table.
alpha particle The nucleus of a helium atom ($^{4}_{2}He$).
anion A negatively charged ion.
atom A neutral particle having one nucleus; the smallest representative sample of an element.
atomic mass The average mass (in u) of the atoms of the isotopes of a given element as they occur naturally.
atomic mass unit (u) The mass ($1.660\,54 \times 10^{-27}$ kg) equal to $\frac{1}{12}$ of the mass of one atom of ^{12}C.
atomic number (*Z*) The number of protons in a nucleus.
cation A positively charged ion.
chalcogens The elements in group 16 of the periodic table.
chemical equation A form of notation used to describe chemical reactions, in which the reactants and products of the reaction are separated by a directional arrow, with the reactants appearing on the left-hand side.
chemical formula A formula written using chemical symbols and subscripts that describes the composition of a chemical compound or element.
chemical reaction A process involving transformation of chemical species into different chemical species, usually involving the making and/or breaking of chemical bonds.
chemical symbol The formula of an element.
compound A chemical substance containing two or more elements in a definite and unchanging proportion.
covalent bond A chemical bond in which two atoms share one or more pairs of electrons.
***d*-block elements** Collective name for the elements in groups 3 to 12 of the periodic table.
Dalton's atomic theory Matter consists of tiny, indestructible particles called atoms and all atoms of one element are identical. The atoms of different elements have different masses. Atoms combine in definite ratios of atoms when they form compounds.
electron A subatomic particle (e, $^{0}_{-1}e$), with a charge of -1 and mass of 5.4858×10^{-4} u (9.1094×10^{-31} kg), that is outside an atomic nucleus; the particle that moves when an electric current flows.
electronic transition The movement of electrons between states of different energies.
element A chemical species consisting of atoms of a single type.
excited state Any state in which a chemical system is not in its lowest possible energy state.
***f*-block elements** A collective name for the lanthanoid and actinoid elements.
ground state The lowest possible energy state of a chemical system.
group A vertical column of elements in the periodic table.
halogens The elements in group 17 of the periodic table.
ion A charged chemical species.
isotopes Atoms of the same element having different numbers of neutrons in their nuclei.
lanthanoids Elements 57 to 71 of the periodic table.
law of conservation of mass No detectable gain or loss in mass occurs in chemical reactions. Mass is conserved.
law of definite proportions In a given chemical compound, the constituent elements are always combined in the same proportion by mass.
law of multiple proportions Whenever two elements form more than one compound, the different masses of one element that combine with the same mass of the other are in a ratio of small whole numbers.
mass number (*A*) The numerical sum of the protons and neutrons in an atom of a given isotope.

matter Anything that has mass and occupies space.
metalloids Elements with properties that lie between those of metals and nonmetals, and that are found in the periodic table around the diagonal line running from boron, B, to astatine, At.
metals Elements that are good conductors of heat and electricity, are malleable (can be beaten into a thin sheet) and ductile (can be drawn out into a wire), and have the usual metallic lustre.
molecule An uncharged collection of atoms bonded together in a definite structure.
neutron A subatomic particle (n, ${}^{1}_{0}n$), with a charge of 0 and a mass of 1.0086 u (1.6749×10^{-27} kg), that exists in all atomic nuclei except those of the ${}^{1}H$ isotope.
noble gases The elements in group 18 of the periodic table.
nonmetals Nonductile, nonmalleable, nonconducting elements.
nucleon A proton or a neutron.
nucleus The dense core of an atom that comprises protons and neutrons.
nuclide A particular atom of specified atomic number and mass number.
orbital A three-dimensional wave describing a bound electron.
***p*-block elements** A collective name for the elements in groups 13 to 18 of the periodic table.
period A horizontal row of elements in the periodic table.
periodic table of the elements A table in which symbols for the elements are displayed in order of increasing atomic number and arranged so that elements with similar properties lie in the same column.
pnictogens The elements in group 15 of the periodic table.
product The chemical species obtained as the result of a chemical reaction.
proton A subatomic particle (${}^{1}_{1}p$), with a charge of +1 and a mass of 1.0073 u (1.6726×10^{-27} kg), that is found in atomic nuclei.
quantisation A phenomenon whereby the energy of a chemical system is not continuous but is restricted to certain definite values.
radioactive Able to emit various atomic radiations or gamma rays.
radionuclide A radioactive isotope.
rare earth elements An alternative name for the lanthanoid elements.
reactant A chemical species that is transformed in a chemical reaction.
redox reaction A reaction involving the transfer of one or more electrons between chemical species.
***s*-block elements** A collective name for the elements in groups 1 and 2 of the periodic table.
spin The intrinsic angular momentum of electrons and protons that gives them magnetism.
stable isotopes Isotopes that do not undergo any decay processes.
subatomic particles Electrons, protons and neutrons.
transition metals The elements in groups 3 to 12 of the periodic table.

REVIEW QUESTIONS

The essential concepts in brief — LO1

1.1 Define the following terms: matter, atom, covalent bond, ion, cation, anion, element, compound, chemical formula, reactant, chemical reaction, product.

The atomic theory — LO2

1.2 Name and state the three laws of chemical combination discussed in this chapter.
1.3 Balanced chemical equations have the same number of atoms of each type on either side of the arrow. Which of the three laws discussed in this chapter require this to be the case, and why?
1.4 Which of the laws of chemical combination are used to define the term ‘compound’?
1.5 How did Dalton’s theory explain the law of conservation of mass?

The structure of the atom **LO3**

1.6 How did the discovery of X-rays and radioactivity support the idea that atoms were not indivisible, but were composed of discrete particles?

1.7 Why did most of the alpha particles in Rutherford's gold foil experiment pass straight through the foil undeflected? Name the force that resulted in the deflection of some of the alpha particles.

1.8 Which component particles contribute most to the mass of an atom? Where in an atom are these particles situated?

1.9 When we calculate the mass of an atom, we generally neglect any contribution to this from electrons in the atom. Why is this?

1.10 Define the term 'nucleon'.

1.11 What is an isotope? Why do isotopes of an element exhibit similar chemical behaviour?

1.12 Consider the symbol ${}^{A}_{Z}X$, where X stands for the chemical symbol for an element. What information is given by (a) A and (b) Z?

1.13 Write the symbols (mass number, atomic number and chemical symbol) of the following isotopes. (Consult a table of atomic numbers or a periodic table, as needed.)

(a) an isotope of gold which contains 118 neutrons
(b) an isotope of fluorine which contains 9 neutrons
(c) an isotope of neodymium which contains 83 neutrons
(d) an isotope of osmium which contains 108 neutrons

The periodic table of the elements **LO4**

1.14 What is the chemical symbol for each of the following elements?

(a) potassium
(b) sodium
(c) arsenic
(d) yttrium
(e) tin
(f) antimony
(g) tungsten
(h) gold
(i) mercury
(j) lead

1.15 What is the name of each of the following elements?

(a) Be
(b) Ru
(c) Pu
(d) Tc
(e) V
(f) Po
(g) Ge
(h) Es
(i) Rf
(j) Ag

1.16 On what basis did Mendeleev construct his periodic table? On what basis are the elements arranged in the modern periodic table?

1.17 The element francium, Fr, is one of the rarest elements that occurs naturally on Earth. It is formed by radioactive decay of heavier elements and there is thought to be only 20–30 g of francium present on Earth at any one time. From its position in the periodic table, would you expect this element to undergo a vigorous reaction with water?

1.18 Why did Mendeleev leave gaps in his periodic table?

1.19 Why does the atomic number of an element allow better prediction of its chemical properties than does its mass number?

1.20 On the basis of their positions in the periodic table, why is it not surprising that ^{90}Sr, a dangerous radioactive isotope of strontium, replaces calcium in newly formed bones?

1.21 When nickel-containing ores are refined, commercial amounts of palladium and platinum are also often obtained. Why is this not unexpected?

1.22 Why would you reasonably expect cadmium to be a contaminant in zinc but not in silver?

1.23 Scientists can produce new heavy elements, with atomic numbers greater than 92. Explain why it is very unlikely that a completely new element with an atomic number of less than 92 will ever be discovered.

1.24 In each of the following sets of elements, state which fits the description in parentheses.
(a) Sm, Cu, Nb, Ba, Ga (*s*-block element)
(b) Bi, Mt, Co, Mg, H (*p*-block element)
(c) At, P, Zr, Ca, Se (transition metal)
(d) Rg, S, Sc, Eu, Al (lanthanoid element)
(e) Yb, Cr, Au, Np, Cl (actinoid element)

1.25 Calculations show that a rod of platinum 10 cm long and 1 cm in diameter can theoretically be drawn out into a wire nearly 28 000 km long. What is this property of metals called?

1.26 Gold can be hammered into sheets so thin that some light can pass through them. Which property of gold allows such thin sheets to be made?

1.27 Name the elements that exist as diatomic gases (gases that exist as molecules containing two atoms) at 25 °C (room temperature) and 1.013×10^5 Pa (atmospheric pressure).

1.28 Which two elements exist as liquids at room temperature and atmospheric pressure?

1.29 Weighable amounts of the very heavy elements, with atomic numbers greater than 112, have not yet been prepared, and so their bulk physical properties are as yet unknown. Which element, Fl (element 114) or Lv (element 116), would be more likely to exhibit properties of a metalloid?

1.30 Sketch the shape of the periodic table and mark off those areas where we find each of the following.
(a) metals
(b) nonmetals
(c) metalloids

Electrons in atoms **LO5**

1.31 What is the name given to the most probable region of space in which an electron might be found?

1.32 When electrons of opposite spin occupy an orbital, we say that their spins are paired. Molecules with odd numbers of electrons, therefore, cannot have all of the electron spins paired, and we say that they have unpaired spins. Which of the following molecules *must* have unpaired spins: N_2, F_2, CO, NO, NO_2?

1.33 An atom in an excited state has a higher energy than the same atom in its ground state. Given that neon lights involve neon atoms in excited states, suggest a method by which the excited state atoms might lose the excess energy they have.

1.34 Quantisation is very important on the atomic scale but, in the large scale of our everyday lives, we barely notice it. Why do you think this might be so?

REVIEW PROBLEMS

1.35 Methane is the simplest of a series of compounds collectively called the alkanes, which consist of only carbon and hydrogen and have the general chemical formula C_nH_{2n+2}. For every 1.000 g of C in a sample of methane there is 0.336 g of hydrogen. Which of the following compositions corresponds to that of methane?
(a) 7.317 g carbon, 8.295 g hydrogen
(b) 2.618 g carbon, 5.228 g hydrogen
(c) 3.884 g carbon, 1.305 g hydrogen
(d) 6.911 g carbon, 4.003 g hydrogen
(e) 9.352 g carbon, 7.417 g hydrogen

1.36 One of the substances used to melt ice on footpaths and roads in cold climates is calcium chloride. In this compound, calcium and chlorine are combined in a ratio of 1.00 g of calcium to 1.77 g of chlorine. Which of the following calcium–chlorine mixtures will produce calcium chloride with no calcium or chlorine left over after the reaction is complete? **LO2**

(a) 3.65 g calcium, 4.13 g chlorine
(b) 0.856 g calcium, 1.56 g chlorine
(c) 2.45 g calcium, 4.57 g chlorine
(d) 1.35 g calcium, 2.39 g chlorine
(e) 5.64 g calcium, 9.12 g chlorine

1.37 Germanium tetrachloride is a dense liquid that is used in the production of fibre-optic cables. Any sample of germanium tetrachloride is composed of germanium and chlorine in the mass ratio of 1.00 : 1.95. If a sample of germanium tetrachloride contains 5.00 g of germanium, how much chlorine does it contain? **LO2**

1.38 A compound of phosphorus and chlorine used in the manufacture of a flame-retardant treatment for fabrics contains 1.20 g of phosphorus for every 4.12 g of chlorine. Suppose a sample of this compound contains 6.22 g of chlorine. What mass of phosphorus does it contain? **LO2**

1.39 With reference to problem 1.37, if 2.00 g of germanium combined completely with chlorine to form germanium tetrachloride, what mass of germanium tetrachloride would be formed? **LO2**

1.40 Refer to the data about the phosphorus–chlorine compound in problem 1.38. If 12.5 g of phosphorus combined completely with chlorine to form this compound, what mass of the compound would be formed? **LO2**

1.41 Combustion of any carbon compound in air forms two major compounds containing only carbon and oxygen. Molecules of one of these compounds contain one atom each of C and O, with the mass ratio of C to O being 1 : 1.332. Molecules of the second compound of carbon and oxygen contain one atom of C and two atoms of O. What mass of oxygen would be combined with each 1.000 g of carbon in this compound? **LO2**

1.42 Tin forms two compounds with chlorine. In one of them (compound 1), there are two Cl atoms for each Sn atom; in the other (compound 2), there are four Cl atoms for each Sn atom. When combined with the same mass of tin, what would be the ratio of the masses of chlorine in the two compounds? In compound 1, 0.597 g of chlorine is combined with each 1.000 g of tin. What mass of chlorine would be combined with 1.000 g of tin in compound 2? **LO2**

1.43 The atomic mass unit is defined in terms of the mass of the ^{12}C atom. Given that 1 atomic mass unit corresponds to $1.660\,54 \times 10^{-24}$ g, calculate the mass of 1 atom of ^{12}C in grams. **LO3**

1.44 Use the mass corresponding to the atomic mass unit given in problem 1.43 to calculate the mass of 1 atom of sulfur. **LO3**

1.45 One of the earliest anaesthetics — its first recorded use was in 1844 — was a compound called nitrous oxide, or, more commonly, laughing gas. Molecules of nitrous oxide are composed of two atoms of nitrogen and one atom of oxygen. In this compound, 1.7513 g of nitrogen is combined with 1.0000 g of O. If the atomic mass of O is 16.00 u, use the above information to calculate the atomic mass of the nitrogen. **LO3**

1.46 Element X forms a compound with oxygen in which there are two atoms of X for every three atoms of O. In this compound, 1.125 g of X is combined with 1.000 g of oxygen. Use the average atomic mass of oxygen to calculate the average atomic mass of X. Use your calculated atomic mass to identify element X. **LO3**

1.47 If an atom of ^{12}C had been assigned a relative mass of 24.0000 u, determine the average atomic mass of hydrogen relative to this mass. **LO3**

1.48 The short-lived radioactive ^{11}C isotope is used to prepare radiolabelled molecules that are used in positron emission tomography (PET), a medical imaging technique. An atom of ^{11}C has a mass that is 0.917 58 times that of a ^{12}C atom. What is the atomic mass of this isotope of carbon expressed in atomic mass units? **LO3**

1.49 Antimony (Sb) has two stable isotopes. ^{121}Sb has a mass of 120.9038 u and an abundance of 57.36%, while ^{123}Sb has a mass of 122.9042 u and an abundance of 42.64%. Use these data to calculate the average atomic mass of antimony. **LO3**