Chapter 1 Materials in our world

Section 1.1 Materials science

1.1 KEY QUESTIONS

1. A material is a substance used to construct an object.
2. Metals generally have high tensile strength, are ductile, are malleable, have a shiny lustre, have high melting points and are good thermal and electrical conductors.
3. An alloy is a mixture of a metal with other metals or small amounts of non-metals.
4. Polymers generally are less dense than metals, are corrosion resistant, possess excellent electrical resistance and some are compatible with human tissue.
5. Ceramics generally are hard, demonstrate high compressive strength and are able to withstand high temperatures.

Section 1.2 Nanotechnology

1.2 KEY QUESTIONS

1. **a**

$$\begin{aligned} 8.35\text{ cm} &= 0.0835\text{ m} \\ 0.0835\text{ m} &= 8.35 \times 10^{-2}\text{ m} \\ (8.35 \times 10^{-2}) \times 10^{9} &= 8.35 \times 10^{-2+9} \\ &= 8.35 \times 10^{7}\text{ nm} \end{aligned}$$

   b

$$\begin{aligned} 1.35\text{ cm} &= 1.35 \times 10^{-2}\text{ m} \\ 1.35 \times 10^{-2}\text{ m} \times 10^{9} &= 1.35 \times 10^{7}\text{ nm} \end{aligned}$$

   c

$$\begin{aligned} 4.2\text{ mm} &= 4.2 \times 10^{-3}\text{ m} \\ 4.2 \times 10^{-3} \times 10^{9} &= 4.2 \times 10^{6}\text{ nm} \end{aligned}$$

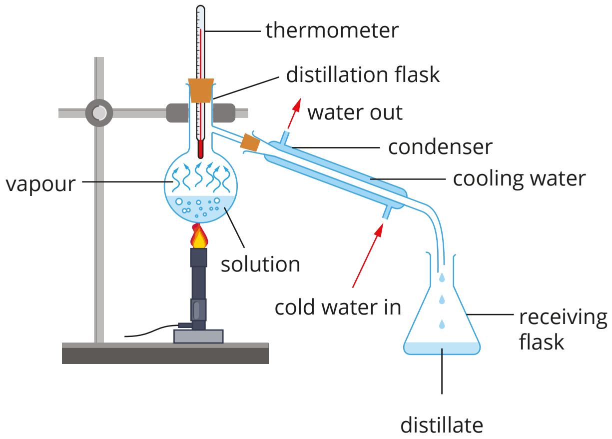
2. Due to the size of nanoparticles, they can travel freely throughout the body, even into cells themselves.
3. At the nanoscale, the ratio of surface area to volume of the particles becomes such that surface area effects, or quantum effects, begin to significantly influence the properties of the material.

Section 1.3 Purifying materials

1.3 KEY QUESTIONS

1. The particle size of salt and pepper is too similar and will likely both be either passed or blocked by a sieve.
2. No, you cannot. The solute in a solution will not settle out of the solution no matter how long you wait.
3. No. Water is not a mixture of hydrogen and oxygen, but a pure compound composed of molecules of hydrogen chemically bonded to oxygen.

4



5 **a** Use a separating funnel or distillation.
  b Use evaporation or distillation.
  c Use sieving.
  d Add water then filter or decant.
  e Add water then decant (the wood chips float).

6 pentane, benzene, heptane, octane

CHAPTER 1 REVIEW

1 A composite material is a combination of two or more distinct materials that creates a substance that has a unique combination of properties of its components.

2 Papier mâché is composed of paper and glue.

3 Metals = aluminium and stainless steel. Ceramics = porcelain. Polymers = cotton and nylon.

4 There are two possible reasons. Pure gold is very soft, so alloys of gold are used in jewellery as they are more durable than pure gold. Alloys of gold are also cheaper, using less gold in total but still maintaining the attractiveness of pure gold.

5 The properties demonstrated by elements and compounds depend on the arrangement of the atoms and molecules in the substance. Diamond, graphite and carbon nanotubes must have different arrangements of particles/atoms.

6 Reinforced concrete gains the benefit of steel reinforcing bars with high tensile strength, while maintaining the high compressive strength of concrete.

7 Iron corrodes quite readily and would rust in a relatively short period of time. Stainless steel is much more corrosion-resistant and durable.

8 **a** copper and tin
  b tin and copper, antimony, bismuth or lead
  c gold and nickel, manganese, platinum or palladium

9 **a** $5\,\text{cm} = 0.05\,\text{m} = 5.0 \times 10^{-2}\,\text{m} = 5.0 \times 10^{-2} \times 10^{9} = 5.0 \times 10^{7}\,\text{nm}$
  b $12\,\text{mm} = 0.012\,\text{m} = 1.2 \times 10^{-2}\,\text{m} = 1.2 \times 10^{-2} \times 10^{9} = 1.2 \times 10^{7}\,\text{nm}$
  c $0.02\,\text{mm} = 0.00002\,\text{m} = 2.0 \times 10^{-5}\,\text{m} = 2.0 \times 10^{-5} \times 10^{9} = 2.0 \times 10^{4}\,\text{nm}$

10 Zinc oxide nanoparticles are colourless.

11 **a** Top-down fabrication involves the removal of material from particles until the desired material is produced, whereas bottom-up fabrication is the building up of the material, one atom or molecule at a time.
  b Advantages of top-down are that large quantities can be produced relatively cheaply, with good uniformity and with existing technology. Disadvantages are that it is limited to relatively simple structures and by the scale of the tools available. Advantages of bottom-up is that far more complicated materials can be manufactured, but the disadvantages are that only small amounts of material can be produced economically.

12 A colloid is a mixture of an insoluble substance dispersed throughout another substance. Unlike suspensions, the two components in a colloid will not separate (based on density) over a period of time as a suspension will.

13 In distillation, all components of the mixture are retained, whereas in evaporation the solvent is lost to the atmosphere as vapour.

ISBN 978 1 4886 1772 0

14 The heat from the Sun causes water to evaporate, forming a vapour. This water vapour then condenses on the plastic and drips into the container located at the centre of the hole. Any salt or contaminants present in the original water are left behind.

15 Evaporation of the solution will only remove the solvent from the solution, leaving the multiple different salts still mixed but as a solid instead of a solution.

16 **a** Add water, decant off sawdust then filter out sand as the residue, then evaporate off water to leave sugar.

b Use a separating funnel to remove oil (less dense and floats) then use fractional distillation to separate ethanol and water.

c Dissolve the instant coffee in water, filter to collect tea leaves as a residue, evaporate off water to obtain instant coffee.

17 Carbon fibres are usually used to reinforce polymer or plastic and produce a material that is very strong but still light.

18 The ease with which nanoparticles enter the body and the lack of evidence for their long-term safety are causes for concern.

19 Grey goo is an apocalyptic scenario where self-assembling/self-replicating nanobots consume all the available resources on Earth by uncontrolled replication of themselves.

ISBN 978 1 4886 1772 0

Chapter 2 Atoms: structure and mass

Section 2.1 Atomic theory

2.1 KEY QUESTIONS

1 proton, neutron and electron

2 protons and neutrons; found in the nucleus

3 The electrostatic attraction of the protons; the negative electrons are attracted to the positive protons and pulled towards them.

Section 2.2 Describing atoms

Worked example: Try yourself 2.2.1

CALCULATING THE NUMBER OF SUBATOMIC PARTICLES

Calculate the number of protons, neutrons and electrons for the atom with this atomic symbol: ${}^{235}_{92}U$	
Thinking	**Working**
The atomic number is equal to the number of protons.	Number of protons = $Z = 92$
Find the number of neutrons. Number of neutrons = mass number – atomic number	The number of neutrons = $A - Z$ $= 235 - 92$ $= 143$
Find the number of electrons. The number of electrons is equal to the atomic number because the total negative charge is equal to the total positive charge.	The number of electrons = $Z = 92$

2.2 KEY QUESTIONS

1 **a** Na
  b Cu
  c Ne
  d K
  e Be
  f Ca

2 atomic number

3 mass number

4 protons and electrons

5 $Z = 15$: protons = 15, electrons = 15, neutrons = 31 – 15 = 16

6 It has seven protons so its atomic number is 7. The element that contains seven protons is nitrogen.

ISBN 978 1 4886 1772 0

Section 2.3 Isotopes

Worked example: Try yourself 2.3.1

CALCULATING THE RELATIVE ATOMIC MASS OF SILVER: METHOD 1

Use the data from Table 2.3.1 to calculate the relative atomic mass of silver.

Thinking	**Working**
Calculate the relative atomic mass contribution by the silver-107 isotope, which is the relative isotopic mass of the Ag-107 isotope, times the percentage abundance, divided by 100.	$106.9 \times 51.8 \div 100 = 55.3742$
Calculate the relative atomic mass contribution by the silver-109 isotope.	$108.9 \times 48.2 \div 100 = 52.4898$
Add the relative atomic mass contributions of all the isotopes.	$55.3742 + 52.4898 = 107.864$ $= 107.9$

Worked example: Try yourself 2.3.1

CALCULATING THE RELATIVE ATOMIC MASS OF SILVER: METHOD 2

Use the data from Table 2.3.1 to calculate the relative atomic mass of silver.

Thinking	**Working**
Identify the relative isotopic masses and the abundance of each isotope.	^{107}Ag relative isotopic mass = 106.9 relative abundance = 51.8% ^{109}Ag relative isotopic mass = 108.9 relative abundance = 48.2%
Use the expression for calculating the weighted mean (A_r) from the percentage abundance and the relative isotopic masses: $A_r = \frac{(\% \times \text{relative isotopic mass}) + (\% \times \text{relative isotopic mass})}{100}$	$A_r = \frac{(51.80 \times 106.9) + (48.2 \times 108.9)}{100}$
Calculate the relative atomic mass (Give answer to 4 significant figures.)	$A_r = 107.9$

ISBN 978 1 4886 1772 0

Worked example: Try yourself 2.3.2

CALCULATING THE ABUNDANCE OF ISOTOPES FROM RELATIVE ATOMIC MASSES AND ISOTOPIC MASSES

The relative atomic mass of copper is 63.54. The relative isotopic masses of its two isotopes, ^{63}Cu and ^{65}Cu, are 62.95 and 64.95. Calculate the relative abundances of the isotopes.

Thinking	**Working**
Use ***a*** to represent the percentage abundance of the lighter isotope. The percentage abundance of the heavier isotope must therefore equal (100 – ***a***).	Abundance of ^{63}Cu = ***a*** Abundance of ^{65}Cu = (100 – ***a***)
Substitute the values into the following expression: $A_r = \frac{(\% \times \text{relative isotopic mass}) + (\% \times \text{relative isotopic mass})}{100}$	$63.54 = \frac{(62.95 \times \mathbf{a}) + (64.95 \times (100 - \mathbf{a}))}{100}$
Expand the top line of the equation.	$63.54 = \frac{62.95\mathbf{a} + 6495 - 64.95\mathbf{a}}{100}$
Solve the equation to find ***a*** (the abundance of the lighter isotope).	$6354 = 62.95\mathbf{a} + 6495 - 64.95\mathbf{a}$ $6354 - 6495 = 62.95\mathbf{a} - 64.95\mathbf{a}$ $-141 = -2.00\mathbf{a}$ $\mathbf{a} = \frac{141}{2.00} = 70.50$ Abundance of ^{63}Cu = 70.50%
Determine the abundance of the heavier isotope using (100 – ***a***).	Abundance of ^{65}Cu = 100 – 70.50 = 29.50%

2.3 KEY QUESTIONS

1 An isotope is an atom of an element that has a different number of neutrons and therefore a different mass number.

2 Isotopes of the same element have the same atomic number and therefore the same number of protons and electrons. The number of neutrons is different between isotopes of the same element, therefore they have different mass numbers.

3 Carbon-14 has a mass number $A = 14$. The number of neutrons is:
mass number – the atomic number = 14 – 6 = 8
Carbon-12 has a mass number of 12 and therefore has six neutrons.
Therefore the difference in the number of neutrons is 8 – 6, which is two neutrons.

4 **a** $A_r(\text{O}) = \frac{(15.995 \times 99.76) + (16.999 \times 0.04) + (17.999 \times 0.20)}{100}$

$= 15.999$

b $A_r(\text{H}) = \frac{(1.008 \times 99.986) + (2.014 \times 0.014) + (3.016 \times 0.0001)}{100}$

$= 1.008$

5 Let the percentage abundance of the lighter isotope be x%.
∴ percentage abundance of the heavier isotope will be $(100 - x)$%

$$6.94 = \frac{(6.02 \times x) + (7.02 \times (100 - x))}{100}$$

$$694 = 6.02x + 702 - 7.02x$$

$$x = 8\%$$

Percentage abundance of the lighter isotope is 8%.

ISBN 978 1 4886 1772 0

Section 2.4 Mass spectrometry

Worked example: Try yourself 2.4.1

CALCULATING PERCENTAGE ABUNDANCE OF EACH ISOTOPE FROM THE MASS SPECTRUM

Figure 2.4.5 shows a simplified mass spectrum of lead. From this mass spectrum, calculate the percentage abundances of each of its three isotopes.

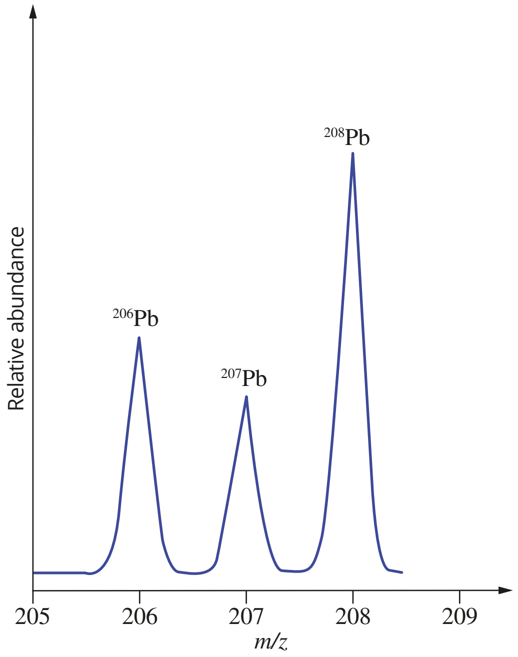


FIGURE 2.4.5 Mass spectrum of lead

Thinking	**Working**
Measure the peak height for each isotope using a ruler.	From the spectrum, the height of each peak is: ^{206}Pb = 3.1 cm ^{207}Pb = 2.4 cm ^{208}Pb = 5.6 cm
Calculate the total peak height for the three isotopes by adding the individual peak heights.	Total peak height = 3.1 + 2.4 + 5.6 = 11.1 cm
Substitute the peak height for each isotope into the formula: $\% \text{ abundance} = \frac{\text{peak height}}{\text{total peak height}} \times 100$	$\% \text{ abundance } ^{206}\text{Pb} = \frac{3.1}{11.1} \times 100 = 28\%$ $\% \text{ abundance } ^{207}\text{Pb} = \frac{2.4}{11.1} \times 100 = 22\%$ $\% \text{ abundance } ^{208}\text{Pb} = \frac{5.6}{11.1} \times 100 = 50\%$

2.4 KEY QUESTIONS

1 1 The sample is vaporised and then ionised using high-energy electrons.
  2 The ions are separated and accelerated according to their mass-to-charge $\left(\frac{m}{z}\right)$ ratios in an electric/magnetic field.
  3 The ions that have a particular mass-to-charge ratio are detected by a device that counts the number of ions that strike it.

2 the number of different isotopes in the sample

3 the mass-to-charge ratio, which equates to the relative isotopic mass

4 the relative abundance of the isotopes

5 **a** % abundance ^{90}Zr = $\frac{\text{peak height}}{\text{total peak height}} \times 100 = 51\%$

% abundance ^{91}Zr = $\frac{\text{peak height}}{\text{total peak height}} \times 100 = 11\%$

% abundance ^{92}Zr = $\frac{\text{peak height}}{\text{total peak height}} \times 100 = 17\%$

% abundance ^{94}Zr = $\frac{\text{peak height}}{\text{total peak height}} \times 100 = 17\%$

% abundance ^{96}Zr = $\frac{\text{peak height}}{\text{total peak height}} \times 100 = 4\%$

b $A_r(\text{Zr}) = \frac{(90 \times 51) + (91 \times 11) + (92 \times 17) + (94 \times 17) + (96 \times 4)}{100} = 91$

CHAPTER 2 REVIEW

1 Atoms are hard, indivisible structures.

2 the nucleus

3 The protons and neutrons form the nucleus. The electrons are grouped in shells and occupy the space around the nucleus.

4 The mass of a proton is approximately equal to the mass of a neutron and is about 1840 times the mass of an electron. The proton and electron have equal but opposite charges and the neutron has no charge.

5 Thomson discovered the electron, Rutherford discovered the proton and Chadwick discovered the neutron.

6 The periodic table is organised by atomic number. Although it appears to also be organised by mass number, it is not consistent throughout the whole table.

7 **a** atomic number is 24; mass number is 52

b 24 electrons, 24 protons, 52 – 24 = 28 neutrons

8 **a** atomic number is 25; mass number is 55

b 25 electrons, 25 protons, 55 – 25 = 30 neutrons

9 Atoms are electrically neutral. The positive charge on one proton balances the negative charge on one electron. Therefore, for electrical neutrality, there must be an equal number of protons and electrons.

10 Most elements have more than one isotope, so they will have more than one mass number. All bromine atoms have 35 protons in their nuclei. No other type of atom has 35 protons in its nucleus (i.e. no other atom has an atomic number of 35). Isotopes of bromine, however, differ in their mass numbers, so mass number is not fixed for an element (except for those elements such as sodium, which have only one naturally occurring isotope). In addition, an isotope of one element may have the same mass number as an isotope of another element.

11 No. Isotopes have the same number of protons (atomic number) but different numbers of neutrons (and therefore different mass numbers). These atoms have different atomic numbers and different mass numbers.

12 The relative atomic mass of carbon is the weighted average of the isotopic masses of all carbon isotopes (i.e. ^{12}C, ^{13}C and ^{14}C). Small amounts of ^{13}C and ^{14}C make this average slightly greater than 12, the relative isotopic mass of the ^{12}C isotope.

13 **a** Let the percentage abundance of the lighter isotope be x%.

$\therefore$ percentage abundance of the heavier isotope will be $(100 - x)\%$

$\therefore 69.72 = \frac{(x \times 68.95) + ((100 - x) \times 70.95)}{100}$

$\therefore x = 61.5$

$\therefore$ proportion of the lighter isotope = 61.5%

b 20.2% (calculated by using the same process as for part **a**.)

14 The isotopes of an element all have the same chemical properties, they only differ by the physical properties.

15 Nothing. Argon has 18 protons, 18 electrons and 22 neutrons whereas calcium has 20 protons, 20 electrons and 20 neutrons.

16 $A_r(\text{Pd}) = \frac{(101.9049 \times 0.96) + (103.9036 \times 10.97) + (104.9046 \times 22.23) + (105.9032 \times 27.33) + (107.9039 \times 26.71) + (109.9044 \times 11.800)}{100}$

$= 106.4$

ISBN 978 1 4886 1772 0

17 **a** $A_r(\text{Ar}) = \dfrac{(35.978 \times 0.307) + (37.974 \times 0.060) + (39.974 \times 99.633)}{100}$

$= 39.96$

$A_r(\text{K}) = \dfrac{(38.975 \times 93.3) + (39.976 \times 0.011) + (40.974 \times 6.69)}{100}$

$= 39.11$

b Although potassium atoms have one more proton than argon atoms, the most abundant isotope of argon has 22 neutrons, giving it a relative atomic mass close to 40. The most abundant isotope of potassium has only 20 neutrons, giving it a relative atomic mass close to 39.

18 **a** peak heights $^{50}\text{Cr} = 0.3$ units, $^{52}\text{Cr} = 12$ units, $^{53}\text{Cr} = 1$ unit, $^{54}\text{Cr} = 0.2$ units, Total height = 13.5 units

percentages $^{50}\text{Cr} = 2.2\%$, $^{52}\text{Cr} = 88.9\%$, $^{53}\text{Cr} = 7.4\%$, $^{54}\text{Cr} = 1.5\%$

b $A_r(\text{Cr}) = \dfrac{(50 \times 2.2) + (52 \times 88.9) + (53 \times 7.4) + (54 \times 1.5)}{100}$

$= 52$

19 Let the percentage abundance of the lighter isotope be $x\%$.

$\therefore$ percentage abundance of the heavier isotope will be $(100 - x)\%$

$\therefore\ 151.96 = \dfrac{(x \times 150.92) + ((100 - x) \times 152.92)}{100}$

$\therefore\ x = 48.0$

$\therefore$ proportions of the isotopes are 48.0% and 52.0%.

20 Since the relative atomic mass of an element is a weighted average of the isotopic masses of the ions of that element and since the most abundant isotope occurs in far greater quantities than the other two, the relative atomic mass will be close to 20.

21 **a** The mass number is the sum of the protons and neutrons present in the nucleus of an atom. For example, the mass number of the lighter of the two chlorine isotopes (see Worked example 2.3.1, page 33) is 35.

b The relative isotopic mass of an isotope is the mass of an atom of that isotope relative to the mass of an atom of ^{12}C, taken as 12 units exactly. For example, the relative isotopic mass of the lighter of the two chlorine isotopes (see Worked example 2.3.1) is 34.969.

c The relative atomic mass of an element is the weighted average of the relative masses of the isotopes of the element, based on a natural sample, on the ^{12}C scale. For example, the relative atomic mass of chlorine (see Worked example 2.3.1) is 35.45.

22 **a** $(454\,212.6 + 56.984\,73 + 32\,779.37) = 487\,049$

$\text{K-40} = \dfrac{56.98473}{487049} \times 100 = 0.011\,70\%$

b $38.96 \times \dfrac{454212.6}{(454212.6 + 56.98473 + 32779.37)} = 36.333\,355\,66$

$39.96 \times \dfrac{56.98473}{(454212.6 + 56.98473 + 32779.37)} = 0.004\,675\,32$

$40.96 \times \dfrac{32799.37}{(454212.6 + 56.98473 + 32779.37)} = 2.756\,690\,025$

$36.333\,355\,66 + 0.004\,675\,32 + 2.756\,690\,025 = 39.09$

ISBN 978 1 4886 1772 0

Chapter 3 Electrons and the periodic table

Section 3.1 Electronic structure of atoms

3.1 KEY QUESTIONS

1 Each line in an emission spectrum corresponds to a specific amount of energy. This energy is emitted when electrons from higher energy electron shells transition to a lower energy shell. Different lines indicate that there are differences in energy between shells. This is evidence that electrons are found in shells with discrete energy levels.

2 Electrons revolve around the nucleus in fixed, circular orbits.

Electron orbits correspond to specific energy levels in the atom.

Electrons can only occupy fixed energy levels and cannot exist between two energy levels.

Orbits of larger radii correspond to higher energy levels.

3 Energy is emitted as light or electromagnetic radiation.

Section 3.2 Electron arrangement in the periodic table

Worked example: Try yourself 3.2.1

DETERMINING THE ELECTRON CONFIGURATION OF AN ATOM

Determine the electron configuration of argon.

Thinking	Working
Locate argon (Ar) on the periodic table and find its atomic number. This is the number of protons in the nucleus.	18 protons
As atoms are uncharged, they have an equal number of positive charges (protons) as negative charges (electrons).	18 electrons
Start allocating electrons from lowest energy to highest; ensure each shell is complete before progressing to the next shell. The first shell can accommodate a maximum of two electrons. There are 16 electrons remaining to be allocated.	2
The second shell can accommodate a maximum of eight electrons. There are eight electrons remaining to be allocated.	2,8
The remaining eight electrons are allocated into the third (valence) shell.	2,8,8

3.2 KEY QUESTIONS

1 **a** 2,3
  b 2,8,2
  c 2,8,8,2

2 **a** 2,2
  b 2,8,6
  c 2,8,8
  d 2,8,2
  e 2,8

ISBN 978 1 4886 1772 0

3 **a** helium, He
  b fluorine, F
  c aluminium, Al
  d nitrogen, N
  e chlorine, Cl

Section 3.3 Trends in the periodic table

3.3 KEY QUESTIONS

1 Carbon has 6 protons and 6 electrons. Its electron configuration is 2,4. Carbon has two inner-shell electrons.
Core charge = 6 – 2 = +4.

2 The atomic radii can be considered the distance from the centre of the nucleus to the valence shell. As the number of protons increases, the attraction on the valence electrons increases, thus pulling the electrons closer to the nucleus.

3 **a** Ionisation energy is the least amount of energy needed to remove an electron from an atom or ion in the gas phase.

b The factors that affect ionisation energy across a period are the size of the atom (i.e. the distance of the outermost (highest energy) electron from the nucleus), and the charge on the nucleus.

4 Across a period, the number of occupied shells in the atoms remains constant but core charge increases. The valence electrons become more strongly attracted to the nucleus, so more energy is required to remove an electron from an atom. Therefore, the first ionisation energy increases across a period.

5 **a** **i** fluorine, F
  ii francium, Fr
b **i** group 17
  ii group 1
c Elements in group 18, the noble gases, have a very stable electron configuration and so are unreactive.

6 As core charge increases, electronegativity increases.

Section 3.4 Quantisation of energy

3.4 KEY QUESTIONS

1 Each line in an emission spectrum corresponds to a specific amount of energy. This energy is emitted when electrons from higher-energy electron shells transition to a lower-energy shell. Different lines indicate that there are differences in energy between shells. This is evidence that electrons are found in shells with discrete energy levels.

2 Different wavelengths and frequencies correspond to different colours of visible light.

3 A. There is a total of six electrons, equal to the atomic number of carbon.

4 C. Carbon still only has six electrons, equal to its atomic number, but here we have an electron in the higher third energy level rather than the lower second energy level, which is still unfilled.

5 No. Although they share the same noble gas electron configuration, they are different in terms of the arrangement of electrical charges both in the electron shells and the nucleus, therefore they will demonstrate different emission spectra.

6 No. Although they share the same electron configuration, differences in the mass of the nucleus (due to different number of neutrons) affect the system of electrical charges in the atom and therefore will alter their emission spectra slightly.

ISBN 978 1 4886 1772 0

Section 3.5 **Spectroscopy**

3.5 KEY QUESTIONS

1. Spectroscopy is the study of the interaction between matter and electromagnetic radiation.
2. The advantage of a flame test is it is a quick and easy test for metal atoms. Its weaknesses are it only provides qualitative data, only a small range of metals are detectable using the flame test, metals in low concentrations may be difficult to observe and mixtures of metals will produce confusing results.
3. The light source that produces the light characteristic of the metal being analysed is the source of energy for the excitation of the atoms. The flame merely vaporises the sample to produce free gaseous atoms in their ground state.
4. Atomic absorption spectroscopy (AAS) provides qualitative data whereas a flame test does not. A larger range of elements can be analysed by AAS whereas those suitable to be analysed by the flame test are fewer. AAS is very sensitive and can detect low concentrations whereas a flame test is not as sensitive. AAS is highly selective and can analyse mixtures well whereas the flame test is only applicable to pure samples.
5. Maritime flares traditionally use strontium or calcium compounds to produce their characteristic red colour.

CHAPTER 3 **REVIEW**

1. Bohr's idea was that the energy possessed by electrons is quantised, that is, can only exist as certain discrete values.
2. Nothing is between those electrons and the nucleus.
3. **a** **i** group 13
      ii group 17
      iii group 1
      iv group 18
      v group 14
      vi group 14
   b **i** 4
      ii 2
      iii 1
      iv 1
      v 7
      vi 3
   c **i** silicon, Si, 2,8,4
      ii beryllium, Be, 2,2
      iii argon, Ar, 2,8,8
4. Add the total number of electrons in the electron configuration. All atoms have an equal number of protons and electrons. The number of protons is equal to the atomic number. Find the relevant atomic number on the periodic table.
   a beryllium
   b fluorine
   c silicon
   d calcium
5. **a** silicon (or any other group 14 element)
   b calcium (or any other group 2 element)
   c bromine (or any other group 17 element)
   d nitrogen (or any other group 15 element)
6. groups 1, 2 and 13–18
7. **a** 1
   b 15 – 10 = 5
   c 17 – 10 = 7
   d 2
8. Moving from left to right across groups 1, 2 and 13–17, the charge on the nucleus increases. Each time the atomic number increases by one, the electrons are attracted to an increasingly more positive nucleus. Within a period, the outer electrons are in the same shell—that is, they have the same number of inner-shell electrons shielding them from the nucleus. Therefore, the additional nuclear charge attracts the electrons more strongly, drawing them closer to the nucleus and so decreasing the size of the atom.

ISBN 978 1 4886 1772 0

9 **a** Magnesium and phosphorus, with outer electrons in the third shell, are in the same period. Magnesium has a nuclear charge of +12 but, with two completed inner shells, the outer electrons experience the attraction of a core charge of +2. The outer-shell electrons of phosphorus, which has a nuclear charge of +15 and the same number of inner shells as magnesium, are attracted by a core charge of +5. The stronger attraction of the phosphorus electrons to the core means that more energy is required to remove an electron from a phosphorus atom than from a magnesium atom.

b Both fluorine and iodine are in group 17, so the outer electrons of each atom experience the attraction of the same core charge. Because the outer-shell electrons of a fluorine atom are closer to the nucleus than those of an iodine atom, they are attracted more strongly and so more energy is needed to remove one.

10 Moving across period 2 from lithium to fluorine:

a The radius of the atoms decreases as the core charge increases.

b There is a trend from metals (lithium, beryllium) to non-metals (boron, carbon, nitrogen, oxygen and fluorine).

c Electronegativity increases as the core charge increases and size of the atoms decreases.

11 **a** nitrogen

b chlorine

c chlorine

12 Fluorine. Both oxygen and fluorine will have higher first ionisation energy than chlorine due the reduced distance between the valence shell and the nucleus, and the additional distance between the nucleus and chlorine's valence (third) shell. Out of oxygen and fluorine, fluorine will be highest due to the increased core charge.

13 Ionisation energy of an atom is measured from gaseous atoms so that the energy value is a true indication of only the energy required to ionise the atoms (remove electrons) rather than the energy required to separate atoms during the melting and/or vaporisation of the element.

14 Smaller. Sodium, with an electron configuration of 2,8,1 has a single valence electron in the third shell that is lost when it becomes a cation. The resultant sodium ion, Na^+, has an electron configuration of 2,8. With only two shells now, the sodium ion is significantly smaller.

15 Although the electron that fluorine gains to become an ion enters the same valence shell as fluorine's existing seven valence electrons, the increased repulsion within the (now crowded) valence shell causes the shell to enlarge.

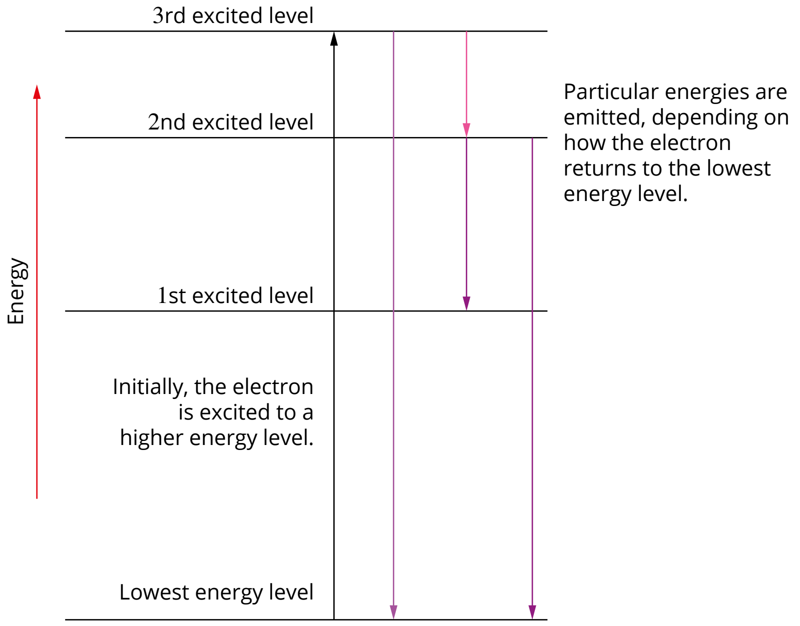
16 $n = 1$

17 The ionisation of an atom requires more energy as this involves the removal of an electron from the atom, completely overcoming the electrostatic attraction to the nucleus, whereas the excitation of an atom does not completely remove an electron, but only raises it temporarily to a higher energy level.

18 The blue colour is a result of the emission spectrum of various chemicals generated during the combustion of the gas that have been excited due to the temperature of the flame.

19 The absorption lines characteristic of oxygen that Fraunhofer measured were the result of absorption taking place within Earth's atmosphere.

20 When atoms absorb energy, it is often possible for electrons to be promoted to various higher energy levels. Electrons can return to the ground state from these excited states by undergoing a number of transitions of different energy. (This is shown in the diagram.) Each transition results in a line of specific energy in the emission spectrum.



Particular energies are emitted, depending on how the electron returns to the lowest energy level.

Initially, the electron is excited to a higher energy level.

21 Using a luminous flame will obscure any colour produced by the metal atoms with the orange colour of the flame itself. Also, an orange flame may not be hot enough to achieve excitation of the metal atoms.

22 Excite a pure sample of the element itself.

23 The excitation and emission of light via the flame test will result in a mixture of different colours that will be hard to identify individually. In atomic absorption spectroscopy, the process relies on using the specific characteristic radiation of the element being tested for, and as no other elements will absorb the same frequencies of light, they should not interfere with the determination.

24 **a** Until Rutherford's work, the plum pudding model of the atom was widely accepted. However, his discovery that a beam of alpha particles directed at thin gold foil caused a few particles to be deflected through high angles led to the development of a new atomic model.

b Although Rutherford's atomic model accounted for a number of atomic properties, it was not able to account for the characteristic emission spectrum of each element.

25 Possible answer:

Periodic table **a** shows the atoms arranged in increasing atomic number (starting from the centre of the 'spiral') but also aligned with other atoms with the same number of valence electrons moving out from this centre. The transition metals and the lanthanide and actinide series are shown as an expansion of the periodic table, corresponding to an expansion of the electron shells within the atoms themselves.

Periodic table **b** shows the expanding shells of the atoms as more electrons are added to atoms of increasing atomic number, but keeping the 'main group' elements in the eight columns to the right of the table.

ISBN 978 1 4886 1772 0

Chapter 4 Metals

Section 4.1 Properties of metals

Worked example: Try yourself 4.1.1

DETERMINING CHARGES

Determine the charge of an aluminium cation.	
Thinking	**Working**
Unreacted aluminium atoms have the same number of protons and electrons.	Atomic number (*Z*) of aluminium is 13: Number of protons is 13, number of electrons is 13.
The electrons in an atom are in shells.	Shell configuration of calcium: 2,8,3
Only the outer-shell electrons will be lost.	Outer shell contains three electrons, 13 – 3 = 10 electrons remaining.
Cation charge = number of protons – number of electrons	Cation charge = 13 – 10 = +3

4.1 KEY QUESTIONS

1 **a** Li atoms have three electrons. The electronic structure is 2,1. There is one electron in the outer shell. The charge of the cation will therefore be +1.
  b Mg atoms have 12 electrons. The electronic structure is 2,8,2. There are two electrons in the outer shell. The charge of the cation will therefore be +2.
  c Al atoms have 13 electrons. The electronic structure is 2,8,3. There are three electrons in the outer shell. The charge of the cation will therefore be +3.

2 **a** Both potassium and gold have good thermal and electrical conductivity. However, gold has a higher density, and higher melting and boiling temperatures than potassium.
  b sodium
  c silver
  d Sodium and potassium are in group 1. Gold and silver are transition metals.

3 **a** silver, copper, gold, aluminium
  b Availability and cost need to be considered; also properties such as malleability and ductility.

4 Sodium belongs to the alkali metals, which have relatively low melting and boiling points, relatively low density and are relatively soft. Iron is a transition metal, which have relatively high melting and boiling points, relatively high density and are relatively hard.

5 tensile strength, cost, availability

ISBN 978 1 4886 1772 0

Section 4.2 Metallic bonding

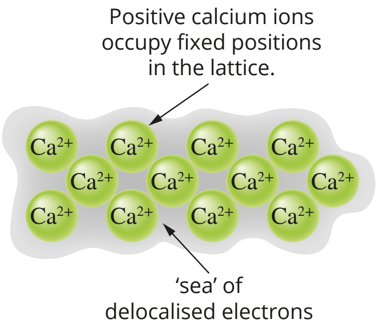
Worked example: Try yourself 4.2.1

EXPLAINING THE CONDUCTIVITY OF MAGNESIUM

With reference to the electron configuration of magnesium, explain why solid magnesium can conduct electricity.	
Thinking	**Working**
Using the atomic number of the element, determine the electron configuration of its atoms. (You may need to refer to a periodic table.)	Mg has an atomic number of 12. This means a neutral atom of magnesium has 12 electrons. The electron configuration is 2,8,2.
From the electron configuration, find how many outer-shell electrons are lost to form cations that have a stable, noble gas electron configuration. These electrons become delocalised.	Mg has two electrons in its outer shell. Mg atoms will tend to lose these two valence electrons to form a cation with a charge of +2. The outer-shell electrons become delocalised and form the sea of delocalised electrons within the metal lattice.
An electric current occurs when there are free-moving charged particles.	If the Mg is part of an electric circuit, the delocalised electrons are able to move through the lattice towards a positively charged electrode.

4.2 KEY QUESTIONS

1 **a**



b strong electrostatic forces of attraction between Ca^{2+} ions and the delocalised valence electrons

2 Barium has a high melting temperature because there are strong attractive forces between the positive ions and the delocalised electrons. Barium conducts electricity because the delocalised electrons from the outer shell are free to move.

3 **a** Graphite and other metals are lustrous and conduct heat and electricity.

b These properties are explained by the presence of free-moving electrons. Graphite and other metals must contain delocalised electrons.

Section 4.3 Reactivity of metals

4.3 KEY QUESTIONS

1 $2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$ potassium + water → potassium hydroxide + hydrogen

2 **a** Metals react more energetically further down the group.

b Metal atoms further down the group have more electron shells, which provide more shielding of the valence electrons from the core charge, reducing the electronegativity and allowing more spontaneous reaction.

3 zinc > iron > gold

4 Calcium. Of the three metals, calcium is the highest in the series so is the most reactive.

ISBN 978 1 4886 1772 0

Section 4.4 **Modifying metals**

4.4 KEY QUESTIONS

1 **a** A 20-cent coin contains copper and nickel. High-carbon steel contains iron and carbon.
  b The 20-cent coin is a substitutional alloy similar to Figure 4.4.3. High-carbon steel is an interstitial alloy, similar to Figure 4.4.2.

2 The metal in the hooks becomes work hardened and brittle.

3 **a** Aluminium contains small areas of regular metallic lattice called crystals. When the aluminium is annealed, the crystal structure is changed to contain more large crystals. Larger crystals are more flexible and easier to shape than smaller crystals and the metal is less likely to break along crystal boundaries during shaping.
  b Quenching the aluminium by heating to the critical temperature and then rapidly cooling causes the growth of small crystals. These crystals make the metal stiffer and do not allow the metal to deform as easily.

4 An alloy with properties appropriate to the task is selected. The metal is annealed to allow shaping of the chisel. The chisel is tempered to make the shaft flexible and strong. The tip of the chisel could then be hardened either by work hardening or local quenching.

5 A is a nanoparticle. B is a nanowire, as the length is more than five times the diameter. C is a nanorod, as the length is between three and five times the diameter. D is not a nanoparticle as all of its dimensions are above 100 nm.

6 iron + oxygen → iron oxide

CHAPTER 4 **REVIEW**

1 Mg, Ca, Sr

2 **a** silver
  b It is too expensive and tarnishes readily.
  c aluminium, copper (combined with stainless steel)

3 electrical conductivity

4 **a** low density
  b high electrical conductivity
  c high tensile strength

5 20 in Ca, 18 in Ca^{2+}

6 Al: 2,8,3; Al^{3+}: 2,8

7 It is able to be drawn into a wire.

8 **a** When a current is applied to the copper wire, the free-moving, delocalised electrons move from one end to the other and so the copper wire conducts electricity.
  b The delocalised electrons in the metal spoon obtain energy from the boiling mixture and move more quickly. These electrons move freely throughout the spoon, colliding with other electrons and metal ions, transferring energy so that the spoon becomes warmer and, eventually, too hot to hold.
  c A lot of energy is required to overcome the strong forces of attraction between the iron ions and the delocalised electrons in the metal lattice, so that the iron changes from a solid to a liquid.
  d Because of the strong forces of attraction between them, the lead ions and the delocalised electrons form a closely packed three-dimensional structure. Also, the lead atom itself has a higher mass to volume ratio than the sulfur atom. This means that the density—the mass per volume—is high.
  e As the copper is drawn out, the copper ions are forced apart and the delocalised electrons rearrange themselves around these ions and re-establish strong forces of attraction.

9 **a** **i** valence electrons not restricted to a region between two atoms
    ii a regular three-dimensional arrangement of a very large number of positive ions or cations
    iii the electrostatic attraction between a lattice of cations and delocalised electrons
  b valence electrons

10 A metal wire contains an extended lattice of metal cations surrounded by a ‘sea’ of delocalised electrons. The electrons are charged and free to move and so can conduct electricity.

11 In a metal lattice, metal cations are in a regular three-dimensional arrangement and have a positive charge. The positive cations are surrounded by a mobile sea of delocalised electrons. This is shown in Figure 4.2.1.

ISBN 978 1 4886 1772 0

12 **a** any of the group 2 metals, e.g. magnesium
  b Magnetic metals are found in the transition metals.

13 aluminium

14 The bubbles are hydrogen gas, which is produced when a reactive metal reacts with water.

15 The reaction on the left is more vigorous and the metal must be more reactive. Iron is a more reactive metal than silver and so iron must be on the left. Silver is less reactive than iron and so silver must be on the right.

16 Metal A is copper. Metal B is sodium. Metal C is aluminium.

17 **a** false
  b true
  c false
  d false
  e true

18 magnesium + oxygen → magnesium oxide

19 **a** copper and nickel; harder, more corrosion-resistant and a silver colour
  b tin and lead; lower melting temperature
  c gold, silver and copper; harder
  d iron, nickel and chromium; resists rusting, stronger
  e mercury and zinc (sometimes a little silver is added); harder, non-toxic

20 needle 2 < needle 3 < needle 1

Needle 2 has been quenched, producing a hard but brittle metal, so it is the least malleable.

Needle 3 has been tempered, producing a hard but more malleable metal.

Needle 1 has been annealed, producing a soft, malleable metal.

21 Steel is used instead of iron as this alloy is stronger, more flexible and resistant to corrosion. Heating the horseshoe during the process allows the worker to change the crystal structure through heat treatment. The final shaping and hammering is an example of work hardening, which aligns the crystals and increases strength.

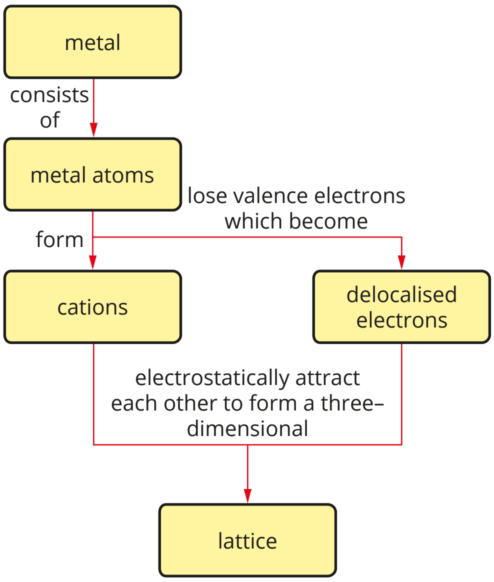
22 nanowire

23 8.34×10^{-7} m is 834 nm. This particle would be too big to be classified as a nanoparticle.

24 The width of a nanowire is in the range of standard nanomaterials. This changes the properties of the metal atoms as they are not exposed to delocalised electrons in the same way as in bulk metals. The long length of a nanowire is not enough to give it electrons that behave in the same way as electrons in bulk metals.

25 **a** **i** iron (steel) or aluminium
    ii Iron and steel are strong. Aluminium has a low density (light) and can be easily coloured.
    iii Iron rusts easily. Aluminium is soft and lacks strength.
  b **i** Copper
    ii It is a good conductor of electricity and is ductile.
  c **i** Gold, silver and platinum
    ii They are non-reactive, malleable, ductile and lustrous.

26 A variety of answers is possible. An example of a possible answer is shown.



27 **a** aluminium Al, copper Cu, gold Au, iron Fe, silver Ag

b aluminium: group 13, period 3
copper: 1st transition series, period 4
gold: 3rd transition series, period 6
iron: 1st transition series, period 4
silver: 2nd transition series, period 5

c gold and silver

d copper, gold, iron and silver

e gold

28 **a** The positive ions are arranged in a regular, three-dimensional lattice.

b Stress corrosion cracking can occur between the crystals.

29 **a** Na: group 1, period 3
K: group 1, period 4
Ca: group 2, period 4

b Na: 2,8,1
K: 2,8,8,1
Ca: 2,8,8,2

c **i** The atoms of Na are smaller than those of K, so the delocalised valence electrons of Na are closer to the positive nuclear charge than those of K. The electrostatic forces of attraction between delocalised electrons and cations are stronger in Na, so Na requires more energy to overcome the metallic bonding to boil the metal.

ii Valence electrons are in the fourth shell in the atoms of both Ca and K. However, there are twice as many valence electrons in the atoms of Ca. Also, the charge on a calcium cation is +2 as opposed to +1 on the potassium cation. So the electrostatic forces of attraction between delocalised electrons and cations are stronger in Ca and so it requires more energy to overcome the metallic bonding to boil the metal.

30 Aluminium is extracted from its ore by electrolysis. There was no process available until 1886.

ISBN 978 1 4886 1772 0

Chapter 5 Ionic bonding

Section 5.1 Properties and structures of ionic compounds

5.1 KEY QUESTIONS

1 B. Substance B has a melting point of 308°C. B will not conduct electricity at 250°C when it is solid but will when it is at 350°C when it is molten, which is characteristic of ionic compounds.

2 **a** The diagram shows that in solid sodium chloride, the sodium and chloride ions are held in fixed positions in the crystal lattice and are not free to move and conduct electricity.

b Molten sodium chloride contains sodium and chloride ions that are free to move and, therefore, it can conduct electricity.

3 Aluminium is a metal. It has an electron configuration of 2,8,3. It would lose three electrons to have a valence shell with eight electrons and therefore become a cation (positive ion).

4 The negative ions are slightly further away from each other than they are from the positive ions in the lattice and the attractive force of the oppositely charged ions outweighs the repulsive force of two positively charged or two negatively charged ions near each other.

Section 5.2 Using the ionic bonding model to explain properties

5.2 KEY QUESTIONS

1 The electrostatic forces of attraction between the positive and negative ions holding the lattice together are very strong and a lot of energy is required to break them apart.

2 **a** When hit with a hammer or hard object, the ions move within the lattice so that like-charged ions line up adjacent to each other.

b When like-charged particles are near each other, they repel due to electrostatic repulsion and this causes the ionic compound to shatter.

3 In solid form, the ionic compound forms a crystal lattice. This is a very strong structure as the strong electrostatic forces of attraction between the positively charged cations and negatively charged anions means that the ions are not free to move. For a substance to be able to conduct electricity, the particles not only need to be charged but also free to move. In solid form, these particles cannot move, but when heated so the ionic compound is now molten, they are able to conduct electricity.

4 salad; salami; saline; expressions such as ‘salt of the Earth’, ‘take with a pinch of salt’, ‘worth one’s salt’; and superstitions, such as throwing salt over one’s left shoulder to keep away evil spirits

ISBN 978 1 4886 1772 0

Section 5.3 Formation of ionic compounds

Worked example: Try yourself 5.3.1

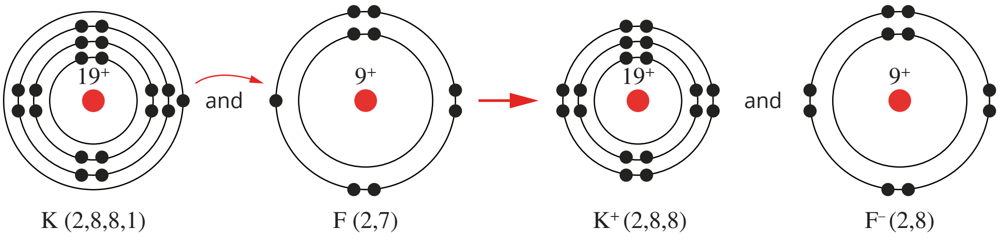
WRITING EQUATIONS FOR REACTIONS BETWEEN METALS AND NON-METALS ATOMS

Write an equation for the reaction between calcium and phosphorus atoms. Show the electron configurations for each element before and after the reaction.	
Thinking	**Working**
Write the symbol and the electron configuration for the metal atom.	Ca (2,8,8,2)
How many electrons will the metal atom lose from its outer shell when it reacts?	2
Write the symbol and the electron configuration of the metal ion that will be formed.	Ca^{2+} (2,8,8)
Write the symbol and the electron configuration for the non-metal atom.	P (2,8,5)
How many electrons will the non-metal atom gain in its outer shell when it reacts?	3
Write the symbol and the electron configuration of the non-metal ion that will be formed.	P^{3-} (2,8)
The total number of electrons lost by metal atoms must equal the total number of electrons gained by non-metal atoms. What is the lowest number ratio of metal atoms to non-metal atoms that will allow this to happen?	metal atom : non-metal atom = 3:2
Using the ratio of metal ion : non-metal ion calculated above, write a balanced equation for the reaction. Show the electron configurations for both the reactant atoms and the product ions.	3Ca (2,8,8,2) + 2P (2,8,5) $\rightarrow$3Ca^{2+} (2,8,8) + 2P^{3-} (2,8,8)

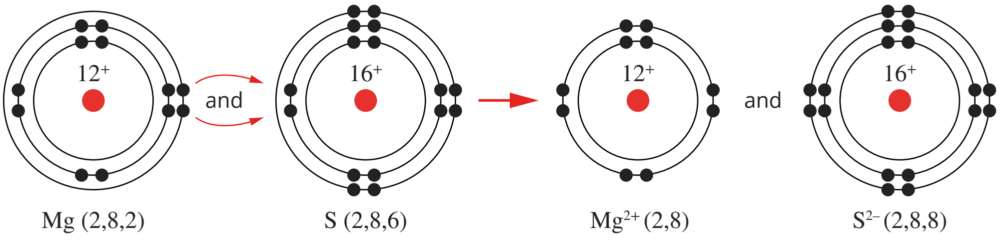
5.3 KEY QUESTIONS

1 Cations: calcium, aluminium. Anions: nitrogen, fluorine and phosphorus. Metals form cations and non-metals form anions. Metals have lower ionisation energies than non-metals so they tend to lose valence electrons more readily. Non-metals have higher electronegativity than metals, so they tend to gain electrons more readily.

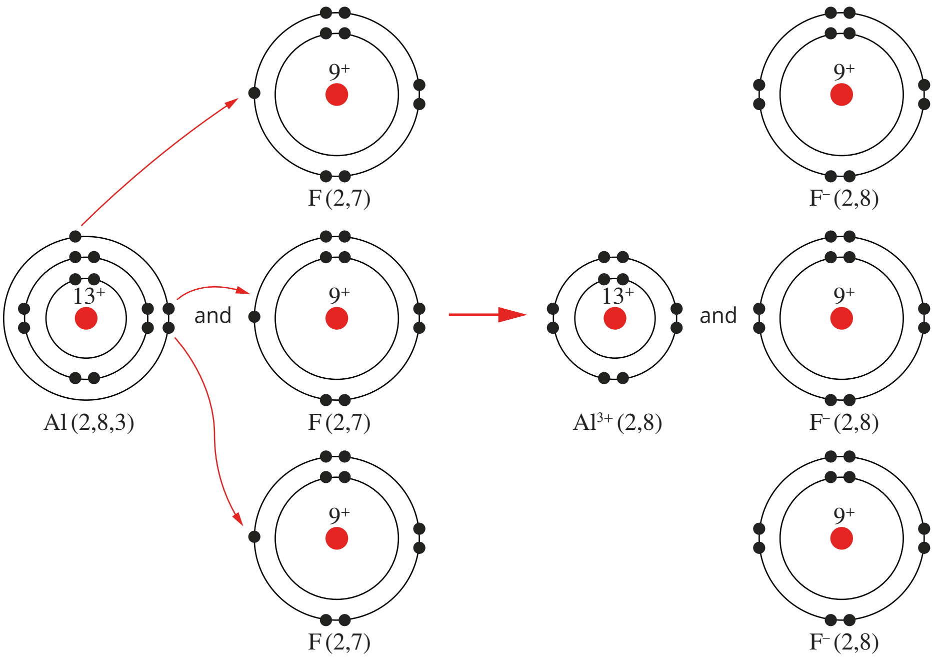
2 **a**



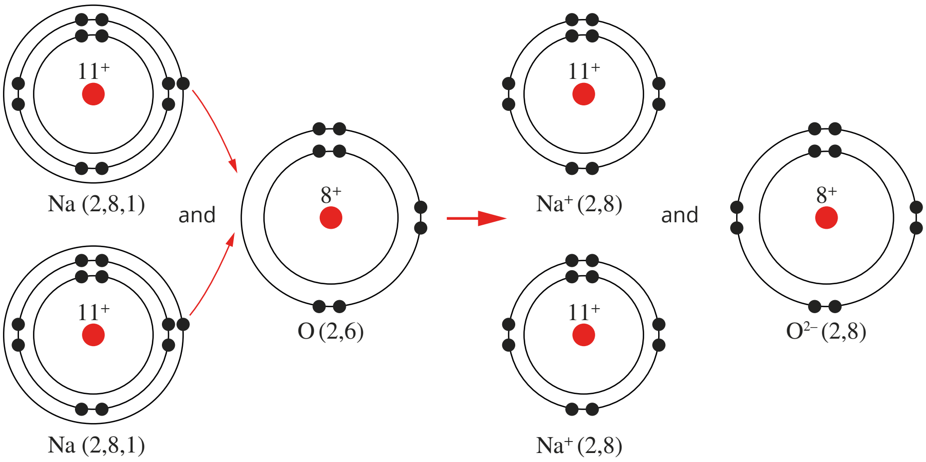
b



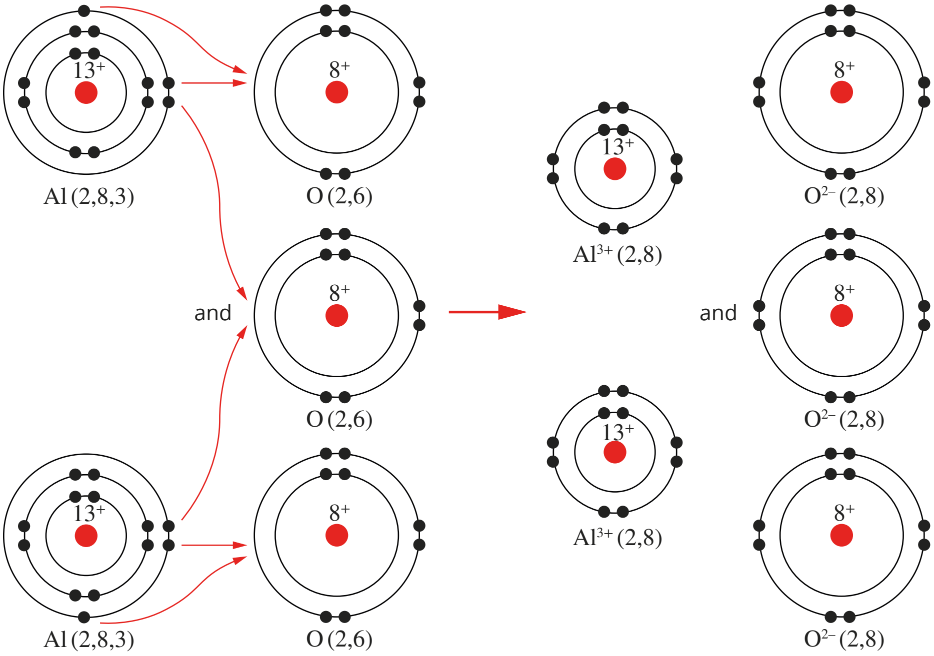
c



d



e



ISBN 978 1 4886 1772 0

3 Group 2 metals have two electrons in their valence shell. They lose these two electrons and therefore become positively charged, as they still have the original number of protons but have lost two electrons. Cations are ions with a positive charge.

4 The electron configuration of an atom of potassium is 2,8,8,1 and that of an atom of chlorine is 2,8,7. Because an atom of K has one more electron than an atom of a noble gas and an atom of Cl has one less, in a reaction one K atom can donate one electron to one Cl atom to give K^+ (2,8,8) and Cl^- (2,8,8), both of which have the electronic configuration of the noble gas argon. The formula of this compound is, therefore, KCl.

Calcium, however, has the electronic configuration of 2,8,8,2. A Ca atom will lose two electrons to gain a noble gas configuration. Because each Cl atom will gain only one electron, there will be two chlorine atoms for each Ca atom. This reaction will therefore produce Ca^{2+} (2,8,8) ions and Cl^- (2,8,8) ions. The formula of the compound is, therefore, $CaCl_2$.

5 **a** Na (2,8,1) + Cl (2,8,7) → Na^+ (2,8) + Cl^- (2,8,8)
b Mg (2,8,2) + O (2,6) → Mg^{2+} (2,8) + O^{2-} (2,8)
c 2Al (2,8,3) + 3S (2,8,6) → $2Al^{3+}$ (2,8) + $3S^{2-}$ (2,8,8)

Section 5.4 Chemical formulae of simple ionic compounds

Worked example: Try yourself 5.4.1

STEPS IN WRITING A CHEMICAL FORMULA

Determine the chemical formula of the ionic compound formed between barium and fluoride ions. You may need to refer to Tables 5.4.1 and 5.4.2 on page 118.

Thinking	Working
Write the symbol and charge of the two ions forming the ionic compound.	Ba^{2+} and F^-
Calculate the lowest common multiple of the two numbers in the charges of the ions.	2 × 1 = 2
Calculate how many positive ions are needed to equal the lowest common multiple.	one Ba^{2+} ion
Calculate how many negative ions are needed to equal the lowest common multiple.	two F^- ions
Use the answers from the previous two steps to write the formula for the ionic compound. Write the symbol of the positive ion first. (Note that 1 is not written as a subscript.)	BaF_2

5.4 KEY QUESTIONS

1 B. Potassium sulfide. Atoms A and B are combining in a ratio of 2:1. Therefore the charge on the A ion must be +1 and the charge on the B ion must be –2. If A^+ has the same electron configuration as argon (2,8,8), then A^+ must be the potassium ion. B^{2-} has the same electron configuration so therefore must be a sulfide ion.

2 **a** 2:1
b 1:3
c 3:2
d 1:1
e 1:2

ISBN 978 1 4886 1772 0

3 **a** NaCl
  b KBr
  c $ZnCl_2$
  d K_2O
  e $BaBr_2$
  f AlI_3
  g AgBr
  h ZnO
  i BaO
  j Al_2S_3

4 **a** potassium chloride
  b calcium oxide
  c magnesium sulfide
  d potassium oxide
  e sodium fluoride

Section 5.5 Writing formulae of more complex ionic compounds

5.5 KEY QUESTIONS

1 **a** Na_2CO_3
  b $Ba(NO_3)_2$
  c $Al(NO_3)_3$
  d $Ca(OH)_2$
  e $Zn(SO_4)_2$
  f KOH
  g KNO_3
  h $ZnCO_3$
  i K_2SO_4
  j $Ba(OH)_2$

2 **a** CsCl
  b Fe_2O_3
  c CuO
  d $Cr_2(SO_4)_3$
  e FeO
  f Al_2O_3
  g $Ca(NO_3)_2$

3 **a** magnesium hydroxide
  b sodium carbonate
  c iron(II) sulfate
  d copper(II) sulfate
  e barium nitrate
  f copper(I) sulfate
  g iron(III) sulfate
  h ammonium nitrate
  i sodium hydrogenphosphate

ISBN 978 1 4886 1772 0

CHAPTER 5 REVIEW

1 **a** Assemble equipment to test conductivity. Add a globe to the circuit. When the electrodes are touching the solid magnesium chloride, the globe will not light up.

b Using the same equipment with molten sodium chloride, the globe will glow. Care is needed, as sodium chloride melts at 801°C.

c If a crystal of sodium chloride was hit firmly with a hammer, it would shatter. Again, care is needed—safety glasses must be worn.

2 **a** both metallic and ionic lattices

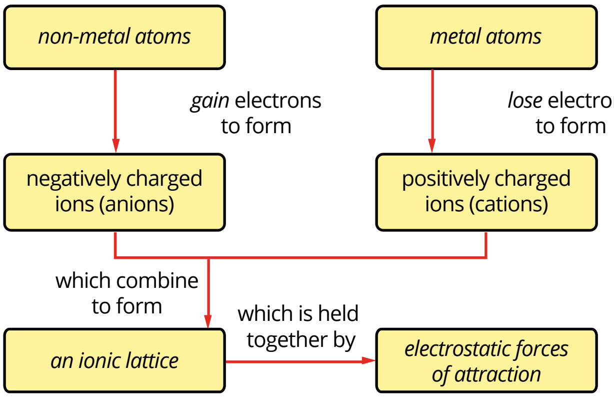
b both metallic and ionic lattices

c both metallic and ionic lattices

d ionic lattices only

e both metallic and ionic lattices

3



4 **a** The electrostatic forces of attraction between the positive and negative ions are strong and will be overcome only at high temperatures.

b The strong electrostatic forces of attraction between the ions mean that a strong force is needed to break up the lattice, giving the ionic crystals the property of hardness. However, the crystal lattice will shatter when a strong force is applied, suddenly causing ions of like charge to become adjacent to each other and be repelled.

c In the solid state, the ions are not free to move. However, when the solid melts or dissolves in water, the ions are free to move and conduct electricity.

5 **a** Na^+, Cl^-; Mg^{2+}, O^{2-}

b MgO. More energy is required to overcome the stronger forces. The higher melting temperature therefore reflects the solid with stronger forces between particles.

c The strength of electrostatic attraction between ions will depend on the size of the ions and on their charge. The Mg^{2+} ion is slightly smaller than the Na^+ ion, and the O^{2-} ion is much smaller than the Cl^- ion. More importantly, the Mg^{2+} ion and the O^{2-} ion each have twice the charge of the Na^+ ion and the Cl^- ion. The attraction between the ions in MgO is therefore much stronger than in NaCl. Magnesium oxide therefore has a much higher melting temperature.

6 The strength of the forces remains unchanged, but the kinetic energy of the ions increases until the forces can no longer hold the ions in the solid lattice, and the lattice breaks up as the solid melts.

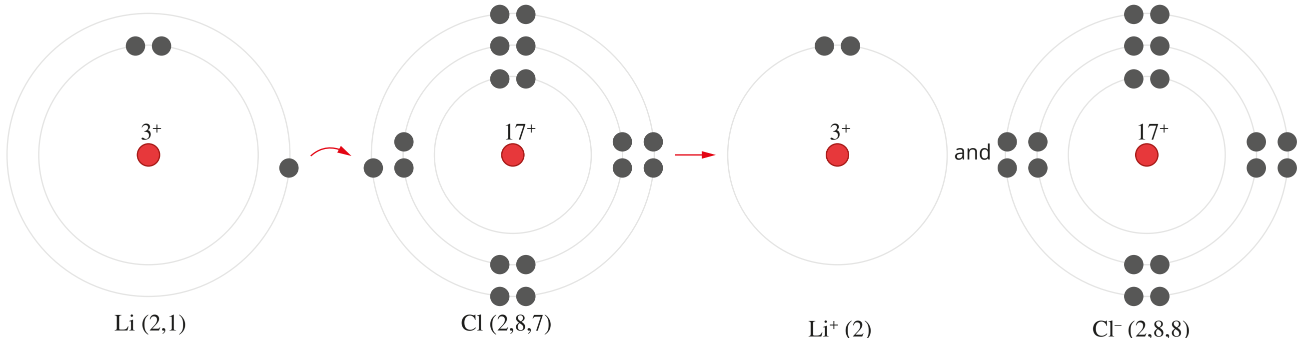
7 **a** Solid ionic compounds do not conduct electricity.

b Ionic compounds are hard.

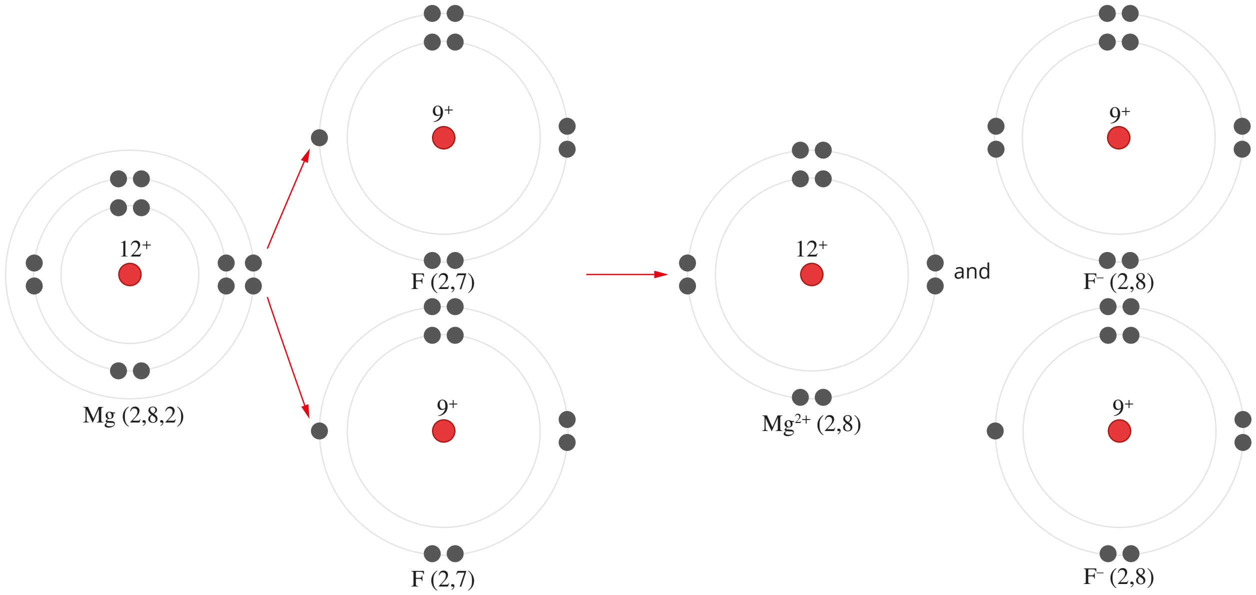
c In solution, ionic compounds conduct electricity.

ISBN 978 1 4886 1772 0

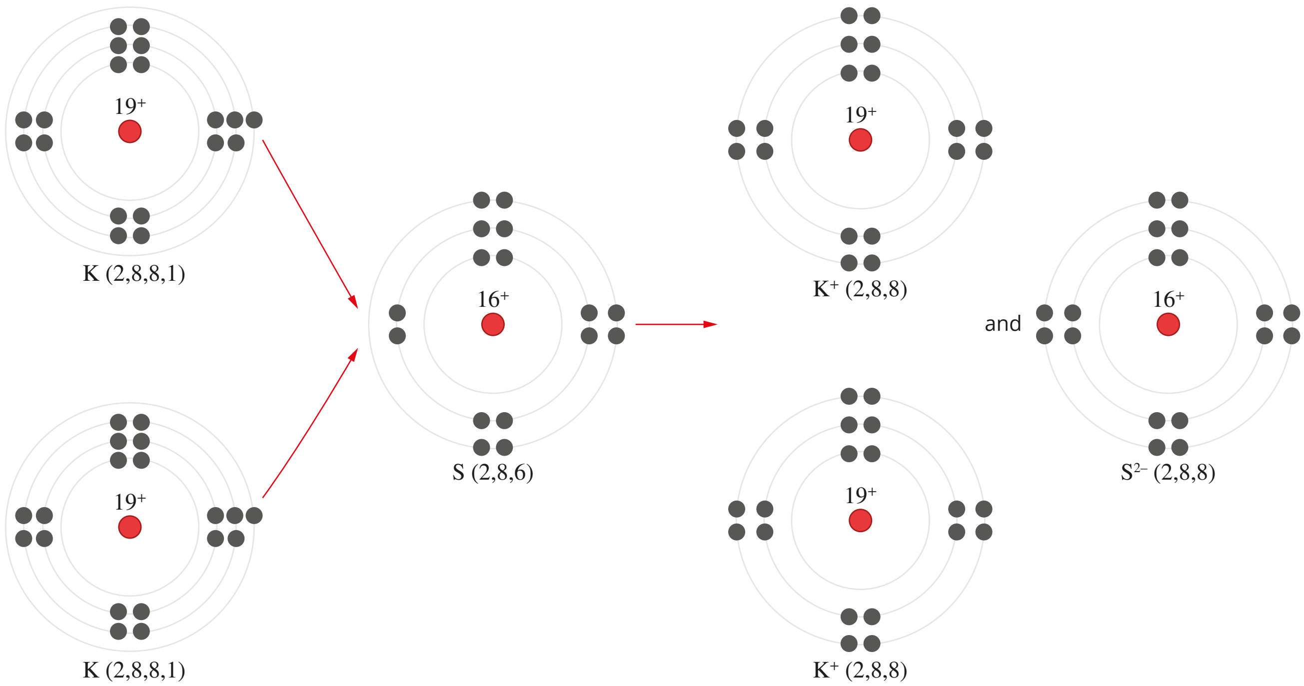
8 **a**



b

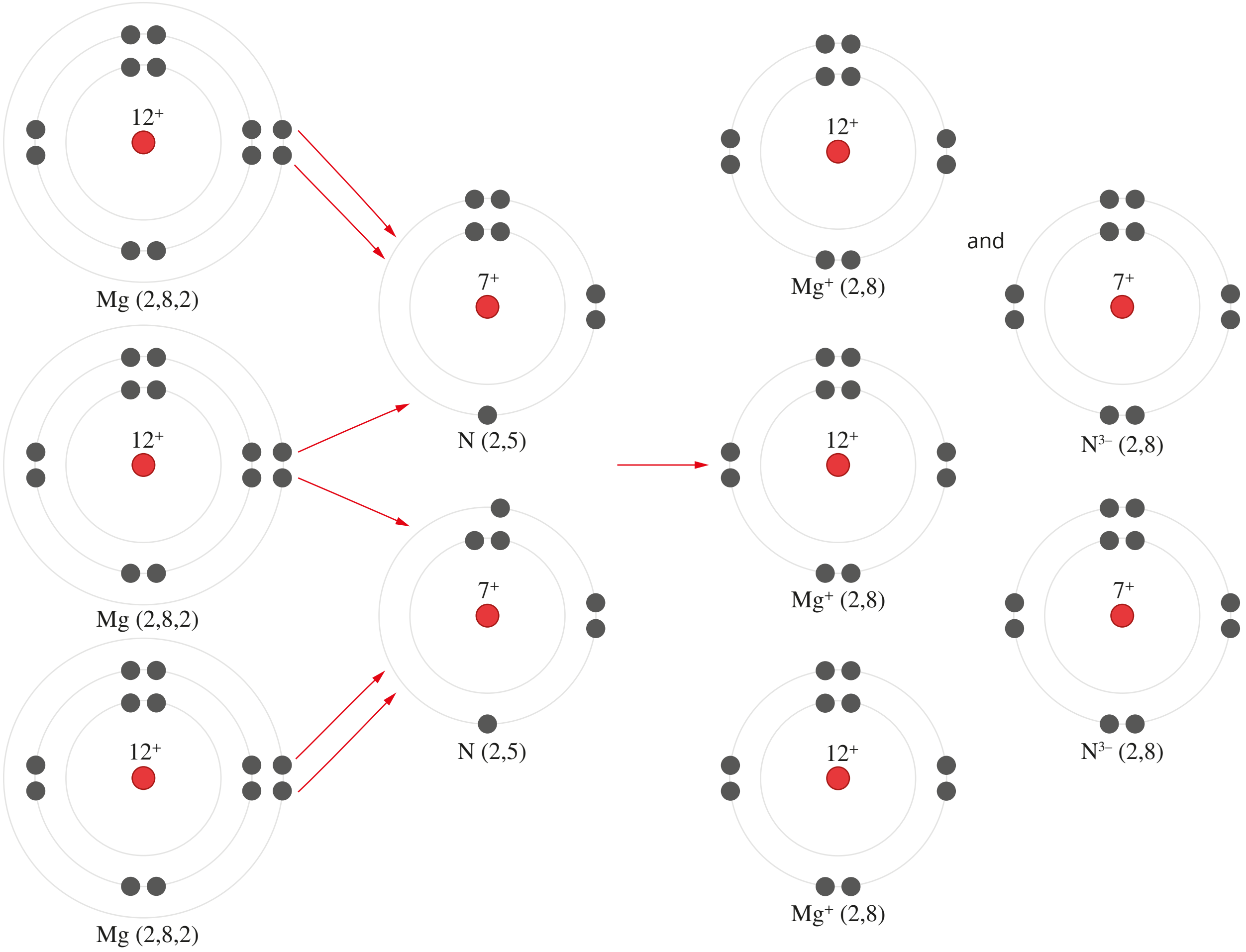


c



ISBN 978 1 4886 1772 0

d



9 **a** 2,8
b 2,8
c 2,8
d 2,8

10 **b** CD_3
c EF
d G_3H
e KL

11 **a** $MgCl_2$ or MgF_2
b NaCl or CaS
c Na_2O or K_2S
d Na_3N or Li_3N
e $AlCl_3$ or AlF_3
f Mg_3N_2

12 Elements in group 17 of the periodic table have seven electrons in their outer shell so only need to gain one electron to satisfy the octet rule. This means they become negative by gaining one electron.

13 **a** Mg (2,8,2) + 2Cl (2,8,7) → Mg^{2+} (2,8) + $2Cl^-$ (2,8,8)
b 2Al (2,8,3) + 3O (2,6) → $2Al^{3+}$ (2,8) + $3O^{2-}$ (2,8)

14 **a** KBr. Potassium ion has a charge of +1, bromide ion has –1.
b MgI_2. Magnesium ion has a charge of +2, iodide ion has –1.
c CaO. Calcium ion has a charge of +2, oxide ion has –2.
d AlF_3. Aluminium ion has a charge of +3, fluoride ion has –1.
e Ca_3N_2. Calcium ion has a charge of +2, nitride ion has –3.

15 **a** CuCl
b Ag_2O
c Li_3N
d KI

16 The subscripts represent the ratio of metal to non-metal ions in the ionic compound.

17 **a** **i** –3
ii +1
iii –2
b **i** Y_2SO_4
ii K_2Z
iii Y_3X
iv Y_2Z

18 **a** $CuNO_3$
b CrF_2
c K_2CO_3
d $Mg(HCO_3)_2$
e $Ni_3(PO_4)_2$

19 **a** ammonium carbonate
b copper(II) nitrate
c chromium(III) bromide

20 **a** Agree. Both metallic and ionic lattices do contain positive ions in a regular arrangement.
In a metallic lattice, the positive ions are surrounded by delocalised electrons; in an ionic lattice, negative ions alternate with the positive ions.

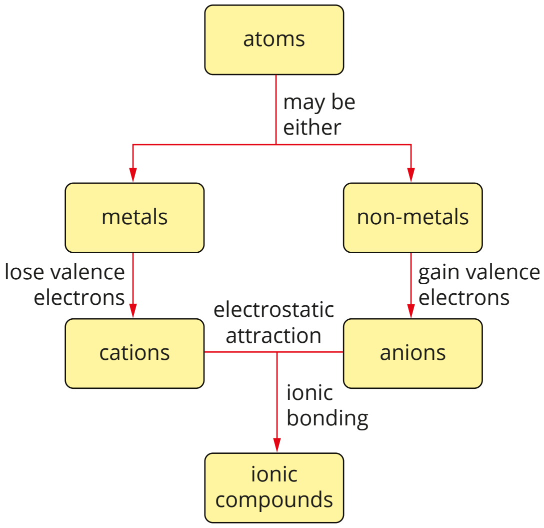
b Agree. In a metallic lattice, each positive ion attracts the delocalised electrons in its region, and each delocalised electron is attracted to all neighbouring positive ions. In an ionic lattice, each positive ion is attracted to the negative ions that surround it, and vice versa.

c Agree. In a metallic solid, there will be repulsion between the positive ions, and between the delocalised electrons. The particles are arranged to minimise these repulsions. In an ionic lattice, the arrangement of alternating positive and negative ions also minimises repulsion between like charges.

d Agree. The energy required to remove the outer electron(s) is known as the ionisation energy, with each electron removed having a specific ionisation energy. Metals with low ionisation energy, such as sodium and potassium, form positive ions more readily than metals with high ionisation energies.

e Disagree. Although this is true, it is not the reason ionic solids do not conduct. In a metal, delocalised electrons are free to move so it conducts electricity; in an ionic solid, the ions are not free to move so it does not conduct electricity.

21 A possible answer is shown.



ISBN 978 1 4886 1772 0

Chapter 6 Materials made of molecules

Section 6.1 Properties of non-metallic substances

6.1 KEY QUESTIONS

1 Non-metallic elements: H_2 and Br_2 as they consist of only one type of atom. Non-metallic compounds: NO_2 as it consists of more than one type of atom (nitrogen and oxygen).

2 A molecule is a discrete group of atoms of known formula, bonded together.

3 **a** Non-metals do not conduct electricity because they do not contain free-moving charged particles (neither delocalised electrons nor ions).

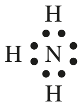
  b Non-metals have low melting and boiling temperatures because they have weak intermolecular forces between molecules.

4 When sugar turns to a liquid, it is melting; the intermolecular forces between sugar molecules are broken. When the liquid turns black and a gas is produced, the bonds between atoms in the sugar molecules are broken, allowing new substances to be produced.

Section 6.2 Covalent bonding

Worked example: Try yourself 6.2.1

LEWIS STRUCTURE DIAGRAMS

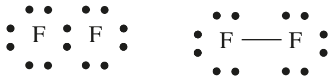
Draw a Lewis structure diagram of ammonia (NH_3).	
Thinking	**Working**
Write the electronic configuration of the atoms in the molecule.	N electronic configuration: 2,5 H electronic configuration: 1
Determine how many electrons each atom requires for a stable valence shell.	N requires 3 electrons. H requires 1 electron.
Draw a Lewis structure diagram of the likely molecule, ensuring that each atom has a stable valence shell. Electrons not involved in bonding will be in non-bonding (lone) pairs.	Draw a Lewis structure diagram of the molecule. 

ISBN 978 1 4886 1772 0

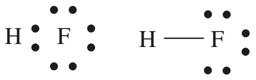
6.2 KEY QUESTIONS

1 **a** One. Hydrogen requires one electron to complete its valence shell, so it forms a single covalent bond.
  b Three. Nitrogen requires three electrons to complete its valence shell, so it forms a triple covalent bond.
  c Two. Oxygen requires two electrons to complete its valence shell, so it forms a double covalent bond.
  d One. Fluorine requires one electron to complete its valence shell, so it forms a single covalent bond.

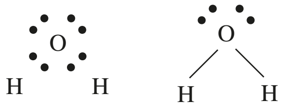
2 **a** fluorine (F_2)



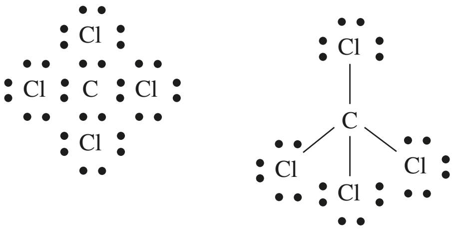
b hydrogen fluoride (HF)



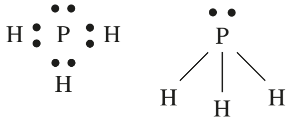
c water (H_2O)



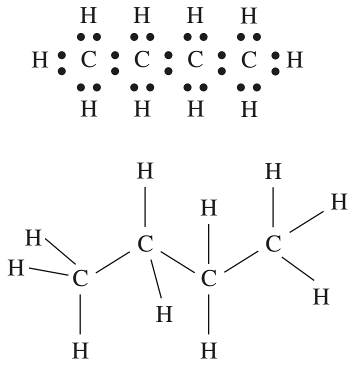
d tetrachloromethane (CCl_4)



e phosphine (PH_3)



f butane (C_4H_{10})



g carbon dioxide (CO_2)

3 **a** one
  b two
  c three
  d four
  e one
  f zero

4 To complete its valence shell, the oxygen atom uses two of its valence-shell electrons to form two single bonds or a double bond with suitable non-metal atoms. The remaining four electrons in the valence shell are not required for bonding, as the valence shell is now complete, and they arrange themselves as two lone pairs around the oxygen atom.

5 **a** CCl_4
  b NBr_3
  c SiO_2
  d HF
  e PF_3

6 **a** True. Methane is formed from two types of non-metal atoms (carbon and hydrogen).
  b False. Methane contains five atoms (one carbon and four hydrogen atoms).
  c True. There are five atoms per molecule and so four molecules will have 20 atoms.
  d False. The forces of attraction between the carbon and hydrogen atoms in the molecule are bonds.

ISBN 978 1 4886 1772 0

7 **a** ball-and-stick model
  b Lewis structure (or electron dot diagram)
  c space-filling model

CHAPTER 6 REVIEW

1 The strength of the intermolecular forces in pure hydrogen chloride must be relatively weak. Since the pure hydrogen chloride exists as a gas at room temperature, it must have a low boiling temperature, which indicates that not much energy is required to break the intermolecular forces between molecules.

2 Non-metallic substances do not conduct electricity.

3 Bonds are the forces that hold the atoms within a molecule together. In carbon dioxide molecules they are the bonds between the carbon atoms and the oxygen atoms. Intermolecular forces are the forces between one molecule and its neighbouring molecules. It is the intermolecular forces that are broken when carbon dioxide sublimes.

4 For an odour to be detected, gas molecules need to enter our nose. Because covalent molecular substances have weak forces of attraction between their molecules some molecules at the surface of a solid or liquid will have enough energy to leave the surface and become gaseous.

5 A is naphthalene. It is a soft, covalent molecular compound.

B is magnesium chloride. As it is a conductor in aqueous solution, brittle and has a high melting point, it is an ionic compound.

C is copper. As it is not soluble in water, has a high melting point and is hard, it is a metal.

D is phosphorus trihydride. As it is a non-conductor in aqueous solution, has a low melting point and is soft as a solid, it is a covalent molecular compound.

6 D. Each hydrogen atom has one electron only. If two atoms share an electron each they both obtain a complete valence shell. A single covalent bond is formed.

7 **a** Six. Each nitrogen atom shares three electrons with the other nitrogen atom.
  b Three. Each nitrogen atom shares three electrons with the other nitrogen atom, so three bonding pairs are formed, which is equal to a triple covalent bond.
  c Four. Each nitrogen atom has two non-bonding electrons, making a total of four when both atoms are counted.

8 B. To have a formula XY_4, atom X must require four more valence-shell electrons and atom Y must require one. The likely molecule is CH_4 as carbon has four valence-shell electrons and hydrogen requires one more. Oxygen has six valence electrons and needs to form two covalent bonds. Hydrogen has one valence electron and needs to form one bond. For OH_4, each hydrogen would need to form a single covalent bond, giving a total of four, but oxygen only needs to form two, so this option is incorrect. Bromine has seven valence electrons. Each bromine needs to form a single covalent bond, giving a total of four, but hydrogen only needs to form one bond, so this option is incorrect. Carbon has four valence-shell electrons and needs to form four bonds. Oxygen has six valence electrons and needs to form two covalent bonds, to have a total of eight, but carbon only needs to form four, so this option is incorrect.

9 Atom X requires two electrons to form a stable valence shell. Each atom of element Y can share one electron. Therefore the molecular formula will be XY_2.

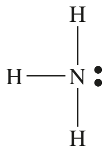
10 X will have seven valence electrons. If X has seven valence electrons, each X atom can share one electron with one of the four Si valence electrons so giving Si its octet. As well, each X atom will have an octet.

11 Neon will not form bonds to other atoms as it has a stable valence shell.

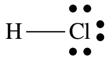
12 **a** zero

$\mathrm{H{-}H}$

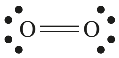
  b one



  c three

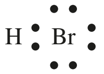


  d four

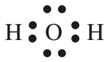


13 The diagrams of each molecule below show the number of electron pairs.

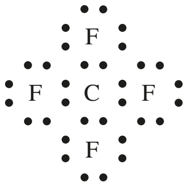
a two bonding electrons, six non-bonding electrons



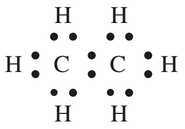
b four bonding electrons, four non-bonding electrons



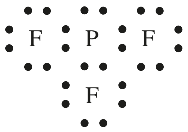
c eight bonding electrons, 24 non-bonding electrons



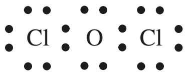
d 14 bonding electrons, zero non-bonding electrons



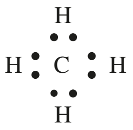
e six bonding electrons, 20 non-bonding electrons



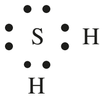
f four bonding electrons, 16 non-bonding electrons



g eight bonding electrons



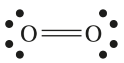
h four bonding electrons, four non-bonding electrons

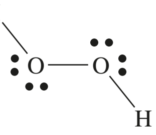


14 a The bonds are similar in that they all involve the sharing of electron pairs between two atoms; that is, they are covalent bonds.

b They differ in the number of electron pairs shared: one pair (fluorine), two pairs (oxygen) and three pairs (nitrogen).

15 a





b In a single bond there are two electrons that are shared between the atoms, whereas in a double bond there are four electrons. The net attraction between the shared electrons and the protons in the nuclei of the atoms in the bond is stronger and more energy is needed to break the double bond than the single bond.

c Oxygen has six valence-shell electrons and completes its valence shell by forming two single covalent bonds or a double bond with another non-metal. There is no need for the oxygen to form a triple bond or three single covalent bonds as this exceeds its requirement of two electrons to complete its valence shell.

16 Metallic: Ag, Cu; ionic: $CuCl_2$, CaS; molecular: NH_3, HCl, H_2O

Ag and Cu are both metals. This can be determined from general knowledge, their position in the periodic table or from their electronic configuration.

$CuCl_2$ and CaS are both formed from a combination of a metal and non-metal(s). This is characteristic of ionic compounds.

NH_3, HCl, H_2O are compounds made from non-metals. They are composed of small molecules.

17 Dissolve a small quantity of both substances in water and test for electrical conductivity. The solution that conducts electricity will be calcium chloride and the non-conductor will be glucose.

18 The nuclei of the two atoms involved in the bond each have an (electrostatic) attraction for the shared electrons due to their opposite charges. This holds the atoms together in the molecule.

19 **a** True. Carbon dioxide is formed from two types of non-metal atoms (carbon and oxygen).

- **b** True. Carbon dioxide contains three atoms (one carbon and two oxygen atoms).
- **c** False. There are two oxygen atoms per molecule and so 33 molecules will have 66 oxygen atoms.
- **d** True. The forces of attraction in the molecule are covalent bonds.
- **e** False. There are two double covalent bonds in a molecule of carbon dioxide.
- **f** True. Each oxygen atom has two lone pairs of electrons. As there are two oxygen atoms, there are four lone pairs of electrons in a molecule of carbon dioxide.

ISBN 978 1 4886 1772 0

Chapter 7 Carbon

Section 7.1 Carbon lattices

7.1 KEY QUESTIONS

1 It can form four covalent bonds. Carbon atoms can form single, double and triple covalent bonds. Carbon atoms can bond to each other.

2 **a** To sublime is to turn from a solid directly into a gas.

b Diamond and graphite contain extended networks of strong covalent bonds, which must be overcome to allow the material to sublime.

3 **a** Diamond is hard because it has strong covalent bonds throughout the lattice, with all atoms being held in fixed positions.

Graphite is soft because there are weak dispersion forces between the layers in graphite, so layers can be made to slide over each other easily.

b Diamond is a non-conductor of electricity because all of its electrons are localised in covalent bonds and are not free to move.

Graphite is able to conduct electricity because it has delocalised electrons between its layers of carbon atoms.

4 **a** Graphite is used as a dry lubricant because the dispersion forces between the layers in graphite enable the layers to slide over each other easily and to reduce the friction between moving parts.

b The strong covalent bonding throughout the lattice means that the carbon atoms are fixed in place. This makes diamond very hard and suitable as a material for cutting softer materials.

Section 7.2 Carbon nanomaterials

7.2 KEY QUESTIONS

1 Fullerenes are similar to graphite in that they are allotropes of carbon in which each carbon atom has bonds to three other carbon atoms. Fullerenes conduct electricity and heat, similar to graphite. Fullerenes differ from graphite in that they are nanomaterials, which give them different physical properties.

2 Each carbon atom in a buckyball is covalently bonded to three other carbon atoms. Each carbon atom has one free electron, which is shared between the other carbon atoms in the sphere. The structure consists of hexagonal and pentagonal rings of atoms.

3 Each carbon atom in a graphene sheet is covalently bonded to three other carbon atoms. Each carbon atom has one free electron that is shared with the other carbon atoms in the sheet. The structure consists of hexagonal rings of atoms. The sheet is one atom thick but may be of any length and width.

4 Carbon nanotubes are immensely strong for their mass. They can theoretically be produced with any length, allowing for very long cables to be produced. A space elevator would require an extremely strong, long and uniform cable to be produced.

CHAPTER 7 REVIEW

1 Carbon exists in different forms with different arrangements of atoms.

2 Refer to Figure 7.1.10. Carbon atoms bond covalently to three other carbon atoms in graphite to form layers. The strong covalent bonds between the atoms in each layer explain graphite’s resistance to heat and, hence, why graphite sublimes at high temperature. These layers consist of hexagonal rings connected to each other. The fourth electron in each carbon atom is delocalised, which explains its electrical conductivity. There are weak dispersion forces between layers of graphite, allowing the layers to slide over each other and enabling it to act as a lubricant.

3 Charcoal has many pores and pockets giving it a very large surface area. The surface of charcoal may form dispersion forces with the coloured contaminants. The coloured contaminants may also be trapped in the internal pores of the charcoal.

ISBN 978 1 4886 1772 0

4 Graphite can be used under very low temperature conditions because it does not freeze (unlike the potential for oil-based lubricants). It can be used under very high temperature conditions because it has a very high sublimation temperature.

5 Diamond consists of an extended three-dimensional network of strong covalent bonds, which must be overcome for sublimation to occur.

6 Silicon dioxide forms a covalent network lattice with each silicon atom bonded tetrahedrally to four oxygen atoms and each oxygen atom bonded to two silicon atoms. The high melting point and hardness are due to the large amount of energy needed to break these strong covalent bonds. Its non-conductivity is due to the absence of any mobile charged particles in the structure. All valence electrons are involved in forming covalent bonds and so are not free to move through the structure.

7 **a** Methane is an example of a molecular substance. It has strong, covalent intramolecular bonds and very weak intermolecular bonds. Diamond is an example of a covalent network lattice. It has strong covalent bonds throughout its structure.

b Due to the weak dispersion forces between molecules, methane will have extremely low boiling and melting points. If it were a solid it would be crystalline, brittle and soft. Due to its extended covalent lattice, diamond is extremely hard, does not exist as a liquid and has a very high sublimation point.

8 Carbon atoms in diamond have a tetrahedral bond geometry; carbon atoms in graphene are arranged with one atom at the centre and three atoms at the corners of an equilateral triangle.

9 It is composed of single layers and the covalent bonding in each layer is very strong.

10 Fullerenes are an allotrope of carbon and carbon atoms have four valence electrons. In fullerenes, each carbon atom forms single covalent bonds with three neighbouring carbon atoms. This leaves one delocalised valence electron in each carbon atom that can move throughout the fullerene, allowing it to conduct electricity.

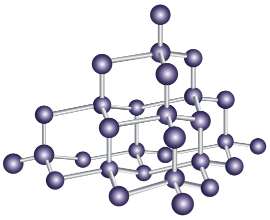
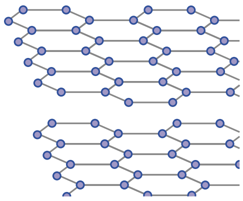
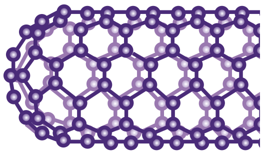
11 A carbon nanotube is like a cylinder of graphene with half a buckyball on each end.

12 Carbon nanotubes are strong due to their chain of unbroken covalent carbon–carbon bonds. As well, each nanotube is a single large molecule which means it does not have any weak points such as boundaries between crystalline grains.

13 Carbon nanotubes have a very high surface area. Thus coating a nanotube with a catalyst would give a very high surface area of catalyst for reactants to come into contact with.

14 O_2 and O_3; S_6 and S_8

15

				
allotrope name	diamond	graphite	carbon nanotube	graphene
bonding feature	Each carbon is bonded to four others in a network lattice structure.	Each carbon is bonded to three others in a layered structure.	Each carbon is bonded to three others in a cylindrical shape.	Each carbon is bonded to three others in a layer one atom thick.
properties	hard, brittle	conducts electricity, soft	very strong, conducts electricity	very strong, conducts electricity

16 The common structural feature of graphene and graphite is their layered structure. Both contain layers of carbon atoms with each carbon bonded to three neighbouring atoms, with three covalent bonds. Each carbon atom still has another valence electron that is delocalised in the layer, explaining the electrical conductivity of both allotropes. The bonding between the atoms in the layer is strong. Graphene is just a single layer but graphite contains layer upon layer.

This gives rise to the different properties of graphite and graphene. Graphene is very strong. There are dispersion forces between layers in graphite, allowing the layers to slide past each other. Graphite can therefore act as a lubricant. Both allotropes have a high sublimation point.

17 **a** An electric circuit is used that contains a light globe and a power supply. When the power supply is switched on, the light globe will glow if the circuit is complete. The diamond must conduct electricity for the circuit to be complete and the light globe to glow. For this to occur, the diamond must contain charged particles that are free to move. Since the light globe does not glow, the diamond is not an electrical conductor as it does not contain free-moving charged particles.

b Graphite, graphene and nanotubes are all conductive of electricity. Diamond is non-conductive of electricity.

18 Germanium, like carbon, is a group 14 element so has four valence electrons. This enables germanium atoms to bond to other germanium atoms to form covalent network bonds in the same way as carbon.

19 **a** Silicon and carbon are both group 14 elements so have four valence electrons. This enables them to form four covalent bonds to other carbon and/or silicon atoms to form a covalent network lattice.

b Silicon carbide would be expected to have a high melting point, and to be hard, insoluble in water and most solvents, and a non-conductor of electricity.

20 The models of the structures of graphite, solid copper and an aqueous solution of sodium chloride all have mobile charged particles enabling them to conduct an electrical current. In an aqueous solution of sodium chloride, it is sodium ions and chloride ions that carry the charge, while in graphite and solid copper it is delocalised valence electrons. In solid copper, the valence electrons are not fixed to one atom but can move freely within the structure of the copper. The copper atoms are described as existing in a ‘sea’ of delocalised valence electrons. In graphite, each carbon atom has three of its four valence electrons covalently bonded to three other carbon atoms with the fourth valence electron delocalised, enabling an electric current to be carried by these delocalised electrons.

ISBN 978 1 4886 1772 0

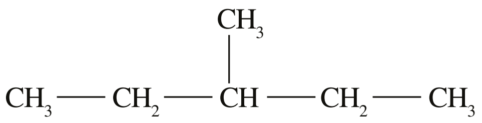
Chapter 8 Organic compounds

Section 8.1 Alkanes

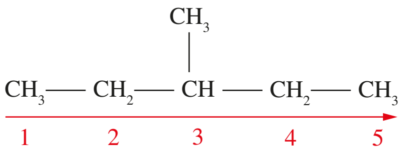
Worked example: Try yourself 8.1.1

IUPAC NAMING SYSTEM FOR ALKANES

Write the systematic name for the following molecule.



$CH_3-CH_2-\overset{\overset{CH_3}{|}}{CH}-CH_2-CH_3$

Thinking	**Working**
Identify the longest carbon chain in the molecule. The name of the molecule is based on this longest chain.	There are five carbons in the longest chain. The stem name is based on pentane.
Number the carbons starting from the end closest to the branch.	 $\underset{1}{CH_3}-\underset{2}{CH_2}-\underset{3}{\overset{\overset{CH_3}{|}}{CH}}-\underset{4}{CH_2}-\underset{5}{CH_3}$ Note: In this molecule, it does not matter which end the numbering begins at as the side branch comes off the central carbon, which is the third carbon.
Identify the branch.	The side chain is a methyl group.
Combine all components.	The name of the molecule is 3-methylpentane.

8.1 KEY QUESTIONS

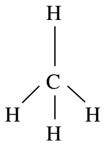
1 A, D

2 **a** CH_4

b Methane is a hydrocarbon because it is a compound of carbon and hydrogen.

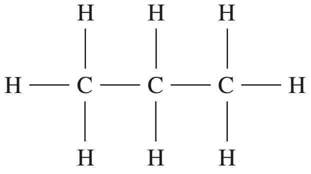
c Carbon has the electron configuration of 2,4. Each carbon atom needs four electrons to complete its outer shell. Hydrogen has an electron configuration of 1. Each hydrogen atom needs one electron to complete its outer shell. Because both atoms need electrons, they will share electrons, that is, form covalent bonds. It will take four hydrogen atoms to provide the four electrons required by each carbon atom.

d The hydrogen atoms are arranged around the central carbon atom in a tetrahedral configuration. This tetrahedral arrangement gives minimum electrostatic repulsion between the four pairs of bonding electrons.



3 **a** Propane. This hydrocarbon contains three carbon atoms so the stem name is prop-. Its formula fits the general formula of an alkane, C_nH_{2n+2} ($n = 3$, so $2 \times 3 + 2 = 8$). Hence the name of the hydrocarbon is propane.

b



c $CH_3CH_2CH_3$

4 **a** Butane. The longest chain of carbon atoms is four and there are no branches.

b Heptane. The longest chain of carbon atoms is seven and there are no branches. Note that the prefix hept- rather than sept- is used to denote seven.

c 2-methylpentane. The longest chain of carbon atoms is five, giving the stem name of pentane, and there is one methyl, $-CH_3$, branch on carbon number two when numbered from the end closest to the branch.

d 2-methylbutane. The longest chain of carbon atoms is four, giving the stem name of butane, and there is one methyl, $-CH_3$, branch on carbon number 2 when numbered from the end closest to the branch.

e 2,4-dimethylpentane. The longest chain of carbon atoms is five, giving the stem name of pentane, and there are two methyl, $-CH_3$, branches: one on carbon number 2 and the other on carbon number 4 when numbered in either direction. Dimethyl is used in addition to the two numbers to indicate the presence of two separate methyl groups.

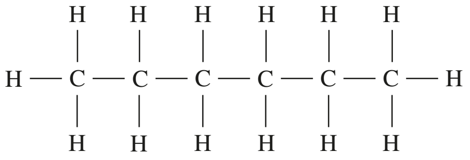
5 **a** butane

b methylpropane

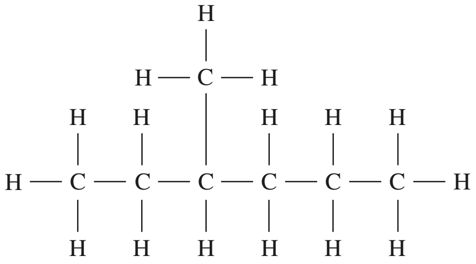
c 2,4-dimethylpentane

d 3-methylhexane

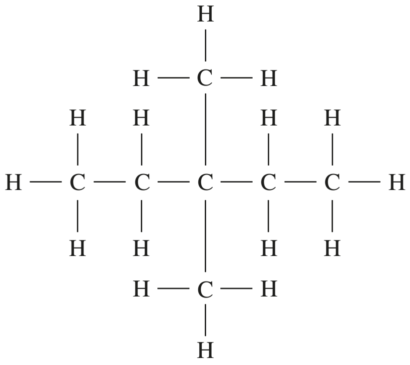
6 **a** hexane



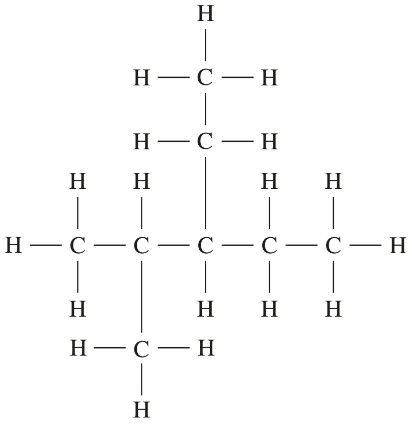
b 3-methylhexane



c 3,3-dimethylpentane



d 3-ethyl-2-methylpentane

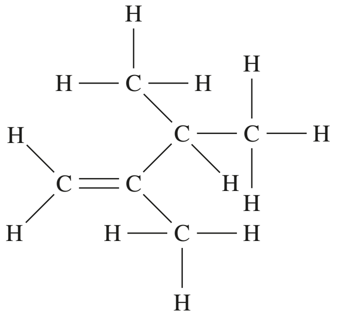


Section 8.2 Alkenes

Worked example: Try yourself 8.2.1

IUPAC NAMING SYSTEM FOR ALKENES

Write the systematic name for the following molecule.



Thinking	Working
Identify the longest carbon chain in the molecule that contains the double or triple bond. The name of the molecule is based on this longest chain.	There are four carbons in the longest chain. The name is based on butene.
Identify whether the compound exhibits geometric isomerism.	It does not exhibit geometric isomerism.
Number the carbon atoms starting from the end closest to the double bond. Note the position of any double bond.	There is a double bond on carbon number 1, so the longest chain is but-1-ene.
Identify each branch and the number carbon that it is on.	The side chains are methyl groups and they are on carbon number 2 and carbon number 3.
Combine all components.	The name of the molecule is 2,3-dimethylbut-1-ene.

8.2 KEY QUESTIONS

1 ethene (C_2H_4), propene (C_3H_6), methylpropene (C_4H_8), pentene (C_5H_{10}), octene (C_8H_{16})

2 **a** Four. The structure of methylpropene is based on propene with a methyl group attached.

b eight

c Methylpropene is an isomer of butene (C_4H_8).

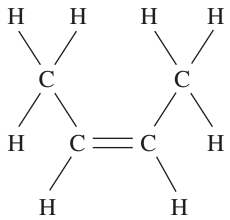
3 B. The condensed structural formula should show all atoms, but not the bonds or lone pairs of electrons.

4 **a** 4-methylpent-1-ene

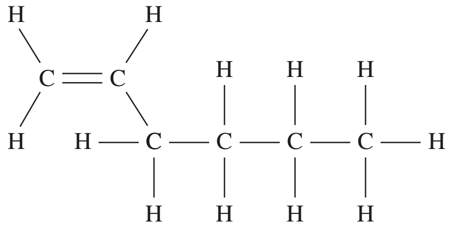
b pent-2-ene

c 2-methylbut-1-ene

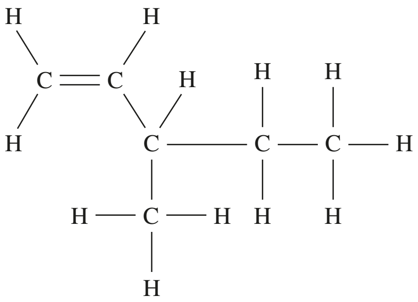
5 **a**



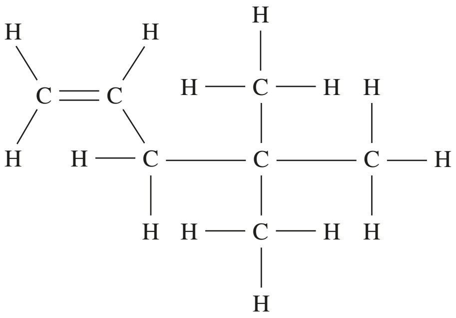
b



c



d



Section 8.3 Benzene

8.3 KEY QUESTIONS

1 D. All bonds are the same length, in-between a single and double bond.

2 C. The six carbon atoms are bonded in a hexagonal structure, with one delocalised electron per atom. These electrons are free to move within the ring structure.

Section 8.4 Reactions of hydrocarbons

8.4 KEY QUESTIONS

1 B

2 $CH_4 + Br_2 \rightarrow CH_3Br + HBr$

3 $CH_2CHCH_2CH_3 + Cl_2 \rightarrow CH_2ClCHClCH_2CH_3$

4 $C_6H_6 + Br_2 \xrightarrow{AlBr_3} C_6H_5Br + HBr$

5 It must be an unsaturated compound, so B is the correct answer.

CHAPTER 8 REVIEW

1 Carbon can form a large number of compounds for several reasons. Carbon atoms can use two electrons each to bond to form very long chains. This leaves the other two valence electrons able to bond other non-metal atoms onto the chain. In addition, there can be double and triple bonds between carbon atoms as well as ring structures. All these possible variations mean that an enormous number of different compounds containing carbon can be formed.

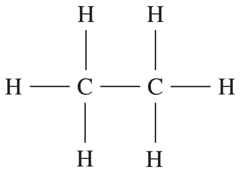
2 A

3 **a** alkanes

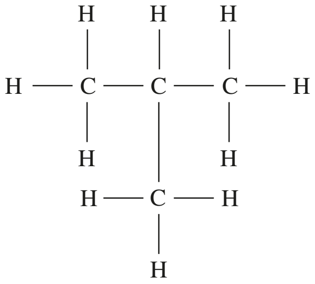
- **b** $C_{17}H_{36}$
- **c** $C_{15}H_{32}$

ISBN 978 1 4886 1772 0

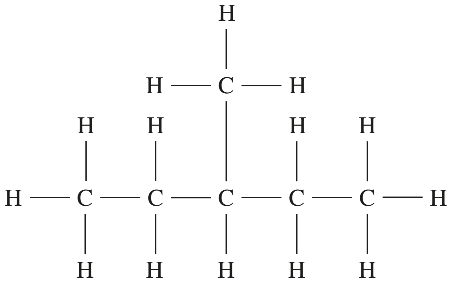
4 **a** ethane



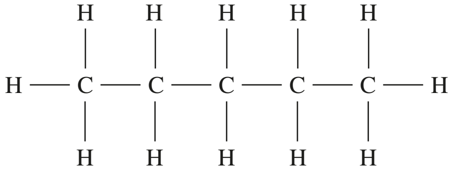
b methylpropane



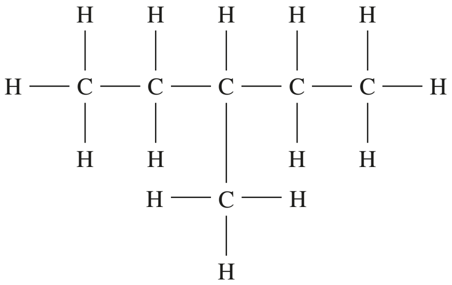
c 3-methylpentane



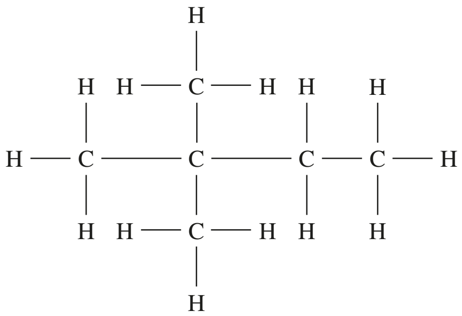
d pentane



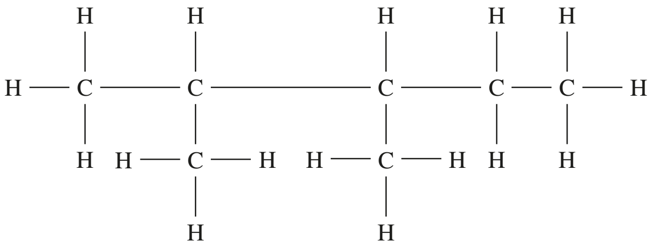
5 **a**



b



c



ISBN 978 1 4886 1772 0

6 Hexane molecules are larger than propane molecules, so hexane has a higher boiling point than propane.

7 **a** alkane

b alkene

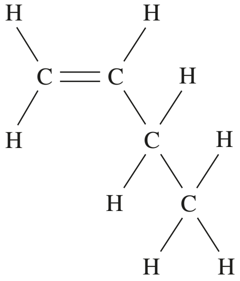
c alkane

d alkene

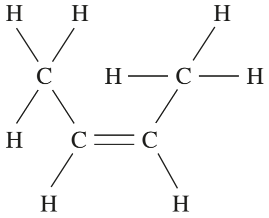
e alkane

8

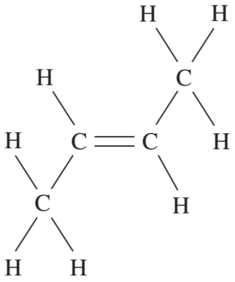
but-1-ene



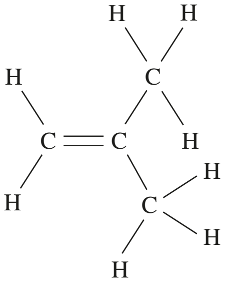
cis-but-2-ene



trans-but-2-ene



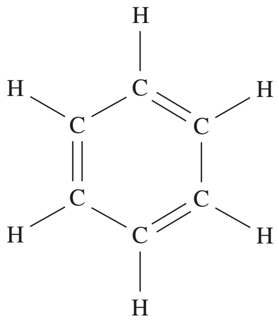
methylpropene



9 **a** *cis*-hex-2-ene

b 4-methylpent-2-ene

c propene

10 

The Kekulé structure of benzene depicts the six-carbon ring as containing alternating double and single bonds. This is incorrect as experimental results show that each carbon–carbon bond is in fact of equal length. Furthermore, benzene does not readily undergo addition reactions, as it would if it did contain three double bonds. Instead, it has a similar reactivity to saturated compounds. The benzene ring in fact contains six single bonds and six delocalised electrons (one per atom).

11 B. It adheres to the general formula of C_6H_6 for an aromatic compound, with a bromine atom substituted for a hydrogen.

12 **a** $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

b $2C_6H_{14}(l) + 19O_2(g) \rightarrow 12CO_2(g) + 14H_2O(g)$

c Methylpropane is an isomer of butane, C_4H_{10}: $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$

d $C_{31}H_{64}(s) + 47O_2(g) \rightarrow 31CO_2(g) + 32H_2O(g)$

13 **a** $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$

b $C_3H_6(g) + Br_2(aq) \rightarrow C_3H_6Br_2(aq)$

c $C_4H_8(g) + H_2O(l) \rightarrow C_4H_9OH(l)$

d $C_2H_4(g) + HCl(g) \rightarrow C_2H_5Cl(l)$

14 You could distinguish between them by adding a few drops of orange bromine water to each. The hex-2-ene will rapidly decolourise the bromine water, whereas the hexane will not readily react.

15 Substitution: $C_2H_6(g) + Cl_2(g) \rightarrow C_2H_5Cl(g) + HCl(g)$

Addition: $C_2H_4(g) + HCl(aq) \rightarrow C_2H_5Cl(g)$

16 10 Br_2 molecules

17 $C_2H_4(g) + H_2O(g) \rightarrow CH_3CH_2OH(g)$

18 **a** An alkene contains one double carbon–carbon bond, which requires two carbon atoms. The first alkene is therefore ethene.

b The carbon atom has four electrons in the outer shell, which are available for sharing with other atoms to produce four covalent bonds.

19 Polyunsaturated: contains many double bonds; mono-unsaturated: contains one double bond; saturated: contains only single bonds between carbon atoms.

20 The compound would be an alkene, as it has the general formula C_nH_{2n}. Alkenes are unsaturated and react readily with bromine water via an addition reaction. The orange bromine water would rapidly decolourise.

Chapter 9 The mole

Section 9.1 Masses of particles

Worked example: Try yourself 9.1.1

CALCULATING THE RELATIVE MOLECULAR MASS OF MOLECULES

Calculate the relative molecular mass of nitric acid (HNO_3).	
Thinking	**Working**
Use the periodic table to find the relative atomic mass for the elements represented in the formula.	$A_r(H) = 1.008$ $A_r(N) = 14.01$ $A_r(O) = 16.00$
Determine the number of atoms of each element present, taking into consideration any brackets in the formula.	1 × H atom 1 × N atom 3 × O atoms
Determine the relative molecular mass by adding the appropriate relative atomic masses.	$M_r = 1 \times A_r(H) + 1 \times A_r(N) + 3 \times A_r(O)$ $= 1 \times 1.008 + 1 \times 14.01 + 3 \times 16.00$ $= 63.02$

Worked example: Try yourself 9.1.2

CALCULATING THE RELATIVE FORMULA MASS OF IONIC COMPOUNDS

Calculate the relative formula mass of copper(II) nitrate ($Cu(NO_3)_2$).	
Thinking	**Working**
Use the periodic table to find the relative atomic mass for the elements represented in the formula.	$A_r(Cu) = 63.55$ $A_r(N) = 14.01$ $A_r(O) = 16.00$
Determine the number of atoms of each element present, taking into consideration any brackets in the formula.	1 × Cu atom 1 × 2 = 2 N atoms 3 × 2 = 6 O atoms
Determine the relative formula mass by adding the appropriate relative atomic masses.	Relative formula mass $= 1 \times A_r(Cu) + 2 \times A_r(N) + 6 \times A_r(O)$ $= 63.55 + 2 \times 14.01 + 6 \times 16.00$ $= 187.6$

9.1 KEY QUESTIONS

1 **a** $M_r = 2 \times A_r(H) + A_r(S) + 4 \times A_r(O)$
$= 2 \times 1.008 + 32.07 + 4 \times 16.00$
$= 98.09$

b $M_r = A_r(N) + 3 \times A_r(H)$
$= 14.01 + 3 \times 1.008$
$= 17.03$

c $M_r = 2 \times A_r(C) + 6 \times A_r(H)$
$= 2 \times 12.01 + 6 \times 1.008$
$= 30.07$

ISBN 978 1 4886 1772 0

2 **a** $M_r = A_r(\text{K}) + A_r(\text{Cl})$
$= 39.10 + 35.45$
$= 74.55$

b $M_r = 2 \times A_r(\text{Na}) + A_r(\text{C}) + 3 \times A_r(\text{O})$
$= 2 \times 22.99 + 12.01 + 3 \times 16.00$
$= 106.0$

c $M_r = 2 \times A_r(\text{Al}) + 3 \times A_r(\text{S}) + 12 \times A_r(\text{O})$
$= 2 \times 26.98 + 3 \times 32.07 + 12 \times 16.00$
$= 342.2$

Section 9.2 Introducing the mole

Worked example: Try yourself 9.2.1

CALCULATING THE NUMBER OF MOLES OF ATOMS GIVEN THE NUMBER OF MOLES OF MOLECULES

Calculate the amount, in mol, of hydrogen atoms in 0.75 mol of water (H_2O).	
Thinking	**Working**
List the data given in the question next to the appropriate symbol. Include units.	The number of mol of hydrogen atoms is the unknown so: $n(\text{H}) = ?$ $n(H_2O) = 0.75\,\text{mol}$
Calculate the amount, in mol, of hydrogen atoms from the amount of water molecules and the molecular formula.	$n(\text{H}) = n(H_2O) \times 2$ $= 0.75 \times 2$ $= 1.5\,\text{mol}$

Worked example: Try yourself 9.2.2

CALCULATING THE NUMBER OF MOLECULES

Calculate the number of molecules in 1.6 moles of carbon dioxide (CO_2).	
Thinking	**Working**
List the data given in the question next to the appropriate symbol. Include units.	The number of carbon dioxide molecules is the unknown, so: $N(CO_2) = ?$ $n(CO_2) = 1.6\,\text{mol}$ $N_A = 6.022 \times 10^{23}$
Rearrange the formula to make the unknown the subject.	$n = \frac{N}{N_A}$ so $N(CO_2) = n \times N_A$
Substitute in data and solve for the answer.	$N(CO_2) = n \times N_A$ $= 1.6 \times 6.022 \times 10^{23}$ $= 9.6 \times 10^{23}$ molecules

ISBN 978 1 4886 1772 0

Worked example: Try yourself 9.2.3

CALCULATING THE NUMBER OF ATOMS

Calculate the number of hydrogen atoms in 0.35 mol of methane (CH_4).	
Thinking	**Working**
List the data given in the question next to the appropriate symbol. Include units.	The number of hydrogen atoms is the unknown so: $N(\text{H}) = ?$ $n(\text{CH}_4) = 0.35\,\text{mol}$ $N_A = 6.022 \times 10^{23}$
Calculate the amount, in mol, of hydrogen atoms from the amount of methane molecules and the molecular formula.	$n(\text{H}) = n(\text{CH}_4) \times 4$ $= 0.35 \times 4$ $= 1.4\,\text{mol}$
Rearrange the formula to make the unknown the subject.	$n = \frac{N}{N_A}$ so $N(\text{H}) = n \times N_A$
Substitute in data and solve for the answer.	$N(\text{H}) = n \times N_A$ $= 1.4 \times 6.022 \times 10^{23}$ $= 8.4 \times 10^{23}$ atoms

Worked example: Try yourself 9.2.4

CALCULATING THE NUMBER OF MOLES OF PARTICLES GIVEN THE NUMBER OF PARTICLES

Calculate how many moles of magnesium are in 8.1×10^{20} of magnesium atoms.	
Thinking	**Working**
List the data given in the question next to the appropriate symbol. Include units.	The number of mol of magnesium atoms is the unknown so: $n(\text{Mg}) = ?$ $N(\text{Mg}) = 8.1 \times 10^{20}$ atoms $N_A = 6.022 \times 10^{23}$
Rearrange the formula to make the unknown the subject.	$n = \frac{N}{N_A}$ n is the unknown so rearrangement is not required.
Substitute in data and solve for the answer.	$n(\text{Mg}) = \frac{N}{N_A}$ $= \frac{8.1 \times 10^{20}}{6.022 \times 10^{23}}$ $= 0.0013\,\text{mol}$ ($= 1.3 \times 10^{-3}\,\text{mol}$)

9.2 KEY QUESTIONS

1 Number of particles = amount (mol) × N_A

- **a** $N(\text{Na}) = 2.0 \times 6.022 \times 10^{23}$
  $= 1.2 \times 10^{24}$ atoms
- **b** $N(\text{N}_2) = 0.10 \times 6.022 \times 10^{23}$
  $= 6.0 \times 10^{22}$ molecules
- **c** $N(\text{C}) = 20.0 \times 6.022 \times 10^{23}$
  $= 1.20 \times 10^{25}$ atoms
- **d** $N(\text{H}_2\text{O}) = 4.2 \times 6.022 \times 10^{23}$
  $= 2.5 \times 10^{24}$ molecules
- **e** $N(\text{Fe}) = 1.0 \times 10^{-2} \times 6.022 \times 10^{23}$
  $= 6.02 \times 10^{21}$ atoms

ISBN 978 1 4886 1772 0

f $N(CO_2) = 4.62 \times 10^{-5} \times 6.022 \times 10^{23}$

$= 2.78 \times 10^{19}$ molecules

2 Remember: Avogadro's number, N_A, is 6.022×10^{23}.

$$\text{Amount (mol)} = \frac{\text{number of particles}}{N_A}$$

a $n(H_2O \text{ molecules}) = \dfrac{3.0 \times 10^{23}}{N_A}$

$= 0.50\,\text{mol}$

b $n(\text{Ne atoms}) = \dfrac{1.5 \times 10^{23}}{N_A}$

$= 0.25\,\text{mol}$

c $n(\text{Fe atoms}) = \dfrac{4.2 \times 10^{25}}{N_A}$

$= 70\,\text{mol}$

d $n(C_2H_5OH \text{ molecules}) = \dfrac{4.2 \times 10^{25}}{N_A}$

$= 70\,\text{mol}$

3 Remember: Avogadro's number, N_A, is 6.022×10^{23}.

$$\text{Amount (mol)} = \frac{\text{number of particles}}{N_A}$$

a $n(\text{Na atoms}) = \dfrac{1.0 \times 10^{20}}{N_A}$

$= 1.7 \times 10^{-4}\,\text{mol}$

b $n(\text{Al atoms}) = \dfrac{1.0 \times 10^{20}}{N_A}$

$= 1.7 \times 10^{-4}\,\text{mol}$

c $n(Cl_2 \text{ molecules}) = \dfrac{1.0 \times 10^{20}}{N_A}$

$= 1.7 \times 10^{-4}\,\text{mol}$

4 **a** Each Cl_2 molecule has two Cl atoms.

$\therefore n(\text{Cl atoms}) = 0.4 \times 2$

$= 0.8\,\text{mol}$

b Each methane molecule has four H atoms.

$\therefore n(\text{H atoms}) = 4 \times 1.2$

$= 4.8\,\text{mol}$

c Each ethane molecule has six H atoms.

$\therefore n(\text{H atoms}) = 6 \times 0.12$

$= 0.72\,\text{mol}$

d Each sulfate ion has four O atoms.

$\therefore n(\text{O atoms}) = 4 \times 1.5$

$= 6.0\,\text{mol}$

Section 9.3 Molar mass

Worked example: Try yourself 9.3.1

CALCULATING THE MASS OF A SUBSTANCE

Calculate the mass of 4.68 mol of sodium carbonate (Na_2CO_3).	
Thinking	**Working**
List the data given to you in the question. Remember that whenever you are given a chemical formula, you can calculate the molar mass.	$m(Na_2CO_3) = ?$ g $n(Na_2CO_3) = 4.68\,\text{mol}$ $M(Na_2CO_3) = 2 \times 22.98 + 12.01 + (3 \times 16.0) = 106.0\,\text{g mol}^{-1}$

Calculate the mass of sodium carbonate using: $n = \frac{m}{M}$	$n = \frac{m}{M}$, so $m = n \times M$ $m(Na_2CO_3) = 4.68 \times 106.0$ $= 496\,g$

Worked example: Try yourself 9.3.2

CALCULATING THE NUMBER OF MOLECULES

Calculate the number of sucrose molecules in a teaspoon (4.2 g) of sucrose ($C_{12}H_{22}O_{11}$).	
Thinking	**Working**
List the data given to you in the question. Convert mass to grams if required. Remember that whenever you are given a formula you can calculate the molar mass.	$N(C_{12}H_{22}O_{11}) = ?$ $M(C_{12}H_{22}O_{11}) = (12 \times 12.01) + (22 \times 1.008) + (11 \times 16.00)$ $= 342.3\,g\,mol^{-1}$ $m(C_{12}H_{22}O_{11}) = 4.2\,g$
Calculate the amount, in mol, of $C_{12}H_{22}O_{11}$, using: $n = \frac{m}{M}$	$n(C_{12}H_{22}O_{11}) = \frac{m}{M}$ $= \frac{4.2}{342.3}$ $= 0.012\,mol$
Calculate the number of $C_{12}H_{22}O_{11}$ molecules using: $n = \frac{N}{N_A}$	$n = \frac{n}{N_A}$, so $N = n \times N_A$ $N(C_{12}H_{22}O_{11}) = 0.012 \times 6.022 \times 10^{23}$ $= 7.4 \times 10^{21}$ molecules

9.3 KEY QUESTIONS

1 **a** $M_r = 2 \times A_r(N)$

$= 2 \times 14.01$

$= 28.0$

b $M_r = A_r(N) + 3 \times A_r(H)$

$= 14.01 + 3 \times 1.008$

$= 17.0$

c $M_r = 2 \times A_r(H) + A_r(S) + 4 \times A_r(O)$

$= 2 \times 1.008 + 32.07 + 4 \times 16.00$

$= 98.1$

d $M_r = A_r(Fe) + 3 \times A_r(N) + 9 \times A_r(O)$

$= 55.85 + 3 \times 14.01 + 9 \times 16.00$

$= 241.9$

e $M_r = 2 \times A_r(C) + 4 \times A_r(H) + 2 \times A_r(O)$

$= 2 \times 12.01 + 4 \times 1.008 + 2 \times 16.00$

$= 60.1$

f $M_r = A_r(S)$

$= 32.1$

g $M_r = 6 \times A_r(C) + 8 \times A_r(H) + 6 \times A_r(O)$

$= 6 \times 12.01 + 8 \times 1.008 + 6 \times 16.00$

$= 176.1$

h $M_r = A_r(Cu) + A_r(S) + 4 \times A_r(O) + 5 \times M_r(H_2O)$

$= 63.55 + 32.07 + 4 \times 16.00 + 5 \times 18.016$

$= 249.7$

ISBN 978 1 4886 1772 0

2 It is useful to remember the formula $m = nM$, where m is the mass in grams, n the amount of substance in mol, and M the molar mass. Use the periodic table to work out the molar masses.

a $m(\text{Na atoms}) = 1.0\,\text{mol} \times 22.99\,\text{g mol}^{-1}$
$= 23.0\,\text{g}$

b $m(O_2) = 2.0\,\text{mol} \times 32.00\,\text{g mol}^{-1}$
$= 64.0\,\text{g}$

c $m(CH_4) = 0.10\,\text{mol} \times 16.04\,\text{g mol}^{-1}$
$= 1.60\,\text{g}$

d $m(Al_2O_3) = 0.25\,\text{mol} \times 101.96\,\text{g mol}^{-1}$
$= 25.5\,\text{g}$

3 It is useful to remember the formula $n = \frac{m}{M}$, where m is the mass in grams, n the amount of substance in mol, and M the molar mass. Use the periodic table to work out the molar masses.

a $n(\text{H atoms}) = \frac{5.0}{1.008} = 4.96\,\text{mol}$

b $n(H_2) = \frac{5.0}{2.016} = 2.48\,\text{mol}$

c $n(\text{Al atoms}) = \frac{2.7}{26.982} = 0.10\,\text{mol}$

d $n(CH_4) = \frac{0.4}{16.043} = 0.025\,\text{mol}$

e $n(O_2) = \frac{0.10}{31.998} = 0.0031\,\text{mol}$

f $n(\text{O atoms}) = \frac{0.10}{15.999} = 0.0063\,\text{mol}$

g $n(P_4) = \frac{1.2 \times 10^{-3}}{123.896} = 9.7 \times 10^{-6}\,\text{mol}$

h $n(\text{P atoms}) = \frac{1.2 \times 10^{-3}}{30.974} = 3.87 \times 10^{-5}\,\text{mol}$

4 Use the formulae: number of particles $= n \times N_A$, where $N_A = 6.02 \times 10^{23}$ and $n = \frac{m}{M}$, where m is the mass in grams, n the amount of substance in mol, and M the molar mass. Use the periodic table to work out the molar masses.

a $n(\text{Na}) = \frac{23}{22.99} = 1.0\,\text{mol}$
number of Na atoms $= 1.0 \times 6.02 \times 10^{23} = 6.02 \times 10^{23}$ atoms

b $n(\text{Ar}) = \frac{4.0}{39.948} = 0.10\,\text{mol}$
number of Ar atoms $= 0.10 \times 6.02 \times 10^{23} = 6.03 \times 10^{22}$ atoms

c $n(\text{Mg}) = \frac{0.243}{24.305} = 0.010\,\text{mol}$
number of Mg atoms $= 0.01 \times 6.02 \times 10^{23} = 6.02 \times 10^{21}$ atoms

d $n(\text{Au}) = \frac{10.0}{196.967} = 0.051\,\text{mol}$
number of Mg atoms $= 0.051 \times 6.02 \times 10^{23} = 3.05 \times 10^{22}$ atoms

5 Use the formulae: number of particles $= n \times N_A$, where $N_A = 6.02 \times 10^{23}$ and $n = \frac{m}{M}$, where m is the mass in grams, n the amount of substance in mol, and M the molar mass. Use the periodic table to work out the molar masses.

a i $n(O_2) = \frac{16}{32} = 0.50\,\text{mol}$
number of O_2 molecules $= 0.5 \times 6.02 \times 10^{23} = 3.01 \times 10^{23}$ molecules

ii $n(N_2) = \frac{2.8}{28.014} = 0.10\,\text{mol}$
number of N_2 molecules $= 0.1 \times 6.02 \times 10^{23} = 6.02 \times 10^{22}$ molecules

b $n(SO_2) = \frac{3.2}{64.07} = 0.050\,\text{mol}$
number of SO_2 molecules $= 0.05 \times 6.02 \times 10^{23} = 3.01 \times 10^{22}$ molecules
Each molecule contains two oxygen atoms.
So, number of oxygen atoms $= 6.02 \times 10^{22}$ atoms.

ISBN 978 1 4886 1772 0

c $n(\text{NH}_3) = \frac{288}{17.03} = 16.91\ \text{mol}$

number of NH_3 molecules $= 16.91 \times 6.02 \times 10^{23} = 1.018 \times 10^{25}$ molecules

Each molecule contains four atoms (1 of N and 3 of H).

So, total number of atoms $= 4.0 \times 1.018 \times 10^{25} = 4.07 \times 10^{25}$ atoms.

Section 9.4 Percentage composition

Worked example: Try yourself 9.4.1

CALCULATING PERCENTAGE COMPOSITION

Calculate the percentage by mass of nitrogen in ammonium nitrate (NH_4NO_3).	
Thinking	**Working**
Find the molar mass of the compound.	$M(\text{NH}_4\text{NO}_3) = (2 \times 14.01) + (4 \times 1.008) + (3 \times 16.00)$ $= 80.05\ \text{g mol}^{-1}$
Find the total mass of the element in 1 mole of the compound.	mass of N in 1 mol $= 2 \times M(\text{N})$ $= 2 \times 14.01$ $= 28.02\ \text{g}$
Find the percentage by mass of the element in the compound.	% by mass of N in $\text{NH}_4\text{NO}_3 = \frac{\text{mass of N in 1 mol of NH}_4\text{NO}_3}{\text{molar mass of NH}_4\text{NO}_3} \times 100$ $= \frac{28.02}{80.05} \times 100$ $= 35.0\%$

9.4 KEY QUESTIONS

1 **a** Percentage by mass of an element $= \frac{\text{mass of 1 element in 1 mol of compound}}{\text{mass of 1 mol of the compound}} \times 100$

Use the periodic table to work out the molar masses. For example, $M(\text{Fe}_2\text{O}_3) = 159.7\ \text{g mol}^{-1}$.

$$\therefore\ \%(\text{Fe}) = \frac{2 \times 55.85}{159.7} \times 100$$
$$= 69.9\%$$

b $\%(\text{U}) = \frac{3 \times 238.0}{842.0} \times 100$

$= 84.8\%$

c $\%(\text{N}) = \frac{14.01}{53.49} \times 100$

$= 26.2\%$

d $\%(\text{O}) = \frac{6 \times 16.0}{187.57} \times 100$

$= 51.2\%$

2 **a** urea: percentage nitrogen = 46.6%

b ammonium nitrate: percentage nitrogen = 35.0%

c ammonium sulfate: percentage nitrogen = 21.2%

Therefore urea has the highest nitrogen content.

3 **a** silicon carbide (SiC): Si = 70.0%, C = 30.0%

b water (H_2O): H = 11.2%, O = 88.8%

c potassium phosphate (K_3PO_4): K = 55.3%, P = 14.6%, O = 30.1%

d sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$): C = 42.1%, H = 6.50%, O = 51.4%

e gold sulfate ($\text{Au}_2(\text{SO}_4)_3$): Au = 57.8%, S = 14.1%, O = 28.1%

CHAPTER 9 REVIEW

1 **a** The relative molecular mass, M_r, is the sum of the relative atomic masses, A_r, of the elements in the compound.

$\therefore M_r(H_2O) = 2.016 + 16$
$= 18.02$

b $M_r(P_4) = 4 \times 30.97$
$= 123.9$

c $M_r(CO) = 12.01 + 16.00$
$= 28.01$

2 **a** relative formula mass zinc bromide $= 65.38 + 2 \times 79.90 = 225.2$

b relative formula mass barium hydroxide $= 137.3 + 2 \times 16.00 + 2 \times 1.008 = 171.3$

c relative formula mass iron(III) carbonate $= 2 \times 55.85 + 3 \times 12.01 + 9 \times 16.00 = 291.7$

3 It is useful to remember the formula $n = \frac{m}{M}$, where m is the mass in grams, n the amount of substance in mol, and M the molar mass. Remember also that the number of particles in 1 mol, Avogadro's number, $N_A = 6.022 \times 10^{23}$.

Use the formula: $n = \frac{\text{number of particles}}{N_A}$. Use the periodic table to work out the molar masses.

a $n(H_2O) = \frac{4.50 \times 10^{23}}{6.022 \times 10^{23}} = 0.747\,\text{mol}$

b $n(CH_4) = 14.9\,\text{mol}$

c $n(Cl_2) = 3.8 \times 10^4\,\text{mol}$

d $n(C_{12}H_{22}O_{11}) = \frac{1}{6.022 \times 10^{23}} = 1.7 \times 10^{-24}\,\text{mol}$

4 **a** **i** $N(NH_3) = n \times N_A = 1.45 \times 6.022 \times 10^{23} = 8.73 \times 10^{23}$ molecules

ii $N(\text{atoms}) = N(NH_3) \times 4 = 8.73 \times 10^{23} \times 4 = 3.49 \times 10^{24}$ atoms

b **i** $N(H_2S) = n \times N_A = 0.576 \times 6.022 \times 10^{23} = 3.47 \times 10^{23}$ molecules

ii $N(\text{atoms}) = N(H_2S) \times 3 = 3.47 \times 10^{23} \times 3 = 1.04 \times 10^{24}$ atoms

c **i** $N(HNO_3) = n \times N_A = 0.0153 \times 6.022 \times 10^{23} = 9.21 \times 10^{21}$ molecules

ii $N(\text{atoms}) = N(HNO_3) \times 5 = 9.21 \times 10^{21} \times 5 = 4.61 \times 10^{22}$ atoms

d **i** $N(C_{12}H_{22}O_{11}) = n \times N_A = 2.5 \times 6.022 \times 10^{23} = 1.5 \times 10^{24}$ molecules

ii $N(\text{atoms}) = N(C_{12}H_{22}O_{11}) \times 45 = 1.5 \times 10^{24} \times 45 = 6.8 \times 10^{25}$ atoms

5 The molar mass, M, has the same numerical value as the relative molecular mass, M_r, which is the sum of the relative atomic masses, A_r, of the elements in the compound. The molar mass, M, is the actual mass of 1 mole and so has the unit $\text{g}\,\text{mol}^{-1}$.

6 **a** Fe : $55.85\,\text{g}\,\text{mol}^{-1}$

b H_2SO_4 : $98.09\,\text{g}\,\text{mol}^{-1}$

c Na_2O : $61.98\,\text{g}\,\text{mol}^{-1}$

d $Zn(NO_3)_2$:$189.4\,\text{g}\,\text{mol}^{-1}$

e H_2NCH_2COOH : $75.07\,\text{g}\,\text{mol}^{-1}$

f $Al_2(SO_4)_3$: $342.2\,\text{g}\,\text{mol}^{-1}$

g $FeCl_3{\cdot}6H_2O$: $270.3\,\text{g}\,\text{mol}^{-1}$

7 It is useful to remember the formula $m = nM$, where m is the mass in grams, n the amount of substance in mol, and M the molar mass.

a $m(C_2H_6) = 0.060 \times (24.02 + 6.048) = 1.8\,\text{g}$

b $m(C_6H_{12}O_6) = 0.32 \times ((6 \times 12.01) + (12 \times 1.008) + (6 \times 16.00)) = 58\,\text{g}$

c $m((NH_2)_2CO) = 6.8 \times 10^{-3} \times ((2 \times 14.01) + (4 \times 1.008) + 12.01 + (16.00)) = 0.41\,\text{g}$

d $m(Cu) = 6.12 \times 63.55 = 389\,\text{g}$

8 **a** It is useful to remember the formula $n = \frac{m}{M}$, where m is the mass in grams, n the amount of substance in mol, and M the molar mass in g mol^{-1}. Use the periodic table to work out the molar masses.

a $n(C) = \frac{1.201}{12.01} = 0.10\,\text{mol}$

b $n(S_8) = \frac{10.0}{256.56} = 0.0390\,\text{mol}$

c $n(CH_4) = \frac{20.0}{16.04} = 1.25\,\text{mol}$

d $n(C_6H_4(OCOCH_3)COOH) = \frac{0.300}{180.15} = 0.00167\,\text{mol}$

e $n(Al_2O_3) = \frac{3500000}{101.96} = 3.4 \times 10^4\,\text{mol}$

9 **a** $\text{mass of one atom} = \frac{\text{mass of 1 mole}}{\text{number of particles in a mole}} = \frac{\text{molar mass}}{N_A}$

$\text{mass of one calcium atom} = \frac{40.1}{6.022 \times 10^{23}} = 6.66 \times 10^{-23}\,\text{g}$

b $\text{mass of one water molecule} = \frac{18.0}{6.022 \times 10^{23}} = 2.99 \times 10^{-23}\,\text{g}$

c $\text{mass of one } CO_2 \text{ molecule} = \frac{44.01}{6.022 \times 10^{23}} = 7.31 \times 10^{-23}\,\text{g}$

10 **a** **i** $n(P_4) = \frac{m}{M} = \frac{4.2}{4 \times 30.97} = 0.034\,\text{mol}$

ii $N(P_4) = n \times 6.022 \times 10^{23} = 2.05 \times 10^{22}$ molecules

iii total number of atoms $= 4 \times 2.05 \times 10^{22} = 8.2 \times 10^{22}$ atoms

b **i** $n(S_8) = \frac{m}{M} = \frac{75.0}{8 \times 32.07} = 0.292\,\text{mol}$

ii $N(S_8) = n \times 6.022 \times 10^{23} = 1.76 \times 10^{23}$ molecules

iii total number of atoms $= 0.292 \times 8 \times 6.022 \times 10^{23} = 1.41 \times 10^{24}$ atoms

c **i** $n(HCl) = \frac{m}{M} = \frac{0.32}{1.008 + 35.45} = 0.0088\,\text{mol}$

ii $N(HCl) = n \times 6.022 \times 10^{23} = 5.3 \times 10^{21}$ molecules

iii total number of atoms $= 0.0088 \times 2 \times 6.022 \times 10^{23} = 1.1 \times 10^{22}$ atoms

d **i** $n(C_6H_{12}O_6) = \frac{m}{M} = \frac{2.2 \times 10^{-2}}{(6 \times 12.01) + (12 \times 1.008) + (6 \times 16.0)} = 1.22 \times 10^{-4}\,\text{mol}$

ii $n(C_6H_{12}O_6) = n \times 6.022 \times 10^{23} = 7.3 \times 10^{19}$ molecules

iii total number of atoms $= 1.22 \times 10^{-4} \times 24 \times 6.022 \times 10^{23} = 1.8 \times 10^{21}$ atoms

11 It is useful to remember the formula $n = \frac{m}{M}$, where m is the mass in grams, n the amount of substance in mol, and M the molar mass in g mol^{-1}. Use the periodic table to work out the molar masses of iron and water.

$M(Fe) = 55.85\,\text{g mol}^{-1}$ and $M(H_2O) = 18.0\,\text{g mol}^{-1}$

$\therefore n(H_2O) = \frac{20}{18} = 1.11\,\text{mol}$

$\therefore n(Fe)$ needed $= 1.11$ mol

$\therefore m(Fe)$ needed $= 1.11 \times 55.8 = 62.0$ g

12 It is useful to remember the formula $n = \frac{m}{M}$, where m is the mass in grams, n the amount of substance in mol, and M the molar mass in g mol^{-1}. Use the periodic table to work out the molar masses.

a **i** $n(NaCl) = \frac{5.85}{58.44} = 0.100\,\text{mol}$

ii $n(Na^+) = n(NaCl) = 0.100\,\text{mol}$

$n(Cl^-) = n(NaCl) = 0.100\,\text{mol}$

b **i** $n(CaCl_2) = \frac{45.0}{110.98} = 0.405\,\text{mol}$

ii $n(Ca^{2+}) = n(CaCl_2) = 0.405\,\text{mol}$

$n(Cl^-) = 2 \times n(CaCl_2) = 2 \times 0.405 = 0.810\,\text{mol}$

c **i** $n(Fe_2(SO_4)_3) = \frac{1.68}{399.91} = 0.00420\,\text{mol}$

ii $n(Fe^{3+}) = 2 \times n(Fe_2(SO_4)_3) = 0.00840\,\text{mol}$

$n(SO_4^{2-}) = 3 \times n(Fe_2(SO_4)_3) = 0.0126\,\text{mol}$

ISBN 978 1 4886 1772 0

13 It is useful to remember the formula $M = \frac{m}{n}$, where m is the mass in grams, n the amount of substance in mol, and M the molar mass in g mol^{-1}. Remember also that the number of particles in 1 mol is Avogadro's number, $N_A = 6.022 \times 10^{23}$. Use the formula: $n = \frac{\text{number of particles}}{N_A}$

a $M(\text{substance}) = \frac{72.0}{0.5} = 144\,\text{g mol}^{-1}$

b $n(\text{substance}) = \frac{6.0 \times 10^{22}}{N_A} = 0.1\,\text{mol}$

$M(\text{substance}) = \frac{10.0}{0.1} = 100\,\text{g mol}^{-1}$

14 It is useful to remember the formula $M = \frac{m}{n}$, where m is the mass in grams, n the amount of substance in mol, and M the molar mass in g mol^{-1}.

a $M(\text{substance}) = \frac{80.0}{2.0} = 40\,\text{g mol}^{-1}$

All other parts follow the same process.

b 98 g mol^{-1}

c 44 g mol^{-1}

d 106 g mol^{-1}

15 It is useful to remember the formula $m = nM$, where m is the mass in grams, n the amount of substance in mol, and M the molar mass. Use the periodic table to find the molar masses of iron and silver, $M = 55.85\,\text{g mol}^{-1}$ and $107.9\,\text{g mol}^{-1}$, respectively.

Remember also that the number of particles in 1 mol is N_A.

$m(\text{Cu}) = 100\,\text{g}$

$m(\text{Fe}) = 4.0 \times 55.85 = 223\,\text{g}$

$n(\text{Ag}) = \frac{1.2 \times 10^{24}}{N_A} = 2.0\,\text{mol}$

$\therefore m(\text{Ag}) = 2.0 \times 108 = 216\,\text{g}$

$\therefore$ B is correct: 4.0 mol of iron atoms has the greatest mass.

16 It is useful to remember the formula $m = nM$, where m is the mass in grams, n the amount of substance in mol, and M the molar mass in g mol^{-1}. Remember also that the number of particles in 1 mol is, Avogadro's number, $N_A = 6.02 \times 10^{23}$.

a $M(\text{antibiotic}) = 12\,500\,\text{g mol}^{-1} = 1.25 \times 10^4\,\text{g mol}^{-1}$

b $n(\text{antibiotic}) = \frac{2.0 \times 10^{-3}}{1.25 \times 10^4} = 1.6 \times 10^{-7}\,\text{mol}$

c number of molecules = nN_A

$= 1.6 \times 10^{-7} \times 6.022 \times 10^{23} = 9.6 \times 10^{16}$ molecules

17 Percentage by mass of an element = $\frac{\text{mass of 1 element in 1 mol of compound}}{\text{mass of 1 mol of the compound}} \times 100$

Use the periodic table to work out the molar masses. A useful check of these answers is provided by seeing that they add up to 100%, or somewhere close to that value.

a $M(\text{Al}) = 26.98\,\text{g mol}^{-1}$, $M(\text{O}) = 16.0\,\text{g mol}^{-1}$, $M(\text{Al}_2\text{O}_3) = 101.96\,\text{g mol}^{-1}$

$\%(\text{Al}) = \frac{2 \times 27}{102} \times 100 = 52.9\%$

$\%(\text{O}) = \frac{3 \times 16}{102} \times 100 = 47.1\%$

b Cu 65.1%; O 32.8 %; H 2.1 %

c $M(\text{Mg}) = 24.31\,\text{g mol}^{-1}$, $M(\text{Cl}) = 35.45\,\text{g mol}^{-1}$, $M(\text{MgCl}_2{\cdot}6\text{H}_2\text{O}) = 203.3\,\text{g mol}^{-1}$

$\%(\text{Mg}) = \frac{24.31}{203.3} \times 100 = 12.0\%$

$\%(\text{Cl}) = \frac{2 \times 35.45}{203.3} \times 100 = 34.9\%$

$\%(\text{H}) = \frac{12 \times 1.008}{203.3} \times 100 = 5.9\%$

$\%(\text{O}) = \frac{6 \times 16.0}{203.3} \times 100 = 47.2\%$

ISBN 978 1 4886 1772 0

d Fe 27.9%; S 24.1%; O 48.0%

e H 1.0%; Cl 35.3%; O 63.7%

18 percentage by mass of an element = $\frac{\text{mass of 1 element in 1 mol of compound}}{\text{mass of 1 mol of the compound}} \times 100$

Use the periodic table to work out the molar masses.

a $M(C_{10}H_8) = 128.16\,\text{g mol}^{-1}$, $M(H) = 1.008\ \text{g mol}^{-1}$, $M(C) = 12.01\,\text{g mol}^{-1}$

$\%(C) = \frac{10 \times 12.01}{128.16} \times 100 = 93.7\%$

b %C = 40.0%

c %C = 19.9%

d %C = 60.0%

19 a 3.0×10^{24} water molecules

b 4.0 mol

c Mg atom = 4.04×10^{-23} g

d H_2O molecule = 2.99×10^{-23} g

20 a i 2 mol Au

ii 3 mol S

iii 12 mol O

b $(2 \times 197.0) + (3 \times 32.07) + (12 \times 16.00) = 682.2\,\text{g}$

c $n = \frac{m}{M} = \frac{27.5}{682.2} = 0.403\,\text{mol}$

d $m = nM = 0.66 \times 682.2 = 450\,\text{g}$

e $\%\text{Au} = \frac{(2 \times 197.0)}{682.2} \times 100 = 57.8\%$

21 a $Cu(NO_3)_2{\cdot}6H_2O = 63.55 + (2 \times 14.01) + (6 \times 16.00) + (6 \times 18.016) = 295.7\,\text{g}$

b $n = \frac{m}{M} = \frac{30.0}{295.7} = 0.101\,\text{mol}$

c $\%H_2O = \frac{(6 \times 18.016)}{295.7} \times 100 = 36.5\%$

22 $n = \frac{m}{M} \therefore M = \frac{m}{n} = \frac{4.2}{0.05} = 84.0\,\text{g}$

From periodic table: Kr

23 a atomic mass of X = $\frac{2.2 \times 16.0}{1.0} = 35.2\,\text{g}$

b Some gases, such as neon, are monatomic (composed of single atoms), and some gases, such as oxygen, are diatomic (formed from two atoms). This will cause results to be incorrect if the scientists did not know whether the gases being tested were monatomic or diatomic.

ISBN 978 1 4886 1772 0

Chapter 10 Energy changes in chemical reactions

Section 10.1 Exothermic and endothermic reactions

10.1 KEY QUESTIONS

1 D. Combustion reactions result in a production of heat; therefore, energy is lost to the surroundings and the reaction is exothermic.

2 **a** 180 kJ
  b 1.5×10^3 kJ
  c 0.0100 kJ or 1.00×10^{-2} kJ
  d 2.0×10^{-6} kJ

3 In chemistry, the system is usually the chemical reaction, whereas the surroundings refer to everything else. For example, the beaker or test-tube in which the reaction takes place are the surroundings.

4 In any reaction, the total amount of chemical energy of the reactants is made up of the bonds between atoms within the reactants. If the total amount of chemical energy within the reactants is less than the total amount of chemical energy within the products, energy must be supplied to the system, and the reaction is said to be endothermic.

Section 10.2 Thermochemical equations, energy profile diagrams and enthalpy

Worked example: Try yourself 10.2.1

INTERPRETING A THERMOCHEMICAL REACTION

<table>
<tr><td colspan="2">The ΔH of the photosynthesis reaction shown below is 2803 kJ mol^{-1}.
$6CO_2(g) + 6H_2O(g) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$
Rewrite the reaction, including the energy term.</td></tr>
<tr><td>Thinking</td><td>Working</td></tr>
<tr><td>The reaction is endothermic so show energy is absorbed in the reaction.</td><td>Energy term goes on left-hand side.</td></tr>
<tr><td>$\Delta H = +2803$ kJ mol^{-1} so add this value.</td><td>$6CO_2(g) + 6H_2O(g) + 2803\ kJ\ mol^{-1} \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$</td></tr>
</table>

Worked example: Try yourself 10.2.2

IDENTIFYING THE ENTHALPY CHANGE IN A THERMOCHEMICAL REACTION

<table>
<tr><td colspan="2">Rewrite the reaction, showing the enthalpy change separately.
$4C_3H_5N_3O_9(l) \rightarrow 12CO_2(g) + 10H_2O(g) + 6N_2(g) + O_2(g) + 1456\ kJ\ mol^{-1}$</td></tr>
<tr><td>Thinking</td><td>Working</td></tr>
<tr><td>The reaction is exothermic—energy is lost to the surroundings.</td><td>Enthalpy change is negative.</td></tr>
<tr><td>Energy released = 1456 kJ mol^{-1}</td><td>$4C_3H_5N_3O_9(l) \rightarrow 12CO_2(g) + 10H_2O(g) + 6N_2(g) + O_2(g)$
$\Delta H = -1456\ kJ\ mol^{-1}$</td></tr>
</table>

10.2 KEY QUESTIONS

1 A negative Δ*H* value indicates a reaction is *exothermic*. This is because the enthalpy of the reactants is *greater than* the enthalpy of the products. Energy is being *released to* the surroundings.

2 $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \quad \Delta H = -890\,kJ\,mol^{-1}$

3 It would be lower because the change of state of the H_2O from liquid to gas will require energy to be absorbed.

4 **a** endothermic
   b The total enthalpy of the product (HI) is greater than that of the reactants (hydrogen gas and iodine gas).
   c The activation energy is greater than the Δ*H* value.

5 $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g) \quad \Delta H = +2803\,kJ\,mol^{-1}$

CHAPTER 10 REVIEW

1 **a** 2.21 kJ
   b 152 J
   c 1.89 MJ
   d 12.5 kJ

2 **a** Exothermic, because heat and light energy are released to the surrounding environment by the combustion of wood.
   b Endothermic, because thermal energy is absorbed from the surrounding environment to melt the ice.
   c Endothermic, because electrical energy is consumed from a power supply as the battery is recharged.
   d Exothermic, because heat energy is released to the surrounding environment as organisms in the compost heap decompose the plant material. The temperature of the heap rises as a consequence.

3 During chemiluminescence, energy is given off mainly in the form of light. Although the temperature of the surroundings may not increase, energy is still being lost to the surroundings, and the enthalpy of the chemicals decreases. Therefore it is an exothermic process, with a negative enthalpy change.

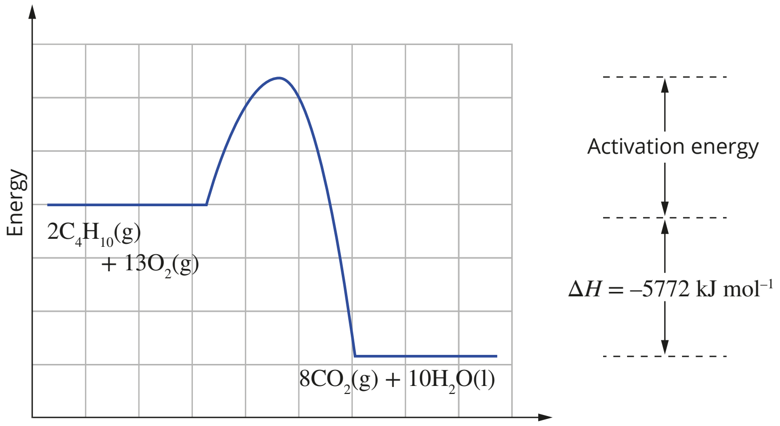
4 C. Some energy is always absorbed to break bonds in the reactants.

5 **a** true
   b false
   c false
   d true

6 If a chemical equation is written for an endothermic reaction, Δ*H* is positive, telling you that energy is absorbed as the reaction proceeds. The enthalpy of the products must be higher than the enthalpy of the reactants.

If this reaction is reversed, the enthalpy of the reactants is now higher than the enthalpy of the products. The reaction releases energy as the reaction proceeds, so is exothermic. Δ*H* becomes negative.

7 **a** The energy in the bonds of the reactants is higher than the energy in the bonds of the products.
   b



8 $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$ $\Delta H = -286\,kJ\,mol^{-1}$

or

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ $\Delta H = -572\,kJ\,mol^{-1}$

9 **a** 1 mol of CO(g) and 0.5 mol of $O_2(g)$

b **i** As the coefficients are twice those in the given equation, ΔH will also be doubled.

$\Delta H = 2 \times (-283)\,kJ\,mol^{-1}$

$= -566\,kJ\,mol^{-1}$

ii As the coefficients are twice those in the given equation and the equation is reversed, ΔH will also be doubled, and have the opposite sign. $\Delta H = +566\,kJ\,mol^{-1}$.

ISBN 978 1 4886 1772 0

Chapter 11 Fuels and introduction to stoichiometry

Section 11.1 Types of fuels

11.1 KEY QUESTIONS

1 A non-renewable fuel cannot be replenished at the rate at which it is consumed. Renewable fuels are those that can be replenished at a rate similiar to that at which they are consumed.

2 Renewable: bioethanol, biogas, biodiesel. Non-renewable: coal, oil, LPG, natural gas, coal seam gas.

3 Coal has the longest expected lifespan based on reserves and current rates of use. Carbon dioxide emission restrictions may reduce long-term dependence on coal.

4 **a** The rate of global energy use is more than can be supplied by wood. Wood has a relatively low energy density and is unsuitable for many portable/transport applications.

b Using a non-renewable energy source cannot be sustained indefinitely, but moderate and careful use now can increase the likelihood that it will meet the needs of future generations.

5 Crude oil consists of a range of hydrocarbons with different boiling points. The use to which a fraction is put is influenced by its boiling point and so fractional distillation is needed to separate them.

6 Some CO_2 is consumed in the production of the plant materials that biodiesel is made from.

Section 11.2 Combustion reactions

Worked example: Try yourself 11.2.1

WRITING EQUATIONS FOR COMPLETE COMBUSTION OF HYDROCARBON FUELS

Write the equation, including state symbols, for the complete combustion of hexane (C_6H_{14}).

Thinking	Working
Add oxygen as a reactant and carbon dioxide and water as the products.	$C_6H_{14} + O_2 \rightarrow CO_2 + H_2O$
Balance carbon and hydrogen atoms, based on the formula of the hydrocarbon.	$C_6H_{14} + O_2 \rightarrow 6CO_2 + 7H_2O$
Find the total number of oxygen atoms on the product side.	Total O $= (6 \times 2) + 7$ $= 19$
If this is an odd number, multiply all the coefficients in the equation by two, except for the coefficient of oxygen.	$2C_6H_{14} + O_2 \rightarrow 12CO_2 + 14H_2O$
Balance oxygen by adding the appropriate coefficient to the reactant side of the equation.	$2C_6H_{14} + 19O_2 \rightarrow 12CO_2 + 14H_2O$
Add state symbols.	$2C_6H_{14}(l) + 19O_2(g) \rightarrow 12CO_2(g) + 14H_2O(g)$

ISBN 978 1 4886 1772 0

Worked example: Try yourself 11.2.2

WRITING EQUATIONS FOR COMBUSTION REACTIONS OF ALCOHOLS

Write the equation, including state symbols, for the complete combustion of liquid methanol (CH_3OH).

Thinking	Working
Add oxygen as a reactant and carbon dioxide and water as the products.	$CH_3OH + O_2 \rightarrow CO_2 + H_2O$
Balance carbon and hydrogen atoms, based on the formula of the alcohol.	$CH_3OH + O_2 \rightarrow CO_2 + 2H_2O$
Find the total number of oxygen atoms on the product side. Then subtract the one oxygen atom in the alcohol molecule from the total number of oxygen atoms on the product side.	Total O on product side $= (1 \times 2) + (2 \times 1)$ $= 4$ Total O on product side – 1 in alcohol $= 4 - 1 = 3$
If this is an odd number, multiply all of the coefficients in the equation by two, except for the coefficient of oxygen.	$2CH_3OH + O_2 \rightarrow 2CO_2 + 4H_2O$
Balance oxygen by adding the appropriate coefficient to the reactant side of the equation.	$2CH_3OH + 3O_2 \rightarrow 2CO_2 + 4H_2O$
Add state symbols.	$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(g)$

Worked example: Try yourself 11.2.3

WRITING EQUATIONS FOR INCOMPLETE COMBUSTION OF FUELS

Write an equation, including state symbols, for the incomplete combustion of liquid methanol (CH_3OH) to form carbon monoxide and water vapour.

Thinking	Working
Add oxygen as a reactant and carbon monoxide and water as the products.	$CH_3OH + O_2 \rightarrow CO + H_2O$
Balance the carbon and hydrogen atoms, based on the formula of the hydrocarbon.	$CH_3OH + O_2 \rightarrow CO + 2H_2O$
Balance oxygen by adding the appropriate coefficient to the reactant side of the equation	$CH_3OH + O_2 \rightarrow CO + 2H_2O$
If oxygen gas has a coefficient that is half of a whole number, multiply all of the coefficients in the equation by two.	$CH_3OH + O_2 \rightarrow CO + 2H_2O$
Add state symbols.	$CH_3OH(l) + O_2(g) \rightarrow CO(g) + 2H_2O(g)$

11.2 KEY QUESTIONS

1 $2C_6H_6(l) + 15O_2(g) \rightarrow 12CO_2(g) + 6H_2O(g)$

2 $C_2H_5OH(l) + 2O_2(g) \rightarrow 2CO(g) + 3H_2O(g)$

3 $2H_2S(g) + 3O_2(g) \rightarrow 2SO_2(g) + H_2O(g)$

4 11

5 **a** Burning biofuels produces carbon dioxide, however the net effect is less than the burning of fossil fuels.
   b Burning petrol produces carbon dioxide.
   c Carbon dioxide will increase over time, as less photosynthesis can take place which will reduce the amount of CO_2 being absorbed from the atmosphere.

ISBN 978 1 4886 1772 0

- **d** Carbon dioxide deceases as the process of growth requires photosynthesis, which uses atmospheric carbon dioxide.
- **e** Using windpower rather than fossil fuels reduces the amount of carbon dioxide produced.
- **f** Increased industrialisation increases carbon dioxide as industrial processes generally create carbon dioxide.

6
- **a** Carbon dioxide is a greenhouse gas that makes rainwater naturally acidic.
- **b** Carbon monoxide is a toxic gas.
- **c** Methane is a greenhouse gas.
- **d** Water vapour is a greenhouse gas that does not affect water acidity.
- **e** Sulfur dioxide is responsible for high levels of acidity in rainwater.

Section 11.3 Calculations involving fuels

Worked example: Try yourself 11.3.1

CALCULATING THE CARBON CONTENT IN COAL

Calculate the mass of carbon present in 1300 tonnes of black coal that has a carbon content of 90.0%.	
Thinking	**Working**
Change the percentage to a fraction and multiply by the mass of the substance.	$m(\text{C}) = \left(\frac{90.0}{100}\right) \times 1300 = 1170\text{ tonnes}$

Worked example: Try yourself 11.3.2

CALCULATING THE CARBON CONTENT IN PURE COMPOUNDS

Calculate the mass of carbon in 25.0 kg of butanol, C_4H_9OH.	
Thinking	**Working**
Write the formula of the compound and calculate the relative molecular mass.	$M(C_4H_9OH) = (4 \times 12.01) + (10 \times 1.008) + (1 \times 16.00)$ $= 74.12$
Calculate the percentage of carbon in this total.	$\%(\text{C}) = \left(\frac{4 \times 12.01}{74.12}\right) \times 100$ $= 64.81\%$
Calculate the unknown mass of carbon using this percentage.	$m(\text{C}) = \left(\frac{64.81}{100}\right) \times 25.0\text{ kg}$ $= 16.2\text{ kg}$

Worked example: Try yourself 11.3.3

CALCULATING ENERGY RELEASED BY A SPECIFIED MASS OF A PURE FUEL

Calculate the amount of energy released when 5.40 kg of propane (C_3H_8) is burnt in an unlimited supply of oxygen.	
Thinking	**Working**
Calculate the number of moles of the compound using $n = \frac{m}{M}$.	$n(C_3H_8) = \frac{m}{M}$ $= \frac{5.4 \times 10^3}{44.09}$ $= 122.5$
Multiply the number of moles by the heat of combustion. (Give answer to 3 significant figures.)	$\text{energy} = n \times \Delta H_c$ $= 122.5 \times 2220$ $= 2.72 \times 10^5\text{ kJ}$

Worked example: Try yourself 11.3.4

USING MOLE RATIOS

How many moles of carbon dioxide are generated when 0.50 moles of butane are burnt completely in oxygen?	
Thinking	**Working**
Write a balanced equation for the reaction.	$2C_4H_{10} + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$
Note the number of moles of the known substance (in this case butane).	$n(C_4H_{10}) = 0.50\,\text{mol}$
Write a mole ratio for the chemicals being investigated. $\frac{\text{coefficient of unknown}}{\text{coefficient of known}}$	$\frac{n(CO_2)}{n(C_4H_{10})} = \frac{8}{2}$
Calculate the number of moles of the unknown substance using: $n(\text{unknown}) = \text{mole ratio} \times n(\text{known})$	$n(CO_2) = \frac{8}{2} \times 0.50$ $= 2.0\,\text{mol}$

Worked example: Try yourself 11.3.5

SOLVING MASS–MASS STOICHIOMETRIC PROBLEMS

Calculate the mass of carbon dioxide produced when 3600 g of butane (C_4H_{10}) is burnt completely in oxygen. (The molar mass of butane is 58.12 g mol^{-1}.)	
Thinking	**Working**
Write a balanced equation for the reaction.	$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$
Calculate the number of moles of the known substance (in this case butane) using: $n = \frac{m}{M}$	$n(C_4H_{10}) = \frac{3600}{58.12}$ $= 61.94\,\text{mol}$
Find the mole ratio: $\frac{\text{coefficient of unknown}}{\text{coefficient of known}}$	$\frac{n(CO_2)}{n(C_4H_{10})} = \frac{8}{2}$
Calculate the number of moles of the unknown substance (in this case carbon dioxide) using: $n(\text{unknown}) = \text{mole ratio} \times n(\text{known})$	$n(CO_2) = \frac{8}{2} \times 61.94$ $= 247.8\,\text{mol}$
Calculate the mass of the unknown substance using: $m = n \times M$	$m(CO_2) = 247.8 \times 44.01\,\text{g}$ $= 10.9\,\text{kg}$

11.3 KEY QUESTIONS

1 250 mol

$C_4H_9OH(g) + 6O_2(g) \rightarrow 4CO_2(g) + 5H_2O(g)$

$$\frac{n(H_2O)}{n(C_4H_9OH)} = \frac{5}{1}$$

Therefore, for 50 moles of C_4H_9OH:

$n(H_2O) = 50 \times 5 = 250\,\text{mol}$

2 propane: heat of combustion $= \frac{-2220}{44.0} = -50.5\,\text{kJ g}^{-1}$

octane: heat of combustion $= \frac{-5450}{114.0} = -47.8\,\text{kJ g}^{-1}$

Propane produces more energy per kilogram.

ISBN 978 1 4886 1772 0

3 $M(C_4H_{10}) = 58.1\,g\,mol^{-1}$

$n(C_4H_{10}) = 17.2\,mol$

$n(C) = 68.8\,mol$

$m(C) = 826\,g$

$M(C_4H_{10}) = 4 \times 12.01 + 10 \times 1.008$

$= 58.1\,g\,mol^{-1}$ (to 3 significant figures)

$n(C_4H_{10}) = \frac{1000}{58.1}$

$= 17.2\,mol$

$n(C) = 4 \times n(C_4H_{10})$

$= 4 \times 17.2$

$= 68.8\,mol$

$m(C) = 68.8 \times 12.01$

$= 826\,g$

4 **a** $n = \frac{m}{M} = \frac{250}{16.0} = 15.6\,mol$

energy $= n \times \Delta H_c = 15.6 \times 890 = 1.39 \times 10^4\,kJ$

b $n = \frac{m}{M} = \frac{9.64 \times 10^3}{44.0} = 219\,mol$

energy $= n \times \Delta H_c = 219 \times 2220 = 4.86 \times 10^5\,kJ$

c $n = \frac{m}{M} = \frac{403 \times 10^3}{46.0} = 8.76 \times 10^3\,mol$

energy $= n \times \Delta H_c = 8.76 \times 10^3 \times 1367 = 1.20 \times 10^7\,kJ$

d energy $= m \times \Delta H_c = 573 \times 10^6 \times 30 = 1.7 \times 10^{10}\,kJ$

5 The balanced equation shows that 25 mol of oxygen reacts with 2 mol of octane, producing 16 mol of carbon dioxide. The amount of each is found using $n = \frac{m}{M}$, where m is the mass in grams and M is the molar mass. A periodic table is used to calculate the molar masses of $CO_2 = 44.0\,g\,mol^{-1}$, $O_2 = 32.0\,g\,mol^{-1}$ and $C_8H_{18} = 114.0\,g\,mol^{-1}$.

a $n(C_8H_{18}) = \frac{200}{114.0} = 1.754\,mol$

$\therefore \frac{n(O_2)}{(C_8H_{18})} = \frac{25}{2}$

$\therefore n(O_2) = \frac{25}{2} \times 1.754 = 21.93\,mol$

$\therefore m(O_2) = 21.93 \times 32.0 = 702\,g$

b $n(C_8H_{18}) = \frac{200}{114.0} = 1.7544\,mol$

$\therefore \frac{n(CO_2)}{n(C_8H_{18})} = \frac{16}{2} = 8$

$n(CO_2) = 8 \times 1.7544 = 14.035\,mol$

$\therefore m(CO_2) = 14.035 \times 44.0 = 618\,g$

6 The equation for the reaction is:

$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$

a $n(C_4H_{10}) = \frac{m}{M} = \frac{1.00 \times 10^3}{58.0} = 17.241\,mol$

$\frac{n(O_2)}{n(C_4H_{10})} = \frac{13}{2}$

$n(O_2) = \frac{13}{2} \times 17.241 = 112.07\,mol$

$m(O_2) = n \times M = 112.07 \times 32.0 = 3586\,g = 3.59\,kg$

b $\frac{n(CO_2)}{n(C_4H_{10})} = \frac{8}{2}$

$n(CO_2) = \frac{8}{2} \times 17.341 = 68.96\,mol$

$m(CO_2) = n \times M = 68.96 \times 44.0 = 3034\,g = 3.03\,kg$

7 **a** advantage
b disadvantage
c advantage
d advantage
e disadvantage

8 **a** Energy will be absorbed by the evaporation of the water. The energy content per gram will be less than that for coal with a lower water content.
b removal of some water by drying and/or crushing
c Use of off-peak (e.g. night-time) electricity to dry the coal would reduce the cost.
d Potential atmospheric pollution from oxides of sulfur require either pre-treatment and/or post-treatment to remove or reduce sulfur from the brown coal, and post-treatment to remove or reduce it from exhaust gases.

CHAPTER 11 **REVIEW**

1 Apart from the difficulties in daily travel, lack of crude oil and natural gas will make transport of manufactured goods difficult and costly. It will also stop the production of all the products that are derived from crude oil—plastics, synthetic fibres, dyes, paints, solvents, detergents and pharmaceuticals.

2 **a** oil
b coal, oil, natural gas, biofuels, nuclear energy, new technologies, hydroelectricity

3 The formation of fossil fuels is a process that occurs over millions of years. The organic matter produced by plants and animals undergoes complex changes as it is subjected to heat and pressure under tonnes of mud and sand. Once the current reserves of fossil fuels have been used, they will not be replaced in the foreseeable future.

4 **a** black coal
b peat
c black coal

5 Natural gas. Australia has a large supply of natural gas, but globally, other fuel resources will last longer based on current patterns of usage.

6 Photosynthesis uses six molecules of carbon dioxide to produce one molecule of glucose:

$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$

Fermentation of glucose produces ethanol and two molecules of CO_2:

$C_6H_{12}O_6(aq) \rightarrow 2CO_2(g) + 2CH_3CH_2OH(aq)$

Combustion of ethanol releases heat and four molecules of CO_2:

$2CH_3CH_2OH(l) + 3O_2(g) \rightarrow 3H_2O(g) + 2CO_2(g)$

Therefore, all of the carbon dioxide released into the atmosphere by the fermentation and combustion of ethanol has been taken in by plants in the process of photosynthesis.

7 Biogas is formed through the action of anaerobic bacteria on organic matter. Anaerobic bacteria operate in the absence of oxygen. Methane is the most abundant component of biogas, followed by carbon dioxide and then a mix of other gases in low percentages.

8 $C_4H_9OH(g) + 6O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$

9 $2C_4H_{10}(g) + 9O_2(g) \rightarrow 8CO(g) + 10H_2O(l)$

10 All of the points are interrelated. Many other valid responses are possible.
a Polar ice caps are shrinking, causing sea levels to rise worldwide. This has local effects on the wildlife adapted to polar conditions. Melting ice adds cold, fresh water to the ocean with the possibility of altering the major ocean currents, resulting in climate changes, e.g. to Western Europe.
b Changed weather patterns causing droughts, floods, hurricanes and resulting in changed growth patterns in plants that might lead to both plant and animal extinctions.
c The changed weather patterns can result in crops failing, affecting the economy and driving some people to starvation.
d Some plants and animals that are dependent on particular weather conditions for propagation or for key parts of their life cycle may become extinct.

11

$$\begin{aligned} n &= \frac{m}{M} \\ &= \frac{5000}{16.0} \\ &= 312.5\,\text{mol} \end{aligned}$$

energy = $n \times \Delta H_c$
= 312.5×890
= 278 125 kJ
= $278\,125 \times 10^{-3}$ = 278 MJ

12 **a** The balanced equation shows that 1 mol of propane produces 3 mol of carbon dioxide. The amount of each is found using $n = \frac{m}{M}$, where *m* is the mass in grams and *M* is the molar mass in g mol^{-1}. Use a periodic table to calculate the molar masses. $M(C_3H_8) = 44.0\text{ g mol}^{-1}$, $M(CO_2) = 44.0\text{ g mol}^{-1}$, $M(H_2O) = 18.0\text{ g mol}^{-1}$ and $M(O_2) = 32.0\text{ g mol}^{-1}$.

$\frac{n(CO_2)}{(C_3H_8)} = \frac{3}{1}$

$n(C_3H_8) = \frac{6.70}{44.0} = 0.152\text{ mol}$

$n(CO_2)$ produced = 3×0.152 mol = 0.457 mol

∴ $m(CO_2)$ produced = 0.457×44.0 = 20.1 g

b $\frac{n(O_2)}{n(C_3H_8)} = \frac{5}{1}$

$n(C_3H_8) = \frac{6.70}{44.0} = 0.1522\text{ mol}$

$n(O_2)$ consumed = 5×0.1522 mol = 0.7610 mol
$m(O_2)$ consumed = 0.7610×32.0 = 24.4 g

c $\frac{n(H_2O)}{n(C_3H_8)} = \frac{4}{1}$

$n(C_3H_8) = \frac{6.70}{44.0} = 0.1522\text{ mol}$

$n(H_2O)$ produced = 4×0.1522 mol = 0.6088 mol
∴ $m(H_2O)$ produced = 0.6088×18.0 = 11.0 g

13 **a** $CH_3CH_2OH(g) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$

b $M(CH_3CH_2OH) = 46.068\text{ g mol}^{-1}$

$n(CH_3CH_2OH) = \frac{5.00}{46.068} = 0.1085\text{ mol}$

energy released = 1367×0.1085 = 148 kJ

c $\frac{n(CO_2)}{n(CH_3CH_2OH)} = \frac{2}{1}$

$n(CO_2)$ produced = $\frac{2}{1} \times 0.1085 = 0.2170$ mol

$m(CO_2)$ produced = 0.2170×44.01 = 9.55 g

14 **Methane**

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$
$M(CH_4) = 16.042\text{ g mol}^{-1}$
1.00 mol releases 890 kJ

moles to release 100 000 kJ = $\frac{100\,000}{890}$ = 112.4 mol

mass to release 100 000 kJ = 112.4×16.042 = 1803 g = 1.80 kg
$n(CO_2) = n(CH_4)$ = 112.4 mol
$M(CO_2) = 44.01\text{ kJ mol}^{-1}$
$m(CO_2) = 112.4 \times 44.01$ = 4947 g = 4.95 kg

Propane

$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$
$M(C_3H_8) = 44.094\text{ kJ mol}^{-1}$
1.00 mol releases 2220 kJ mol^{-1}

mol to release 100 000 kJ = $\frac{100\,000}{2220}$ = 45.04 mol

mass to release 100 000 kJ = 45.04×44.094 = 1986 g = 1.99 kg
$n(CO_2) = 3 \times n(C_3H_8) = 3 \times 45.04$ = 135.12 mol
$m(CO_2) = 135.12 \times 44.01$ = 5947 g = 5.95 kg

A greater mass of propane is required to produce the 100 000 kJ of energy and a higher mass of carbon dioxide is released in the process.

ISBN 978 1 4886 1772 0

15 **a** E10 petrol is a mixture of 10% ethanol and 90% conventional petrol.

b Use of E10 petrol will extend the availability of petrol as a fuel and will allow more of the larger fractions from crude oil to be used as a feedstock for other uses rather than being burnt as a fuel.

16 heat losses to the surroundings, friction, release of light and sound

17 **a** 1 mol of CO(g) and 0.5 mol of $O_2(g)$

b **i** As the coefficients are twice those in the given equation, ΔH will also be doubled.

$$\Delta H = 2 \times (-283)\,\text{kJ mol}^{-1}$$
$$= -566\,\text{kJ mol}^{-1}$$

ii As the coefficients are twice those in the given equation and the equation is reversed, ΔH will also be doubled, and have the opposite sign. $\Delta H = +566\,\text{kJ mol}^{-1}$.

18 42.0 mg

$$2C_8H_{18}(g) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$$

$$M(C_8H_{18}) = 8 \times 12.01 + 18 \times 1.008 = 114.224\,\text{g mol}^{-1}$$

$$n(C_8H_{18}) = \frac{0.0120}{114.224} = 1.0497 \times 10^{-4}\,\text{mol}$$

$$\frac{n(O_2)}{n(C_{14}H_{10})} = \frac{25}{2}$$

$$n(O_2) = \frac{25}{2} \times n(C_4H_{10}) = \frac{25}{2} \times 1.0497 \times 10^{-4} = 1.313 \times 10^{-3}\,\text{mol}$$

$$m(O_2) = 1.313 \times 10^{-3} \times 32.00 = 0.0420\,\text{g} = 42.0\,\text{mg}$$

19 **a** $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$

For each minute of operation:

$$n(C_3H_8) = \frac{12.7}{44.094} = 0.2880\,\text{mol}$$

$$\frac{n(O_2)}{n(C_3H_8)} = \frac{5}{1}$$

$$n(O_2)\text{ required} = 0.2880 \times \frac{5}{1} = 1.440\,\text{mol}$$

$$m(O_2)\text{ required} = 1.440 \times 32.00 = 46.1\,\text{g}$$

b Incomplete combustion of the propane may occur. This will likely result in the formation of carbon monoxide instead of carbon dioxide. Note that carbon monoxide is a highly toxic gas that can cause death if inhaled in large enough quantities.

20 **a** Biochemical fuels are derived from renewable resources such as plants. Ethene, the feedstock used for the industrial production of ethanol, is derived from the distillation and cracking of crude oil, a non-renewable resource.

b Ethanol can be produced from the fermentation of sugar by yeasts. Sugar may be derived from sugar cane or by the hydrolysis of starch from grains such as maize.

21 **a** methane, octane, biodiesel, ethanol

b B

22 Individual student response required.

23 **a** Burning coal to generate electricity produces greenhouse gas emissions, which would not help Australia meet its Paris Agreement obligations. However, the cost of electricity could rise if other energy sources were required to replace coal, which is relatively abundant and cheap.

b Biofuels could potentially be carbon neutral—the carbon dioxide produced from the combustion of the fuels is offset by the carbon dioxide absorbed to grow the crops that are used to make the fuels. This would help Australia meet its Paris Agreement obligations.

Unit 1 Review

Section 1: Multiple choice

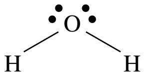
1 B
2 C
3 D
4 D
5 C
6 A
7 C
8 A
9 D
10 C
11 B
12 B

Section 2: Short answer

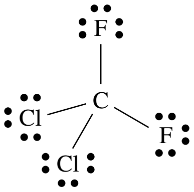
1 **a** **i** Au_2S_3
ii $Sr(OH)_2$
iii $(NH_4)_3PO_4$
b **i** iron(III) sulfate **ii** diphosphorus pentoxide

2 **a** 180.2 g
b 3.0×10^{-22} g
c Fructose molecules are attracted to each other due to intermolecular forces between the molecules.
d hydrogen bonds and dispersion forces

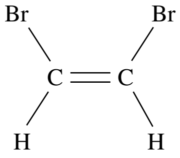
3 **a** **i** water (H_2O)



ii dichlorodifluoromethane (CCl_2F_2)



b **i** *cis*–dibromoethene ($C_2H_2Br_2$)



ii carbon disulfide (CS_2)
$S{=}C{=}S$

ISBN 978 1 4886 1772 0

4 **a** **i** methylpropene

ii *trans*-difluoroethene

iii 3-ethyl-2,5,5-trimethylheptane

b **i** and **ii** only as they are unsaturated

c $C_{12}H_{26}$

5 $M(Al(NO_3)_3 \cdot 6H_2O) = 321.1\,g$

a $\%H_2O = \frac{6 \times 18.016 \times 100}{321.1} = 33.7\%$

b $n = \frac{m}{M} = \frac{1000.0}{M} = 3.11\,mol$

6

Reaction	$\Delta H_{reaction}$	**Exothermic or endothermic**
$2H_2(g) + O_2(g) \rightarrow 2H_2O + 287\,kJ\,mol\,L^{-1}$	$-287\,kJ\,mol\,L^{-1}$	exothermic
$H_2S + 90\,kJ\,mol\,L^{-1} \rightarrow H_2 + S(s)$	$+90\,kJ\,mol\,L^{-1}$	endothermic

Section 3: Extended answer

1 **a** metallic bond; a lattice of positive ions surrounded by a ‘sea’ of mobile delocalised electrons; conductor

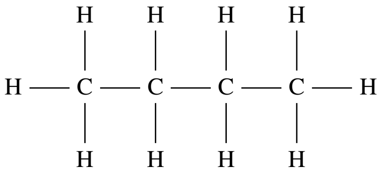
b ionic bonds between positive sodium ions and negative chloride ions; no free ions or electrons; non-conductor

c covalent bonds between the hydrogen atoms and the oxygen atoms within the H_2O molecules; weak intermolecular bonds (hydrogen bonds) between the molecules; no free charged particles; non-conductor

d two-dimensional covalent network with free mobile electrons; conductor

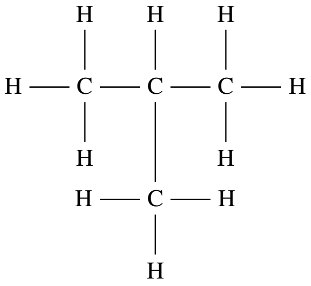
e three-dimensional covalent network in which all electrons are covalently bonded; no free electrons; non-conductor

2 **a**



$$\begin{array}{ccccccccccc} & & H & & H & & H & & H & & \\ & & | & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & C & - & H \\ & & | & & | & & | & & | & & \\ & & H & & H & & H & & H & & \end{array}$$

b



methylpropane

c $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$

d exothermic

e $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g) + 5750\,kJ$

f $M(C_4H_{10}) = 58.21\,g$

$n(C_4H_{10}) = \frac{m}{M} = \frac{1000.0}{58.12} = 17.21\,mol$

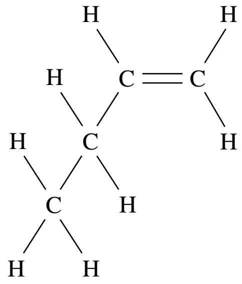
$\therefore n(CO_2) = 4 \times 17.21 = 68.84\,mol$

$\therefore m(CO_2) = nM = 68.84 \times 44.01 = 3.03 \times 10^3 g$ (3 sig. figs)

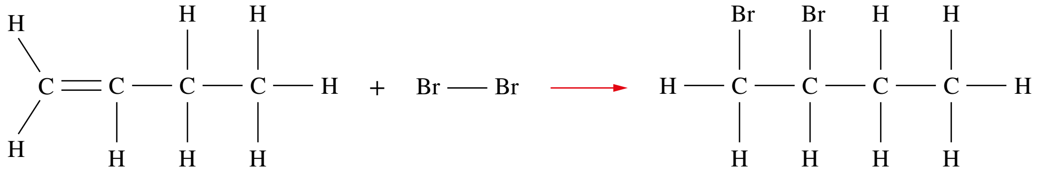
g $CH_3–CH_2–CH_2–CH_3 + F–F \rightarrow CH_3–CH_2–CH_2–CH_2F + H–F$

h fluorobutane

i



j



k 1,2-dibromobutane

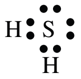
ISBN 978 1 4886 1772 0

Chapter 12 Intermolecular forces

Section 12.1 Shapes of molecules

Worked example: Try yourself 12.1.1

PREDICTING THE SHAPE OF MOLECULES

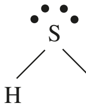
Predict the shape of a molecule of hydrogen sulfide (H_2S).	
Thinking	**Working**
Draw the electron dot diagram for the molecule.	
Count the number of bonds and lone pairs on the central atom.	There are two bonds and two lone pairs.
Determine how the groups of electrons will be arranged to get maximum separation.	Because there are four electron pairs, the groups will be arranged in a tetrahedral arrangement.
Deduce the shape of the molecule by considering the arrangement of just the atoms.	The sulfur and hydrogen atoms are a bent or V-shape.

12.1 KEY QUESTIONS

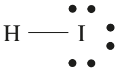
1 The VSEPR theory is based on the principle that negatively charged electron pairs in the outer shell of an atom repel each other. As a consequence, these electron pairs are arranged as far away from each other as possible.

2 four pairs, one bonding pair and three lone pairs

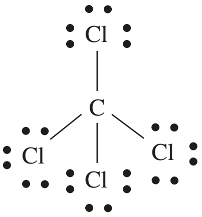
3 **a**



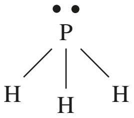
b



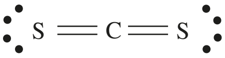
c



d



e



4 **a** V-shaped
  b linear
  c tetrahedral
  d pyramidal
  e linear

5 **a** tetrahedral
  b pyramidal
  c V-shaped

ISBN 978 1 4886 1772 0

Section 12.2 Properties of covalent molecular substances

Worked example: Try yourself 12.2.1

COMPARING THE POLARITY OF COVALENT BONDS

Compare the polarity of the covalent bonds in nitrogen monoxide (NO) and hydrogen chloride (HCl).	
Thinking	**Working**
Use the table of electronegativity values in Figure 12.2.5 on page 270 to find the electronegativities of the atoms in each molecule.	NO: nitrogen 3.04, oxygen 3.44 HCl: hydrogen 2.20; chlorine 3.16
For each molecule, subtract the lowest electronegativity value from the highest value.	NO: 3.44 – 3.04 = 0.40 HCl: 3.16 – 2.20 = 0.96
Determine which molecule has the biggest difference in electronegativity in order to determine the more polar molecule.	HCl is more polar than NO.

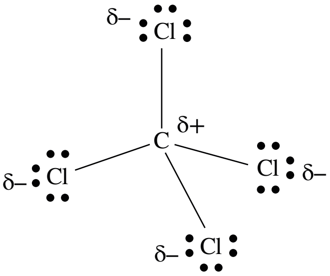
12.2 KEY QUESTIONS

1 **a** O
  b C
  c N
  d N
  e F
  f F

2 **a** P–F
  b C–H

3 hydrogen–nitrogen bond, carbon–nitrogen bond, sulfur–nitrogen bond, oxygen–nitrogen bond, nitrogen–nitrogen bond

4



5 **a** polar
  b polar
  c polar
  d polar
  e non-polar

6 C

ISBN 978 1 4886 1772 0

Section 12.3 Types of intermolecular forces

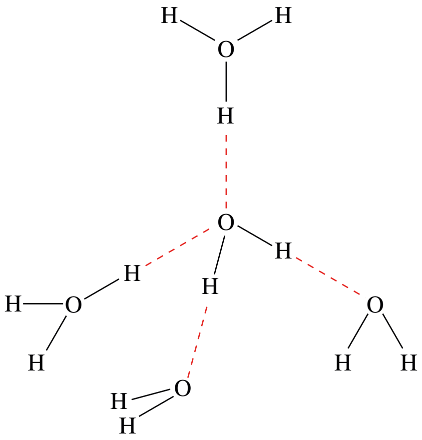
12.3 KEY QUESTIONS

1 Hydrogen chloride and chloromethane would form dipole–dipole forces.

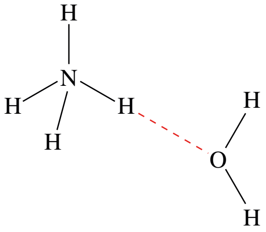
2 C. hydrogen chloride

3 dipole–dipole forces: a to h; hydrogen bonds: a, g, h

4



5



6 **a** dispersion forces only
- **b** dispersion forces and hydrogen bonding
- **c** dispersion and dipole–dipole forces
- **d** dispersion forces and hydrogen bonding
- **e** dispersion forces only

CHAPTER 12 REVIEW

1 PCl_3 – pyramidal, HOCl – V-shaped, $CHCl_3$ – tetrahedral, HF – linear

2 V-shaped

3 Beryllium does not obey the octet rule in this molecule and forms only two single bonds. Each bond has two electrons, so there are four bonding electrons involved in bonding.

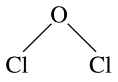
4 **a** tetrahedral
- **b** pyramidal
- **c** tetrahedral
- **d** V-shaped
- **e** pyramidal

ISBN 978 1 4886 1772 0

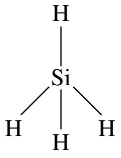
5 **a** non-polar (polar bonds but symmetrical)

$S=C=S$

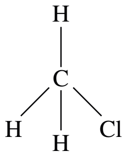
b polar (polar bonds and asymmetrical)



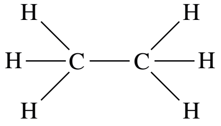
c non-polar (polar bonds but symmetrical)



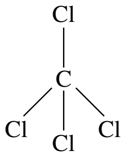
d polar (one polar bond)



e non-polar (non-polar bonds)



f non-polar (polar bonds but symmetrical)



6 B. The O–H bond in water is the most polar bond.

7 **a** non-polar
b polar
c non-polar
d polar
e polar

8 Si–O, H–Br, N–O, O–Cl, F–F

9

a SO_3	**i** non-polar	**ii** dispersion forces
b $SiCl_4$	**i** non-polar	**ii** dispersion forces
c CF_4	**i** non-polar	**ii** dispersion forces
d NF_3	**i** polar	**ii** dipole–dipole attraction
e CH_3NH_2	**i** polar	**ii** hydrogen bonding

10 A and C. Compounds CCl_4 and CO_2 are both non-polar, and so intermolecular forces operating between these molecules will be dispersion forces.

11 Melting temperatures increase down the table because the molecules increase in mass and size and there are more electrons in the molecules; therefore, the strength of the dispersion forces increases.

12 CF_4 has a slightly higher boiling temperature (–128°C) than OF_2 (–145°C), indicating that the forces between molecules in CF_4 are stronger. OF_2 is slightly polar; CF_4 is non-polar. OF_2 molecules are held together by forces of dipole–dipole attraction and dispersion forces. Although CF_4 molecules are attracted by dispersion forces only, the much larger size of CF_4 molecules makes the dispersion forces stronger than the sum of the dipole–dipole forces and the dispersion forces between OF_2 molecules.

13 Neon exists as single atoms, with the only forces of attraction being dispersion forces. As neon atoms have very few electrons, the dispersion forces are extremely weak. Neon therefore has a very low boiling temperature. Hydrogen fluoride molecules, however, are very polar and so are held together by electrostatic attraction between permanent dipoles. Because hydrogen is bonded to the very electronegative fluorine, the forces between molecules are hydrogen bonds. These are relatively strong intermolecular bonds and hydrogen fluoride, therefore, has a much higher boiling temperature than neon. (The dispersion forces operating between hydrogen fluoride molecules are extremely weak.)

ISBN 978 1 4886 1772 0

14 **a** CCl_4

b CH_4 and CCl_4 are both non-polar and so are held together in a lattice only by dispersion forces. CCl_4 is the larger of these two molecules and has more electrons, so the dispersion forces between CCl_4 molecules will be greater than those between CH_4 molecules. As there are stronger dispersion forces between molecules of CCl_4 than between molecules of CH_4, it takes more energy to vaporise CCl_4.

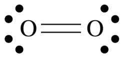
15 Iodine is a much larger molecule, so it has more electrons and therefore the dispersion forces are much stronger in iodine than in fluorine.

16 A permanent dipole is formed if there is a difference in electronegativity between the two atoms that form a bond. The more electronegative atom has a partial negative charge and the less electronegative atom has a partial positive charge. In larger molecules, individual bonds may be polar, but if the molecule is symmetrical the molecule will not have a permanent dipole overall and is non-polar. Asymmetry in the molecule causes an asymmetry in the electron distribution around the molecule, causing one end of the molecule to develop a partial negative charge while the other end develops a partial positive charge. The positive and negative ends of neighbouring molecules attract each other, forming dipole–dipole bonds.

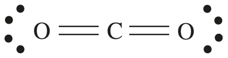
A temporary dipole is caused by random fluctuations in the electron distributions around the molecule. The electrons are constantly moving and can occasionally concentrate at one end of the molecule, causing that end to develop a temporary negative charge while the other end develops a temporary positive charge. This temporary dipole can then induce dipoles in the neighbouring molecules. The induced dipoles attract each other. Such attractions are known as dispersion forces and are present between all molecules.

17 **a** **i** N_2

ii O_2



iii CO_2



b **i** NH_3, HCl, H_2O, $CHCl_3$

ii N_2, Cl_2, O_2, CH_4, CO_2, CCl_4

iii NH_3, H_2O

18 If water were a linear molecule, the polarity of the two O–H bonds would cancel each other out and make the molecule non-polar. As water *is* polar, it cannot be a linear molecule.

19 The intermolecular bonds that hold molecules together in covalent molecular substances are much weaker (100 times) than the chemical bonds holding the atoms together in ionic, metallic and covalent network substances. As a result, it takes much less heat energy to break the intermolecular bonds holding covalent molecular solids and liquids together and these substances have relatively low melting and boiling points.

20 A. Hydrogen fluoride has the highest boiling point because there is hydrogen bonding between its molecules, which is much stronger than the dipole–dipole and dispersion forces between the molecules of the other substances. The other substances will not have hydrogen bonding because they don’t contain hydrogen atoms bonded to oxygen, nitrogen or fluorine atoms.

21 C. Molecular substances do not conduct electricity because they do not contain mobile charges such as electrons or ions that can create a current. On the other hand, boiling points increase as the intermolecular forces increase. Vapour pressure decreases as the intermolecular forces in a liquid increase, preventing the molecules from entering the gas phase. The solubility of a substance depends on the polarity of the solvent.

22 D. The molecules of icosane are larger than the other molecules, so there are stronger dispersion forces between the molecules in this compound.

Chapter 13 Chromatography

Section 13.1 Principles of chromatography

Worked example: Try yourself 13.1.1

CALCULATING R_f VALUES

Calculate the R_f value of the blue component in Figure 13.1.7.	
Thinking	**Working**
Record the distance the component has moved from the origin and the distance the solvent front has moved from the origin.	distance from origin of blue component = 3 cm distance from origin of solvent front = 10 cm
$R_f = \frac{\text{distance of component from origin}}{\text{distance of solvent front from origin}}$	$R_f \text{ (blue component)} = \frac{3}{10} = 0.3$

13.1 KEY QUESTIONS

1 B. Ethanol is the mobile phase in this example. The pigments contain the components to be separated and the paper is the stationary phase. There will be ethanol vapour in the air in the jar, but this is not the mobile phase.

2

Term	Description
adsorption	the attraction of one substance to the surface of another
desorption	the breaking of the attraction between a substance and the surface to which the substance is adsorbed
components	the different compounds in the mixture, which can be separated by chromatography
polar molecule	a molecule that acts as a dipole; it has one or more polar covalent bonds, with the charge being distributed asymmetrically
mobile phase	the solvent that moves over the stationary phase in chromatography
stationary phase	the components of a mixture undergo adsorption to this phase

3 For parts **a**, **b**, **c**, it would be useful to set your answers out in a table.

Band	a Distance from origin (mm)	b R_f	c Compound
light green	20	0.33	chlorophyll b
dark green	27	0.45	chlorophyll a
orange	40	0.67	xanthophyll
yellow	50	0.83	ß-carotene
solvent front	60	–	–

d The chromatogram would be likely to be different because separation of components depends on their solubility in the mobile phase (as well as strength of adsorption to the stationary phase). The polarity of the solvent used in TLC and paper chromatography will affect the R_f of the sample components. A polar solvent will dissolve polar samples readily; a non-polar solvent will dissolve non-polar samples readily.

4
1. Dissolve a sample of pure phenacetin in a volume of chloroform. This is the standard solution.
2. Dissolve a tablet of the analgesic in chloroform. This is the sample solution.
3. Place a small spot of the sample solution near the bottom of a thin-layer plate.
   Place a spot of the standard solution next to it, at the same distance from the bottom of the plate.
4. When the spots are dry, place the plate in a container with a small volume of solvent, such as chloroform. The lower edge of the plate, but not the spots, should be immersed.

ISBN 978 1 4886 1772 0

  5 Allow the solvent to rise until it almost reaches the top of the plate and then remove the plate from the container.

  6 Let the plate dry and examine it under ultraviolet light. If a spot from the sample appears at the same distance from the origin as the spot from the standard solution, the tablet is likely to contain phenacetin.

5 During column chromatography, the components of the sample *adsorb* onto the stationary phase and *desorb* into the liquid mobile phase. A component that adsorbs most strongly to the *stationary* phase and is least soluble in the *mobile* phase would be expected to take the *longest* time to pass through the column.

Section 13.2 Advanced applications of chromatography

Worked example: Try yourself 13.2.1

CONCENTRATION OF A PESTICIDE

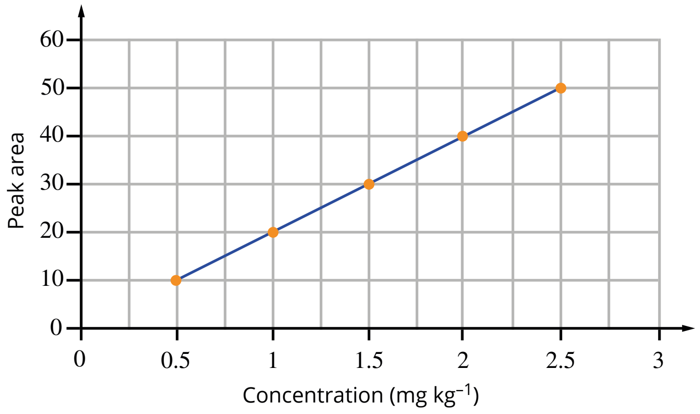
Procymidone is a pesticide that can be used to prevent disease in oranges. Tests have been performed using HPLC to determine the concentration of this pesticide in a sample of orange juice.

The chromatograms of a series of standards with accurately known concentrations of procymidone were obtained under the same conditions as the sample.

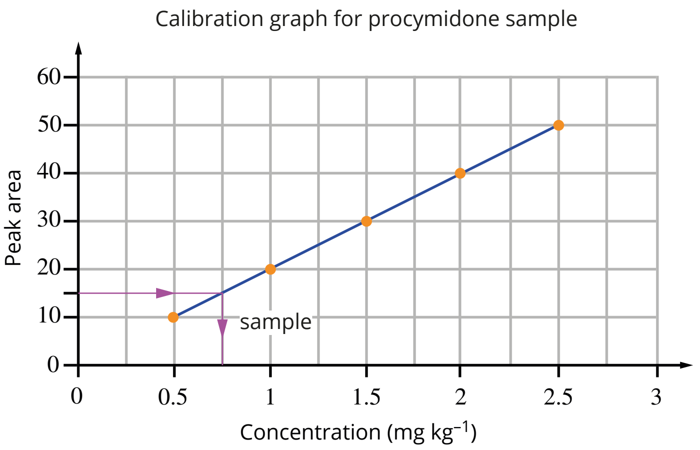
The results from the chromatograms of the sample and the standards are shown in the following table.

Determine the concentration of procymidone in the sample of orange juice.

	Peak area	Concentration (mg kg^{-1})
Standard 1	10	0.5
Standard 2	20	1.0
Standard 3	30	1.5
Standard 4	40	2.0
Standard 5	50	2.5
Sample	15	?

Thinking	**Working**
Construct a calibration graph by plotting peak area versus concentration.	Calibration graph for concentration of procymidone in orange juice sample 

ISBN 978 1 4886 1772 0

Mark where the peak area of the sample lies on the calibration curve by tracing a horizontal line to the curve and then a vertical line down to the *x*-axis.	
Determine the concentration of the sample by reading the concentration value for the unknown solution from the *x*-axis.	The concentration of procymidone in the orange juice sample is 0.75 mg kg^{-1}.

13.2 KEY QUESTIONS

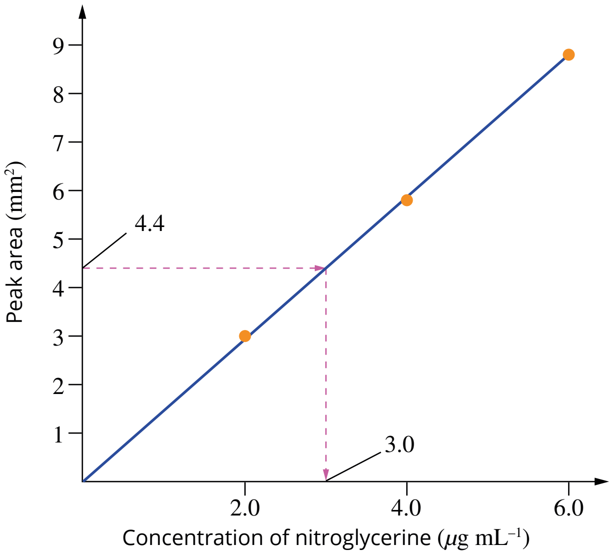
1 A

2 A

3 B, A, C

A and B are equally soluble in water, but B will move at a faster rate since it is less strongly adsorbed to the column. C will move slowest because it is least soluble in water.

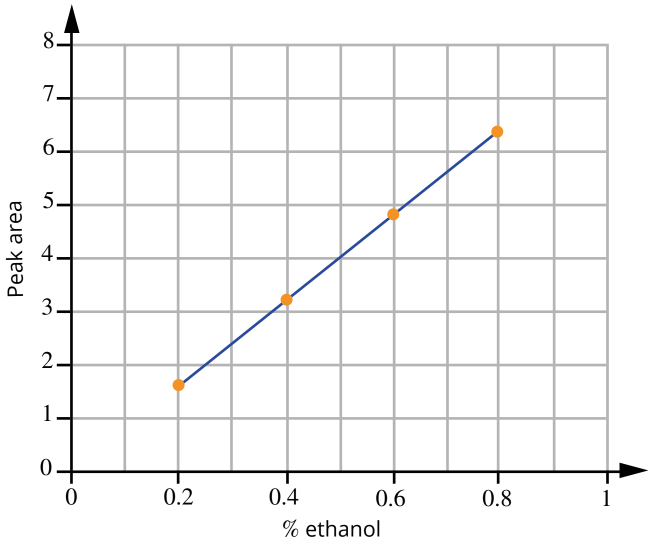
4 **a**



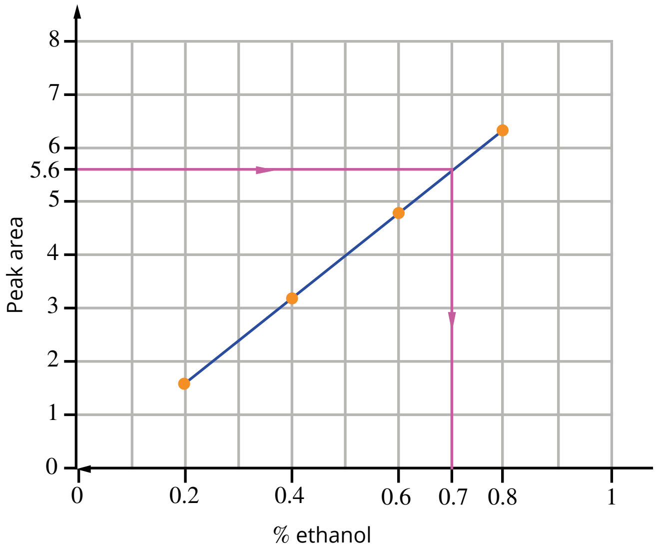
b 3.0 μg mL^{-1}

ISBN 978 1 4886 1772 0

5 **a** Calibration graph for ethanol sample



b Using the above calibration graph, the concentration of ethanol that corresponds with a peak area of 5.6 is 0.7%. To obtain this value, find the value of the peak area of the wine sample (5.6) on the vertical axis. Draw a horizontal line from this location until it intersects the calibration graph. From this point, draw a line vertically to the horizontal axis. Determine the percentage concentration of ethanol in the wine sample (0.7%).



CHAPTER 13 **REVIEW**

1 In paper chromatography, the paper acts as the *stationary* phase. A small spot of a solution is placed at one end of the paper, called the *origin*. A small spot of a solution is placed at one end of the paper, called the *sample*. The sample solution contains a number of different coloured compounds, the *components.*

The paper is suspended so that the end with the spot is *above* the surface of the solvent. The solvent or *mobile* phase moves up the *stationary* phase. Different coloured spots are observed at various places on the paper, due to the separation of different *components.*

2 **a** Water was *absorbed* by the towel as the wet swimmer dried himself.
A thin layer of grease *adsorbed* onto the cup when it was washed in the dirty water.

b Absorb: Atoms or molecules are taken *into* the material. Adsorb: Atoms or molecules accumulate and bond weakly to the *surface* of a solid or liquid.

3 The component at the top of the chromatogram has a greater rate of adsorption and desorption compared to a component at the bottom of the chromatogram.

ISBN 978 1 4886 1772 0

4 D. The rate at which a component moves down a chromatography column depends on the degree of adsorption of the component to the stationary phase and its solubility in the mobile phase. The length of the column is not a factor in the speed at which components move down the column.

5 The R_f value is the ratio of the distance a component has moved from the origin to the distance from the origin to the solvent front.

$R_f(\text{blue}) = \frac{8}{10} = 0.8$

$R_f(\text{purple}) = \frac{6}{10} = 0.6$

$R_f(\text{yellow}) = \frac{2}{10} = 0.2$

6 R_f = distance dye has moved/distance solvent front has moved.

$R_f(\text{blue}) = \frac{7.5}{9.0} = 0.83$

$R_f(\text{red}) = \frac{5.2}{9.0} = 0.58$

7 **a** If the solvent were above the level of the origin, the compounds under test would dissolve and disperse throughout the solvent.

b Components in a mixture undergoing chromatography cannot move faster than the solvent that is carrying them over the stationary phase. R_f values must therefore be less than one.

c three

d B: blue; C: green. They can be identified on the basis of their colour and R_f values.

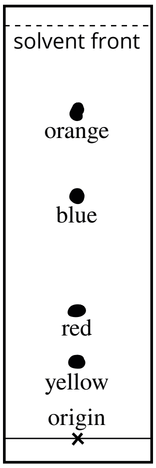
e blue

f 0.63; 0.13

8 **a** 3.2 cm

b 15 cm

c



9 **a** Taurine, glycine and an unknown. Two of the three spots produced by the medicine match those produced by taurine and glycine. The third spot does not match any of the amino acid standards and represents an unknown substance.

b There are several ways of visualising the spots.

- View the chromatogram under UV light.
- Spray the finished chromatogram with a compound that causes the amino acids to fluoresce.
- Spray the finished chromatogram with sulfuric acid that causes a brown spot for organic compounds.

c $R_f = \frac{\text{distance from origin by component}}{\text{distance moved from origin by solvent}} = \frac{0.6}{5} = 0.12$

d Leucine. The component in a mixture bonded least strongly to the stationary phase will move the greatest distance from the origin.

10 The component most strongly adsorbed to the stationary phase is band *A*.

As band B begins to emerge from the column, it appears to separate into two bands. These two components would be more readily separated if you *increased the column length.*

HPLC differs from column chromatography because the particles in the stationary phase are *smaller* and *high* pressure is applied to the mobile phase.

ISBN 978 1 4886 1772 0

11 A. HPLC does not directly produce measures on concentration. Standard solutions are used to find the relationship between peak area and concentration.

12 A. All forms of chromatography are based on the adsorption of the component particles to the stationary phase and desorption back into the mobile phase.

13 A. A carrier gas should be used that is unlikely to react with the components of samples under analysis. Of the gases listed, oxygen is much more reactive than the others.

14 B Obtain chromatograms of the stream water and of standard solutions of the herbicide.

D Use R_t values to identify the herbicide peak on the chromatogram of stream water.

A Measure the herbicide peak areas from the chromatograms of the standard solutions and the stream water sample.

C Construct a calibration curve and mark the herbicide peak area from the stream water sample on it.

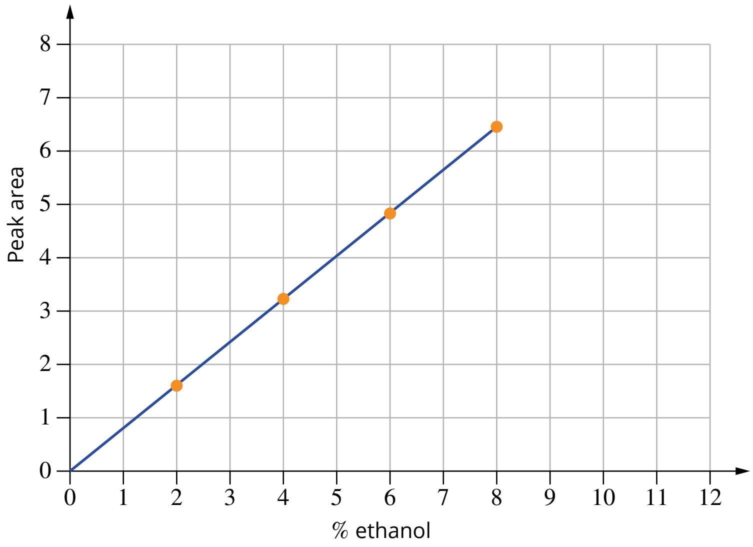
E Determine the herbicide concentration in the stream water sample.

15 **a** Chemists can determine the number of components and the absorption of components, from which concentration can be determined.

b four (The peaks of two components are overlapping.)

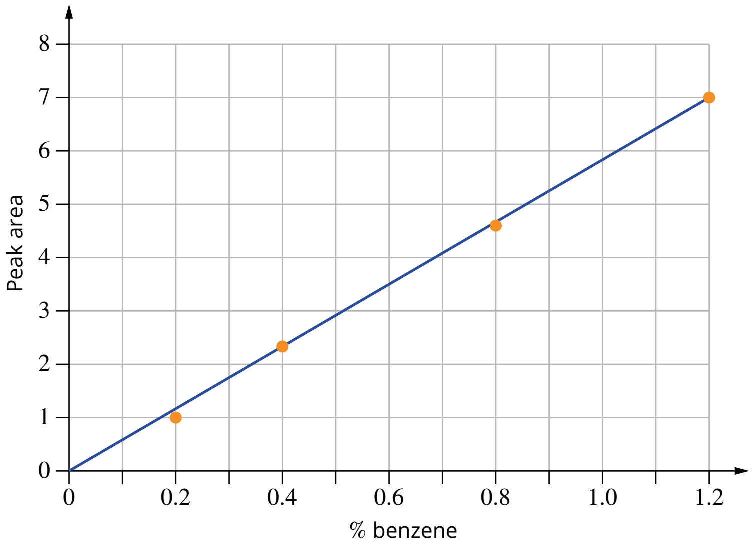
c Solid samples are dissolved in a suitable solvent. The liquid sample is injected into the top of an HPLC column. The stationary and mobile phases are chosen to achieve a good separation of the components in the sample. The sample components alternately adsorb onto the stationary phase and then desorb into the solvent as they are swept forward. The time taken to exit the column increases if the component strongly absorbs onto the stationary phase and has a low solubility in the mobile phase.

16 **a**



b 7.0%

17



The concentration of benzene is 0.6%.

ISBN 978 1 4886 1772 0

18 B. The results obtained from HPLC would provide identification and accurate concentrations of the dioxins present.

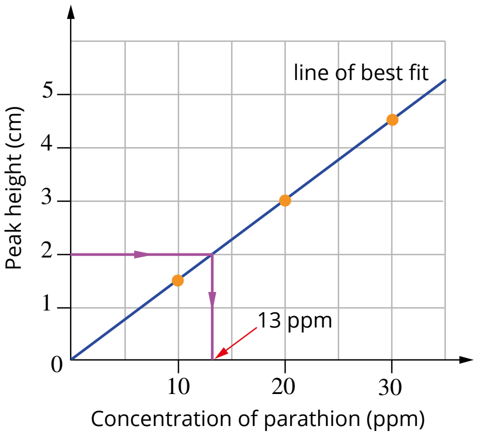
19

Technique	**Problem**
paper or thin-layer chromatography	B It is difficult or impossible to obtain quantitative data. D Samples must be able to be dissolved in solvent.
high performance chromatography (HPLC)	A It requires large amounts of solvents to operate. C Expensive equipment is needed. D Samples must be able to be dissolved in solvent.
gas chromatography (GC)	C Expensive equipment is needed.

20 **a**

Standard	**Peak height (cm)**
Standard 0 ppm parathion	0.2
Standard 10 ppm parathion	1.5
Standard 20 ppm parathion	3.1
Standard 30 ppm parathion	4.4
Reservoir water	2.0

b



c concentration of parathion in water sample = 13 ppm

d 13 ppm ≡ 13 μg/mL ≡ 13 mg/L

No, this is more than 1000 times the allowable concentration.

e LD_{50} of 8 mg/kg ≡ (8 mg/1000g) × 150 g = 1.2 mg. A mouse of 150 g has a 50% chance of being killed by a 1.2 mg dose of parathion.

volume of water that contains 1.2 mg parathion = $\frac{1.2}{13}$ L = 0.092 L = 92 mL

ISBN 978 1 4886 1772 0

Chapter 14 Gases

Section 14.1 Introducing gases

Worked example: Try yourself 14.1.1

CONVERTING VOLUME UNITS

A gas has a volume of 700 mL.

What is its volume in:

a cubic centimetres (cm^3)?

b litres (L)?

c cubic metres (m^3)?

Thinking	**Working**
Recall the conversion factors for each of the units of volume. Apply the correct conversion to each situation.	**a** The units of mL and cm^3 are equivalent. $1\,mL = 1\,cm^3$ $700\,mL = 700\,cm^3$ **b** $1000\,mL = 1\,L$ Divide volume in mL by 1000 to convert to L. $700\,mL = \frac{700}{1000}$ $= 0.700\,L$ **c** $1 \times 10^6\,mL = 1\,m^3$ Divide volume in mL by 1×10^6 to convert to m^3. $700\,mL = \frac{700}{1 \times 10^6}$ $= 7.00 \times 10^{-4}\,m^3$

Worked example: Try yourself 14.1.2

CONVERTING PRESSURE UNITS

Cyclone Yasi was one of the biggest cyclones in Australian history.

a The atmospheric pressure in the eye of Cyclone Yasi was measured as 0.902 bar. What was the pressure in kilopascals (kPa)?

b What was the pressure in the eye of Cyclone Yasi in atmospheres (atm) if it was known to be 677 mmHg?

c If the atmospheric pressure in the eye of Cyclone Yasi was 0.891 atm, what was the pressure in kilopascals (kPa)?

d The atmospheric pressure in the eye of Cyclone Yasi was 677 mmHg. What was the pressure in bars?

Thinking	**Working**
a To convert bars to kilopascals, use the conversion relationship: 1 bar = 100 kPa To change bar to kPa, multiply the value by 100.	$0.902\,bar = 0.902 \times 100$ $= 90.2\,kPa$
b To convert millimetres of mercury to atmospheres, use the relationship: 1 atm = 760 mmHg To change mmHg to atm, divide the value by 760.	$677\,mmHg = \frac{677}{760}$ $= 0.891\,atm$
c To convert atmospheres to kilopascals, use the conversion relationship: 1 atm = 101.3 kPa To change atm to kPa, multiply the value by 101.3.	$0.891\,atm = 0.891 \times 101.3$ $= 90.2\,kPa$

d	This can be done in two steps. First, convert millimetres of mercury to atmospheres. Use the conversion relationship: 760 mmHg = 1 atm To change mmHg to atm, divide the value by 760. Keep the answer in your calculator and proceed to the next step. Next, convert atmospheres to bar. Use the conversion relationship: 1 atm = 1.013 bar To change atm to bar, multiply the quotient from the previous step by 1.013.	$677\,\text{mmHg} = \frac{677}{760}$ $677\,\text{mmHg} = \frac{677}{760} \times 101.3$ $= 0.902\,\text{bar}$

14.1 KEY QUESTIONS

1 **a** Molecules of gases are in constant, rapid, random motion and the forces between molecules are negligible. They continue to move outwards until stopped by the walls of the container, filling all the space available.

b Most of the volume occupied by a gas is space, so compression can be achieved by reducing the space between the particles.

c The molecules in a gas are spread much further apart than those of a liquid. A given mass of gas would occupy a much greater volume than the same mass of the liquid phase. Therefore, the density of the gas is less.

d Gases mix easily together because of the large amount of space between the molecules.

e The pressure exerted by a gas depends on the number of collisions of gas particles and the wall of the container. The pressure is independent of the type of gas involved. The total pressure exerted by a mixture of gases will depend on the total number of collisions each gas has with the container.

2 **a** Tyres have a recommended maximum pressure to give a comfortable ride as well as good traction on the road. If the pressure in a tyre is too high, the gas inside cannot be compressed as easily and passengers will be more aware of bumps on the road.

b During a long journey on a hot day, the air in a tyre warms up. This means the air molecules have increased kinetic energy, and collisions with the walls of the tyres will increase in frequency and exert more force, and so the pressure will increase.

c Particles from the cooking food escape the pot and move randomly through the house. If the food has an odour, and if there are enough particles in the air, you will detect the odour as you enter the house.

d As air is pumped into a balloon, air molecules collide with the rubber of the balloon, forcing it to expand. If too much air is pumped in, the balloon reaches a stage where it cannot stretch any further. If the number of collisions by molecules per given surface area is increased still further, the rubber will break.

3 The order should be from gases with the smallest molecular mass to largest molecular mass.

He < Ne < Ar < Kr

This is because, at the same temperature, gases have the same kinetic energies. As the mass of the gas increases, its velocity decreases. Therefore, it will diffuse more slowly.

4 **a** As temperature increases, the average kinetic energy of gas molecules in the can will increase. This will lead to an increase in the frequency and force of collisions of gas molecules with the inside walls of the aerosol cans. This will cause an increase in pressure.

b As the syringe is compressed, the inside surface area of the syringe will decrease. The number of collisions of molecules per unit area per second with the inside walls of the syringe will increase. This will cause a pressure increase.

5 These conversions should be applied as needed: $1.00\,\text{atm} = 760\,\text{mmHg} = 101.3\,\text{kPa} = 1.013 \times 10^5\,\text{Pa} = 1.013\,\text{bar}$

a $140\,\text{kPa} = 140 \times 1000\,\text{Pa} = 1.40 \times 10^5\,\text{Pa}$

b $92\,000\,\text{Pa} = \frac{92\,000}{1000}\,\text{kPa} = 92\,\text{kPa}$

c $1.00\,\text{atm} = 760\,\text{mmHg}$

So $4.24\,\text{atm} = 760 \times 4.24\,\text{mmHg} = 3.22 \times 10^3\,\text{mmHg}$

$1.00\,\text{atm} = 1.013 \times 10^5\,\text{Pa}$

So $4.24\,\text{atm} = 1.013 \times 10^5 \times 4.24\,\text{Pa} = 4.30 \times 10^5\,\text{Pa}$

d 101.3 kPa = 760 mmHg

So 120 kPa = $\frac{760}{101.3} \times 120$ kPa = 900 mmHg

101.3 kPa = 1.00 atm

So 120 kPa = $\frac{1.00}{101.3} \times 120$ atm = 1.18 atm

101.3 kPa = 1.013 bar

So 120 kPa = $\frac{1.013}{101.3} \times 120$ atm = 1.20 bar

6 These conversions should be applied as needed: 1 mL = 1 cm^3, 1 L = 1000 mL = 1000 cm^3, 1 m^3 = 1 × 10^6 cm^3 = 1 × 10^6 mL = 1000 L

a 2 L = 2 × 1000 mL = 2 × 10^3 mL

b 4.5 L = $\frac{4.5}{1000}$ = 4.5 × 10^{-3} m^3

c 2250 mL = $\frac{2250}{1000}$ = 2.250 L

d 120 mL = $\frac{120}{1000}$ = 0.120 L

Section 14.2 Molar volume of a gas

Worked example: Try yourself 14.2.1

CALCULATING THE VOLUME OF A GAS FROM ITS AMOUNT (IN MOL)

Calculate the volume, in L, occupied by 3.50 mol of oxygen gas at STP. Assume that oxygen behaves like an ideal gas.	
Thinking	**Working**
Rearrange $n = \frac{V}{V_m}$ to make volume the subject.	$n = \frac{V}{V_m}$ $V = n \times V_m$
Substitute in the known values where V_m = 22.71 L mol^{-1} (at STP) and solve.	$V_{(STP)} = n \times V_m$ $= 3.5 \times 22.71$ = 79.485 L
Consider the units and significant figures. The answer should be given to the smallest number of significant figures in the measurement.	$V_{(STP)}$ = 79.5 L

Worked example: Try yourself 14.2.2

CALCULATING THE AMOUNT OF A GAS FROM ITS VOLUME

Calculate the amount (in mol) of oxygen gas in a volume of 3.5 L at STP. Assume that oxygen behaves like an ideal gas.	
Thinking	**Working**
Use $n = \frac{V}{V_m}$ where V_m = 22.71 L mol^{-1} (at STP)	$n = \frac{V}{V_m}$ $= \frac{3.5}{22.71}$ = 0.1541 mol
Consider the units and significant figures. The answer should be given to the smallest number of significant figures in the measurement.	n = 0.15 mol

14.2 KEY QUESTIONS

1 **a** 32 L
  b 23 mL
  c 1.1 L

2 **a** 0.123 mol
  b 2.20 mol
  c 6.16×10^{-3} mol

3 D. Under standard conditions (STP), 1 mol of any gas has a volume of 22.71 L.

Use the formula: $n = \frac{V}{V_m}$ where n is the amount in moles, V is the volume in litres and V_m is the molar volume in mol L^{-1}.

To calculate V, the formula is rearranged to $V = n \times V_m$

Section 14.3 Calculations involving reactions with gases

Worked example: Try yourself 14.3.1

USING MOLE RATIOS

How many moles of carbon dioxide are generated when 2.55×10^3 moles of butane (C_4H_{10}) are burned completely in oxygen?	
Thinking	**Working**
Write a balanced equation for the reaction.	$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 5H_2O(g)$
Note the number of moles of the known substance.	$n(C_4H_{10}) = 2.55 \times 10^3$ mol
Write a mole ratio for: $\frac{\text{coefficient of unknown}}{\text{coefficient of known}}$	$\frac{n(CO_2)}{n(C_4H_{10})} = \frac{4}{1}$
Calculate the number of moles of the unknown substance using: $n(\text{unknown}) = \text{mole ratio} \times n(\text{known})$	$n(CO_2) = \frac{4}{1} \times 2.55 \times 10^3$ $= 1.02 \times 10^2$ mol

Worked example 14.3.2: Try yourself

MASS–VOLUME STOICHIOMETRIC CALCULATIONS

A sample of calcium carbonate ($CaCO_3$) is heated until it has decomposed completely to form calcium oxide (CaO) and carbon dioxide (CO_2). A mass of 3.000 g of calcium oxide is produced. Calculate the volume of carbon dioxide produced, measured at STP. Give your answer to 3 significant figures.	
Thinking	**Working**
Write a balanced equation for the reaction.	$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
Identify the known quantity and the unknown.	Known: mass of CaO = 3.000 g Unknown: What is the volume of CO_2 at STP?
Calculate the number of moles, n, of the known substance using: $n = \frac{m}{M}$	$n(CaO) = \frac{3.000}{56.08}$ $= 0.05350$ mol
Find the mole ratio: $\frac{\text{coefficient of unknown}}{\text{coefficient of known}}$	$\frac{n(CO_2)}{n(CaO)} = \frac{1}{1}$
Calculate the number of moles of the unknown substance using: $n(\text{unknown}) = \text{mole ratio} \times n(\text{known})$	$n(CO_2) = n(CaO)$ $= 0.05350$ mol
Calculate the volume of the unknown substance using: $V = n \times 22.71$	$V(CO_2) = n(CO_2) \times V_m$ $= 0.05350 \text{ mol} \times 22.71 \text{ L mol}^{-1} = 1.21$ L

ISBN 978 1 4886 1772 0

Worked example: Try yourself 14.3.3

GAS VOLUME–VOLUME CALCULATIONS

Methane gas (CH_4) is burned in a gas stove according to the following equation:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

If 50 mL of methane is burned in air, calculate the volume of CO_2 gas produced under constant temperature and pressure conditions.

Thinking	**Working**
Use the balanced equation to find the mole ratio of the two gases involved.	1 mol of CH_4 gas produces 1 mol of CO_2 gas.
The temperature and pressure are constant, so volume ratios are the same as mole ratios.	1 volume of CH_4 produces 1 volume of CO_2 gas, so 50 mL of CH_4 reacts with 50 mL of CO_2.

Worked example: Try yourself 14.3.4

EXCESS REACTANT CALCULATIONS

Calculate the volume of carbon dioxide, in L, produced when 65.0 g of butane is burned completely in 200 L of oxygen. The gas volume is measured at STP. The equation for the reaction is:

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$$

a Which reactant is the limiting reactant?

b What is the volume of carbon dioxide formed at STP?

Thinking	**Working**
a Calculate the number of moles of each reactant using $n = \frac{m}{M}$ or $n = \frac{V}{V_m}$ as appropriate.	$n(C_4H_{10}) = \frac{m}{M}$ $n(C_4H_{10}) = \frac{65}{58.12}$ $= 1.12$ mol $n(O_2) = \frac{200}{22.71}$ $= 8.81$ mol
Use the coefficients of the equation to find the limiting reactant.	The equation shows 2 mol of C_4H_{10} reacts with 13 mol of O_2. So to react all of the O_2 will require $\frac{2}{13} \times n(O_2)$ of C_4H_{10} $= \left(\frac{2}{13}\right) \times 8.81$ $= 1.36$ mol As there is 1.12 mol of C_4H_{10}, the O_2 is in excess. The C_4H_{10} is the limiting reactant (it will be completely consumed).
b Find the mole ratio using: $\frac{\text{coefficient of unknown}}{\text{coefficient of known}}$ The limiting reactant is the known substance.	$\frac{n(CO_2)}{n(C_4H_{10})} = \frac{8}{2}$
Calculate the number of moles of the unknown substance using: $n(\text{unknown}) = \text{mole ratio} \times n(\text{known})$	$n(CO_2) = 4 \times 1.12$ $= 4.48$ mol
Calculate the required quantity of the unknown using: $m = n \times M$ or $V = n \times 22.71$ as appropriate.	$V(CO_2) = 4.48 \times 22.71$ $= 101.7$ L

ISBN 978 1 4886 1772 0

14.3 KEY QUESTIONS

1 **a** $\frac{n(CH_3OH)}{n(O_2)} = \frac{2}{3}$

b $\frac{n(O_2)}{n(H_2O)} = \frac{3}{4}$

c $\frac{n(CH_3OH)}{n(CO_2)} = \frac{2}{2} = 1$

2
- Write a balanced equation for the reaction.
- Identify the known and unknown substances in the question.
- Calculate the amount, in mol, of the known substance using $n = \frac{m}{M}$.
- Use mole ratios from the equation to calculate the amount of the unknown.
- Calculate the mass of the unknown substance using $m = n \times M$.

3 6.00 g of NH_3

4 **a** 0.0200 mol

b 0.0400 mol

c 0.909 L

5 2.04 L of O_2

6 7.93 g of H_2O

CHAPTER 14 REVIEW

1 B

2 A. Temperature is a measure of the average kinetic energy of the particles. As the temperature increases, the average kinetic energy of the particles also increases. Average kinetic energy is proportional to the average speed of the particles. As average kinetic energy increases, the average speed of the particles also increases.

3 A

4 smaller, straight-line, weak, elastic, directly

5 **a** As volume is reduced, there is an increase in the frequency of molecular collisions per unit wall area. This is measured as an increase in pressure.

b When the temperature of a gas is lowered, the average kinetic energy of the particles decreases. The rate of collisions between particles and the walls of the container decreases and particles collide with less force. As pressure is a measure of the force of molecular collisions per unit wall area of the container, pressure is found to decrease.

c In a mixture of gases, the particles of each gas are moving and colliding with the walls of the container, independently of each other. Each gas therefore exerts a pressure. As the gases behave independently of each other, total pressure is simply the sum of the individual gas (or partial) pressures.

d When more gas is added to a container, the total number of particles in the container increases. Provided that the volume of the container and the temperature have not changed, the collisions of these additional particles means that the total pressure in the container has increased.

6 **a** The pressure inside the container is reduced when some of the gas escapes.

b There are fewer gas molecules to collide with each other and the walls of the container. Pressure is the force exerted by the molecules over a defined area, so this will decrease.

7 B

8 $V = \frac{500}{1000} = 0.500\ L$

$$n = \frac{V}{V_m} = \frac{0.500}{22.71} = 0.022\ mol$$

ISBN 978 1 4886 1772 0

9 $n = \frac{0.1100}{44.01} = 0.0025\,\text{mol}$

$V = 0.0025 \times 22.71$

$= 0.0568\,\text{L}$

10 C

11 **a** $n = \frac{m}{M}$

$= \frac{8.0}{32}$

$= 0.25\,\text{mol}$

$V = n \times V_m$

$= 0.25 \times 22.71$

$= 5.7\,\text{L}$

b $n = \frac{V}{V_m}$

$= \frac{10}{22.71}$

$= 0.44\,\text{mol}$

$m = n \times M$

$= 0.44 \times 46$

$= 20\,\text{g}$

12 **a** $m = n \times M$

$= 1.00 \times 44.0$

$= 44.0\,\text{g}$

b $V = n \times V_m$

$= 1.00 \times 22.71$

$= 22.7\,\text{L}$

c $p = \frac{m}{V}$

$= \frac{44.0}{22.7}$

$= 1.94\,\text{g}\,\text{L}^{-1}$

13 **a** $n = \frac{V}{V_m}$

$= \frac{220\,000}{22.71}$

$= 9687\,\text{mol}$

$= 9.69 \times 10^3\,\text{mol}$

b mol of nitrogen $= 9.69 \times 10^3\,\text{mol} \times \frac{80}{100}$

$= 7.75 \times 10^3\,\text{mol}$

$m = n \times M$

$= 7.75 \times 10^3 \times 14$

$= 108\,494\,\text{g}$

$= 108\,\text{kg}$

14 The balanced equation shows that 1 mol of glucose produces 2 mol of ethanol and carbon dioxide. The number of moles of glucose is found using $n = \frac{m}{M}$ where m is the mass in grams and M is the molar mass in $\text{g}\,\text{mol}^{-1}$. Use a periodic table to calculate the molar masses. $M(\text{glucose}) = 180.0\,\text{g}\,\text{mol}^{-1}$ and $M(\text{ethanol}) = 46.068\,\text{g}\,\text{mol}^{-1}$.

$\frac{n(\text{ethanol})}{n(\text{glucose})} = \frac{2}{1}$

$n(\text{glucose}) = \frac{80.0}{180.156} = 0.44406\,\text{mol}$

$n(\text{ethanol}) = \frac{2}{1} \times n(\text{glucose}) = 0.8881\,\text{mol}$

$\therefore m(\text{ethanol}) = 0.8881 \times 46.068 = 40.9\,\text{g}$

ISBN 978 1 4886 1772 0

15 The balanced equation shows that 1 mol carbon produces 1 mol of carbon dioxide. The number of moles of each is found using $n = \frac{m}{M}$ where m is the mass in grams and M is the molar mass in g mol^{-1}. Use a periodic table to calculate the molar masses. M(carbon) = 12.0 g mol^{-1} and $M(CO_2)$ = 44.0 g mol^{-1}.

$$\frac{n(\text{carbon dioxide})}{n(\text{carbon})} = \frac{1}{1}$$

$$n(\text{carbon}) = \frac{1000000}{12.0} = 8.3 \times 10^4\,\text{mol}$$

$$n(\text{carbon dioxide}) = n(\text{carbon}) = 8.3 \times 10^4\,\text{mol}$$

$$\therefore m(\text{carbon dioxide}) = 8.3 \times 10^4 \times 44.0 = 3.7 \times 10^6\,\text{g}$$
$$= 3.7\,\text{tonne}$$

16 a i $n(C_3H_8) = \frac{m}{M}$

$= \frac{22}{44}$

$= 0.50$ mol

$n(O_2)$ = mole ratio × $n(C_3H_8)$

$= \frac{5}{1} \times 0.50$

$= 2.50$ mol

$V = n \times V_m$

$= 2.50 \times 22.71$

$= 57$ L

ii $n(CO_2)$ = mole ratio × $n(C_3H_8)$

$= \frac{3}{1} \times 0.50$

$= 1.50$ mol

$V = n \times V_m$

$= 1.50 \times 22.71$

$= 34$ L

b i $n(C_3H_8) = \frac{m}{M}$

$= \frac{5.0}{44}$

$= 0.11$ mol

$n(O_2)$ = mole ratio × $n(C_3H_8)$

$= \frac{5}{1} \times 0.11$

$= 0.57$ mol

$V = n \times V_m$

$= 0.57 \times 22.71$

$= 13$ L

ii $n(CO_2)$ = mole ratio × $n(C_3H_8)$

$= \frac{3}{1} \times 0.11$

$= 0.34$ mol

$V = n \times V_m$

$= 0.34 \times 22.71$

$= 7.7$ L

ISBN 978 1 4886 1772 0

c **i** $n(C_3H_8) = \frac{m}{M}$

$= \frac{0.145}{44}$

$= 3.30 \times 10^{-3}$ mol

$n(O_2) =$ mole ratio $\times n(C_3H_8)$

$= \frac{5}{1} \times 3.30 \times 10^{-3}$

$= 1.65 \times 10^{-2}$ mol

$V = n \times V_m$

$= 1.65 \times 10^{-2} \times 22.71$

$= 0.374$ L

ii $n(CO_2) =$ mole ratio $\times n(C_3H_8)$

$= \frac{3}{1} \times 3.30 \times 10^{-3}$

$= 9.89 \times 10^{-3}$ mol

$V = n \times V_m$

$= 9.89 \times 10^{-3} \times 22.71$

$= 0.225$ L

17 A. The pressure of a fixed amount of gas is independent of the identity of the gas, so it doesn't matter what the gas is, as long as the sample is at the same temperature and the amount is the same, the pressure will be the same.

18 **a** $n(Fe_2O_3) = \frac{m}{M}$

$= \frac{350000}{159.69}$

$= 2.19 \times 10^3$ mol

b $\frac{n(CO_2)}{n(Fe_2O_3)} = \frac{3}{1}$

$n(CO_2) = 3 \times n(Fe_2O_3)$

$= 6.58 \times 10^3$ mol

c $V(CO_2) = n \times V_m$

$= 6.58 \times 10^3 \times 22.71$

$= 1.49 \times 10^5$ L

19 **a** Equal. With pressure, volume and temperature the same, *n* will be the same.

b Carbon dioxide. Each CO_2 molecule contains three atoms and each O_2 molecule contains two atoms. As there is an equal number of molecules of each gas, there are more atoms in the CO_2 sample.

c Carbon dioxide. Density = mass ÷ volume. The volume is the same for each gas, but the mass of CO_2 is greater, so it has the greater density.

20 **a** $\frac{n(NO_2)}{n(HNO_3)} = \frac{4}{4} = 1$

$n(NO_2) = 1 \times n(HNO_3)$

$= 0.500$ mol

b $\frac{n(O_2)}{n(HNO_3)} = \frac{1}{4} = 0.25$

$n(O_2) = 0.25 \times n(HNO_3)$

$= 0.125$ mol

c The total moles of gases produced is the sum of the moles of the individual gases.

$n(\text{gases}) = n(NO_2) + n(O_2)$

$= 0.500 + 0.125$

$= 0.625$ mol

$V(\text{gases}) = n(\text{gases}) \times V_m$

$= 0.625 \times 22.71$

$= 14.2$ L

ISBN 978 1 4886 1772 0

21 **a** $n(O_2) = \frac{V}{V_m}$

$= \frac{350}{22.71}$

$= 15.4\,\text{mol}$

b $n(CuFeS_2) = \text{mole ratio} \times n(O_2)$

$= \frac{2}{5} \times 15.4$

$= 6.16\,\text{mol}$

c $m(CuFeS_2) = n(CuFeS_2) \times M(CuFeS_2)$

$= 6.16 \times 183.53$

$= 1130\,\text{g}$

$= 1.13\,\text{kg}$

22 **a** As the air is heated, the molecules move faster. Because a hot-air balloon has an opening at the bottom, some of the gas molecules escape. The balloon filled with warmer air has fewer molecules in the same volume as a balloon filled with cooler air. Fewer molecules have less mass. Therefore, the warmer balloon has less mass per unit volume, or density, than the cooler balloon.

b The molecules in a sample of gas at any temperature have a range of kinetic energies. Some have low energy and some have high energy. As the temperature of a gas increases, the average kinetic energy of molecules increases. However, as the temperature increases, there will still be some molecules with low kinetic energy. So the student is mistaken, as not all of the kinetic energies will have increased.

c When the temperature of a gas is lowered, the average kinetic energy of the particles decreases. The frequency of collisions between particles and the walls of the container decreases and particles collide with less force. As the pressure is a measure of the force of molecular collisions per unit surface area of the container, the pressure of the gas will decrease.

ISBN 978 1 4886 1772 0

Chapter 15 Properties and uses of water

Section 15.1 Essential water

15.1 KEY QUESTIONS

1 Surface water has the highest risk of contamination since it can be easily polluted. It is not a protected water source like mains water where the water is tested and treated if needed.

2 Most of the freshwater on Earth is locked in ice caps, glaciers or the soil.

3 Australia is the driest inhabited continent. Rainfall in Australia is extremely variable. Most of the rain that falls evaporates before it can enter rivers and reservoirs.

4 To determine if the water is mixed and therefore what depth to take the samples from. If the temperature is consistent, it can be assumed that the water is thoroughly mixed and samples can be taken halfway down. If there is a temperature variation, then accurate information can only be obtained if a sample is taken in the middle of each temperature region.

5 A, B and C are all correct and can be seen in Figure 15.1.13.

6 Fluoride in drinking water helps to reduce the incidence of tooth decay. It does this by interacting with tooth enamel. The fluoride ion replaces the hydroxyl ion in tooth enamel, forming fluorapatite, which is stronger and more resistant to decay.

Section 15.2 Properties of water

Worked example: Try yourself 15.2.1

CALCULATING THE AMOUNT OF ENERGY REQUIRED TO HEAT A SPECIFIED MASS OF A SUBSTANCE USING SPECIFIC HEAT CAPACITY

Calculate the heat energy, in kJ, needed to increase the temperature of 375 g of water by 45.0°C.	
Thinking	**Working**
Find the specific heat capacity (C) of the substance from the data in Table 15.2.4.	The specific heat capacity of water is $4.18\,\text{J g}^{-1}\,°\text{C}^{-1}$.
To calculate the quantity of heat energy in joules, use the formula: $q = C \times m \times \Delta T$	$q = 4.18 \times 375 \times 45$ $= 7.05 \times 10^4\,\text{J}$
Express the quantity of energy in kJ. Remember that to convert from J to kJ, you multiply by 10^{-3}.	$q = 7.05 \times 10^4 \times 10^{-3}$ $= 70.5\,\text{kJ}$

15.2 KEY QUESTIONS

1 **a** Physical properties that are unusual to water include:
- relatively high melting and boiling temperatures for its molecular size
- decrease in density on freezing
- high heat capacity
- high latent heat of fusion and evaporation for a substance of its molecular size.

b The bond between H and O atoms in water is highly polar. As a result, hydrogen bonds exist between water molecules. Hydrogen bonds are stronger than other intermolecular bonds (although still weaker than the covalent intramolecular bonds) and so require more energy to break. Thus, water has relatively high melting and boiling temperatures. Hydrogen bonding between water molecules in ice results in a very open arrangement of molecules so ice is less dense than liquid water.

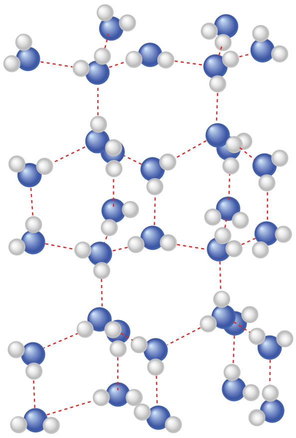
ISBN 978 1 4886 1772 0

2 Each water molecule has two hydrogen atoms and one oxygen atom. The oxygen atom has two pairs of non-bonding electrons, each of which can form one hydrogen bond.

So the maximum number of water molecules with which one water molecule can form hydrogen bonds is four: up to two hydrogen bonds involving the two hydrogen atoms and up to two hydrogen bonds involving the two pairs of non-bonding electrons on the oxygen atom.

3 **a** In ice, the distribution of hydrogen bonds in water is almost tetrahedral in shape. In order for these hydrogen bonds to form as the water solidifies, the water molecules must move into these positions, so occupying more volume than in liquid water.

b



c As it freezes, water expands, unlike most liquids. This is because of hydrogen bonding. Each water molecule is surrounded by four others in what is almost a crystal-type situation. Therefore, ice is less dense than liquid water, and it floats on liquid water. (For most liquids, the solid is denser than the liquid.)

4 D. Specific heat capacity is the energy needed to increase the temperature of 1g of a substance by 1 degree, therefore D is correct.

5 B. A water molecule at the surface of a sample of water is not completely surrounded by other water molecules. Hydrogen bonding occurs between molecules on the sides and below, but not with the atmosphere above.

6 At the melting point the flat region of the graph represents the substance *changing from a solid to a liquid*. The energy change is equal to *the latent heat of fusion*. At the boiling point the flat region of the graph represents the substance *changing from a liquid to a gas*. The energy change is equal to *the latent heat of vaporisation.*

7 hydrogen bonding

Section 15.3 **Water as a solvent**

15.3 KEY QUESTIONS

1 **a** solvent
   b solution
   c solute
   d solute

2 They all have water as the solvent.

3 **a** soluble
   b insoluble
   c soluble
   d soluble
   e insoluble
   f soluble
   g insoluble

4 The solute–solvent forces are stronger than the solute–solute and solvent–solvent forces.

5

a	capable of dissolving	soluble
b	incapable of dissolving	insoluble
c	the minor component of a solution	solute
d	the process by which a substance is dissolved in a solvent	dissolution
e	a liquid mixture in which a solute is dispersed throughout a solvent	solution
f	the major component of a solution	solvent

6 CH_4 (methane). Methane is a non-polar molecule. According to the principle of 'like dissolves like', a non-polar substance will dissolve in a non-polar solvent.

Section 15.4 Water as a solvent of molecular substances

15.4 KEY QUESTIONS

1 B and E

An –OH group is polar and can form hydrogen bonds with water. As a consequence, methanol and ethylene glycol dissolve readily in water.

Hydrogen chloride contains polar covalent molecules; however, its molecules are not able to form hydrogen bonds with water molecules.

Ammonium sulfate and sodium hydroxide are ionic compounds containing polyatomic ions.

Diamond has a covalent lattice structure and is insoluble in water.

Hydrogen gas consists of non-polar molecules and will not dissolve in water.

2 $CH_3OH(l) \xrightarrow{H_2O(l)} CH_3OH(aq)$

$C_6H_{12}O_6(s) \xrightarrow{H_2O(l)} C_6H_{12}O_6(aq)$

3
- Hydrogen bonds between water molecules are broken.
- Hydrogen bonds between ethanol molecules are broken.
- New hydrogen bonds form between ethanol molecules and water molecules.

4 $HI(s) + H_2O(l) \rightarrow H_3O^+(aq) + I^-(aq)$

5 *Covalent bonds* within hydrogen chloride molecules are broken.

Hydrogen bonds between water molecules are broken. The HCl *ionises* and produces Cl^- and H^+ ions. *Covalent bonds* form between *H^+* ions and water to produce *hydronium* ions. *Ion–dipole bonds* form between the Cl^- and H_3O^+ ions and polar water molecules.

Section 15.5 Water as a solvent of ionic compounds

Worked example: Try yourself 15.5.1

DETERMINING IF IONIC COMPOUNDS ARE SOLUBLE OR INSOLUBLE IN WATER

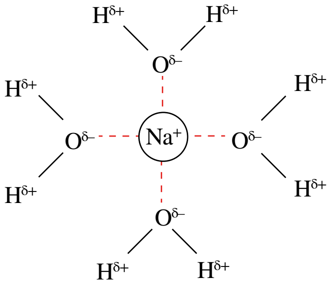
<table>
<tr><td colspan="2">Is ammonium phosphate ($(NH_4)_3PO_4$) soluble or insoluble in water? You will need to refer to the solubility table to complete this question.</td></tr>
<tr><th>Thinking</th><th>Working</th></tr>
<tr><td>Identify the ions that are present in the ionic compound.</td><td>Ammonium (NH_4^+) and phosphate (PO_4^{3-})</td></tr>
<tr><td>Check the solubility tables to see if compounds containing the cation are usually soluble or insoluble in water.</td><td>Compounds containing ammonium ions are soluble.</td></tr>
<tr><td>Check the solubility tables to see if compounds containing the anion are usually soluble or insoluble in water.</td><td>Ammonium phosphate is a noted water-soluble, phosphate ion–containing compound. So, ammonium phosphate will be soluble in water.</td></tr>
</table>

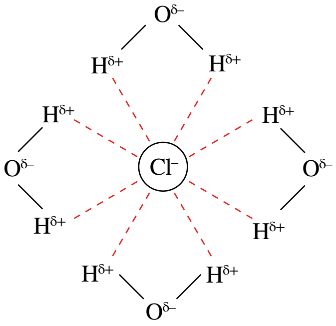
15.5 KEY QUESTIONS

1 $NaNO_3(s) \xrightarrow{H_2O(l)} Na^+ (aq) + NO_3^- (aq)$

$Ca(OH)_2(s) \xrightarrow{H_2O(l)} Ca^{2+} (aq) + 2OH^- (aq)$

2





- - - - represents ion–dipole interaction

The positive sodium ion attracts the partial negative charges on the oxygen atoms in the water molecule. The negative chloride ion attracts the partial positive charges on the hydrogen atoms in the water molecule.

3 A, B, D, E, H

4 A, C, D, E, F, H

5
- **a** Na^+/CO_3^{2-}
- **b** Ca^{2+}/NO_3^-
- **c** K^+/Br^-
- **d** Fe^{3+}/SO_4^{2-}
- **e** Cu^{2+}/Cl^-

6
- **a** Nitrates are highly soluble in water. If found on Earth, they would dissolve in rainwater and wash into the oceans. Therefore, they are found only in areas of low rainfall.
- **b** The high solubility of sodium, chloride and sulfate ions results in them dissolving and flowing into the world's oceans.

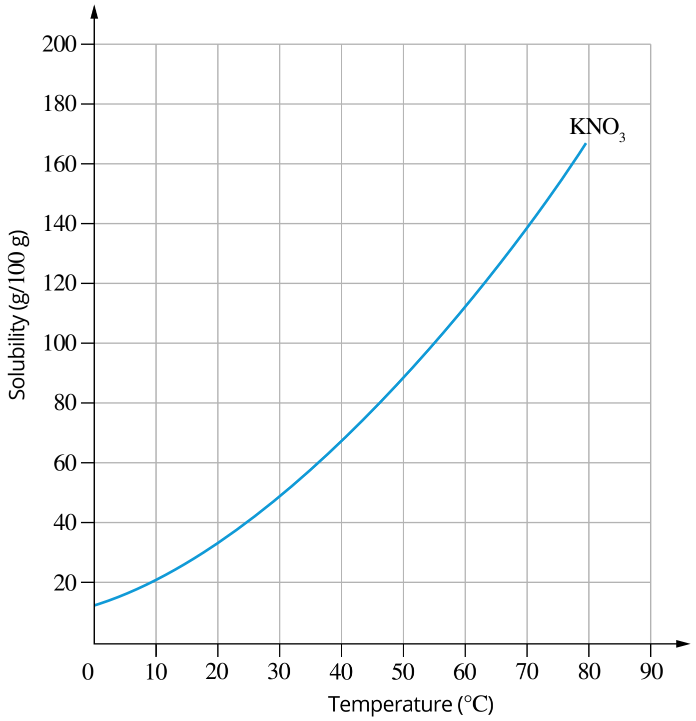
ISBN 978 1 4886 1772 0

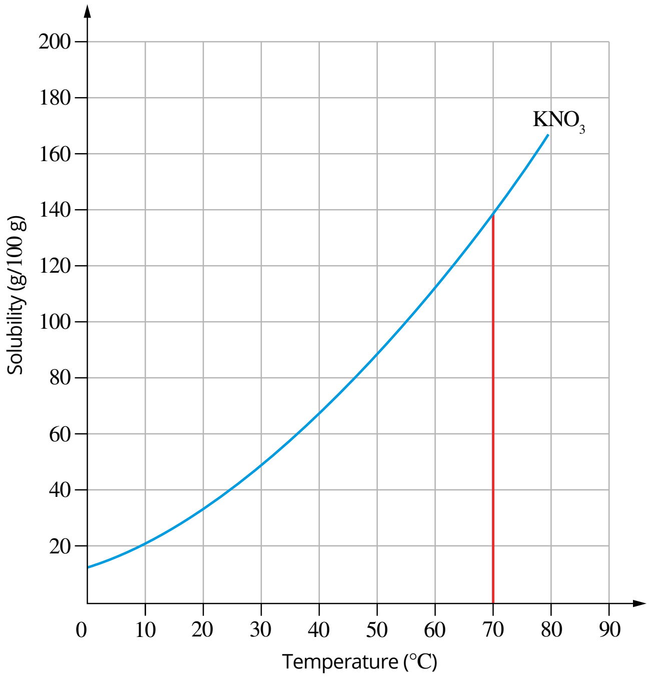
Section 15.6 **Solubility**

Worked example: Try yourself 15.6.1

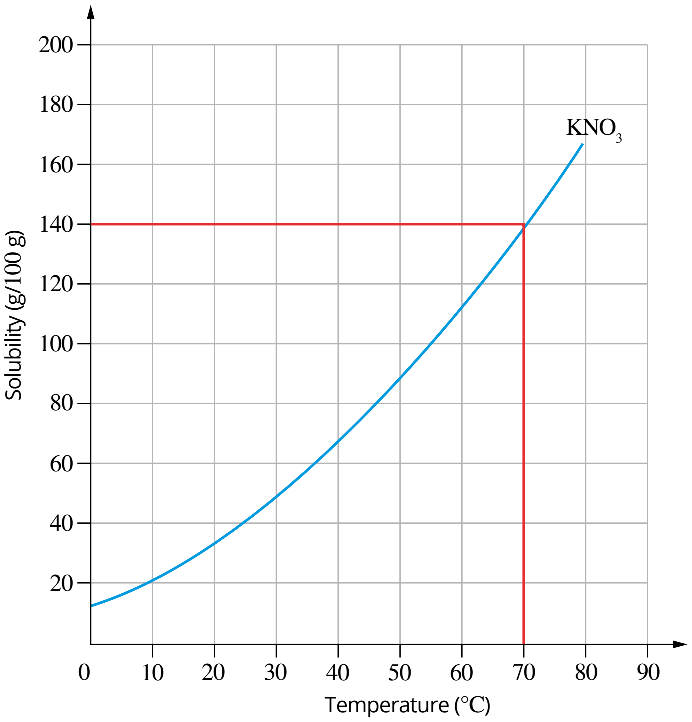
SOLUBILITY CURVE CALCULATIONS

Use the solubility curve below to find how many grams of potassium nitrate (KNO_3) will dissolve in 100 g of water at 70°C.



Thinking	**Working**
Draw a vertical line from the required temperature on the horizontal axis to intersect with the appropriate solubility curve.	

ISBN 978 1 4886 1772 0

Draw a horizontal line from the intersection point of the vertical line drawn in the previous step. The point where this horizontal line intersects with the vertical axis will give the mass of dissolved substance.	
Read the solubility from the graph.	The horizontal line intersects with the vertical axis at 140 g. Therefore, the mass of KNO_3 that will dissolve in 100 g of water at 70°C is 140 g.

Worked example: Try yourself 15.6.2

SOLUBILITY CURVES

A 120 g sample of sodium nitrate ($NaNO_3$) is added to 300 mL of H_2O at 40°C. Use Figure 15.6.2 to calculate how much more $NaNO_3$ needs to be added to make the solution saturated with $NaNO_3$ at 40°C.	
Thinking	**Working**
Use the solubility curve to find the mass of solute in a saturated solution of 100 g of H_2O at the required temperature.	Imagine a line drawn from 40°C on the horizontal axis to the solubility curve for $NaNO_3$ and find the corresponding value on the vertical axis. The value is 110 g.
Use the amount of solute that will dissolve in 100 g of H_2O to find the mass of solute to make a saturated solution in the mass of H_2O for this question.	The density of water is 1.0 g mL^{-1}, so 300 mL of water will weigh 300 g. So three times the mass of solute can dissolve in 300 g of water as in 100 g. $m(NaNO_3) = 3 \times 110\,g$ $= 330\,g$
To find out how much extra solute you need to add, find the difference between the mass of solute needed to make a saturated solution and how much has already been added.	120 g of $NaNO_3$ has already been added to 300 g H_2O. So the extra mass of $NaNO_3$ needed: $= 330\,g - 120\,g$ $= 210\,g$

Worked example: Try yourself 15.6.3

CRYSTALLISATION

If 200 g of sucrose is dissolved in 100 g water at 20°C, what mass of sucrose crystals will form if the temperature is reduced to 10°C?	
Thinking	**Working**
Identify the mass of solute dissolved in the original solution.	mass of sucrose in original solution = 200 g
Find the maximum mass of solute that will remain dissolved in 100 g water at the final temperature.	The solubility curve of sucrose shows that the maximum mass that will dissolve at 10°C is 190 g.
Calculate the mass of solute crystals that will form in the solution at the final temperature.	mass of crystals formed = original mass – remaining = 200 – 190 = 10 g

15.6 KEY QUESTIONS

1 C. A supersaturated solution is one that contains more solute than would normally be dissolved in a saturated solution. At 18°C, 35 g of sodium chloride would form a saturated solution in 100 g of water, so 17.5 g dissolves in 50 g of water. If 50 g of water at 18°C contained 20 g dissolved sodium chloride, it would be supersaturated.

2 The solubility curve represents the maximum mass of solute that can be dissolved at a set temperature. The green dot represents a supersaturated solution as it is above the solubility curve. The orange dot represents a saturated solution as it lies on the solubility curve. The blue dot represents an unsaturated solution as it lies below the solubility curve.

3 **a** 40 g
b 160 g
c 20 g

4 The mass that will dissolve in each case is the mass found from the solubility curve multiplied by 2 as the question asks for the amount in 200 g of solvent, not 100 g.
a 48 g
b 200 g
c 70 g

5 B. At 55°C the maximum amount of KNO_3 that will dissolve in 100 g of water is 100 g. Therefore the mass of KNO_3 that will crystallise out of the solution at this temperature is 140 – 100 = 40 g.

6 20 g. The solubility curve for $CuSO_4{\cdot}5H_2O$ shows that the maximum amount of solute that will dissolve in 100 g of solution at 60°C is 40 g. Therefore the mass of crystals will be 60 – 40 = 20 g.

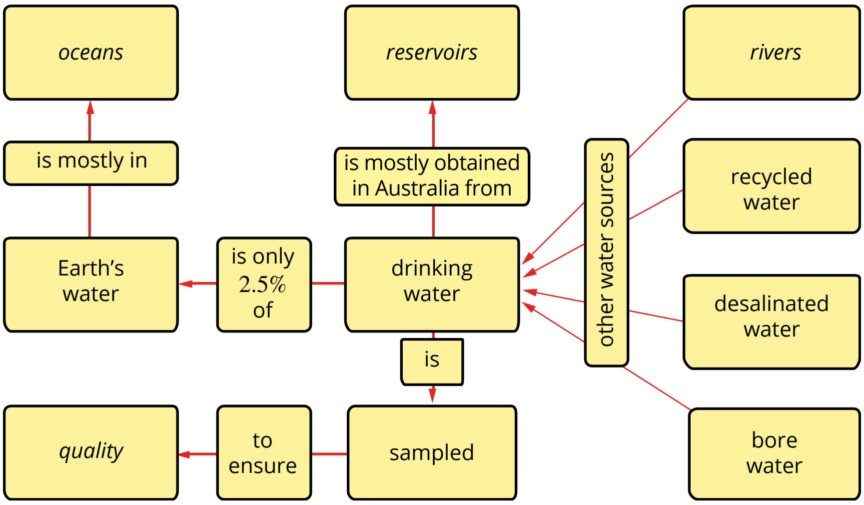
7 55 g. The solubility curve for $NaNO_3$ shows that the maximum amount of solute that will dissolve in 100 g of water at 50°C is 120 g. Therefore the mass that may dissolve in 50 g of water will be 60 g. As this is greater than the amount present, it will all remain dissolved.

8 Granite is a rock that is formed from magma cooling beneath the Earth's surface. This means the crystals are formed slowly. Crystals formed from slow cooling are larger than those formed from cooling quickly. Magma that cools above the surface forms basalt, the cooling process is faster, and therefore the crystals in basalt are smaller than those in granite.

ISBN 978 1 4886 1772 0

CHAPTER 15 REVIEW

1



2 A. Most of the drinkable water is found in the ice caps, although it is not readily accessible.

3 Bore water is water collected in aquifers (underground water-bearing rock) below the Earth's surface.

4 Most Australians live in the capital cities and their drinking water is supplied from reservoirs, as cities in Australia obtain their water from protected sources (although Perth sources a large amount of its water from desalination).

5 oceans, ice caps and glaciers, groundwater, ground ice and permafrost, lakes, soil moisture, atmosphere as water vapour, rivers

6 Considerations include:

- which chemical is to be analysed?
- why the testing is to take place?
- what are the health risks associated with the sampling?
- which equipment is required?
- what is the sample size required for the selected test?
- what method is to be used to record the measurements?
- where and at what depths should the samples be taken?
- how to obtain a representative sample?
- what are the labelling, storage and transport requirements for the sample?

7 D. Temperature, depth and location must be recorded and future measurements made at the same depth and location. The humidity has no significant effect on the result.

8 Levels of arsenic, cadmium and copper are below the Australian Drinking Water Guideline levels and are therefore acceptable. Lead and mercury are above the Guideline levels and are therefore unacceptable.

9 **a** unsafe position

b sample taken from surface of groundwater

c container not sterile

d container not protected from light

e water sample not secured safely

10 Water is a *polar* molecule. Within a single molecule, hydrogen and oxygen atoms are held together by strong *covalent bonds*. Between different molecules, the most significant forces are *hydrogen bonds.*

It is the relatively *high* strength of the intermolecular forces that give water its unique properties of:

- relatively *high* boiling point, *100*°C
- relatively *high* latent heat values, *6.0* kJ mol^{-1} and *44.0* kJ mol^{-1}
- relatively *high* specific heat capacity, *4.18* $\text{J g}^{-1}\,{}^{\circ}\text{C}^{-1}$.

11 **a** oxygen atom

b hydrogen atom

c hydrogen bond (and dispersion forces)

d covalent bond

12 **a** Intermolecular forces are those between one molecule and other molecules. For water, these are hydrogen bonds. Intramolecular forces are those holding the atoms together within a molecule. For water, these are covalent bonds.

b Covalent bonds are stronger. Evidence for this is the high temperatures required to break the bonds between the oxygen and hydrogen atoms inside the water molecule and so decompose it into its constituent gases. Changing liquid water into gaseous water involves breaking hydrogen bonds to separate one molecule from another. The lower temperatures needed to do so indicate that hydrogen bonds are weaker.

13 Water has a significantly higher melting point than hydrogen sulfide due to the hydrogen bonds between water molecules. Hydrogen sulfide cannot form hydrogen bonds.

14 It is the high polarity of the water molecule that allows relatively strong hydrogen bonding to occur between molecules. As a consequence, a relatively large quantity of energy is required to break the hydrogen bonds between water molecules when water changes from a liquid to a gas. This gives water a high boiling point.

15 Water has a high latent heat of vaporisation. Water is effective as a coolant because it absorbs a relatively large amount of energy when it evaporates, giving it a high latent heat of vaporisation.

16 D

$$\Delta T = \frac{q}{C \times m}$$

So the change in temperature of a substance (ΔT) is dependent upon the amount of heat energy transferred to the substance (q, option C), the mass of the substance (m, option A) and the heat capacity of the substance (C), which is related to the types of bonds present in the substance (option B).

17 dispersion forces, hydrogen bonds, covalent bonds

18 A solution is most likely to form when the polarity of bonding of the solute is similar to that of the solvent. The bonds formed between solute and solvent are then similar to those that existed between solute particles and between solvent particles. Water, being polar, is therefore a good solvent for ionic and polar substances.

19 Using the 'like dissolves like' rule, only polar substances will dissolve in water. Nitrogen is a non-polar molecule so will not dissolve well. Ethene is a non-polar hydrocarbon that will not be expected to dissolve in water. Ethanol, however, is a small polar molecule that can be expected to dissolve in water.

Full explanation:

- Nitrogen, N_2, is a covalent molecular substance. It is composed of non-polar molecules, so nitrogen would not be expected to dissolve well in water.
- Ethane, C_2H_6, is also a molecular compound. Ethane molecules contain slightly polar covalent bonds that are distributed symmetrically across the molecule. This makes the entire molecule non-polar, so ethane would be expected to not dissolve in water.
- Ethanol, C_2H_5OH, is a molecular compound that dissolves in water by forming hydrogen bonds with water molecules.

20 Carbon is slightly more electronegative than hydrogen, so each C–H bond is slightly polar. However, the resulting partial charges are distributed symmetrically across the octane molecule, making the molecule non-polar overall.

The energy released in the formation of solute–solvent bonds is not enough to overcome the intermolecular bonds between solute molecules and the intermolecular bonds between solvent molecules, so octane will not dissolve in water.

21 $C_6H_{12}O_6$ and C_3H_7OH are polar molecules. They contain the polar –OH group and so are able to form hydrogen bonds with water.

HI and HNO_3 contain polar molecules but they are unable to form hydrogen bonds with water. They dissolve by ionising.

I_2, CH_4 and C_2H_4 are non-polar covalent molecules. They do not dissolve well in water.

22 Y, Z, X

As the CH_3OH molecule is polar, it will dissolve readily in polar solvents. Because CH_3OH dissolves in Y, it can be ascertained that Y is a polar solvent. The non-polar nature of a CH_4 molecule means CH_4 will dissolve readily in non-polar solvents. It can be concluded that X is a non-polar solvent because CH_4 completely dissolves in it. CH_3OH molecules and CH_4 molecules partially dissolve in Z, which indicates it is more polar than X. No information is given about methanol dissolving in Z, although it can be implied.

23 DDT is most likely non-polar because it is soluble in fats, which are non-polar, and insoluble in water, which is polar.

24 **a** dissociation

b **i** $Cu^{2+}(aq)$, $NO_3^-(aq)$

ii $Zn^{2+}(aq)$, $SO_4^{2-}(aq)$

iii $NH_4^+(aq)$, $PO_4^{3-}(aq)$

ISBN 978 1 4886 1772 0

25 **a** K^+/CO_3^{2-}

b Pb^{2+}/NO_3^-

c Na^+/OH^-

d Na^+/SO_4^{2-}

e Mg^{2+}/Cl^-

f Zn^{2+}/NO_3^-

g K^+/S^{2-}

h Fe^{3+}/NO_3^-

26 **a** $MgSO_4(s) \xrightarrow{H_2O(l)} Mg^{2+}(aq) + SO_4^{2-}(aq)$

b $Na_2S(s) \xrightarrow{H_2O(l)} 2Na^+(aq) + S^{2-}(aq)$

c $KOH(s) \xrightarrow{H_2O(l)} K^+(aq) + OH^-(aq)$

d $(CH_3COO)_2Cu(s) \xrightarrow{H_2O(l)} 2CH_3COO^-(aq) + Cu^{2+}(aq)$

e $Li_2SO_4(s) \xrightarrow{H_2O(l)} 2Li^+(aq) + SO_4^{2-}(aq)$

27 Hydrated hydronium ions and chloride ions. Hydrogen chloride is a polar molecule that cannot form hydrogen bonds. When added to water, the molecule ionises to form a hydronium ion, H_3O^+, and a chloride ion, Cl^-. These ions become hydrated by ion–dipole attractions. There will also be water molecules and a very, very few HCl molecules.

28 Magnesium ions are cations; they have a positive charge. This means the negatively charged non-bonding electron pairs on the oxygen atoms in the water molecules are attracted to them.

The water molecules arrange around the magnesium ion with their oxygen atoms, rather than the hydrogen atoms, closest to the chloride ion.

Chloride ions are anions; they have a negative charge. This means the hydrogen atoms in the water molecules, which have a partial positive charge, are attracted to them. The water molecules arrange around the chloride ion with their hydrogen atoms, rather than the oxygen atoms, closest to the chloride ion.

29 There are a number of possible answers to this question. Use the information in Tables 15.5.1 and 15.5.2. For example, Na_2CO_3, Li_2CO_3 and K_2CO_3 are soluble, whereas $CaCO_3$, $MgCO_3$ and Ag_2CO_3 are insoluble.

30 There are a number of possible answers to this question. Use the information in Tables 15.5.1 and 15.5.2. For example, Na_2SO_4, K_2SO_4 and $(NH_4)_2SO_4$ are soluble, whereas $CaSO_4$, $BaSO_4$ and $PbSO_4$ are insoluble.

31 Potassium ions and bromide ions are held in the ionic lattice by ionic bonds that are based on electrostatic attraction. These ionic bonds in the solute break when it dissolves in water.

Water, the solvent, has hydrogen bonds between water molecules.

When potassium bromide is added to water, the hydrogen atoms of the water molecules are attracted to the negative bromide ions, and the oxygen atoms of the water molecules are attracted to the positive potassium ions.

Ion–dipole bonds form between the ions and water molecules and the surface ions are pulled into the solution. Gradually, the ionic lattice dissociates and a solution is formed.

32 **a** 50 g

b 10.5 g

c 25 g

33 **a** $\frac{200\,\text{g}}{100\,\text{g}} \times 500\,\text{g} = 1000\,\text{g}$

b $\frac{40\,\text{g}}{100\,\text{g}} \times 40\,\text{g} = 16\,\text{g}$

34 The maximum amount of potassium nitrate that will dissolve in 100 g water at 30°C is 50 g. Therefore the maximum amount that will dissolve in 200 g water is 100 g. There was initially 160 g of potassium nitrate dissolved, so the amount that will crystallise out is 160 – 100 = 60 g.

35 As the temperature increases, the solubility of most gases in water decreases. This means that less oxygen and other gases will be available to aquatic life.

36 0.5 mmol in 500 mL gives 1.0 mmol in 1 L. According to Figure 15.6.11, 23°C and 1.0 mmol L^{-1} corresponds to carbon monoxide, CO.

37 Table 15.6.2 shows that at 20°C, 1 kg of water can dissolve 529 g of ammonia.

The mass in 2 kg of water is $529\,\text{g} \times \frac{2\,\text{kg}}{1\,\text{kg}} = 1058\,\text{g}$.

The mass of ammonia that will leave the solution at 20°C is 1200 – 1058 = 142 g.

ISBN 978 1 4886 1772 0

38

a	does not dissolve in water because a large proportion of the molecule is non-polar	**v** compound composed of covalent molecules with a large non-polar end and one –OH group
b	dissolves in water by ionising, then forming ion–dipole bonds with water	**i** ionic compound
c	does not dissolve in water	**iv** non-polar molecular compound
d	dissolves in water by forming hydrogen bonds with water molecules	**ii** compound composed of polar molecules with –OH groups
e	dissolves in water by dissociating, then forming ion–dipole bonds with water	**iii** compound composed of small polar molecules in which a hydrogen atom is covalently bonded to an atom of a group 17 element

39 **a** Ammonia is a highly polar molecule and forms hydrogen bonds with water. It is therefore very soluble in water. Methane, however, is non-polar. Weak (dispersion) forces would occur between methane and water, and these are unable to disrupt the stronger hydrogen bonds between water molecules. Therefore, methane does not dissolve in water.

b Glucose dissolves in water because it has very polar –OH groups that can form hydrogen bonds with water molecules. Sodium chloride is ionic; hence, there are ion–dipole attractions between the ions and water. This is strong enough to overcome the attraction between the sodium ions and chloride ions in the solid NaCl lattice.

ISBN 978 1 4886 1772 0

Chapter 16 Aqueous solutions

Section 16.1 Precipitation reactions

Worked example: Try yourself 16.1.1

PREDICTING THE PRODUCTS OF A PRECIPITATION REACTION

<table>
<tr><td colspan="2">What precipitate, if any, will be produced when solutions of sodium sulfide (Na_2S) and copper(II) nitrate ($Cu(NO_3)_2$) are added together? You will need to refer to the solubility tables (Tables 15.5.1 and 15.5.2) to complete this question.</td></tr>
<tr><td>Thinking</td><td>Working</td></tr>
<tr><td>Identify which ions are produced by each of the ionic compounds in the mixture.</td><td>
<table>
<tr><th></th><th>Ions in solution</th><th>Possible precipitates</th><th>Actual precipitate, if any, from solubility tables</th></tr>
<tr><td>solution 1</td><td>Na^+ and S^{2-}</td><td></td><td></td></tr>
<tr><td>solution 2</td><td>Cu^{2+} and NO_3^-</td><td></td><td></td></tr>
</table>
</td></tr>
<tr><td>Identify which two new combinations of cations and anions are possible in the mixture of the solutions.</td><td>
<table>
<tr><th></th><th>Ions in solution</th><th>Possible precipitates</th><th>Actual precipitate, if any, from solubility tables</th></tr>
<tr><td>solution 1</td><td>Na^+ and S^{2-}</td><td>$NaNO_3$
or
CuS</td><td></td></tr>
<tr><td>solution 2</td><td>Cu^{2+} and NO_3^-</td><td></td><td></td></tr>
</table>
</td></tr>
<tr><td>Use the solubility tables to check which, if any, of these combinations will produce an insoluble compound.</td><td>
<table>
<tr><th></th><th>Ions in solution</th><th>Possible precipitates</th><th>Actual precipitate, if any, from solubility tables</th></tr>
<tr><td>solution 1</td><td>Na^+ and S^{2-}</td><td>$NaNO_3$
or
CuS</td><td>CuS</td></tr>
<tr><td>solution 2</td><td>Cu^{2+} and NO_3^-</td><td></td><td></td></tr>
</table>
Compounds containing sodium ions are usually soluble, so sodium nitrate will not form a precipitate.
Compounds containing sulfide ions are usually insoluble, so copper(II) sulfide will form a precipitate.
</td></tr>
</table>

ISBN 978 1 4886 1772 0

Worked example: Try yourself 16.1.2

WRITING EQUATIONS FOR PRECIPITATION REACTIONS

Write a balanced ionic equation for the reaction between copper(II) sulfate and sodium hydroxide, in which the precipitate is copper(II) hydroxide. Identify the spectator ions in this reaction.	
Thinking	**Working**
Identify the ions present in each solution.	Copper(II) sulfate solution has $Cu^{2+}(aq)$ and $SO_4^{2-}(aq)$ ions. Sodium hydroxide solution has $Na^+(aq)$ and $OH^-(aq)$ ions.
Write down the formula of the precipitate on the right-hand side. Include the state symbol (s). Place an arrow to the left of it.	$\rightarrow Cu(OH)_2(s)$
To the left of the arrow, add the formulae of the ions that form the precipitate, in the ratio shown by the formula of the precipitate. Include state symbols. Check the equation is balanced.	$Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$
Write the formulae of the ions that do not form a precipitate in the reaction. These are the spectator ions.	$Na^+(aq)$ and $SO_4^{2-}(aq)$ are spectator ions.

Worked example: Try yourself 16.1.3

IDENTIFYING PRECIPITATES AND WRITING EQUATIONS FOR PRECIPITATION REACTIONS

Identify the precipitate, if any, when solutions of potassium carbonate and copper(II) sulfate are mixed. Write the balanced ionic equation for the formation of any precipitate. Identify the spectator ions in any reaction.	
Thinking	**Working**
Identify the ions present in each solution.	Potassium carbonate solution has $K^+(aq)$ and $CO_3^{2-}(aq)$ ions. Copper(II) sulfate solution has $Cu^{2+}(aq)$ and $SO_4^{2-}(aq)$ ion.
Identify the possible products.	potassium sulfate and copper(II) carbonate
Use the solubility rules to determine if one of the products is insoluble.	Compounds containing potassium ions are usually soluble, so potassium sulfate will not form a precipitate. Compounds containing carbonate ions are usually insoluble, so copper(II) carbonate will form a precipitate.
Write down the formula of the precipitate on the right-hand side. Include the state symbol (s). Place an arrow to the left of it.	$\rightarrow CuCO_3(s)$
To the left of the arrow, add the formulae of the ions that form the precipitate in the ratio shown by the formula of the precipitate. Include state symbols. Check the equation is balanced.	$Cu^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CuCO_3(s)$
Write the formulae of the ions that do not form a precipitate in the reaction. These are the spectator ions.	$K^+(aq)$ and $SO_4^{2-}(aq)$ are spectator ions.

ISBN 978 1 4886 1772 0

16.1 KEY QUESTIONS

1 **a** a precipitate of silver carbonate

b a precipitate of lead(II) hydroxide

c a precipitate of magnesium sulfide

d no precipitate

Silver, lead and magnesium ions are not found in the solubility table. However, the anions that each of these are combined with in these questions all form compounds that are usually insoluble. It is also worth noting that compounds containing sodium and nitrate ions are usually soluble.

The iron(II) ion is also not found in the solubility table. However, iron(II) does not appear in the 'exceptions' column of the table for either nitrates or sulfates, both of which generally form soluble compounds.

2 Chemicals containing the ions Na^+, K^+, NH_4^+ and NO_3^- almost never form a precipitate.

a **i** magnesium sulfide

ii silver chloride

iii aluminium hydroxide

iv magnesium hydroxide

b **i** $Mg^{2+}(aq) + S^{2-}(aq) \rightarrow MgS(s)$

ii $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

iii $Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$

iv $Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$

3 **a** $Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$

b No precipitate is formed.

c $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

d $Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$

e $3Ca^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$

f $Pb^{2+}(aq) + 2OH^-(aq) \rightarrow Pb(OH)_2(s)$

4 **a** $Na^+(aq)$ and $NO_3^-(aq)$

b no spectator ions

c $NH_4^+(aq)$ and $Cl^-(aq)$

d $K^+(aq)$ and $NO_3^-(aq)$

e $Na^+(aq)$ and $Cl^-(aq)$

f $Na^+(aq)$ and $NO_3^-(aq)$

Section 16.2 Concentration of solutions

Worked example: Try yourself 16.2.1

CALCULATING CONCENTRATION IN $g\,L^{-1}$ (GRAMS PER LITRE)

Calculate the concentration, in $g\,L^{-1}$, of a solution containing 5.00 g of glucose in 250 mL of solution.	
Thinking	**Working**
Change the volume of solution so it is expressed in litres.	$250\text{ mL} = \frac{250}{1000}$ $= 0.250\text{ L}$
Calculate the concentration in $g\,L^{-1}$.	$c = \frac{\text{mass of solute (in g)}}{\text{volume of solution (in L)}}$ $= \frac{5.00}{0.250}$ $= 20.0\text{ g L}^{-1}$

Worked example: Try yourself 16.2.2

CALCULATING CONCENTRATION IN PARTS PER MILLION (PPM)

A sample of tap water was found to contain 0.0537 g of NaCl per 250.0 g of solution. Calculate the concentration of NaCl in parts per million (ppm).

Thinking	**Working**
Calculate the mass of solute in mg. Remember: mass (in mg) = mass (in g) × 1000	mass of solute (NaCl) in mg = 0.0537 × 1000 = 53.7 mg
Calculate the mass of solution in kg. Remember: mass (in kg) = $\frac{\text{mass (in g)}}{1000}$	mass of solution in kg = $\frac{250.0}{1000}$ = 0.2500 kg
Calculate the concentration of the solution in mg kg^{-1}. This is the same as concentration in ppm.	concentration of calcium carbonate in ppm = $\frac{\text{mass of solute (in mg)}}{\text{mass of solution (in kg)}}$ = $\frac{53.7}{0.2500}$ = 215 mg kg^{-1} = 215 ppm

16.2 KEY QUESTIONS

1 D. Volume in litres = $\frac{60\text{ mL}}{1000}$ = 0.060 L. Concentration of sugar = $\frac{5.0\text{ g}}{0.060}$ = 83 g L^{-1}

2 **a** $c(\text{CaCl}_2) = \frac{m(\text{CaCl}_2)\text{ in mg}}{m(\text{solution})\text{ in kg}} = \frac{25}{5.0} = 5.0\text{ ppm}$

b $c(\text{Pb(NO}_3)_2) = \frac{m(\text{Pb(NO}_3)_2)\text{ in mg}}{m(\text{solution})\text{ in kg}}$

$m(\text{Pb(NO}_3)_2) = 1.25\text{ g} \times 1000 = 1250\text{ mg}$

$m(\text{solution}) = \frac{2000\text{ g}}{1000} = 2.000\text{ kg}$

$c(\text{Pb(NO}_3)_2) = \frac{1250}{2.000} = 625\text{ ppm}$

c $c(\text{MgSO}_4) = \frac{m(\text{MgSO}_4)\text{ in mg}}{m(\text{solution})\text{ in kg}}$

$m(\text{MgSO}_4) = 4.0 \times 10^{-3}\text{ g} \times 1000 = 4.0\text{ mg}$

$m(\text{solution}) = \frac{150\text{ g}}{1000} = 0.150\text{ kg}$

$c(\text{MgSO}_4) = \frac{4.0}{0.150} = 27\text{ ppm}$

3 To find the % (m/v) the mass of each nutrient in grams is divided by the volume in millilitres and multiplied by 100.

sugar = $\frac{35.0}{250} \times 100 = 14.0\%$ (m/v)

fat = $\frac{7.5}{250} \times 100 = 3.0\%$ (m/v)

4 concentration (ppm) = $\frac{\text{mass of solute (in mg)}}{\text{mass of solution (in kg)}}$

$= \frac{1.096}{1.26}$

= 0.870 ppm

The concentration of mercury is less than the maximum allowable concentration so the fish can be sold.

ISBN 978 1 4886 1772 0

Section 16.3 Molar concentration

Worked example: Try yourself 16.3.1

CALCULATING MOLAR CONCENTRATIONS (MOLARITY)

Calculate the molar concentration of a solution that contains 0.24 mol of glucose dissolved in 500 mL of solution.

Thinking	Working
Convert the given volume to litres.	$V(\text{glucose}) = \frac{500}{1000}$ $= 0.500\text{ L}$
Calculate the molar concentration using the formula: $c = \frac{n}{V}$	$c(\text{glucose}) = \frac{n}{V}$ $= \frac{0.24}{0.500}$ $= 0.48\text{ mol L}^{-1}$

Worked example: Try yourself 16.3.2

CALCULATING MOLARITY GIVEN THE MASS OF SOLUTE

Calculate the concentration, in mol L^{-1}, of a solution that contains 4000 mg of ethanoic (acetic) acid (CH_3COOH) dissolved in 100 mL of solution.

Thinking	Working
Convert the volume to litres.	$V(CH_3COOH) = \frac{100}{1000}$ $= 0.100\text{ L}$
Convert the mass to grams.	$m(CH_3COOH) = \frac{4000}{1000}$ $= 4.000\text{ g}$
Calculate the molar mass of the solute. To do this, add up the atomic masses of all the atoms in the compound.	$M(CH_3COOH) = (2 \times 12.0) + (4 \times 1.008) + (2 \times 16.0)$ $= 60.032\text{ g mol}^{-1}$
Calculate the number of mol of solute using the formula: $n = \frac{m}{M}$	$n(CH_3COOH) = \frac{m}{M}$ $= \frac{4.000}{60.032}$ $= 6.66 \times 10^{-2}\text{ mol}$
Calculate the molar concentration using the formula: $c = \frac{n}{V}$	$c(CH_3COOH) = \frac{n}{V}$ $= \frac{6.66 \times 10^{-2}}{0.100}$ $= 6.66 \times 10^{-1}\text{ mol L}^{-1}$

Worked example: Try yourself 16.3.3

CALCULATING THE NUMBER OF MOLES OF SOLUTE IN A SOLUTION

Calculate the amount, in moles, of potassium permanganate ($KMnO_4$) in 100 mL of a 0.0250 mol L^{-1} solution of the compound.

Thinking	Working
Convert the given volume to litres.	$V(KMnO_4) = \frac{100}{1000}$ $= 0.100\text{ L}$
Calculate the amount of compound, in moles, using the formula: $n = c \times V$	$n(KMnO_4) = c \times V$ $= 0.0250 \times 0.100$ $= 2.50 \times 10^{-3}\text{ mol}$

ISBN 978 1 4886 1772 0

16.3 KEY QUESTIONS

1 B. Convert the mass of H_2O_2 to moles:

$M(H_2O_2) = (1.008 \times 2) + (16.00 \times 2) = 34.016\,g\,mol^{-1}$

$n(H_2O_2) = n(H_2O_2) = \frac{m(H_2O_2)}{M(H_2O_2)} = \frac{5.00}{34.016} = 0.147\,mol$

$V = \frac{120\,mL}{1000} = 0.120\,L$

$c = \frac{0.147}{0.120} = 1.23\,mol\,L^{-1}$

2 **a** $V = \frac{25\,mL}{1000} = 0.025\,L$

$c = \frac{2.0 \times 10^{-3}}{0.025} = 8.0 \times 10^{-2}\,mol\,L^{-1}$

b $c = \frac{1.23}{4.1} = 0.30\,mol\,L^{-1}$

c $c = \frac{1.8 \times 10^{3}}{9.3 \times 10^{3}} = 0.19\,mol\,L^{-1}$

3 **a** $c = \frac{2.0}{8.0} = 0.25\,mol\,L^{-1}$

b $V = \frac{500\,mL}{1000} = 0.5\,L$

$c = \frac{0.25}{0.5} = 0.50\,mol\,L^{-1}$

c $V = \frac{200\,mL}{1000} = 0.2\,L$

$c = \frac{0.0876}{0.2} = 0.438\,mol\,L^{-1}$

4 Convert the mass of $AgNO_3$ to moles:

$M(AgNO_3) = 107.9 + 14.01 + (3 \times 16.00) = 169.9\,g\,mol^{-1}$

$n(AgNO_3) = \frac{m(AgNO_3)}{M(AgNO_3)} = \frac{5.09}{169.9} = 0.0300\,mol$

$V = \frac{250\,mL}{1000} = 0.250\,L$

$c = \frac{0.0300}{0.250} = 0.120\,mol\,L^{-1}$

5 Convert the mass of $CaCl_2$ to moles:

$M(CaCl_2) = 40.08 + (2 \times 35.45) = 111.1\,g\,mol^{-1}$

$n(CaCl_2) = \frac{m(CaCl_2)}{M(CaCl_2)} = \frac{1.223}{111.1} = 1.101 \times 10^{-2}\,mol$

$c = \frac{1.101 \times 10^{-2}}{1.55} = 7.10 \times 10^{-3}\,mol\,L^{-1}$

ISBN 978 1 4886 1772 0

6 **a** $n = c \times V = 0.1 \times 0.22 = 2.2 \times 10^{-2}$ mol

b $V = \frac{10\,\text{mL}}{1000} = 0.01$ L

$n = c \times V = 0.64 \times 0.01 = 6.4 \times 10^{-3}$ mol

c $V = \frac{15.6\,\text{mL}}{1000} = 0.0156$ L

$n = c \times V = 0.0150 \times 0.0156 = 2.34 \times 10^{-4}$ mol

d $V = \frac{0.15\,\text{mL}}{1000} = 1.5 \times 10^{-4}$ L

$n = c \times V = 5.2 \times 1.5 \times 10^{-4} = 7.8 \times 10^{-4}$ mol

Section 16.4 Dilution

Worked example: Try yourself 16.4.1

QUESTIONS INVOLVING DILUTION

Calculate the concentration of the solution formed when 95.0 mL of water is added to 5.00 mL of 0.500 mol L^{-1} HCl.

Thinking	**Working**
Write down the value of c_1 and V_1. Note: c_1 and V_1 refer to the original solution, before water is added.	$c_1 = 0.500$ mol L^{-1} $V_1 = 5.00$ mL
Write down the value of V_2. Note: V_2 is the total volume of the original solution plus the added water.	$V_2 = 95.0 + 5.00$ $= 100.0$ mL
Transpose the equation $c_1V_1 = c_2V_2$ to allow calculation of the concentration, c_2, of the new solution.	$c_1V_1 = c_2V_2$ $c_2 = \frac{c_1V_1}{V_2}$
Calculate the concentration of the diluted solution.	$c_2 = \frac{0.500 \times 5.00}{100}$ $= 0.0250$ mol L^{-1}

Worked example: Try yourself 16.4.2

QUESTIONS INVOLVING DILUTION

Calculate the volume of 11.5 mol L^{-1} hydrochloric acid solution required to prepare 2.5 L of a 0.50 mol L^{-1} solution. What volume of water must be added to the 11.5 mol L^{-1} hydrochloric acid solution to give the required volume of the diluted solution?

Thinking	**Working**
Write down the value of c_1 and V_1. Note: c_1 and V_1 refer to the diluted solution, after water is added.	$c_1 = 0.50$ mol L^{-1} $V_1 = 2.5$ L
Write down the value of c_2. Note: c_2 is the concentration of the concentrated solution.	$c_2 = 11.5$ mol L^{-1}
Transpose the equation $c_1V_1 = c_2V_2$ to allow calculation of the volume, V_2, of the concentrated solution.	$c_1V_1 = c_2V_2$ $V_2 = \frac{c_1V_1}{c_2}$
Calculate the volume of the concentrated solution.	$V_2 = \frac{0.50 \times 2.5}{11.5}$ $= 0.109$ L
Calculate the volume of water required by finding the difference between the final volume of the diluted solution and the volume of concentrated solution.	$V(\text{water}) = 2.5 - 0.109$ $= 2.4$ L

ISBN 978 1 4886 1772 0

Worked example: Try yourself 16.4.3

CONCENTRATION UNIT CONVERSIONS

Calculate the concentration, in ppm, of a 0.0100 mol L^{-1} solution of NaOH. Note that concentration in ppm is the same as mg L^{-1}, if the density of the solution is 1 kg L^{-1}, which can be assumed for solutions of low concentrations.

Thinking	**Working**
Calculate the number of moles of solute in 1.00 L of the solution.	$n(\text{NaOH}) = c \times V$ $= 0.0100 \times 1.00$ $= 0.0100$ mol
Calculate the mass, in grams, of solute in 1.00 L of the solution.	$M(\text{NaOH}) = 23.0 + 16.0 + 1.008$ $= 40.008$ g mol^{-1} $m(\text{NaCl}) = n \times M$ $= 0.0100 \times 40.008$ $= 0.40008$ g
Calculate the mass, in mg, of solute in 1.00 L of the solution.	$m(\text{NaOH}) = 0.40008 \times 1000$ $= 400.08$ mg
Express the concentration of the solute in ppm.	$c(\text{NaOH}) = 400$ ppm

16.4 KEY QUESTIONS

1 **a** $V_2 = 10.0\text{ mL} + 5.0\text{ mL} = 15.0\text{ mL}$

$$c_2 = \frac{c_1V_1}{V_2} = \frac{1.2 \times 5.0}{15.0} = 0.4\text{ mol L}^{-1}$$

b $V_2 = 3.0\text{ L} + 1.0\text{ L} = 4.0\text{ L}$

$$c_2 = \frac{c_1V_1}{V_2} = \frac{0.10 \times 3.0}{4.0} = 0.075\text{ mol L}^{-1}$$

c $V_2 = 95.0\text{ mL} + 5.0\text{ mL} = 100.0\text{ mL}$

$$c_2 = \frac{c_1V_1}{V_2} = \frac{0.50 \times 5.0}{100.0} = 0.025\text{ mol L}^{-1}$$

2 D. $V_2 = \frac{250\text{ mL}}{1000} = 0.250\text{ L}$

$$V_1 = \frac{c_1V_1}{c_2} = \frac{0.30 \times 0.250}{10} = 0.0075\text{ L} = 7.5 \times 10^{-3}\text{ L}$$

3 mass of ammonia in 1 L of original solution = 15 g

$M(\text{NH}_3) = 14.01 + (1.008 \times 3) = 17.03\text{ g mol}^{-1}$

$$n(\text{NH}_3) = \frac{m(\text{NH}_3)}{M(\text{NH}_3)} = \frac{15}{17.03} = 0.881\text{ mol}$$

$$c_1 = \frac{n(\text{NH}_3)}{V(\text{NH}_3)} = \frac{0.881}{1} = 0.881\text{ mol L}^{-1}$$

$V_2 = 25.0 + 250 = 275\text{ mL}$

$$c_2 = \frac{c_1V_1}{V_2} = \frac{0.881 \times 25.0}{275.0} = 8.0 \times 10^{-2}\text{ mol L}^{-1}$$

4 In 1 L of solution there will be 170 g of NaOH.

$M(\text{NaOH}) = 22.99 + 16.00 + 1.008 = 40.0\text{ g mol}^{-1}$

$$n(\text{NaOH}) = \frac{m(\text{NaOH})}{M(\text{NaOH})} = \frac{170}{40.0} = 4.25\text{ mol}$$

$$c(\text{NaOH}) = \frac{n(\text{NaOH})}{V(\text{solution})} = \frac{4.25}{1.0} = 4.25\text{ mol L}^{-1}$$

ISBN 978 1 4886 1772 0

Section 16.5 Calculations involving reactions in solutions

Worked example: Try yourself 16.5.1

STOICHIOMETRY PROBLEMS INVOLVING SOLUTIONS

Calculate the mass of aluminium hydroxide precipitated when 45.0 mL of a 0.200 mol L^{-1} solution of sodium hydroxide reacts completely with a solution of aluminium chloride.	
Thinking	**Working**
Write a balanced equation for the reaction.	$3NaOH(aq) + AlCl_3(aq) \rightarrow Al(OH)_3(s) + 3NaCl(aq)$
Calculate the number of moles of the known substance using $n = c \times V$. The 'known substance' is the one you are provided information about in the question.	$V(NaOH) = 0.0450\,L$ $n(NaOH) = c \times V$ $= 0.200 \times 0.0450$ $= 0.009\,mol$
Write the mole ratio for the known and unknown substances.	mole ratio $= \frac{n(Al(OH)_3)}{n(NaOH)} = \frac{1}{3}$
Calculate the number of moles of the unknown substance using: $n(\text{unknown}) = n(\text{known}) \times$ mole ratio The 'unknown substance' is the one whose mass you are required to calculate.	$n(Al(OH)_3) = 0.009 \times \frac{1}{3} = 0.003$ mol
Calculate the mass of the unknown substance using: $m(\text{unknown}) = n(\text{unknown}) \times$ molar mass	$M(Al(OH)_3) = 78.0\,g\,mol^{-1}$ $m(Al(OH)_3) = 0.003 \times 78.0$ $= 0.234\,g$

Worked example: Try yourself 16.5.2

STOICHIOMETRY PROBLEMS INVOLVING SOLUTIONS

Calculate the concentration of a potassium sulfate solution if 7.88 g of barium sulfate is precipitated when 55.0 mL of the potassium sulfate solution is reacted with excess barium nitrate solution.	
Thinking	**Working**
Write a balanced equation for the reaction.	$K_2SO_4(aq) + Ba(NO_3)_2(aq) \rightarrow BaSO_4(s) + 2KNO_3(aq)$
Calculate the number of moles of the known substance using $n = \frac{m}{M}$. The 'known substance' is the one you are provided information about in the question.	$M(BaSO_4) = 233.4\,g\,mol^{-1}$ $n(BaSO_4) = \frac{m(BaSO_4)}{M(BaSO_4)} = \frac{7.88}{233.4} = 0.0338$ mol
Write the mole ratio for the known and unknown substances.	mole ratio $= \frac{n(K_2SO_4)}{n(BaSO_4)} = \frac{1}{1}$
Calculate the number of moles of the unknown substance using: $n(\text{unknown}) = n(\text{known}) \times$ mole ratio The 'unknown substance' is the one whose mass you are required to calculate.	$n(K_2SO_4) = 0.0338 \times \frac{1}{1} = 0.0338$ mol
Calculate the quantity (volume in this example) of the unknown substance by rearranging: $n = c \times V$	$c(K_2SO_4) = \frac{n}{V} = \frac{0.0338}{0.055} = 0.615$ mol L^{-1}

ISBN 978 1 4886 1772 0

16.5 KEY QUESTIONS

1 The precipitate produced is calcium sulfide and the balanced equation for the reaction is:

$CaCl_2(aq) + Na_2S(aq) \rightarrow CaS(s) + 2NaCl(aq)$

$n(CaCl_2) = c \times V = 0.075 \times 0.120 = 9.00 \times 10^{-3}$ mol

$n(CaS) = n(CaCl_2) = 9.00 \times 10^{-3}$ mol

$m(CaS) = n \times M = 9.00 \times 10^{-3} \times 72.2 = 0.65$ g

2 The balanced equation for the reaction is:

$Pb(NO_3)_2(aq) + 2NaI(aq) \rightarrow PbI_2(s) + 2NaNO_3(aq)$

$$n(PbI_2) = \frac{m}{M} = \frac{0.984}{461} = 2.13 \times 10^{-3}\text{ mol}$$

$n(Pb(NO_3)_2) = n(PbI_2) = 2.13 \times 10^{-3}$ mol

$$V(Pb(NO_3)_2) = \frac{n}{c} = \frac{2.13 \times 10^{-3}}{0.050} = 4.27 \times 10^{-2}\text{ L}$$

3 $$n(AgNO_3) = \frac{m}{M} = \frac{2.40}{169.9} = 1.41 \times 10^{-2}\text{ mol}$$

$n(AgCl) = n(AgNO_3) = 1.41 \times 10^{-2}$ mol

$m(AgCl) = n \times M = 1.41 \times 10^{-2} \times 143.4 = 2.03$ g

4 $n(Na_3PO_4) = c \times V = 0.155 \times 0.030 = 4.65 \times 10^{-3}$ mol

$$n(Ca(NO_3)_2) = 4.65 \times 10^{-3} \times \frac{3}{2} = 6.975 \times 10^{-3}\text{ mol}$$

$$V(Ca(NO_3)_2) = \frac{n}{c} = \frac{6.975 \times 10^{-3}}{0.125} = 5.58 \times 10^{-2}\text{ L}$$

CHAPTER 16 REVIEW

1 **a** true
b true
c false
d false
e true
f false

2

	NaOH	**KBr**	**NaI**	**$MgSO_4$**	**$BaCl_2$**
$Pb(NO_3)_2$	$Pb(OH)_2$	$PbBr_2$	PbI_2	$PbSO_4$	$PbCl_2$
KI					
$CaCl_2$	$Ca(OH)_2$			$CaSO_4$	
Na_2CO_3				$MgCO_3$	$BaCO_3$
Na_2S				MgS	BaS

3 **a** barium sulfate
b none
c lead(II) sulfate
d none

4 **a** $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
b $Fe^{2+}(aq) + S^{2-}(aq) \rightarrow FeS(s)$
c $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$
d $Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$
e $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

5 **a** $Cu^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CuCO_3(s)$
spectator ions: Na^+, SO_4^{2-}

b $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
spectator ions: K^+, NO_3^-

c $Pb^{2+}(aq) + S^{2-}(aq) \rightarrow PbS(s)$
spectator ions: Na^+, NO_3^-

d $Fe^{3+}(aq) + PO_4^{3-}(aq) \rightarrow FePO_4(s)$
spectator ions: Na^+, Cl^-

e $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$
spectator ions: K^+, SO_4^{2-}

6 **a** concentration of lead (ppm) $= \frac{m(\text{Pb}) \text{ in mg}}{m(\text{solution}) \text{ in kg}} = \frac{12}{6} = 2\,\text{ppm}$

b percentage by mass $= \frac{m(\text{Pb}) \text{ in mg}}{m(\text{solution}) \text{ in mg}} \times 100$

mass of solution = $6 \times 1\,000\,000 = 6\,000\,000\,\text{mg}$

$\%\ (\text{m/m}) = \frac{12}{6\,000\,000} \times 100 = 2.0 \times 10^{-4}\%$

7 mass of sodium hypochlorite $= \frac{0.75\,\text{L}}{1\,\text{L}} \times 42 = 31.5\,\text{g}$

8 $M(\text{NaOH}) = 23.0 + 16.0 + 1.0 = 40.0\,\text{g mol}^{-1}$

$m(\text{NaOH}) = \frac{30}{40.0} = 0.75\,\text{mol}$

$c(\text{NaOH}) = \frac{0.75}{2.0} = 0.375\ \text{mol L}^{-1} = 0.38\,\text{mol L}^{-1}$

9 **a** $V = \frac{12}{1000} = 0.012\,\text{L}$

$n = cV = 0.22 \times 0.012 = 2.6 \times 10^{-3}\,\text{mol}$

b $V = \frac{150}{1000} = 0.150\,\text{L}$

$n = cV = 0.0250 \times 0.150 = 3.75 \times 10^{-3}\,\text{mol}$

c $n = cV = 3.15 \times 10^{-3} \times 7.2 = 2.3 \times 10^{-2}\,\text{mol}$

10 **a** $V = \frac{100}{1000} = 0.100\,\text{L}$

$n(\text{NH}_3) = cV = 1.2 \times 0.100 = 0.120\,\text{mol}$
$M(\text{NH}_3) = 14.01 + (1.008 \times 3) = 17.0\,\text{g mol}^{-1}$
$m(\text{NH}_3) = n(\text{NH}_3) \times M(\text{NH}_3)$
$= 0.120 \times 17.0 = 2.04\,\text{g}$

b $V = \frac{20}{1000} = 0.020\,\text{L}$

$n(\text{AgNO}_3) = cV = 0.50 \times 0.020 = 0.010\,\text{mol}$
$M(\text{AgNO}_3) = 107.9 + 14.01 + (16.00 \times 3) = 169.9\,\text{g mol}^{-1}$
$m(\text{AgNO}_3) = n(\text{AgNO}_3) \times M(\text{AgNO}_3) = 0.010 \times 169.9 = 1.7\,\text{g}$

11 $V_2 = \frac{c_1 V_1}{c_2} = \frac{4.0 \times 25}{1.6} = 62.5\,\text{mL}$, so volume to be added = 62.5 – 25 = 37.5 mL

12 $n(\text{CaCl}_2)$ in concentrated solution $= c \times V = 0.0500 \times 0.0015 = 7.50 \times 10^{-5}\,\text{mol}$
$n(\text{Cl}^-) = 2 \times n(\text{CaCl}_2) = 2 \times 7.50 \times 10^{-5} = 1.50 \times 10^{-4}\,\text{mol}$
$\text{M}(\text{Cl}^-) = 35.5\,\text{g mol}^{-1}$
$m(\text{Cl}^-) = n \times M = 1.50 \times 10^{-4} \times 35.5 = 5.33 \times 10^{-3}\ \text{g} = 5.33\,\text{mg}$
mass of diluted solution = $1.0 \times 10.0 = 10.0\,\text{kg}$

concentration of diluted solution (ppm) $= \frac{m(\text{Cl}^-)\ (\text{in mg})}{\text{mass of solution (in kg)}} = \frac{5.33}{10.0} = 0.533\ \text{ppm}$

13 The precipitate produced is copper(II) carbonate and the balanced equation for the reaction is:

$CuCl_2(aq) + K_2CO_3(aq) \rightarrow CuCO_3(s) + 2KCl(aq)$

$n(CuCl_2) = c \times V = 0.550 \times 0.0256 = 1.408 \times 10^{-2}\,mol$

$n(CuCO_3) = n(CuCl_2) = 1.408 \times 10^{-2}\,mol$

$m(CuCO_3) = n \times M = 1.408 \times 10^{-2} \times 123.5 = 1.74\,g$

14 The precipitate produced is lead(II) iodide and the balanced equation for the reaction is:

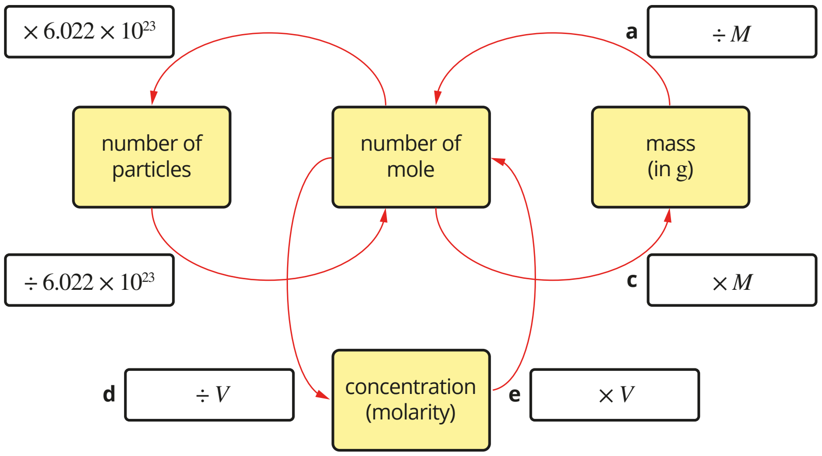
$Pb(NO_3)_2(aq) + 2NaI(aq) \rightarrow PbI_2(s) + 2NaNO_3(aq)$

$n(Pb^{2+}) = n(Pb(NO_3)_2) = c \times V = 0.085 \times 0.150 = 1.275 \times 10^{-2}\,mol$

$$n(NaI)\text{needed} = \frac{2}{1} \times 1.275 \times 10^{-2} = 2.55 \times 10^{-2}\,mol$$

$$V(NaI) = \frac{n(NaI)}{c(NaI)} = \frac{2.55 \times 10^{-2}}{0.752} = 3.39 \times 10^{-2}\,L = 34\,mL$$

15



16 a The precipitate produced is barium sulfate and the balanced equation for the reaction is:

$Na_2SO_4(aq) + Ba(NO_3)_2(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$

$n(BaSO_4) = n(Na_2SO_4) = c \times V = 0.150 \times 0.1000 = 0.015\,mol$

$M(BaSO_4) = 137.3 + 32.1 + (4 \times 16.0) = 233.4\,g\,mol^{-1}$

$m(BaSO_4) = n \times M = 0.0150 \times 233.4 = 3.50\,g$

b $V(\text{total}) = 100.0 + 25.0 = 125\,mL = 0.125\,L$

$n(Na^+) = 2 \times n(Na_2SO_4) = 2 \times c \times V = 2 \times 0.150 \times 0.1000 = 0.030\,mol$

$$c(Na^+) = \frac{n(Na^+)}{V(\text{total})} = \frac{0.030}{0.125} = 0.24\ mol\ L^{-1}$$

Chapter 17 Acids and bases

Section 17.1 Properties of acids and bases

17.1 KEY QUESTIONS

1 A. Hydrochloric acid is a strong acid and corrosive so is not found in drinks. The others are all weak organic acids and are therefore potentially found in nature and drinks.

2 B. Alkaline solutions are those that feel soapy and slippery to touch. A sour taste is a property related to acids.

3 Litmus just tells you whether the solution is acidic or basic. Universal indicator will give an indication of the level of acidity.

4 Although CuO is a base, it is insoluble, so will not affect the pH of the water, i.e. copper oxide is a base, but not an alkali.

Section 17.2 The Arrhenius theory of acids and bases

Worked example: Try yourself 17.2.1

CALCULATING THE pH OF AN AQUEOUS SOLUTION FROM $[H^+]$

Calculate the pH of a solution in which the concentration of H^+ is 6.0×10^{-9} mol L^{-1}. Express your answer to 2 significant figures.

Thinking	**Working**
Write down the concentration of H^+ ions in the solution.	$[H^+] = 6.0 \times 10^{-9}$ mol L^{-1}
Substitute the value of $[H^+]$ into: $pH = -\log_{10}[H^+]$ Use the logarithm function on your calculator to determine the answer.	$pH = -\log_{10}[H^+]$ $= -\log_{10}[6.0 \times 10^{-9}]$ (use your calculator) $= 8.2$

17.2 KEY QUESTIONS

1 $HBr(g) \rightarrow H^+(aq) + Br^-(aq)$

2 $H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$

3 C. Arrhenius bases dissociate to produce hydroxide ions in solution. Ionic compounds dissociate and ionise in water, therefore C is the most correct answer.

4 Stronger acids more readily ionise, forming ions in solution. As perchloric acid is stronger, more hydronium ions would be present in the solution than in a solution of ethanoic acid, making it a better conductor of electricity.

5 $pH = -\log_{10}[H^+]$
$= -\log_{10}0.01$
$= 2$

6 As nitric acid is a strong acid, the concentration of H_3O^+ ions is 0.001 mol L^{-1}.
$pH = -\log_{10}[H^+]$
$= -\log_{10}0.001$
$= 3$

ISBN 978 1 4886 1772 0

Section 17.3 Reactions of acids and bases

Worked example: Try yourself 17.3.1

WRITING AN IONIC EQUATION FOR AN ACID–BASE REACTION

Write an ionic equation for the reaction that occurs when sulfuric acid is added to potassium hydroxide solution.

Thinking	Working
What is the general reaction? Identify the products formed.	acid + metal hydroxide → salt + water A solution of potassium sulfate and water is formed.
Identify the reactants and products. Indicate the state of each, i.e. (aq), (s), (l) or (g).	Reactants: H_2SO_4 is ionised in solution, forming $H^+(aq)$ and $SO_4^{2-}(aq)$. KOH is dissociated in solution, forming $K^+(aq)$ and $OH^-(aq)$. Products: Potassium sulfate is dissociated and exists as $K^+(aq)$ and $SO_4^{2-}(aq)$. Water is a molecular compound and its formula is $H_2O(l)$.
Write the equation showing all reactants and products. (There is no need to balance the equation yet.)	$H^+(aq) + SO_4^{2-}(aq) + K^+(aq) + OH^-(aq) \rightarrow K^+(aq) + SO_4^{2-}(aq) + H_2O(l)$
Identify the spectator ions: the ions that have an (aq) state both as a reactant and as a product.	$K^+(aq)$ and $SO_4^{2-}(aq)$
Rewrite equation without the spectator ions. Balance the equation with respect to number of atoms of each element and charge.	$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

Worked example: Try yourself 17.3.2

WRITING IONIC EQUATIONS FOR REACTIONS BETWEEN ACIDS AND METAL CARBONATES

What products are formed when a solution of hydrochloric acid is added to a solution of sodium hydrogencarbonate? Write an ionic equation for this reaction.

Thinking	Working
What is the general reaction? Identify the products.	acid + metal hydrogencarbonate → salt + water + carbon dioxide Products of this reaction are sodium chloride in solution, water and carbon dioxide gas.
Identify the reactants and products. Indicate the state of each, i.e. (aq), (s), (l) or (g).	Reactants: Hydrochloric acid is ionised in solution, forming $H^+(aq)$ and $Cl^-(aq)$ ions. Sodium hydrogencarbonate dissociates in solution and exists as $Na^+(aq)$ and $HCO_3^-(aq)$ ions. Products: Sodium chloride is dissociated into $Na^+(aq)$ and $Cl^-(aq)$ ions. Water has the formula $H_2O(l)$. Carbon dioxide has the formula $CO_2(g)$.
Write the equation showing all reactants and products. (There is no need to balance the equation yet.)	$H^+(aq) + Cl^-(aq) + Na^+(aq) + HCO_3^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l) + CO_2(g)$
Identify the spectator ions.	$Na^+(aq) + Cl^-(aq)$
Rewrite the equation without the spectator ions. Balance the equation with respect to number of atoms of each element and charge.	$H^+(aq) + HCO_3^-(aq) \rightarrow H_2O(l) + CO_2(g)$ Note that if hydronium ions are represented as $H_3O^+(aq)$, rather than as $H^+(aq)$, this reaction would be written as: $H_3O^+(aq) + HCO_3^-(aq) \rightarrow 2H_2O(l) + CO_2(g)$

ISBN 978 1 4886 1772 0

Worked example: Try yourself 17.3.3

WRITING IONIC EQUATIONS FOR REACTIONS BETWEEN ACIDS AND REACTIVE METALS

Write an ionic equation for the reaction that occurs when aluminium is added to a dilute solution of hydrochloric acid.	
Thinking	**Working**
What is the general reaction? Identify the products formed.	acid + reactive metal → salt + hydrogen Hydrogen gas and aluminium chloride solution are produced.
Identify the reactants and products. Indicate the state of each, i.e. (aq), (s), (l), (g).	Reactants: Aluminium is a solid, Al(s). Hydrochloric acid is ionised, forming $H^+(aq)$ and $Cl^-(aq)$ ions. Products: Hydrogen gas, H_2, forms. Aluminium chloride is dissociated into $Al^{3+}(aq)$ and $Cl^-(aq)$ ions.
Write the equation showing all reactants and products. (There is no need to balance the equation yet.)	$H^+(aq) + Cl^-(aq) + Al(s) \rightarrow Al^{3+}(aq) + Cl^-(aq) + H_2(g)$
Identify the spectator ions.	$Cl^-(aq)$
Rewrite the equation without the spectator ions. Balance the equation with respect to the number of atoms of each element and charge.	$6H^+(aq) + 2Al(s) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

17.3 KEY QUESTIONS

1 a $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$
  $Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$
 b $Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$
  $Ca(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2(g)$
 c $Zn(s) + 2CH_3COOH(aq) \rightarrow Zn(CH_3COO)_2(aq) + H_2(g)$
  $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$
 d $2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$
  $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

2 a magnesium sulfate
 b calcium chloride
 c zinc acetate or ethanoate
 d aluminium chloride

3 a i $ZnO(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2O(l)$
  ii $ZnO(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2O(l)$
 b i $Ca(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2(g)$
  ii $Ca(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2(g)$
 c i $Cu(OH)_2(s) + 2HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2H_2O(l)$
  ii $Cu(OH)_2(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + 2H_2O(l)$
 d i $Mg(HCO_3)_2(s) + 2HCl(aq) \rightarrow MgCl_2(s) + 2H_2O(l) + 2CO_2(g)$
  ii $Mg(HCO_3)_2(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2O(l)$
 e i $SnCO_3(s) + H_2SO_4(aq) \rightarrow SnSO_4(s) + H_2O(l) + CO_2(g)$
  ii $SnCO_3(s) + 2H^+(aq) \rightarrow Sn^{2+}(s) + H_2O(l) + CO_2(g)$

ISBN 978 1 4886 1772 0

4 The first step, as in the development of any equation, is to write the correct chemical formulae for each of the chemicals involved.

- **a** $2KOH(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$
  $OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$
- **b** $NaOH(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l)$
  $OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$
- **c** $Mg(OH)_2(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$
  $Mg(OH)_2(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + 2H_2O(l)$
- **d** $CuCO_3(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l) + CO_2(g)$
  $CuCO_3(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l) + CO_2(g)$
- **e** $KHCO_3(aq) + HCl(aq) \rightarrow KCl(aq) + H_2O(l) + CO_2(g)$
  $HCO_3^-(aq) + H^+(aq) \rightarrow H_2O(l) + CO_2(g)$
- **f** $Zn(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2(g)$
  $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$
- **g** $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$
  $CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$
- **h** $NaHCO_3(s) + CH_3COOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l) + CO_2(g)$
  $NaHCO_3(s) + H^+(aq) \rightarrow Na^+(aq) + H_2O(l) + CO_2(g)$
- **i** $(NH_4)_2CO_3(aq) + Ca(OH)_2(aq) \rightarrow 2NH_3(g) + CaCO_3(s) + 2H_2O(l)$
  $2NH_4^+(aq) + Ca^{2+}(aq) + CO_3^{2-} + 2OH^-(aq) \rightarrow CaCO_3(s) + 2NH_3(g) + 2H_2O(l)$

Section 17.4 Calculations involving acids and bases

Worked example: Try yourself 17.4.1

CALCULATING MOLAR CONCENTRATION AFTER DILUTION

Calculate the molar concentration of nitric acid when 80.0 mL of water is added to 20.0 mL of 5.00 mol L^{-1} HNO_3.	
Thinking	**Working**
The number of moles of solute does not change during a dilution. So $c_1V_1 = c_2V_2$, where c is the concentration in mol L^{-1} and V is the volume of the solution. (Each of the volume units must be the same, although not necessarily litres.)	$c_1V_1 = c_2V_2$
Identify given values for concentrations and volumes before and after dilution. Identify the unknown.	Remember that 80.0 mL was added to 20.0 mL, so the final volume is 100.0 mL. $c_1 = 5.00$ mol L^{-1} $V_1 = 20.0$ mL $V_2 = 100.0$ mL You are required to calculate c_2, the concentration after dilution.
Transpose the equation and substitute the known values into the equation to find the required value.	$c_2 = \frac{c_1 \times V_1}{V_2}$ $= \frac{5.00 \times 20.0}{100.0}$ $= 1.00$ mol L^{-1}

ISBN 978 1 4886 1772 0

Worked example: Try yourself 17.4.2

CALCULATING THE VOLUME OF WATER TO BE ADDED IN A DILUTION

How much water must be added to 15.0 mL of 10.0 mol L^{-1} NaOH to dilute the solution to 2.00 mol L^{-1}?	
Thinking	**Working**
The number of moles of solute does not change during a dilution. So $c_1V_1 = c_2V_2$, where c is the concentration in mol L^{-1} and V is the volume of the solution. (Each of the volume units must be the same, although not necessarily litres.)	$c_1V_1 = c_2V_2$
Identify given values for concentrations and volumes before and after dilution. Identify the unknown.	$c_1 = 10.0$ mol L^{-1} $V_1 = 15.0$ mL $c_2 = 2.00$ mol L^{-1} You are required to calculate V_2, the volume of the diluted solution.
Transpose the equation and substitute the known values into the equation to find the required value.	$V_2 = \frac{c_1 \times V_1}{c_2}$ $= \frac{10.0 \times 15.0}{2.00}$ $= 75.0$ mL
Calculate the volume of water to be added.	volume of dilute solution = 75.0 mL initial volume of NaOH = 15.0 mL So 75.0 – 15.0 = 60.0 mL of water must be added.

Worked example: Try yourself 17.4.3

CALCULATING THE pH OF A DILUTED ACID

If 10.0 mL of 0.1 mol L^{-1} HCl is diluted to 30.0 mL, what is the pH of the diluted solution?	
Thinking	**Working**
Identify given values for concentrations and volumes before and after dilution.	$c_1 = 0.1$ mol L^{-1} $V_1 = 10.0$ mL $V_2 = 30.0$ mL $c_2 = ?$
Calculate c_2, which is the concentration of H$^+$ after dilution, by transposing the formula: $c_1V_1 = c_2V_2$	$c_2 = \frac{c_1 \times V_1}{V_2}$ $= \frac{0.1 \times 10.0}{30.0}$ $= 0.03$ mol L^{-1}
Calculate pH using: $\text{pH} = -\log_{10}[\text{H}^+]$ Use the logarithm function on your calculator to calculate pH.	$\text{pH} = -\log_{10}[\text{H}^+]$ $= -\log_{10}(0.03)$ $= 1.5$

ISBN 978 1 4886 1772 0

Worked example: Try yourself 17.4.4

DETERMINING THE VOLUME OF AN UNKNOWN SOLUTION FROM A NEUTRALISATION REACTION

Calculate the volume of 0.500 mol L^{-1} hydrochloric acid (HCl) that reacts completely with 25.0 mL of 0.100 mol L^{-1} calcium hydroxide ($Ca(OH)_2$) solution.

Thinking	**Working**
Write a balanced chemical equation.	$Ca(OH)_2(aq) + 2HCl(aq) \rightarrow CaCl_2(aq) + 2H_2O(l)$
Calculate the amount, in mol, of the solution of known concentration and volume.	$n(Ca(OH)_2) = cV = 0.100 \times 0.0250$ $= 0.00250$ mol
Determine the mole ratio as given by the balanced chemical equation.	mole ratio $= \frac{n(\text{unknown})}{n(\text{known})} = \frac{2}{1}$
Determine the amount, in mol, of the unknown solution.	$n(HCl) = 0.00250 \times \frac{2}{1} = 0.00500$ mol
Determine the volume of the unknown solution needed to react.	$V(HCl) = \frac{n}{c} = \frac{0.00500}{0.500} = 0.0100$ L $= 10.0$ mL Therefore, 10.0 mL of HCl is required to react with 25.0 mL of 0.100 mol L^{-1} $Ca(OH)_2$.

Worked example: Try yourself 17.4.5

A LIMITING REAGENT ACID–BASE PROBLEM

A sample of 30.0 mL of a 0.100 mol L^{-1} H_2SO_4 solution is mixed with 40.0 mL of a 0.200 mol L^{-1} KOH solution.

a Which reactant is the limiting reagent?

b What will be the mass of the K_2SO_4 produced by this reaction?

Thinking	**Working**
Write a balanced chemical equation.	$H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$
Calculate the amount, in mol, of the first reactant.	$n(H_2SO_4) = cV = 0.100 \times 0.0300 = 0.00300$ mol
Calculate the amount, in mol, of the second reactant.	$n(KOH) = cV = 0.200 \times 0.0400 = 0.00800$ mol
Use the coefficients of the reaction to determine the limiting reagent.	The equation shows that 1 mole of H_2SO_4 reacts with 2 mole of KOH. So the H_2SO_4 is the limiting reagent.
Find the mole ratio of the unknown substance to the limiting reactant from the equation coefficients.	mole ratio $= \frac{n(\text{unknown})}{n(\text{known})} = \frac{1}{1}$
Determine the amount, in mol, of the unknown substance.	$n(K_2SO_4) = 0.00300$ mol
Determine the mass, in g, of the unknown substance.	$m(K_2SO_4) = nM = 0.00300 \times 174.27 = 0.5228$ g $= 0.523$ g

17.4 KEY QUESTIONS

1 $c_1V_1 = c_2V_2$, $c_1 = 0.01$ mol L^{-1}, $V_1 = 3.0$ L, $V_2 = 4.0$ L, $c_2 = ?$

$0.01 \times 3.0 = c_2 \times 4.0$

$c_2 = \frac{0.10 \times 3.0}{4.0} = 0.075$ mol L^{-1}

2 $c_1V_1 = c_2V_2$, $c_1 = 2.0$ mol L^{-1}, $V_1 = 10$ mL, $V_2 = ?$, $c_2 = 0.50$ mol L^{-1}

$2.0 \times 10 = 0.50 \times V_2$

$V_2 = \frac{2.0 \times 10}{0.50} = 40$ mL

Final volume − initial volume = 40 − 10 = 30 mL, so 30 mL of water needs to be added.

3 $c_1V_1 = c_2V_2$, $c_1 = 0.60\,\text{mol L}^{-1}$, $V_1 = 20\,\text{mL}$, $V_2 = ?$, $c_2 = 0.10\,\text{mol L}^{-1}$

$0.60 \times 20 = 0.10 \times V_2$

$V_2 = \frac{0.60 \times 20}{0.10} = 120\,\text{mL}$

Final volume – initial volume = 120 – 20 = 100 mL, so 100 mL of water needs to be added.

4 This increases the pH by one unit.

5 **a** $HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)$

b $n(KOH) = cV = 0.09927 \times 0.02000 = 0.001985\,\text{mol}$

c $n(HNO_3) = n(KOH) = 0.001985\,\text{mol}$

d $c(HNO_3) = \frac{n}{V} = \frac{0.001985}{0.01826} = 0.1087\,\text{mol L}^{-1}$

6 **a** $n(H_2SO_4) = cV = 0.200 \times 0.0100 = 0.00200\,\text{mol}$

b $n(Na_2CO_3) = cV = 0.100 \times 0.0160 = 0.00160\,\text{mol}$

c According to the reaction, $n(H_2SO_4) = n(Na_2CO_3)$, therefore Na_2CO_3 is the limiting reactant.

CHAPTER 17 **REVIEW**

1 Dip pieces of the red litmus paper into the two solutions. The paper will remain red in the acidic solution and turn blue in the alkaline solution.

2 **a** Sulfuric acid (H_2SO_4) is a diprotic acid because each molecule can produce two hydrogen ions when it ionises:

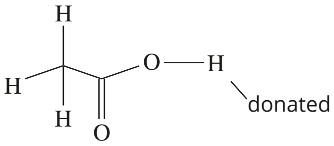
$H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$

Hydrochloric acid (HCl) is a monoprotic acid as each molecule can only produce one hydrogen ion when the molecule ionises:

$$HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$$

b A strong acid is one that ionises completely in solution (e.g. HCl). A concentrated acid is one in which there is a large amount of acid dissolved in a given volume of solution; for example, 5 mol L^{-1} HCl and 5 mol L^{-1} CH_3COOH are concentrated acids.

3



4 A scientific model is a representation that can be used to describe and explain phenomena or observations. It is often used when the phenomena cannot be experienced directly (first hand).

5 Strong acids are completely ionised; weak acids are partially ionised. Lower levels of ionisation mean there are fewer charged hydrogen ions in solution.

6 $HClO_4(aq) \rightarrow ClO_4^-\,(aq) + H^+(aq)$

7 $HClO_3(aq) \rightleftharpoons ClO_3^-\,(aq) + H^+(aq)$

8 $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+\,(aq) + OH^-(aq)$

9 A solution with a pH of 7.4 is basic.

10 Remember: $pH = -\log_{10}[H^+]$

pH = 3

$\therefore [H^+] = 10^{-3}\,\text{mol L}^{-1}$

pH = 5

$\therefore [H^+] = 10^{-5}\,\text{mol L}^{-1}$

∴ the difference is a factor of 100.

11 **a** 2

b 0.7

c –0.6

12 **a** $HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)$

b $H_2SO_4(aq) + K_2CO_3(aq) \rightarrow K_2SO_4(aq) + H_2O(l) + CO_2(g)$

c $2H_3PO_4(aq) + 3Ca(HCO_3)_2(s) \rightarrow Ca_3(PO_4)_2(s) + 6H_2O(l) + 6CO_2(g)$

d $2HF(aq) + Zn(OH)_2(s) \rightarrow ZnF_2(aq) + 2H_2O(l)$

13 E. An acid reacts with a base to produce a salt plus water. In this case, the salt produced is magnesium chloride, which is soluble in water and exists as magnesium ions and chloride ions.

14 $2H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)$

15 a carbonate or hydrogencarbonate

16 **a** $2Al(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$

b $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

17 Remember to use $pH = -\log_{10}[H^+]$.

a $pH = 2$, $\therefore$ $[H^+] = 10^{-2}$ or $0.01\ \text{mol}\,L^{-1}$

b The formula for amounts of substance in solution is $n = c \times V$ (where n is the amount in moles, c the concentration in $\text{mol}\,L^{-1}$ and V the volume of the solution in litres).

$n = 0.01 \times 0.500 = 0.005\ \text{mol}$

18 These three questions have two steps.

Step 1. Dilution questions are best answered using the formula $c_1V_1 = c_2V_2$, where c is the concentration in $\text{mol}\,L^{-1}$ and V is the volume of the solution. Each of the volume units needs to be the same, although not necessarily litres.

Step 2. Also remember, $pH = -\log_{10}[H^+]$

a $c_2 = 0.25 \times \frac{10}{50} = 0.05\ \text{mol}\,L^{-1}$

$\therefore\ [H^+] = 0.05\ \text{mol}\,L^{-1}$

$pH = -\log_{10}(0.050\ \text{mol}\,L^{-1}) = 1.3$

b $c_2 = 0.15 \times \frac{10}{1500}$ (1.5 L has been converted to mL to maintain identical units)

$= 0.001\ \text{mol}\,L^{-1}$

$\therefore\ [H^+] = 0.001\ \text{mol}\,L^{-1} = 10^{-3}\ \text{mol}\,L^{-1}$

$\therefore\ pH = 3$

19 $c_1V_1 = c_2V_2$, $c_1 = 18.0\ \text{mol}\,L^{-1}$, $V_1 = ?$, $c_2 = 2.00\ \text{mol}\,L^{-1}$, $V_2 = 1.00\ L$

$V_1 = \frac{2.00 \times 1.00}{18.0} = 0.111\ L = 111\ mL$

20 $pH(\text{initial}) = -\log_{10}[H^+] = 1.0$

$c_1V_1 = c_2V_2$, $c_1 = 0.10\ \text{mol}\,L^{-1}$, $V_1 = 40.0\ mL$, $c_2 = ?$, $V_2 = 500\ mL$

$c_2 = \frac{0.10 \times 40.0}{500} = 8 \times 10^{-3}\ \text{mol}\,L^{-1}$

$pH(\text{final}) = -\log_{10}(8 \times 10^{-3}) = 2.10$

Therefore the pH will increase.

21 solution A: weaker base, few freely moving charged particles—ammonia

solution B: neutral, no freely moving charged particles—glucose

solution C: strong base, many freely moving charged particles—sodium hydroxide

solution D: strong acid, many freely moving charged particles—hydrochloric acid

solution E: weaker acid, few freely moving charged particles—ethanoic acid

22 **a** An Arrhenius base is a substance that dissociates in water to form hydroxide (OH^-) ions.

b A strong acid is a substance that completely ionises (in water) to form hydrogen ions (H^+) ions.

c Molarity is a measure of concentration expressed in moles per litre of solution ($\text{mol}\,L^{-1}$).

ISBN 978 1 4886 1772 0

Chapter 18 Rates of chemical reactions

Section 18.1 Investigating the rate of chemical reactions

18.1 KEY QUESTIONS

1 C

2 **a** increasing
  b increasing
  c increasing
  d increasing
  e adding

3 **a** $CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + CO_2(g) + H_2O(l)$
  b Because the gradient of the graph is decreasing, the reaction rate must be decreasing over time.

4 **a** using smaller pieces of wood with a larger surface area
  b using a brick cleaner with a higher concentration
  c increasing the temperature of the oven

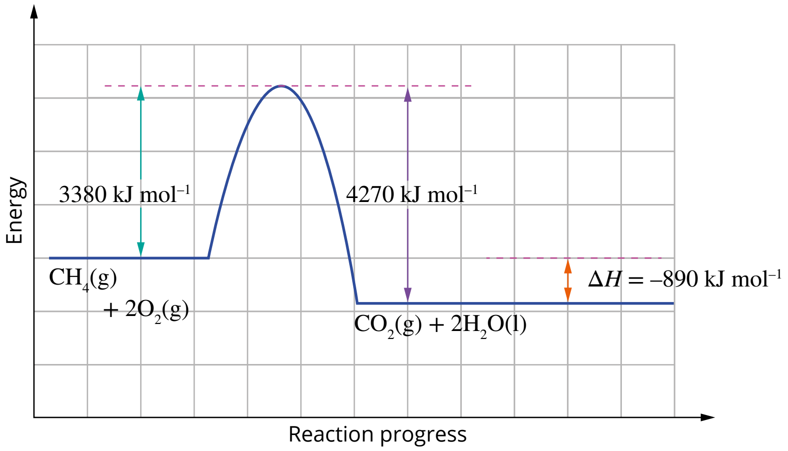
Section 18.2 Collision theory

18.2 KEY QUESTIONS

1 D. D is incorrect because reactant particles need to collide with energy equal to or greater than the activation energy for the reaction.

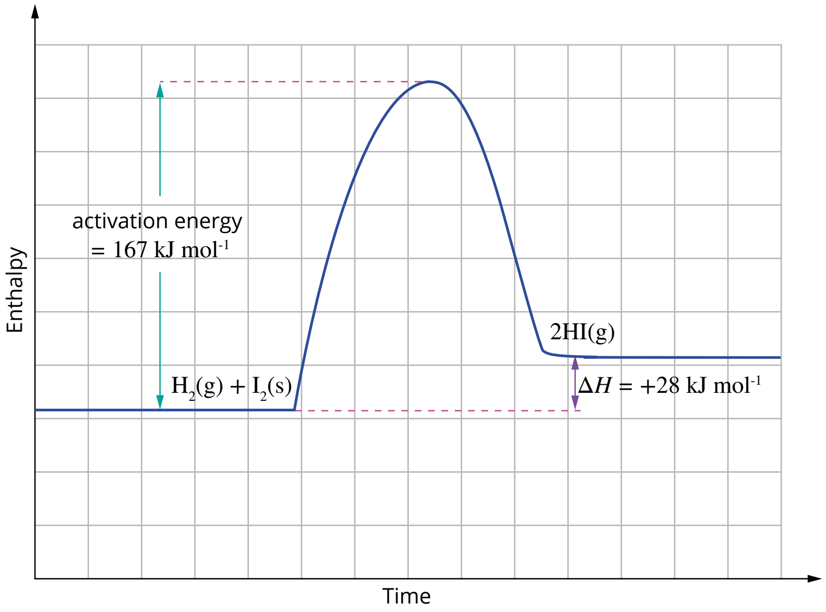
2 A

3 **a** $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$
  b $-890\,kJ\,mol^{-1}$
  c



ISBN 978 1 4886 1772 0

4 **a**



b Endothermic. Δ*H* is positive therefore the enthalpy of the products is greater than the enthalpy of the reactants.

c $-28\,\text{kJ mol}^{-1}$

d $139\,\text{kJ mol}^{-1}$

Section 18.3 Applying collision theory

Worked example: Try yourself 18.3.1

USING COLLISION THEORY TO EXPLAIN CHANGES IN RATES OF REACTIONS

Iron anchors recovered from shipwrecks at considerable depths can show little corrosion after years in the sea, whereas anchors recovered from shallow water are badly corroded. Explain this observation in terms of collision theory.

Thinking	**Working**
Premise: What is the key knowledge, law or theory that you are using?	The rate of reaction is dependent on the number of successful collisions between reacting particles in a certain time.
Reasoning: Consider the state of the reactants. Relate the state of the reactant to the factor that affects the reaction rate and explain in terms of collision theory.	The iron anchor is a solid. The oxygen is a gas dissolved in the seawater. The surface area of the iron anchor is relatively small. The concentration of oxygen at greater depths in the ocean is much lower. The frequency of collisions with reacting particles that could cause corrosion would be low so the rate of corrosion would also be low.
Outcome: Return to the question to complete your answer.	The surface area of the iron anchor is relatively small so the frequency of collisions with reacting particles would be low. The concentration of oxygen at greater depths is also low so the frequency of collisions is further reduced. Therefore, the rate of corrosion is reduced.

ISBN 978 1 4886 1772 0

18.3 KEY QUESTIONS

1 D. An increase in concentration increases collision frequency between reactant particles.

2 A, B and C. A is true because a reaction with a large activation energy would have a low proportion of reactant particles colliding with energy equal to or greater than the activation energy for the reaction.

B is true because a low temperature means a low proportion of reactant particles collide with energy equal to or greater than the activation energy for the reaction.

C is true because for a reaction to occur, reactant particles need to collide with the correct orientation.

3 D, F, B, E, C and A.

D is slowest because it has the lowest surface area, lowest concentration and lowest temperature.

F is faster (than D) because it has a larger surface area.

B is faster (than F) because it has a higher temperature.

E is faster (than B) because it has a larger surface area.

C is faster (than E) because it has a higher temperature.

A is the fastest because it has the largest surface area, highest concentration and highest temperature.

4 **a** At higher temperatures, the molecules that react to form fibreglass plastics have greater energy. They collide more frequently and are more likely to have a total energy exceeding the activation energy of the reaction involved, increasing the rate of reaction.

b Fine particles have a large surface area, resulting in a high frequency of collisions of aluminium particles with gas molecules (such as oxygen) in the air and hence rapid reaction rate. The aluminium can burn vigorously and release a large quantity of heat.

c At high altitude, such as in Nepal, air pressure is considerably lower than at any location in the Australian bush and so the water boils at a lower temperature in Nepal (up to 30°C lower). Thus, the average kinetic energy of the molecules in the potato is lower, so they are less likely to have a total energy exceeding the activation energy of the reactions involved in cooking a potato, so the potato cooks slower.

5 **a** A is activation energy, B is ΔH, C is total enthalpy of reaction.

b A catalyst would lower the value of A and C, while the value of B would remain unchanged.

CHAPTER 18 REVIEW

1 D

2 C. Powder has a higher surface area than chips, and 2 mol L^{-1} is a higher concentration than 1 mol L^{-1}.

3 B

4 B. The rate of reaction is a measure of the change in concentration of a substance over time. Only B has units that correspond to concentration (mol L^{-1}) per unit time (s^{-1}).

5 **a** A gas is produced so mass is lost from the mixture.

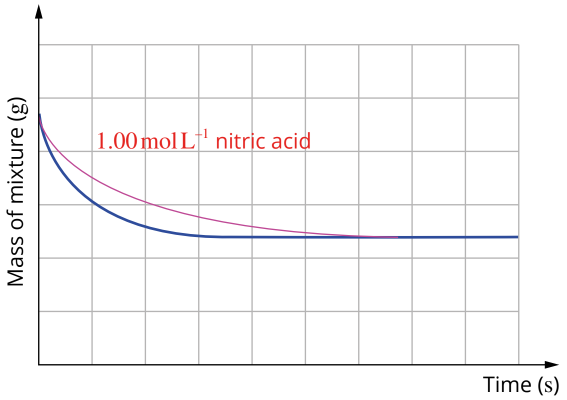
b
$$\begin{aligned} n(\text{Cu}) &= \frac{m}{M} \\ &= \frac{5.00}{63.5} \\ &= 0.0787\ \text{mol} \\ n(\text{HNO}_3) &= cV \\ &= 2.00 \times 0.500 = 1.00\ \text{mol} \end{aligned}$$

$$\begin{aligned} n(\text{HNO}_3)\ \text{react with Cu} &= \frac{8}{3} \times n(\text{Cu}) \\ &= \frac{8}{3} \times 0.0787 \\ &= 0.210\ \text{mol} \end{aligned}$$

0.210 < 1.00; therefore HNO_3 is in excess. So Cu is limiting.

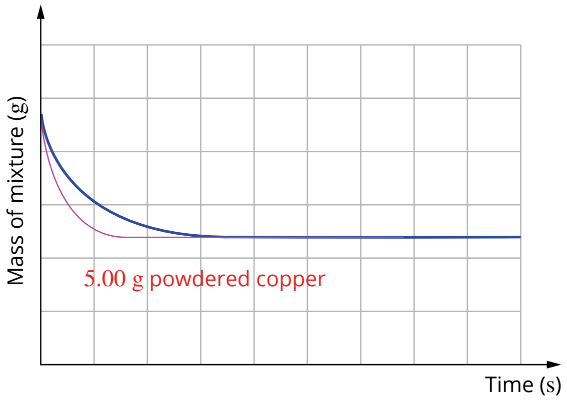
ISBN 978 1 4886 1772 0

c



Decreased rate of mass loss is due to lower nitric acid concentration.
($n(HNO_3) = cV = 1.00 \times 0.500 = 0.500$ mol)
Copper is still limiting, so final mass remains the same.

d

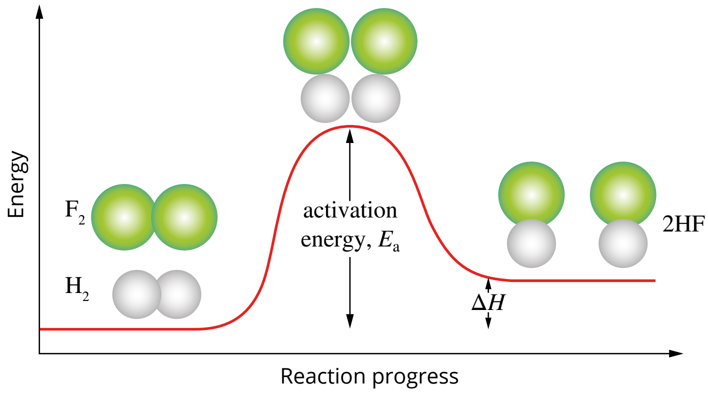


Increased rate of mass loss is due to increased copper surface area.
Copper is still limiting, so final mass remains the same.

6 Reactant particles must collide with each other, collide with sufficient energy to break the bonds within the reactants, collide with the correct orientation to break the bonds within the reactants and so allow the formation of new products.

7 Collision 1 has the correct collision orientation, allowing bonds to break within the reactants and bonds to form within the products.

8 **a**

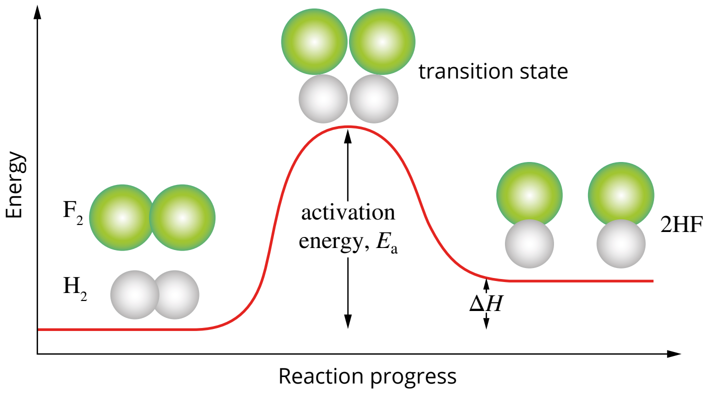


b Activation energy is the minimum amount of energy required to break bonds in the reactants in order to form products in a reaction.

c The reaction is endothermic.

ISBN 978 1 4886 1772 0

d



e F–F and H–H bonds are being broken. H–F bonds are beginning to form.

9 **a** The single H–H bond in each hydrogen molecule and the double O=O bond in each oxygen molecule are broken in the course of this reaction.

b Two H–O bonds are formed in each new water molecule during the reaction.

c The energy change for the reaction is the difference between the energy absorbed to break the bonds in the H_2 and O_2 reactants, and the energy released when the bonds in the H_2O product are made.

d No reaction occurs until sufficient energy is supplied to overcome the activation energy.

10 B. B is correct because an increase in surface area increases the frequency of collisions between reactant particles.

11 D

12 **a** At lower temperatures, the molecules that react to cause the apple to brown have less energy. They collide less frequently and are less likely to have a total energy exceeding the activation energy of the reaction involved, decreasing the rate of reaction.

b Using pure oxygen gas instead of air increases the concentration of oxygen. This results in a higher frequency of collisions between hydrogen and oxygen molecules and hence an increased reaction rate.

13 A. Reactant is not gaseous so increasing pressure of oxygen gas would have no effect on the rate of reaction.

14 C

15 **a** surface area of a solid reactant, concentration of reactants in a solution, pressure of any gaseous reactants, temperature of the reaction, presence of a catalyst

b **i** surface area of a solid reactant, concentration of reactants in a solution, pressure of any gaseous reactants

ii temperature of the reaction (also increases collision frequency), presence of a catalyst

16 **a** $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

b $n(CaCO_3) = 10.0\,g/100.1\,g\,mol^{-1} = 0.0999\,mol$

$n(HCl) = 0.1\,mol\,L^{-1} \times 0.1\,L = 0.01\,mol$

$\therefore$ $CaCO_3$ is in excess

c The rate of reaction can be measured by:

- a decrease in mass of reaction mixture as $CO_2(g)$ escapes to the atmosphere
- an increase in pH with a pH probe as acid is consumed.

d The rate of reaction with the smaller lumps will be faster. The smaller lumps have a larger surface area so more collisions can occur per second.

e increase temperature; increase concentration of hydrochloric acid. Both increase the frequency of collisions between reactants, and temperature also increases the proportion of successful collisions.

17 **a** Higher body temperature increases the rate of reactions. Increased pulse and breathing rate increases the concentration of reactants, which increases the frequency of successful collisions.

b Lower body temperature decreases rate of metabolic reactions in the body, which protects vital organs from damage.

18 You could do any of the following to increase the rate of dissolution without ruining the toffee.

- Grind up the sugar crystals or use caster sugar.
- Use a cup of hot water to dissolve the sugar.
- Gently heat the sugar and water mixture while the sugar was dissolving.
- Stir the sugar and water mixture while the sugar was dissolving.

ISBN 978 1 4886 1772 0

Chapter 19 **Catalysts**

Section 19.1 **Catalysts**

19.1 KEY QUESTIONS

1 D. A and C are decreased because a catalyst decreases the activation energy required for a reaction.

2 **a** A catalyst is a substance that increases the rate of a chemical reaction without itself undergoing permanent change.

b Activation energy is the minimum amount of energy required to break the bonds in reactants in order to form products in a reaction.

3 When salt is mixed with sugar, the salt acts as a catalyst and lowers the activation energy of the combustion reaction between sugar and oxygen.

4 **a** Reactions involving a heterogeneous catalyst take place at the surface of the catalyst. Reactants form bonds with the catalyst, lowering the activation energy of reactions and allowing them to proceed more rapidly.

b A porous pellet has a much larger surface area than a solid lump. More reactants may be in contact with the surface of a porous pellet at any instant, producing a faster rate of reaction.

CHAPTER 19 **REVIEW**

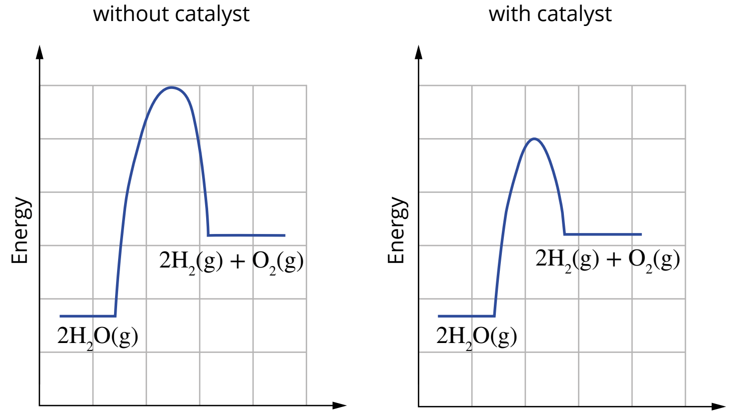
1 **a** Increasing the pressure of the gases would cause an increase in the number of collisions between the reactant molecules in a given time, so more collisions would occur with the correct orientation to react and with energy that is greater than or equal to the activation energy. As a result, the rate of reaction would increase.

b Adding a catalyst would allow the reaction to occur by a different pathway with a lower activation energy. The proportion of collisions with energy greater than the activation energy would thus be increased. As a result, the rate of reaction would increase.

2 **a** 1370 kJ mol^{-1}

b –572 kJ mol^{-1}

c



without catalyst

with catalyst

d +572 kJ mol^{-1}

3 **a** Homogeneous. Hydrogen peroxide is a liquid and the dichromate ion catalyst is in an aqueous solution, thus they are the same phase.

b Heterogeneous. The reactants are gaseous while the catalyst is solid (V_2O_5).

c Heterogeneous. The reactants are gaseous while the catalyst is solid (nickel metal).

d Homogeneous. Both the reactant (ethanol) and catalyst (enzymes) are in aqueous solution.

4 Heterogeneous catalysts are:

- more easily separated from the products of a reaction
- much easier to reuse
- able to be used at high temperatures.

5 **a** False. A catalyst has no effect on the quantity of product; it only makes the reaction faster.

b True. Catalysts are not consumed in the reaction so can be recovered unchanged from the reaction mixture.

c False. Catalysts do not affect the energy of the particles. A catalyst increases the rate of reaction by providing an alternative reaction pathway that has a lower activation energy than the uncatalysed pathway.

d False. The enthalpy of a reaction is a constant. A catalyst lowers only the activation energy.

6 C. Catalysts affect the rate of reaction by providing an alternative reaction pathway with a lower activation energy; the difference in the enthalpies of the products and reactants remain unchanged.

7 The rapid effervescence on addition of the cobalt(II) ions is evidence that CO_2 is being rapidly formed in the reaction, which suggests the cobalt(II) ions have increased the rate of reaction. The colour change from pink to green suggests that the cobalt(II) ions are involved in the reaction and are changed to an intermediate during the course of the reaction, but the intermediate is changed back to cobalt(II) ions at the end of the reaction.

8 **A** is correct. Lactase is an enzyme and enzymes are proteins.

B is correct. The substance that is interacting with the enzymes is called a substrate.

C is incorrect. Active sites are on the enzyme not the substrate.

D is incorrect. Glucose is a product of the reaction. The substrate is the reactant interacting with enzyme.

9 The intricate three-dimensional structure of their active site is destroyed at pH values outside these ranges. The enzyme and substrate will be unable to form the enzyme–substrate complex that allows the reaction to proceed.

10 Catalytic converters are able to convert pollutants formed during combustion of hydrocarbon fuels in an internal combustion engine into CO_2, H_2O and N_2. A catalyst made of rhodium and platinum is used in catalytic converters. It is able to catalyse the removal of CO and NO gases:

$$2NO(g) \rightarrow N_2(g) + O_2(g)$$

$$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$$

The catalyst also converts unburnt hydrocarbon into carbon dioxide and water, e.g.

$$2C_8H_{18}(g) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$$

11 The order for the reaction pathway is X, Z, W, Y.

12 The platinum gauze is acting as a catalyst. It provides an alternative reaction pathway with a lower activation energy than the reaction pathway without the gauze. This allows hydrogen and oxygen gases to react at a lower temperature.

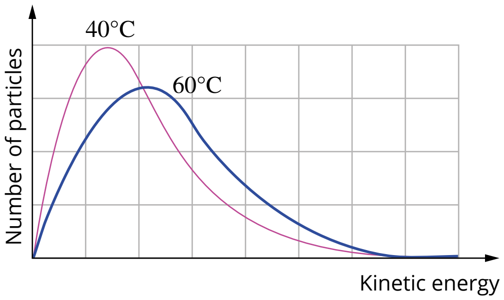
13 Nanomaterials have a very large surface area so there is much greater contact between reactant particles and catalyst, which means more catalytic reactions can occur in any given time.

The larger surface area of nanomaterials means it is possible to use less material to achieve the same level of catalytic activity.

14 C. Enzymes are biological catalysts. Enzymes have an active site that the substrate (reactant) attaches to. In this example, the substrate forms an attachment to the enzyme through the positively polarised H atom, forming a hydrogen bond intermolecular attraction to the lone pair of electrons on the oxygen atom.

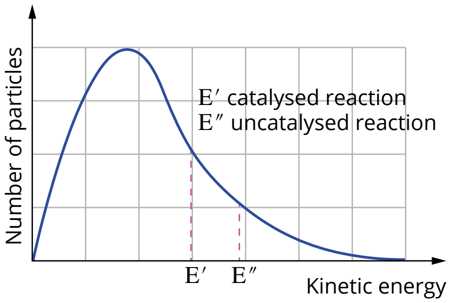
15 Nitrogen molecules, N_2, consist of two nitrogen *atoms* held together by a triple covalent bond. This is a particularly *strong* covalent bond and so the reaction has a large activation energy in the absence of a catalyst. Iron is a suitable catalyst for this reaction, capable of providing a new reaction pathway and so *decreasing* the activation energy. Nitrogen and hydrogen molecules *adsorb* onto the iron surface and form bonds with the iron surface. The *covalent* bonds in the N_2 and H_2 molecules are *broken* and individual neighbouring nitrogen and hydrogen atoms *form covalent bonds* and become ammonia molecules. The ammonia molecules then leave the iron surface.

16 **a**



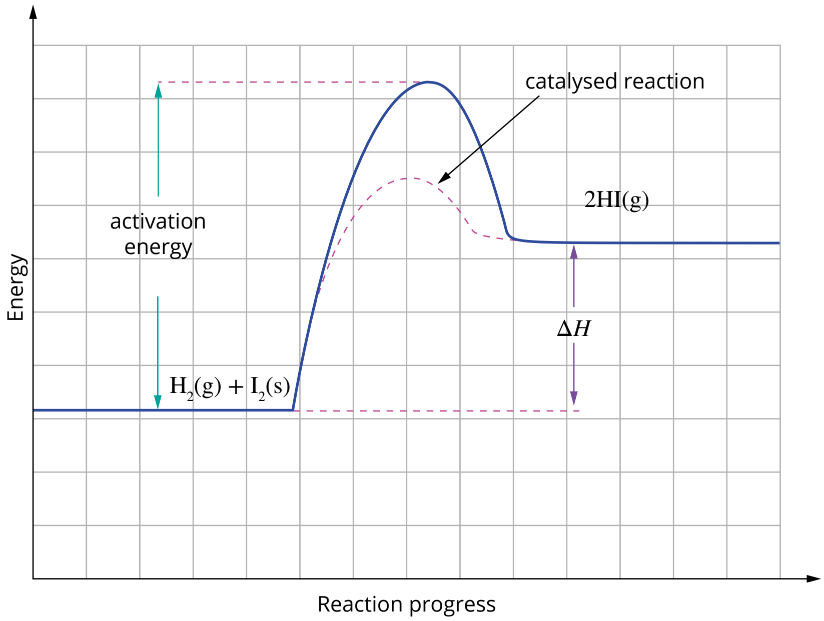
Copyright © Pearson Australia 2018 (a division of Pearson Australia Group Pty Ltd) ISBN 978 1 4886 1772 0

b



c When a catalyst is present, the reaction proceeds by an alternative reaction pathway with a lower activation energy (*E*′) than in the uncatalysed reaction (*E*″). At a given temperature, there will be a greater proportion of reacting particles that have a kinetic energy equal to or greater than the activation energies (*E*′ than *E*″). As more reactants have sufficient energy to react, the rate of reaction increases.

17 **a, c**



b The reaction is endothermic.

ISBN 978 1 4886 1772 0

Chapter 20 Science inquiry skills in chemistry

Section 20.1 Questioning

Worked example: Try yourself 20.1.1

DEVELOPING A HYPOTHESIS

Develop a hypothesis and identify variables for investigating how the molecular mass of a hydrocarbon affects the energy released per gram of the hydrocarbon.

Thinking	Working
Write an inquiry question.	What is the relationship between the molecular mass of a hydrocarbon and the energy released per gram of the hydrocarbon?
Identify the independent variable.	molecular mass of hydrocarbon
Identify the dependent variable.	energy (heat) released
Identify the controlled variables.	type of hydrocarbon (e.g. all alkanes) mass of water height of water above spirit burner amount of stirring of water
Write a potential hypothesis.	The energy released per gram of the hydrocarbon increases as the molecular mass of the hydrocarbon fuel increases.

20.1 KEY QUESTIONS

1 **a** A hypothesis is a statement that can be tested. This involves making a prediction based on previous observations.

b A theory is a hypothesis that is supported by a great deal of evidence from a wide variety of sources. A principle is a theory that is so strongly supported by evidence that it is unlikely to be shown to be untrue in the future.

2 A

3 **a** electrical conductivity

b concentration of lead

c electrical conductivity

d pH

4 qualitative

5 C

6 C. A hypothesis should test only one independent variable and it should predict the relationship between the independent and dependent variable. Hypothesis 1 tests two independent variables. Hypothesis 2 does not predict the type of relationship between the independent and dependent variables.

ISBN 978 1 4886 1772 0

Section 20.2 Planning investigations

20.2 KEY QUESTIONS

1 B

2 **a** In a control experiment, two groups of subjects are tested; the groups, or the tests performed on them, are identical except for a single factor (the variable), which is not changed in the control experiment.

b The dependent variable is the variable that is measured to determine the effect of changes in the independent variable. The independent (experimental) variable is the variable that is changed in an experiment. For example, in an experiment testing the effect of soil pH on flower colour, the independent variable would be soil pH and the dependent variable would be flower colour.

3 **a** valid

b reliable

c accurate

4 **a** type of soft drink

b pH

c temperature of solutions

5 **a** Litmus paper and universal indicator give qualitative information about pH through colour.

b A calibrated pH meter will give quantitative information and is more accurate than using litmus paper or universal indicator.

Section 20.3 Uncertainty and error in data

20.3 KEY QUESTIONS

1 **a** systematic errors

b random errors

2 **a** systematic (Evaporation will lead to a value that is always higher than the true value after evaporation takes place.)

b random

c mistake

3 There could be many reasons why the same experimental results cannot be obtained. The experimental design may be poor because of:

- a lack of objectivity
- a lack of clear and simple instructions
- a lack of appropriate equipment
- a failure to control variables.

Other problems not specifically related to the experiment could be:

- a poor hypothesis that could not be tested objectively
- conclusions that do not agree with the results
- interpretations that are subjective.

Section 20.4 Processing data and information

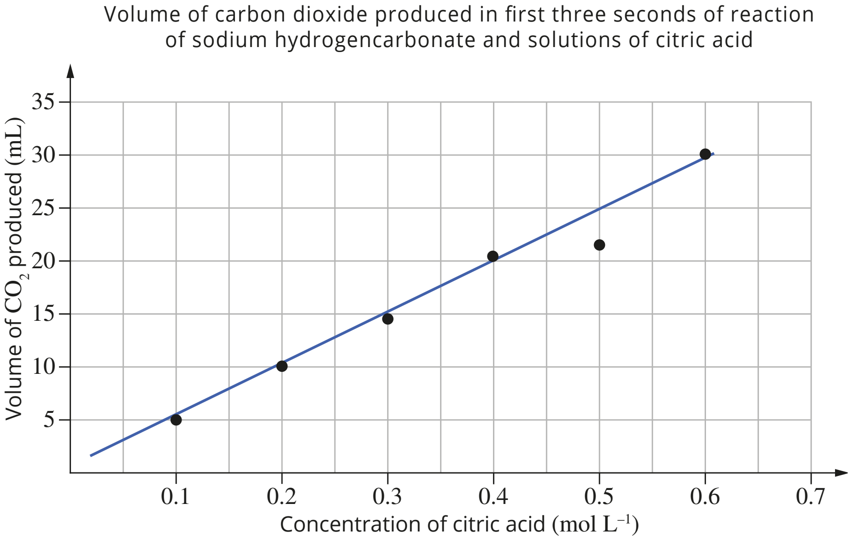
20.4 KEY QUESTIONS

1 mean = (21 + 28 + 19 + 19 + 25 + 24) ÷ 6 = 22.7

2 Add a trend line or line of best fit.

ISBN 978 1 4886 1772 0

3 **a**



b Data point 5 (0.5, 21.5) is an outlier.

c An outlier is a point in the data that does not fit the trend.

Section 20.5 Analysing data and information

20.5 KEY QUESTIONS

1 A linear graph shows the proportional relationship between two variables.

2 an inversely proportional relationship

3 directly proportional

4 time restraints and limited resources

Section 20.6 Conclusions

20.6 KEY QUESTIONS

1 B. A scientific article should always be written using objective language. A concluding paragraph must summarise the information presented in the article and connect it with the title. It should also include limitations, possible applications of the research and potential future research.

2 The statement ‘Many repeats of the procedure were conducted’ is unquantified. ‘Thirty repeats of the procedure were conducted’ is better because the number of trials is quantified.

Section 20.7 Communicating

20.7 KEY QUESTIONS

1 B. Scientific writing should not use biased or absolute language.

2 D. First-person narrative uses the pronoun ‘I’.

3 **a** $\mathrm{g\,mol^{-1}}$

b $\mathrm{mol\,L^{-1}}$

4 Divide the value in g by 1000.

5 to reflect how we use quantities in the laboratory or to make the numbers easier to comprehend

ISBN 978 1 4886 1772 0

CHAPTER 20 REVIEW

1 **a** dependent variable
  b controlled variable
  c independent variable

2 **a** reflect
  b create
  c analyse
  d investigate
  e apply
  f identify
  g describe

3 independent variable = source of the water; dependent variable = heavy metal concentration; controlled variables = temperature, time of testing

4 **a** bar graph
  b line graph
  c scatter graph (with line of best fit)
  d pie diagram

5 **a** It can dissolve or eat away at substances including tissue such as your skin or lungs.
  b It is toxic (poisonous) if inhaled.
  c It is a highly combustible liquid that could catch on fire.

6 Accuracy refers to the ability of the method to obtain the correct measurement close to a true or accepted value.

Validity refers to whether an experiment or investigation is in fact testing the set hypothesis and aims.

7 **a** mistake
  b random error
  c systematic error

8 an exponential relationship/logarithmic relationship

9 **a** reliability
  b validity
  c accuracy
  d precision

10 Evaluate the method; identify issues that could affect validity, accuracy, precision and reliability of data; state systematic sources of error and uncertainty; and recommend improvements to the investigation if it is to be repeated.

11 A trend is a pattern or relationship that can be seen between the dependent and independent variables. It may be linear, in which the variables change in direct proportion to each other to produce a straight trend line. The relationship may be in proportion but non-linear, giving a curved trend line. The relationship may also be inverse—one in which a variable decreases in response to the other variable increasing. This could be linear or non-linear.

12 any issues that could have affected the validity, accuracy, precision or reliability of the data plus any sources of error or uncertainty

13 Bias is a form of systematic error resulting from a researcher's personal preferences or motivations.

14 C

15 The purpose of referencing and acknowledgments is to avoid plagiarism and ensure creators and sources are properly credited for their work.

16 A. In APA style, the in-text citation should list the first author's last name and the year of publication.

17 **a** 0.0300 L or 3.00×10^{-2} L
  b 34 500 g or 3.45×10^{4} g

ISBN 978 1 4886 1772 0

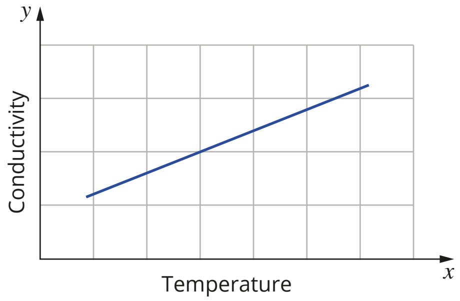
18 **a** Aim: To determine the effect of increasing water temperature on the electrical conductivity of water

b independent variable—water temperature; dependent variable—electrical conductivity of water; controlled variables—pH, water source, sampling container

c The data collected would be conductivity using a probe and, therefore, it would be quantitative.

d 10 mL measuring cylinder ±0.1 mL, pH probe ±0.02, alcohol-filled glass thermometer ±0.1°C, electrical conductivity probe ±2

e Raw data is data collected from experiments and recorded as measurements are taken. Processed data tabulates this into a form in which the reader can clearly see the temperature of the water and the conductivity at each separate temperature value. This can also then be processed and graphed with the independent variable on the *x*-axis (temperature) and dependent variable on the *y*-axis (conductivity). If the hypothesis is correct, the graph for this experiment would look like the one below:



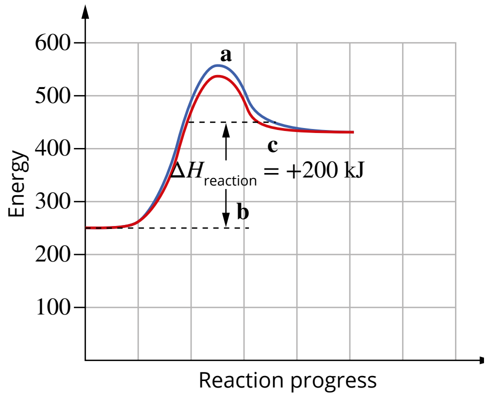
Unit 2 Review

Section 1: Multiple choice

1 B
2 C
3 C
4 A
5 C
6 D
7 D
8 C
9 A
10 D
11 B
12 C

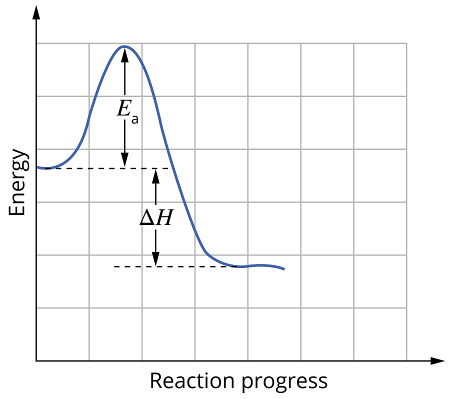
Section 2: Short answer

1

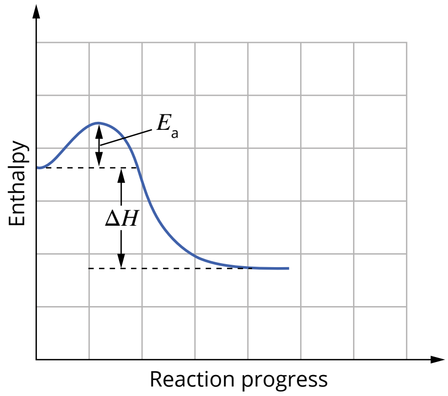


2 **a** $2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$ + heat

b



c



3 **a** $H_2SO_4(aq) + 2KOH(aq) \rightarrow H_2O(l) + K_2SO_4(aq)$

b $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

c $6HCl(aq) + 2Al(s) \rightarrow 2AlCl_3\ (aq) + 3H_2(g)$

d $6H^+(aq) + 2Al(s) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

4 **a** $M(CaBr_2) = 199.9\,g$

$n(CaBr_2) = \frac{m}{M} = \frac{20.0}{199.9} = 0.100\,mol\ L^{-1}$

$c(CaBr_2) = \frac{n}{V} = \frac{0.100}{0.400} = 0.250\,mol\ L^{-1}$

b $M(NaCl) = 58.44\,g$

$n(NaCl) = cV = 0.600 \times 1000.0 = 600.0\,mol$

$m(NaCl) = nM = 600.0 \times 58.44 = 3.51 \times 10^5\,g$ (3 sig. figs)

5 The molecules are all polar and bond in the liquid state by dipole–dipole forces between the molecules. As molar mass increases, the dispersion force of attraction between the molecules increase, therefore the boiling point increases.

Between the HF molecules there is a particularly strong dipole–dipole force called a hydrogen bond. This causes the liquid HF to have a higher BP than the other hydrogen halides despite it having a lower molar mass.

6 **a** Boyle's law: As pressure increases the gaseous particles (atoms or molecules) are forced closer together and therefore volume decreases as pressure increases.

b Charles' law. As temperature increases, the molecules (or atoms) gain more kinetic energy and will move faster and further causing a gas to expand when heated.

Section 3: Extended answer

1 $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$

a $M(NH_3) = 17.03\,g$

i $n(NH_3) = \frac{m}{M} = \frac{1.00 \times 10^6}{17.03} = 5.872 \times 10^4\,mol$

$\therefore n(O_2) = \frac{5}{4} \times 5.872 \times 10^4 = 7.340 \times 10^4\,mol$

$\therefore m(O_2) = n \times M = 7.340 \times 10^4 \times 32.0 = 2.35 \times 10^6\ g$ (3 sig. figs)

ii $n(NH_3) = 5.872 \times 10^4\,mol$

$\therefore n(NO) = 5.872 \times 10^4\,mol$

$V(NO) = 5.872 \times 10^4 \times 22.71 = 1.33 \times 10^6\,L$ at STP (3 sig. figs)

b The rate of a chemical reaction depends on the frequency of successful collisions between the reaction particles (atoms, ions and molecules). The higher the number of collisions per second, the more chance of a successful collision and the faster the rate.

i High concentration of reacting molecules causes more frequent collisions per second, and so there is more chance of a successful collision and the reaction rate will increase.

ii High temperature increases the kinetic energy of the molecules and therefore of collision frequency. Higher temperature also means that more reacting molecules have enough energy to overcome the activation energy.

iii High pressure causes more molecules to occupy a space or volume and so collide more frequently, giving more chance of a successful collision and so a higher rate of reaction.

iv A catalyst provides another reaction path that has a lower activation energy and reaction will continue more rapidly. The finely divided platinum provides a much greater surface area on which a reaction may occur and therefore speeds up the reaction.

ISBN 978 1 4886 1772 0

2 **a** as a percentage (parts per 100) and most commonly as ppm

b $n(\text{NaCl}) = \frac{m}{M} = \frac{36.0}{58.43} = 0.606\,\text{mol}$

$\therefore$ 0.616 mol per 100.0 g

c 0.616 mol per 100.0 g = 6.16 mol per 1000.0 g = 6.16 mol L^{-1}

d 15.0 g of NaCl in 25 g of water = 60.0 g in 100 g of water
The solubility is 36.0 g per 100.0 g of water so some would remain undissolved and the solution will be saturated.

e The solubility changes as temperature changes so temperature must be stated for comparison.

f The beaker would contain a clear, colourless liquid with a white solid on the bottom.

g Solid ionic crystals such as NaCl have a strong ionic bond between the ions and there are no free charged species (ions or electrons) to allow conduction.

However, when dissolved in water, the ionic crystal dissociates into freely moving ions, which are attracted to the positive and negative electrodes. This movement of ions allows the solution to conduct a current.

$$\underset{\text{ionic solid}}{\text{NaCl(s)}} \xrightarrow{H_2O} \underset{\text{free mobile ions}}{\text{Na}^+\text{(aq)} + \text{Cl}^-\text{(aq)}}$$

h Sugar is a molecular solid. All the electrons are covalently bonded within the molecule and there are no charged particles. When sugar dissolves in water, the weak molecular bonds (hydrogen bonds) between the molecules are broken by the interaction (bonding) with the water molecules. As there are no charged particles, the solution is a non-conductor.
There will, however, be a very low conductivity due to the weak ionisation of water.

$$\underset{\text{covalent molecular solid}}{C_{12}H_{22}O_{11}\text{(s)}} \xrightarrow{H_2O} \underset{\text{neutral molecules}}{C_{12}H_{22}O_{11}\text{(aq)}}$$

Salt solutions do conduct electricity because of the freely moving ions.

ISBN 978 1 4886 1772 0