

MODULE 2

metals

Metals play a very important part in our lives. Think about some of your daily activities. The taps and shower head you used in the bathroom this morning were chrome-plated brass (an alloy or mixture of two metals), the pipes the water flowed through were copper, the electricity used for lighting or heating flowed through copper wires, the mirror you used while combing your hair was a film of silver on the back of a sheet of glass, the cutlery used for eating breakfast and the sink you washed it up in were probably stainless steel, the bus, train, car or bicycle that brought you to school was made largely of iron with some parts being aluminium, the drink you bought at recess was probably in an aluminium can, the coins used to pay for it were alloys of copper with nickel or aluminium, the roof of your classroom may be iron coated with zinc, the window frames are probably aluminium, your ballpoint pen has a steel tip and ball, your watch probably has a steel case, your keeper earrings may be gold, your dinner will be cooked in either aluminium or stainless steel saucepans, and the TV or computer you use tonight contains many metals (copper wires, gold contacts, aluminium castings, brass pins on plugs).

The discovery of certain metals and the ability to form them into various tools and weapons have dramatically affected the way civilisation has developed over the past 5000 years—so much so that terms like the Bronze Age and the Iron Age are commonly used in archaeology and ancient history.

The most widely used metals are iron (and its common alloys called steels), copper, aluminium, zinc, lead, tin, chromium, nickel, titanium, magnesium, silver, gold and platinum.

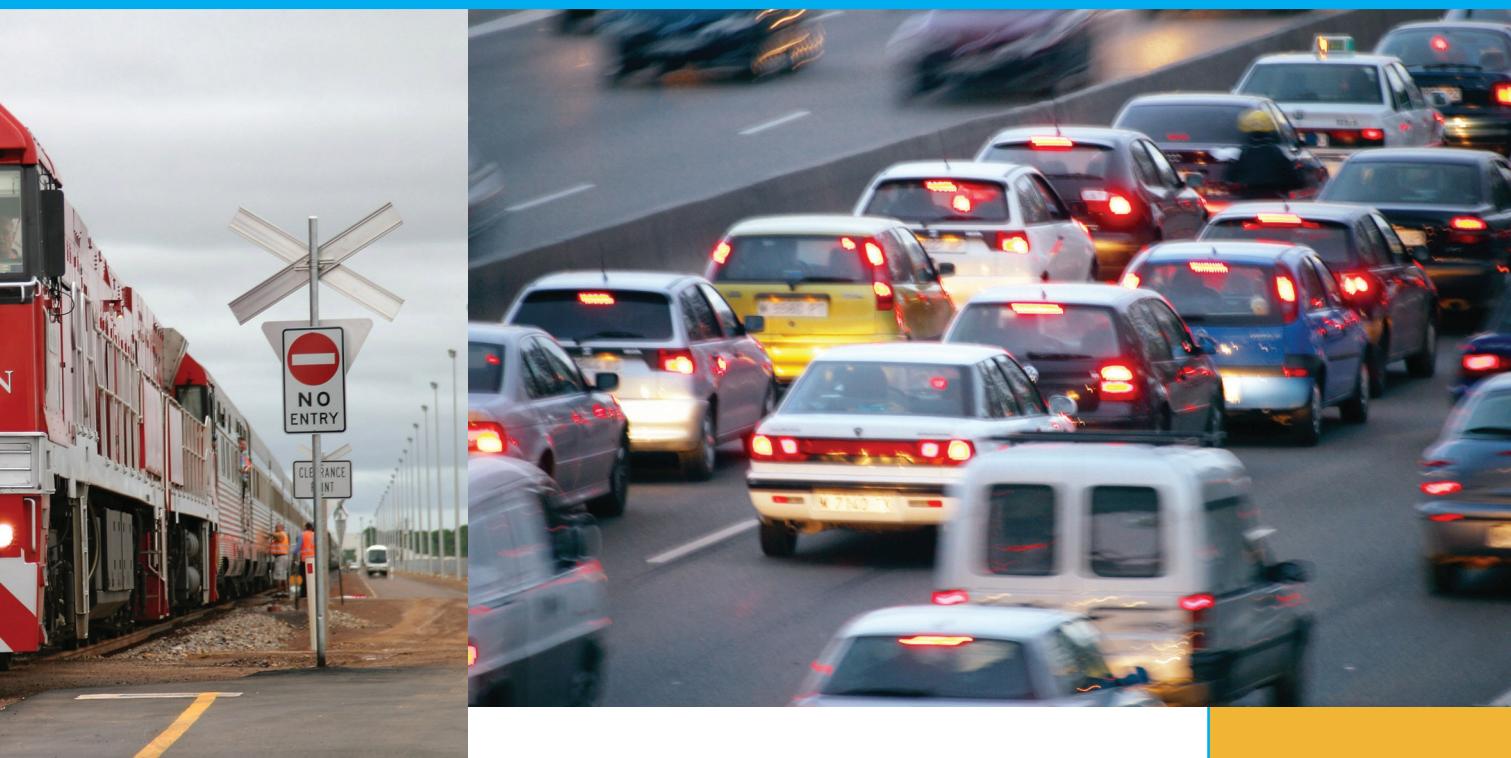


Because metals have been so intimately involved in the development of human civilisations and because they are so very important in our current lifestyles, we shall undertake a study of them—their properties and how their uses relate to these properties, their chemical reactions, and how we extract them from their ores. Because alloys (mixtures) of metals are equally as important as the pure metals, we shall also consider them. This is in Chapter 4.

Metals are generally extracted from their ores by reacting them with some other substances. For efficient use of resources and good economy we should use just enough (and no more) of these ‘other substances’ to bring about the desired change. This means that we need to be able to calculate quantities involved in chemical reactions. So we shall study quantitative aspects of chemical formulae and reactions in Chapter 5.

‘Because metals make up the majority of elements, an examination of the physical and chemical properties of metals is also an appropriate context in which to consider the organisation of the common Periodic Table.’<sup>†</sup> So in the last part of this module we will discuss the Periodic Table. Chapter 6 will outline its historical development, look at how the modern table is structured and survey the type of information that we can get from it.

<sup>†</sup> from *Chemistry Stage 6 Syllabus*, NSW Board of Studies, 2002, p. 28.



## PLANES, TRAINS AND AUTOMOBILES

are all examples of everyday use of metals such as aluminium, steel and copper



# CHAPTER 4

## Metals: properties, uses and reactions

### IN THIS CHAPTER

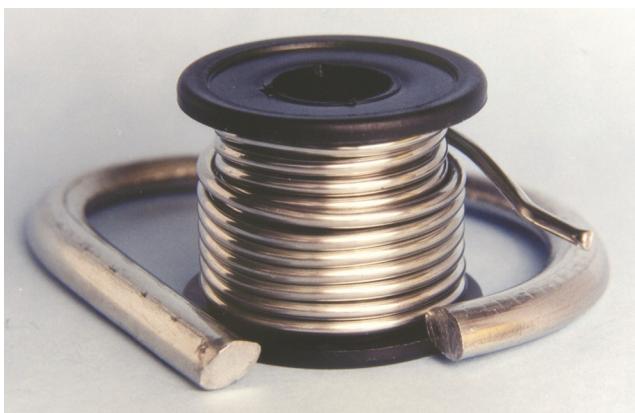
Nature of alloys  
Uses of metals and alloys  
Properties of metals  
How uses of metals relate to properties  
Metals through history  
Extraction of metals from ores  
Extraction of copper  
Chemical equations for the extraction of copper  
Energy considerations

Reactions of metals with oxygen, water and dilute acids  
An activity series for metals and its use in selecting metals for particular purposes  
How ease of extraction and use through history relate to reactivity  
Economic aspects of metal use  
Recycling of metals, particularly aluminium

As the preamble indicates, metals are widely used in our industrialised society. However, today many metals are used not as pure substances but rather as alloys.

Generally the ‘other elements’ in an alloy are metals. An alloy is a *solid solution*. Because alloys are mixtures, they do not have constant compositions (as compounds do). Brass can contain between 45 and 65% copper and still be brass.

An **alloy** is a homogeneous mixture of a metal with one or more other elements.



Common alloys are brass (copper and zinc), bronze (copper and tin), solder (tin and lead) and stainless steel (mainly iron with 12 to 18% chromium)

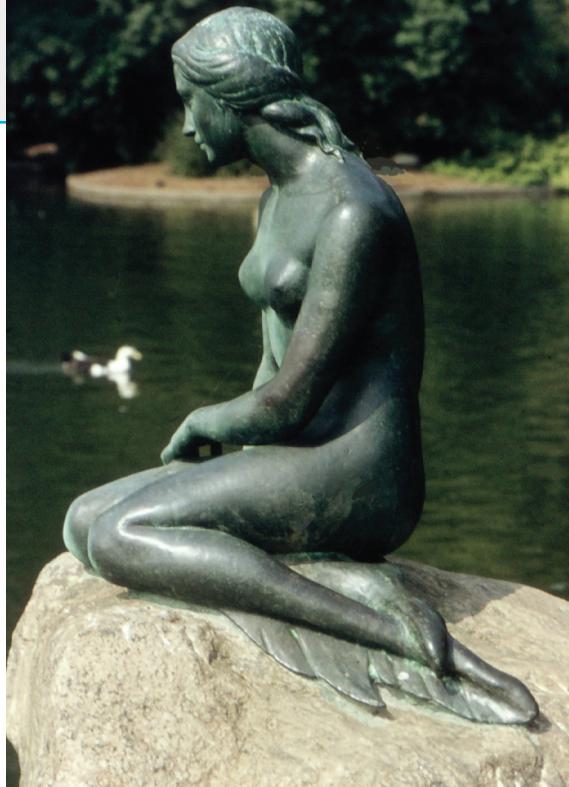
LEFT Solders are alloys of tin and lead  
BELOW Our silver coins are a copper, nickel alloy while the gold ones are a copper, aluminium, nickel alloy



ABOVE Brass is used for many household items, both functional and decorative

RIGHT Bronze is widely used for statues

and generally other metals, particularly nickel, as well). Steel is an alloy of less than 2% carbon in iron.



## 4.1 USES OF METALS AND ALLOYS

Some uses of the more common metals are shown in Table 4.1.

In many of the alloys in common use the predominant metal is iron; these are called **steels** (or sometimes ferrous alloys). Table 4.2 lists common steels and their uses. The composition and uses of non-ferrous alloys are shown in Table 4.3.

**TABLE 4.1 Major uses of common metals (or alloys which are more than 90% of the named metals)**

### Iron and steel (an alloy of iron with <2% carbon)

- railways, bridges, buildings (particularly roofing)
- motor car bodies, ships and trains
- engine blocks, fire hydrants, drainage pipes and grates
- reinforcing in concrete (roads, bridges, high-rise buildings)
- refrigerators, washing machines and other domestic appliances
- heavy machinery in industry
- containers (drums and tin cans)
- Pipes, nails, nuts and bolts

### Aluminium

- buildings (window and door frames, panelling)
- aeroplanes
- motor car parts (cylinder heads, radiator cores)
- high voltage transmission lines
- domestic pots and pans and wrapping foil
- drink containers

**TABLE 4.1 Major uses of common metals (or alloys which are more than 90% of the named metals) (continued)**

**Copper**

- electrical wiring (household and street cables)
- pipes and plumbing fittings
- electroplating, jewellery and household decorations

**Zinc**

- galvanising iron (surface coating)
- protective paints and electroplating
- diecast alloys (car carburettors) and brass
- casings for dry cells (batteries)

**Lead**

- in car batteries
- plumbing (flashings) and in solder
- in 'crystal' glass and as glaze for pottery

Steel is widely used for bridges, cranes and building reinforcing, and numerous hardware items



Many household items are made of aluminium.  
Zinc is used for battery casings and for galvanising, as on the nails here

**TABLE 4.2 Types of steel (alloys of iron) and their uses**

Type of steel	Composition (balance is iron)	Particular properties	Uses
<b>Carbon steels</b>			
mild steel	<0.2% carbon	soft, malleable	car bodies, pipes, nuts and bolts, roofing
structural steel	0.3 to 0.6% carbon	hard, high tensile strength	beams and girders, railways, reinforcing
high-carbon steel	0.6 to 1.5% carbon	very hard	knives and tools such as drill bits, chisels, hammers
<b>Alloy steels</b>			
chrome steel	2 to 4% chromium	hard and shock resistant	safes, files, ball bearings
tungsten steel	10 to 20% tungsten, 0 to 5% chromium and vanadium	hard at high temperatures	cutting and grinding tools
silicon steel	2 to 5% silicon	easily magnetised and demagnetised	electromagnets and transformer cores
<b>Stainless steels</b>			
	10 to 20% chromium, 5 to 20% nickel	hard, resist corrosion	food processing machinery, kitchen sinks and appliances, cutlery, surgical and dental instruments, some razor blades

**TABLE 4.3 Common non-ferrous alloys and their uses**

Alloy	Composition	Special properties	Uses
brass	50 to 60% copper with zinc	lustrous gold appearance, hard but easily machined	plumbing fittings, musical instruments, decorations
bronze	80 to 90% copper with tin	hard, resists corrosion, easily cast	ships' propellers, casting statues
duralumin	95% aluminium, 4% copper, 1% manganese	low density, very strong	aircraft parts, racing bicycles
solder	30 to 60% tin with lead	low melting point, adheres firmly to other metals when molten	joining metals together, particularly in plumbing and electronics
copper-nickel coinage alloy	75% copper with nickel	silvery appearance, resists corrosion	5, 10, 20, 50 cent coins
'gold' coinage alloy	92% copper, 6% aluminium, 2% nickel	gold appearance, resists corrosion	\$1 and \$2 coins
18 carat gold	75% gold	harder than pure gold (24 carat)	jewellery
9 carat gold	37.5% gold for both, balance is roughly equal amounts of silver and copper		

A particular metal is used for a particular purpose because its properties make it suitable for that purpose. Copper is used for electrical wiring because copper has a high electrical conductivity; steel is used for car and ship bodies because it is strong and malleable; and aluminium for window frames because it is easily extruded and resists corrosion. Let us then look at common properties of metals.

## FOR INVESTIGATION

Gather information and prepare a table of alloys and their properties and uses. Select alloys that illustrate the wide range of metals used. Include four alloys that are predominantly one metal (with small amounts of other elements) and at least six alloys that are mixtures of comparable amounts of two or more metals. Your table should include twelve to fifteen alloys in total. Include duralumin, alnico, nichrome and at least one titanium alloy in addition to the older alloys of brass, bronze, pewter and steels. In addition to the headings of Table 4.3 include the headings, *Why it was developed* and *When it was first used*.



## 4.2 COMMON PROPERTIES OF METALS

In Section 1.23 we defined metals as solid elements that have a lustrous appearance and good electrical and thermal conductivity, and are malleable and ductile. However, metals display a wide range of properties. Some, such as gallium and potassium, melt at quite low temperatures ( $30^{\circ}\text{C}$  and  $63^{\circ}\text{C}$  respectively) while others have very high melting points (tungsten  $3410^{\circ}\text{C}$  and molybdenum  $2610^{\circ}\text{C}$ ). Some have low densities (lithium  $0.53\text{ g/mL}$  and calcium  $1.55\text{ g/mL}$ ), others have high densities (gold  $19.3\text{ g/mL}$  and uranium  $18.9\text{ g/mL}$ ). Some are good conductors of heat and electricity (silver and copper) while others are more mediocre performers (barium and titanium); some are soft (lead and sodium) while others are hard (chromium and osmium); some have a high tensile strength (iron) while others are quite brittle (tin). Some corrode easily (magnesium and iron), others do not (platinum and gold). All these varying properties affect our choice of metal for a particular job.

Table 4.4 shows common properties of several of the metals in widespread use today. Some of these properties need explanation.

**Electrical conductivity** is the current passing through a 1 metre cube of the substance when a voltage of exactly 1 volt is applied across opposite faces of the cube: it is measured in megohm $^{-1}$  m $^{-1}$ . (An ohm $^{-1}$  is called a siemens, symbol S, so you will sometimes see conductivity given as MS m $^{-1}$ , meaning megasiemens per metre.)

**Thermal conductivity** is the energy transmitted per second through a 1 metre cube of the substance when there is a  $1^{\circ}\text{C}$  temperature difference across opposite faces: it is measured in J s $^{-1}$  m $^{-1}$  K $^{-1}$ .

**Hardness** can be measured in several ways. The Brinell Hardness Number (shown in Table 4.4) is based upon the size of the indentation made by forcing a hard sphere into the surface of the material. Another measure of hardness (more commonly used for minerals) is the Mohs scale: this assigns numbers to materials based on a scale on which talc (extremely soft) is given a value of 1, quartz, 7 and diamond (extremely hard), 10.

**Tensile strength** is a measure of how well a material resists bending, twisting or stretching and indicates the suitability of the material for structural purposes. Tensile strength is measured by stretching a sheet or rod of the material of

**TABLE 4.4 Common properties of some metals**

Metal	Aluminium	Calcium	Copper	Gold	Iron	Lead	Magnesium
melting point (°C)	660	839	1085	1063	1535	327	650
density (g/mL)	2.70	1.55	8.96	19.3	7.86	11.4	1.74
electrical conductivity (megohm <sup>-1</sup> m <sup>-1</sup> )	38	29	58	45	10.3	4.8	23
thermal conductivity (J s <sup>-1</sup> m <sup>-1</sup> K <sup>-1</sup> )	237	200	400	317	80	37	156
hardness (Brinell scale)	17	17	50	33	40	4	65
tensile strength (MPa)	50	55	210	130	340	15	160
resistance to corrosion	high	low	high	very high	moderate	high	low
Metal	Platinum	Silver	Sodium	Tin	Titanium	Tungsten	Zinc
melting point (°C)	1770	962	98	232	1660	3410	419
density (g/mL)	21.4	10.5	0.97	7.30	4.50	19.3	7.14
electrical conductivity (megohm <sup>-1</sup> m <sup>-1</sup> )	9.3	63	21	8.5	2.3	19	16
thermal conductivity (J s <sup>-1</sup> m <sup>-1</sup> K <sup>-1</sup> )	73	429	141	67	22	174	116
hardness (Brinell scale)	38	25	1	50	25	225	38
tensile strength (MPa)	120	160		17	540	120	140
resistance to corrosion	very high	high	low	high	high	high	high

known cross-sectional area until it breaks; the measured force needed to break the sample divided by the original cross-sectional area of the sample is the tensile strength. It has the units newton per square metre ( $N\ m^{-2}$ ) which is a pascal (Pa), the unit for pressure (though here it is a ‘stretch’, not a ‘squeeze’).

*Resistance to corrosion* is assessed for Table 4.4 on a four-category qualitative scale: very high, high, moderate and low. A *practical* basis is used: a ‘high’ resistance to corrosion is given to metals that are fairly unreactive under normal atmospheric conditions *and* to metals such as aluminium and zinc that are very reactive but that become coated with a protective layer so that they do not corrode from a practical viewpoint.

Some properties of relevance to metals are difficult to measure quantitatively. Table 4.5 explains the meanings of the common ones.

## WEBSITES

<http://www.webelements.com>

(presents a great variety of properties for each element; just click on the desired element in the displayed Periodic Table and its properties will appear, including in most cases a photograph of a sample of the element)



<http://www.chemicool.com>

(presents a wide range of properties of elements; click on the required element in the displayed Periodic Table)

**TABLE 4.5 Common mechanical properties of solids**

Property	Meaning	Examples
hard	If a force is applied, does the solid distort? Can it be cut with a knife or can it be scratched?	Diamond, sodium chloride and tungsten are hard; lead, calcium, phosphorus and paraffin wax are soft
brittle	If hit with a hammer on a hard surface, does it shatter?	Diamond and sodium chloride are brittle; tungsten and copper are not
malleable	Able to be beaten or rolled into thin sheets	Iron, aluminium and copper are malleable; sucrose, sulfur and graphite are not
ductile	Able to be drawn into thin wires	Copper, gold, aluminium and iron are ductile; sulfur, phosphorus and sodium chloride are not

Let us now look at how the uses of some metals relate to their properties.

## 4.3 RELATING USES TO PROPERTIES

The metal that is used for a particular purpose is the one that has the best properties for that purpose. Examples will illustrate.

### Iron

Iron is rarely used in pure form. Generally we use iron which contains about 0.2% carbon; this is called *mild steel*. Mild steel is very malleable so it is easily rolled into sheets, which can be stamped into special shapes. Mild steel has fairly good tensile strength, so it is used to make motor car bodies, corrugated sheets for roofing and the outer cases of refrigerators, washing machines and numerous other household appliances.

The higher tensile strength of *structural steel*, 0.3 to 0.6% carbon (which means that it resists bending), lends it to being used for beams and girders in buildings and factories, and as reinforcing in concrete. This steel is relatively easy to machine (cut and shape in a lathe). This, combined with its high tensile strength, makes it very suitable for making machinery and screws, nuts and bolts.

Iron containing 3 or 4% carbon, the primary product from a blast furnace (for extracting iron from haematite), is very hard, but quite brittle and not at all malleable. However, when molten and poured into a suitable mould it makes good castings (takes the shape of the mould very well). So it is used for casting engine blocks, fire hydrants and drainage grates. It is called *cast iron* or *pig iron*.

Another significant factor in choosing iron or steel for these purposes is its relative cheapness. We do not use an expensive metal such as aluminium or titanium if a cheaper one such as steel will do the job just as well.

A disadvantage with using iron or steel is that it rusts (corrodes). This means that we have to protect steel products by covering them with paint or tin or by galvanising them (coating them with zinc).

## Aluminium

Aluminium, particularly when alloyed with small amounts of copper, magnesium or manganese, has low density (light weight), good tensile strength, and high resistance to corrosion. These properties make it very suitable for the following uses.

### Aeroplanes

Aluminium is used for the outer cladding of bodies and wings of aircraft because of the above properties.

### Window and door frames

Aluminium can easily be extruded (forced through a die) into quite complex shapes. This, combined with its strength and corrosion resistance, makes it very suitable for window and door frames.

### Household utensils

Aluminium is malleable and can easily be stamped into shapes, so it has been widely used for household utensils such as cake tins, cooking trays and muffin moulds (though today stainless steel and non-stick, surface-coated, tin plate are more common). In addition its high thermal conductivity makes it good for saucepans and cooking pots (though, again, stainless steel and coated iron are becoming more common).

### Drink cans

Lightness, malleability and ease of stamping into shapes have led to aluminium being the currently favoured material for drink cans, despite a higher cost per kilogram than steel.

### Cylinder heads for motor cars

Low density and high thermal conductivity mean that cylinder heads for car engines made of aluminium (or one of its alloys) perform more efficiently (and reach operating temperature more quickly) than the heavier, less-conducting steel heads of a couple of decades ago.

### High voltage electrical transmission lines

High electrical conductivity combined with low density and reasonable tensile strength mean that aluminium is widely used for high voltage electrical transmission lines; even when a few strands of steel wire are added to give the required tensile strength, the aluminium-steel cable has a lower mass than a copper cable of the same electrical conductivity. Cable mass is particularly important for these transmission lines because it determines the spacing of the expensive towers needed to support the cables. For 240 volt cables around streets and in houses mass is less of a consideration so cheaper copper is generally used.

Aluminium is much more expensive than steel or copper so there have to be real advantages in using aluminium rather than steel to justify the higher

outlay. Typical of the choices that have to be made is the decision whether to use galvanised steel guttering on a house (low cost but susceptible to rusting after a decade or two) or to use aluminium guttering (higher initial cost but much longer lasting).

## Exercises

- \*1 For iron (or steel) and copper, take each of the major uses listed in Table 4.1 and explain why that metal or particular steel is used for each of the listed uses rather than some other metal. Information in Table 4.4 may be helpful.
- 2 For each of the metals listed below, what properties make it particularly suited for the use indicated?
  - a mild steel for fencing wire
  - b tungsten for filaments in electric light bulbs
  - c gold for wedding rings
  - \*d lead for rifle bullets
  - \*e aluminium for small runabout boats (dinghies)
  - \*f liquid sodium as a heat transfer liquid in some nuclear power stations
  - \*g chromium for electroplating (coating) motor car parts.
- 3 a What properties would make a metal suitable for saucepans? From Table 4.4, choose two metals that would make good saucepans.  
b Stainless steel has a thermal conductivity of  $30 \text{ J s}^{-1} \text{ m}^{-1} \text{ K}^{-1}$ . Does this suggest that it is a good material for making saucepans? Stainless steel is a very common material for saucepans; how is this thermal conductivity problem overcome?  
c What are the advantages and disadvantages (if any) of making saucepans out of copper?
- 4 Suggest reasons why aluminium is less used today for cake tins, cooking trays, muffin moulds and saucepans than it was 20 years ago.
- 5 Why is a copper–nickel alloy used for ‘silver’ coins instead of real silver?
- 6 Of the metals in Table 4.4:
  - a which has the lowest melting point and which the highest?
  - b which has the lowest density and which the highest?
  - c which three are the best conductors of electricity?
  - d which two have the highest tensile strength?



## 4.4 METALS THROUGH HISTORY

The first metal to be extracted from an ore was copper. This occurred in the Middle East in about 6000 BCE<sup>†</sup>. Copper oxide was heated with charcoal (mainly carbon) and globules of copper resulted. This copper was used to make ornaments and possibly a few domestic utensils. However, working copper was difficult (its melting point was very high for simple wood fires) and the finished products were fairly soft (pure copper is a soft metal). So copper tools did not replace the stone ones then in common use. Copper was actively mined on Cyprus from about 3000 BCE.

<sup>†</sup> BCE: before the common era; CE common era. BCE used to be called BC and CE was called AD. CE has the same start date as AD.



Bronze Age axe and lance heads

Subsequently it was discovered that heating copper and tin ores with charcoal produced a much harder metal alloy, bronze. Bronze has a lower melting point than copper and so could be melted, moulded and worked much more easily in wood fires. By 2000 BCE bronze was commonly used for tools and weapons throughout Europe and Asia. Nations that first developed the technology to make bronze weapons soon conquered their less-developed neighbours. There were tin mines in Cornwall from about 2000 BCE. The ox-drawn plough was probably invented at about this time. Bronze was the dominant material for tools and weapons from about 2000 BCE until 1000 BCE. Hence this period is called the *Bronze Age*.

Extraction of iron from iron oxide using charcoal requires a higher temperature than does extraction of copper. It was not until about 1200 BCE that humans developed this method of obtaining iron and of forging iron into tools and weapons. It was necessary to blow air into the fire to get a sufficiently high temperature, first to extract the iron and then to soften it enough to forge it. Iron had

been known since about 3000 BCE but only because it had been found in nearly pure form as meteorites.

The technology for extracting and processing iron spread throughout Europe and Asia. Tools and weapons made of iron were harder than bronze ones and so by about 1000 BCE iron had largely replaced bronze for tools and weapons. Hence the period from about 1000 BCE to 1 CE is called the *Iron Age*.

Anthropologists and ancient historians usually ‘end’ the Iron Age at about 1 CE and start talking about the Modern Age or Modern Era after that (see Table 4.6). However, there was no dramatic movement away from iron as the dominant metal in subsequent centuries.

**TABLE 4.6 The anthropological or historical ‘ages’ of human development**

Stone Age	up to about 3000 BCE
Copper Age	3200 to 2300 BCE
Bronze Age	2300 to 1000 BCE
Iron Age	1000 BCE to 1 CE
Modern Age	1 CE to present

## Iron and steel in the Modern Age (after 1 CE)

The technologies for producing iron and steel improved over the centuries and the quantities being used increased markedly, especially after the Industrial Revolution of the eighteenth century. However, there was no significant move towards the use of new alloys or metals until late in the nineteenth century. This is despite metals such as cobalt, nickel, zinc, manganese, tungsten, titanium and chromium being isolated (on the laboratory scale) in the eighteenth century, and sodium, magnesium, cadmium, aluminium and vanadium in the nineteenth century.

Alloy steels began to be made from about the 1880s onwards; tungsten steel (for cutting tools) was first, followed by manganese steels (for railway lines and digging tools) and then many others (using for example silicon, chromium, nickel and vanadium). Stainless steel came into use after about 1920.

## Newer metals

Metals that have come into common use (as pure metal or as the major component of alloys) in the last 100 years include aluminium, tungsten, magnesium and titanium.

### Aluminium

The commercial use of aluminium began towards the end of the nineteenth century and has expanded rapidly ever since. Aluminium has displaced iron and steel for many uses for which it was better suited (Table 4.1). Aluminium is often used alloyed with small amounts of other metals such as copper, manganese, magnesium and titanium.

### Tungsten

Although tungsten had been used in steel alloys from the 1880s, its use as a pure metal for filaments in light bulbs (and later in television picture tubes) did not begin until about 30 years later because of the difficulty of getting the metal into a ductile form (because of its high melting point).

### Magnesium

Magnesium came to prominence during World War II when it was used as the major component of alloys for aircraft bodies and castings for car parts (particularly in the VW 'Beetle' of 1939 to the 1990s). It is still widely used for alloy ('mag') wheels for cars and in alloy casings for mobile phones and lap-top computers, though for most other uses it has been largely replaced by alloys that are predominantly aluminium.

### Titanium

Titanium has the highest strength to weight ratio of any metal; it is very hard and is very resistant to corrosion. Alloyed with small amounts of aluminium and vanadium it is even stronger and harder. Titanium is widely used in the aerospace industry (both for outer cladding and in engines), for medical implants (artificial hips and knees) and for bicycles and other sporting appliances.

#### WEBSITES

<http://www.uow.edu.au/commerce/econ/modbusiness/special%20steels.pdf>

(gives a short history of steel making, particularly of the development of alloy steels)

<http://titaniumera.com/facts.html>

(a good account of the extraction, properties and uses of titanium)

<http://pubs.acs.org/cen/80th/magnesium.html>

(presents useful information about the uses of magnesium; you can get information about other metals also by clicking on their names in the drop-down menu in the 'Element name' box near the top of the page.)



#### Exercises

- 7 Historically, gold and silver were used before copper and bronze, but mainly for ornaments. Why were they used before copper and bronze and why were they not used for tools and weapons?
- 8 In early times pewter was 75% tin and 25% lead. In more recent times it has been 85% tin, 7% copper, 6% bismuth and 2% antimony. The physical properties of both types of pewter are very similar. Suggest two reasons why there was a change in the composition of pewter.



None of the metals in widespread use today (apart from gold and platinum) occur in nature as free metals. Let us then look at how metals are extracted from their naturally occurring ores.

## 4.5 EXTRACTION OF METALS FROM ORES

Before turning to the extraction of metals from naturally occurring substances, let us clarify the meanings of the two terms, mineral and ore.

A **mineral** is a pure (or nearly pure) crystalline compound that occurs in the Earth's crust.

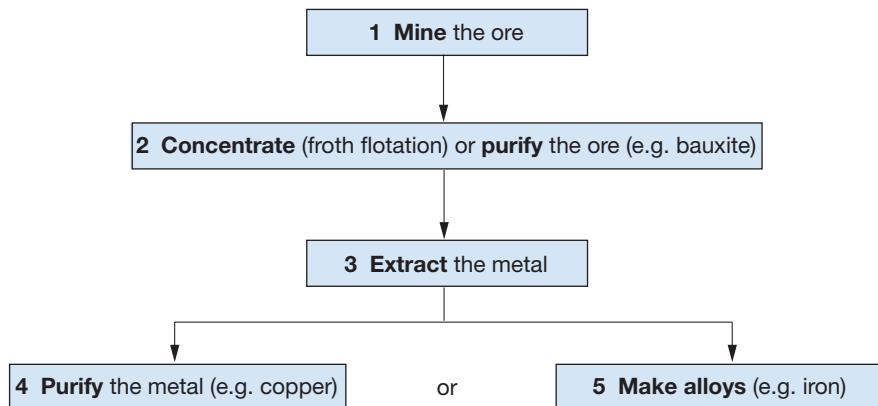
An **ore** is a compound or mixture of compounds from which it is economic (or commercially profitable) to extract a desired substance such as a metal.

Haematite (iron(III) oxide) is a mineral; it is the common ore of iron. The ore for aluminium is bauxite: it is a mixture of compounds, mainly the minerals gibbsite and boehmite (which are hydrates of aluminium oxide), iron(III) oxide, silica and various clays. There are various minerals called alumino-silicates, but these minerals are not ores of aluminium because it is not economic to extract aluminium from them.

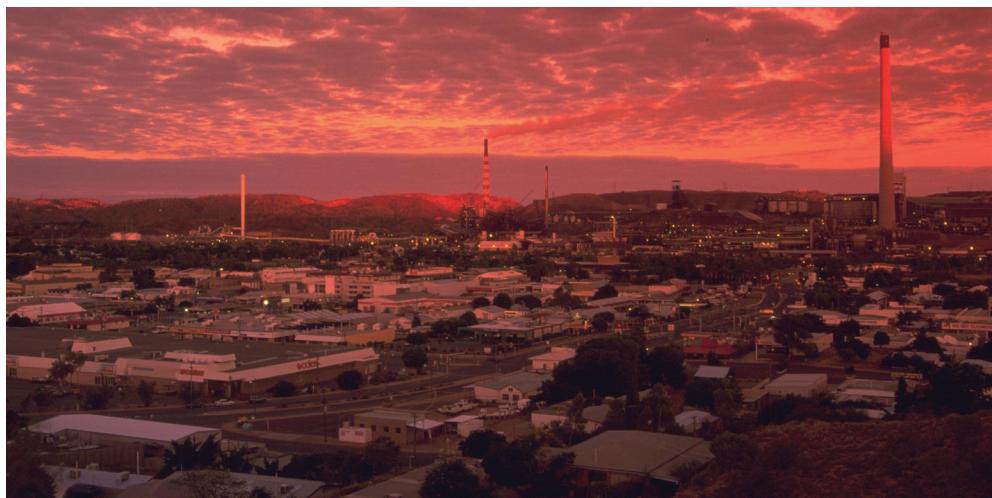
### Extracting metals from ores—the general process

The general procedure for extracting a metal from its ore is shown in Figure 4.1.

FIGURE 4.1  
A flow chart for extracting a metal from its ore



Mount Isa in north-west Queensland is Australia's largest copper producer



Let us see how this generalised procedure applies to copper.

## 4.6 EXTRACTION OF COPPER

Common ores of copper contain the minerals chalcopyrite,  $\text{CuFeS}_2$ , chalcocite,  $\text{Cu}_2\text{S}$ , malachite,  $\text{Cu}(\text{OH})_2 \cdot \text{CuCO}_3$ , or azurite,  $\text{Cu}(\text{OH})_2 \cdot 2\text{CuCO}_3$ , mixed with a variety of other minerals. Typical ores contain only 1 to 5% copper. This means that the ores need to be processed to remove most of the unwanted material before the copper compounds can be reduced to metal.

In Australia the main copper ores are sulfides. The common method for concentrating these ores is **froth flotation**. When air is blown through tiny jets into a detergent solution, lots of bubbles (froth) form. If the solution contains small suspended mineral particles, then the particles that are wetted by the solution will sink to the bottom of the container while particles that are not wetted by the solution will be drawn inside the bubbles and so float to the surface; the froth that is skimmed off the surface is much richer in the non-wettable particles than was the starting material (Fig. 4.2). By adding certain oils to a mixture of finely crushed ore, sulfide particles which adsorb the oil can be made non-wettable by the detergent solution, while other particles remain wettable, and so sulfide particles can be separated out.

With sulfide ores, this froth flotation method produces a copper concentrate which is about 25 to 30% copper: typically it is chalcopyrite with other metal sulfides as impurities.

For Step 3 of Figure 4.1, the copper concentrate is heated with sand (silica,  $\text{SiO}_2$ ). This produces two immiscible liquids, one containing mainly copper(I) sulfide, the other containing most of the unwanted iron as iron silicate,  $\text{FeSiO}_3$ , which is discarded. The remaining iron is removed by re-heating the retained liquid with more sand, but now with air being bubbled through it.

The copper(I) sulfide liquid is then heated on its own to a higher temperature while air is bubbled through it. This reduces the sulfide to copper metal. Sulfur dioxide is also produced.

While most of the sulfur dioxide escapes quickly, some does not leave the cooling copper until it is almost solid. This gives the solid a blistery appearance, which is why the product of this smelting process is called *blister copper*. It is about 98% pure.

The term **smelting** is used for the extraction of a metal by heating substances to sufficiently high temperatures to produce a molten material from which the metal can be obtained.

In the present example we say that the copper concentrate is smelted with sand (silica,  $\text{SiO}_2$ ) to produce copper.

### FOR INVESTIGATION (non-syllabus)

What is (are) the main source(s) of pollution in the extraction of copper from its ore? How are these managed at (a) Mount Isa (a remote location) and (b) Port Kembla (a heavily populated area)?

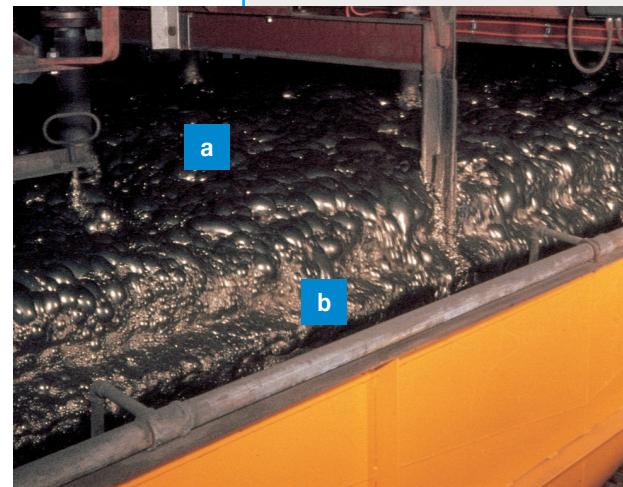


FIGURE 4.2  
Concentrating ores by froth flotation. Bubbles of black froth (a) bring the required ore to the top of the separation tank and overflow into the collector trough (b)



For most purposes (particularly for electricity conduction) copper of at least 99.9% purity is required. Blister copper is generally purified electrolytically. This will be described in *Conquering Chemistry HSC Course*, Section 2.11.

Let us now look at the chemical equations for the processes just described.

## 4.7 CHEMICAL EQUATIONS FOR THE EXTRACTION OF COPPER

The steps for writing a chemical equation were described in Section 3.6. They are summarised in Box 4.1 below and are illustrated by writing the equations for the extraction of copper from chalcopyrite.

### BOX 4.1 TO WRITE A CHEMICAL EQUATION

- 1 Write a word equation that includes all reactants and products.
- 2 Write the formulae for all substances involved (to make an unbalanced equation).
- 3 Balance the equation by putting numbers in front of symbols or formulae.  
Start with elements that occur in one place only on both sides of the equation.
- 4 Add symbols for the physical states, (s), (l), (g), (aq).

Step 1 may be omitted when you have gained more experience.



### Example 1

Write a balanced chemical equation for the reaction of copper(I) sulfide with oxygen.

- 1 copper(I) sulfide + oxygen → copper + sulfur dioxide
  - 2  $\text{Cu}_2\text{S} + \text{O}_2 \rightarrow \text{Cu} + \text{SO}_2$
  - 3  $\text{Cu}_2\text{S} + \text{O}_2 \rightarrow 2\text{Cu} + \text{SO}_2$
- Equation is now balanced: 2 Cu atoms, 1 S and 2 O on each side.
- 4  $\text{Cu}_2\text{S(l)} + \text{O}_2\text{(g)} \rightarrow 2\text{Cu(l)} + \text{SO}_2\text{(g)}$  ... (4.1)

This equation was easy to balance. Let us consider a more difficult one, namely the first step of the extraction of copper from chalcopyrite.



### Example 2

Write an equation for the reaction between chalcopyrite, silica (sand) and oxygen to form copper(I) sulfide, iron(II) silicate and sulfur dioxide.

Following the steps in Box 4.1:

- 1 Chalcopyrite + silica + oxygen → copper(I) sulfide + iron(II) silicate + sulfur dioxide
- 2  $\text{CuFeS}_2 + \text{SiO}_2 + \text{O}_2 \rightarrow \text{Cu}_2\text{S} + \text{FeSiO}_3 + \text{SO}_2$
- 3 a Balance by starting with Cu, Fe and Si. Balance Cu by putting a 2 in front of  $\text{CuFeS}_2$ . Now balance Fe by putting a 2 in front of  $\text{FeSiO}_3$  and balance Si by writing  $2\text{SiO}_2$ :  
$$2\text{CuFeS}_2 + 2\text{SiO}_2 + \text{O}_2 \rightarrow \text{Cu}_2\text{S} + 2\text{FeSiO}_3 + \text{SO}_2$$
- b Now look at S and O, the elements that occur in more than one place on one or both sides of the equation. Balance S by writing  $3\text{SO}_2$ :



For O we have  $2 \times 2 + 2 = 6$  atoms on the left and  $2 \times 3 + 3 \times 2 = 12$  on the right. Hence write  $4\text{O}_2$  on the left (then  $2 \times 2 + 4 \times 2 = 12$  as required):



This is the required balanced equation. The numbers in front of the formulae or symbols are called the **coefficients** of the equation. Box 4.2 summarises the steps.

#### BOX 4.2 TO BALANCE A CHEMICAL EQUATION

- 1 Start from the left and balance each element that occurs in only one place on each side of the equation (Cu, Fe and Si in this example).
- 2 Balance the remaining element(s); this may require further adjusting of the coefficients of the once-a-side elements.
- 3 Remember that you balance an equation by putting numbers (called coefficients) in front of formulae: you must not alter the formulae. If you are having difficulty balancing, check that you have included *all* of the reactants and products.

After this diversion into balancing equations, perhaps we should summarise the process for extracting copper.

*To extract copper from chalcopyrite ores:*

- 1 Concentrate the ore by froth flotation (from 2% Cu to 25%).
- 2 Heat the ore with sand and air to remove iron as a silicate slag, leaving  $\text{Cu}_2\text{S}$ .
- 3 Heat  $\text{Cu}_2\text{S}$  with air to form blister copper (98% pure).
- 4 Purify the copper (to 99.95%) by electrolysis.

Another consideration in extracting metals from ores is energy.

## 4.8 ENERGY CONSIDERATIONS

For nearly all metals a chemical reaction is used to extract the metal from its ore. As we saw in Sections 3.1 to 3.4, chemical reactions are accompanied by either release or absorption of heat, a form of energy.

The chemical reactions involved in extracting some metals from their ores absorb heat, while for other metals the extraction reactions liberate heat. For example:

- Extracting copper from sulfide ores liberates heat.
- Extracting aluminium from alumina absorbs heat (in quite large amounts).
- Extracting iron from iron(III) oxide with carbon in a blast furnace liberates heat.

However, the energy involved in the extraction reaction is only one part of the total energy budget of the extraction process. Energy has to be supplied:

- to mine the ore
- to purify or concentrate the ore
- to maintain the high temperatures needed to make the extraction reactions go
- to purify the raw metal or to form it into useful alloys (such as steel).

Overall a large amount of energy is needed to extract a metal from its ore. This is true even for copper where the extraction reaction actually releases energy. Table 4.7 shows the amount of energy required for three common metals. The amount of energy required to produce a metal is a factor in deciding which metal to use for a particular purpose.

**TABLE 4.7 Energy required for the production of 1 kilogram of certain metals**

Metal	Energy required (in megajoules)	
	from natural ores	from recycled material
aluminium	200	7
copper	70	4
mild steel	40	8

## Exercises

- 9 a** In the Copper and Bronze Ages copper was extracted from copper oxide by heating it with charcoal (carbon), liquid copper and carbon monoxide being the products. Write a balanced chemical equation for the reaction.
- b** In the Bronze Age tin was extracted from tin(IV) oxide in the same way. Write a balanced chemical equation for this reaction.
- c** In the Iron Age iron was extracted from haematite, iron(III) oxide, by reacting it with carbon; liquid iron and carbon monoxide were the products. Write an equation.
- d** Nowadays the main reaction for extracting iron from haematite in a blast furnace is the reaction with carbon monoxide (to form liquid iron and carbon dioxide). Write a balanced equation for this reaction.
- 10** Equations 4.2 and 4.1 show the conversion of chalcopyrite to copper(I) sulfide, then to copper. Add these equations up (just as you add algebraic equations) to get an overall equation for the conversion of chalcopyrite to copper.
- 11** Two non-sulfide ores of copper are malachite,  $\text{Cu}(\text{OH})_2 \cdot \text{CuCO}_3$ , and azurite,  $\text{Cu}(\text{OH})_2 \cdot 2\text{CuCO}_3$ . Copper is extracted from these ores by first heating them to decompose the hydroxide and carbonate into oxide (the other products being water and carbon dioxide respectively). The oxide is then heated with coke (carbon) to form copper. Write chemical equations for these reactions.
- 12** Titanium is extracted from rutile ( $\text{TiO}_2$ ) first by reacting it with chlorine gas and carbon at high temperature to form gaseous titanium(IV) chloride and carbon monoxide. The titanium(IV) chloride is then reacted with liquid magnesium, again at high temperature, to form titanium and magnesium chloride, both as solids. Write equations for these two reactions.



### WEBSITE

[http://www.minerals.org.au/education/secondary/secondary\\_resources/factsheets](http://www.minerals.org.au/education/secondary/secondary_resources/factsheets)

(presents information about mining, processing and the use of metals in Australia; click on the metal you want)

One of the properties listed in Table 4.4 was ‘resistance to corrosion’. It is a serious factor in selecting a metal for a particular job. Corrosion is related to chemical reactivity, so let us take a closer look at the reactivity of metals.

## 4.9 REACTIVITY OF METALS

Metals display a wide range of reactivity with many other substances, varying from extremely reactive to no reaction at all. The ‘other substances’ that most influence the choice of metal for a particular purpose are oxygen (from air), water and acids. In discussing the reactions with these three substances we shall restrict ourselves to the following metals:

Li, Na, K (from Group 1 of the Periodic Table), Mg, Ca, Ba (from Group 2), Al, Sn, Pb, Fe, Cu, Zn, Ag, Au, Pt.

### Reaction with oxygen

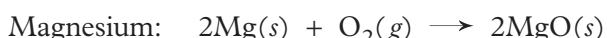
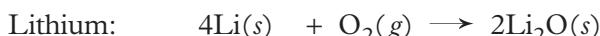
All metals except silver, platinum and gold react with oxygen to form oxides:

- Li, Na, K, Ca, Ba react rapidly at room temperature.
- Mg, Al, Fe, Zn react slowly at room temperature but burn vigorously if heated in air or pure oxygen.
- Sn, Pb, Cu react slowly and then only if heated.

All the oxides formed are ionic compounds (Section 2.15).

Those metals which burn in air or oxygen form crystalline white solids that have none of the physical properties of the original metal (lustre, strength, malleability, conductivity). When metals slowly react at room temperature, they lose their shiny lustrous appearance. Some, such as Al and Zn, become coated with a dull layer of tightly adhering oxide, which prevents further reaction. Others such as iron form a powdery surface layer of oxide which does not impede further reaction. When copper is heated in air it forms a black surface layer of copper oxide.

Chemical equations for some of these reactions are as follows.



### Reaction with water

Some metals react with water or steam while others do not:

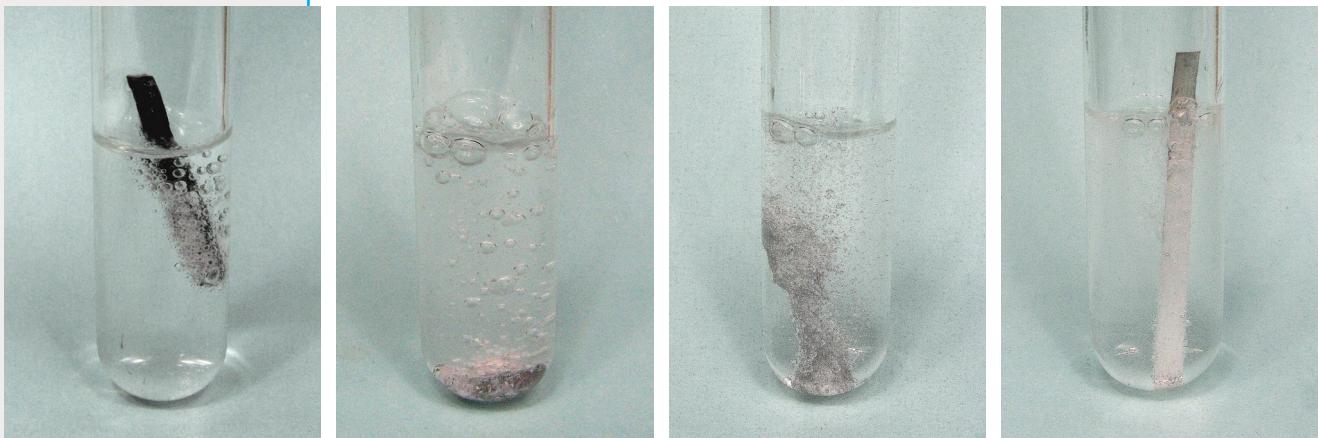
- Li, Na, K, Ca, Ba react with water at room temperature.
- Mg, Al, Zn, Fe react with steam at elevated temperatures.
- Sn, Pb, Cu, Ag, Au, Pt do not react at all.

When reaction occurs with water the products are hydrogen gas and the metal hydroxide. The chemical equations for sodium and calcium are:



When pieces of lithium, calcium and barium are dropped into water, bubbles of colourless gas form: with lithium and barium a colourless solution of lithium or barium hydroxide results, but with calcium a suspension of insoluble white

calcium hydroxide forms. Sodium and potassium react very vigorously (often with an explosion as the hydrogen ignites in air).†



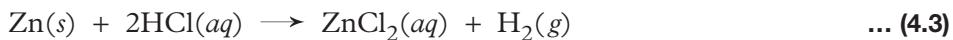
Reaction of (left to right) lithium with water, calcium with water, zinc with sulfuric acid, magnesium with hydrochloric acid

When a piece of freshly cleaned magnesium ribbon is held in the mouth of a flask of vigorously boiling water, a white deposit (of magnesium oxide) forms on the ribbon as the magnesium reacts with steam. Steam needs to be heated significantly above 100°C for the reaction with aluminium, iron and zinc to occur. With steam the product is oxide, not hydroxide.

## Reaction with dilute acid

All metals (on our list) except copper, silver, gold and platinum react with dilute hydrochloric and sulfuric acids to form hydrogen gas, though for tin and lead the reactions are quite slow unless the acid solutions are heated.

Zinc, for example, reacts with dilute hydrochloric acid to form bubbles of colourless hydrogen gas and a clear solution of zinc chloride. We can write the equation for this reaction as:



From previous studies you will recall that:

acids are substances which in solution produce hydrogen ions, H<sup>+</sup>.

It is actually the hydrogen ion from hydrochloric or sulfuric acid which reacts with the metal. Recognising this, we can write an alternative equation:



We call this a **net ionic equation**. It shows the actual ionic species that undergo change in the reaction. Equations such as 4.3 which show the actual chemical substances used in the reaction are called **neutral species equations**.‡

We can write a third type of equation for such reactions by recognising that hydrochloric acid, HCl(aq), is actually H<sup>+</sup> + Cl<sup>-</sup>, and ZnCl<sub>2</sub>(aq) is Zn<sup>2+</sup> + 2Cl<sup>-</sup>:



† The reaction between potassium and water should *not* be performed in a school laboratory. The reaction between sodium and water should be performed *only* by an experienced teacher, and the piece of sodium used should not be much bigger than a grain of rice.

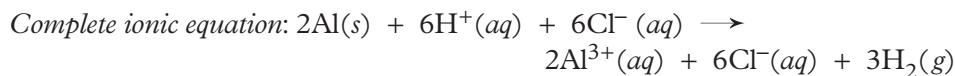
‡ The Board of Studies syllabus document refers to these as *balanced formulae equations*.

This is called a **complete ionic equation** in that it shows all the ions involved in the solutions used for the reaction. Notice that the chloride ions do not undergo any chemical change: there are two  $\text{Cl}^-$ s on each side of the equation. Ions that do not undergo any change during a reaction are called **spectator ions**. In the reaction between zinc and hydrochloric acid, chloride ions are spectator ions. It is the spectator ions that we omit when we write a net ionic equation.

When zinc reacts with dilute sulfuric acid, hydrogen gas is formed. The net ionic equation is still Equation 4.4. The neutral species and complete ionic equations are:



The reactions of the other metals are completely similar to those of zinc. For example for aluminium with hydrochloric acid we can write:



## A common feature

A common feature of all the reactions of metals with oxygen, water and dilute acids is that *atoms of the metals lose electrons to become positive ions*.

- With oxygen, ionic oxides are formed. For example magnesium formed magnesium oxide,  $\text{MgO}$ ; this is an ionic compound containing the ions  $\text{Mg}^{2+}$  and  $\text{O}^{2-}$ .
- With water, ionic hydroxides, either solutions or solids, are formed. For example, lithium hydroxide in solution is  $\text{Li}^+$  and  $\text{OH}^-$ , while solid calcium hydroxide contains  $\text{Ca}^{2+}$  and  $\text{OH}^-$ .
- With dilute hydrochloric and sulfuric acids, solutions of metallic chlorides and sulfates are formed. For example, magnesium with hydrochloric acid forms magnesium chloride containing  $\text{Mg}^{2+}$  and  $\text{Cl}^-$ , while iron with sulfuric acid forms iron(II) sulfate,  $\text{Fe}^{2+}$  and  $\text{SO}_4^{2-}$ .

If the metal atoms have lost electrons, then some other species (such as  $\text{O}_2$ ,  $\text{H}_2\text{O}$  or  $\text{H}^+$ ) must have gained electrons. Such reactions are called *electron-transfer reactions*. Chemists use the terms *oxidation* and *reduction* in connection with such reactions.

## 4.10 OXIDATION AND REDUCTION AND HALF REACTIONS<sup>†</sup>

When an atom loses one or more electrons we say that it has been *oxidised*. If an atom gains one or more electrons we say that it has been *reduced*. In other words:

**Oxidation** means loss of electrons

**Reduction** means gain of electrons

<sup>†</sup> Oxidation and reduction are not in the Preliminary Course syllabus, though a discussion of electron-transfer reactions and half equations is an ideal place to introduce these terms.

In normal chemical reactions there can be no overall loss or gain of electrons. Hence oxidation and reduction occur simultaneously in complete chemical reactions. We call such reactions **redox reactions**, that is **reduction–oxidation reactions**. Another name for redox reactions is **electron-transfer reactions**.

We have already seen that magnesium metal reacts with oxygen gas to form white solid magnesium oxide:



Mg loses two electrons to become  $\text{Mg}^{2+}$  (in ionic  $\text{MgO}$ ) so Mg is *oxidised* to  $\text{Mg}^{2+}$ . Oxygen gains two electrons to become oxide  $\text{O}^{2-}$ , so oxygen is *reduced* to oxide. This is therefore a redox reaction.

The reaction of a metal with dilute acid to form hydrogen gas is an example of a redox reaction. As discussed above zinc reacts with hydrochloric acid, the net ionic equation being:



Zn has been oxidised to  $\text{Zn}^{2+}$  while  $\text{H}^+$  has been reduced to  $\text{H}_2$ .

Frequently, to identify oxidation and reduction chemists write what are called *half reactions* or *half equations*.

## Half reactions (or half equations)

**Half reactions** (or **half equations**) are reactions that describe the oxidation and reduction processes separately in terms of electrons lost or gained.

For the reaction of zinc with dilute acid, the overall reaction is Equation 4.4. It is made up of the half reactions:



Equation 4.6 is an oxidation half reaction while Equation 4.7 is a reduction half reaction. The overall Reaction 4.4 is made up of an oxidation half reaction and a reduction half reaction.

In combining half reactions into complete reactions, it is necessary to balance the number of electrons. This is because there can be no electrons left over on either side of a complete reaction. For the reaction of aluminium with dilute acid the half reactions are:



To combine these into an overall equation we multiply Equation 4.8 by 2 and Equation 4.9 by 3 (to give 6 electrons in each) then add to get:



## Exercises

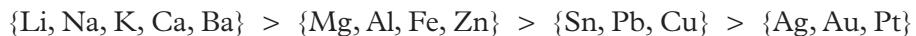
- 13 Write balanced chemical equations for the reaction of **(a)** calcium **(b)** iron, with oxygen. Iron forms iron(III) oxide.
- 14 Write balanced chemical equations for the reactions of **(a)** lithium with cold water **(b)** barium with cold water and **(c)** magnesium with steam.

- 15** For the reaction of **(a)** iron with dilute sulfuric acid solution and **(\*b)** aluminium with dilute hydrochloric acid solution, write **(i)** neutral species equations, **(ii)** complete ionic equations and **(iii)** net ionic equations. Identify any spectator ions in these reactions. What (if any) extra information does the neutral species equation tell us that the net ionic equation does not?
- 16** **a** In each of the chemical reactions in Exercises 13, 14 and 15, which species (if any) has been oxidised? Explain why you think this is so.
- b** For each of the reactions in Exercise 15, where applicable, write half reactions (half equations) showing the electron transfers that are occurring.

## 4.11 AN ACTIVITY SERIES

We can use the reactivity of metals with oxygen, water and dilute acids as described in Section 4.9 to draw up a list of the metals in order of decreasing reactivity. We call such a list an **activity series**.

Based upon reaction with oxygen, we conclude that the order of reactivity is:



Metals within curly brackets {} have equal activity.

Based upon reaction with water, we conclude that the order of reactivity is:



The order, based upon reaction with dilute acids, is:



(We omit consideration of Li, Na, K, Ca, Ba from the acid reaction because of their reaction with the water in dilute acid solutions.)

Combining these three orders of reactivity we get:



This is as far as we can go based upon reactions with oxygen, water and dilute acids. To separate the metals with equal reactivities on this list we need to use either displacement reactions or voltage measurements from galvanic cells (devices which generate electricity from chemical reactions): these will be discussed in Chapter 2 of *Conquering Chemistry HSC Course*. Use of such reactions or measurements *to separate metals rated equally in the above list* gives:

**K Na Li Ba Ca Mg Al Zn Fe Sn Pb Cu Ag Pt Au ... (4.9)<sup>†</sup>**

This Sequence 4.9 is an *activity series* for the common metals.

As explained above, all the reactions under consideration here involve the metal atoms losing electrons to become metal ions. Our activity series then lists the metals in order of *decreasing ease of losing electrons*: metals to the left lose electrons more easily than metals to the right.

Also as explained above, loss of electrons means oxidation, so our activity series lists the metals in order of *decreasing ease of oxidation*: K, Na, Li, Ba ... are more easily oxidised than Cu, Ag, Pt, Au.

<sup>†</sup> This is slightly different from the order we would get if we based it *solely* upon measurements from galvanic cells. We shall use Sequence 4.9 in this chapter because it better relates to the practical reactivity of the metals.

FIGURE 4.3  
Reactions of common metals

Conversely the activity series list metals in order of *increasing ease of reduction of the metal ions*:  $\text{Ag}^+$  and  $\text{Cu}^{2+}$  ions are easily reduced while  $\text{Zn}^{2+}$ ,  $\text{Al}^{3+}$  and  $\text{Na}^+$  are difficult to reduce.

The reactions of metals with oxygen, water and dilute acids are summarised in terms of the activity series in Figure 4.3.

<b>Reaction</b>	K	Na	Li	Ba	Ca	Mg	Al	Zn	Fe	Sn	Pb	Cu	Ag	Pt	Au
with cold water					yes										
with steam						yes									
burn in $\text{O}_2$						yes									
react slowly with $\text{O}_2$							yes								
with dilute acids								yes							
															ease of oxidation of metals decreases
															ease of reduction of metal ions increases

## 4.12 RELATIVE REACTIVITY, PERIODIC TABLE, AND FIRST IONISATION ENERGY

An inspection of our activity series shows some relationships to the Periodic Table.

The activity series shows that Group 1 metals are the most reactive followed by Group 2 metals. Group 3 (Al) comes next in reactivity followed by some transition metals (Zn, Fe), then the metals of Group 4 (Sn, Pb). At the end of the series are more transition metals (Cu, Ag, Pt, Au).

The activity series also shows that in Groups 1 and 2 reactivity increases from top to bottom (Li to K, Mg to Ba).

The relative reactivity of metals also correlates well with an important physical property called *first ionisation energy*.

### First ionisation energy and relative reactivity

The **first ionisation energy** of an element is the energy required to remove an electron from a gaseous atom of the element. It is the energy change for the process:



where M is any element.

While ionisation energy can be measured in joules per atom (where a joule is the unit of energy), it is more commonly reported in *kilojoules per mole* of atoms (kJ/mol) where a **mole** is a particular number of atoms ( $6.02 \times 10^{23}$ ). The meaning and significance of the term mole will be explained in the next chapter.

The first ionisation energy measures the ease of removing an electron from a metal atom: the lower the ionisation energy, the easier it is to remove an electron.

As Exercise 19 below will demonstrate:

the reactivity of metals increases as their ionisation energy decreases.

The reactivity of metals is a significant factor in choosing a metal for a particular job. In addition, the ease of extracting metals from their ores and the historical use of metals are also related to the reactivity of the metals.

## 4.13 CHOICE OF METAL BASED UPON REACTIVITY

Some situations where the choice of metal is based heavily upon chemical reactivity (particularly with oxygen, water and dilute acids) are:

- roof guttering for houses—non-reactive but expensive aluminium or cheaper galvanised iron (which eventually corrodes)?
- water pipes—expensive but non-reactive copper or cheaper corrodible iron?
- electrical contacts for replaceable circuit boards in computers and other electronic equipment—cheap copper (which slowly forms a non-conducting oxide layer) or expensive gold (which does not react with oxygen)?
- body implants—expensive but extremely inert titanium alloys or less expensive, but over the long term corrosion-susceptible stainless steel?

### Exercises

\*17 A particular factory has to pump dilute sulfuric acid from one place to another. Of the common metals available (Fe, Al, Cu, Zn, Pb, Sn, Mg, Ag), which would you use for the pipes for moving this sulfuric acid? Give your reasons. Also give your reasons for not using other metals.



- 18 a Which of the metals Al, Ca, Cu, Fe, Sn gives up electrons (is oxidised) most easily and which is the most difficult to oxidise? Give chemical reactions to support your choices.  
b Which of the metal ions  $\text{Ag}^+$ ,  $\text{Mg}^{2+}$ ,  $\text{Pb}^{2+}$ ,  $\text{Zn}^{2+}$  gains electrons (is reduced) most easily and which is the most difficult to reduce?

\*19 Investigate whether there is any correlation between reactivity (as measured by position in the activity series (4.9)) and first ionisation energy. Do this by plotting ‘rank’ in the activity series (e.g. K is 1, Al is 7, Au is 15) against first ionisation energy from the table below. Can you draw a trend line through the points on your graph? What does this tell you?

First ionisation energy (in kJ/mol) (elements listed alphabetically by symbol)

Ag	737	Ca	596	Li	526	Rb	409
Al	584	Cs	382	Mg	744	Sn	715
Au	896	Cu	752	Na	502	Sr	556
Ba	509	Fe	766	Pb	722	Zn	913
Be	905	K	425	Pt	870		

- \*20a For elements in Group 1 of the Periodic Table, plot first ionisation energy versus period number. Use data in Exercise 19. Draw a straight line or smooth curve (whichever is appropriate) through the points. How does first ionisation energy change with period number?  
b On the same graph as in (a) (but using a different symbol for the points) draw a similar plot for Group 2 elements. Is there a similar dependence of first ionisation energy upon period number? By comparing the two plots what can you say about first ionisation energies for Group 2 elements relative to those of Group 1 elements?

## 4.14 REACTIVITY OF METALS AND EASE OF EXTRACTION

Figure 4.3 shows that as we move from left to right across the activity series, metal ions become easier to reduce to metal atoms. Because metals are present in their ores as ions, this means that:

the further to the right in the activity series a metal is, the more easily it can be extracted from its ores.

Copper (as we saw in Section 4.6) and silver can be extracted from sulfide ores simply by roasting the ore in air, for example:

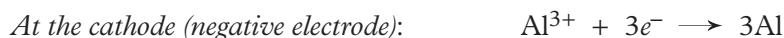


When the sulfides of more reactive metals such as lead, tin and zinc are heated in air, the oxide is formed. This can be reduced by heating with carbon:



Iron which occurs as oxide is obtained industrially from the ore by heating with carbon monoxide, though carbon itself will reduce iron(III) oxide.

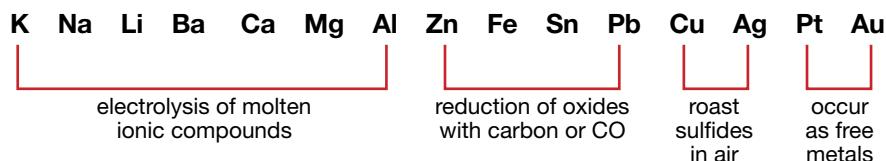
For the most reactive metals, potassium to aluminium in our activity series, there is no simple chemical reagent which will reduce the metal ions to atoms. So electrolysis of molten ionic compounds must be used. For example, aluminium is extracted from aluminium oxide in a molten mixture by electrolysis:



Sodium and magnesium are similarly obtained by electrolysis of molten chlorides.

Figure 4.4 summarises the methods of extraction of metals from their ores.

FIGURE 4.4  
Methods of extracting metals from their ores



## 4.15 CURRENT VERSUS HISTORICAL AVAILABILITY OF METALS

Today there are many more metals available for people to use than there were 200 years ago. Before 1800 only ten metals were in use: copper, tin, iron, lead, gold, silver, zinc, mercury, bismuth and platinum. These were metals that either occurred in the Earth as uncombined elements (Au, Ag, Pt) or could be extracted by heating with carbon. Today in addition to these, there are at least another dozen in widespread industrial use (magnesium, aluminium, titanium, tungsten, manganese, cobalt, nickel, chromium, vanadium, molybdenum, cadmium and uranium), while there are specialised uses for about another twenty or so (such as rhodium, rhenium, palladium, iridium, osmium, zirconium, beryllium, calcium, strontium, sodium, rubidium and gallium).

The availability of metals for commercial and industrial use depends upon one or both of the following factors:

- abundance of the metal in the Earth's crust
- ease of extracting the metal from its naturally occurring ore.

The industrial need for a metal with properties not possessed by any metal in common use was often the incentive for research into developing commercial (as opposed to laboratory) extraction processes. Titanium for jet engines and military aircraft as discussed below is a good example of this. In addition, the development over the last hundred years or so of metallurgical skills for making and testing new alloys led to the incorporation of a wider range of metals into everyday products.

Abundance (or scarcity) in the Earth's crust was a major factor in delaying the use of metals such as palladium, rhodium, rhenium and osmium. However, ease or difficulty of extraction has had a far greater influence over which metals have come into common use, as a brief look at some history will show.

## Historical use of metals

Historically the sequence in which metals started to be used by humans closely parallels the ease of extracting the metals from their ores. As we saw in Section 4.4, copper was the first metal to be extracted from an ore. Bronze was slightly more difficult to prepare because it involved extracting two metals (copper and tin) from their ores with tin being a bit more difficult to extract from its ore than copper. Hence the Bronze Age came after the Copper Age. Iron is somewhat harder to extract from its ores than copper or tin, so the Iron Age followed the Bronze Age.

Lead, being harder to extract than bronze but easier than iron, came into common use during the Roman Empire (during the Iron Age), but being soft, lead was not of much use for tools and weapons. However, it was used for water pipes. As a result of this use, lead poisoning was possibly a factor in the decline and fall of the Roman Empire, because lead poisoning tends to cause a loss of intellectual capacity.

The last century or so has provided striking examples of the use of a metal being related to its ease of extraction.

## Iron and aluminium in recent times

Today aluminium is widely used in a great variety of products from motor cars, ships, aeroplanes, building materials, household goods, packaging materials and electrical cables. Until 1886 there was no aluminium in use at all, despite it being the most abundant metal in the Earth's crust. If a metal was required, it generally had to be iron or steel. There were steel pots and pans; machinery and steam engines were made solely of steel, as were farm tools and weapons. Water and drainage pipes were generally made of iron, and the only metal used for building or structural purposes was steel. Copper, lead and the alloys brass and bronze were available and had their specialised uses. However, iron was by far the dominant metal.

The reason why there was no aluminium in use was that it was extremely difficult to extract aluminium metal from its ore, aluminium oxide. Until 1886 aluminium was a chemist's curiosity: it could be obtained in small quantities with great difficulty and at great expense, but it was far too costly to be used for any practical purposes.

Metals such as magnesium and sodium had been obtained from the naturally occurring chlorides by electrolysis of the molten salts. It was thought that aluminium should be obtainable in a similar way from a molten ionic compound. However, aluminium chloride is not a simple ionic compound and so cannot be electrolysed; the only suitable compound, aluminium oxide, has too high a melting point (over 2000°C) for electrolysis to be possible. Small amounts of aluminium had been prepared in the laboratory by reacting metallic sodium with aluminium chloride.

In 1886 a Frenchman, Paul Heroult, and an American, Charles Hall, simultaneously and independently discovered that if aluminium oxide was mixed with cryolite ( $\text{Na}_3\text{AlF}_6$ ) the mixture melted at about 1000°C. At this temperature it was possible to carry out electrolysis and so an industrial process was developed. This produced aluminium metal in commercial quantities but at prices that were very much higher than for iron. Aluminium started to replace steel in some applications where the properties of aluminium made it the preferred metal, even with its high price.

The major cost in extracting aluminium is the electricity used for the electrolysis and for keeping the electrolyte molten. During most of the twentieth century the price of electricity, particularly for aluminium smelting, fell steadily. This, with refinements to the electrolytic process, improved the competitive position of aluminium relative to steel, so its use increased quite dramatically. Today iron (or steel) is still the most widely used metal, but aluminium has now rapidly grown into second place.

## Titanium

Titanium is the tenth most abundant element in the Earth's crust (or seventh most abundant metal), but it came into practical use only after the 1940s. This is because it was extremely difficult to extract from its ores. The commercial process for its extraction (Exercise 12) was not developed until 1946, and even then it was very expensive. The manufacture of high speed military aircraft and jet engines required alloys that were light, strong, capable of withstanding high temperatures and very corrosion-resistant; cost was not a major factor. Titanium alloys fitted these criteria, and so use of the metal increased rapidly. Increasing use led to decreases in production costs, which in turn resulted in the use of titanium in many other industries where some or all of the above properties were needed. More recently, titanium alloys have been used in human implants such as hips and knees.

These two examples demonstrate that it is not just the physical and chemical properties of a metal that determine what metal will be used for a particular purpose, but also the ease of extraction, availability of the metals and their prices.

Let us end this survey of metals in our society with some economic considerations.

## 4.16 ECONOMIC ASPECTS OF METAL USE

In the early part of this chapter it was implied that when we wanted a metal for a particular purpose, we chose the one with the most suitable physical and chemical properties. However, there are economic considerations as well. We need to consider the price of the metal, and decide whether or not a metal with slightly less suitable properties but with a much lower price would be a better choice overall.

## Factors affecting the price of metals

Factors which affect the price of metals are:

- the abundance and location of ores of the metal (less abundant ores will generally attract higher royalties and so will be more expensive)
- the cost of extracting the metal from the ore (aluminium and titanium are much more expensive to extract than iron or copper)
- the cost of transporting the metal or its ores to the required location (rare metals or their ores may need to be shipped from remote locations, while for abundant metals conveniently located ore deposits can be used)
- the world-wide demand for the metal; if demand is high, the price rises; if it slumps, price falls: the supply and demand factor.

These factors affect the cost of one metal relative to another. General economic conditions such as booming or stagnating economies also affect metal prices, but they generally have a similar effect on all metals. Metal prices have declined somewhat in recent years reflecting decreased economic demand. Table 4.8 gives the prices of some common metals in February 2004.

**TABLE 4.8 Prices (in Australian dollars per tonne) of common metals in February 2004**

Metal	Price	Metal	Price
aluminium	2150	nickel	19600
copper	3250	silver	294 000
gold	17 000 000	tin	10800
lead	1070	zinc	1360
mild steel	350		

Another consideration in choosing a metal is the amount of energy used to produce it. Today we are conscious of the need to conserve energy, both because we are likely to exhaust supplies of oil and natural gas in the near future and because high energy use (in the form of burning fossil fuels such as coal, oil and gas) contributes significantly to global warming. As Table 4.7 shows, the production of aluminium uses about five times the amount of energy (per kilogram of metal) as does steel.

Table 4.7 also points out another aspect of metal use: recycling.

### FOR INVESTIGATION

Search out information to compare the cost in dollars and energy terms of the production of pure aluminium from bauxite and from recycled material. Try to include *all* costs—from bauxite in the ground for virgin material and from houses, shops and factories for recycled material. (The data in Table 4.7 are American and a bit dated, so do not be concerned if your investigation throws up different results.)



## 4.17 RECYCLING OF METALS

There are three advantages in **recycling** metals:

- 1 Less energy is used in recycling metals than in extracting metal from virgin ore (see Table 4.7).
- 2 Finite natural resources (ores) are conserved.
- 3 Less rubbish has to be disposed of (a major consideration in big cities where sites for garbage dumps are becoming hard to find).

Recycling may lead to lower prices for metals, but even without this advantage the above three reasons are strong justification for recycling. Recycling to conserve energy and other natural resources needs to be a worldwide activity, because we are all dependent on resources.

Metal ores are **non-renewable resources**. They were formed when the Earth was formed and there is no way of forming any more of them. While we are unlikely to use up all the known reserves of metal ores in the short term, we nevertheless should use them as sparingly as possible so as to make them last for as long as possible.

The problem with any form of recycling is that the used material has to be collected from very scattered locations. The ore comes from a confined location (the mine site); it is processed into metal and then into useable products at quite specific places or factories; but then as the products are sold and used, they are scattered very widely throughout our communities. Collecting used material for recycling is a major money and energy cost of recycling.

*The steps in recycling a material such as aluminium are:*

- 1 Collect the used products from homes, shopping centres, factories and the like.
- 2 Transport the collected material to a central processing plant.
- 3 Separate the required substance (such as aluminium) from the impurities (labelling, food remnants, dirt).
- 4 Re-smelt the metal into ingots and transport them to product manufacturers.

In recent years attention has focused on recycling aluminium drink containers because they are so commonly used and so widely dispersed throughout the community. There has been (and still is) very significant recycling of iron and steel, copper and lead for many decades, though it has mainly been at the commercial or industrial level rather than by consumers.



### Exercises

- 21 a** Common sources of magnesium are sea water (from which it is precipitated as magnesium hydroxide) and the mineral magnesite, which is magnesium carbonate. Both the hydroxide and the carbonate can be converted to chloride by reacting them with hydrochloric acid. The hydroxide forms water and the carbonate forms carbon dioxide and water. Write equations for these two reactions.
- b** Based upon position in the activity series, suggest a method for extracting magnesium from magnesium chloride. Suggest a reason why magnesium was not used industrially until the middle of the twentieth century.
- \*22** In the activity series, nickel comes between iron and tin. Nickel occurs naturally as a sulfide. How would you extract nickel from this ore?
- 23 a** Mercury can be extracted simply by heating sulfide ores. Mercury does not react with oxygen or water. Where would you place it in the activity series? Explain.

- b** Titanium occurs as titanium(IV) oxide (rutile from beach sands). This oxide cannot be reduced to metal simply by heating with carbon. Titanium tarnishes slowly with oxygen but does not react with water. Where would you place titanium in the activity series? Justify this positioning.
  - c** Which metal, mercury or titanium, would require the greater input of energy to extract it from its ore (after the ore had been concentrated to virtually pure mineral)? Explain.
  - d** Historically, which metal, mercury or titanium, would you expect to have become available earlier? Why?
- \*24a** Which metal do you expect to be the more expensive to produce, iron or magnesium? Explain why.
- b** The final step in extracting titanium from its ore is to react titanium(IV) chloride with magnesium at high temperature. Which metal would you expect to be the more expensive, magnesium or titanium? Explain why.
- 25 a** Why is it more important to recycle aluminium drink containers than steel food containers (tin cans)?
- \*b** What steel products can be recycled relatively easily?
  - c** Why do we not hear much about recycling copper?
  - \*d** Large quantities of lead are recycled. From what products is lead most easily recycled?
  - \*e** Why is it difficult to recycle zinc?

## Important new terms

You should know the meaning of the following terms:

activity series (p. 119)  
 alloy (p. 99)  
 coefficients (in chemical equations)  
     (p. 113)  
 complete ionic equation (p. 117)  
 electrical conductivity (p. 103)  
 electron-transfer reactions (p. 118)  
 first ionisation energy (p. 120)  
 froth flotation (p. 111)  
 half reaction (equation) (p. 118)  
 hardness (p. 103)  
 mineral (p. 110)  
 net ionic equation (p. 116)

neutral species equation (p. 116)  
 non-renewable resource (p. 126)  
 ore (p. 110)  
 oxidation\* (p. 117)  
 recycling (p. 126)  
 redox reaction\* (p. 118)  
 reduction\* (p. 117)  
 smelting (p. 111)  
 spectator ion (p. 117)  
 steel (carbon, alloy, stainless) (p. 100,  
     102)  
 tensile strength (p. 103)  
 thermal conductivity (p. 103)

\* Not in the Preliminary Course syllabus

## Test yourself

- 1 Explain the meaning of each of the items in the ‘Important new terms’ section above.
- 2 Name four common non-ferrous alloys.
- 3 List major uses of steel, aluminium, copper, zinc and lead.
- 4 For each of the uses of the metals you gave in Question 3, describe the properties of the metal that make it particularly suitable for that use.

- 5** Explain the meaning of the terms Bronze Age, Iron Age, and Copper Age. Put them in chronological order and suggest a reason for them occurring in that order.
- 6** What metals have been introduced into widespread use since about 1900 and what have they been used for?
- 7** What are the general steps involved in extracting metals from their ores?
- 8** Illustrate the general steps in Question 7 by explaining how copper is extracted from sulfide ores.
- 9** Write an equation for the reaction of copper(I) sulfide with oxygen.
- 10** Why is a large amount of energy required to extract metals from their ores?
- 11** Write equations for the reaction of lithium, magnesium and iron with oxygen.
- 12** Name three metals which react with water. Write an equation for the reaction of one of them.
- 13** List four metals which react with dilute acids, but not with water. Write equations for the reaction of one of these with sulfuric acid and one with hydrochloric acid.
- 14** Write half equations for the overall reactions you wrote in Question 13, showing the actual electron transfer processes.
- 15** How can we establish an activity series for common metals?
- 16** Arrange the metals copper, gold, iron, lead, magnesium, sodium in order of decreasing reactivity. Explain the basis for your sequence.
- 17** On the Periodic Table, where do the most reactive metals occur?
- 18** List three situations where the choice of which metal to use depends heavily upon chemical reactivity.
- 19** How does ease of extraction of a metal from an ore relate to position in the activity series?
- 20** Why was aluminium not in common use as a metal until the twentieth century? What was the breakthrough that made its widespread use possible?
- 21** List three factors which affect the price of metals.
- 22** Give three reasons for recycling metals.
- 23** What steps are involved in recycling a metal such as aluminium?

# Quantitative aspects of formulae and equations

## IN THIS CHAPTER

Why quantitative calculations are necessary  
Relative atomic mass (atomic weight)  
Relative molecular mass (molecular weight)  
The mole  
Converting between masses, moles and numbers of atoms and molecules  
Per cent composition  
Calculating formulae for compounds

Empirical and molecular formulae  
Chemical equations and moles  
Mass calculations using equations  
Yield and its importance for extracting metals  
The contributions of Dalton, Gay-Lussac and Avogadro to our understanding of formulae, equations and moles  
Volumes of gases in reactions

In Chapter 4 we saw that metals were generally extracted from their ores by reacting them with some other substance. For efficient use of both the ore and the ‘other substance’, we need to be able to calculate the quantities that we should use so as not to waste any of either component. For example, to extract iron from its ore we use carbon in the form of coke; how much coke do we need per tonne of iron ore? Or another example: the last step in extracting titanium from its ore is to react titanium(IV) chloride with magnesium, a very expensive reagent; how much magnesium should we use per kilogram of titanium(IV) chloride?

In this chapter we shall see how chemical formulae and equations allow us to calculate the masses of reactants and/or products involved in chemical reactions.

We shall also see that chemical formulae on their own allow us to calculate the masses of elements present in compounds. For example, we can calculate the amount of iron obtainable from a tonne of the iron ore,  $\text{Fe}_2\text{O}_3$ , or the mass of nitrogen in a kilogram of the fertiliser, sulfate of ammonia,  $(\text{NH}_4)_2\text{SO}_4$ , or the amount of calcium carbonate a person needs to take daily to provide the body’s requirement of the element, calcium.

The study of quantitative aspects of formulae and equations is called **stoichiometry**. The calculations involved are called *stoichiometric calculations*.

In order to perform these calculations we need to introduce a few new terms.

## 5.1 RELATIVE ATOMIC MASS (ATOMIC WEIGHT)

Atoms of one element are different from atoms of any other element. Not surprisingly, then, the atoms of different elements have different masses. However because atoms are extremely small, it is difficult to measure their individual masses. Chemists at first were only able to measure the *relative* masses of atoms; that is, to determine for example that a titanium atom had four times the mass of a carbon atom, that an oxygen atom was 1.33 times the mass of a carbon atom, that a sulfur atom had twice the mass of an oxygen atom. Hence a scale of relative masses was devised.

As a start, let us take the *relative atomic mass* of an element as the mass of an atom of the element relative to the mass of a carbon atom, taken as exactly 12.

A titanium atom is four times the mass of a carbon atom, so its relative atomic mass is  $4 \times 12 = 48$  (Figure 5.1). An oxygen atom is 1.33 times the mass of a carbon atom, so oxygen has a relative atomic mass of  $1.33 \times 12 = 16$ . A sulfur atom has a mass twice that of an oxygen atom; the relative atomic mass of sulfur is therefore  $2 \times 16 = 32$ .

In this way, chemists compiled a table of relative atomic masses of the elements. Some common ones are given in Table 5.1; a more complete (and more accurate) list is printed inside the back cover of this book.

TABLE 5.1 Relative atomic masses (atomic weights) of some common elements

aluminium	Al	27	lead	Pb	207
argon	Ar	40	magnesium	Mg	24.3
calcium	Ca	40	mercury	Hg	201
carbon	C	12	nitrogen	N	14
chlorine	Cl	35.5	oxygen	O	16
copper	Cu	63.5	phosphorus	P	31
fluorine	F	19	potassium	K	39
gold	Au	197	silver	Ag	108
helium	He	4	sodium	Na	23
hydrogen	H	1	sulfur	S	32
iodine	I	127	uranium	U	238
iron	Fe	56	zinc	Zn	65.4

FIGURE 5.1

Two illustrations of the meaning of relative atomic mass (atomic weight)



Three helium atoms have the same mass as one carbon atom. Therefore the relative atomic mass of helium is  $\frac{1}{3} \times 12 = 4$

One titanium atom has the same mass as four carbon atoms. Therefore the relative atomic mass of titanium is  $4 \times 12 = 48$

However there is a problem with the simple definition above. It ignores the existence of isotopes.

## A complication—**isotopes**

For certain elements, some of the atoms have slightly different masses from the other atoms. This is because there are different numbers of neutrons in the nucleus. These different types of atoms of the one element are called **isotopes**. For the element chlorine, 75% of the atoms have 18 neutrons while the other 25% have 20. We say that there are two isotopes of chlorine, one with mass number 35, the other with mass number 37 (*all* chlorine atoms contain 17 protons). Remember (from Section 2.9) that mass number is the sum of the numbers of protons and neutrons. Isotopes are named by their mass numbers. We talk about chlorine-35 and chlorine-37.

In our definition of relative atomic mass then, instead of just ‘mass of an atom’ we need to say ‘*average mass of the atoms in the naturally occurring element*’.

A further complication is that naturally occurring carbon contains two isotopes: 98.89% of the atoms contain 6 neutrons (and 6 protons, so mass number 12), while the other 1.11% contain 7 neutrons (mass number 13). The carbon-12 isotope is used as the standard for relative atomic masses. Hence the strict definition is:

The **relative atomic mass** (or **atomic weight**) of an element is the average mass of the atoms present in the naturally occurring element relative to the mass of an atom of the carbon-12 isotope taken as exactly 12.

Fortunately, the isotopic composition of each element is virtually the same throughout the world.

Of 100 atoms of chlorine, 75 have a mass of 35 while 25 have a mass of 37. The ‘average mass’ in the above definition is therefore:

$$\frac{75 \times 35 + 25 \times 37}{100} = 35.5$$

Similarly, considering 100 atoms of copper which is 69% copper-63 and 31% copper-65, the average mass is:

$$\frac{69 \times 63 + 31 \times 65}{100} = 63.6$$

Naturally occurring carbon is 98.89% carbon-12 and 1.11% carbon-13, so its relative atomic mass is:

$$\frac{98.89 \times 12 + 1.11 \times 13}{100} = 12.01$$

## Why ‘relative to carbon taken as 12’?

If we are compiling a relative scale, it does not matter much which element we use as our reference or what mass we arbitrarily assign to it. Historically, atomic weights were defined relative to the hydrogen atom, which was taken as exactly 1. This was done because hydrogen was the lightest element, so it made sense to let it be unity. However, on that basis nearly all atomic weights turned out to be non-integral: oxygen had a value of 15.9, carbon 11.9, nitrogen 13.9. So it was decided to change the standard to oxygen taken as exactly 16; this gave many elements integer values for their atomic weights and still gave no values less than 1. In 1961 in order to conform to a relative atomic mass scale widely used by physicists, the standard was changed to the carbon-12 isotope, taken as exactly 12.

*It is essential to realise that the relative atomic mass (or atomic weight) of an element is not (repeat, not) the mass of an atom of that element. It is just a relative mass—relative to the mass of a carbon atom. It is not a mass at all, but merely a pure number (that is—it has no units).*

When we say that copper has a relative atomic mass of 63.6, we are really saying that the copper atom is  $63.6/12 = 5.3$  times the mass of the carbon atom. Silver, with a relative atomic mass of 108, has atoms which are  $108/12 = 9$  times the mass of carbon atoms.

## Relative atomic mass or atomic weight?

*Relative atomic mass* is the strictly correct name to use for this term. However when the term was first introduced nearly two centuries ago, it was called *atomic weight* and this name has been widely used ever since. Because of its widespread use, and because it is shorter to say and write, it will be used widely in this book, but you should remember that it is just a synonym for the more correct name, relative atomic mass.

## 5.2 RELATIVE MOLECULAR MASS (MOLECULAR WEIGHT)

While we talk about relative atomic masses of elements, for compounds we refer instead to relative molecular masses. As with elements there is a commonly used historical name, molecular weight.

The **relative molecular mass** (or **molecular weight**) of a compound is the mass of a molecule of the compound relative to the mass of an atom of the carbon-12 isotope taken as exactly 12.

It follows then that:

The molecular weight of a compound is the sum of the atomic weights of the atoms as given in the molecular formula.

In this book the shorter name, molecular weight, will generally be used instead of the more correct but longer name, relative molecular mass.

Using Table 5.1:

- the molecular weight of water,  $\text{H}_2\text{O}$ , is  $2 \times 1 + 16 = 18$
- the molecular weight of nitrogen dioxide,  $\text{NO}_2$ , is  $14 + 2 \times 16 = 46$  and
- the molecular weight of sulfuric acid,  $\text{H}_2\text{SO}_4$ , is  $2 \times 1 + 32 + 4 \times 16 = 98$ .

Again, the molecular weight is not the mass of a molecule: it is a pure number that merely expresses the relative mass of the molecule—relative to the mass of the carbon atom.

### Example 1

Calculate the molecular weight of table sugar, sucrose, which has the formula  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ .

$$\text{Molecular weight} = 12 \times A_{\text{C}} + 22 \times A_{\text{H}} + 11 \times A_{\text{O}}$$

(where  $A$  is the atomic weight of the indicated element)

$$= 12 \times 12.0 + 22 \times 1.01 + 11 \times 16.0$$

(taking atomic weights from inside the back cover)

$$= 342.2$$



We generally use atomic weights rounded off to one decimal place, except for hydrogen where we start with two, though the answer can only have one decimal place. (Why?)

## 5.3 RELATIVE FORMULA MASS (FORMULA WEIGHT)

As we saw in Chapter 2, many compounds, particularly ionic ones, do not exist as discrete molecules. Instead they exist as an orderly array of ions or atoms bound tightly to one another with no recognisable molecules (Figs 2.11, 2.16, 2.17). We still use chemical formulae for these compounds, but instead of representing the composition of discrete molecules, these formulae describe the ratios in which the elements are present in the compounds. Sodium chloride is a typical example: we cannot recognise any discrete NaCl molecules. The formula NaCl instead tells us that in this compound, sodium and chlorine atoms are present in the ratio 1:1. Similarly, in calcium fluoride, CaF<sub>2</sub>, there are no discrete CaF<sub>2</sub> molecules; instead the formula tells us that throughout the sample of calcium fluoride there are two fluorine atoms for every calcium atom present. For such compounds, it is more correct to talk about the relative formula mass (or formula weight) instead of the molecular weight.

The **relative formula mass** (or **formula weight**) of a compound is the sum of the atomic weights of the atomic species as given in the stated formula of the compound.

Again using Table 5.1:

- the formula weight of sodium chloride, NaCl, is  $23.0 + 35.5 = 56.5$
- the formula weight of calcium fluoride, CaF<sub>2</sub>, is  $40 + 2 \times 19 = 76$ .

Formula weight is sometimes called molecular weight. While you are urged to use the more correct term, you should not be unduly worried about the loose use of molecular weight in many textbooks.

### Example 2

Calculate the formula weight of calcium phosphate, Ca<sub>3</sub>(PO<sub>4</sub>)<sub>2</sub>, commonly called rock phosphate and used to make superphosphate fertiliser.

$$\begin{aligned}\text{Formula weight} &= 3 \times A_{\text{Ca}} + 2 \times (A_{\text{P}} + 4 \times A_{\text{O}}) \\ &\quad (\text{where } As \text{ are atomic weights}) \\ &= 3 \times 40.1 + 2 \times (31.0 + 4 \times 16.0) \\ &\quad (\text{taking atomic weights from inside the back cover}) \\ &= \mathbf{310.3}\end{aligned}$$

Note carefully that the subscript 2 refers to everything inside the brackets.



### Exercises

- 1 Atoms of krypton are seven times as heavy as atoms of carbon. What is the atomic weight of krypton?
- 2 The atomic weight of molybdenum is 96. How many times heavier than a carbon atom is a molybdenum atom?



- 3** Calcium has an atomic weight of 40 and neon 20. What does this tell you about the relative masses of calcium and neon atoms?
- 4** Calculate the molecular weight (or formula weight) of:
- |   |  |
|---|--|
| <b>a</b> potassium bromide, KBr                         | <b>i</b> ammonium nitrate, $\text{NH}_4\text{NO}_3$  |
| <b>b</b> carbon dioxide, $\text{CO}_2$                  | <b>j</b> glucose, $\text{C}_6\text{H}_{12}\text{O}_6$                                      |
| <b>c</b> magnesium chloride, $\text{MgCl}_2$            | <b>k</b> calcium hydroxide, $\text{Ca}(\text{OH})_2$                                       |
| <b>d</b> sodium sulfide, $\text{Na}_2\text{S}$          | <b>l</b> copper nitrate, $\text{Cu}(\text{NO}_3)_2$  |
| <b>e</b> iron(III) oxide, $\text{Fe}_2\text{O}_3$       | <b>m</b> ammonium sulfate, $(\text{NH}_4)_2\text{SO}_4$                                    |
| <b>f</b> diphosphorus pentoxide, $\text{P}_2\text{O}_5$ | <b>n</b> aluminium sulfate, $\text{Al}_2(\text{SO}_4)_3$                                   |
| <b>g</b> calcium sulfate, $\text{CaSO}_4$               | <b>o</b> sodium carbonate decahydrate, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ |
| <b>h</b> sodium carbonate, $\text{Na}_2\text{CO}_3$     |  |

## 5.4 THE MOLE

Eventually chemists were able to measure the mass of individual atoms and molecules. It was found that the mass of the carbon atom was  $1.99 \times 10^{-23}$  grams. Expressed differently, there are  $6.02 \times 10^{23}$  atoms in 12 grams of carbon.

Now if  $6.02 \times 10^{23}$  atoms of carbon have a mass of 12 grams, then for titanium (atomic weight 48),  $6.02 \times 10^{23}$  atoms (each of which is four times as heavy as a carbon atom) must have a mass of  $4 \times 12 = 48$  grams. For sulfur  $6.02 \times 10^{23}$  atoms (each of which is  $32/12 = 2.67$  times the mass of a carbon atom) must have a mass of  $2.67 \times 12 = 32$  grams. In general:

If for any element we take the mass which in grams is numerically equal to the atomic weight, then it contains  $6.02 \times 10^{23}$  atoms.

63.6 grams of copper contain  $6.02 \times 10^{23}$  atoms. 108 grams of silver contain  $6.02 \times 10^{23}$  atoms.

The argument also extends to compounds. The molecular weight of water is 18. This means a water molecule is 1.5 times the mass of a carbon atom. Hence  $6.02 \times 10^{23}$  molecules of water must have a mass of 1.5 times the mass of  $6.02 \times 10^{23}$  atoms of carbon, that is  $1.5 \times 12 = 18$  grams.

If for any compound we take the mass which in grams is numerically equal to the molecular weight, then it contains  $6.02 \times 10^{23}$  molecules.

46 grams of nitrogen dioxide,  $\text{NO}_2$ , contain  $6.02 \times 10^{23}$  molecules. 98 grams of sulfuric acid,  $\text{H}_2\text{SO}_4$ , contain  $6.02 \times 10^{23}$  molecules. 342 grams of sucrose,  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ , contain  $6.02 \times 10^{23}$  molecules.

We call this quantity—this  $6.02 \times 10^{23}$  atoms or molecules—a mole. To define terms:

A **mole** of a substance is that quantity which contains as many elementary units (e.g. atoms, ions or molecules) as there are atoms in exactly 12 grams of the carbon-12 isotope.

The **Avogadro constant** (for which we use the symbol  $N_A$ ) is the number of atoms in exactly 12 grams of the carbon-12 isotope.

$$N_A = 6.02 \times 10^{23} \text{ particles per mole}$$

It follows then that the Avogadro constant is the number of atoms or molecules in a mole of any substance.

A mole of an element is the mass that in grams is numerically equal to the atomic weight. A mole of a compound is the mass that in grams is numerically equal to the molecular weight.

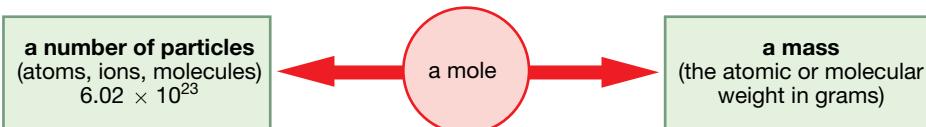
Stated less precisely, a mole of a substance is simply the atomic or molecular weight in grams. A mole of copper (atomic weight 63.6) is 63.6 grams. A mole of nitrogen dioxide (molecular weight 46) is 46 grams.

Figure 5.2 shows a mole of each of several common substances.



FIGURE 5.2  
A mole of each of (clockwise from bottom left) sulfur (32.1 g), zinc (65.4 g), copper sulfate pentahydrate (249.7 g), mercury(II) oxide (216.6 g), sodium chloride (58.5 g), carbon (12.0 g) and, in the centre, copper (63.6 g) and ethanol (46.0 g)

There are two ways of looking at a mole:



A mole of silver (atomic weight 108) is  $6.02 \times 10^{23}$  atoms of silver or 108 grams of silver. A mole of sucrose (molecular weight 342) is  $6.02 \times 10^{23}$  molecules of sucrose or 342 grams of sucrose.

The accepted abbreviation for mole is mol, as in 3.9 mol of zinc or  $4.3 \times 10^{-3}$  mol of carbon. The word 'molecule' is always written in full (to avoid any confusion).

A commonly used term is molar mass.

The **molar mass** is the mass of a mole of the substance. It can be used for both elements and compounds.

Molar mass has units—mass per mole, most commonly grams per mole, usually written as  $\text{g mol}^{-1}$  or g/mol. The molar mass of copper is 63.6 g/mol. The molar mass of nitrogen dioxide is 46 g/mol.

## 5.5 MOLES OF GASEOUS ELEMENTS

When we talk about a mole of oxygen (atomic weight 16), do we mean 16 grams or 32 grams? (Oxygen gas is O<sub>2</sub> with a molecular weight of 32.) This highlights another factor involved in using moles. Sometimes to avoid ambiguity we need to specify which ‘elementary units’ (atoms, ions or molecules) we are talking about.

A mole of oxygen *atoms* is quite clearly 16 grams. A mole of oxygen *molecules* is 32 grams. This mole of oxygen molecules contains 2 moles of oxygen atoms because each molecule of oxygen contains two atoms. Similarly, a mole of chlorine gas (that is molecules) is 2 × 35.5 = 71 grams, whereas a mole of chlorine atoms is 35.5 grams.

## 5.6 CONVERTING BETWEEN MASS, MOLES AND NUMBERS OF ATOMS OR MOLECULES

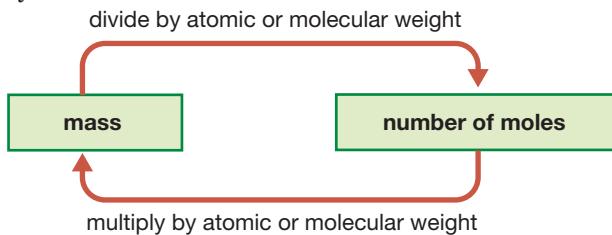
We often need to calculate the number of moles present in a given mass of a substance (element or compound). At other times we want to calculate the number of atoms or molecules in a sample. Even if our final goal is the number of atoms or molecules, we must first convert mass to number of moles and then go on to atoms or molecules.

### Mass to moles

The key equation for converting mass into number of moles<sup>†</sup> is:

$$\begin{aligned}\text{Number of moles} &= \frac{\text{mass}}{\text{mass of one mole}} \\ &= \frac{\text{mass}}{\text{molar mass}} \\ &= \frac{\text{mass}}{\text{atomic or molecular weight in grams}}\end{aligned}\quad \dots (5.1)$$

Schematically:



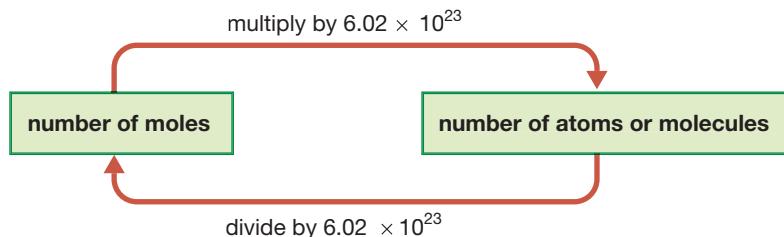
### Moles to number of atoms or molecules

To convert number of moles to number of atoms or molecules we use the Avogadro constant:

$$\text{Number of atoms} = \text{number of moles} \times \text{the Avogadro constant} \quad \dots (5.2)$$

<sup>†</sup> Strictly speaking we should use *amount of substance (in moles)* instead of *number of moles*, because mass in grams divided by molar mass in grams per mole gives the answer in moles. A *number* should not have any units. However, *number of moles* makes the meaning quite clear and is easier to say and write, so it will be used throughout this book, despite its lack of precision.

Again schematically:



To go from mass to number of atoms or molecules (or vice versa) we need to go through number of moles; that is, we have to use both Equations 5.1 and 5.2. Schematically:



Some examples will illustrate the use of these important equations.

### Example 3

How many atoms are there in a pure copper coin weighing 2.56 g? The atomic weight of copper is 63.6.



From Equation 5.1:

$$\begin{aligned}\text{Number of moles of copper} &= \frac{2.56}{63.6} \\ &= 0.0403 \text{ mol}\end{aligned}$$

Now from Equation 5.2:

$$\begin{aligned}\text{Number of atoms of copper} &= 0.0403 \times 6.02 \times 10^{23} \\ &= 2.43 \times 10^{22}\end{aligned}$$

### Example 4

What is the mass of  $10^{12}$  molecules of sulfur dioxide,  $\text{SO}_2$  (a pollutant formed when copper is extracted from sulfide ores)?



First we need to convert from number of molecules to number of moles (using the Avogadro constant and Equation 5.2), then we can use Equation 5.1 to calculate the mass.

$$\begin{aligned}\text{Number of moles} &= \frac{10^{12}}{6.02 \times 10^{23}} \\ &= 1.66 \times 10^{-12}\end{aligned}$$

$$\text{Molar mass of } \text{SO}_2 = 32.1 + 2 \times 16.0 = 64.1 \text{ g/mol}$$

$$\begin{aligned}\text{Mass of } \text{SO}_2 &= 1.66 \times 10^{-12} \times 64.1 \\ &= 1.1 \times 10^{-10} \text{ g} \\ &= 0.11 \text{ ng} \quad (\text{ng} = \text{nanogram} = 10^{-9} \text{ g})\end{aligned}$$



## Exercises

- 5 How many moles are there in:
- a 13 g zinc      c 15 g sulfur      e 5.0 kg iron  
b 0.50 g gold      d 42 g carbon      f 1 tonne lead?
- 6 What is the mass of:
- a 3.0 mol copper      b 0.044 mol sulfur      c  $2.3 \times 10^{-3}$  mol magnesium?
- 7 How many atoms are there in each of (a) to (f) of Exercise 5?
- 8 Calculate the mass of one atom of:
- a zinc      b gold      c helium
- 9 A one carat diamond (pure carbon) weighs 0.10 g. How many atoms does it contain?
- 10 Using results from Exercise 4, calculate the number of moles in 20 g of:
- a sodium sulfide      c glucose  
b diphosphorus pentoxide      d copper nitrate
- 11 Use your results for Exercise 4 to calculate the mass of:
- a 4.6 mol magnesium chloride      c  $6.5 \times 10^{-4}$  mol ammonium nitrate  
b 0.037 mol calcium sulfate
- 12 How many molecules are there in 10 g of:
- a chlorine gas,  $\text{Cl}_2$       d hydrogen peroxide,  $\text{H}_2\text{O}_2$   
b phosphorus pentachloride,  $\text{PCl}_5$       e propane,  $\text{C}_3\text{H}_8$   
c nitric acid,  $\text{HNO}_3$       f ethanol,  $\text{C}_2\text{H}_5\text{OH}$ ?
- 13 What is the mass of:
- a 5.18 mol nitrogen gas      c  $6.24 \times 10^{-2}$  mol hydrogen gas?  
b 0.082 mol argon gas
- 14 How many molecules are there in 7.0 g of nitrogen gas,  $\text{N}_2$ ? How many nitrogen atoms are there in this sample?
- 15 Consider 0.374 g aluminium chloride,  $\text{AlCl}_3$ . It contains how many moles of:
- a aluminium chloride      b aluminium      c chlorine?
- 16 Calculate the mass of one molecule of:
- a ammonia,  $\text{NH}_3$       b chlorine gas,  $\text{Cl}_2$       c sucrose,  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$
- 17 A crystal of table sugar (sucrose,  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ ) weighs 1.4 mg. How many molecules are in it? How many atoms in total are present?

## 5.7 PER CENT COMPOSITION

The chemical formula of a compound tells us directly the ratios in which the atoms are present. In aluminium oxide,  $\text{Al}_2\text{O}_3$ , the Al and O atoms are present in the ratio 2:3. In sulfuric acid,  $\text{H}_2\text{SO}_4$ , the H, S, O atoms are present in the ratio 2:1:4. Sometimes we need to know the ratio by mass in which the elements are present, that is the per cent composition by weight.

Knowing the per cent composition, an industrialist can determine how much aluminium can be extracted from a tonne of aluminium oxide, or a farmer can decide which of several fertilisers contains the most nitrogen per kilogram of compound.

## Calculating per cent composition from a chemical formula

The method of calculating per cent composition from a chemical formula is best illustrated by an example.

### Example 5

Calculate the percentage of iron in the common ore, haematite,  $\text{Fe}_2\text{O}_3$ . Atomic weights are Fe, 55.9 and O, 16.0.

One mole of  $\text{Fe}_2\text{O}_3$  is  $2 \times 55.9 + 3 \times 16$  g = 159.8 g.

One formula unit of  $\text{Fe}_2\text{O}_3$  contains two atoms of iron. Hence one *mole* of  $\text{Fe}_2\text{O}_3$  contains 2 *moles* of iron. Two moles of iron is  $2 \times 55.9$  g = 111.8 g. Hence 159.8 g  $\text{Fe}_2\text{O}_3$  contains 111.8 g Fe.

$$\begin{aligned}\text{Fraction of } \text{Fe}_2\text{O}_3 \text{ which is Fe} &= \frac{111.8}{159.8} \\ &= 0.700\end{aligned}$$

$$\% \text{ iron in } \text{Fe}_2\text{O}_3 = 70.0\%$$



This example illustrates the general method for calculating per cent composition:

$$\% \text{ A in a compound} = \frac{\text{mass of A in 1 mole of the compound}}{\text{mass of one mole of the compound}} \times 100$$

or more specifically,

in the compound of formula  $\text{A}_x\text{B}_y\text{C}_z$ :

$$\% \text{ A} = \frac{w \times (\text{atomic weight of A}) \times 100}{\text{molecular weight of } \text{A}_x\text{B}_y\text{C}_z} \quad \dots (5.3)$$

### Exercises

- 18** **a** Calculate the percentage of sulfur in sodium sulfate,  $\text{Na}_2\text{SO}_4$ .  
**b** Calculate the percentage of aluminium in aluminium oxide,  $\text{Al}_2\text{O}_3$ .  
**c** Calculate the percentage of barium in barium hydroxide,  $\text{Ba}(\text{OH})_2$ .

- 19** Calculate the percentage of each element in:

- a** silver nitrate,  $\text{AgNO}_3$       **c** ammonium phosphate,  $(\text{NH}_4)_3\text{PO}_4$   
**b** aluminium hydroxide  $\text{Al}(\text{OH})_3$       **d** iron(III) sulfate,  $\text{Fe}_2(\text{SO}_4)_3$

- 20** Calculate the percentage of nitrogen in:

- a** ammonium sulfate,  $(\text{NH}_4)_2\text{SO}_4$       **b** urea,  $\text{CH}_4\text{ON}_2$

Both of these compounds are used as nitrogen fertilisers. Which is better on a 'yield of nitrogen per gram of fertiliser' basis?

- 21** The explosive, TNT, has the molecular formula,  $\text{C}_7\text{H}_5\text{N}_3\text{O}_6$ . Calculate the percentage of nitrogen in the compound.
- 22** The common ore of copper, chalcopyrite, has the formula  $\text{CuFeS}_2$ . What is the percentage of copper in this compound?



- 23** Rutile,  $\text{TiO}_2$ , and ilmenite,  $\text{FeTiO}_3$ , are the common minerals from which titanium is extracted. What mass of ilmenite yields the same amount of titanium as 1.00 tonne of rutile does?
- 24** Chalcocite,  $\text{Cu}_2\text{S}$ , malachite,  $\text{Cu}(\text{OH})_2 \cdot \text{CuCO}_3$ , and azurite,  $\text{Cu}(\text{OH})_2 \cdot 2\text{CuCO}_3$ , are minerals from which copper is extracted (in addition to the chalcopyrite,  $\text{CuFeS}_2$ , discussed in Section 4.7 and Exercise 22). Arrange these four minerals in order of decreasing percentage copper content.

So far in this chapter, it has been assumed that formulae are known for all the compounds concerned. Actually, the determination of formulae for compounds was a far from trivial task, and it is not merely a task of the past. Whenever chemists prepare a new compound today, they must determine its formula.

## 5.8 CALCULATING FORMULAE FROM EXPERIMENTAL DATA

To determine the formula of a compound, a chemist must perform an analysis that gives the mass of each element present in a given mass of the compound or, alternatively, which gives the per cent composition of the compound. An example will illustrate.



### Example 6

When 5.12 g lithium metal was allowed to react with excess air (oxygen), it was completely converted to lithium oxide. 11.05 g of oxide was formed. Calculate the formula of the compound.

The experiment showed that:

$$5.12 \text{ g lithium combined with } (11.05 - 5.12) = 5.93 \text{ g oxygen}$$

The formula of a compound tells how many of each type of atom are present in the molecule; alternatively it tells how many *moles* of each element are present in 1 mole of the compound. Let us then convert our masses of Li and O into moles (1 mole of Li is 6.9 g, 1 mole of O is 16.0 g).

$$\frac{5.12}{6.9} \text{ mol Li combines with } \frac{5.93}{16.0} \text{ mol O}$$

$$0.74 \text{ mol Li with } 0.37 \text{ mol O}$$

Because we want the ratio by moles to be in simple whole numbers, let us divide through by the smallest number, that is by 0.37:

$$2.0 \text{ mol Li with } 1.0 \text{ mol O}$$

This means that the formula of the compound is  $\text{Li}_2\text{O}$ . It is called an empirical formula.

### Two types of formulae

The **empirical formula** of a compound is the formula that tells us the ratio in which the atoms are present in the compound.

For ionic and covalent lattice compounds this is the only formula we can write (Sections 2.15 and 2.23). However for molecular compounds, the usual formula we write is what we call a molecular formula.

The **molecular formula** of a compound is the formula that tells us how many of each type of atom are present in a molecule of the compound.

Some molecular compounds of carbon (with their molecular formulae) are ethene, C<sub>2</sub>H<sub>4</sub>, propene, C<sub>3</sub>H<sub>6</sub>, pentene, C<sub>5</sub>H<sub>10</sub>, octene, C<sub>8</sub>H<sub>16</sub> and cyclohexane, C<sub>6</sub>H<sub>12</sub>. For all of these the carbon and hydrogen atoms are present in the ratio of 1:2. This means that all of these compounds have the empirical formula, CH<sub>2</sub>.

Chemical analysis of a compound gives us its empirical formula, even for molecular compounds. To get the molecular formula (of a molecular compound) we need extra information. However, we will not go into this matter here.

The method for calculating an empirical formula is summarised in Box 5.1 and is further illustrated in Example 7.

#### BOX 5.1 TO CALCULATE AN EMPIRICAL FORMULA

- 1 Write down the masses of all of the elements present in a given sample of the compound (call percentages masses in 100 g of compound).
- 2 Convert masses to moles (by dividing by atomic weights in grams).
- 3 Divide through by the smallest number of moles to get a simple ratio.
- 4 If the numbers in step 3 are not close to whole numbers, multiply through by a suitable factor to make them whole numbers (for example multiply 1.3:2.7:1 by 3 to get 3.9:8.1:3 which in step 5 becomes 4:8:3).
- 5 Round off the numbers in step 3 or 4 to get whole numbers and use these to write the empirical formula (round-off errors of up to 10% are quite normal).

### Example 7

The chlorofluorocarbon called CFC-113 is one of the compounds recently banned from industrial use because of its contribution to the ozone hole. Upon analysis a 2.37 g sample of this compound was found to contain 1.36 g chlorine and 0.71 g fluorine, with the balance being carbon. Calculate the empirical formula of the compound.



Following the procedure of Box 5.1:

1 1.36 g Cl is combined with 0.71 g F and with  $(2.37 - 1.36 - 0.71) = 0.30$  g C

2  $\frac{1.36}{35.5}$  mol Cl ...       $\frac{0.71}{19.0}$  mol F ...       $\frac{0.30}{12.0}$  mol C

0.0383 mol ...      0.0374 mol F      0.0250 mol C

3  $\frac{0.0383}{0.0250}$  mol Cl ...       $\frac{0.0374}{0.0250}$  mol F ...       $\frac{0.0250}{0.0250}$  mol C

1.53 mol Cl ...      1.49 mol F ...      1 mol C

4 3.06 mol Cl ...      2.98 mol F ...      2 mol C

5 3 mol Cl ...      3 mol F ....      2 mol C

Hence the empirical formula is C<sub>2</sub>F<sub>3</sub>Cl<sub>3</sub>.



## Exercises

- 25 1.00 g of a compound of copper and chlorine contained 0.524 g chlorine. Calculate the empirical formula of the compound.
- 26 A compound of silver contained 69.2% silver, 10.3% sulfur and 20.5% oxygen. Calculate the empirical formula.
- 27 Magnesium is one of only a few metals that react with nitrogen. When 2.55 g magnesium was heated in nitrogen gas until all the magnesium had reacted, the mass of the solid nitride formed was 3.51 g. Calculate the empirical formula of magnesium nitride.
- 28 a When samples of aluminium, tin, zinc and lead were heated in air and completely converted to oxide, the masses of the metals increased by 89% for aluminium, 27% for tin, 24.5% for zinc and 7.7% for lead. Calculate the mass of oxygen that is combined with 5.00 g of each metal.  
b For each of the metals in (a), use the mass of oxygen combined with 5.00 g of the metal to calculate the empirical formula of the oxide of the metal.
- 29 If 1.00 g samples of (a) lithium (b) beryllium (c) magnesium and (d) copper were heated in oxygen until they were completely converted to oxides, what increase in mass would you expect for each metal? Identify two factors that determine the percentage increase.

## 5.9 MOLES AND CHEMICAL EQUATIONS

We have already seen that the chemical equation:



tells us that:

- 1 qualitatively, solid phosphorus reacts with oxygen gas to form solid diphosphorus pentoxide
- 2 four atoms of phosphorus react with five molecules of oxygen to form two molecules of diphosphorus pentoxide

Because atoms and molecules have such small masses, we usually have to work with very large numbers of them in order to have masses that we can conveniently weigh in the laboratory. Since Statement 2 above is true, then multiplying through by the Avogadro constant we can say:

- 3 four moles of phosphorus react with five moles of oxygen to form two moles of diphosphorus pentoxide.

That is, chemical equations can be read in terms of *moles* as well as in terms of atoms and molecules.



## Example 8

- (a) How many moles of oxygen are needed to react with 1 mole of aluminium (to form  $\text{Al}_2\text{O}_3$ )?

First we need a balanced equation:



From this equation we see that 4 moles of Al react with 3 moles of  $\text{O}_2$ , so we can write:

$$\frac{\text{number of moles of O}_2 \text{ required}}{\text{number of moles of Al given}} = \frac{3}{4}$$

So number of moles of  $\text{O}_2$  required = 0.75 mol (for 1 mol Al)

- (b)** Acetylene burns with oxygen (in oxyacetylene torches for cutting and welding metals) to form carbon dioxide and water. How many moles of  $\text{CO}_2$  are formed when  
**i** 1 mol   **ii** 0.25 mol acetylene burns?

The balanced equation is:



From this equation:

$$\frac{\text{number of moles of CO}_2 \text{ formed}}{\text{number of moles of C}_2\text{H}_2 \text{ used}} = \frac{4}{2}$$

$$\begin{aligned} \text{In i number of moles CO}_2 \text{ formed} &= \frac{4}{2} \times 1 \\ &= \mathbf{2 \text{ mol CO}_2} \end{aligned}$$

$$\begin{aligned} \text{In ii number of moles CO}_2 \text{ formed} &= \frac{4}{2} \times 0.25 \\ &= \mathbf{0.5 \text{ mol CO}_2} \end{aligned}$$

## Equations in terms of masses

We can convert a moles statement about an equation into a mass statement. We saw above that Equation 5.4 can be read as:

*four moles of phosphorus react with five moles of oxygen to form two moles of diphosphorus pentoxide*

Using the fact that a mole is the atomic or molecular weight in grams, we can convert this moles statement into a grams one:

$4 \times 31.0$  grams of phosphorus react with  $5 \times 32.0$  grams of oxygen to form  $2 \times (2 \times 31.0 + 5 \times 16)$  grams of diphosphorus pentoxide.

That is, 124.0 g P react with 160.0 g  $\text{O}_2$  to form 284.0 g  $\text{P}_2\text{O}_5$ .

A chemical equation tells us the ratios by mass in which substances react or are formed in a reaction.

If we wish to calculate the mass of diphosphorus pentoxide formed from 0.25 g phosphorus, we first write the equation (5.4 above), then say:

4 moles of P  $\longrightarrow$  2 moles of  $\text{P}_2\text{O}_5$

or writing it another way:

$$\frac{\text{number of moles of P}_2\text{O}_5 \text{ produced}}{\text{number of moles of P used}} = \frac{2}{4}$$

$$\text{mass of P used} = 0.25 \text{ g}$$

$$\begin{aligned} \text{number of moles of P used} &= \frac{0.25}{31.0} \\ &= 0.0081 \end{aligned}$$

$$\begin{aligned} \text{number of moles of P}_2\text{O}_5 \text{ produced} &= \frac{2}{4} \times 0.0081 \\ &= 0.0040 \end{aligned}$$

$$\begin{aligned} \text{mass of P}_2\text{O}_5 \text{ produced} &= 0.0040 \times (2 \times 31.0 + 5 \times 16.0) \\ &= 0.57 \text{ g} \end{aligned}$$

That is, **0.25 g P produces 0.57 g  $\text{P}_2\text{O}_5$** .

In this calculation we have ignored the oxygen because we did not need to do any calculations on it. If instead we had been asked to calculate the mass of oxygen needed to react with 0.25 g phosphorus, we should have proceeded as follows:

From Equation 5.4:

4 moles of P react with 5 moles of O<sub>2</sub>. Or

$$\frac{\text{number of moles of O}_2 \text{ needed}}{\text{number of moles of P used}} = \frac{5}{4}$$

number of moles of P used = 0.0081 (from above)

$$\text{so number of moles of O}_2 \text{ needed} = \frac{5}{4} \times 0.0081$$

$$= 0.010 \text{ mol}$$

$$\text{and mass of O}_2 \text{ needed} = 0.010 \times 32$$

$$= 0.32 \text{ g}$$

That is, **0.25 g P reacts with 0.32 g O<sub>2</sub>**.

## 5.10 MASS CALCULATIONS FROM CHEMICAL EQUATIONS

The basic ideas explained above for the phosphorus–oxygen reaction can be used to calculate:

- 1 the mass of one reactant needed to react completely with a given mass of another reactant, or
- 2 the mass of a product that can be formed from a given mass of a reactant.

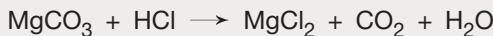
An example will illustrate.



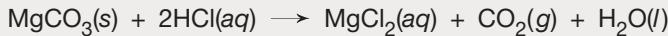
### Example 9

The first step in extracting magnesium from one of its ores, magnesite (magnesium carbonate), is to dissolve the ore in hydrochloric acid. Calculate the mass of hydrochloric acid needed to dissolve 5.0 g of magnesium carbonate.

First we write the equation in skeletal form:



Then we balance it:



From the balanced equation one mole MgCO<sub>3</sub> reacts with 2 moles HCl, or expressed differently: ... (5.5)

$$\frac{\text{number of moles of HCl required}}{\text{number of moles of MgCO}_3 \text{ given}} = \frac{2}{1}$$

That is, number of moles HCl required = 2 × (number of moles of MgCO<sub>3</sub> given)

Hence we need to calculate the number of moles of magnesium carbonate given. To do this we first want:

$$\begin{aligned}\text{Molar mass of MgCO}_3 &= 24.3 + 12.0 + 3 \times 16.0 \\ &= 84.3 \text{ g}\end{aligned}$$

From Equation 5.1:

$$\begin{aligned}\text{number of moles of MgCO}_3 \text{ given} &= \frac{5.0 \text{ g}}{84.3 \text{ g}} \\ &= 0.059\end{aligned}$$

$$\begin{aligned}\text{number of moles of HCl needed} &= 2 \times 0.059 \\ &= 0.118\end{aligned}$$

As Equation 5.1 can be rearranged to:

$$\text{mass} = \text{number of moles} \times \text{mass of one mole} \dots (5.6)$$

$$\begin{aligned}\text{mass of HCl} &= 0.118 \times (1.01 + 35.5) \\ &= 4.3 \text{ g}\end{aligned}$$

That is, we need **4.3 g** HCl to react completely with 5.0 g MgCO<sub>3</sub>.

This method of doing what we call **mass–mass calculations** is summarised in Box 5.2 and shown schematically in Figure 5.3. It is illustrated again in Example 10.

#### BOX 5.2 TO PERFORM A MASS–MASS CALCULATION

- 1 Write a *balanced chemical equation*.
- 2 Calculate the *number of moles of the given substance* (from Equation 5.1).
- 3 Use the chemical equation to write an expression for:

$$\frac{\text{number of moles of required substance}}{\text{number of moles of given substance}}$$

This is simply equal to the ratio of the stoichiometric coefficients in the equation as shown in Examples 9 and 10.

- 4 Use this ratio to calculate the *number of moles of the required substance*.
- 5 Calculate the *mass of the required substance* from rearranged Equation 5.1.

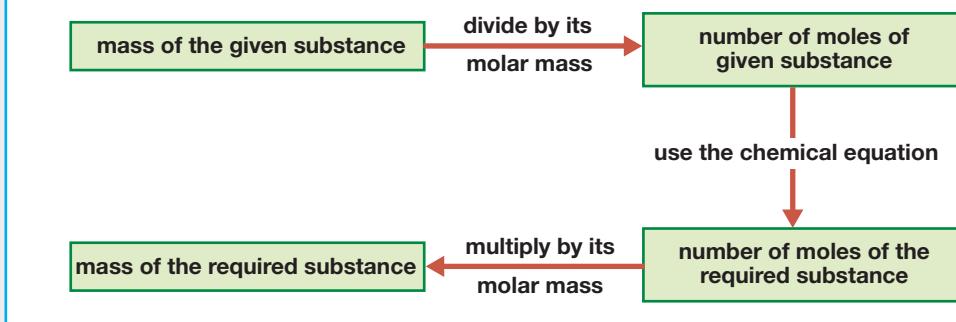


FIGURE 5.3  
Steps for a mass–mass calculation involving a chemical reaction

#### Example 10

Calculate the amount of iron needed to react with oxygen to produce 3.5 grams of iron(III) oxide, Fe<sub>2</sub>O<sub>3</sub>. The equation is:



**Step 1:** The equation is given.



$$\begin{aligned}
 \text{Step 2: no. of moles of iron oxide produced} &= \frac{\text{mass of iron(III) oxide}}{\text{molar mass of iron(III) oxide}} \\
 &= \frac{3.5 \text{ g}}{(2 \times 55.9 + 3 \times 16.0)} \\
 &= \frac{3.5}{159.8} \\
 &= 0.0219
 \end{aligned}$$

### **Step 3:**

$$\frac{\text{number of moles of iron needed}}{\text{number of moles of Fe}_2\text{O}_3 \text{ produced}} = \frac{4}{2}$$

## **Exercises**

- 30 a** Ethane ( $C_2H_6$ , a minor constituent of natural gas) burns in oxygen to form carbon dioxide ( $CO_2$ ) and water. How many moles of oxygen are needed to react with one mole of ethane? How many moles of:

(i)  $CO_2$       (ii) water are produced?

**b** Octane ( $C_8H_{18}$ , a constituent of petrol) burns in air to form  $CO_2$  and  $H_2O$ . How many moles of oxygen are needed to react with 10 moles of octane and how many moles of water are produced?

**31 a** Nitrogen and hydrogen combine to form ammonia,  $NH_3$ . How many moles of hydrogen are needed to form 40 moles of ammonia?

**b** Sulfur dioxide,  $SO_2$ , reacts with oxygen to form sulfur trioxide,  $SO_3$ . How many moles of oxygen are needed to form 0.6 mole of sulfur trioxide?

**32 a** What mass of zinc is formed when 0.2 mole zinc oxide reacts with excess carbon?

**b** What mass of magnesium hydroxide,  $Mg(OH)_2$ , is precipitated from solution when excess sodium hydroxide is added to a solution containing 0.050 mol magnesium ion? This reaction can be used to extract magnesium from sea water.

**33** Calcium carbonate,  $CaCO_3$ , when heated decomposes to calcium oxide ( $CaO$ ) and carbon dioxide. Calculate the mass of calcium oxide formed when 2.0 g calcium carbonate is decomposed.

**34** What mass of hydrogen gas is produced when 1.3 g zinc reacts with excess hydrochloric acid,  $HCl$ ? How much zinc chloride,  $ZnCl_2$ , is produced?

**35** Sodium reacts with water to form hydrogen and sodium hydroxide solution. What mass of water is decomposed by 1.32 g sodium and what mass of hydrogen is formed?

**36** The first step in extracting aluminium oxide from bauxite is to dissolve it in sodium hydroxide solution. The reaction is:

$$Al_2O_3(s) + 2NaOH(aq) + 3H_2O(l) \longrightarrow 2Al(OH)_4^-(aq) + 2Na^+(aq)$$

What mass of sodium hydroxide is needed to react with 12.7 g aluminium oxide?

**37** To extract copper from cuprite,  $Cu_2S$ , the ore is roasted in air: copper and sulfur dioxide form. What mass of sulfur dioxide is produced per tonne of copper by this reaction?

## 5.11 YIELD AND ITS IMPORTANCE FOR EXTRACTING METALS

The percentage composition and mass–mass calculations just described are widely used in the mining and minerals industries to calculate or measure the yields of metals from particular minerals or ores.

The **yield** of a metal from a particular mineral or ore is the mass of metal that can be obtained from a particular mass of the mineral or ore.

Yield is often expressed as a percentage. For example, the yield of iron from haematite,  $\text{Fe}_2\text{O}_3$ , is 70%; the yield of copper from chalcopyrite,  $\text{CuFeS}_2$ , is 35%.

If we express yield as a percentage of the ore dug out of the ground (the required mineral plus other minerals, dirt and rock that occur with it), then it will generally be much lower than if it is expressed as a percentage of the pure mineral. For example, copper ores typically have yields of 1 to 5% copper. Gold ores commonly yield 10 to 50 g gold per tonne (yields of 0.001% to 0.005%).

While we can use formulae to calculate (or predict) yields of metals from particular minerals (pure compounds), for ores we have to measure them experimentally. This is because ores are mixtures of the required mineral and unwanted material, and, being mixtures, they have variable composition (from one location to another).

Measurement of the yield from a particular ore body is extremely important in the mining and minerals industry because it determines whether extraction of the metal from that ore is profitable or not. Whenever a new ore body is found, samples must be analysed to determine the likely yield of the metal from it, and hence to decide whether it is economically worthwhile to mine and process the ore. The price that the metal can be sold for (which can fluctuate quite markedly depending upon economic conditions) is also a major factor in making that decision.

### Example 11

Two minerals containing tungsten are scheelite (calcium tungstate,  $\text{CaWO}_4$ ) and tungstenite,  $\text{WS}_2$ :

- (a) Calculate the percentage yield of tungsten from these minerals.  
(b) Suppose that two ores contained 8.2% scheelite and 5.5% tungstenite respectively. Calculate the percentage yield of tungsten from these ores.

(a) In scheelite, using Equation 5.3:

$$\begin{aligned}\% \text{ tungsten} &= \frac{183.9}{40.1 + 183.9 + 4 \times 16.0} \times 100\% \\ &= \mathbf{63.9\%}\end{aligned}$$

In tungstenite, again using Equation 5.3:

$$\begin{aligned}\% \text{ tungsten} &= \frac{183.9}{183.9 + 2 \times 32.1} \times 100\% \\ &= \mathbf{74.1\%}\end{aligned}$$



**(b)** Since scheelite is only 8.2% of the ore:

$$\begin{aligned}\% \text{ tungsten in the ore} &= \frac{8.2}{100} \times 63.9\% \\ &= 5.2\%\end{aligned}$$

Since tungstenite is only 5.5% of the ore:

$$\begin{aligned}\% \text{ tungsten in the ore} &= \frac{5.5}{100} \times 74.1\% \\ &= 4.1\%\end{aligned}$$

Another factor in considering yield is the difference between the yield obtained in a large-scale industrial plant and that obtained in a chemical laboratory: industrial processes generally produce lower yields than the theoretical or laboratory ones. Industrialists need to monitor yields, because a low yield from a chemical process for extracting a metal would have a serious effect upon the economic viability of the extraction operation.

Even when we carry out a chemical reaction in the laboratory, the amount of product we get is often less than what was expected from the chemical equation. This gives rise to another use of the term yield, namely the percentage yield of a chemical reaction.

The **percentage yield** of a chemical reaction is the amount of product obtained expressed as a percentage of the amount expected from the chemical equation.

For example, say that for the reaction of a certain quantity of copper(I) sulfide with oxygen (air), we expected (from the equation) to get 2.33 g, but when we performed the reaction we obtained only 1.97 g, so the percentage yield of the reaction would be  $\frac{1.97 \times 100}{2.33} = 85\%$ .

Some reasons why yields are often less than 100% are:

- The product does not all precipitate out of solution.
- There are physical losses (such as spillages during filtration or leakage of gases).
- Some of the reactant or product is lost by vaporisation.
- Not all the reactant actually reacts (because it becomes covered with a product).
- The reaction becomes too slow for us to wait for it to go to completion.
- There is a side reaction, which uses up some of the reactant without producing the desired product.

## Exercises

**38** Upon analysis it was found that 1.364 g of a particular nickel ore produced 0.0423 g nickel. What is the percentage yield of nickel from this ore?

**39** Magnetite,  $\text{Fe}_3\text{O}_4$ , is a mineral from which iron is sometimes extracted. Compare the yields of iron from magnetite and from haematite,  $\text{Fe}_2\text{O}_3$ . What other factors would affect the practical yield of iron from ores containing these minerals?

**40** If the theoretical yield of aluminium from a particular bauxite ore is 14% and if the extraction process is only 87% efficient, what mass of bauxite needs to be processed to produce 1.00 tonne of aluminium?

**41** What mass of sodium would be produced from the electrolysis of 1.00 tonne of molten sodium chloride if the electrolysis process was 92% efficient?

Let us finish this discussion of quantitative aspects of formulae and equations by looking at some aspects of the historical development of these ideas.

## 5.12 SOME HISTORY—DALTON, GAY-LUSSAC AND AVOGADRO

Based upon quantitative work that had been done over the previous decades, John Dalton in 1803 and in fuller form in 1808 proposed a theory that became known as **Dalton's atomic theory**.

The three postulates of Dalton's atomic theory were:

- 1 Matter is composed of tiny indivisible particles called atoms.
- 2 All atoms of the one element are identical, but different from the atoms of all other elements.
- 3 Chemical reactions consist of combining, separating or rearranging atoms in simple whole number ratios.

This theory led to the use of symbols for atoms of elements, formulae for molecules of compounds, and equations for chemical reactions. It also led to the concept of relative atomic mass (atomic weight). In fact all the material we have been talking about so far in this chapter is based on Dalton's atomic theory.

However, formulae and atomic weights did not flow *directly* from the atomic theory; there was still a problem. In water, for example, it was determined that 2 g hydrogen was combined with 16 g oxygen. Was water HO with H and O having atomic weights of 2 and 16, or was it  $H_2O$  with atomic weights of 1 and 16? It was Gay-Lussac and Avogadro who provided the answer.



Joseph Gay-Lussac (1778–1850) was a French physicist and chemist and a great ballooning enthusiast, which explains his interest in properties of gases. He designed many pieces of laboratory glassware still in routine use today



Amadeo Avogadro (1776–1856) was an Italian lawyer turned scientist. His famous hypothesis, largely ignored during his lifetime partly because of the obscurity of his written expression, became the basis for determining formulae and atomic weights in the latter part of the nineteenth century

## Gay-Lussac's law of combining volumes and Avogadro's hypothesis

After studying the volumes in which gases reacted, Gay-Lussac, in 1808, proposed the **law of combining volumes**:

When measured at constant temperature and pressure, the volumes of gases taking part in a chemical reaction show simple whole number ratios to one another.

For example:

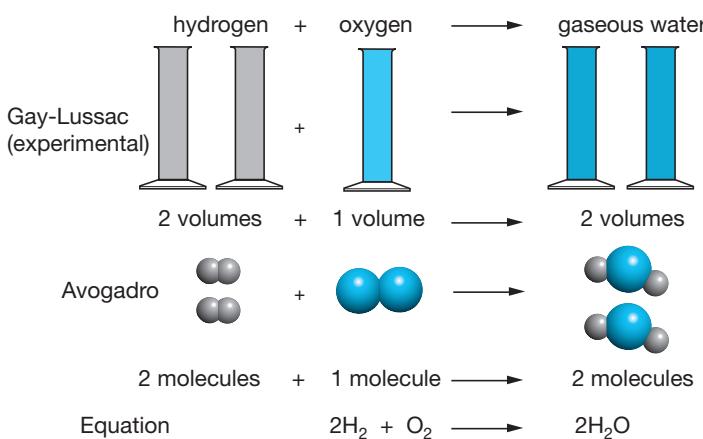
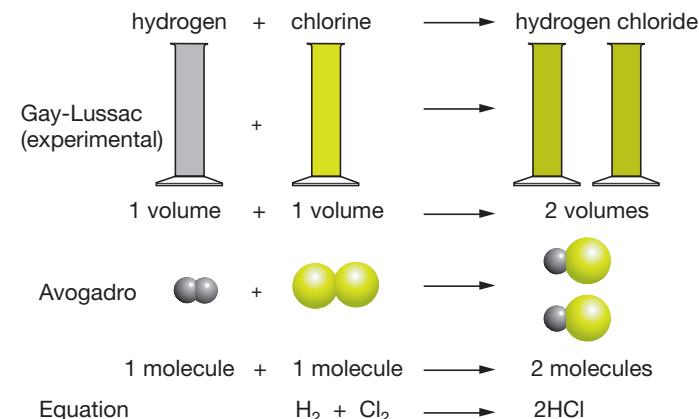
- 1 100 mL of hydrogen reacts with 100 mL of chlorine to form 200 mL of hydrogen chloride (*one* ‘volume’ reacts with *one* ‘volume’ to form *two* ‘volumes’) (Fig. 5.4). ... (5.7)
- 2 100 mL of hydrogen reacts with 50 mL of oxygen to form 100 mL of steam (gaseous water) at temperatures above 100°C (*two* ‘volumes’ react with *one* ‘volume’ to form *two* ‘volumes’) (Fig. 5.4). ... (5.8)

Avogadro noted the similarity between Gay-Lussac’s statement and the third postulate of the atomic theory: that when atoms or molecules react, they do so in simple whole number ratios. This led him, in 1811, to propose the following:

When measured at the same temperature and pressure, equal volumes of gases contain the same number of molecules. ... (5.9)

FIGURE 5.4

The work of Gay-Lussac and Avogadro was important for developing formulae and determining atomic weights



This became known as **Avogadro's hypothesis**. It was not a summary of experimental fact as the traditional laws of science were, but instead it was a theory to explain observations. Like all theories, at first it was called an hypothesis. Today we accept it as a law because there is now indisputable evidence for its truth.

Applying Avogadro's hypothesis to Statement 5.7, we deduce that *one molecule* of hydrogen combines with *one molecule* of chlorine to form *two molecules* of hydrogen chloride. Consequently, a molecule of hydrogen must contain *at least two atoms* of hydrogen (because it contributes to two molecules of hydrogen chloride), and similarly a molecule of chlorine must also contain at least two atoms (Fig. 5.4).

Likewise, from Statement 5.8, *two molecules* of hydrogen combine with *one molecule* of oxygen to form *two molecules* of gaseous water. Because one molecule of oxygen contributes to two molecules of water, it must contain at least two atoms of oxygen. Two molecules of hydrogen (which cannot be monatomic from Statement 5.7) reacting with one molecule of oxygen meant that water could not be HO, and so was most likely to be H<sub>2</sub>O (Fig. 5.4). No reactions could be found which required oxygen, hydrogen or chlorine molecules to split into more than two atoms. It was therefore concluded that they were all diatomic molecules and consequently that water was definitely H<sub>2</sub>O.

## Significance of the work of Gay-Lussac and Avogadro

Gay-Lussac's law of combining volumes and Avogadro's hypothesis allowed chemists to use results from quantitative analyses of compounds and reactions to determine formulae for compounds and hence relative atomic masses for elements.

The development of formulae for substances and the determination of relative atomic masses for elements went hand in hand, often with quite complicated and indirect arguments being needed to decide between alternatives (such as whether red mercury oxide was HgO with the atomic weight of mercury being 200, or Hg<sub>2</sub>O with atomic weight 100). A similar problem with indium was solved by Mendeleev as Exercise 2 in Chapter 6 will show. Such determinations of formulae and atomic weights were major parts of chemistry throughout the nineteenth century.

Because the existence of formulae and atomic weights, and hence the ability to write chemical equations, are essential for talking about moles, we can say that the work of Gay-Lussac and Avogadro was critically important in developing the mole concept.

## 5.13 VOLUMES OF GASES IN REACTIONS

Avogadro's law in the form of Statement 5.9 can be rearranged to read:

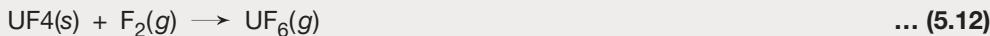
Equal numbers of molecules of different gases occupy the same volume  
(at the same temperature and pressure). ... (5.10)

In this form Avogadro's law allows us to convert statements about numbers of molecules in chemical equations into statements about volumes, *but only for gaseous reactants and products*. An example will illustrate.



## Example 12

As part of the preparation of uranium for isotopic enrichment for use in nuclear reactors, uranium(IV) oxide is reacted with hydrogen fluoride gas to form uranium(IV) fluoride, which is then reacted with fluorine to form uranium hexafluoride:



- a If 25 L hydrogen fluoride reacts with excess uranium(IV) oxide, what volume of steam is formed?
- b If 20 L fluorine then reacts with excess uranium(IV) fluoride, what volume of uranium hexafluoride is formed? All volumes are measured at the same temperature and pressure.
- a From Equation 5.11 we can write:  
4 molecules of HF form 2 molecules of steam

From Statement 5.10, it follows that:

4 volumes of HF form 2 volumes of steam

or 25 L HF form **12.5 L** steam

(if '4 volumes' is 25 L, then '2 volumes' must be 12.5 L)

- b From Equation 5.12:  
1 molecule of F<sub>2</sub> produces 1 molecule of UF<sub>6</sub>

From Statement 5.10:

1 volume of F<sub>2</sub> produces 1 volume of UF<sub>6</sub>

or 20 L F<sub>2</sub> produces **20 L** UF<sub>6</sub>

## Exercises

**42** Measured at constant temperature and pressure, 300 mL carbon monoxide reacted with 100 mL ozone (which contains only the element oxygen) to form 300 mL carbon dioxide. Assuming that you do not know the formulae for carbon monoxide, carbon dioxide and ozone, deduce the number of atoms of oxygen in a molecule of ozone. Carefully explain your reasoning.

**43** An unknown gaseous element, G, reacts with oxygen. When measured at constant temperature and pressure, it was found that 125 mL of G reacted with 250 mL of oxygen to form 250 mL of gaseous oxide. We know that oxygen is diatomic. What is the smallest number of atoms that there can be in one molecule of G? What would this give for the formula of the oxide? This oxide contained 30.4% G. If the formula is as just decided, what is the atomic weight of G? Are other formulae possible for the molecule of G and hence for the oxide? Explain.

- 44 a** Carbon monoxide reacts with iron(III) oxide to form carbon dioxide and iron. Write an equation for this reaction. If 5.0 L carbon monoxide were reacted with excess iron, what volume of carbon dioxide (measured at the same temperature and pressure) would be formed? Explain how you worked this out.
- b** One method of purifying nickel is to react it with carbon monoxide. Gaseous nickel carbonyl, Ni(CO)<sub>4</sub>, forms. This is separated from solid impurities, then heated to a higher temperature; this decomposes the compound back to nickel. Write an equation for the formation of nickel carbonyl. What volume of carbon monoxide, measured at the same temperature and pressure, is needed to form 2.0 L nickel carbonyl?

## Important new terms

You should know the meaning of the following terms:

atomic weight (p. 131)

Avogadro constant (p. 134)

Avogadro's hypothesis (law) (p. 151)

Dalton's atomic theory (p. 149)

empirical formula (p. 140)

formula weight (p. 133)

Gay-Lussac's law of combining volumes  
(p. 150)  
isotopes (p. 131)  
molar mass (p. 135)  
mole (p. 134)  
molecular formula (p. 141)  
molecular weight (p. 132)

percentage yield (p. 148)  
relative atomic mass (p. 131)  
relative formula mass (p. 133)  
relative molecular mass (p. 132)  
stoichiometry (p. 129)  
yield (p.147)

## Test yourself

- 1 Explain the meaning of each of the items in the 'Important new terms' section above.
- 2 Krypton atoms are seven times as heavy as carbon atoms. What is the atomic weight of krypton?
- 3 How is molecular weight of a compound calculated from atomic weights?
- 4 What is the difficulty with using the term *molecular weight* for compounds such as sodium chloride and magnesium oxide? What term should be used?
- 5 The atomic weight of copper is 63.6. What is the mass of a mole of copper?
- 6 What is wrong with the statement: 'The molecular weight of sulfuric acid is 98 g.'?
- 7 How do you calculate the number of moles in a sample of a compound, given its mass and molecular weight?
- 8 What is unclear about the phrase, 'a mole of nitrogen'?
- 9 How do you convert **(a)** a mass to a number of moles, **(b)** a number of moles to a number of molecules, and **(c)** a mass to a number of molecules?
- 10 Explain how you would calculate the percentage sulfur in aluminium sulfate,  $\text{Al}_2(\text{SO}_4)_3$ .
- 11 Explain the difference between empirical and molecular formulae.
- 12 Which formula can be calculated from a chemical analysis of a compound?
- 13 What steps are involved in calculating the mass of a product formed from a given mass of a reactant?
- 14 Why is it necessary to determine the yield of a metal from a newly discovered ore?
- 15 We can talk about the yield of a metal from a mineral or from an ore. Which (if either) can we predict and which (if either) must we determine experimentally? Explain why.
- 16 Which three statements make up Dalton's atomic theory?
- 17 Which, if any, of these statements has to be modified in the light of our present knowledge and why?
- 18 How does Avogadro's hypothesis link Gay-Lussac's law of combining volumes to Dalton's atomic theory?
- 19 Show how Gay-Lussac's law of combining volumes and Avogadro's hypothesis can be used to demonstrate that molecules of hydrogen and chlorine must each contain an even number of atoms and that the formula of water is  $\text{H}_2\text{O}$ , not HO.

# CHAPTER 6

## The Periodic Table

### IN THIS CHAPTER

Historical development of the Periodic Table

Periodic variation in some physical properties

Ionisation energy

Valence and position in the Periodic Table

Further trends:

- metallic character
- reactivity
- nature of compounds
- electronegativity

The Periodic Table shown inside the front cover of this book was first discussed in Section 1.25. There it was introduced as a chart of the elements arranged so that elements with similar properties appeared in vertical columns called groups. In Section 2.13 the relationship between electron configuration of the atoms and the Periodic Table was explained. The relation between chemical bonding and position in the table was discussed in Sections 2.16 and 2.18.

The Periodic Table was not *suddenly* produced in the form shown in those sections. Instead it slowly evolved over nearly a century of efforts to rationalise and systematise the properties of the elements. In this chapter, we shall outline the historical development of this table, and then survey some properties of the elements and see how they vary with position in the table.

### 6.1 HISTORICAL DEVELOPMENT OF THE PERIODIC TABLE

Early in the nineteenth century as more and more elements became known, attempts were made to see patterns in their properties. In 1829 with about 40 elements known, the German chemist, Dobereiner, drew attention to several groups of three elements (which he called *triads*) with very similar properties:

lithium, sodium, potassium  
calcium, strontium, barium  
chlorine, bromine, iodine

In 1864 with over 60 known elements, John Newlands, an Englishman, proposed a '**law of octaves**': when the elements were arranged in order of increasing atomic weight, 'the eighth element starting from a given one is a kind of repetition of the first like the eighth note in an octave of music'. His

'law' identified many similarities among the elements, but erroneously required similarities where none existed. Figure 6.1 lists in order of increasing atomic weights the first twenty-two elements known to Newlands and shows how his law works up to about calcium but then breaks down (because the shaded elements do not resemble those eight places before them).

H	Li (1)	Be (2)	B (3)	C (4)	N (5)	O (6)	F (7)
Na (8)	Mg (9)	Al (10)	Si (11)	P (12)	S (13)	Cl (14)	
K (15)	Ca (16)	Ti (17)	V (18)	Cr (19)	Mn (20)	Fe (21)	

In 1869 Dmitri Mendeleev, a Russian, and Lothar Meyer, a German, independently produced the forerunner of the modern Periodic Table. They arranged the elements in order of increasing atomic weight, and placed elements having similar properties under one another to obtain a table which illustrated what they called the **periodic law**:

Properties of the elements vary periodically with their atomic weights.

The properties available to Mendeleev and Meyer were atomic weights, melting and boiling points (of many but not all elements), atomic volumes<sup>†</sup> and valencies (based mainly on hydrides, oxides and chlorides). In addition classification into metal or non-metal was generally available.

This table was far more successful than that of Newlands, primarily because *Mendeleev recognised that there were probably elements in existence that had not been discovered at that time*. He left gaps in his table in order to place similar elements under one another and proposed that there were undiscovered elements to fill the gaps. And he went further; he predicted the properties of six such elements, three of which he called eka-silicon, eka-boron and eka-aluminium ('eka' meaning 'like'). Subsequently, germanium, scandium and gallium were discovered with properties remarkably similar to his predictions. Figure 6.2 shows Mendeleev's table as originally published in 1872 while Table 6.1 compares Mendeleev's predicted properties for eka-silicon with those of germanium.

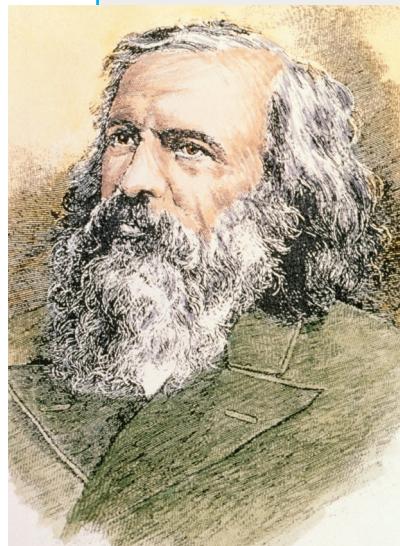


TABELLE II								
REIHEN	GRUPPE I. — R <sup>2</sup> O	GRUPPE II. — RO	GRUPPE III. — R <sup>2</sup> O <sub>3</sub>	GRUPPE IV. RH <sup>4</sup> R <sup>2</sup> O <sub>2</sub>	GRUPPE V. RH <sup>3</sup> R <sup>2</sup> O <sub>5</sub>	GRUPPE VI. RH <sup>2</sup> R <sup>2</sup> O <sub>3</sub>	GRUPPE VII. RH R <sup>2</sup> O <sub>7</sub>	GRUPPE VIII. — R <sup>2</sup> O <sub>4</sub>
I	H = 1							
2	Li = 7	Ba = 9,4	B = 11	C = 12	N = 14	O = 16	F = 19	
3	Na = 23	Mg = 24	Al = 27,3	Si = 28	P = 31	S = 32	Cl = 35,5	
4	K = 39	Ca = 40	= 44	Ti = 48	V = 51	Cr = 52	Mn = 55	Fe = 56, Co = 59, Ni = 59, Cu = 63.
5	(Cu = 63)	Zn = 65	= 68	= 72	As = 75	Se = 78	Br = 80	
6	Rb = 85	Sr = 87	?Yt = 88	Zr = 90	Nb = 94	Mo = 96	= 100	Ru = 104, Rh = 104, Pd = 106, Ag = 108.
7	(Ag = 108)	Cd = 112	In = 113	Sn = 118	Sb = 122	Te = 125	J = 127	
8	Cs = 133	Ba = 137	?Di = 138	?C8 = 140	-	-	-	
9	(-)	-	-	-	-	-	-	
10	-	-	?Er = 178	?La = 180	Ta = 182	W = 184	-	Os = 195, Ir = 197,

FIGURE 6.2 Mendeleev's Periodic Table as published in a German journal in 1872. In German 'Gruppe' means 'group' and 'Reihen' means 'row'. J is the German symbol for iodine. Mendeleev used superscripts in formulae where we now use subscripts.

<sup>†</sup> Atomic volume was atomic weight divided by density of the solid element. Today we would say molar mass divided by density and call it molar volume, with units mL/mol. Density itself does not vary periodically with atomic weight.

FIGURE 6.1  
Some elements arranged to illustrate the success and failure of Newlands' law of octaves. The number in brackets is just the position of the element in the sequence.

Dmitri Mendeleev (1834–1907) was professor of Chemistry at the University of St Petersburg from 1867 to 1890. He developed his Periodic Table in an effort to systematise or organise inorganic chemistry for his students while writing a textbook for them. The table in Figure 6.2 was the second version, the first being published in 1869. He was reputed to have had his hair cut once per year, whether it needed it or not.

**TABLE 6.1 Comparison of properties of germanium: those predicted by Mendeleev in 1871 and those measured in 1886 after its discovery**

	Predicted by Mendeleev	Observed in 1886
atomic weight	72	72.3
density	5.5 g/mL	5.47 g/mL
melting point	very high	960°C
specific heat capacity	$0.31 \text{ J K}^{-1} \text{ g}^{-1}$	$0.32 \text{ J K}^{-1} \text{ g}^{-1}$
formula of oxide	$\text{RO}_2$	$\text{GeO}_2$
density of oxide	4.7 g/mL	4.70 g/mL
formula of chloride	$\text{RCl}_4$	$\text{GeCl}_4$
boiling point of chloride	100°C	86°C

As time passed and atomic weights became more accurately known, some discrepancies emerged in this Periodic Table. In a few cases, in order to fit similar elements under one another, it was necessary to invert the order of atomic weights (argon and potassium, tellurium and iodine, cobalt and nickel). This problem was resolved in 1914 when the British scientist, Henry Moseley, determined what we now call the *atomic number* of each of the elements. He proposed that it, rather than atomic weight, was the basic feature which determined properties. Moseley proposed a **modified periodic law**:

Properties of the elements vary periodically with their atomic numbers.

This law puts argon and potassium, cobalt and nickel, and tellurium and iodine in their right orders.

Once it was recognised that properties were dependent upon the number of protons (atomic number) and hence upon the numbers of electrons, tendencies towards relating the layout of the Periodic Table to electron configuration developed, and so the current form of the table shown inside the front cover was gradually devised.

The Periodic Table developed through attempts to classify the elements—to put similar ones into families or groups. Now that we know that it is electron configuration that is paramount in deciding properties, we can use the Periodic Table and electron configuration to explain why groups of elements have their particular properties.

## Exercises

- \*1 Imagine that you are Mendeleev. You have the data in the table opposite about various elements and you are trying to develop a periodic table. Hydrogen and the transition elements have been omitted from the table (to simplify the exercise).
- a Plot atomic volume against atomic weight. Would you consider that atomic volume varies periodically with atomic weight? Explain your reasoning.
  - b Draw up a periodic table, starting your first period with the first element in the data table and starting a new period at each maximum in your graph (but ignore the small maximum at F). How many complete periods do you have? How many elements are in each period? If there are different numbers in the periods, it is probably because there are elements that have not yet been discovered.

Element <sup>a</sup>	A	B	C	D	E	F	G	J	K	L	M
atomic weight	7	9	11	12	14	16	19	23	24.3	27	28
atomic volume <sup>b</sup>	13.0	4.9	4.4	3.4	13.5	17.4	11.2	23.6	14.0	10.0	12.1
formulae of compounds <sup>c</sup>	AH A <sub>2</sub> O ACl	BH <sub>2</sub> BO BCl <sub>2</sub>	C <sub>2</sub> O <sub>3</sub> CCl <sub>3</sub>	DH <sub>4</sub> DO <sub>2</sub> DCl <sub>4</sub>	EH <sub>3</sub> ECl <sub>3</sub>	FH <sub>2</sub> FCl <sub>2</sub>	GH G <sub>2</sub> O	JH J <sub>2</sub> O JCl	KH <sub>2</sub> KO KCl <sub>2</sub>	L <sub>2</sub> O <sub>3</sub> LCl <sub>3</sub>	MH <sub>4</sub> MO <sub>2</sub> MCl <sub>4</sub>
Element <sup>a</sup>	N	P	Q	R	S	T	U	V	W	X	
atomic weight	31	32	35.5	39	40	75	79	80	86	87.5	
atomic volume <sup>b</sup>	17.0	15.6	17.4	46.0	26.2	13.0	16.4	19.8	55.8	34.0	
formulae of compounds <sup>c</sup>	NH <sub>3</sub> N <sub>2</sub> O <sub>3</sub> NCl <sub>3</sub>	PH <sub>2</sub> Q <sub>2</sub> O PCl <sub>2</sub>	QH R <sub>2</sub> O RCI	RH SO SCl <sub>2</sub>	SH <sub>2</sub> T <sub>2</sub> O <sub>3</sub> TCI <sub>3</sub>	UH <sub>2</sub>	VH		W <sub>2</sub> O WCl	XH <sub>2</sub> XO XCl <sub>2</sub>	

a The elements have been given simple alphabetical labels. These labels are not the chemical symbols of the elements.

b The ‘atomic volume’ of Mendeleev’s day is what we today would call the molar volume, the volume of one mole of the solid element. Units are mL mol<sup>-1</sup>.

c In this row H, O and Cl do have their normal chemical meaning—hydrogen, oxygen and chlorine.

- c Use the formulae of hydrides, oxides and chlorides to align elements so that those with similar formulae fall under one another, leaving gaps for undiscovered elements where necessary.
  - d How many elements are there, or should there be, in each period of your table? Is this in agreement with the modern Periodic Table? Offer an explanation.
  - e Bearing in mind that we have omitted hydrogen and the transition elements, identify the elements in your table; that is, assign the correct chemical symbol to each of the alphabetical labels used in the data table.
- 2 When Mendeleev developed his Periodic Table, indium was thought to have an atomic weight of 76 (based on the observation that the oxide of assumed formula, InO, contained 82.5% indium). Between which pair of elements did this place indium? Indium is definitely a metal, so why did Mendeleev think that this positioning was wrong? He proposed that indium oxide was in fact In<sub>2</sub>O<sub>3</sub>. On this basis, what would be the atomic weight of indium? Between which pair of elements would this locate indium? Why is this a better position for indium? In this way Mendeleev corrected several atomic weights.
- 3 In Mendeleev’s Periodic Table, no noble gases were present and no gaps were left for them. Why did this not cause any problems for his table?

## 6.2 PERIODIC VARIATION IN SOME PHYSICAL PROPERTIES

Some physical properties neatly demonstrate the periodicity inherent in the Periodic Table. In this section we shall consider atomic radius (which is related to atomic volume and density), melting point and boiling point. In the next section we shall discuss ionisation energy.

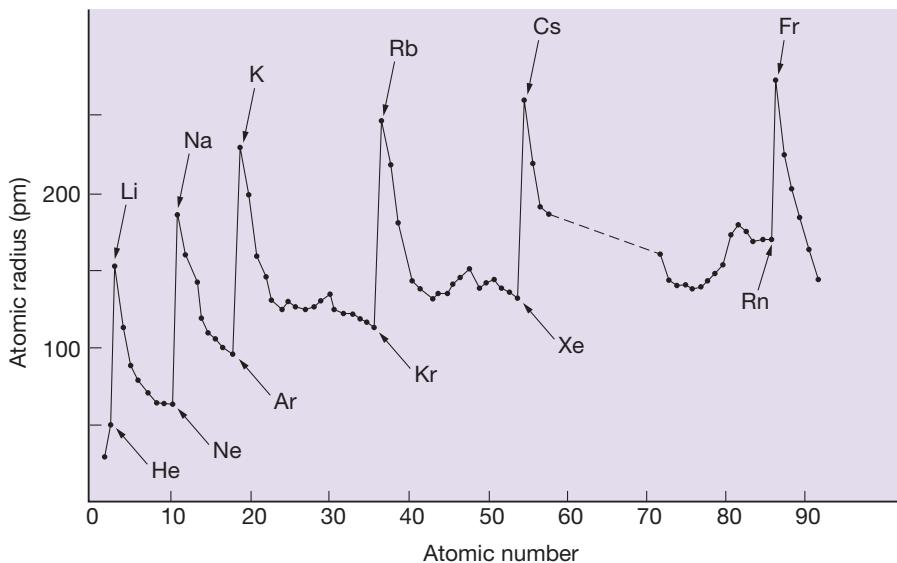
### Atomic radius

When **atomic radius** is plotted against atomic number, as in Figure 6.3, the curve shows a distinctly periodic nature. The atomic radius passes through a set of

sharp maxima corresponding to the alkali metals (Li, Na, K, Rb, Cs—Group 1). The minima occur at the noble gases (Ne, Ar, Kr, Xe, Rn). This means that *atomic radius decreases from left to right across any period of the table*. Values in Table 6.2 illustrate this.

Atomic radius increases in going down any group of the table.

**FIGURE 6.3**  
Atomic radius of the elements as a function of atomic number



**TABLE 6.2** Atomic radii (in picometres)<sup>a</sup> of some elements

		H 37							He 50
Li 152	Be 112		B 88	C 77	N 70	O 66	F 68	Ne 70	
Na 186	Mg 160		Al 143	Si 118	P 110	S 102	Cl 99	Ar 94	
K 231	Ca 197				As 123	Se 116	Br 114	Kr 109	
Rb 244	Sr 215						I 133	Xe 130	
Cs 262	Ba 217							Rn 140	

*a* 1 picometre (pm) =  $10^{-12}$  metre.

## Melting point

Figure 6.4 shows the variation of melting point with atomic number. Again a periodic behaviour results, with the curve passing through a series of minima corresponding to the noble gases. The maxima are less well defined: the first two correspond to the elements carbon and silicon while the others correspond to transition metals.

## Boiling point

When boiling points of elements are plotted against atomic numbers, a graph similar to Figure 6.4 results: boiling point passes through a series of minima

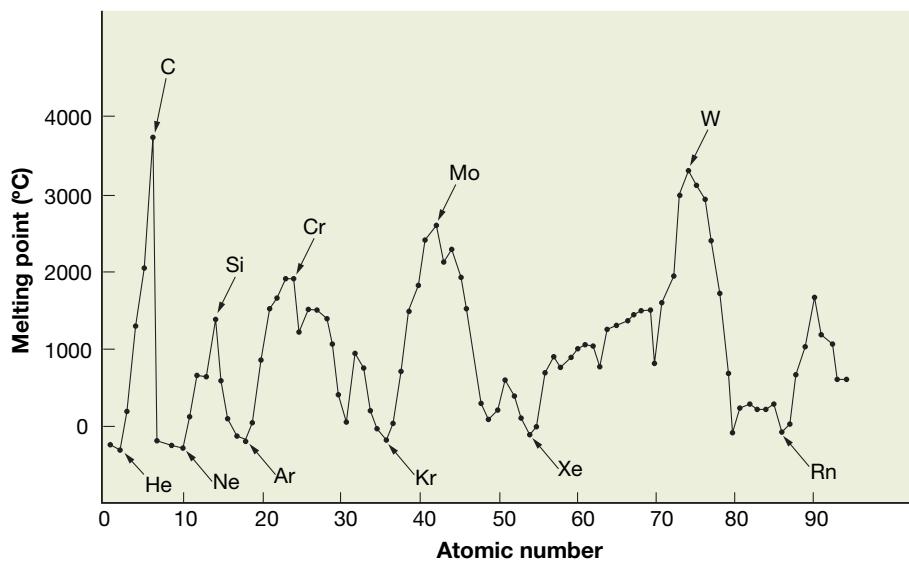


FIGURE 6.4  
Melting point of elements as a function of atomic number

corresponding to the noble gases with maxima occurring roughly half-way between pairs of minima.

### 6.3 IONISATION ENERGY

First ionisation energy,  $I_A$ , was introduced in Section 4.12. You will recall that it is the energy required to remove an electron from a gaseous atom of an element; that is, it is the energy needed to bring about the process:



where M is any element. Ionisation energy is generally measured in kilojoules per mole (kJ/mol). There are actually several ionisation energies for each element, called the first, second, third and so on ionisation energies.

The *first ionisation energy* of an element is the energy required to remove one electron from the neutral atom. The *second ionisation energy* is the energy required to remove a second electron.

The *first ionisation energy* is for  $M(g) \rightarrow M^+(g) + e^-$

The *second ionisation energy* is for  $M^+(g) \rightarrow M^{2+}(g) + e^-$

The *third ionisation energy* is for  $M^{2+}(g) \rightarrow M^{3+}(g) + e^-$

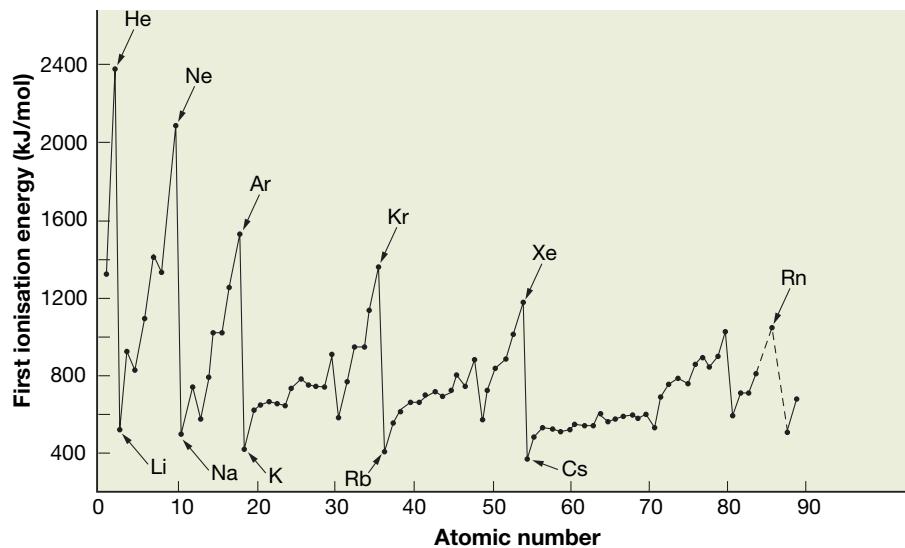
and so on.

The second ionisation energy is always greater than the first. This is because it requires more energy to remove a negative electron from a positive species than from a neutral species. This is due to extra electrostatic attraction. Similarly, the third ionisation energy is greater than the second.

*Ionisation energies provide strong evidence for the periodic law.* In addition they provide striking confirmation of the ‘drive to noble gas configurations’ discussed in Section 2.12. In Figure 6.5, first ionisation energy is plotted against atomic number. Periodicity is again obvious. The minimum values of  $I_A$  are for the alkali metals: it is relatively easy to remove an electron from Li, Na, K, Rb and Cs. The noble gases have the maximum values of  $I_A$ : it is quite difficult to remove an electron from these elements. This confirms that noble gas configurations are extremely stable. Elements with one or two extra electrons tend to lose them to form noble gas configurations.

FIGURE 6.5

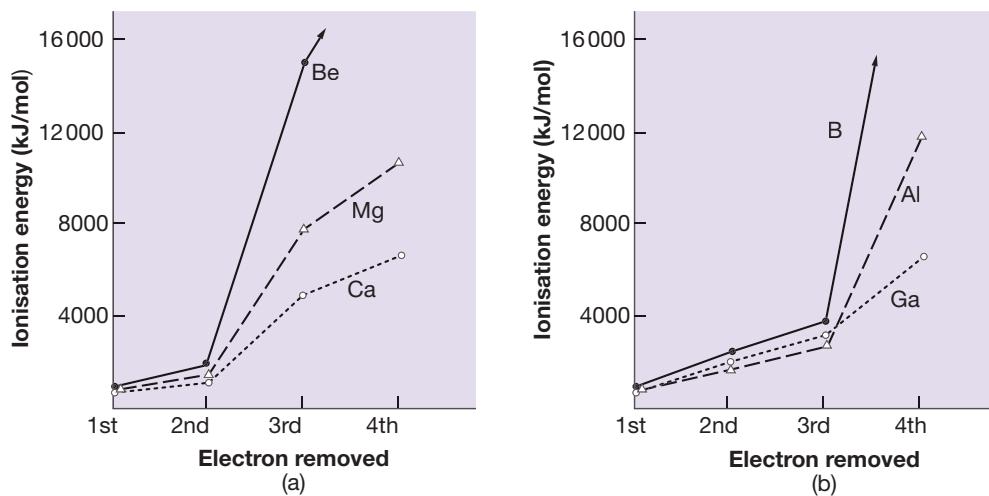
First ionisation energy as a function of atomic number



In Figure 6.6, successive ionisation energies are plotted against the number of the electron being removed for some elements in Groups 2 and 3. In Figure 6.6(a), we note that it is relatively easy to remove one or two electrons from Be, Mg and Ca, but that to remove the third requires vastly more energy. In Figure 6.6(b), it is relatively easy to remove 1, 2 and 3 electrons, but a far larger amount of energy is needed to remove the fourth electron. Table 6.3 shows a similar situation for Group 1: it is easy to remove one electron but quite difficult to remove a second electron. These observations confirm our conclusion in Section 2.16 that *elements in Groups 1, 2 and 3 tend to lose electrons to obtain the electron configuration of the nearby noble gas*. Once one, two or three electrons have been removed from elements in Groups 1, 2 and 3 respectively, the ions have noble gas configurations and it is extremely hard to remove further electrons.

FIGURE 6.6

Successive ionisation energies for (a) Group 2 and (b) Group 3



Elements with low ionisation energies readily form positive ions and therefore such elements form ionic compounds ( $\text{Na}^+$ ,  $\text{Ca}^{2+}$ ,  $\text{Al}^{3+}$ ).

*In going across any period of the table, the first ionisation energy increases, showing that in moving from left to right, there is less tendency to lose electrons.*

**TABLE 6.3 Successive ionisation energies<sup>a</sup> for Group 1**

	<b>First</b>	<b>Second</b>	<b>Third</b>	<b>Fourth</b>
Li	526	7300	11800	
Na	504	4570	6920	9550
K	425	3080	4400	5880
Rb	410	2660	3900	5100
Cs	380	2430	3400	4900

*a* Units are kJ/mol.

In going down any group of the table, ionisation energy decreases; data in Table 6.4 clearly show this.

This means that metal atoms lose electrons to form cations less readily (that is, become less reactive) as we move from left to right across a period (for example, Na, Mg, Al) and lose electrons more readily (become more reactive) as we go down a group (for example Li, Na, K).

**TABLE 6.4 First ionisation energies (in kJ/mol) for some elements**

		H 1320						He 2380
Li 526	Be 905		B 810	C 1090	N 1410	O 1320	F 1690	Ne 2090
Na 504	Mg 740		Al 580	Si 790	P 1020	S 1000	Cl 1260	Ar 1526
K 425	Ca 600				As 953	Se 950	Br 1150	Kr 1360
Rb 410	Sr 560					I 1020	Xe 1180	
Cs 380	Ba 510							Rn 1040

This decrease in  $I_A$  in going down Groups 1 and 2 means that the elements become *more reactive*: they form  $M^+$  or  $M^{2+}$  more easily. For example, lithium reacts fairly slowly with water, sodium reacts quite vigorously and potassium reacts explosively. Calcium reacts more slowly than barium does, and magnesium does not react with cold water at all (Section 4.9).

Trends in first ionisation energies across and down the Periodic Table, and changes in successive ionisation energies for the one element, supply strong evidence that the drive towards noble gas configuration is the basis for chemical bonding.

## Exercises

- 4 a** Using the data below, on the one set of axes draw two graphs of density versus period number, one for Group 1 elements and the other for Group 2 elements. Include period number 7 on the graph scale. Draw appropriate curves or lines through or between the points. Based upon your graphs, what summarising



statement, if any, can you make about the way density changes as we go down a group of the Periodic Table?

period number	2	3	4	5	6
Group 1 element	Li	Na	K	Rb	Cs
density (g/mL)	0.53	0.97	0.86	1.53	1.87
Group 2 element	Be	Mg	Ca	Sr	Ba
density (g/mL)	1.85	1.74	1.55	2.6	3.5

- b** Use your graphs to estimate the density of francium and radium, the Period 7 members of Groups 1 and 2 respectively. Comment on the accuracy or reliability of these estimates.
- 5 \*a** Use the data in the table in Exercise 1 and your results from Exercise 1(e) to complete the table below for the elements in that exercise that belong to Groups 1, 2 and 7. What appeared to be Periods 1, 2, 3 and 4 in Exercise 1 (because H and He were omitted) are normally called Periods 2, 3, 4 and 5.

Period	Group 1		Group 2		Group 7	
	Element	Molar volume <sup>a</sup>	Element	Molar volume <sup>a</sup>	Element	Molar volume <sup>a</sup>
2						
3						
4						
5						

<sup>a</sup> Molar volume is a better name for what Mendeleev called atomic volume.

- b** Plot molar volume against period number for the elements in Groups 1, 2 and 7 (three separate curves on the one graph). Draw suitable lines or curves through the sets of points and summarise how molar volume varies with period number.
- 6** Based on successive ionisation energies in Figure 6.6 and Table 6.3, what charges do you expect the cations of the following elements to carry? **(a)** Be **(b)** Rb **(c)** Ga
- 7** Using only a Periodic Table and basing your answer on general trends (without consulting Tables 6.2 or 6.4), arrange the elements in the following sets in order of:
  - a** increasing atomic radius
  - b** increasing ionisation energy

**(i)** Ca, Ba, Mg    **(ii)** P, Al, Cl    **(iii)** Al, Rb, Na, S    **(iv)** Br, As, F

Answer Exercises 8, 9 and 10 without consulting any tables or diagrams.

- 8** Four elements have the following successive ionisation energies (in kJ/mol):

A	500	4600	6900	9500
B	600	1150	4900	6500
C	2100	3960	6100	9400
D	380	2400	3400	4400

To which group of the Periodic Table does each belong?

- 9** Four elements of atomic numbers 17, 18, 19 and 20 have first ionisation energies of 1500, 600, 400 and 1300 kJ/mol (listed in random order). Which ionisation energy belongs to which element?

\***10** Three elements of Group 1 have first ionisation energies 380, 530 and 430 kJ/mol. List the elements in order of increasing atomic number.

**11** Three elements, X, Y and Z, from the same period of the Periodic Table have one, two and seven electrons in their outer energy levels respectively. Their first ionisation energies (not necessarily in the same order) are 1260, 425 and 740 kJ/mol. Their atomic radii (again in random order) are 160, 230 and 100 pm. Assign the ionisation energies and atomic radii to the elements, X, Y and Z.

## 6.4 VALENCY AND POSITION IN THE PERIODIC TABLE

Valencies (combining powers) of elements are related to position in the Periodic Table as was discussed in Sections 2.16, 2.18, 3.8 and 3.11.

The most common valency of an element is its group number (if it is in Groups 1 to 4) or 8 minus its group number (if it is in Groups 5 to 7).

As the noble gases rarely form compounds, their common valency is zero. Table 6.5 shows the common valencies of the elements in the main groups of the table.

It is not possible to deduce the valency of the transition metals from their position in the Periodic Table. Simple compounds of these metals are ionic. Most of the transition metals form an  $M^{2+}$  cation as well as cations with other charges; examples are  $Cu^{2+}$  and  $Cu^+$ ,  $Fe^{2+}$  and  $Fe^{3+}$ , and  $Cr^{2+}$  and  $Cr^{3+}$ . However, zinc forms only  $Zn^{2+}$ , while silver mainly forms  $Ag^+$  compounds.

**TABLE 6.5 Common valencies of elements by group of the Periodic Table**

Group	Common valency	Type of compound
1 (Li, Na, K, Rb, Cs)	1	ionic
2 (Be, Mg, Ca, Sr, Ba)	2	ionic <sup>a</sup>
3 (B and Al only)	3	B covalent Al ionic <sup>b</sup>
4 (C, Si) (Ge, Sn, Pb)	4 <sup>c</sup> 2 and 4	covalent <sup>d</sup> ionic with valency of 2; can be covalent with valency of 4
5 (N, P)	3 or 5 <sup>e</sup>	covalent <sup>f</sup>
6 (O) (S)	2 <sup>g</sup> 2, 4, 6	ionic or covalent ionic (only for valency of 2) or covalent
7 (F, Cl, Br, I)	1 <sup>h</sup>	ionic or covalent

<sup>a</sup> Be can form some covalent compounds such as  $BeCl_2$ .

<sup>b</sup> Al forms a few covalent compounds such as  $Al_2Cl_6$ .

<sup>c</sup> Except in CO.

<sup>d</sup> C forms the carbide ion,  $C_2^{2-}$  as in calcium carbide.

<sup>e</sup> N shows valencies of 1, 2 and 4 also, in the oxides,  $N_2O$ ,  $NO$ ,  $NO_2$ .

<sup>f</sup> With a few metals N can form nitride ions,  $N^{3-}$ , as in magnesium nitride  $Mg_3N_2$ .

<sup>g</sup> O shows a valency of 1 in peroxides such as hydrogen peroxide,  $H_2O_2$ , and barium peroxide  $BaO_2$ .

<sup>h</sup> Cl, Br and I show higher valencies in oxides and oxyacids such as  $ClO_2$ ,  $HBrO_3$ ,  $HIO_4$ .



## Exercises

- 12 a** What value do you expect for the common valency of each of the following elements: strontium, arsenic, selenium, indium, bromine, caesium, boron, barium, antimony, germanium, gallium, iodine?
- b** Which of these elements do you expect to form:
- i** mainly ionic compounds?
  - ii** mainly covalent compounds?
  - iii** some ionic and some covalent compounds?
- 13 a** Knowing the formulae for the sodium compounds, write down the formulae for rubidium oxide, caesium chloride, lithium sulfate, rubidium nitrate, caesium sulfate and lithium sulfide.
- b** Similarly, knowing the formulae for magnesium and calcium compounds, write down the formulae for barium nitrate, radium oxide, strontium oxide, beryllium sulfate, radium chloride, strontium bromide and barium fluoride.
- \*c** Give formulae for the oxides of caesium, barium, thallium, lead and bismuth.
- 14 a** Based on your knowledge of ammonia, what formulae would you expect for the hydrides of phosphorus (called phosphine), arsenic (arsine) and antimony (stibine)?
- \*b** What formulae do you expect for the hydrides of Se and Te? For each, state whether you expect it to be ionic or covalent. Give your reason.
- 15** Give the formulae and describe the type of compound (ionic or covalent) you would expect for:
- a** the iodides of Group 1 elements, Li, Na, K, Rb, Cs
  - b** the oxides of Group 2 elements, Be, Mg, Ca, Sr, Ba
  - c** the compounds of magnesium with the Group 7 elements, F, Cl, Br, I
  - \*d** the chlorides of Group 4 elements, C, Si, Ge, Sn, Pb
  - \*e** the hydrides of Group 6 elements, O, S, Se, Te

## 6.5 FURTHER TRENDS

### Metallic character

In Section 1.23 the division of the elements into metals and non-metals was discussed. We saw in Table 1.8, reproduced here as Table 6.6, that there were 18 non-metals. They are located in the top right-hand corner of the Periodic Table. In addition there were six elements which we called semi-metals: boron, silicon, germanium, arsenic, antimony and tellurium. The remaining 70 or so elements are metals.

As we go from left to right across a period of the Periodic Table, there is a gradual change from metal to non-metal. Going across Period 3, Na, Mg and Al are metals, Si is a semi-metal, and P, S, Cl and Ar are non-metals.

Going down groups in the middle of the table, there is a trend from non-metal to metal. In Group 4, C is a non-metal, Si and Ge are semi-metals, while Sn and Pb are metals. Similarly in Group 5, N and P are non-metals, As and Sb are semi-metals, and Bi is a metal.

In groups near the edges of the Periodic Table, the elements are either all metals (near the left-hand edge, Groups 1 and 2) or all non-metals (near the right-hand edge, Groups 7 and 0).

**TABLE 6.7 Division of the elements into metals and non-metals**

H																		He			
Li Be										B C N O F Ne											
Na Mg										Al Si P S Cl Ar											
K Ca Sc Ti V Cr Mn Fe Co Ni Cu Zn Ga Ge As Se Br Kr										In Sn Sb Te I Xe											
Rb Sr Y Zr Nb Mo Tc Ru Rh Pd Ag Cd In Sn Sb Te I Xe										Cs Ba La Hf Ta W Re Os Ir Pt Au Hg Tl Pb Bi Po At Rn											
Fr Ra Ac Rf Db																					

Elements to the left of the blue shaded zone are metals; elements to the right (in the pink shaded zone) are non-metals. Elements within the blue shaded zone have properties intermediate between metals and non-metals and are therefore hard to categorise; they are called **semi-metals**. Hydrogen is usually classed as a non-metal though a case can be made for calling it a semi-metal.

## Reactivity

We saw in Section 4.12 that reactivity of metals correlates well with first ionisation energy: the lower the first ionisation energy, the greater the reactivity of the metal. This is because both reactivity of a metal and ionisation energy are measures of the ease with which the metal loses electrons. We can therefore say:

For metallic elements reactivity decreases from left to right across a period of the table (for example Na, Mg, Al) and increases from top to bottom down a group (Be, Mg, Ca, Sr).

Reactivity of non-metals is less clear-cut, largely because there are at least two types of reactivity for non-metals. There is formation of anions or gain of electrons (as in O, S, Cl, Br going to  $O^{2-}$ ,  $S^{2-}$ ,  $Cl^-$ ,  $Br^-$ ) and also formation of covalent compounds (sharing of electrons) as in O, S, P, Cl going to  $H_2O$ ,  $SO_2$ ,  $PCl_3$ ,  $HCl$ .

For formation of anions, reactivity of non-metals increases from left to right across a period (for example,  $N < O < F$ ) and decreases from top to bottom down a group (for example,  $F > Cl > Br > I$ ). For formation of covalent compounds the reactivity generally follows the same trends, though the trends are less clear-cut and there are exceptions.

## Nature of compounds

We saw in Chapter 2 that compounds could be either ionic (made up of ions, such as  $Na^+Cl^-$  and  $Mg^{2+}O^{2-}$ ), covalent molecular (made up of neutral molecules such as  $H_2O$ ,  $NH_3$ ,  $CO_2$ ), or covalent lattices (such as  $SiO_2$ ).

Metals, when they react with non-metals, tend to form ionic compounds ( $MgCl_2$ ,  $CuS$ ). When non-metals combine with other non-metals, they form covalent compounds (molecules such as  $SO_2$ ,  $PBr_3$ ,  $N_2O$  or lattices such as  $SiC$ ).

As there is a trend from metal to non-metal across the Periodic Table, we expect some trend from ionic to molecular compounds across the table. In Period 3, sodium and magnesium (always) and aluminium (mostly) form ionic compounds. Silicon and phosphorus almost always form covalent compounds. Sulfur and chlorine form ionic compounds with metals ( $FeS$ ,  $CaF_2$ ) and covalent ones with non-metals ( $H_2S$ ,  $PCl_3$ ).

Just as there is a trend from non-metal to metal going down Groups 4 and 5, so too there is a trend from covalent to ionic compounds. In Group 4, C and

Si always form covalent compounds ( $\text{CH}_4$ ,  $\text{CO}_2$ ,  $\text{SiCl}_4$ ). Sn and Pb commonly form ionic compounds ( $\text{SnCl}_2$ ,  $\text{SnS}$ ,  $\text{PbO}$ ,  $\text{PbI}_2$ ). In Group 5, N and P almost always form covalent compounds, while bismuth generally forms ionic ones.

There is a physical quantity available that can help us decide whether two elements combine to form an ionic or covalent compound. It is called electronegativity.

## Electronegativity

The **electronegativity** of an element is a measure of the ability of an atom of that element to attract bonding electrons towards itself when it forms compounds.

Electronegativity values for some elements are given in Table 6.7. They range from about 0.7 (for Fr) to 4.0 (for F). As Exercise 18 will show, electronegativity increases from left to right across a period (omitting the noble gases) and decreases from top to bottom down a group.

We can decide whether two elements form an ionic or covalent compound by using the following guideline:

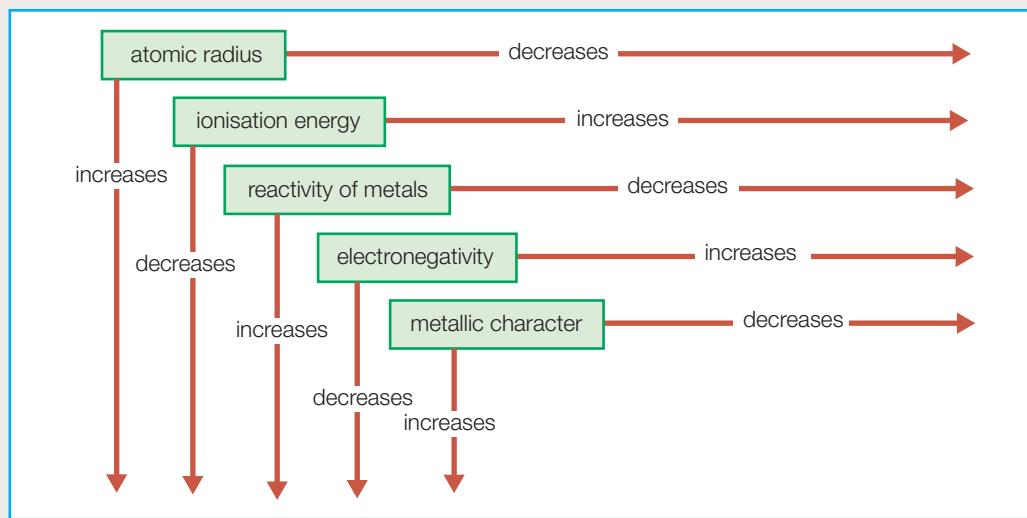
If the difference in electronegativities of two elements is greater than 1.5, the elements will form an ionic compound; otherwise the compound will be covalent.

**TABLE 6.7 Electronegativities for some elements**

Element	H	He	Li	Be	B	C	N	O	F	Ne	Na
atomic number	1	2	3	4	5	6	7	8	9	10	11
electronegativity	2.20	0	0.98	1.57	2.04	2.55	3.04	3.44	3.98	0	0.93
Element	Mg	Al	Si	P	S	Cl	Ar	K	Ca	Sc	Ti
atomic number	12	13	14	15	16	17	18	19	20	21	22
electronegativity	1.31	1.61	1.90	2.19	2.58	3.16	0	0.82	1.00	1.36	1.54
Element	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As
atomic number	23	24	25	26	27	28	29	30	31	32	33
electronegativity	1.63	1.66	1.55	1.83	1.88	1.91	1.90	1.65	1.81	2.01	2.18
Element	Se	Br	Kr	Rb	Sr	Y	Zr	Nb	Te	Cs	Ba
atomic number	34	35	36	37	38	39	40	41	52	55	56
electronegativity	2.55	2.96	0	0.82	0.95	1.22	1.33	1.6	2.1	0.79	0.89

## 6.6 SUMMARY OF TRENDS

Figure 6.7 summarises the way the various properties we have been discussing vary across periods and down groups of the Periodic Table.



**FIGURE 6.7**  
The way some properties vary across and down the Periodic Table

## Exercises

16 Which of the following lists contains only metals?

- a zinc, aluminium, silicon, bismuth
- b titanium, magnesium, lead, palladium
- c strontium, rhodium, selenium, antimony
- d iron, calcium, caesium, boron



17 For each of the elements of Period 3 of the Periodic Table, Na to Ar:

- a Give the formula for the hydride and state whether it is ionic or covalent.
- \*b Give the formula for the chloride and state whether it is ionic or covalent.

18 Use data from Table 6.7 for this exercise.

- \*a Draw a graph of electronegativity versus atomic number and use it to decide whether electronegativity varies periodically with atomic number. Explain your reasoning.
- \*b For Groups 1 and 2, plot electronegativity against period number on the one graph. Draw a suitable curve or lines through the points from each group. Summarise the way that electronegativity varies with period number.
- c Estimate electronegativities for francium and radium. Comment upon the reliability of your estimates.

19 a Hydrogen forms compounds with the elements of Group 5 of the Periodic Table. What formulae do you expect for the hydrides of N, P and As? Do you expect these compounds to be ionic or covalent? Explain why.  
 b Chlorine forms compounds with the elements of Groups 2 and 5. What formulae do you expect for the chlorides of Mg, Ca, Sr, N, P and As? Which (if any) of these do you expect to be ionic and which (if any) to be covalent? Explain your reasoning.

## WEBSITES

The following two sites can be used to display graphs of properties of elements as functions of atomic numbers, of period number and of group number.

<http://www.webelements.com>

(click on an element to get its basic properties. Then click on a property in the left-hand column to get other properties and to be able to display graphs. Click on view to choose the type of graph, then go. You can display a graph of the chosen property for any particular group or period by selecting it in the left-hand column)



<http://www.chemlab.pc.maricopa.edu/periodic/periodic.html>

(can display a graph of a chosen property of a range of elements (specified range of atomic numbers); scroll down the opening page to *Graph*, select the property in the drop-down menu and choose the range of atomic numbers, then click *Graph*. You need to remember which property you are graphing here, because its name does not appear on the graph!)

## Important new terms

You should know the meaning of the following terms:

atomic radius (p. 157)  
electronegativity (p. 166)

law of octaves (p. 154)  
(modified) periodic law (p. 156)  
periodic law (p. 155)  
picometre (p. 158)

# CHAPTER 6

## Test yourself

- 1 Explain the meaning of each of the items in the ‘Important new terms’ section above.
- 2 Why was Mendeleev’s table more successful than that of Newlands?
- 3 What predictions did Mendeleev make from his table? What was the consequence of these being found to be correct?
- 4 What difficulty was there with Mendeleev’s Periodic Table?
- 5 Who overcame this difficulty and how?
- 6 Give three physical properties which vary periodically with atomic number.
- 7 For two of the properties in your answer to Question 6, sketch the graph you would get if you plotted each against atomic number (up to 20), and indicate for each property the elements which correspond to the maxima or minima in each curve.
- 8 How do atomic radii vary in going down any group of the Periodic Table?
- 9 How do values of ionisation energy confirm the theory that in forming compounds, elements try to obtain noble gas configurations?
- 10 How do first ionisation energies vary in going down any group of the Periodic Table?
- 11 Where in the Periodic Table do the non-metals lie? Which elements are classed as semi-metals?
- 12 Name four non-metals which are solids at room temperature.
- 13 How is the common valency of an element related to its position in the Periodic Table?
- 14 Classify each of the elements in Period 3 (Na to Ar) as metal, semi-metal or non-metal.
- 15 How does metallic character change as we go down Group 4 (C to Pb)?

- 16** In the Periodic Table how does electronegativity vary **(a)** across a period, and **(b)** down a group?
- 17** How does electronegativity help us decide whether the compound formed by two elements is ionic or covalent?
- 18** How does reactivity of **(a)** metals **(b)** non-metals vary **(i)** across a period of the Periodic Table and **(ii)** down a group?

# EXTENDED RESPONSE EXAM-STYLE QUESTIONS FOR MODULE 2

For explanations about these New South Wales HSC exam-style questions, see p. 84.

## MARKS

- 1 List two current-day uses for each of the metals, iron (or mild steel), aluminium and copper. **Explain** how the properties of each metal make it particularly suitable for the uses listed. 6
- 2 **Assess** the importance of alloy-making in the use of metals. Include three examples, at least one from ancient times (say, before 1000 BCE) and at least one from modern times (say, from 1900 CE onwards). 6
- 3 Historians and anthropologists date the Iron Age from about 700 BCE to about 1 CE. **Assess** the suitability of this end date from a chemist's viewpoint. 3
- 4 **Explain** why there are more metals available for people to use today than there were 200 years ago. Illustrate your answer with specific examples. 3
- 5 For one metal that has come into widespread use since about 1940, **analyse** the reasons for this happening. 3
- 6 **Discuss** the suitability of iron (in the form of mild steel) and aluminium for making car bodies and drink containers. 4
- 7 **Outline** the chronology of the use of metals through history and relate this chronology to the ease of obtaining the metals. 7
- 8 **Describe** an experiment that you have performed to **construct** an activity series for some common metals. Mention at least five metals that you used and **explain** how you used your results to determine your activity series. 5
- 9 **Describe** two modern-day situations where chemical reactivity is an important factor in deciding which metal to use. **Assess** the importance of reactivity relative to other factors in making the decision. 4
- 10 **Assess** the usefulness of first ionisation energy in estimating the relative reactivities of a set of metals. 2
- 11 **Evaluate** the contributions of Dobereiner and Newlands in developing the Periodic Table. 4
- 12 **Outline** the historical development of the Periodic Table from the early 1800s to the present day. 5
- 13 **Describe** and **evaluate** the contribution of Mendeleev to the development of the Periodic Table. Include mention of the types of data available to Mendeleev. 6
- 14 **Assess** the usefulness of the Periodic Table for working out the valency of a particular element. 3
- 15 In chemical contexts what is a mole? **Explain** why it is useful. 2

- 16** **Describe** an experiment you performed to extract copper from a carbonate ore. Include mention (and a diagram) of the equipment that you used.  
Include relevant chemical equations. **5**
- 17** **Assess**, with examples, the relative importance to the commercial availability of a metal of **(a)** abundance of the metal in the Earth's crust and **(b)** the ease (or difficulty) of extracting the metal. **4**
- 18** **Justify** the present-day worldwide emphasis upon recycling metals. **3**
- 19** **Propose** two reasons why extracting a metal from a particular mine site can be economic at one time but uneconomic at another. **2**
- 20** **Explain** why there is much more emphasis on recycling aluminium than on recycling steel. **3**

# REVISION TEST FOR MODULE 2

**Total marks: 50 Suggested time: 90 minutes**

## MULTIPLE CHOICE QUESTIONS

Select the alternative **a**, **b**, **c**, or **d** which best answers the question.

- 1 Titanium was not widely used until the last half century or so because:
  - a It occurs in the Earth's crust in only very small amounts.
  - b It was not needed until the world's supplies of tungsten began to run out.
  - c It was difficult to develop a method of extracting the metal from its ores.
  - d It is quite toxic and needs special care in its handling.
- 2 Which of the following statements is correct?
  - a The molar mass of zinc is 65.4.
  - b The atomic weight (relative atomic mass) of zinc is 65.4 g.
  - c The mass of an atom of zinc is  $1.09 \times 10^{-22}$  g.
  - d The mass of a mole of atoms of zinc is  $65.4 \times 6.02 \times 10^{23}$  g.
- 3 The order of decreasing chemical reactivity of the four metals listed is:

a Cu, Fe, Mg, Li	c Li, Mg, Cu, Fe
b Li, Fe, Mg, Cu	d Li, Mg, Fe, Cu
- 4 One component of Portland cement is dicalcium silicate,  $(\text{CaO})_2\text{SiO}_2$ . The formula weight of this compound is:

a 132.2	c 172.3
b 200.4	d 160.3
- 5 Of the metals, Mg, Al, Zn, Fe, Pb, Cu, the ones that *cannot* be extracted by heating with carbon or carbon monoxide are:

a Mg and Al	c Al and Zn
b Mg, Al and Zn	d Fe, Pb and Cu
- 6 Based on positions in the Periodic Table, the most likely formula for the compound formed between strontium and fluorine is:

a $\text{SrF}$	c $\text{Sr}_2\text{F}_2$
b $\text{SrF}_2$	d $\text{Sr}_2\text{F}$
- 7 The only correctly balanced equation in the following list is:

a $2\text{Li}(s) + \text{O}_2(g) \rightarrow \text{Li}_2\text{O}(s)$	c $\text{Al}(s) + 2\text{H}^+ \rightarrow \text{Al}^{3+} + \text{H}_2(g)$
b $\text{Ca}(s) + 2\text{H}_2\text{O}(l) \rightarrow \text{Ca}(\text{OH})_2(aq) + \text{H}_2(g)$	d $2\text{Fe}(s) + 3\text{O}_2(g) \rightarrow \text{Fe}_2\text{O}_3(s)$
- 8 A group of the Periodic Table in which there is a marked increase in metallic character going down the group is:

a Group 4	c Group 7
b Group 2	d Group 1
- 9 A common oxide of titanium contains 40.1% oxygen. The empirical formula of this oxide is:

a $\text{TiO}$	c $\text{Ti}_2\text{O}$
b $\text{TiO}_2$	d $\text{Ti}_2\text{O}_4$

- 10** The equation for reducing iron(III) oxide with heated carbon (coke) is:



The mass of carbon needed to produce 100 g of iron is:

- a** 311 g      **c** 21.4 g  
**b** 14.3 g      **d** 32.2 g

#### SHORT ANSWER AND EXTENDED RESPONSE QUESTIONS

The number of marks for each part of each question is shown at the right-hand end of the question. Marks allocated are based on the HSC scale of 1 mark per 1.8 minute.

- |   | MARKS    |
|---|----------|
| <b>11</b> Outline the procedure for extracting copper from a common sulfide ore.<br>Include at least one chemical equation.   | <b>4</b> |
| <b>12</b> Describe what you would observe when a small sample of <b>(a)</b> lithium<br><b>(b)</b> magnesium <b>(c)</b> copper was added to a test tube containing<br><b>(i)</b> cold water <b>(ii)</b> dilute hydrochloric acid. Write equations for any<br>chemical reactions that occur.  | <b>5</b> |
| <b>13</b> The electron configurations of three elements are A 2, 8, 8; B 2, 8, 3;<br>and C 2, 8, 7. The atomic radii and first ionisation energies of these elements<br>(not necessarily in the same order) are 143, 94, 99 pm and 1526, 580,<br>1260 kJ/mol.<br><br><b>a</b> Identify the elements A, B and C. <b>1</b><br><b>b</b> Assign the correct atomic radii and ionisation energies to A, B and C.<br>Explain the basis for your choices. <b>2</b> |          |
| <b>14</b> Nickel dimethylglyoxime is one of only a few molecular compounds<br>of metals. It has the molecular formula $\text{NiC}_8\text{H}_{14}\text{N}_4\text{O}_4$ .<br><br><b>a</b> Calculate the molecular weight of nickel dimethylglyoxime. <b>1</b><br><b>b</b> What is the total number of atoms in a molecule of nickel<br>dimethylglyoxime? <b>1</b><br><b>c</b> How many nitrogen atoms are there in 1.00 g nickel dimethylglyoxime? <b>2</b>   |          |
| <b>15</b> Describe an experiment that you have performed to measure the<br>percentages of metal and non-metal in a binary compound and hence<br>to determine the empirical formula of the compound. List the actual<br>measurements that you made and explain how you used them to determine<br>the percentages. What was the major source of error in your experiment<br>and how did you try to minimise it? <b>5</b>                                      |          |
| <b>16</b> Outline the steps involved in the recycling of aluminium and explain why<br>there is much more emphasis upon recycling aluminium than upon<br>recycling copper or zinc. <b>5</b>  |          |
| <b>17</b> <b>a</b> 1.00 g samples of lithium, calcium and zinc were heated in air to form<br>oxides. Calculate the percentage increase in mass you would expect for<br>each metal (assuming complete reaction). <b>3</b><br><b>b</b> Which two of the following factors determine the percentage increase in<br>mass when a metal is converted to its oxide?<br><b>(i)</b> reactivity of the metal<br><b>(ii)</b> atomic weight of the metal                |          |

- (iii) mass of metal used
  - (iv) the formula of the oxide
  - (v) state of division of the metal
  - (vi) whether pure oxygen or air is used
- 2

**18** Evaluate the contribution of Gay-Lussac and Avogadro to the determination of formulae for gaseous elements and compounds. 3

**19** Select three metals widely used in the everyday world. List two common uses for each of these metals and explain why the properties of the metal make it particularly suitable for those uses. 6

# MODULE 2 AND THE NEW SOUTH WALES HSC SYLLABUS

MODULE  
**2**

This section allows you to check that *Conquering Chemistry* has covered all necessary material for Module 2 of the New South Wales HSC Preliminary Course syllabus.

## Syllabus content

The following table lists (for Module 2) the items from the *students learn to* column of the HSC syllabus and shows where they are treated in *Conquering Chemistry Preliminary Course (CCPC)*.

Location of HSC syllabus Preliminary Course material in *Conquering Chemistry* for Module 2

Syllabus reference	Sections where found in CCPC
Students learn to:	
<b>8.3.1 Metals have been extracted and used for many thousands of years</b>	
■ outline and examine some uses of different metals through history, including contemporary uses, as uncombined metals or as alloys	4.1, 4.2
■ describe the use of common alloys including steel, brass and solder and explain how these relate to their properties	4.1, 4.2, 4.3
■ explain why energy input is necessary to extract a metal from its ore	4.8
■ identify why there are more metals available for people to use now than there were 200 years ago	4.15
■ construct word and balanced formulae equations of chemical reactions as they are encountered	4.7, 4.9, 4.10
<b>8.3.2 Metals differ in their reactivity with other chemicals and this influences their uses</b>	
■ describe observable changes when metals react with dilute acid, water and oxygen	4.9
■ describe and justify the criteria used to place metals into an order of activity based on their ease of reaction with oxygen, water and dilute acids	4.11
■ identify the reaction of metals with acids as requiring the transfer of electrons	4.9, 4.10
■ outline examples of the selection of metals for different purposes based on their reactivity, with a particular emphasis on current developments in the use of metals	4.13
■ outline the relationship between the relative activities of metals and their positions on the Periodic Table	4.12
■ identify the importance of first ionisation energy in determining the relative reactivity of metals	4.12

Syllabus reference Students learn to:	Sections where found in CCPC
<b>8.3.3 As metals and other elements were discovered, scientists recognised that patterns in their physical and chemical properties could be used to organise the elements into a Periodic Table</b>	
<ul style="list-style-type: none"> <li>■ identify an appropriate model that has been developed to describe atomic structure</li> </ul>	2.7 to 2.11
<ul style="list-style-type: none"> <li>■ outline the history of the development of the Periodic Table including its origins, the original data used to construct it and the predictions made after its construction</li> </ul>	6.1
<ul style="list-style-type: none"> <li>■ explain the relationship between the position of elements in the Periodic Table, and             <ul style="list-style-type: none"> <li>– electrical conductivity</li> <li>– ionisation energy</li> <li>– atomic radius</li> <li>– melting point</li> <li>– boiling point</li> <li>– combining power (valency)</li> <li>– electronegativity</li> <li>– reactivity</li> </ul> </li> </ul>	6.2 to 6.5
<b>8.3.4 For efficient resource use, industrial chemical reactions must use measured amounts of each reactant</b>	
<ul style="list-style-type: none"> <li>■ define the mole as the number of atoms in exactly 12 g of carbon-12 (Avogadro's number)</li> </ul> <p>(relative atomic, molecular and formula masses, Sections 5.1, 5.2, 5.3)</p>	5.4, 5.5, 5.6
<ul style="list-style-type: none"> <li>■ compare mass changes in samples of metals when they combine with oxygen</li> </ul>	Chapter 5 Exercises 28, 29; Revision test Question 17
<i>(per cent composition, calculation of formulae, mass-mass calculations, Sections 5.7 to 5.10)</i>	
<ul style="list-style-type: none"> <li>■ describe the contribution of Gay-Lussac to the understanding of gaseous reactions and apply this to an understanding of the mole concept</li> </ul>	5.12
<ul style="list-style-type: none"> <li>■ recount Avogadro's law and describe its importance in developing the mole concept</li> </ul>	5.12
<ul style="list-style-type: none"> <li>■ distinguish between empirical formulae and molecular formulae</li> </ul>	5.8
<b>8.3.5 The relative abundance and ease of extraction of metals influence their value and breadth of use in the community</b>	
<ul style="list-style-type: none"> <li>■ define the terms mineral and ore with reference to economic and non-economic deposits of natural resources</li> </ul>	4.5
<ul style="list-style-type: none"> <li>■ describe the relationship between the commercial prices of common metals, their actual abundances and relative costs of production</li> </ul>	4.16
<ul style="list-style-type: none"> <li>■ explain why ores are non-renewable resources</li> </ul>	4.17

Syllabus reference	Sections where found in CCPC
Students learn to:	
■ describe the separation processes, chemical reactions and energy considerations involved in the extraction of copper from one of its ores	4.5 to 4.8
■ recount the steps taken to recycle aluminium	4.17

## Compulsory experiments

The table below lists the compulsory experiments for Module 2 (from the right-hand column of the syllabus), along with places where you can find procedures for performing them (in *CCPC BLM*) and the location of relevant information in *CCPC*.

*CCPC* is an abbreviation for this book *Conquering Chemistry Preliminary Course*.

*CCPC BLM* is *Conquering Chemistry Preliminary Course Blackline Masters* by Debra Smith, McGraw-Hill Australia, Sydney, 2003

### Information about compulsory experiments for Module 2

Experiment	Location in CCPC BLM	Related material in CCPC
1 Determine a metal activity series by carrying out tests and if necessary incorporating information from other sources	Module 2 Worksheet 4 (page 59), but this uses displacement reactions (HSC syllabus) instead of reactions with oxygen, water and dilute acid. <sup>a</sup>	Section 4.9 Exam-style question 8 Revision Test question 12
2 Extract copper from copper carbonate (or from malachite or azurite)	Module 2 Worksheet 10 (page 86)	Chapter 4 Exercise 11 Exam-style question 16
3 Measure the mass ratio of metal to non-metal in a binary compound and calculate its empirical formula	Module 2 Worksheet 13 (page 96)	Chapter 5 Example 6 and Exercises 25, 27 Revision Test question 15

<sup>a</sup> An experiment that uses reactions with water and dilute acid in addition to displacement reactions is Experiment E2.1 in Deteric, G. and Ware, G., Senior Chemistry Practical Manual, Heinemann, Melbourne, 1995.

## Other items in the student activity (right-hand) column

The table below lists (in abbreviated form) the non-experimental items in the student activity (right-hand) column of the syllabus and indicates where in *Conquering Chemistry* relevant information is located.

### Location of material for other Module 2 activities

Syllabus item	Relevant material in CCPC
<b>8.3.1</b>	■ gather, process analyse ...range of alloys ... ■ analyse information ... Bronze Age ...
	For Investigation in Section 4.1 Section 4.4
<b>8.3.2</b>	■ Experiment 1 (see above)

	■ construct ... equations ... metals ...	Section 4.9 and Exercises 13, 14
	■ construct half-equations ... electron transfer ...	Section 4.10 and Exercise 15
<b>8.3.3</b>	■ process information ... develop a periodic table ...	Chapter 6 Exercise 1 CCPC BLM Module 2 Worksheet 8 (page 75)
	■ use computer-based technologies ...	Websites listed on pages 167–8
<b>8.3.4</b>	■ Experiment 2 (see above)	
	■ process information ... to interpret balanced equations in terms of mole ratios	Section 5.9 and Exercises 30, 31
	■ Experiment 3 (see above)	
	■ solve problems ... involving Avogadro's number and ... number of moles of a substance	Section 5.6 and Exercises 6 to 19
	■ process information ... volumes of gases ...	Section 5.13 and Exercises 42, 43, 44
<b>8.3.5</b>	■ discuss the importance ... yield ... extraction of commercial ore deposits	Section 5.11 and Exercises 38 to 41
	■ justify the increased recycling of metals ...	Section 4.17
	■ analyse information to compare the cost and energy ...extraction ... recycling of aluminium	<i>For Investigation</i> in Section 4.17

## Prescribed focus areas

The five prescribed focus areas of the syllabus are listed on page 94.

The material of Module 2 (in this book and in the syllabus) provides good scope for discussing some aspects of the history of chemistry (Area 1), in particular the development of the extraction and use of metals from about 3000 BCE to the present, the historical development of the Periodic Table, and the contributions of Gay-Lussac and Avogadro to the problem of working out formulae for molecules of gaseous elements and compounds.

There is some scope for discussing Area 5, current issues, research and development, in terms of newer metals and alloys that have been introduced in recent years, particularly titanium alloys.

Uses of metals and recycling provide opportunities for involving Area 4, implications for society and the environment, though the full implications for the environment of mining and extracting metals are not in the syllabus.

With the module's heavy emphasis on properties and uses of metals, there are ample opportunities to discuss the applications and uses of chemistry, Area 3, while the sections on stoichiometry and the Periodic Table provide good insights into the nature and practice of chemistry (Area 2).

Module 2 therefore allows all five prescribed focus areas to receive significant attention.