Chapter 2 Elements and the periodic table

2.1 The atomic world

Worked example: Try yourself 2.1.1

CALCULATING THE NUMBER OF SUBATOMIC PARTICLES

Calculate the number of protons, neutrons and electrons for the atom with this nuclide symbol: $^{235}_{92}U$

Thinking	Working
The atomic number is equal to the number of protons.	The 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.	Number of electrons = Z = 92

Worked example: Try yourself 2.1.2

CALCULATING THE NUMBER OF SUBATOMIC PARTICLES IN AN ION

Calculate the number of protons, neutrons and electrons for the ion with this nuclide symbol: $^{14}_{7}N^{3-}$

Thinking	Working
The atomic number is equal to the number of protons.	The number of protons = Z = 7
Find the number of neutrons. Number of neutrons = mass number – atomic number	The number of neutrons = A – Z = 14 – 7 = 7
Find the number of electrons in an uncharged atom. The number of electrons is equal to the atomic number.	Number of electrons in balanced atom = Z = 7
Find the number of electrons in the ion. The number of electrons is equal to the atomic number plus three, because the total negative charge is three more than the total positive charge.	Number of electrons in ion = 7 – (charge) = 7 – (–3) = 10

KEY QUESTIONS

Knowledge and understanding

1 10 000–100 000 times larger, depending on the element.

2 Protons and neutrons found in the nucleus.

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

4 Mass number

Analysis

5 30 – 2 = 28 electrons

6 **a** $^{90}_{39}Y$

b 51

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7 **a** and **c** as they have the same atomic number but different mass numbers.

8 It would be easier to separate $^{46}_{20}Ca$ and $^{46}_{22}Ti$ as they are different elements, with different chemical properties. $^{46}_{20}Ca$ and $^{40}_{20}Ca$ would be difficult to separate as they are different types of the same element and have identical chemical properties.

2.2 Emission spectra and the Bohr model

Worked example: Try yourself 2.2.1

ELECTRONIC CONFIGURATION FOR UP TO 36 ELECTRONS

Apply the order of filling of the shell model to determine the electronic configuration of an atom with 34 electrons.

Thinking

Recall the maximum number of electrons that each shell can hold.

Working

Shell (*n*)	Maximum number of electrons
1	2
2	8
3	18
4	32

Thinking

Place the first 18 electrons in the shells from the lowest energy to the highest energy. Do not exceed the maximum number of electrons allowed.

Working

Shell (*n*)	Electrons in atom
1	2
2	8
3	8
4	

Thinking

Place the next two electrons in the fourth shell.

Working

Shell (*n*)	Electrons in atom
1	2
2	8
3	8
4	2

Thinking

Continue filling the third shell until it holds up to 18 electrons. Put any remaining electrons in the fourth shell.

Working

Shell (*n*)	Electrons in atom
1	2
2	8
3	18
4	6

Of the 14 remaining electrons, 10 electrons from the previous step have gone into the third shell, and 4 have gone into the fourth shell.

Thinking

Write the electronic configuration by listing the number of electrons in each shell separated by commas.

Working

2,8,18,6

KEY QUESTIONS

Knowledge and understanding

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 Energy is emitted as coloured light or electromagnetic radiation.

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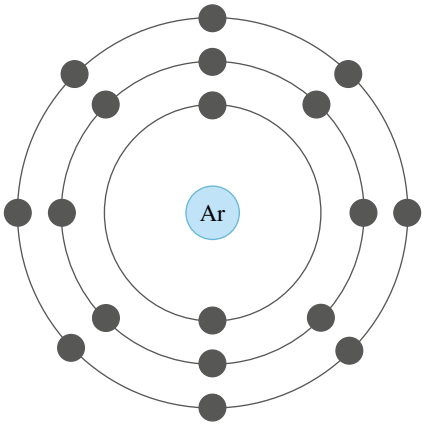
3 18: Argon 2,8,8



Analysis

4 $2n^2 = 2 \times 5^2 = 50$

5 The number of shells corresponds to the row (period) number in the periodic table. The number of valence electrons determines the group number (column).

6 The atom is sulfur because it has 16 electrons. This electronic arrangement is unexpected because the second shell is not filled and electrons have been placed in the third shell. The rules for electronic configuration are that the lower energy shells are filled first; however, the second shell (lower energy) has not been filled. A possible reason for this is that this atom has been excited by an energy source. Two electrons have 'jumped' from the second shell to the third shell. It is expected that, moments after this, the two electrons will return to the second shell and, at the same time, energy in the form of heat or light will be released from the atom.

2.3 The Schrödinger model of the atom

Worked example: Try yourself 2.3.1

WRITING ELECTRONIC CONFIGURATIONS USING THE SUBSHELL MODEL

Write the subshell model of electronic configuration for a vanadium atom with 23 electrons.

Thinking	Working
Recall the order in which the subshells fill by listing them from lowest energy to highest energy and the number of orbitals in each.	1*s*, 1 orbital 2*s*, 1 orbital 2*p*, 3 orbitals 3*s*, 1 orbital 3*p*, 3 orbitals 4*s*, 1 orbital 3*d*, 5 orbitals 4*p*, 3 orbitals
Fill the subshells by assigning two electrons per orbital, starting from the lowest energy subshells until you have reached the total number of electrons in your atom.	**Subshell** | **Electrons in subshell** | **Progressive total of electrons** 1*s* | 2 | 2 2*s* | 2 | 4 2*p* | 6 | 10 3*s* | 2 | 12 3*p* | 6 | 18 4*s* | 2 | 20 3*d* | 3 | 23
Write the electronic configuration by writing each subshell with the number of electrons as a superscript. Remember to group subshells from the same shell.	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^3 4s^2$

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KEY QUESTIONS

Knowledge and understanding

1

Element (atomic number)	Electronic configuration using the shell model	Electronic configuration using the subshell model
boron (5)	2,3	$1s^2 2s^2 2p^1$
lithium (3)	2,1	$1s^2 2s^1$
chlorine (17)	2,8,7	$1s^2 2s^2 2p^6 3s^2 3p^5$
sodium (11)	2,8,1	$1s^2 2s^2 2p^6 3s^1$
neon (10)	2,8	$1s^2 2s^2 2p^6$
potassium (19)	2,8,8,1	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
scandium (21)	2,8,9,2	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^1 4s^2$
copper (29)	2,8,18,1	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$
bromine (35)	2,8,18,7	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$

2
a Ga
b Br
c Cu
d Mn
e Sn
f Sb
g Xe
h K

Analysis

3 The Bohr model was only able to accurately predict the emission spectrum for hydrogen, whereas the Schrödinger model accurately predicts emission spectra for more complex atoms. The Bohr model was unable to explain why the third shell would stop filling after 8 electrons, then would be back filled later. This is explained by the more complex arrangement of subshells and orbitals in the Schrödinger model.

4 The Schrödinger model is a refinement of the Bohr model. The Bohr model proposed that all electrons in the one shell were of equal energy. Evidence from emission spectra indicated that there were different electronic energy levels (called subshells) within a shell. The Schrödinger model includes these subshells.

5 $1s^2 2s^2 2p^6 3s^1 3p^5$ (Other answers are possible that show inconsistent filling of the subshells.)

2.4 The periodic table

CASE STUDY: ANALYSIS

Naming elements on the periodic table

Analysis

1
a iron (Fe)
b potassium (K)
c tungsten (W)
d lead (Pb)
e mercury (Hg)

2 IUPAC guidelines state that any new element must be named after either:
- a mythological concept or character (including an astronomical object),
- a mineral or similar substance
- a place, or geographical region
- a property of the element, or
- a scientist.

Names are proposed by the teams that discovered them, and then the name is chosen and approved by IUPAC.

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3 113 – nihonium (Nh): Nihon is one way to say 'Japan'. Named after Japan, where it was discovered. The first element discovered by scientists working in an Asian country.

114 – flerovium (Fl): Honours the Flerov Laboratory of Nuclear Reactions, itself named after Georgiy N. Flerov (1913 – 1990), a renowned physicist.

115 – moscovium (Mc): Named after the Moscow region, where the discovery experiments were conducted.

116 – livermorium (Lv): Honours the Lawrence Livermore National Laboratory, California.

117– tennessine (Ts): Recognises the laboratories in the Tennessee region of the United States that were involved in the discovery.

118 – oganesson (Og): Named after the Russian nuclear physicist Yuri Oganessian, who is credited with three earlier confirmed elemental discoveries.

KEY QUESTIONS

Knowledge and understanding

1 row

2 Periods are the horizontal rows in the periodic table. Periods are numbered 1–7. Groups are the vertical columns in the periodic table and are labelled 1–18. The periodic table has four main blocks. The elements in each block have the same type of subshell (*s*, *p*, *d* or *f*) as their highest energy subshell.

3 main group

4 **a** alkali metals
b alkaline earth metals
c halogens
d noble gases

5 **a** 2
b 13 – 10 = 3
c 15 – 10 = 5
d 18 – 10 = 8

6 The element is in period 4 and therefore has four occupied shells. As the element is in group 1, it will have one valence electron. This gives an electron configuration of 2,8,8,1 or $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$.

Analysis

7 **a** **i** Group 13
ii Group 17
iii Group 1
iv Group 18
v Group 14
vi Group 14
b **i** Silicon, Si, 2,8,4 or $1s^2 2s^2 2p^6 3s^2 3p^2$
ii Beryllium, Be, 2,2 or $1s^2 2s^2$
iii Argon, Ar, 2,8,8 or $1s^2 2s^2 2p^6 3s^2 3p^6$
c **i** 4
ii 2
iii 1
iv 1
v 7
vi 3

8 A critical element is an element heavily relied on by industry and society, which faces some form of supply uncertainty. Examples include endangered elements like osmium and iridium, conflict elements such as tin, tungsten and gold and critical raw materials such as the lanthanides. (Many other examples are possible.)

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2.5 Trends in the periodic table

Worked example: Try yourself 2.5.1

EFFECTIVE NUCLEAR CHARGE

Determine the effective nuclear charge of an atom of fluorine.

Thinking	Working
Determine the number of electrons in an atom of the element, using the periodic table as a reference.	The atomic number of fluorine is 9. Therefore, an atom of fluorine has 9 protons and 9 electrons.
Use the number of electrons to determine the electronic configuration.	With 9 electrons the electronic configuration is $1s^2 2s^2 2p^5$.
Determine the effective nuclear charge. Effective nuclear charge = number of protons – number of inner-shell electrons	There are 2 inner-shell electrons. Effective nuclear charge = 9 – 2 = +7

KEY QUESTIONS

Knowledge and understanding

1 Effective nuclear charge is the pull from the nucleus felt by each valence electron. Effective nuclear charge = number of protons – number of inner-shell electrons. Carbon has 6 protons and 6 electrons. Its electronic configuration is $1s^2 2s^2 2p^2$. Carbon has two inner-shell electrons. Effective nuclear charge = 6 – 2 = +4.

2 As effective nuclear charge increases, electronegativity increases.

3 **a** metal
b non-metal
c metalloid
d metal
e metal
f non-metal

Analysis

4 **a** **i** F
ii Fr
b **i** Group 17
ii Group 1

5 N, B, Cl, Ga, Al

6 As you move 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.

7 K, Na, Al, Ca, S, P, He

8 The reactivity of the alkali metals with water is related to the ease with which the metal ion is ionised. This is based on the first ionisation energy of the elements. Down the group the effective nuclear charge stays constant; however, the number of shells increases. Therefore, the valence electrons are less attracted to the nucleus the further they are from the nucleus. As a result, the energy required to overcome the attraction between the nucleus and the valence electrons is less, and the first ionisation energy decreases down a group. This means that less energy is required to ionise caesium than lithium, and so caesium is more reactive than lithium.

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Chapter 2 **Review**

REVIEW QUESTIONS

Knowledge and understanding

1 D. The maximum number of electrons a shell can hold is given by $2n^2$, where n is the shell number. D is the only option that fits this rule.

2 B and C. Isotopes are the same element, therefore they contain the same number of protons and electrons as one another. The number of neutrons, and therefore nucleons, is different.

3 C. This is an atom of potassium, but the electron configuration shown does not follow the order of filling: one of the electrons in the 3*p* subshell has been excited to the 4*s* subshell. Options A, B and D all follow the expected order of filling for an atom in its ground state.

4 B. Elements are ordered according to atomic number, which is the number of protons in their nucleus.

5 C. Fluorine is the most reactive non-metal on the periodic table.

6 **a** Atomic number is 24; mass number is 52
b 24 electrons, 24 protons, 52 – 24 = 28 neutrons

7 Magnesium. The total of 12 electrons means the element has 12 protons. The element that has 12 protons and atomic number 12 is magnesium.

8 $1s^2 2s^2 2p^6 3s^2 3p^4$

9 **a** Period 1, *s*-block
b Period 2, *p*-block
c Period 3, *p*-block
d Period 4, *d*-block
e Period 7, *f*-block

10 **a** The force of attraction between the nucleus and valence electrons **increases** in a period from left to right.
b Atomic radii of elements **decrease** in a period from left to right.
c Atomic radii of elements **increase** in a group from top to bottom.
d Metallic character of elements **increases** from top to bottom in a group.

Application and analysis

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 **a** G and H, D and F. These pairs of atoms have the same atomic number and different mass numbers, therefore they belong to isotopes of the same element.
b B, C, D, I all have a mass number that is twice the mass number.
c C has an atomic number of 16, thus it is sulfur.
d A. A magnesium atom has 12 electrons. This atom has a mass number of 13, so it has 13 protons and 13 electrons.
e 7. The symbols represent atoms with seven different atomic numbers: 13, 20, 16, 7, 3, 9, 2.

13 $n = 3$. The first 2 electrons fill the first shell, the next 8 the second shell, the next 8 the third shell. The next 2 electrons go into the fourth shell and the remaining 10 electrons go into the third shell.

14 **a** $1s^2 2s^2 2p^6$
b $1s^2 2s^2 2p^6 3s^2 3p^6$
c $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$ – note that when an ion forms, electrons are lost from the outer-shell first, not necessarily the highest energy subshell.

15 In the Schrödinger model of the atom, electron shells are divided into subshells, and each subshell can have a different energy level. According to the Schrödinger model, the 4*s*-subshell is lower in energy than the 3*d*-subshell. Therefore, the 4*s*-subshell begins filling after the 3*s*- and 3*p*-subshells but before the 3*d*-subshell.

16 **a** $1s^2 2s^2 2p^3$
b Period 2 and group 15
c 5
d +5

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17 **a** Period 2, group 2
b Period 3, group 14
c Period 4, group 13

18 As you move across period 2 from lithium to fluorine:
a the radius of the atoms decreases as the effective nuclear charge increases.
b there is a trend from metals (lithium, beryllium) to non-metals (boron, carbon, nitrogen, oxygen and fluorine); therefore, metallic character decreases.
c electronegativity increases as the effective nuclear charge increases and size of the atoms decreases.

19 **a** Chlorine is on the right-hand side of the periodic table and sodium is on the left. Atomic radius decreases across a period because the increasing effective nuclear charge pulls the outer-shell electrons more tightly to the nucleus, causing the volume of the atom to decrease.
b Fluorine is further to the right on the periodic table than lithium, and effective nuclear charge increases from left to right across the periodic table. As effective nuclear charge increases, the electrons are held more tightly to the nucleus and more energy is required to remove the first one.
c Barium and beryllium are in the same group, with beryllium higher than barium. Going down a group the atom size is increasing, meaning the outer-shell electrons are further from the nucleus. The outer electrons of beryllium are, therefore, held more tightly and are less readily released.
d The *s*-block elements have an *s*-subshell as their outer occupied electron subshell. The *s*-subshell can take one or two electrons, so the block is only two groups wide.

20 **a** Ca
b Ar
c C
d Na or Mg
e Li
f N
g F

21 Ordered elements in groups:
Group 1 Li, K, Cs, Fr
Group 2 Mg, Ca, Sr, Ba
Group 13 B, Al, Ga
Group 14 C, Ge
Group 15 As, Sb
Group 16 S, Te
Group 17 Br, I, Ts
Group 18 He, Kr, Rn

22 **a** First ionisation energy increases across period 3 (due to increasing effective nuclear charge), so it will take more energy to remove an electron from phosphorus.
b First ionisation energy decreases down group 17 (due to increasing number of shells), so it will take more energy to remove an electron from fluorine.

23 K is more reactive than Li (as it is further down group 1). Cl_2 is more reactive than Br_2 (as it is higher in group 17). Therefore, the most reactive combination will be K and Cl_2.

24 **a** A and G both have seven valence electrons as they are in group 17. D and F are both in group 2, therefore they have two valence electrons.
b A and B are in period 2. D and E are in period 4, C and G are in period 5, F and I are in period 6.
c A has two shells, and the radius decreases left to right across the table due to the increase in effective nuclear charge.
d B forms an ion by losing one electron. It then has only one electron shell.
e F. Electronegativity decreases down a group, and increases left to right across a period.
f I. Noble gases are found in group 18.
g A, G and I are found to the right of the 'staircase' that separates metals and non-metals. E is a metalloid.
h E is one of the elements located on the 'staircase' that separates metals and non-metals, known as metalloids.

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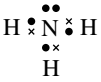
Chapter 3 Covalent substances

3.1 Covalent bonding model

Worked example: Try yourself 3.1.1

LEWIS STRUCTURES

Draw the Lewis structure of ammonia (NH_3).

Thinking	Working
Write the shell 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 outer shell.	N requires 3 electrons H requires 1 electrons
Draw a Lewis structure of the likely molecule, ensuring that each atom has a stable outer shell. Electrons not involved in bonding will be in non-bonding pairs.	Draw a Lewis structure of the molecule. 

KEY QUESTIONS

Knowledge and understanding

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

2 **a** 1. Hydrogen requires 1 electron to complete its outer shell, so it forms a single covalent bond.
b 3. Nitrogen requires 3 electrons to complete its outer shell, so it forms a triple covalent bond.
c 2. Oxygen requires 2 electrons to complete its outer shell, so it forms a double covalent bond.
d 1. Fluorine requires 1 electron to complete its outer shell, so it forms a single covalent bond.

3 **a** 1. Hydrogen requires 1 electron to complete its outer shell, so it forms a single covalent bond.
b 2. Sulfur requires 2 electrons to complete its outer shell, so it forms two covalent bonds.
c 3. Phosphorus requires 3 electrons to complete its outer shell, so it forms three covalent bonds.
d 4. Silicon requires 4 electrons to complete its outer shell, so it forms four covalent bonds.
e 1. Bromine requires 1 electron to complete its outer shell, so it forms a single covalent bond.
f 0. Argon has a stable outer shell octet of electrons, so it does not form covalent bonds.

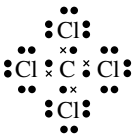
Analysis

4 **a** fluorine (F_2)

b hydrogen fluoride (HF)

c water (H_2O)

d tetrachloromethane (CCl_4)



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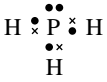
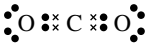
e phosphine (PH_3)



f carbon dioxide (CO_2)



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

6 **a** CF_4. A carbon atom has four valence electrons, so forms four covalent bonds. A fluorine atom has seven valence electrons, so forms one covalent bond. So four fluorine atoms each form one covalent bond with one carbon atom.

b PCl_3. A phosphorus atom has five valence electrons, so forms three covalent bonds. A chlorine has seven valence electrons, so forms one covalent bond. So three chlorine atoms each form one covalent bond with one phosphorus atom.

c CS_2. A carbon atom has four valence electrons, so forms four covalent bonds. A sulfur atom has six valence electrons, so forms two covalent bonds. So two sulfur atoms each form two covalent bonds with one carbon atom.

d SiH_4. A silicon atom has four valence electrons, so forms four covalent bonds. A hydrogen atom has one valence electron, so forms one covalent bond. So four hydrogen atoms each form one covalent bond with one silicon atom.

e NBr_3. A nitrogen atom has five valence electrons, so forms three covalent bonds. A bromine atom has seven valence electrons, so forms one covalent bond. So three bromine atoms each form one covalent bond with one nitrogen atom.

3.2 Shapes of molecules

Worked example: Try yourself 3.2.1

PREDICTING THE SHAPE OF MOLECULES

Predict the shape of a molecule of hydrogen sulfide (H_2S).	
Thinking	**Working**
Draw the Lewis structure for the molecule.	
Count the number of electron groups around the central atom.	There are four electron groups (two single bonds and two non-bonding pairs).
Determine how the electrons groups will be arranged to get maximum separation.	Because there are four electron groups, they 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 arranged in a bent shape.

KEY QUESTIONS

Knowledge and understanding

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

2 Four. One electron group is the single bond between the fluorine and hydrogen atoms; the other three groups exist as non-bonding pairs of electrons.

Analysis

3 **a**

b

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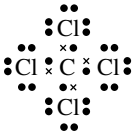
c



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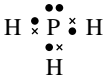

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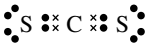

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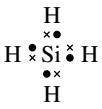
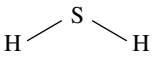


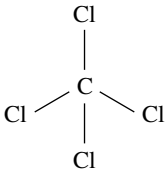
4 **a** Bent
b Linear
c Tetrahedral
d Pyramidal
e Linear
f Tetrahedral

5 **a**



b

c



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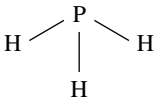
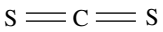


e



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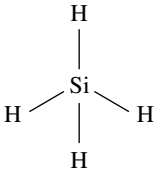

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3.3 Polarity in molecules

Worked example: Try yourself 3.3.1

COMPARING THE POLARITY OF COVALENT BONDS

Compare the polarity of the bond in nitrogen monoxide (NO) and hydrogen chloride (HCl).

Thinking	Working
Use the table of electronegativity values in Figure 3.3.2 to find the electronegativities of the atoms in each bond.	NO: nitrogen 3.0; oxygen 3.4 HCl: hydrogen 2.2; chlorine 3.2
For each bond, subtract the lowest electronegativity value from the highest value.	NO: 3.4 – 3.0 = 0.4 HCl: 3.2 – 2.2 = 1.0
Determine which bond has the biggest difference in electronegativity to determine the more polar bond.	The bond in HCl is more polar than in NO.

KEY QUESTIONS

Knowledge and understanding

1 **a** Non-polar bond: a covalent bond with an even distribution of bonding electrons. Non-polar bonds occur between atoms with the same electronegativity.
b Polar bond: a covalent bond with an uneven distribution of bonding electrons. Polar bonds occur between atoms of different electronegativity.

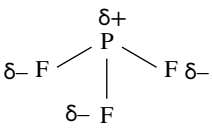
2 **a** O
b C
c N
d N
e F
f F

3 **a** P–F. Greatest difference in electronegativity values: fluorine 4.0 – phosphorus 2.2 = 1.8
b C–H. Least difference in electronegativity values: carbon 2.6 – hydrogen 2.2 = 0.4
or N–H. Least difference in electronegativity values: nitrogen 3.0 – carbon 2.6 = 0.4

Analysis

4 H = 2.2, Cl = 3.2. Therefore, HCl = 3.2 – 2.2 = 1.0
N = 3.0. Therefore, N_2 = 3.0 – 3.0 = 0
H = 2.2, Br = 3.0. Therefore, HBr = 3.0 – 2.2 = 0.8
N = 3.0, O = 3.4. Therefore, NO = 3.4 – 3.0 = 0.4
In order of increasing polarity: N_2 < NO < HBr < HCl

5



6 **a** Non-polar. CF_4 is symmetrical.
b Polar. CHF_3 is asymmetrical.
c Polar. CH_2F_2 is asymmetrical.
d Polar. CH_3F is asymmetrical.
e Non-polar. CH_4 is symmetrical.

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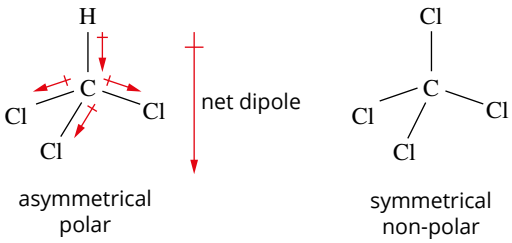
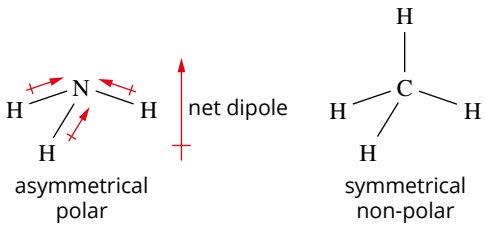
7 **a**



b



3.4 Intermolecular forces

KEY QUESTIONS

Knowledge and understanding

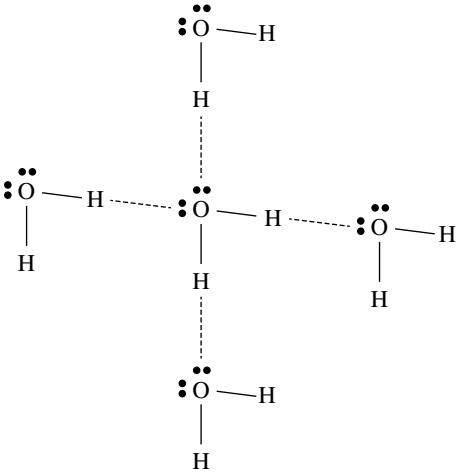
1 **a** Dipole–dipole attraction or hydrogen bonding, and dispersion forces exist between polar molecules.

b Only dispersion forces exist between non-polar molecules.

2 Hydrogen iodide (HI) and fluoromethane (CH_3F) would contain dipole–dipole attraction between their molecules. This is because they are polar molecules. They are polar because they are both asymmetrical.

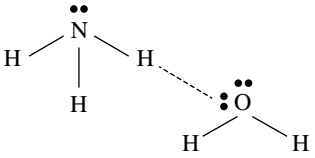
Analysis

3 **a**



b Hydrogen bonding is the strongest type of intermolecular force in ice.

4



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5 Dispersion forces: **e** and **j**; dipole–dipole attraction: **a**, **b**, **c**, **d**, **f**, **g**; hydrogen bonds: **h**, **i**

6 **a** CHF_3 has a higher boiling point. CHF_3 is a polar molecule so dipole–dipole attraction is the strongest force between molecules. CF_4 is a non-polar molecule, so there are only dispersion forces between molecules. Dipole–dipole attraction is stronger than dispersion forces.

b CO_2 has a higher boiling point. O_2 and CO_2 are non-polar molecules, so there are only dispersion forces between molecules. CO_2 molecules have a higher molecular mass so stronger dispersion forces.

c NH_3 has a higher boiling point. NH_3 is a polar molecule capable of forming hydrogen bonds between molecules. CH_4 is a non-polar molecule, so there are only dispersion forces between molecules. Hydrogen bonding is stronger than dispersion forces.

7 When sugar turns to a liquid, it is melting; the intermolecular forces are disrupted. When the liquid turns black and a gas is produced, a chemical reaction is taking place. The intramolecular bonds are broken, allowing new substances to be produced.

3.5 Covalent lattices

CASE STUDY: ANALYSIS

Mined versus synthetic diamonds

1 Similarities: Both contain carbon atoms arranged in a three-dimensional covalent network lattice structure. They are chemically the same.

Differences: Synthetic diamonds are purer with less flaws (inclusion of atoms of other elements in the lattice structure).

2 Environmental: Large open-cut mines have a huge impact on the physical environment. Land is cleared, destroying the habitats of plants and animals. Mines generate pollution that could escape into the local surrounding environment. Mining operations are energy intensive, which generates carbon pollution, contributing to global warming.

Social: In some areas of the world, the mines are located on land belonging to traditional indigenous owners. Traditional owners may become displaced from their lands, interrupting their cultural practices. So called 'blood or conflict diamonds' are sourced from war zones and sold to fund military conflicts.

3 The idea that something is 'real' and 'fake' implies there is a quality difference between the two. However, the chemical and structural composition of natural and synthetic diamonds is identical—they are both made of carbon atoms arranged into a three-dimensional covalent network lattice. Visually, both types of diamonds look identical. There is no significant difference between the two. Experts can only identify minute differences in chemical composition using advanced analytical techniques.

KEY QUESTIONS

Knowledge and understanding

1 Diamond: each carbon atom forms four covalent bonds. Graphite: each carbon atom forms three covalent bonds.

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 Any three of: printer toner, ink, reinforcement of rubber, art pencils, charcoal briquettes.

Analysis

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

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

c Diamond is a good heat conductor because the carbon atoms are strongly bonded together in the lattice.

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

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Chapter 3 **Review**

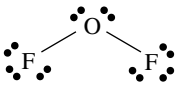
REVIEW QUESTIONS

Knowledge and understanding

1 D. Each hydrogen atom has one electron only. If two atoms share an electron each they both obtain a complete outer shell. A single covalent bond is formed. A is incorrect because it describes an ionic bond. B is incorrect because it describes metallic bonding. C is incorrect because hydrogen atoms only need two electrons for a complete outer shell.

2 B. To have a formula XY_4, atom X must require four more outer-shell electrons and atom Y must require one. The likely molecule is CH_4 as carbon has four outer-shell electrons and hydrogen requires one more.

3 A. The Lewis structure of OF_2 is:



Around the central oxygen atom, there are four electron groups and two non-bonding pairs of electrons. Therefore, OF_2 is bent.

4 C. Graphite is a covalent layer lattice, so strong covalent bonds need to be broken for sublimation to occur. Carbon dioxide, methane and methanol are covalent molecular substances with relatively weak intermolecular forces which are easily broken when these substances boil.

5 Intramolecular bonds are the forces that hold the atoms within a molecule together. In ammonia molecules they are the covalent bonds between the nitrogen and hydrogen atoms. Intermolecular forces are between one molecule and its neighbouring molecules. These are much weaker forces. It is the intermolecular forces that are disrupted when ammonia melts, allowing the molecules to move more freely around each other.

6 Neon will not form bonds to other atoms as it has a stable outer shell containing eight electrons.

7 **a** tetrahedral (no non-bonding pairs of electrons on central C atom)
b pyramidal (one non-bonding pair of electrons on central N atom)
c tetrahedral (no non-bonding pairs of electrons on central P atom)
d bent (two non-bonding pairs of electrons on central S atom)
e pyramidal (one non-bonding pair of electrons on central C atom)

8 NBr_3 – pyramidal, H_2O – bent, CH_2F_2 – tetrahedral, HCN – linear

9 F–F, O–Cl, N–O, H–Br, Si–O

10 The O–H bond in water (H_2O) is the most polar bond (answer **b**).

11 If water was a linear molecule, the two polar O–H bonds would cancel each other out and make the molecule non-polar. As water is polar, it cannot be a linear molecule, it is in fact bent.

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

13 **a** CCl_4
b CH_4 and CCl_4 are both non-polar and so there are only dispersion forces between their molecules. CCl_4 has a higher molecular mass of these two molecules, 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 for CH_4, CCl_4 has a higher boiling point so it exists as a liquid at room temperature.

14 A permanent molecular dipole is formed if there is asymmetry in the molecule. This causes an asymmetry in the electron distribution around the molecule, causing one end of the molecule to have a partial negative charge while the other end has a partial positive charge. The positive and negative ends of neighbouring molecules attract each other, forming dipole–dipole attractions.

A temporary molecular 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 have a temporary negative charge while the other end has 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.

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

16 Diamond has a high sublimation point because it has a covalent network lattice structure. Many strong covalent bonds need to be broken for sublimation to occur.

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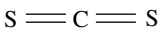
17 Sublimes at a high temperature—covalent layer structure so many strong covalent bonds need to be broken for sublimation to occur. Conducts electricity—graphite contains delocalised electrons. Lubricant—weak dispersion forces between layers allows the layers to easily slide over each other.

18 Carbon atoms in diamond have a tetrahedral bond geometry; carbon atoms in graphene within graphite have a trigonal planar bond geometry.

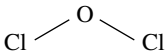
Application and analysis

19 Bent. Around the central oxygen atom, there are four electron groups and two non-bonding pairs of electrons. Therefore, HOBr is V-shaped or bent.

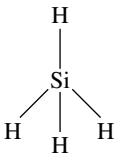
20 **a** Non-polar. CS_2 is a symmetrical linear molecule.



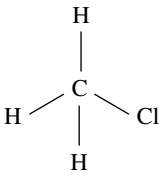
b Polar. Cl_2O is an asymmetric bent molecule.



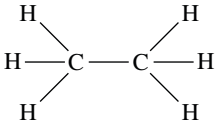
c Non-polar. SiH_4 is a symmetrical tetrahedral molecule.



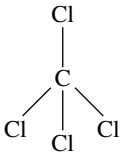
d Polar. CH_3Cl is an asymmetrical tetrahedral molecule.



e Non-polar. CH_3CH_3 is a symmetrical molecule.

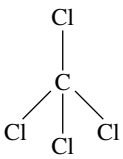


f Non-polar. CCl_4 is a symmetrical tetrahedral molecule.



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

22 **a** Non-polar. CCl_4 is a symmetrical tetrahedral molecule. Only dispersion forces between molecules.



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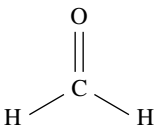
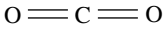
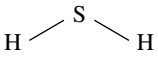
b Polar. CH_2O is an asymmetric trigonal planar molecule. Dipole-dipole attraction and dispersion forces between molecules.



c Non-polar. CO_2 is a symmetrical linear molecule. Only dispersion forces between molecules.

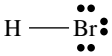


d Polar. H_2S is an asymmetrical bent molecule. Dipole–dipole attraction and dispersion forces between molecules.

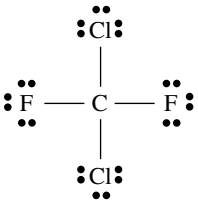


23 The formulas of each molecule below show the number of electron pairs.

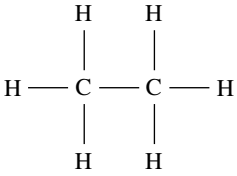
a 2 bonding electrons, 6 non-bonding electrons



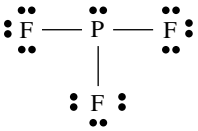
b 8 bonding electrons, 24 non-bonding electrons



c 14 bonding electrons



d 6 bonding electrons, 20 non-bonding electrons



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

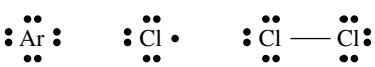
b True. A molecule of carbon dioxide contains 3 atoms (1 carbon and 2 oxygen atoms).

c True. Bonds in the molecule are intramolecular bonds.

d False. There are 2 double covalent bonds in a molecule of carbon dioxide.

e True. Each oxygen atom has 2 non-bonding pairs of electrons. As there are 2 oxygen atoms, there are 4 non-bonding pairs of electrons in a molecule of carbon dioxide.

25 Argon has the electronic configuration of 2,8,8 and therefore has eight electrons in its outer shell. The Lewis structure for an argon atom is shown (see below left), and you can see that it obeys the octet rule and will therefore exist as stable as single atoms. Chlorine, on the other hand, has the electronic configuration of 2,8,7 and therefore has seven electrons in the outer shell. The Lewis structure for chlorine is shown below (in the middle) and you can see that a single chlorine atom requires another electron to complete its octet. It can do this by bonding to another chlorine atom to form the molecule Cl_2 where both Cl atoms have a complete octet in their valence shell (see below right).



26 Melting points increase down the table because the molecules increase in molecular mass. As they are all diatomic non-polar molecules, the strength of the dispersion forces increases.

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27 CF_4 has a slightly higher boiling point (–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 dipole–dipole attraction and dispersion forces. Although CF_4 molecules are attracted by dispersion forces only, the much larger molecular mass of CF_4 molecules makes the dispersion forces stronger than the sum of the dipole–dipole and dispersion forces between OF_2 molecules.

28 Neon exists as single atoms, with the only forces of attraction being dispersion forces; therefore, neon has a very low boiling point. Hydrogen fluoride molecules, however, are very polar as a hydrogen atom is bonded to the very electronegative fluorine atom. The forces between molecules are hydrogen bonds. These are relatively strong intermolecular forces and hydrogen fluoride, therefore, has a much higher boiling point than neon.

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

b The differences in properties all relate to the fact that methane is a covalent molecular substance whereas diamond is a covalent network lattice. All bonding within diamond is strong, whereas methane has weak intermolecular dispersion forces. The many covalent bonds throughout diamond's lattice structure makes it a very hard substance with a very high sublimation point. The dispersion forces between methane molecules means it has a relatively low boiling point and exists as a gas at room temperature. The covalent bonds in methane are all contained within each individual methane molecule.

30 a Diamond has a much higher hardness value than graphite because diamond has a covalent network lattice structure so there are strong covalent bonds throughout its structure. Graphite has a covalent layer lattice structure with weak dispersion forces between layers which makes it a relatively soft, slippery substance. Graphite has a much higher electrical conductivity value because it contains delocalised electrons. Diamond does not conduct electricity because its valence electrons are all localised in covalent bonds.

b It needs to be clear whether you are referring to thermal conductivity or electrical conductivity as the comparison is different for each.

c The higher the quality of the diamond, the higher the thermal conductivity. High-quality diamonds will cool faster than low-quality ones.

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Chapter 4 Metals

4.1 Metallic properties and bonding

Worked example: Try yourself 4.1.1

DETERMINING CHARGES

Write the electronic configuration of an aluminium atom and hence 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 aluminium: 2,8,3
Only the outer-shell electrons will be lost.	Outer shell contains 3 electrons, 13 – 3 = 10 electrons remaining
Cation charge = number of protons – number of electrons	Cation charge = 13 – 10 = 3+ *(Since Al is in group 13 of the periodic table, its cation charge will be 3+.)*

CASE STUDY: ANALYSIS

Colourful transition metal compounds

1 **a** Co, Al, Cd, Fe, Cr, Pb, Sn
b Co, Cd, Fe, Cr

2 Different transition metal ions are present in rubies and sapphires. Chromium ions are present in rubies, whereas titanium and iron ions are present in sapphires.

3 Diamonds are composed almost entirely of carbon. Since transition metal ions are not present they are usually colourless.

Worked example: Try yourself 4.1.2

ELECTRONIC CONFIGURATION OF MAGNESIUM

With reference to the electronic configuration of magnesium, explain why solid magnesium can conduct electricity.

Thinking	Working
Using the atomic number of the element, determine the electronic configuration of its atoms. (You may need to refer to the periodic table.)	Mg has an atomic number of 12. This means a neutral atom of magnesium has 12 electrons. The electronic configuration is 2,8,2.
From the electronic configuration, find how many outer-shell electrons are lost to form cations that have a stable, noble gas electronic configuration. These electrons become delocalised.	Mg has 2 electrons in its outer shell. Mg atoms will tend to lose these 2 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.

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KEY QUESTIONS

Knowledge and understanding

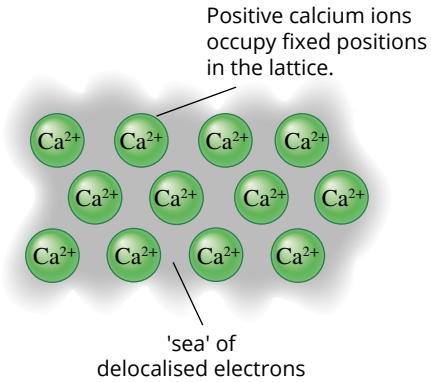
1 Any three of: dense, malleable/ductile, good conductors of heat and electricity, lustrous.

2 **a** silver and gold

b Availability and cost need to be considered; also properties such as malleability and ductility.

3 Ductile means the material is able to be drawn into a wire, while malleable means the material can be shaped by beating or rolling.

4 **a**



b Strong electrostatic forces of attraction are present between Ca^{2+} ions and the delocalised valence electrons.

5 Barium has a high melting point 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 through the entire metal, enabling the transfer of a charge throughout the metal.

Analysis

6 **a** Both potassium and gold have good thermal and electrical conductivity. However, gold has a higher density, and higher melting and boiling points than potassium.

b sodium

c silver

d Sodium and potassium are in group 1. Gold and silver are transition metals.

7 **a** Li atoms have 3 electrons. The electron 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 electron structure is 2,8,2. There are two electrons in the outer shell. The charge of the cation will therefore be 2+.

c Ga atoms have 31 electrons. The electron structure is 2,8,18,3. There are three electrons in the outer shell. The charge of the cation will therefore be 3+.

d Ba atoms have 56 electrons. The electron structure is 2,8,18,18,8,2. There are two electrons in the outer shell. The charge of the cation will therefore be 2+.

Another way to find the answers to this question is by using the group number of the element: Li is in group 1, so it forms 1+ charged cations, Mg and Ba are in group 2, so they form 2+ charged cations, and Ga is in group 13, so it forms 3+ charged cations.

8 The smaller the first ionisation energy of an element, the greater the metallic character of that element.

4.2 Reactivity of metals

KEY QUESTIONS

Knowledge and understanding

1 **a** calcium oxide

b sodium hydroxide and hydrogen

2 B. Sodium is higher in the reactivity series than the other metals, so it is more likely to react.

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

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Analysis

4 Zn > Fe > Au

5 **a** **i** In general, the reactivity of main group metals increases going down a group in the periodic table.

ii In general, the reactivity of main group metals decreases across a period.

b Transition metals tend to be less reactive than the elements in groups 1 and 2

c The trend in metal reactivity corresponds with the periodic trends in first ionisation energy.

6 Potassium is more reactive than sodium because it is lower in group 1 (its valence electrons are further from the protons in the nucleus and less strongly attracted to it, so they are more easily lost in a reaction). Potassium is on the left of calcium in period 4 (so its nucleus has a lower effective nuclear charge and its valence electron is less strongly held within the atom) and so potassium is more reactive than calcium.

4.3 Producing and recycling metals

CASE STUDY: ANALYSIS

E-waste

Analysis

1 Mercury is a neurotoxin. Symptoms of prolonged exposure include tremors, headaches, short-term memory loss, incoordination, weakness, loss of appetite, altered sense of taste and smell, numbness and tingling in the hands and feet, insomnia, and excessive sweating.

Lead is a cumulative toxin which can affect multiple body systems. Lead exposure can permanently damage the brain and impair intellectual development.

Cadmium poisoning symptoms include anaemia and kidney failure. Exposure increases the chance of developing cancer.

2 **a** In a circular economy the ideal is to eliminate waste so that the products in use today are used as raw materials tomorrow. This creates a closed loop.

b While the recycling process described for mobile phones is arguably an improvement on current practice, it does not create closed loops. The process will create waste, including CO_2, and all components of the e-waste, such as plastics, are not being recycled.

3 The aims of a circular economy could be better realised if mobile phones were used for longer before being replaced and if they were designed so that they could be more readily repaired and, ultimately, more readily recycled. Improved rates of recycling of the different components in the phones, perhaps encouraged by legislation, as well as improved recycling techniques would also be desirable.

KEY QUESTIONS

Knowledge and understanding

1 In a linear economy raw materials are used to make a product, and after its use the product is thrown away. A circular economy is based on a model of production and consumption that aims to design out waste and pollution, keep products and materials in use, and regenerate natural systems.

2 Metal recycling meets some of the aims of a circular economy by:

- returning materials to the production cycle
- avoiding using metals as landfill
- saving energy (less energy is used than if the metals are directly extracted from their ores)
- reducing greenhouse emissions compared to mining
- minimising the impact of ore extraction on the environment.

3 Not all products containing metals are recycled; some products are discarded in landfill. It can be uneconomic to recover all the metallic material from some items that contain many different components and only small concentrations of metals. Some waste items are also hard to disassemble. The market price of the metal is a factor, with higher recovery rates for precious metals such as gold and platinum. Furthermore, at the national level, if mineral ores or manufactured products are exported to other countries these materials are not recycled within the country of origin.

4 The market price of gold and platinum is higher than for copper and aluminium, so the economic incentive for their recovery is greater.

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Analysis

5 Use of high-temperature open furnaces can allow lead to enter the immediate environment. In fact, a study has shown that people in developing countries who are engaged in recycling spent batteries in this manner have developed neurological disorders, with high blood lead levels and intellectual disabilities.

Chapter 4 Review

REVIEW QUESTIONS

Knowledge and understanding

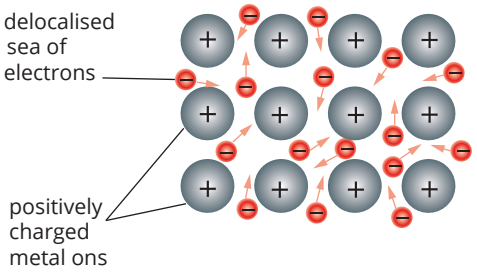
1 **a** aluminium Al, copper Cu, gold Au, iron Fe, silver Ag
b aluminium: period 3, group 13
copper: period 4, group 11
gold: period 6, group 11
iron: period 4, group 8
silver: period 5, group 11
c gold and silver
d copper, gold, iron and silver
e gold

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

3 electrical conductivity

4 C. Mg, Ca and Sr are in the same group (group 2) as Be, so they will have similar properties.

5 The diagram below shows a two-dimensional view of a metallic lattice. The metal atoms are arranged in an ordered manner.
The metal atoms lose their valence electrons and form an 'electron sea'. Electrostatic attraction between the electrons and the metal cations hold the lattice together.



6 **a** **i** valence electrons that are not restricted to a region between two atoms
ii a regular three-dimensional arrangement of a very large number of positive ions (cations)
iii the electrostatic attraction between a lattice of cations and delocalised electrons
b valence (outer-shell) electrons

7 **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, in order for the iron to change from a solid to a liquid.
d 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.

8 Reactive metals, such as sodium and potassium, react with water to form the metal hydroxide and hydrogen gas. The hydrogen gas is observed as bubbles.

9 **a** false (Gold is an example of a metal that does not react with most acids.)
b true
c false (The most reactive metals are located at the bottom of a group in the periodic table.)
d true

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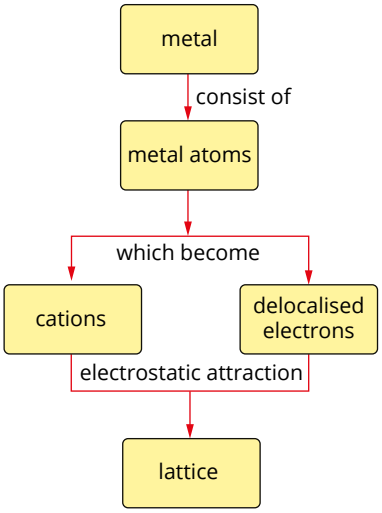
Application and analysis

10 B. 20 in Ca, 18 in Ca^{2+}. Calcium is in group 2, so it has 2 outer-shell electrons which are readily lost to form Ca^{2+} ions.

11 D. Al: 2,8,3; Al^{3+}: 2,8. Aluminium has 13 electrons, so B and D are possibilities on this basis. When Al forms ions, it loses its 3 outer-shell electrons, so there must be 10 electrons in the electronic configuration of Al^{3+}.

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



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

b Na: $1s^2 2s^2 2p^6 3s^1$
K: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
Ca: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

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. Therefore 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.

14 D. Chromium is a transition metal and is less reactive than the group 1 and 2 metals in the other options given.

15 B. Potassium is the element above rubidium in the periodic table. It would therefore be expected to have a similar reactivity (though it will be a little less reactive).

16 The reaction on the left is more vigorous, so 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.

17 Based on the order of metals in the reactivity series, metal A is copper, metal B is sodium and metal C is aluminium.

18 When metals are heated in a smelter at high temperatures they melt without decomposing. When subsequently cooled, they reform a metallic lattice that has the same properties as the original metal. Plastics, however, are likely to burn or decompose at high temperatures. While some plastics (thermoplastics) will melt when heated and can be remoulded into new products, other plastics (thermosetting plastics) do not melt and cannot be recycled in this way.

19 **a** A circular economy aims to reduce waste, pollution (including greenhouse gas emissions), use of natural resources, environmental damage and energy use.

b Recycling metals reduces the cost of raw materials for manufacturers. Furthermore, with a reduced need for mining natural resources there is less environmental damage, and greenhouse gas and pollutant emissions are reduced.

c Some metals are less valuable so there is less financial incentive to recycle products made of these metals. It can be difficult to extract metals that are used in products that are complex and use many other components, such as smartphones. Other metals, such as lithium, are not easily extracted from wastes and considerable energy is required.

20 Metals at the top of the reactivity series, such as sodium, are rarely used in their pure form. They are mainly used in the form of their compounds, e.g. as sodium chloride (salt). Many of these metals are also found as compounds in large quantities in nature. So there is little financial incentive to extract the pure elements. Furthermore, the extraction of the highly reactive metals requires larger quantities of energy and is more expensive.

On the other hand, metals at the bottom of the series, such as platinum and gold, tend to be valuable and require less energy to recycle. They are almost completely recovered after use and then reused.

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Chapter 5 Ionic compounds

5.1 Properties of ionic compounds

CASE STUDY: ANALYSIS

How fluoride ions make tooth enamel harder

Analysis

1 **a** Ionic compounds are hard.
b Ionic compounds are hard.

2 The overall charge on fluoroapatite can be calculated by multiplying the charge on each ion by how many of that ion there are, then finding a total charge for all ions. That is:

Ca^{2+}: $10 \times 2+$ $= 20+$
PO_4^{3-}: $6 \times 3-$ $= 18-$
F^-: $2 \times 1-$ $= 2-$
Total charge: $= 0$

3 SiF_6^{2-}. The formula for sodium fluorosilicate is Na_2SiF_6. The overall charge on two sodium ions is 2+, so the charge on the fluorosilicate ion must be 2–.

KEY QUESTIONS

Knowledge and understanding

1 Ionic compounds are formed when non-metal atoms react with metal atoms. In this process, non-metal atoms **gain** electrons to form **negatively** charged ions called **anions** and metal atoms **lose** electrons to form **positively** charged ions called **cations**. The ions formed pack together in a three-dimensional lattice held strongly together by **electrostatic** forces of attraction.

2 In an ionic compound like sodium chloride, positively and negatively charged ions are held together in a three-dimensional lattice by strong electrostatic forces of attraction. This gives the compound a high melting point, but because the ions in the lattice cannot move, the compound will not conduct electricity in the solid state. If the compound is melted or dissolved in water, however, the lattice breaks down, the ions can now move freely and so the compound can conduct an electric current.

3 **a** Ionic compounds have high melting points and are hard.
b Ionic compounds conduct electricity in the molten state or in solution.
c Ionic compounds are brittle.

4 When an ionic compound is hit with a hammer, the layers of ions within the ionic lattice move relative to each other. This causes ions with like charges to be adjacent to each other and they repel. This causes the lattice to shatter, as seen in the diagrams in Figure 5.1.7 on page 164.

Analysis

5 **a** ionic compounds only
b metals and ionic compounds
c metals only
d metals and ionic compounds
e metals only

6 Substance A is an ionic compound because it has a high melting point and will not conduct electricity at 100°C because it is a solid at that temperature. It will conduct electricity at 1000°C because that is above its melting point and it will be in a liquid state. Substance C is also an ionic compound for similar reasons.

Substance B is a metal because it will conduct electricity both in the solid and molten state. The fact that it is insoluble in water also indicates that it could be a metal.

Some ionic compounds are soluble in water and some are not, so solubility alone cannot be used to decide whether substances are ionic compounds.

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5.2 Formation of ionic compounds

Worked example: Try yourself 5.2.1

WRITING EQUATIONS FOR REACTIONS BETWEEN METALS AND NON-METALS ATOMS

Write an equation for the reaction between calcium and phosphorus atoms. Show the electronic configurations for each element before and after the reaction.

Thinking	Working
Write the symbol and the electronic 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 electronic configuration of the metal ion that will be formed.	Ca^{2+} (2,8,8)
Write the symbol and the electronic 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 electronic configuration of the non-metal ion that will be formed.	P^{3-} (2,8,8)
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) $\longrightarrow$ $3Ca^{2+}$ (2,8,8) + $2P^{3-}$ (2,8,8)

Worked example: Try yourself 5.2.2

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 5.2.1 and 5.2.2 on page 171.

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.	$2 \times 1 = 2$
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

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KEY QUESTIONS

Knowledge and understanding

1 **a** $KBr(s) \longrightarrow K^+(aq) + Br^-(aq)$

b $Ca(NO_3)_2(s) \longrightarrow Ca^{2+}(aq) + 2NO_3^-(aq)$

c $Na_2S(s) \longrightarrow 2Na^+(aq) + S^{2-}(aq)$

d $FeCl_3(s) \longrightarrow Fe^{3+}(aq) + 3Cl^-(aq)$

e $Al_2(SO_4)_3(s) \longrightarrow 2Al^{3+}(aq) + 3SO_4^{2-}(aq)$

2 A. Magnesium hydroxide. Most hydroxides are insoluble. Exceptions do not include $Mg(OH)_2$.

D. Zinc sulfide. Most sulfides are insoluble. Exceptions do not include ZnS.

E. Lead sulfate. Most sulfates are soluble, but $PbSO_4$ is an exception.

3 Potassium and sulfur will react together to form a compound, potassium sulfide. During this process each sulfur atom will **gain two** electron(s) to form a **negatively** charged sulfide ion with the symbol $\mathbf{S^{2-}}$. Each sulfide ion will have the same stable electron configuration as an atom of **argon**, which is the **noble gas** element nearest to it on the periodic table. Also during the reaction, each potassium atom will **lose one** electron(s) to achieve the same stable electron configuration as an atom of **argon**.

4 **a** $ZnCl_2$

b K_2O

c Sr_3N_2

d Na_2CO_3

e $Al_2(SO_4)_3$

f $Zn_3(PO_4)_2$

g CuCl

h Fe_2O_3

i $Cr_2(SO_4)_2$

5 **a** magnesium sulfide

b potassium oxide

c iron(II) sulfate

d barium nitrate

e copper(I) sulfate

f iron(III) cyanide

g gold(III) dichromate

h lead(IV) phosphate

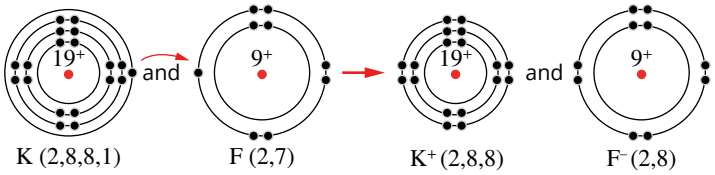
Analysis

6 Cations: calcium, Ca^{2+}, aluminium, Al^{3+}.

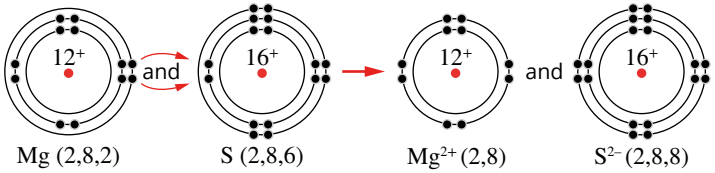
Anions: nitrogen (nitride ion, N^{3-}), fluorine (fluoride ion, F^-) and phosphorus (phosphide ion, P^{3-}).

Metals form cations and non-metals form anions. Metals have low electronegativities and so it is easier for metals to lose electrons than it is for non-metals. Non-metals have high electronegativities and therefore gain electrons more readily than metal atoms.

7 **a**



b



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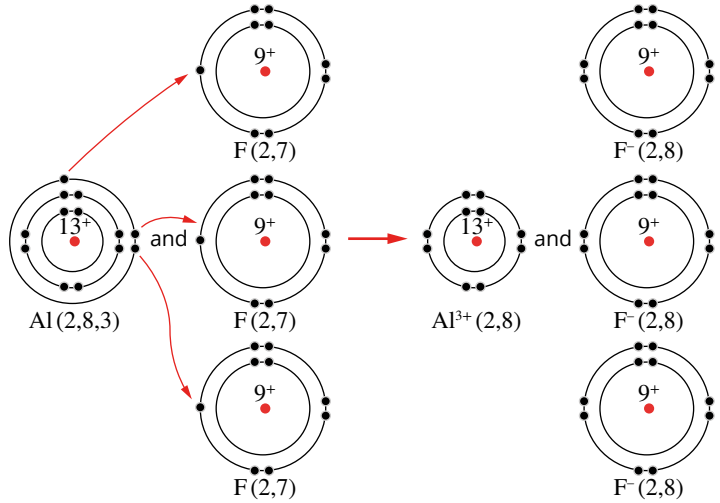
c



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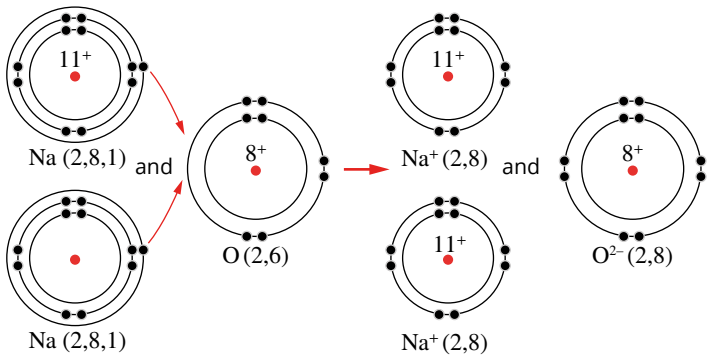
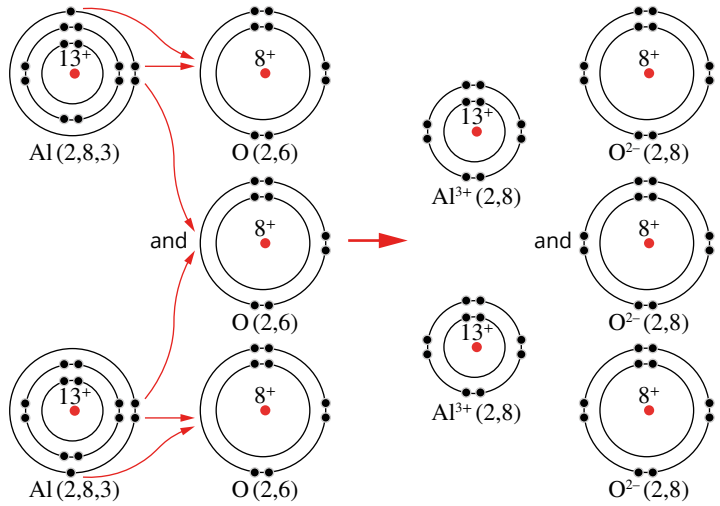


e



8 **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)

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5.3 Precipitation reactions

Worked example: Try yourself 5.3.1

PREDICTING THE PRODUCTS OF A PRECIPITATION REACTION

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 5.3.1 and 5.3.2 on page 178) to complete this question.

Thinking	Working
Identify which ions are produced by each of the ionic compounds in the mixture.	$Na^+(aq)$, $S^{2-}(aq)$, $Cu^{2+}(aq)$ and $NO_3^-(aq)$
Identify which two new combinations of positive and negative ions are possible in the mixture of the solutions.	$Na^+(aq)$ and $NO_3^-(aq)$ $Cu^{2+}(aq)$ and $S^{2-}(aq)$
Use the solubility table to check which, if any, of these combinations will produce an insoluble compound.	Compounds containing sodium ions or nitrate 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 as a precipitate.

Worked example: Try yourself 5.3.2

WRITING EQUATIONS FOR PRECIPITATION REACTIONS

Write a balanced 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
Write an incomplete, unbalanced equation showing the reactants and the precipitate product. Include symbols of state.	$CuSO_4(aq) + NaOH(aq) \longrightarrow Cu(OH)_2(s)$
Add to the equation above the formula of the other compound formed in the reaction.	$CuSO_4(aq) + NaOH(aq) \longrightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$
Balance the equation.	$CuSO_4(aq) + 2NaOH(aq) \longrightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$
Write the formulas 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 5.3.3

WRITING IONIC EQUATIONS FOR PRECIPITATION REACTIONS

Write an ionic equation for the precipitation reaction between solutions of sodium hydroxide and barium nitrate in which the precipitate is barium hydroxide.

Thinking	Working
Write the formula of the precipitate on the right-hand side of the page. Place an arrow to the left of the formula.	$\longrightarrow Ba(OH)_2$
To the left of this formula, add the formulas of the two types of ions that form the precipitate, using the ratio of ions shown in the precipitate formula.	$Ba^{2+} + 2OH^- \longrightarrow Ba(OH)_2$
Add symbols of state to the equation and check that it is balanced.	$Ba^{2+}(aq) + 2OH^-(aq) \longrightarrow Ba(OH)_2(s)$

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CASE STUDY: ANALYSIS

The chemistry of colour

Analysis

1 Sn^{4+}

2 $Na_2CrO_4(aq) + Pb(NO_3)_2(aq) \longrightarrow PbCrO_4(s) + 2NaNO_3(aq)$

3 any soluble cadmium compound, for example: $CdCl_2$, $Cd(NO_3)_2$, $CdSO_4$, $Cd(CH_3COO)_2$

KEY QUESTIONS

Knowledge and understanding

1 **a** $KBr(s) \longrightarrow K^+(aq) + Br^-(aq)$

b $Ca(NO_3)_2(s) \longrightarrow Ca^{2+}(aq)\ 2NO_3^-(aq)$

c $Na_2S(s) \longrightarrow 2Na^+(aq) + S^{2-}(aq)$

d $FeCl_3(s) \longrightarrow Fe^{3+}(aq) + 3Cl^-(aq)$

e $Al_2(SO_4)_3 \longrightarrow 2Al^{3+}(aq) + 3SO_4^{2-}(aq)$

2 **a** **i** $CaCO_3$

ii no precipitate

iii MgS

iv $Fe(OH)_2$

v Ag_3PO_4

b **i** $NO_3^-(aq)$ and $K^+(aq)$

ii no spectator ions

iii $Na^+(aq)$ and $SO_4^{2-}(aq)$

iv $Cl^-(aq)$ and $NH_4^+(aq)$

v $Na^+(aq)$ and $NO_3^-(aq)$

3 **a** **i** magnesium sulfide

ii silver chloride

iii aluminium hydroxide

iv magnesium hydroxide

b **i** $K_2S(aq) + MgCl_2(aq) \longrightarrow MgS(s) + 2KCl(aq)$

ii $CuCl_2(aq) + 2AgNO_3(aq) \longrightarrow 2AgCl(s) + Cu(NO_3)_2(aq)$

iii $AlCl_3(aq) + 3KOH(aq) \longrightarrow Al(OH)_3(s) + 3KCl(aq)$

iv $MgSO_4(aq) + 2NaOH(aq) \longrightarrow Mg(OH)_2(s) + Na_2SO_4(aq)$

4 **a** **i** $NH_4Cl(aq) + AgNO_3(aq) \longrightarrow AgCl(s) + NH_4NO_3(aq)$

ii $Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$

b **i** $Cu(NO_3)_2(aq) + K_2CO_3(aq) \longrightarrow CuCO_3(s) + 2KNO_3(aq)$

ii $Cu^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow CuCO_3(s)$

c **i** $2K_3PO_4(aq) + 3MgSO_4(aq) \longrightarrow Mg_3(PO_4)_2(s) + 3K_2SO_4(aq)$

ii $3Mg^{2+}(aq) + 2PO_4^{3-}(aq) \longrightarrow Mg_3(PO_4)_2(s)$

d **i** $Ca(OH)_2(aq) + FeCl_2(aq) \longrightarrow Fe(OH)_2(s) + CaCl_2(aq)$

ii $Fe^{2+}(aq) + 2OH^-(aq) \longrightarrow Fe(OH)_2(s)$

e **i** $Ba(NO_3)_2(aq) + (NH_4)_2SO_4(aq) \longrightarrow BaSO_4(s) + 2NH_4NO_3(aq)$

ii $Ba^{2+}(aq) + SO_4^{2}(aq)^- \longrightarrow BaSO_4(s)$

f **i** $Pb(CH_3COO)_2(aq) + Na_2SO_4(aq) \longrightarrow PbSO_4(s) + 2NaCH_3COO(aq)$

ii $Pb^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow PbSO_4(s)$

5 **a** NH_4^+, NO_3^-

b K^+, NO_3^-

c K^+, SO_4^{2-}

d Ca^{2+}, Cl^-

e NH_4^+, NO_3^-

f Na^+, CH_3COO^-

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Analysis

6 **a** The precipitate was barium sulfate. Barium and sulfate ions are the only ions present that will combine to form an insoluble compound.

b $Ba(CH_3COO)_2(aq) + K_2SO_4(aq) \longrightarrow BaSO_4(s) + 2KCH_3COO(aq)$

c $Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)$

Chapter 5 Review

REVIEW QUESTIONS

Knowledge and understanding

1 D. Substance A is malleable. Ionic compounds are brittle, not malleable, so response A is incorrect. Substance B does not conduct electricity in the solid or molten state, so is neither a metal or an ionic compound. Substance C will conduct electricity below its melting point, that is, in the solid state, so is probably a metal. It is not an ionic compound. Substance D has a high melting point and will dissolve in water, as some ionic compounds do. D is the correct response.

2 A. The term polyatomic means 'more than one atom'. There are three ions in the responses, but the OH^- ion, which contains an oxygen and a hydrogen atom, is the only polyatomic one.

3 D. The two new compounds that could form when solutions of barium nitrate and sodium sulfate are mixed are barium sulfate and sodium nitrate. An inspection of the solubility tables (Tables 5.3.1 and 5.3.2) show that only barium sulfate is insoluble in water, so this is the precipitate. The equation for the precipitation reaction between barium nitrate and sodium sulfate is:

$$Ba(NO_3)_2(aq) + Na_2SO_4(aq) \longrightarrow BaSO_4(s) + 2NaNO_3(aq)$$

Sodium ions and nitrate ions are both in solution, so they are the spectator ions in this reaction.

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 Statement 1: Ionic compounds are able to conduct electricity in the molten state.

Statement 2: Ionic compounds are not able to conduct electricity in the solid state.

Statement 3: Ionic compounds are hard and have high melting points..

Statement 4: Ionic compounds are brittle.

6 **a** 2,8

b 2,8

c 2,8

d 2,8

7 **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–.

8 According to Tables 5.2.1 and 5.2.2, Fe^{2+} would react in the presence of PO_4^{3-} and S^{2-} to produce the insoluble compounds $Fe_3(PO_4)_2$ and FeS. Iron(II) bromide and iron(II) sulfate are both soluble in water so would not form a precipitate.

9 **a** $CuNO_3$

b CrF_2

c K_2CO_3

d $Mg(HCO_3)_2$

e $Ni_3(PO_4)_2$

10 **a** ammonium carbonate

b copper(II) nitrate

c copper(II) nitrite

d chromium(III) bromide

e tin(II) dihydrogen phosphate

f lead(IV) hydrogen sulfite

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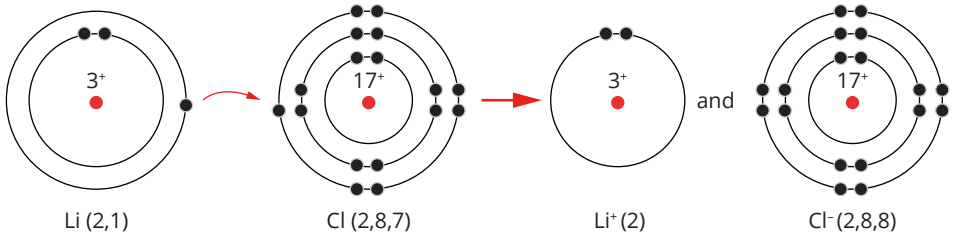
11 a



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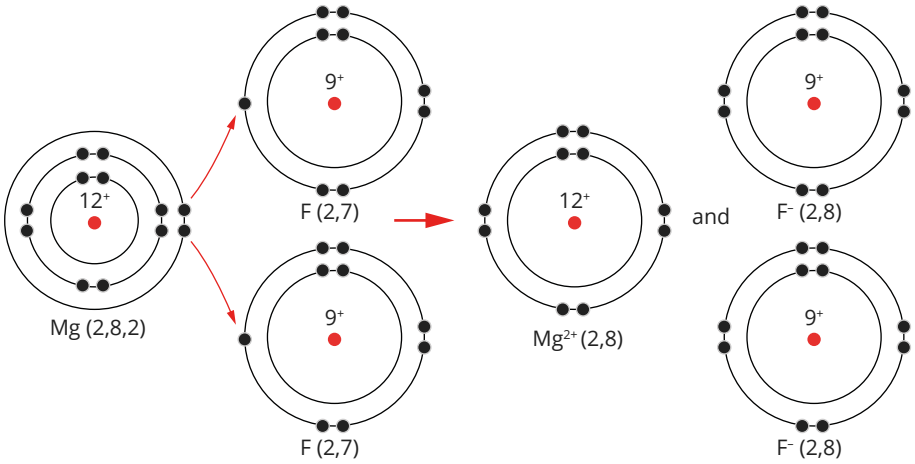

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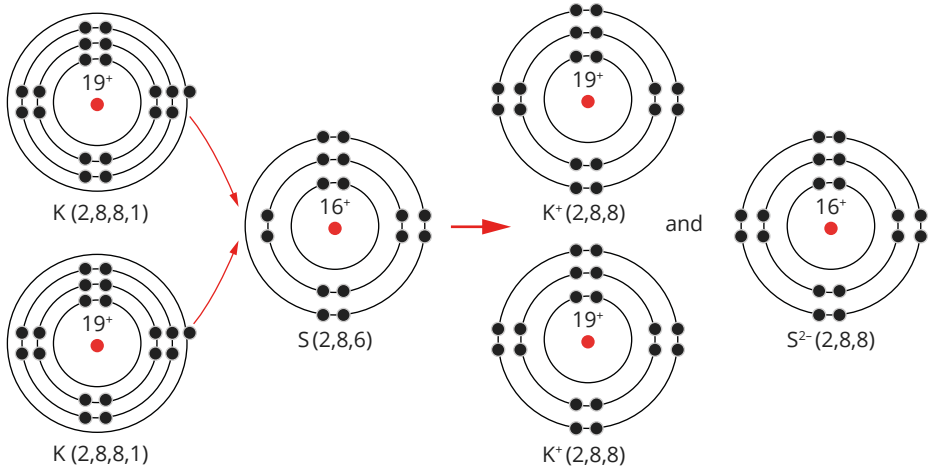

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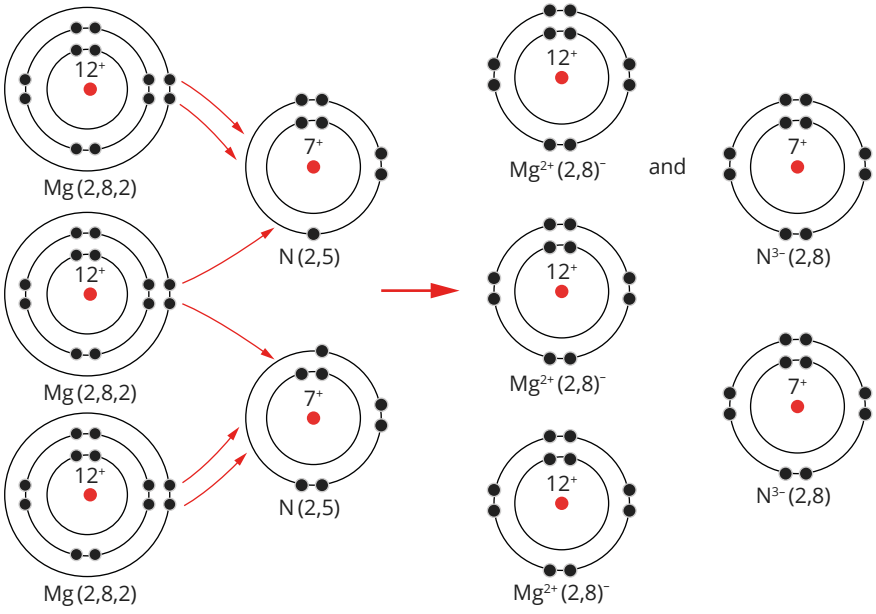

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12 **a** $Mg(2,8,2) + 2Cl(2,8,7) \longrightarrow Mg^{2+}(2,8) + 2Cl^{-}(2,8,8)$

b $2Al(2,8,3) + 3O(2,6) \longrightarrow 2Al^{3+}(2,8) + 3O^{2-}(2,8)$

c $3Na(2,8,1) + P(2,8,5) \longrightarrow 3Na^{+}(2,8) + P^{3-}(2,8,8)$

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

14 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.

15 **a** barium sulfate

b none

c lead(II) sulfate

d none

Application and analysis

16 B. Atom X, being a non-metal will gain electrons. It already has 6 electrons in its outermost shell and, according to the octet rule, will gain two, giving it a charge of 2–. Atom Y, being a metal will lose electrons. It has 3 electrons in its outermost shell and according to the octet rule, will lose three, giving it a charge of 3+. When Y^{3+} combines with X^{2-}, they will do so in a ratio of 2 : 3, giving a compound with the formula Y_2X_3. Metals are always written first in a chemical formula.

17 B. In the chemical formula of an ionic compound the symbol for the metal is always written first, so in this case A is a metal. The ratio of metal to non-metal is 2 : 1, so the charge on A is probably 1+ and on B, 2–. (It could also be a ratio of 2+ and 4–, etc). Both responses B and D meet this requirement, but only a sulfide ion and a potassium ion will have the same electron configuration as an argon atom.

18 **a** K^{+}, F^{-}, Ca^{2+} and O^{2-}

b Both ions in potassium fluoride are singly charged. Both ions in calcium oxide have double charges. The forces of attraction between the two double charged ions in calcium oxide will be much stronger than that between the single-charged ions of potassium fluoride. The melting point of calcium oxide (2572°C) will therefore be higher than that of potassium fluoride (858°C)

19 **a** Agree. In metals the outermost shell electrons are delocalised throughout the metallic lattice. In ionic compounds the metal atom loses its outermost electron(s) to the non-metal atoms.

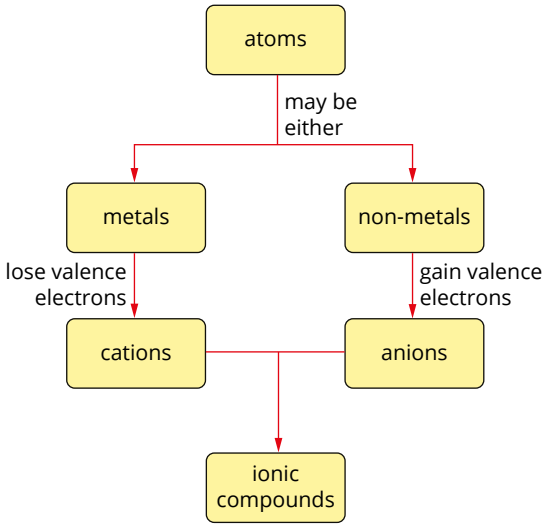
b Disagree. Particles of opposite charge will attract, not repel each other.

c Disagree. In ionic compounds the negatively charged ions are held in fixed positions between positively charged ions. In metals, however, the negatively charged outer shell electrons are free to move, not held in fixed positions.

d Disagree. In metals it is the freely moving delocalised electrons that conduct the electric current. In molten ionic compounds, however, both positively and negatively charged ions move, therefore conducting the electric current.

e Agree. When metals are drawn into a wire layers of positively charged metal ions slide over each other as the metal 'stretches'. This means that positively charged metal ions are moving past each other. They do not repel each other, however, because electrons from the freely moving delocalised electron 'sea' move in between the ions, thus holding them together.

20 A possible answer is shown.



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21

Element	Electrons lost or gained when forming an ion?	Noble gas with same electron configuration as ion formed
phosphorus	gained	argon
lithium	lost	helium
oxygen	gained	neon
aluminium	lost	neon
potassium	lost	argon
bromine	gained	krypton
sulfur	gained	argon

22 **a** CD_3
b EF
c G_3H
d KL

23 **a** $NH_4Cl(aq) + AgNO_3(aq) \longrightarrow NH_4NO_3(aq) + AgCl(s)$
$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$
b $FeCl_2(aq) + Na_2S(aq) \longrightarrow FeS(s) + 2NaCl(aq)$
$Fe^{2+}(aq) + S^{2-}(aq) \longrightarrow FeS(s)$
c $Fe(NO_3)_3(aq) + 3KOH(aq) \longrightarrow 3KNO_3(aq) + Fe(OH)_3(s)$
$Fe^{3+}(aq) + 3OH^-(aq) \longrightarrow Fe(OH)_3(s)$
d $CuSO_4(aq) + 2NaOH(aq) \longrightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$
$Cu^{2+}(aq) + 2OH^-(aq) \longrightarrow Cu(OH)_2(s)$
e $Ba(NO_3)_2(aq) + Na_2SO_4(aq) \longrightarrow BaSO_4(s) + 2NaNO_3(aq)$
$Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)$

24

	NaOH	**$(NH_4)_3PO_4$**	**NaI**	**$MgSO_4$**	**$BaCl_2$**
$Pb(NO_3)_2$	$Pb(OH)_2$	$Pb_3(PO_4)_2$	PbI_2	$PbSO_4$	$PbCl_2$
KI					BaI_2
$CaCl_2$	$Ca(OH)_2$	$Ca_3(PO_4)_2$		$CaSO_4$	
Na_2CO_3				$MgCO_3$	$BaCO_3$
Na_2S				MgS	BaS

25 Note that in the answers to parts **b** and **d** of this question, square brackets have been placed around the ferrocyanide, $Fe(CN)_6^{4-}$ ion. This has been done to help distinguish this ion from the CN^- ion which is part of the ferrocyanide ion. It is acceptable to use only curved brackets in your answers to parts **b** and **d** of the question.
a $Ca(ClO_4)_2$
b $Al_4[Fe(CN)_6]_3$
c $Fe(ClO_4)_3$
d $(NH_4)_4[Fe(CN)_6]$

26 **a** **i** 2+
ii 2–
iii 3+
b **i** X_3N_2
ii PbY_2
iii $Z_2(Cr_2O_7)_3$
iv Z_2Y_3

27 Possible combinations are:
a $MgCl_2$ or MgF_2 (2+ cation with 1– anion)
b NaCl or CaS (cation and anion charge cancels out; e.g. 1+ cation with a 1– anion)
c Na_2O or K_2S (1+ cation with 2– anion)
d Na_3N or Li_3P (1+ cation with 3– anion)
e $AlCl_3$ or AlF_3 (3+ cation with 1– anion)

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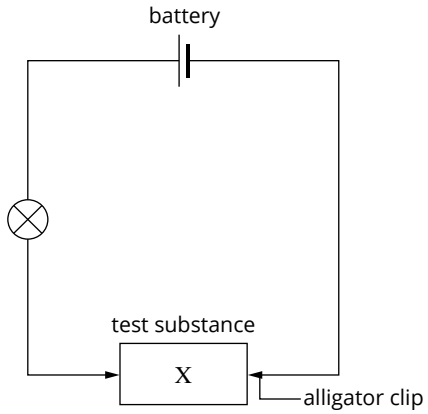
28 **a** Assemble equipment to test conductivity. Add a globe to a circuit containing a power source, such as a battery. When the electrodes are touching the solid magnesium chloride, the globe will not light up.



b Dissolve some solid sodium chloride (about 1 teaspoon per 200 mL) in deionised water. Using the same equipment, place the two electrodes in the solution, but don't allow them to touch; the globe will glow.

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

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Chapter 6 Separation and identification of components of mixtures

6.1 How substances dissolve

CASE STUDY: ANALYSIS

Vitamin C and vitamin A: Similar but different

Analysis

1 Vitamin D molecules would be largely non-polar, whereas those of vitamin B would be polar.

2 Whereas vitamin C is excreted in urine on a daily basis, vitamin A is not excreted in urine, but is stored in body fat. The concentration of vitamin A will therefore increase over time if relatively large quantities of it are being consumed and hypervitaminosis can develop.

3 Since vitamin E molecules have a long non-polar hydrocarbon chain and only one polar –OH group, the vitamin is likely to be relatively non-polar overall and fat soluble.

KEY QUESTIONS

Knowledge and understanding

1 **a** solute
b solution
c solvent
d solute

2 Nitrogen gas (**b**), ethane (**e**) and tetrachloromethane (**f**) molecules are all symmetrical and non-polar and so will be insoluble in water. Molecules in canola oil (**d**) have long, non-polar hydrocarbon chains so it will also be insoluble in water. Glucose (**c**) and ethanoic acid (**g**) are all polar molecules and will be soluble in water. Nitric acid (**a**) is a strong acid that ionises in water.

3 D and F. Glycerol and propanol molecules both contain polar hydroxyl groups, which can form hydrogen bonds with water and so they will dissolve. Metals such as lead do not dissolve in water. Some metals, such as sodium, can react with water, but that is not just a dissolving process. Nitrogen gas and ethane are non-polar molecules, so do not form hydrogen bonds with water and are insoluble. Like sodium chloride, potassium chloride is soluble in water, but it dissolves in water by dissociating, not by forming hydrogen bonds with water.

4 Methanol dissolving: $CH_3OH(l) \xrightarrow{H_2O(l)} CH_3OH(aq)$

Nitric acid dissolving: $HNO_3(l) + H_2O(l) \longrightarrow H_3O^+(aq) + NO_3^-(aq)$

5 Sodium chloride is an ionic compound consisting of sodium and chloride **ions**. In solid sodium chloride, the two different ions form a **lattice**, which is held together by an **electrostatic** force of attraction called an **ionic bond**. When water is added to solid sodium chloride, water molecules attach themselves to ions in the solid by forces of **ion–dipole attraction**. When the lattice breaks up and the solid dissolves, the sodium and chloride ions are now surrounded by water molecules and are said to be **hydrated**.

6 **a** $MgSO_4(s) \xrightarrow{H_2O} Mg^{2+}(aq) + SO_4^{2-}(aq)$
b $Cu(NO_3)_2(s) \xrightarrow{H_2O} Cu^{2+}(aq) + 2NO_3^-(aq)$
c $(NH_4)_2S(s) \xrightarrow{H_2O} 2NH_4^+(aq) + S^{2-}(aq)$
d $Al_2(SO_4)_3(s) \xrightarrow{H_2O} 2Al^{3+}(aq) + 3SO_4^{2-}(aq)$
e $Na_3PO_4(s) \xrightarrow{H_2O} 3Na^+(aq) + PO_4^{3-}(aq)$

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Analysis

7

	Ionising molecular compounds	Ionic compounds	Non-ionising molecular compounds
Examples	sulfuric acid, H_2SO_4	$Ca(OH)_2$	propanol, C_3H_7OH
Type of particles present before dissolving occurs	molecules	ions	molecules
Type of particles present after dissolving occurs	ions	ions	molecules
Equation for dissolving process	$H_2SO_4(l) + 2H_2O(l) \rightarrow 2H_3O^+(aq) + SO_4^{2-}(aq)$	$Ca(OH)_2(s) \xrightarrow{H_2O} Ca^{2+}(aq) + 2OH^-(aq)$	$C_3H_7OH(l) \xrightarrow{H_2O} C_3H_7OH(aq)$

8 **a** Hydrogen chloride only
b Neither methanol nor hydrogen chloride
c Both methanol and hydrogen chloride
d Hydrogen chloride only.

6.2 Principles of chromatography

Worked example: Try yourself 6.2.1

CALCULATING R_f VALUES

Calculate the R_f value of the blue component in Figure 6.2.3 on page 204.

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 (blue component) $= \frac{3}{10}$ $= 0.3$

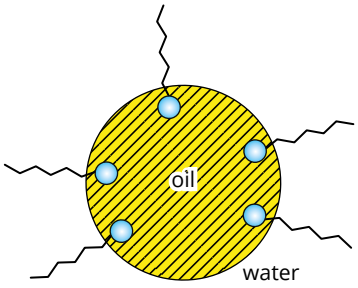
CASE STUDY: ANALYSIS

Investigating the ingredients of whipped cream

Analysis

1 **a** The stationary phase is silica and the mobile phase is a mixture of liquids, pentane, hexane and diethyl ether.
b Ultraviolet light is used to detect the spots on the TLC plates because the components are colourless and invisible. They fluoresce and become visible under ultraviolet light.
c It is probable that the scientists found they obtained better separation of the components on the TLC plates using a mixture of solvents.

2



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KEY QUESTIONS

Knowledge and understanding

1 Ethanol is the mobile phase in this example. The pigments contain the components to be separated and the paper is the stationary phase.

2

Components	the different compounds in the mixture, which can be separated by chromatography
Mobile phase	the solvent that moves over the stationary phase in chromatography
Stationary phase	the components of a mixture undergo adsorption to this phase
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

Analysis

3 For parts **a**, **b**, **c**, it would be useful to set your answers out in a table. Measurements may vary slightly depending on the screen or book you are measuring from, so this is a model answer only. Yours may vary slightly.

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 probably 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.
(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 (phenacetin is colourless but fluoresces 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 probably contains phenacetin.

Chapter 6 Review

REVIEW QUESTIONS

Knowledge and understanding

1 A solution is most likely to form when the polarity 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.

2 Propanol is a polar molecule with a hydroxyl group. The hydroxyl group is able to form hydrogen bonds with water molecules and will therefore dissolve in water. Propane is a non-polar hydrocarbon and so can only interact with other molecules by weak dispersion forces. It is not able to form hydrogen bonds with other molecules and so will not dissolve in water.

3 **a** dissociation
b **i** $Cu^{2+}(aq)$, $NO_3^-(aq)$
ii $Zn^{2+}(aq)$, $SO_4^{2-}(aq)$
iii $NH_4^+(aq)$, $PO_4^{3-}(aq)$

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4 **a** $CH_3OH(l) \xrightarrow{H_2O(l)} CH_3OH(aq)$

b $C_{12}H_{22}O_{11}(l) \xrightarrow{H_2O(l)} C_{12}H_{22}O_{11}(aq)$

5 **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 $Li_2SO_4(s) \xrightarrow{H_2O(l)} 2Li^{+}(aq) + SO_4^{2-}(aq)$

6 Ion–dipole interaction. A potassium ion has a positive charge and so will attract the negative part (the oxygen atom) of a polar water molecule. Several water molecules will orient themselves around the potassium ion so that their oxygen atoms, which carry a partial negative charge, are closer to the potassium ion than the hydrogen atoms

7 C. Methanol, CH_3OH, contains a polar hydroxyl group which will form hydrogen bonds with water molecules. Dispersion forces will also occur between all of the molecules in the solution. Methanol is a molecular compound and its atoms bond to each other by covalent bonds. Methanol does not ionise when it dissolves in water so no ion–dipole interactions will occur.

8

i Ionic compound	**e** Dissolves in water by dissociating, then forming ion–dipole bonds with water
ii Compound composed of polar molecules with –OH groups	**d** Dissolves in water by forming hydrogen 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	**b** Dissolves in water by ionising then forming ion–dipole bonds with water
iv Non-polar molecular compound	**c** Does not dissolve in water
v Compound composed of covalent molecules with a large non-polar end and one –OH group.	**a** Does not dissolve in water due to the size of the molecule

9 Propan-1-ol has a polar hydroxyl group which will form hydrogen bonds to water molecules. These hydrogen bonds are strong enough to overcome the intermolecular bonds holding propan-1-ol molecules together, so the two liquids will mix. Once this happens, each propan-1-ol molecule will be surrounded by water molecules.

10 Thin layer chromatography is a technique that allows you to determine the **composition** and **purity** of different types of substances. In this technique, a thin layer of a solid **stationary phase** is applied to a plate. The **components** of the sample are carried over the surface of the stationary phase by the solvent, or **mobile phase**. The components separate, depending on the relative attractions of compounds towards the two phases. The individual components are seen as **spots** on the plate, which can be identified by calculating their $\boldsymbol{R_f}$ **values**.

11 **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.

Application and analysis

12 **a** I_2, CH_4 and C_2H_4 are non-polar covalent molecules. They do not dissolve well in water, which has polar molecules.

b $C_6H_{12}O_6$ and C_3H_7OH are polar molecules. They contain the polar –OH group and so are able to form hydrogen bonds with water. They dissolve in water without ionising.

c HI and HNO_3 contain polar molecules which dissolve by ionising.

13 Hexane. Since benzene is a non-polar solvent, it will dissolve non-polar solutes best. KCl is an ionic soid which therefore would not dissolve well. Glycerol and ethanol are relatively small molecules with polar –OH groups, so they would also not dissolve well in benzene.

14 **a** methanol, butanol, pentane

b pentane, butanol, methanol

c Methanol is a small molecule with a hydroxyl group. Methanol can form hydrogen bonds to other molecules and so will be readily soluble in water. It will not dissolve in non-polar solvents such as hexane, because the forces of attraction between methanol molecules will be too strong for the hexane molecules to break them. Pentane is a non-polar molecule and is unable to form hydrogen bonds to other molecules, so it is not soluble in water. It is, however, soluble in compounds that are non-polar like itself, and so will be readily soluble in hexane. Butanol has a hydroxyl group and so can form hydrogen bonds to water, but it also has quite a long non-polar hydrocarbon chain, which reduces its solubility in water but enhances its solubility in non-polar solvents such as hexane. Butanol will be partially soluble in both water and hexane.

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15 **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, but 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. These attractions are strong enough to overcome the attraction between the sodium ions and chloride ions in the solid NaCl lattice.

16 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.

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

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

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

17 R_f = distance dye has moved/distance solvent front has moved.

Blue dye $R_f = \frac{7.5}{9.0} = 0.83$

Red dye $R_f = \frac{5.2}{9.0} = 0.58$

18 **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 2

d B: blue; C: green. They can be identified on the basis of their colour and R_f values.

e purple

f 0.63; 0.13

19 A component that appears near the top of a paper chromatogram is likely to be more soluble in the mobile phase and adsorb less strongly to the stationary phase than a component near the bottom of the chromatogram.

20 **a** $R_f\,(\text{purple}) = \frac{\text{distance from origin of component}}{\text{distance moved from origin by solvent}} = \frac{\text{distance from origin of component}}{8} = 0.61$

So distance from origin of purple component = $0.61 \times 8 = 4.9\,\text{cm}$

$R_f\,(\text{yellow}) = \frac{\text{distance from origin of component}}{\text{distance moved from origin by solvent}} = \frac{\text{distance from origin of component}}{8} = 0.19$

So distance from origin of yellow component = $021 \times 8 = 1.7\,\text{cm}$

Distance of the two components apart = $4.9 - 1.7 = 3.2\,\text{cm}$

b $R_f\,(\text{green}) = \frac{\text{distance from origin of component}}{\text{distance moved from origin by solvent}} = \frac{5}{\text{distance moved from origin by solvent}} = 0.30$

So distance moved from origin by component = $\frac{5}{0.30} = 16.7\,\text{cm}$

Substitute this value in the R_f expression for the orange component.

$R_f\,(\text{orange}) = \frac{\text{distance from origin of component}}{\text{distance moved from origin by solvent}} = \frac{\text{distance from origin of component}}{16.7} = 0.80$

So distance moved from origin by component = $0.80 \times 16.7 = 13\,\text{cm}$

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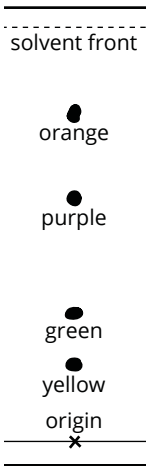
c



21 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 Two ways of visualising the spots are:
- viewing the chromatogram under UV light
- spraying the finished chromatogram with a compound that causes the amino acids to fluoresce.

c $R_f = \frac{\text{distance from origin of component}}{\text{distance moved from origin by solvent}} = \frac{0.6}{5} = 0.12$

d Leucine. The component in a mixture bonded least strongly to the stationary phase will move the greatest distance from the origin.

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Unit 1 Area of Study 1

How do the chemical structures of materials explain their properties and reactions?

Multiple-choice questions

1 C. All atoms of the same element have the same number of protons in their nuclei. A is incorrect as elements can have isotopes. B is incorrect for the same reason – isotopes of an element are not identical. D is incorrect as all atoms of zinc contain more neutrons than protons.

2 B. The reactivity of metals increases down a group and decreases across a period. Therefore, the most reactive metal will be the one closest to the bottom left of the periodic table, and this is K.

3 C. Ionic bonding involves the transfer of electrons, not the sharing of electrons, from a metal to a non-metal. Therefore, answer C is the incorrect choice. A, B and D are all correct descriptions of ionic compounds.

4 D. Mass number is the sum of the number of protons and neutrons. So for this ion the mass number is 11 + 12 = 23. The ion has 1 fewer electrons than protons, so it has a charge of 1+. A and B are incorrect because the mass number is incorrectly given as the number of protons. C is incorrect, as the ion has lost a negative charge and therefore has more positive than negative charges.

5 C. A sulfur atom has 16 electrons, so a S^{2-} ion has 18 electrons. A is not correct as it shows the ground state electronic configuration for an uncharged atom of sulfur. B is not correct as it contains the wrong number of electrons. D is not correct because the 3*p*-subshell contains 6 electrons in the ground state.

6 C. Any *d*-subshell is made up of 5 orbitals. Each orbital can contain a maximum of 2 electrons. So a *d*-subshell can contain a maximum of 10 electrons in total. The other options are all incorrect as they do not contain the correct number of orbitals and electrons.

7 B. Isotopes have the same atomic number but different mass numbers. None of the other options share an atomic number, therefore they cannot be the same element.

8 B. X has an atomic number of 19, so it has 19 protons in the nucleus. Its electronic configuration is $1s^2 2s^2 2p^6 3s^2 3p^6$, so it has 18 electrons. Therefore, as it has 1 more proton than electrons, it will have a positive charge. A is incorrect as W represents chlorine, which is not a noble gas. C is incorrect as Y would be in group 2. D is incorrect as the number of protons and electrons in Z is equal, therefore it is not an ion.

9 D. Element X would have 4 valence electrons, so needs 4 electrons to achieve a stable outer shell configuration. Element Y has 6 valence electrons, hence needs 2 to achieve stability. Therefore, one X atom would bond with 2 Y atoms. Options A, B and C are all incorrect as they would not achieve a stable compound.

10 A. The electrostatic forces of attraction between ions will determine the melting point of an ionic solid. These electrostatic forces of attraction are proportional to the charge on the ions. MgO is the only compound listed in which both ions are doubly charged (Mg^{2+} and O^{2-}) and these will exert the strongest attraction. So MgO will be expected to have the highest melting point.

11 C. Sucrose is a solid that dissolves in water by forming hydrogen bonds with water. It does not dissociate or ionise, but forms an aqueous solution with hydrogen bonds between sucrose molecules and water molecules. A is incorrect as it shows sugar acting as a base. B is incorrect as it shows sugar acting as an acid. D is incorrect as dissolved sugar is not in the (l) state.

12 A. Diatomic molecules, e.g. HF and H_2 can only be linear. CO_2 is also linear because it has two electron groups (each a double bond). These repel and so are oriented as far from one another as possible, i.e. at 180°.

H_2S and H_2O are bent. The sulfur or oxygen atom has two bonding pairs of electrons and two non-bonding pairs. The four pairs of electrons take on a distorted tetrahedral arrangement that results in a bent structure. CH_4 has 4 C–H bonds in a tetrahedral structure. NH_3 is a triangular pyramid. The only option that matched these shapes is A.

13 A. F_2 is a non-polar molecule and so only has dispersion forces between molecules.

CH_3F has a polar C–F bond, so it has intermolecular dipole–dipole interactions as well as dispersion forces. However, these are not as strong as hydrogen bonds in this case because the F is not bonded directly to the H atom as in HF.

14 D. Hydrogen chloride is a covalent molecular compound. Its atoms are held together by covalent bonds that break when HCl dissolves in water. The new bonds formed are ion–dipole bonds between the chloride ions and water. The protons released from HCl covalently bond with water molecules to form H_3O^+ ions, which then form hydrogen bonds with neighbouring water molecules.

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15 B. No precipitate would be formed. All ion species will remain in solution.
Reaction A would form a precipitate of $BaSO_4$.
Reaction C would form a precipitate of $CuCO_3$.
Reaction D would form a precipitate of $PbSO_4$.

Short-answer questions

16 a $1s^2 2s^2 2p^6 3s^2$

b i 1 pm = 10^{-12} m, 1 nm = 10^{-9} m, so 160 pm = 0.160 nm

ii Na atom would have a larger radius as its effective nuclear charge is lower, pulling its electrons less strongly toward the nucleus.

c i It is a lattice of positively charged magnesium ions surrounded by a 'sea' of valence electrons. The lattice is held together by the electrostatic attraction between the valence electrons and cations.

ii The electrons are not localised; they are free to move and so can conduct an electric current.

d i Observations: bubbles of gas evolved and increase in temperature
$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$
or $Mg(s) + 2H^+(aq) \longrightarrow Mg^{2+}(aq) + H_2(g)$

ii Any suitable example such as K, Na, Ca.

17 a i Ca

ii Ar

iii C

iv Na or Mg

v Li

vi N

vii F

b Elements in group 1 are metals, which react by giving away electrons. Chemical reactivity increases because the outer-shell electrons are further from the nucleus as one moves down the group, and so are more readily released in a reaction, which makes the metals more reactive.

18 a Chlorine is on the right side of the periodic table and sodium is on the left. Atomic radius decreases across a period because the increasing effective nuclear charge pulls the outer-shell electrons more tightly to the nucleus, causing the radius of the atom to decrease.

b Fluorine is further to the right on the periodic table than lithium and effective nuclear charge increases from left to right across the periodic table. As effective nuclear charge increases, the electrons are held more tightly to the nucleus and more energy is required to remove the first one.

c Ba and Be are in the same group, with Be higher than Ba. Going down a group, the atom size is increasing, meaning the outer-shell electrons are further from the nucleus. The outer electrons of Be are, therefore, held more tightly and are less readily released.

d The *s*-block elements have an *s*-subshell as their outer occupied electron subshell. The *s*-subshell can take 1 or 2 electrons, so the block is only 2 groups wide.

19 a $Pb(NO_3)_2(aq) + 2KCl(aq) \rightarrow PbCl_2(s) + 2KNO_3(aq)$

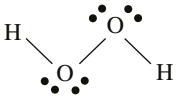
b $Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$

c i lead(II) chloride

ii potassium ions and nitrate ions

20 a The forces between the molecules of ice are intermolecular hydrogen bonds. The bonds between the H and O atoms within the water molecules are covalent bonds. Covalent bonds are much stronger than hydrogen bonds and so require much more energy, and thus a higher temperature, to break.

b Ethyne has the structure CHCH. The C atoms have a triple bond between them, each using three of their four valence electrons to form the triple bond. The fourth valence electron of each carbon atom forms a covalent bond with a hydrogen atom. There are no lone pairs and only two electron groups on the carbons. So these adopt a linear arrangement.
Hydrogen peroxide has the following structure:



Each oxygen has six valence electrons. Two are involved in bonding leaving two pairs of non-bonding electrons. So there are four electron groups around each oxygen atom. These assume a tetrahedral arrangement to minimise repulsion. So the molecule is not linear.

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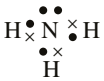
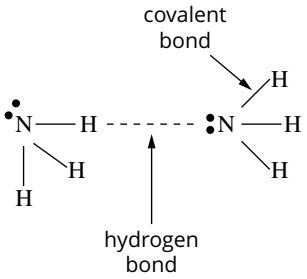
21 a

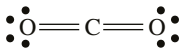


b The four electron pairs form a tetrahedral arrangement around the atom due to the repulsion of the electron pairs. The result is a triangular pyramid shaped molecule.



c i

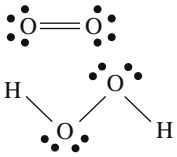
ii



d i N_2 has no polar bonds, so the intermolecular bonds are weak dispersion forces.
Although the bonds within CO_2 are polar, the molecule overall is symmetrical, so the dipoles cancel and the overall molecule is non-polar. Therefore, the only intermolecular forces are dispersion forces.

ii Nitrogen is a highly electronegative element; consequently, the bonds between the atoms of nitrogen and hydrogen are highly polarised. The ammonia molecule is a dipole because its shape is not symmetrical. There is an electrostatic attraction between the nitrogen atom of one ammonia molecule and the hydrogen atom of a nearby ammonia molecule. These attractions between these dipoles are known as hydrogen bonds.

22 a



b The intramolecular bonds in oxygen are stronger because they are double covalent bonds, as opposed to the single covalent bonds in hydrogen peroxide.

c i dispersion forces

ii Hydrogen bonds, which are significantly stronger than the dispersion forces between oxygen molecules. Dispersion forces also exist between molecules of hydrogen peroxide. Because the hydrogen peroxide molecule is larger than that of oxygen, the dispersion forces between hydrogen peroxide molecules are stronger than those between oxygen molecules.

d Oxygen has six valence electrons, so achieves stability when it forms two covalent bonds. Nitrogen only has five valence electrons, so forms three covalent bonds to achieve stability.

23 a Allotropes are different physical forms of the same element.

b Both consist of carbon atoms covalently bonded to other carbon atoms.

c Diamond is a three-dimensional lattice in which each carbon atom is covalently bonded to four other carbon atoms in a tetrahedral configuration. So strong bonding extends throughout the lattice.
Graphite consists of layers of carbon atoms in which each atom is covalently bonded to three other carbon atoms, making strong layers. There are weaker dispersion forces between the layers. The one electron of each carbon atom not involved in bonding is delocalised.

d The delocalised electrons in graphite are free to move and conduct electricity. In diamond, each carbon atom is bonded with four other carbon atoms so there are no free electrons.

e Because of the weak bonding between the layers of graphite, the layers can slide over one another and thus can slide onto a page.

24 a i $MgCl_2(s) \xrightarrow{H_2O(l)} Mg^{2+}(aq) + 2Cl^-(aq)$

ii ionic bonds

iii ion–dipole interactions

b i $C_2H_5OH(l) \xrightarrow{H_2O(l)} C_2H_5OH(aq)$

ii hydrogen bonds and dispersion forces

iii hydrogen bonds and dispersion forces

c The magnesium chloride solution would be the better conductor as many charged particles in the form of $Mg^{2+}(aq)$ and $Cl^-(aq)$ ions are present in the solution after the $MgCl_2$ has dissolved; the dissolved ethanol does not contain charged particles.

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25 a i no change
ii A white precipitate, magnesium carbonate, is formed.
iii A green precipitate, copper (II) carbonate, is produced.
iv no change
v no change
vi no change

b ii $Mg^{2+}(aq) + CO_3^{2-}(aq) \rightarrow MgCO_3(s)$
iii $Cu^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CuCO_3(s)$

c

	Sodium carbonate	Potassium nitrate	Magnesium nitrate	Copper(II) nitrate
Sodium carbonate	x	x	white precipitate	green precipitate
Potassium nitrate	x	x	x	x
Magnesium nitrate	white precipitate	x	x	x
Copper(II) nitrate	green precipitate	x	x	x

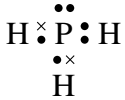
Sodium carbonate solution will react with two of the other solutions. It will form a white precipitate with one solution, and a green precipitate with another. No reaction will be observed with the third solution.
Potassium nitrate solution will not react with any of the other three solutions.
Magnesium nitrate solution will only react with one of the other solutions, forming a white precipitate.
Copper(II) nitrate solution will only react with one of the other solutions to form a green precipitate.

26 The R_f value for the blue colour is:

$$\frac{\text{distance travelled by component}}{\text{distance travelled by solvent front}} = \frac{9}{12} = 0.75$$

When the blue spot has moved 15 cm, the solvent will have moved $\frac{15.0}{0.75} = 20$ cm

27 a



b PH_3 is a polar molecule. It is asymmetrical, so it has an overall molecular dipole. (The P–H bond is not particularly polar, but the overall asymmetry of the molecule results in an overall molecular dipole.)

c In a PH_3 molecule, there are four pairs of electrons around the central P atom. These electron pairs adopt a tetrahedral geometry. Since there is one lone pair, the molecular shape is a triagonal pyramid.

d $M_r(PH_3) = 31.0 + 3 \times 1.0 = 34.0$. $M_r(NH_3) = 14.0 + 3 \times 1.0 = 17.0$.

e Since H is bonded to N in ammonia, NH_3 molecules are able to form hydrogen bonds between molecules. Between PH_3 molecules there are dipole–dipole forces. Hydrogen bonding is stronger than dipole–dipole forces, so NH_3 has the higher melting point.

f ${}^{32}_{15}P$

g $1s^2 2s^2 2p^6 3s^2 3p^3$

28 a A: 0.60
B: 0.48
C: 0.38
D: 0.20
E: 0.10

b A: leucine and/or isoleucine
B: β-phenylalanine
C: proline and/or valine and/or tyrosine
D: threonine and/or hydroxyproline and/or serine and/or glycine
E: lysine and/or arginine and/or taurine

c A: leucine and/or isoleucine
B: β-phenylalanine
C: proline
D: serine
E: arginine

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d In the second run, the two solvents could not clarify the identity of A, but did help identify C, D and E. A two-way chromatogram produces better separation of components of complex mixtures, permitting easier isolation and identification.

29 a Sugar ($C_{12}H_{22}O_{11}$) and ethanol (CH_3CH_2OH) dissolve by forming hydrogen bonds with water. Salt (NaCl) dissolves by dissociating into ions, and forming ion–dipole interactions with water. Hydrochloric acid dissolves by ionising to form H_3O^+ ions and Cl^- ions.

b the conductivity of the solution

c the solute used

d Hydrochloric acid is a strong acid, and ionises completely. Thus, when in solution it will produce many ions, which will lead to a higher conductivity. Vinegar is a weak acid that partially ionises. Fewer ions in solution will lead to a lower conductivity.

e Ethanol and sugar both dissolve by forming hydrogen bonds with water. No ions are formed in this process, therefore conductivity gives no indication of solubility for these substances.

f The design of the experiment does not investigate the stated aim, therefore the conclusion is not valid. The aim of the experiment is to investigate the different *ways* substances dissolve, yet the conclusion is about the relationship between solubility and conductivity.

g For example:

– Change the aim and hypothesis of the investigation.

– Use the same amount of each solute.

– Use all ionic substances, as these all dissolve in the same way.

30 a $1s^2$

b Group 18, period 1. Helium is the first of the noble gases.

c Helium would be expected to be highly unreactive, due to its position at the top of the noble gases group.

d Helium-3 and helium-4 both contain 2 protons and 2 electrons. However, helium-3 contains 1 neutron, whereas helium-4 contains 2 neutrons.

e Helium exists as single atoms. The only forces between these atoms are dispersion forces. Furthermore, as helium is very small, these dispersion forces are very weak. As the forces between helium atoms are so weak, it has both a low melting point and a low boiling point.

f Low density makes it good for buoyancy; Low reactivity means it provides a safe atmosphere and it can be a carrier gas without interfering with the substance being analysed; liquid helium exists at temperatures below –269°C, making it useful in medical and scientific research that requires extremely low temperatures.

g i A linear economy operates on a 'take-make-dispose' model. There is little attempt to recapture, reuse or recycle materials. The circular economy model focuses on the optimal use and reuse of resources from the extraction of raw materials to production then consumption.

ii Low reactivity means that it is not possible to capture the used helium in a compound; its low density means that once released, helium continues to rise until it leaves Earth's atmosphere. These properties make it difficult to use helium in a way described by the circular economy model.

h i Hydrogen is the only gas with a density lower than helium, thus it would be a good replacement in balloons and airships (blimps); hydrogen can be produced from other substances, making it more sustainable.

ii Hydrogen is highly reactive and explosive. It would need to be kept away from oxygen and sources of ignition.

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Chapter 7 Quantifying atoms and compounds

7.1 Relative mass

Worked example: Try yourself 7.1.1

CALCULATING RELATIVE ATOMIC MASS FROM ISOTOPIC MASSES AND PERCENTAGE ABUNDANCES

Boron has two isotopes. Their relative isotopic masses and percentage abundances are provided. Calculate the relative atomic mass of boron.

Isotope	Relative isotopic mass	Relative abundance (%)
^{10}B	10.013	19.91
^{11}B	11.009	80.09

Thinking	Working
Determine the relative isotopic masses and abundances for each isotope.	First isotope: relative isotopic mass 10.013; abundance 19.91% Second isotope: relative isotopic mass 11.009; abundance 80.09%
Substitute the relative isotopic masses and abundance into the formula for calculating relative atomic mass: $A_r = \frac{(\text{relative isotopic mass} \times \%\ \text{abundance}) + (\text{relative isotopic mass} \times \%\ \text{abundance})}{100}$	$A_r = \frac{(10.013 \times 19.91) + (11.009 \times 80.09)}{100}$
Calculate the relative atomic mass.	$A_r = \frac{199.36 + 881.71}{100} = 10.811$
Express the answer to the appropriate number of significant figures.	A_r(B) = 10.81

Worked example: Try yourself 7.1.2

CALCULATING PERCENTAGE ABUNDANCES FROM RELATIVE ATOMIC MASS AND THE RELATIVE ISOTOPIC MASSES

The relative atomic mass of copper is 63.54. The relative isotopic masses of its two isotopes are 62.95 and 64.95. Calculate the relative abundances of the isotopes in naturally occurring copper.

Thinking	Working
State the relative abundances of the two isotopes in terms of x, where x is the abundance of the lighter isotope. Abundance of lighter isotope = x The abundance of heavier isotope must equal 100 – x.	Abundance of 62.95 isotope = x Abundance of 64.95 isotope = 100 – x
Substitute the relative isotopic masses, relative abundances and relative atomic mass into the formula: $A_r = \frac{(\text{relative isotopic mass} \times \%\ \text{abundance}) + (\text{relative isotopic mass} \times \%\ \text{abundance})}{100}$	$63.54 = \frac{62.95x + (64.95(100 - x))}{100}$
Expand the top line of the equation.	$63.54 = \frac{62.95x + 6495 - 64.95x}{100}$
Solve the equation to find x, the relative abundance of the lightest isotope.	$6354 = 62.95x + 6495 - 64.95x$ $6354 - 6495 = 62.95x - 64.95x$ $-141 = -2x$ $x = 70.50\%$
Determine the abundance of the heavier isotope.	Abundance of 64.95 isotope = 100 – x = 100 – 70.50 = 29.50%

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KEY QUESTIONS

Knowledge and understanding

1 B. A is incorrect because ^{13}C is not the carbon isotope used as the reference. C and D are incorrect because it is an isotope of carbon, not oxygen that is used as the reference standard.

2 ^{39}K would be most abundant. The relative atomic mass of 39.1 is closest to the mass number of 39. There needs to be a majority of isotopes with the mass number 39 for the weighted average to be just above 39.

3 B. A is incorrect because the relative atomic mass is a weighted average of the masses of the two isotopes, taking into account their abundances. C is incorrect because the lighter isotope ^{35}Cl is more abundant than the heavier isotope ^{37}Cl. D is incorrect because isotopes have the same number of protons and different numbers of neutrons.

4 $$A_r(H) = \frac{(1.008 \times 99.986) + (2.014 \times 0.014) + (3.016 \times 0.0001)}{100}$$
$$= 1.008$$

Analysis

5 Let the percentage abundance of the lighter isotope be x%.

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

$$6.94 = \frac{(6.02 \times x) + (7.02 \times (100 - x))}{100}$$
$$694 = 6.02x + 702 - 7.02x$$
$$-8 = -1.00x$$
$$x = 8.00\%$$

Percentage abundance of the lighter isotope is 8.00%.

6 **a** The peak heights can be measured from the mass spectrum. Measurements may vary slightly, depending on the screen or book you are measuring from, so this is a model answer only. Yours may vary slightly.

m/z	peak height
90	5.6
91	1.2
92	1.9
94	1.9
96	0.45

Total peak height = 5.6 + 1.2 + 1.9 + 1.9 + 0.45
= 11.1

$$\% \text{ abundance } ^{90}Zr = \frac{\text{peak height}}{\text{total peak height}} \times 100$$
$$= \frac{5.6}{10.9} \times 100$$
$$= 51\%$$

$$\% \text{ abundance } ^{91}Zr = \frac{\text{peak height}}{\text{total peak height}} \times 100$$
$$= \frac{1.2}{11.1} \times 100$$
$$= 11\%$$

$$\% \text{ abundance } ^{92}Zr = \frac{\text{peak height}}{\text{total peak height}} \times 100$$
$$= \frac{1.9}{11.1} \times 100$$
$$= 17\%$$

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

$$= \frac{1.9}{11.1} \times 100$$

$$= 17\%$$

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

$$= \frac{0.45}{11.1} \times 100$$

$$= 4.1\%$$

b $A_r(\text{Zr}) = \frac{(90 \times 50) + (91 \times 11) + (92 \times 17) + (94 \times 17) + (96 \times 4.1)}{100} = 90$

7.2 Avogadro's constant

Worked example: Try yourself 7.2.1

CALCULATING THE NUMBER OF MOLECULES GIVEN THE AMOUNT OF A SUBSTANCE

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: $N(CO_2) = ?$ $n(CO_2) = 1.6$ mol $N_A = 6.02 \times 10^{23}$
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.	$N(CO_2) = n \times N_A$ $= 1.6 \times 6.02 \times 10^{23}$ $= 9.6 \times 10^{23}$ molecules

Worked example: Try yourself 7.2.2

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: $n(H) = ?$ $n(H_2O) = 0.75$ mol
Calculate the amount, in mol, of hydrogen atoms from the amount of water molecules and the molecular formula.	$n(H) = n(H_2O) \times 2$ $= 0.75 \times 2$ $= 1.5$ mol

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Worked example: Try yourself 7.2.3

CALCULATING THE NUMBER OF ATOMS GIVEN THE AMOUNT OF A SUBSTANCE

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: $N(H) = ?$ $n(CH_4) = 0.35$ mol $N_A = 6.02 \times 10^{23}$
Calculate the amount, in mol, of hydrogen atoms from the amount of methane molecules and the molecular formula.	$n(H) = n(CH_4) \times 4$ $= 0.35 \times 4$ $= 1.4$ mol
Rearrange the formula to make the unknown the subject.	$n = \frac{N}{N_A}$ so $N(H) = n \times N_A$
Substitute in data and solve for the answer.	$N(H) = n \times N_A$ $= 1.4 \times 6.02 \times 10^{23}$ $= 8.4 \times 10^{23}$ atoms

Worked example: Try yourself 7.2.4

CALCULATING THE NUMBER OF ATOMS GIVEN THE AMOUNT OF AN IONIC COMPOUND

Calculate the number of hydrogen atoms in 0.85 mol of aluminium hydroxide ($Al(OH)_3$).

Thinking	Working
List the data given in the question next to the appropriate symbol. Include units.	The number of oxygen atoms is the unknown, so: $N(H) = ?$ $n(Al(OH)_3) = 0.85$ mol $N_A = 6.02 \times 10^{23}$
Calculate the amount, in mol, of oxygen atoms from the amount of oxygen molecules and the molecular formula.	$n(H) = n(Al(OH)_3) \times 3$ $= 0.85 \times 3$ $= 2.55$ mol
Rearrange the formula to make the unknown the subject.	$n = \frac{N}{N_A}$ so $N(H) = n \times N_A$
Substitute in data and solve for the answer.	$N(H) = n \times N_A$ $= 2.55 \times 6.02 \times 10^{23}$ $= 1.5 \times 10^{24}$ atoms

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Worked example: Try yourself 7.2.5

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

Calculate the amount, in mol, of magnesium atoms represented by 8.1×10^{20} 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 $N_A = 6.02 \times 10^{23}$
Rearrange the formula to make the unknown the subject.	$n = \frac{N}{N_A}$ n is the unknown so rearrangement not required
Substitute in data and solve for the answer.	$n(\text{Mg}) = \frac{N}{N_A}$ $= \frac{8.1 \times 10^{20}}{6.02 \times 10^{23}}$ $= 0.0013$ mol

KEY QUESTIONS

Knowledge and understanding

1 $6.022\ 140\ 76 \times 10^{23}$

2 **a** amount of substance counted in moles
b actual number of particles
c Avogadro's constant, i.e. 6.02×10^{23}

3 It could mean one mole of hydrogen atoms (H) or one mole of hydrogen molecules (H_2).

4 Number of particles = amount (mol) $\times N_A$
a $N(\text{Na}) = 3.0 \times 6.02 \times 10^{23}$
$= 1.8 \times 10^{24}$ atoms
b $N(\text{Fe}) = 1.5 \times 10^{-2} \times 6.02 \times 10^{23}$
$= 9.0 \times 10^{21}$ atoms
c $N(CO_2) = 2.85 \times 10^{-5} \times 6.02 \times 10^{23}$
$= 1.72 \times 10^{19}$ molecules

Analysis

5 Remember: Avogadro's constant, N_A, is 6.02×10^{23}.

Amount (mol) $= \frac{\text{number of particles}}{N_A}$ i.e. $n = \frac{N}{N_A}$

a $n(H_2O \text{ molecules}) = \frac{6.0 \times 10^{23}}{N_A}$
$= 1.0$ mol

b $n(\text{Ne atoms}) = \frac{2.5 \times 10^{23}}{N_A}$
$= 0.42$ mol

c $n(\text{ethanol molecules}) = \frac{3.2 \times 10^{25}}{N_A}$
$= 53$ mol

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

$\therefore n(\text{Cl atoms}) = 0.40 \times 2$

$= 0.80 \text{ mol}$

b Each methane molecule has four H atoms.

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

$= 4.8 \text{ mol}$

c Each sulfate ion has four O atoms.

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

$= 6.0 \text{ mol}$

7 Remember: Avogadro's constant, N_A, is 6.02×10^{23}.

$\text{amount (mol)} = \frac{\text{number of particles}}{N_A}$ i.e. $n = \frac{N}{N_A}$

a $n(O_2 \text{ molecules}) = \frac{4.0 \times 10^{23}}{N_A}$

$= 0.66 \text{ mol}$

Each oxygen molecule has two O atoms.

$\therefore n(\text{H atoms}) = 2 \times 0.66$

$= 1.3 \text{ mol}$

b $n(\text{methane molecules}) = \frac{3.5 \times 10^{22}}{N_A}$

$= 0.058 \text{ mol}$

Each methane molecule has four H atoms.

$\therefore n(\text{O atoms}) = 4 \times 0.058$

$= 0.23 \text{ mol}$

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

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

Each chlorine molecule has two Cl atoms.

$\therefore n(\text{Cl atoms}) = 2 \times 1.7 \times 10^{-4}$

$= 3.3 \times 10^{-4} \text{ mol}$

d $n(MgCl_2) = \frac{2.0 \times 10^{23}}{N_A}$

$= 0.33 \text{ mol}$

Each $MgCl_2$ has two chloride ions

$\therefore n(Cl^-) = 2 \times 0.33$

$= 0.66 \text{ mol}$

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7.3 Molar mass

Worked example: Try yourself 7.3.1

CALCULATING THE MOLAR MASS OF MOLECULES

Calculate the molar 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.	$A_r(H) = 1.0$ $A_r(N) = 14.0$ $A_r(O) = 16.0$
Determine the number of atoms of each element present, taking into consideration any brackets in the formula.	$1 \times$ H atom $1 \times$ N atom $3 \times$ O atoms
Determine the molar mass by adding the appropriate relative atomic masses.	$M_r = (1 \times A_r(H)) + (1 \times A_r(N)) + (3 \times A_r(O))$ $= (1 \times 1.0) + (1 \times 14.0) + (3 \times 16.0)$ $= 63.0$ g mol^{-1}

Worked example: Try yourself 7.3.2

CALCULATING THE AMOUNT OF SUBSTANCE FROM MASS

Calculate the amount, in mol, of 100 g of methane gas (CH_4).

Thinking	Working
Find the molar mass of methane using relative atomic masses sourced from the periodic table.	$A_r(C) = 12.0$ $A_r(H) = 1.0$ $M(CH_4) = 12.0 + (4 \times 1.0)$ $= 16.0$ g mol^{-1}
Calculate the amount, in mol, of methane using: $n = \frac{m}{M}$	$n(CH_4) = \frac{m}{M}$ $= \frac{100}{16.0}$ $= 6.25$ mol

Worked example: Try yourself 7.3.3

CALCULATING THE MASS OF A SUBSTANCE FROM NUMBER OF MOLES

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 formula, you can determine the molar mass by referring to the periodic table for relative atomic masses.	$m(Na_2CO_3) = ?$ g $n(Na_2CO_3) = 4.68$ mol $M(Na_2CO_3) = (2 \times 23.0) + 12.0 + (3 \times 16.0)$ $= 106.0$ g mol^{-1}
Calculate the mass of sodium carbonate using: $m = n \times M$	$m = n \times M$ $m(Na_2CO_3) = 4.68 \times 106.0$ $= 496$ g

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CASE STUDY: ANALYSIS

The sting of a bee

Analysis

1 $M(C_7H_{14}O_2) = (7 \times A_r(C)) + (14 \times A_r(H)) + (2 \times A_r(O))$
$= (7 \times 12.0) + (14 \times 1.0) + (2 \times 16.0)$
$= 84.0 + 14.0 + 32.0$
$= 130 \text{ g mol}^{-1}$

2 One bee sting releases 1.0×10^{-6} g
$n(C_7H_{14}O_{2)} = \frac{m}{M}$
$= \frac{1.0 \times 10^{-6}}{130}$
$= 7.7 \times 10^{-9}$ mol

3 $N(C_7H_{14}O_2) = n \times N_A$
$= 7.7 \times 10^{-9} \times 6.02 \times 10^{23}$
$= 4.6 \times 10^{15}$ molecules

Worked example: Try yourself 7.3.4

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.	$N(C_{12}H_{22}O_{11}) = ?$ $M(C_{12}H_{22}O_{11}) = (12 \times 12.0) + (22 \times 1.0) + (11 \times 16.0)$ $= 342.0 \text{ g mol}^{-1}$ $m(C_{12}H_{22}O_{11}) = 4.2$ g
Calculate the amount, in mol, of $C_{12}H_{22}O_{11}$, using: $n = \frac{m}{M}$	$n(C_{12}H_{22}O_{11}) = \frac{m}{M}$ $= \frac{4.2}{342.0}$ $= 0.012$ mol
Calculate the number of $C_{12}H_{22}O_{11}$ molecules using: $n = \frac{N}{N_A}$	$n = \frac{n}{N_A}$, so $N = n \times N_A$ $N(C_{12}H_{22}O_{11}) = 0.012 \times 6.02 \times 10^{23}$ $= 7.4 \times 10^{21}$ molecules

KEY QUESTIONS

Knowledge and understanding

1 The molar mass of a compound can be calculated by adding the relative atomic masses for each atom present in the compound's formula. This value is then expressed in grams per mol (g mol^{-1}).

2 $m = n \times M$

Analysis

3 **a** $M = 2 \times A_r(Cl)$
$= 2 \times 35.5$
$= 71.0 \text{ g mol}^{-1}$

b $M = (6 \times A_r(C)) + (8 \times A_r(H)) + (6 \times A_r(O))$
$= (6 \times 12.0) + (8 \times 1.0) + (6 \times 16.0)$
$= 72.0 + 8.0 + 96.0$
$= 176 \text{ g mol}^{-1}$

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c $M = A_r(\text{Cu}) + A_r(\text{S}) + (4 \times A_r(\text{O})) + (5 \times M_r([2 \times 1.0] + [1 \times 16.0]))$
$= 63.5 + 32.1 + (4 \times 16.0) + (5 \times 18.0)$
$= 249.6 \text{ g mol}^{-1}$

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

a $m(\text{O}_2) = 3.0 \text{ mol} \times 32.0 \text{ g mol}^{-1}$
$= 96.0 \text{ g}$

b $m(\text{CH}_4) = 1.5 \text{ mol} \times 16.0 \text{ g mol}^{-1}$
$= 24.0 \text{ g}$

c $m(\text{Al}_2\text{O}_3) = 2.5 \text{ mol} \times 102 \text{ g mol}^{-1}$
$= 255 \text{ g}$

5 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 a periodic table to work out the molar masses.

a $n(\text{H}_2 \text{ molecules}) = \frac{5}{2.0}$
$= 2.5 \text{ mol}$

b $n(\text{AlCl}_3) = \frac{50}{133.5}$
$= 0.37 \text{ mol}$

c $n(\text{CH}_4) = \frac{4.5}{16.0}$
$= 0.28 \text{ mol}$

6 Use the formulas: 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 a periodic table to work out the molar masses.

a $n(\text{Na}) = \frac{23}{23.0}$
$= 1.0 \text{ mol}$
number of Na atoms $= 1.0 \times 6.0 \times 10^{23}$
$= 6.0 \times 10^{23}$ atoms

b $n(\text{Ar}) = \frac{4.0}{39.9}$
$= 0.10 \text{ mol}$
number of Ar atoms $= 0.10 \times 6.0 \times 10^{23}$
$= 6.0 \times 10^{22}$ atoms

c $n(\text{Mg}) = \frac{0.243}{24.3}$
$= 0.0100 \text{ mol}$
number of Mg atoms $= 0.01 \times 6.0 \times 10^{23}$
$= 6.0 \times 10^{21}$ atoms

7 **a** $n(\text{H}_2 \text{ molecules}) = \frac{5.0}{2.0}$
$= 2.5 \text{ mol}$
$n(\text{H atoms}) = 2 \times 2.5$
$= 5.0 \text{ mol}$

b $n(\text{Mg(NO}_3)_2) = \frac{100}{148.3}$
$= 0.0674 \text{ mol}$
$n(\text{nitrate ions}) = 2 \times 0.0674$
$= 1.35 \text{ mol}$

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c $n(P_4 \text{ molecules}) = \dfrac{1.2 \times 10^{-3}}{124.0}$

$= 9.7 \times 10^{-6}$ mol

$n(\text{P atoms}) = 4 \times 9.7 \times 10^{-6}$

$= 3.9 \times 10^{-5}$ mol

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

a i $n(O_2) = \dfrac{48}{32} = 1.5$ mol

number of O_2 molecules $= 1.5 \times 6.02 \times 10^{23} = 9.00 \times 10^{23}$ molecules

ii $n(N_2) = \dfrac{50}{28} = 1.79$ mol

number of N_2 molecules $= 1.8 \times 6.02 \times 10^{23} = 1.1 \times 10^{24}$ molecules

b $n(SO_2) = \dfrac{3.2}{64} = 0.050$ mol

Number of SO_2 molecules $= 0.05 \times 6.02 \times 10^{23} = 3.0 \times 10^{22}$ molecules
Each molecule contains 2 oxygen atoms.
So, number of oxygen atoms $= 2 \times 3.0 \times 10^{22}$

$= 6.02 \times 10^{22}$ atoms.

c $n(NH_3) = \dfrac{170}{17} = 10$ mol

number of NH_3 molecules $= 10 \times 6.02 \times 10^{23} = 6.02 \times 10^{24}$ molecules
each molecule contains 4 atoms (1 of N and 3 of H).
so total number of atoms $= 4 \times 6.02 \times 10^{24} = 2.4 \times 10^{25}$ atoms.

7.4 Percentage composition, empirical and molecular formulas

Worked example: Try yourself 7.4.1

CALCULATING PERCENTAGE COMPOSITION BY MASS

Calculate the percentage by mass of nitrogen in ammonium nitrate (NH_4NO_3).

Thinking	Working
Find the molar mass of the compound.	$M(NH_4NO_3) = (2 \times 14.0) + (4 \times 1.0) + (3 \times 16.0)$ $= 80.0$ g mol^{-1}
Find the total mass of the element in one mole of the compound.	mass of N in 1 mol $= 2 \times M(N)$ $= 2 \times 14.0$ $= 28.0$ g
Find the percentage by mass of the element in the compound.	% by mass of N in NH_4NO_3 $= \dfrac{\text{mass of N in 1 mol of } NH_4NO_3}{\text{molar mass of } NH_4NO_3} \times 100$ $= \dfrac{28.0}{80.0} \times 100$ $= 35.0\%$

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Worked example: Try yourself 7.4.2

DETERMINING THE EMPIRICAL FORMULA FROM PERCENTAGE COMPOSITION

Chemical analysis of an organic compound present in the gaseous emissions from a factory shows that its percentage composition is 40.0% carbon, 6.7% hydrogen and the remainder is oxygen. Find its empirical formula.

Thinking	Working
Assume that there is 100 g of the compound, so percentages convert directly to masses. Write down the mass, in g, of all elements present in the compound.	C : 40.0 g H : 6.7 g O : 100 – 40.0 – 6.7 = 53.3 g
Calculate the amount, in mol, of each element in the compound using: $n = \frac{m}{M}$	$n(\text{C}) = \frac{40.0}{12.0}$ $= 3.33$ mol $n(\text{H}) = \frac{6.7}{1.0}$ $= 6.7$ mol $n(\text{O}) = \frac{53.3}{16.0}$ $= 3.33$ mol
Simplify the ratio by dividing all of the numbers of moles by the smallest number of moles calculated in the previous step. This gives you a ratio by number of atoms.	$\text{C} = \frac{3.33}{3.33}$ $= 1$ $\text{H} = \frac{6.7}{3.33}$ $= 2$ $\text{O} = \frac{3.33}{3.33}$ $= 1$
Find the simplest whole-number ratio.	C : H : O 1 2 1
Write the empirical formula.	CH_2O

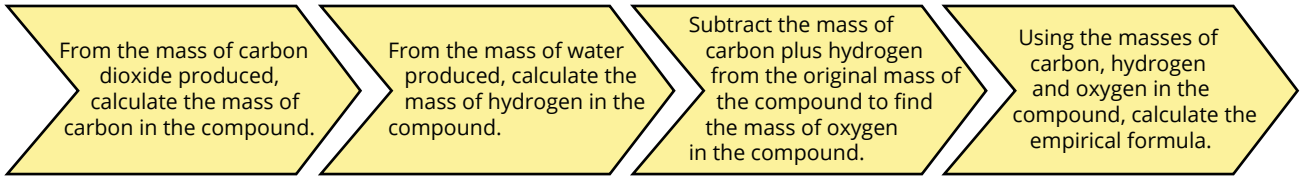
CASE STUDY: ANALYSIS

Analysing a life-saving rat poison

1 The mass of oxygen in the organic compound can be determined by subtracting the mass of carbon plus the mass of hydrogen from the original mass of the compound:

$m(\text{O}) = m(\text{compound}) - m(\text{C}) - m(\text{H})$

2



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3 $n(CO_2) = \frac{m}{M} = \frac{20.900}{44.0} = 0.475$ mol

$n(C) = n(CO_2) = 0.475$ mol

$m(C) = 0.475 \times 12.0 = 5.700$ g

$n(H_2O) = \frac{m}{M} = \frac{3.600}{18.0} = 0.200$ mol

$n(H) = 2 \times n(H_2O) = 0.400$ mol

$m(H) = 0.400 \times 1.0 = 0.400$ g

$m(O) = m(\text{compound}) - (m(C) + m(H))$

$= 7.700 - (5.700 + 0.400)$

$= 1.600$ g

empirical formula:	C	:	H	:	O
mass ratio:	5.700	:	0.400	:	1.600
molar ratio:	$\frac{5.700}{12.0}$	:	$\frac{0.400}{1.0}$	:	$\frac{1.600}{16.0}$
	0.475	:	0.400	:	0.100
	4.75	:	4	:	1
(multiple by 4 to achieve a whole-number ratio)	19	:	16	:	4

The empirical formula is $C_{19}H_{16}O_4$.

Worked example: Try yourself 7.4.3

DETERMINING MOLECULAR FORMULA

A compound has the empirical formula C_2H_5. The molar mass of this compound was determined to be 58 g mol^{-1}. What is the molecular formula of the compound?

Thinking	Working
Calculate the molar mass of one unit of the empirical formula.	Molar mass of a C_2H_5 unit $= (2 \times 12.0) + (5 \times 1.0)$ $= 29.0$ g mol^{-1}
Determine the number of empirical formula units in the molecular formula.	Number of C_2H_5 units $= \frac{58}{29.0}$ $= 2$
Determine the molecular formula of the compound.	Molecular formula $= 2 \times C_2H_5$ $= C_4H_{10}$

KEY QUESTIONS

Knowledge and understanding

1 The percentage composition of a given compound tells you the proportion by mass of the different elements in that compound.

2 $C_6H_{12}O_6$, because the ratio of the atoms could be simplified to CH_2O while still using whole numbers. The molecular formula of a molecule with this molecular formula would be CH_2O.

Analysis

3 **a** Percentage by mass of an element

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

Use a periodic table to work out the molar masses.

$\%(N) = \frac{14}{53.5} \times 100$

$= 26.2\%$

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b $\%(C) = \frac{1.278}{2.450} \times 100$

$= 52.16\%$

c $\%(Si) = \frac{28.1}{60.1} \times 100$

$= 46.8\%$

4 **a** The empirical formula provides the simplest whole-number ratio of atoms in a compound. The number of moles of each atom is found by using $n = \frac{m}{M}$, where m is the mass in grams, and M is the molar mass in g mol^{-1}.

	H	Cl
Mass	2.74 g	97.26 g
Molar mass	1 g mol^{-1}	35.5 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{2.74}{1} = 2.74$ mol	$n = \frac{97.26}{35.5} = 2.74$ mol
Divide all by the smallest amount	$\frac{2.74}{2.74} = 1$	$\frac{2.74}{2.74} = 1$
Round off to whole numbers	1	1

∴ empirical formula is HCl

b

	C	O
Mass	42.9 g	57.1 g
Molar mass	12 g mol^{-1}	16 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{42.9}{12} = 3.575$ mol	$n = \frac{57.1}{16} = 3.57$ mol
Divide all by the smallest amount	$\frac{3.575}{3.57} = 1$	$\frac{3.575}{3.57} = 1$
Round off to whole numbers	1	1

∴ empirical formula is CO

c

	C	H	O
Mass	4.104 g	0.682 g	6.966 – 4.104 – 0.682 = 2.18 g
Molar mass	12.0 g mol^{-1}	1.0 g mol^{-1}	16.0 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{4.104}{12.0} = 0.342$ mol	$n = \frac{0.682}{1.0} = 0.682$ mol	$n = \frac{2.18}{16.0} = 0.136$ mol
Divide all by the smallest amount	$\frac{0.342}{0.136} = 2.51$	$\frac{0.682}{0.136} = 5.01$	$\frac{0.136}{0.136} = 1$
Round off to whole numbers	5	10	2

∴ empirical formula is $C_5H_{10}O_2$

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d

	C	H
Mass	2.4 g	3.2 – 2.4 = 0.8 g
Molar mass	12.0 g mol^{-1}	1.0 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{2.4}{12.0} = 0.20$ mol	$n = \frac{0.8}{1.0} = 0.8$ mol
Divide all by the smallest amount	$\frac{0.20}{0.20} = 1$	$\frac{0.8}{0.20} = 4$
Round off to whole numbers	1	4

∴ empirical formula is CH_4

5

	H	C
Mass	14.3 g	85.7 g
Molar mass	1 g mol^{-1}	12 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{14.3}{1} = 14.3$ mol	$n = \frac{85.7}{12} = 7.14$ mol
Divide all by the smallest amount	$\frac{14.3}{7.14} = 2$	$\frac{7.14}{7.14} = 1$
Round off to whole numbers	2	1

∴ empirical formula is CH_2

6

Empirical formula	Relative mass of one empirical formula unit	Relative molecular mass	Number of empirical formula units in molecular formula	Molecular formula
a CH	13.0	78.0	$\frac{78.0}{13.0} = 6$	C_6H_6
b OH	17.0	34.0	$\frac{34.0}{17.0} = 2$	H_2O_2
c CH_2O	30.0	90.0	$\frac{90.0}{30.0} = 3$	$C_3H_6O_3$
d NO_2	46.0	46.0	$\frac{46.0}{46.0} = 1$	NO_2
e CH_2	14.0	154.0	$\frac{154.0}{14.0} = 11$	$C_{11}H_{22}$

7 **a** The molecular formula is always a whole-number multiple of the empirical formula. The empirical formula provides the simplest whole-number ratio of atoms in a compound. The number of moles of each atom is found by using $n = \frac{m}{M}$, where m is the mass in grams and M is the molar mass in g mol^{-1}.

	C	H	O
Mass	1.8 g	0.3 g	2.4 g
Molar mass	12.0 g mol^{-1}	1.0 g mol^{-1}	16.0 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{1.8}{12.0} = 0.15$ mol	$n = \frac{0.3}{1.0} = 0.30$ mol	$n = \frac{2.4}{16.0} = 0.15$ mol
Divide all by the smallest amount	$\frac{0.15}{0.15} = 1$	$\frac{0.30}{0.15} = 2$	$\frac{0.15}{0.15} = 1$
Round off to whole numbers	1	2	1

∴ empirical formula is CH_2O

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b Molar mass of a CH_2O unit (empirical formula) = 12.0 + (2.0 × 1) + 16.0 = 30.0 g mol^{-1}
Molar mass of the compound (molecular formula) = 180 g mol^{-1}

∴ number of CH_2O units in one molecule = $\frac{180}{30} = 6$
∴ molecular formula is $C_6H_{12}O_6$

Chapter 7 Review

REVIEW QUESTIONS

Knowledge and understanding

1 C. A is incorrect because it is molar mass that is the mass of 1 mol of a compound. B is incorrect because it is the relative isotopic mass that is the mass of one isotope of an atom. D is incorrect because relative atomic mass takes into account percentage abundances of relative isotopic masses.

2 The relative atomic mass of carbon is the weighted average of the isotopic masses of isotopes of an element. Relative isotopic masses are for the mass of individual isotopes. For example, isotopes of carbon have different relative isotopic masses (i.e. ^{12}C, ^{13}C and ^{14}C). Small amounts of ^{13}C and ^{14}C make the relative atomic mass of carbon slightly greater than 12.

3 g mol^{-1}

4 **a** 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 is 34.969.

b The relative atomic mass of an element is the weighted average of the relative masses of the isotopes of the element on the ^{12}C scale. For example, the relative atomic mass of boron is 10.81.

c The molar mass of an element is the mass of 1 mol of the element. It is equal to the relative atomic mass of the element expressed in grams. For example, the molar mass of magnesium is 24.31 g mol^{-1}. Note that relative atomic mass and molar mass of an element are numerically equal. However, relative atomic mass has no units because it is the mass of one atom of the element compared with the mass of one atom of the carbon-12 isotope.

5 The empirical formula of a compound gives the simplest whole-number ratio of elements in that compound. The molecular formula gives the actual number of each type of element in one molecule of the compound.

Application and analysis

6 $A_r(Pd) = \frac{(101.9049 \times 0.9600) + (103.9036 \times 10.97) + (104.9046 \times 22.23) + (105.9032 \times 27.33) + (107.9039 \times 26.71) + (109.9044 \times 11.80)}{100}$

$= 106.4$

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

$= 39.96$

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

$= 39.11$

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.

8 **a** Peak heights: ^{50}Cr = 0.3 units, ^{52}Cr = 12 units, ^{53}Cr = 1 unit, ^{54}Cr = 0.2 units, Total height = 13.5 units
Percentages: ^{50}Cr = 2.2%, ^{52}Cr = 88.9%, ^{53}Cr = 7.4%, ^{54}Cr = 1.5%

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

$= 52$

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9 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, $N_A = 6.02 \times 10^{23}$.

Use the formula: $n = \frac{\text{number of particles}}{N_A}$.

a $n(H_2O) = \frac{4.50 \times 10^{23}}{6.02 \times 10^{23}}$
$= 0.748$ mol

b $n(CH_4) = \frac{9.00 \times 10^{24}}{6.02 \times 10^{23}}$
$= 15.0$ mol

c $n(Cl_2) = \frac{2.3 \times 10^{28}}{6.02 \times 10^{23}}$
$= 3.8 \times 10^4$ mol

d $n(C_{12}H_{22}O_{11}) = \frac{1}{6.02 \times 10^{23}}$
$= 1.7 \times 10^{-24}$ mol

10 B. 1.0 mol of CO_2 is correct as 1.0 mol of CO_2 molecules contains 2×1.0 mol = 2.0 mol of O atoms. A is incorrect as 0.50 mol of O_2 molecules contains 2×0.50 mol = 1.0 mol of O atoms. C is incorrect as 1.0 mol of CO molecules contains 1×1.0 mol = 1.0 mol of O atoms. D is incorrect as 1.0 mol of H_2O molecules contains 1×1.0 mol = 1.0 mol of O atoms.

11 a i $N(NH_3) = n \times N_A$
$= 1.45 \times 6.02 \times 10^{23}$
$= 8.73 \times 10^{23}$ 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 $N(H_2S) = n \times N_A$
$= 0.576 \times 6.02 \times 10^{23}$
$= 3.47 \times 10^{23}$ 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 $N(HNO_3) = n \times N_A$
$= 0.0153 \times 6.02 \times 10^{23}$
$= 9.21 \times 10^{21}$ 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 $N(C_{12}H_{22}O_{11}) = n \times N_A$
$= 2.5 \times 6.02 \times 10^{23}$
$= 1.5 \times 10^{24}$ 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

12 The molar mass of an element or compound is equal to the sum of the relative atomic masses of each atom in the formula. Relative atomic masses can be obtained from a periodic table.

a $M(Na) = 23.0$ g mol^{-1}

b $M(HNO_3) = 1.0 + 14.0 + (3 \times 16.0)$
$= 63.0$ g mol^{-1}

c $M(Mg(NO_3)_2) = 24.3 + (2 \times 14) + (6 \times 16)$
$= 148.3$ g mol^{-1}

d $M(FeCl_3 \cdot 6H_2O) = 55.8 + (3 \times 35.5) + (12 \times 1) + (6 \times 16)$
$= 270.3$ g mol^{-1}

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

a $m(C_2H_6) = 0.080 \times (24 + 6)$
$= 2.4$ g

b $m(C_6H_{12}O_6) = 0.45 \times ((6 \times 12.0) + (12 \times 1.0) + (6 \times 16.0))$
$= 81$ g

c $m(NH_2CONH_2) = 2.8 \times 10^{-3} \times ((2 \times 14) + (4 \times 1) + 12 + 16)$
$= 0.21$ g

d $m(Cu) = 5.35 \times 63.5$
$= 340$ g

14 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 $n(C) = \frac{12}{12.0}$
$= 1.0$ mol

b $n(S_8) = \frac{100}{256.8}$
$= 0.389$ mol

c $n(C_6H_4(OCOCH_3)COOH) = \frac{0.500}{180.0}$
$= 0.002\,78$ mol

d $n(Al_2O_3) = \frac{2\,800\,000}{102.0}$
$= 2.7 \times 10^4$ mol

15 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 a periodic table to work out the molar masses

a $M(C_{10}H_8) = 128.0$ g mol^{-1}, $M(C) = 12.0$ g mol^{-1}

$\%(C) = \frac{10 \times 12.0}{128.0} \times 100$
$= 93.8\%$

b $M(NH_2CONH_2) = 60.0$ g mol^{-1}, $M(C) = 12.0$ g mol^{-1}

$\%(C) = \frac{1 \times 12.0}{60.0} \times 100$
$= 20\%$

c $M(C_6H_4(OCOCH_3)COOH) = 180.0$ g mol^{-1}, $M(C) = 12.0$ g mol^{-1}

$\%(C) = \frac{9 \times 12.0}{180.0} \times 100$
$= 60.0\%$

16 The empirical formula provides the simplest whole-number ratio of atoms in a compound. The number of moles of each atom is found by using $n = \frac{m}{M}$, where m is the mass in grams and M is the molar mass.

a

	C	O
Mass	42.9 g	57.1 g
Molar mass	12.0 g mol^{-1}	16.0 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{42.9}{12.0} = 3.575$ mol	$n = \frac{57.1}{16.0} = 3.57$ mol
Divide all by the smallest amount	$\frac{3.575}{3.57} = 1$	$\frac{3.57}{3.57} = 1$
Round off to whole numbers	1	1

∴ empirical formula is CO

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b

	C	O
Mass	27.2 g	72.8 g
Molar mass	12.0 g mol^{-1}	16.0 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{27.2}{12.0} = 2.27$ mol	$n = \frac{72.8}{16.0} = 4.55$ mol
Divide all by the smallest amount	$\frac{2.27}{2.27} = 1$	$\frac{4.55}{2.27} = 2$
Round off to whole numbers	1	2

∴ empirical formula is CO_2

c

	C	O	H
Mass	54.5 g	36.4 g	9.1 g
Molar mass	12.0 g mol^{-1}	16.0 g mol^{-1}	1.0 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n(C) = \frac{54.5}{12.0} = 4.54$ mol	$n = \frac{36.4}{16.0} = 2.28$ mol	$n = \frac{9.1}{1.0} = 9.1$ mol
Divide all by the smallest amount	$\frac{4.54}{2.28} = 1.99$	$\frac{2.28}{2.28} = 2$	$\frac{9.1}{2.28} = 3.99$
Round off to whole numbers	2	2	4

∴ empirical formula is C_2H_4O

d

	C	Cl	H
Mass	9.4 g	4.7 g	0.65 g
Molar mass	12.0 g mol^{-1}	35.5 g mol^{-1}	1.0 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{9.4}{12.0} = 0.78$ mol	$n = \frac{4.7}{35.5} = 0.13$ mol	$n = \frac{0.65}{1.0} = 0.65$ mol
Divide all by the smallest amount	$\frac{0.78}{0.13} = 6.0$	$\frac{0.13}{0.13} = 1$	$\frac{0.65}{0.13} = 5.00$
Round off to whole numbers	6	1	5

∴ empirical formula is C_6H_5Cl

17 The empirical formula provides the simplest whole-number ratio of atoms in a compound. The amount of each atom is found by 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 find the molar masses of C, H, N and O.

	C	H	N	O
Mass	32 g	6.7 g	18.7 g	42.6 g
Molar mass	12.0 g mol^{-1}	1.0 g mol^{-1}	14.0 g mol^{-1}	16.0 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{32}{12.0} = 2.67$ mol	$n = \frac{6.7}{1.0} = 6.7$ mol	$n = \frac{18.7}{14.0} = 1.34$ mol	$n = \frac{42.6}{16.0} = 2.66$ mol
Divide all by the smallest amount	$\frac{2.67}{1.34} = 1.99$	$\frac{6.7}{1.34} = 5$	$\frac{1.34}{1.34} = 1$	$\frac{2.66}{1.34} = 1.99$
Round off to whole numbers	2	5	1	2

∴ empirical formula is $C_2H_5NO_2$

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18 a The empirical formula provides the simplest whole-number ratio of atoms in a compound. The amount of each atom is found by 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 find the molar masses of C, H, N and O.

	C	H	N	O
Mass	49.48 g	5.15 g	28.87 g	100 – 49.48 – 5.15 – 28.87 = 16.5 g
Molar mass	12.0 g mol^{-1}	1.0 g mol^{-1}	14.0 g mol^{-1}	16.0 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{49.48}{12.0}$ = 4.12 mol	$n = \frac{5.15}{1.0}$ = 5.15 mol	$n = \frac{28.87}{14.0}$ = 2.06 mol	$n = \frac{16.5}{16.0}$ = 1.03 mol
Divide all by the smallest amount	$\frac{4.12}{1.03} = 4$	$\frac{5.15}{1.03} = 5$	$\frac{2.06}{1.03} = 2$	$\frac{1.03}{1.03} = 1$
Round off to whole numbers	4	5	2	1

∴ empirical formula is $C_4H_5N_2O$

b $n = \frac{m}{M}$ is rearranged to show that $M = \frac{m}{n} = \frac{38.8}{0.20} = 194$ g mol^{-1}

c The molar mass of one empirical formula unit $(C_2H_5NO_2) = 4 \times 12.0 + 5 \times 1.0 + 2 \times 14.0 + 16.0$
$= 97.0$ g mol^{-1}

The number of empirical formula units in the molecular formula $= \frac{194}{97.0} = 2$

The molecular formula is $C_8H_{10}N_4O_2$.

d n(caffeine) $= \frac{m}{M}$
$= \frac{1.00}{194}$
$= 0.00515$ mol

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

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

∴ $151.96 = \frac{(x \times 150.92) + ((100 - x) \times 152.92)}{100}$

∴ $x = 48.0$

∴ proportions of the isotopes are 48.0% and 52.0%.

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

Mass of one calcium atom $= \frac{40.1}{6.0 \times 10^{23}}$
$= 6.67 \times 10^{-23}$ g

b Mass of one water molecule $= \frac{18.0}{6.02 \times 10^{23}}$
$= 3.0 \times 10^{-23}$ g

c Mass of one CO_2 molecule $= \frac{44.0}{6.02 \times 10^{23}}$
$= 7.3 \times 10^{-23}$ g

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21 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) = 55.8 g mol^{-1} and $M(H_2O)$ = 18.0 g mol^{-1}

$\therefore n(H_2O) = \frac{20}{18.0} = 1.11$ mol

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

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

22 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 $n(NaCl) = \frac{5.85}{58.5}$

$= 0.100$ mol

ii $n(Na^+) = n(NaCl)$

$= 0.100$ mol

$n(Cl^-) = n(NaCl)$

$= 0.100$ mol

b i $n(CaCl_2) = \frac{45.0}{111.0}$

$= 0.405$ mol

ii $n(Ca^{2+}) = n(CaCl_2)$

$= 0.405$ mol

$n(Cl^-) = 2 \times n(CaCl_2)$

$= 2 \times 0.405$

$= 0.810$ mol

c i $n(Fe_2(SO_4)_3) = \frac{1.68}{399.9}$

$= 0.004\,20$ mol

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

$= 0.008\,40$ mol

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

$= 0.0126$ mol

23 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, $N_A = 6.02 \times 10^{23}$. Use the formula: $n = \frac{\text{number of particles}}{N_A}$

a M(substance) = $\frac{72}{0.5} = 144$ g mol^{-1}

b n(substance) = $\frac{6.02 \times 10^{22}}{N_A} = 0.1$ mol

M(substance) = $\frac{10}{0.1} = 100$ g mol^{-1}

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24 C. It is useful to remember the formula $m = n \times M$, where m is the mass in grams, n the amount of substance in mol, and M the molar mass. Use a periodic table to find the molar masses of iron and silver, $M = 55.8$ g mol^{-1} and 108 g mol^{-1}, respectively.

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

A: $m(\text{Cu}) = 150$ g

B: $m(\text{Fe}) = 2.0 \times 55.8$

$= 111.6$ g

C: $n(\text{Ag}) = \dfrac{1.2 \times 10^{24}}{N_A}$

$= 2.0$ mol

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

$= 216$ g

This is the greatest mass.

D: $m(\text{Na}) = 1.5 \times 23.0$

$= 34.5$ g

25 It is useful to remember the formula $m = n \times M$, 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 $M(\text{antibiotic}) = 12\,500$ g mol^{-1} $= 1.25 \times 10^4$ g mol^{-1}

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

c Number of molecules $= n \times N_A$

$= 1.6 \times 10^{-7} \times 6.02 \times 10^{23}$

$= 9.6 \times 10^{16}$ molecules

26 The empirical formula provides the simplest whole-number ratio of atoms in a compound. The amount of each atom is found by using $n = \dfrac{m}{M}$, where m is the mass in grams and M is the molar mass in g mol^{-1}. Use a periodic table to find the molar mass of O.

	Ni	O
Mass	3.370 g	4.286 – 3.370 = 0.916 g
Molar mass	? g mol^{-1}	16 g mol^{-1}
Amount using $n = \dfrac{m}{M}$	$n = \dfrac{3.370}{M(\text{Ni})}$	$n = \dfrac{0.916}{16.0} = 0.0573$ mol

$\dfrac{n(\text{Ni})}{n(\text{O})} = \dfrac{0.370}{M(\text{Ni})} = \dfrac{1}{0.0573}$

As the empirical formula is NiO, $\dfrac{0.370}{M(\text{Ni})} \times \dfrac{1}{0.0573} = \dfrac{1}{1}$

$\therefore M(\text{Ni}) = \dfrac{0.370}{0.0573} = 58.9$ g mol^{-1}

27 a

	C	O	H
Mass	52.2 g	34.8 g	100 – 52.2 – 34.8 = 13 g
Molar mass	12.0 g mol^{-1}	16 g mol^{-1}	1.0 g mol^{-1}
Amount using $n = \dfrac{m}{M}$	$n(\text{C}) = \dfrac{52.2}{12} = 4.35$ mol	$n(\text{O}) = \dfrac{34.8}{16} = 2.175$ mol	$n(\text{H}) = \dfrac{13}{1} = 13$ mol
Divide all by the smallest amount	$\dfrac{4.35}{2.175} = 2$	$\dfrac{2.175}{2.175} = 1$	$\dfrac{13}{2.175} = 6$

Step 4: Write as empirical formula: C_2H_6O

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b Step 1: Calculate molar mass of one empirical formula unit.

$M = (2 \times 12) + (6 \times 1) + 16$

$= 46 \text{ g mol}^{-1}$

Step 2: Calculate molar mass of compound.

$$M = \frac{m}{M} = \frac{100}{2.17} = 46 \text{ g mol}^{-1}$$

Step 3: Calculate number of empirical formula units in compound.

$$\text{no. of units} = \frac{46}{46} = 1$$

Step 4: Write molecular formula: C_2H_6O

28 a D, F, E, A, C, B

b Step 1: Calculate mass of compound after reaction.

$m = 21.068 - 20.310 = 0.758$ g

Step 2: Calculate mass of oxygen.

$m = 0.758 - 0.542 = 0.216$ g

Step 3: Calculate amount, in mol, of oxygen.

$$n = \frac{m}{M_r}$$

$$= \frac{0.126}{16} = 0.0135 \text{ mol}$$

Step 4: Use mole ratios to determine amount of metal.

Ratio is 1 : 1, so $n = 0.0135$ mol

Step 5: Calculate molar mass of metal.

$$M = \frac{m}{n} = \frac{0.542}{0.0135} = 40.1$$

Step 6: Fill in table.

	Metal	Oxygen
Mass (g)	0.542	0.216
Relative atomic mass	40.1	16.0
Moles	0.0135 mol	0.0135 mol
Ratio	1	1

c Refer to the periodic table to identify the metal as calcium. Calcium has a relative atomic mass of 40.1.

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Chapter 8 Organic compounds

8.1 Organic materials

KEY QUESTIONS

Knowledge and understanding

1 Compounds such as petrol, polymers and cosmetics are **carbon**-based compounds. Many of these compounds are currently produced from **crude oil**, which is a mixture made up of the remains of marine microorganisms, such as **bacteria** and plankton that died millions of years ago. The great age of these deposits explains why petrol is called a **fossil** fuel. Crude oil is a **non-renewable** resource because no more carbon is being added to the environment.

2 Organic compounds are carbon-based compounds which typically also have hydrogen and oxygen in them. They may also contain nitrogen, sulfur, phosphorus and halogens (fluorine, chlorine, bromine, iodine).

3 Photosynthesis is the name of the process by which carbon dioxide is changed by plants into glucose. This involves the following reaction:

$$6CO_2(g) + 6H_2O(l) \xrightarrow{\text{UV light, chlorophyll}} C_6H_{12}O_6(aq) + 6O_2(g)$$

Water and carbon dioxide are converted by a series of complex reactions in the presence of UV light and the molecule chlorophyll into glucose and oxygen.

Analysis

4 Organic chemicals that are made from crude oil are non-renewable because crude oil is a fossil fuel and is not being made at a rate that is anywhere close to the rate at which we use it.

5 The glucose comes from photosynthesis in plants. This is a renewable process as the glucose can be replaced at a rate that is similar to the rate at which we use it, so the polyethene made from ethanol derived from glucose is also renewable.

6 **a** If the polymers are recycled, rather than thrown away, the organic compounds are reused, rather than new hydrocarbons being refined from crude oil.

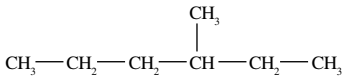
b If the government insists by legislation that petrol and diesel-fuelled cars are replaced by electric cars, then it will have to make this possible by providing infrastructure, while the non-renewable resource, crude oil, will be saved for uses where it cannot be replaced.

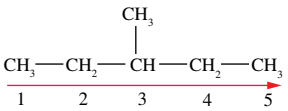
c Since crude oil is a non-renewable resource, it is important to find other ways to make the organic chemicals that we require, so replacement materials, such as plant-based biomass, need to be found.

8.2 Hydrocarbons

Worked example: Try yourself 8.2.1

IUPAC NAMING SYSTEM FOR ALKANES

Write the systematic name for the following molecule. 

Thinking	Working
Identify the longest carbon chain in the molecule. The stem name of the molecule is based on this longest chain.	There are 5 carbons in the longest chain. The stem name is based on pentane.
Number the carbons, starting from the end closest to the side chain.	 Note: In this molecule, it does not matter at which end the numbering begins as the side branch comes off the central carbon, which is the third carbon.
Identify the side chain and its location.	The side chain is a methyl group on carbon number 3.
Combine all components.	The name of the molecule is 3-methylpentane.

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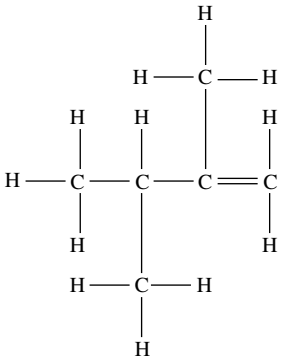
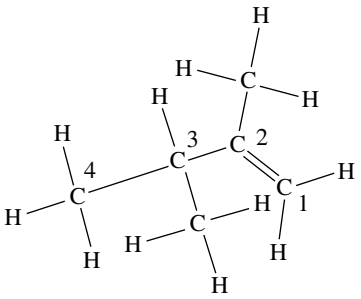
Worked example: Try yourself 8.2.2

IUPAC NAMING SYSTEM FOR ALKENES

Write the systematic name for the following molecule.



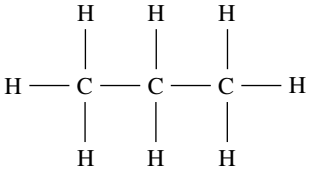
Thinking	Working
Identify the longest carbon chain in the molecule that contains the double bond. The name of the molecule is based on this longest chain.	There are 4 carbons in the longest chain with the double bond. The name is based on butene.
Number the carbon atoms, starting from the end closest to the double bond. Note the position of the double bond.	 There is a double bond on carbon number 1, so the longest chain is but-1-ene.
Identify any side chain and the number carbon that it is on.	The side chains are methyl groups and they are on carbons numbered (from smallest to largest) 2 and 3.
Combine all components.	The name of the molecule is 2,3-dimethylbut-1-ene.

KEY QUESTIONS

Knowledge and understanding

1 **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 $CH_3CH_2CH_3$

d $C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(l)$

2 **a** alkanes: C_nH_{2n+2}; alkenes: C_nH_{2n}

b Alkanes have only C–C single bonds, so have the largest ratio of hydrogen atoms to carbon atoms; alkenes have one C=C double bond, so have fewer hydrogens per carbon than do alkanes.

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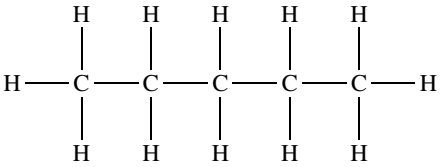
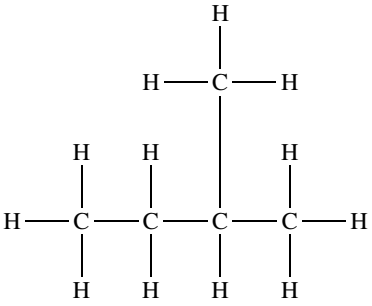
3 **a** Structural isomers of a compound, such as C_5H_{12}, have the same molecular formula, but they differ from each other by having a different arrangement of the carbon atoms in the molecule, and hence a different structure.

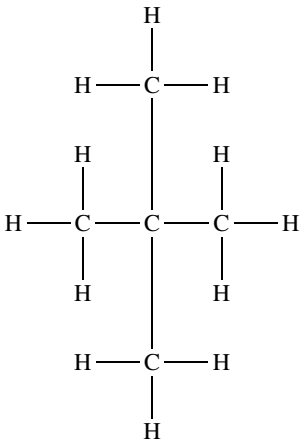
b



pentane



2-methylbutane

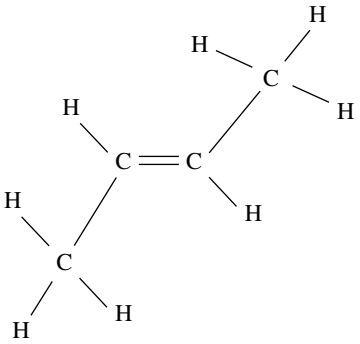


2,2-dimethylpropane

4 **a** pent-1-ene

b $CH_2=CHCH_2CH_2CH_3$ or $CH_3CH_2CH_2CH=CH_2$ or $CH_3CH_2CH_2CHCH_2$ or $CH_2CHCH_2CH_2CH_3$

5 **a** but-2-ene



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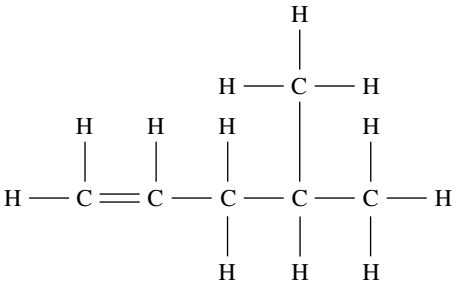
b 4-methylpent-1-ene



Analysis

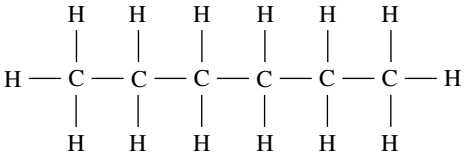
6 **a** hexane. The longest chain of carbon atoms has 6 carbon atoms (hex-) and there are no branches.

b 3-methylhexane. The longest chain of carbon atoms is 6, giving the stem name of hexane, and there is one methyl, $–CH_3$, branch on carbon number 3, when numbered from the end closest to the branch.

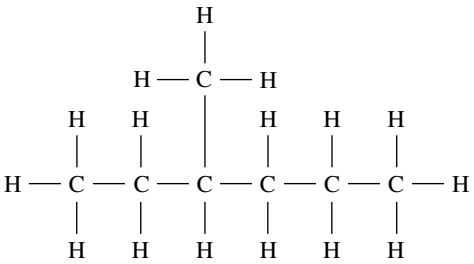
c 2,4-dimethylhexane. The longest chain of carbon atoms is 6, giving the stem name of hexane, and there are two methyl, $–CH_3$, branches on carbons numbered 2 and 4, when numbered from the end closest to the first branch.

d 2,2-dimethylbutane. The longest chain of carbon atoms is 4, giving the stem name of butane, and there are two methyl, $–CH_3$, branches on carbon number 2, when numbered from the end closest to the branches.

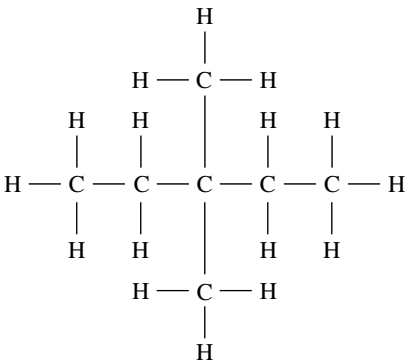
7 **a** hexane



b 3-methylhexane



c 3,3-dimethylpentane



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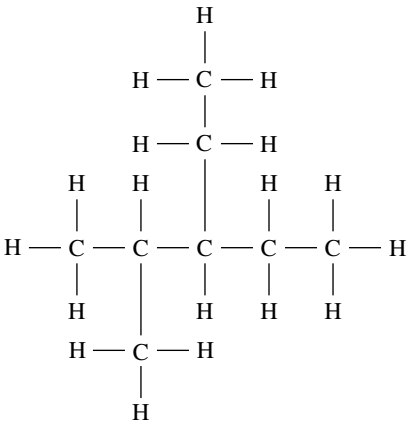
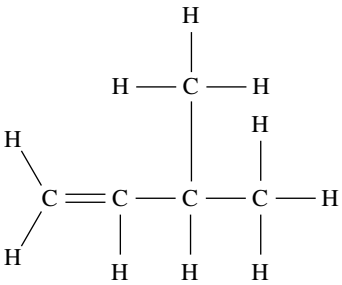
d 3-ethyl-2-methylpentane

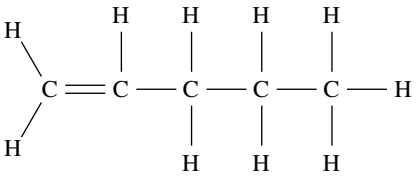


8 **a**

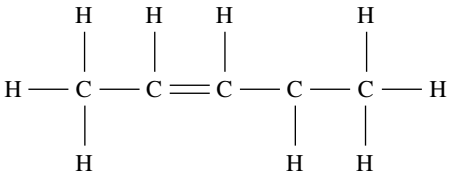


b C_5H_{10}

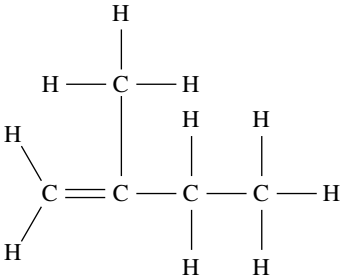
c



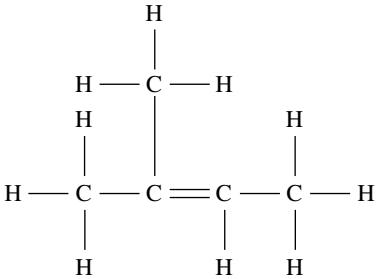
pent-1-ene



pent-2-ene



2-methylbut-1-ene



2-methylbut-2-ene

d 3-methylbut-2-ene is actually the same molecule as 2-methylbut-2-ene, but is just being viewed from the other end of the molecule. IUPAC rules specify the the carbon atoms in the chain are numbered from the end of the chain that will give the smallest numbers to double-bonded carbon atoms.

e The molecular formula of this compound is C_5H_{10}, so all isomers must have the same molecular formula. The proposed isomer, dimethylpropene, would have the molecular formula C_5H_{11}, and it would have 5 bonds around the 2nd carbon atom in the chain, which is not possible.

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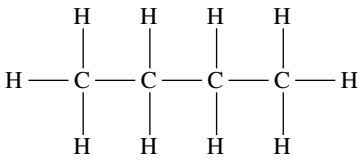
9 a i

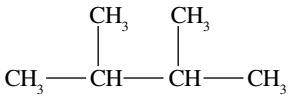


$CH_3(CH_2)_2CH_3$

ii It is not possible to have a methyl branch (or any branch) on the first carbon, so 1-methyl is not a branch, it is the continuation of the main carbon chain.

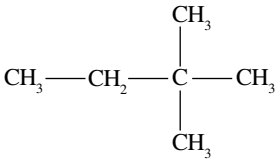
iii butane

b i



2,3-dimethylbutane

And

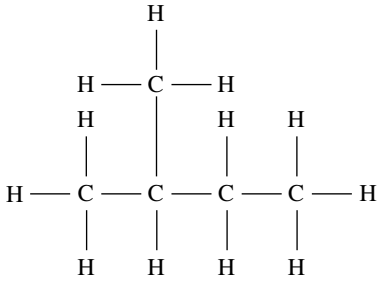


2,2-dimethylbutane

ii The methyl branches are not numbered. These could both be on carbon number 2, or one could be on carbon number 2 and the other on carbon number 3.

iii 2,2-dimethylbutane or 2,3-dimethylbutane

c i



ii The smallest possible number should always be selected for numbering a branch. Carbon number 3 from one end of the chain is actually carbon number 2 from the other end.

iii 2-methylbutane

10

Alkene	Incorrect semi-structural formula	Mistake	Correct semi-structural formula
but-2-ene	$CH_2CHCH_2CH_3$	Double bond is in the wrong place. Formula given is for but-1-ene	$CH_3CHCHCH_3$
2-methylprop-1-ene	$CH_2CH(CH_3)_2$	Five bonds around carbon number 2. Extra H on carbon number 2.	$CH_2C(CH_3)_2$
2,3-dimethylpent-2-ene	$CH_3C(CH_3)CCH_2CH_3$	Missing methyl group on carbon number 3.	$CH_3C(CH_3)C(CH_3)CH_2CH_3$

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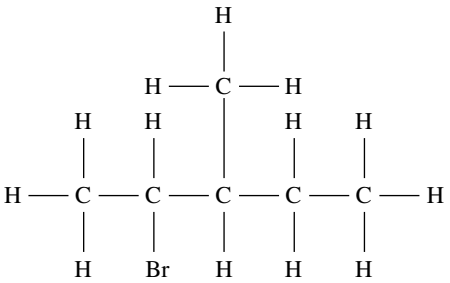
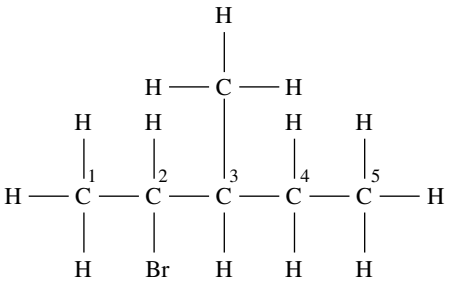
8.3 Haloalkanes

Worked example: Try yourself 8.3.1

IUPAC NAMING SYSTEM FOR HALOALKANES

Write the systematic name for the following molecule.



Thinking	Working
Identify the longest carbon chain in the molecule that contains the halogen atom/s. The name of the molecule is based on this longest chain.	There are 5 carbons in the longest chain with the halogen atom. The name is based on pentane.
Number the carbon atoms, starting from the end closest to the first halogen atom. Note the position of the halogen atom/s.	 There is a bromine atom on carbon number 2, so the longest chain is 2-bromopentane.
Identify each side chain and the number carbon that it is on.	The side chain is a methyl group and it is on carbon number 3.
Combine all components, remembering to list them in alphabetical order.	The name of the molecule is 2-bromo-3-methylpentane.

CASE STUDY: ANALYSIS

Haloalkanes and the ozone layer

Analysis

1 A free radical has an unpaired electron. Covalent bonds are formed when two unpaired electrons are shared between two atoms, so a free radical is very reactive because it has an electron already available to form a bond, whereas before new bonds can be formed in other reactions the existing bonds have to be broken.

2 Because higher energy light is needed to break the bonds in O_2, compared to O_3, this suggests that the bonds in ozone are weaker than the covalent double bond in an oxygen molecule.

3 The preferential breaking of the carbon–chlorine bond suggests that this bond is not as strong as the carbon–fluorine bond, which does not break.

4 CCl_2F_2 is called dichlorodifluoromethane.

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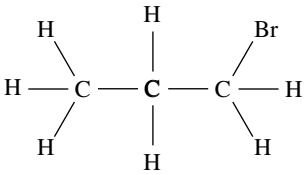
KEY QUESTIONS

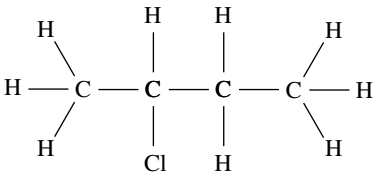
Knowledge and understanding

1 **a** CH_3CH_2Cl
b CH_3Br
c CH_3CHICH_3

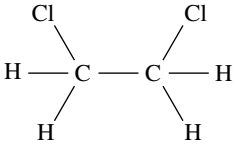
2 **a** 1-bromopropane



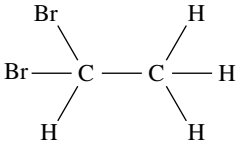
b 2-chlorobutane



c 1,2-dichloroethane



d 1,1-dibromoethane



3 **a** 2-chloro-2-iodopropane
b 1,1,2-tribromoethane
c tetrachloromethane

Analysis

4 The four isomers are 1-chlorobutane, 2-chlorobutane, 1-chloro-2-methylbutane, 2-chloro-2-methylbutane.

5 Because the bromine atom could be bonded to carbon number 1 or carbon number 2, a number must be included in the name of bromobutane. So, it must be either 1-bromobutane or 2-bromobutane.

6 Having two chlorine atoms in the molecule increases the possibilities for isomers significantly. The chlorine atoms may be on the same carbon, on adjacent carbons, or even on the two carbons at either end of the molecule. When there is just one chlorine atom in the molecule, there are only two isomers.

For C_3H_7Cl the isomers are: $CH_3CH_2CH_2Cl$ (1-chloropropane) and $CH_3CHClCH_3$ (2-chloropropane).

For $C_3H_6Cl_2$ the isomers are: $CHCl_2CH_2CH_3$ (1,1-dichloropropane), $CH_2ClCHClCH_3$ (1,2-dichloropropane), $CH_2ClCH_2CH_2Cl$ (1,3-dichloropropane), $CH_2CCl_2CH_3$ (2,2-dichloropropane).

7 **a** The molecule is polar because there is one C–Br bond, which is polar, and so the molecule is asymmetrical.
b The molecule is polar because there is one C–Cl bond, which is polar, and so the molecule is asymmetrical.
c The molecule is polar because there are two C–Br bonds, which are polar, on the same carbon, making the molecule asymmetrical.
d The molecule is polar because the C–Cl bond is more polar than the C–I bond, (chlorine is more electronegative than iodine), so even though there is a polar bond on each carbon, these are not equal, so do not cancel out. The molecule is asymmetrical.

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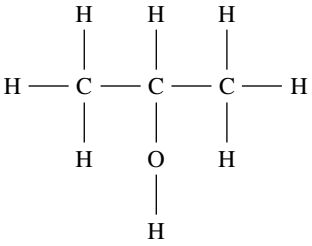
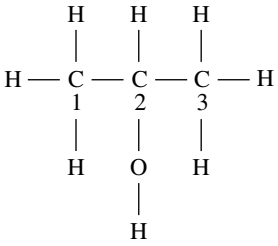
8.4 Alcohols and carboxylic acids

Worked example: Try yourself 8.4.1

IUPAC NAMING SYSTEM FOR ALCOHOLS

Write the systematic name for the following molecule.

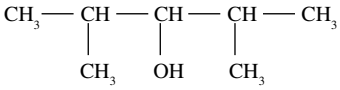


Thinking	Working
Identify the longest carbon chain in the molecule. The name of the molecule is based on this longest chain.	There are 3 carbons in the longest chain. The name is based on propane.
Identify the functional group that is present.	There is a hydroxyl group present.
Number the carbon atoms, starting from the end closest to the functional group.	
Identify the position(s) and the type(s) of side chains.	There are no branches in this molecule.
Combine all components. Place the number for the position of the side chain in front of the prefix, and the number for the position of the hydroxyl group in front of the –ol ending.	The name of the molecule is propan-2-ol.

Worked example: Try yourself 8.4.2

IUPAC NAMING SYSTEM FOR ALCOHOLS

Write the systematic name for the following molecule.



Thinking	Working
Identify the longest carbon chain in the molecule. The name of the molecule is based on this longest chain.	There are 5 carbons in the longest chain. The name is based on pentane.
Identify the functional group that is present.	There is a hydroxyl group present.
Number the carbon atoms, starting from the end closest to the functional group.	The functional group will be on C3.
Identify the position(s) and the type(s) of side chains.	There is a methyl ($-CH_3$) group on C2 and on C4, so the prefix 'dimethyl' will be used.
Combine all components. Place the number for the position of the side chain in front of the prefix, and the number for the position of the hydroxyl group in front of the –ol ending.	The name of the molecule is 2,4-dimethylpentan-3-ol.

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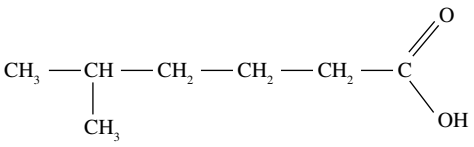
Worked example: Try yourself 8.4.3

IUPAC NAMING SYSTEM FOR CARBOXYLIC ACIDS

Write the systematic name for the following molecule.



Thinking	Working
Identify the functional group that is present.	There is a carboxyl group present.
Identify the longest carbon chain that includes the carboxyl carbon. This atom will be C1. The stem name of the molecule is based on this longest chain.	There are 6 carbons in the longest chain, so the stem name is based on hexane.
Number the carbon atoms, starting from the end incorporating the functional group.	The functional group will be on C1.
Identify the position(s) and the type(s) of side chains.	There is a methyl group on C5.
Combine all components. Place the number for the position of the side chain in front of the prefix and using the ending –oic acid.	The name of the molecule is 5-methylhexanoic acid.

KEY QUESTIONS

Knowledge and understanding

1 The hydroxyl, –OH, functional group forms hydrogen bonds with water molecules. Because methanol and ethanol have only one and two carbon atoms respectively, the non-polar (alkyl) part of the molecule is not large enough to overcome this attraction.

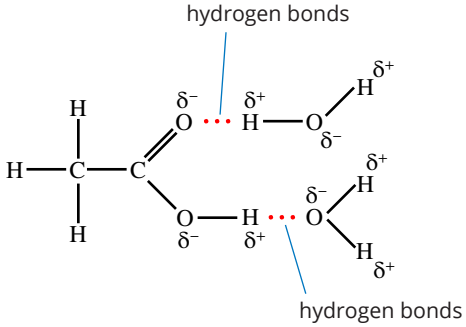
2 **a** The presence of the hydroxyl functional group enables stronger intermolecular forces to be formed between alcohol molecules. This attractive force is stronger than the dispersion forces between alkane molecules, so more energy is required to separate alcohols than alkanes with the same number of carbon atoms, hence the boiling points of alcohols are higher.

b As the length of the carbon chain in alcohols increases, the strength of the dispersion forces between the molecules increases (number of electrons in the molecules increases), so more energy is required to separate the molecules and so the boiling points of the alcohols increase.

3 **a** $C_nH_{2n+1}OH$ or $C_nH_{2n+2}O$

b The molecule $C_5H_{12}OH$ is not a member of the alcohol homologous series because it has too many hydrogen atoms in its formula. With 5 carbon atoms, it should have the formula $C_5H_{11}OH$ to belong to the alcohol homologous series.

4



5 **a** $CH_3CH_2CH_2COOH$

b $CH_3CH(CH_3)COOH$

c $CH_3CH_2CH(CH_3)CH(CH_3)COOH$

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Analysis

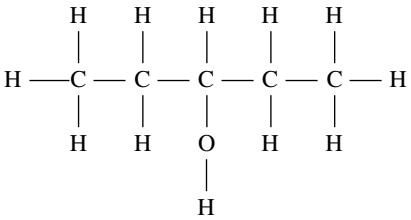
6 **a** propan-2-ol

b 2-methylbutan-2-ol

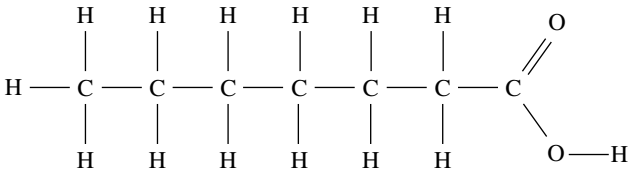
c 2-methylbutanoic acid

d 3,4-dimethylpentanoic acid

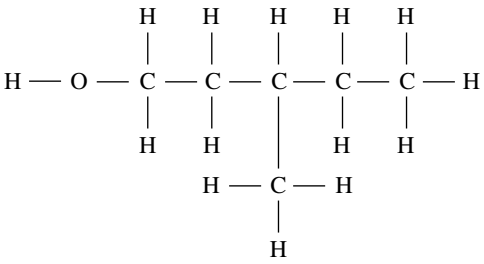
7 **a** pentan-3-ol



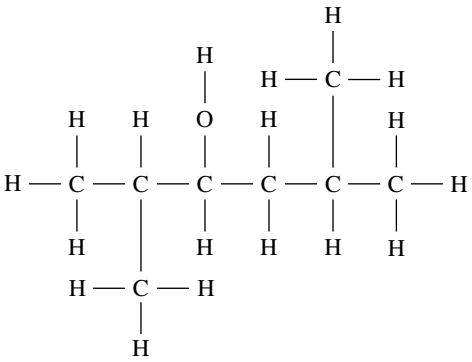
b heptanoic acid



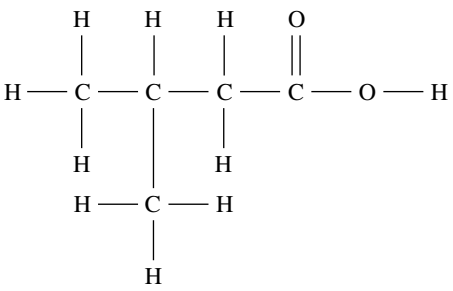
c 3-methylpentan-1-ol



d 2,5-dimethylhexan-3-ol



e 3-methylbutanoic acid



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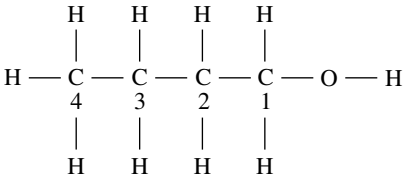
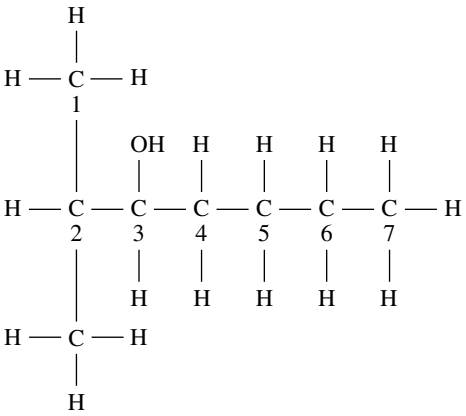
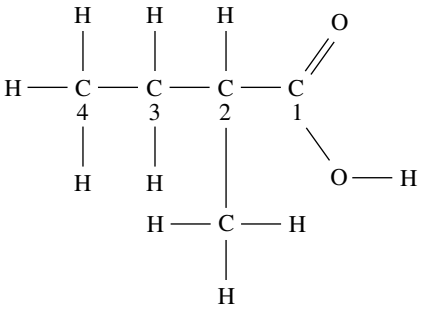
8 **a** The correct name is butan-1-ol. Carbon number 4 should be labelled carbon number 1, as that is the end closest to the hydroxyl group.



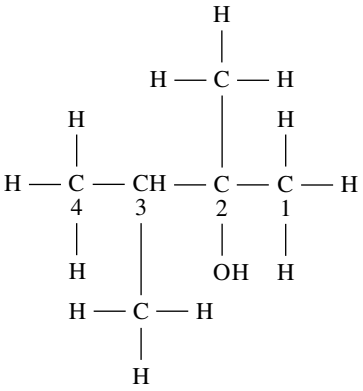
b The correct name is 2-methylheptan-3-ol. One of the mislabeled methyl branches is actually part of the longest carbon chain and the other methyl is actually on carbon number 2. So, what was thought to be carbon number 1 is actually carbon number 2 and, consequently, the hydroxyl group is on carbon number 3.



c The correct name is 3-methylbutanoic acid. The carboxyl group must always be carbon number 1, so the methyl group must be at carbon number 3, since it was 2 carbons away from the carboxyl group in the incorrect name.



d The correct name is 2,3-dimethylbutan-2-ol. Carbon number 3 should be labelled carbon number 2 as that is the end closest to the hydroxyl group.



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Chapter 8 Review

REVIEW QUESTIONS

Knowledge and understanding

1 C. While coal is a fossil fuel, it is found in separate locations to crude oil as a solid deposit, so alternative C is correct. During fractional distillation, petrol (alternative A) is derived from crude oil as part of the gasoline fraction, diesel (alternative B) comes from the kerosene fraction, and asphalt (alternative D) is a product of the final fraction referred to as unvapourised residues.

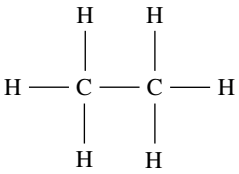
2 D. Unsaturated means that there are one or more carbon–carbon double bonds in the compound. Alternative D, but-2-ene, has one carbon–carbon double bond. Alternative A is a haloalkane, which only has single carbon-carbon bonds, and alternatives B and C are both alkanes, also with only single carbon–carbon bonds.

3 B. This statement is incorrect because it is missing one hydrogen atom. Alkanes have the general formula C_nH_{2n+2}. The correct statements, A, C and D can be found in the chapter in Sections 8.2 and 8.3.

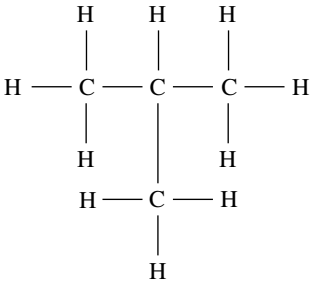
4 D. The weaker the intermolecular forces, the lower the boiling point. The homologous series with the lowest boiling point is the alkanes, because the molecules are non-polar, so only have dispersion forces between them. The next highest is haloalkanes because the presence of the halogen adds a polar bond, enabling some degree of dipole–dipole attraction. The next highest is the alcohol homologous series because there is hydrogen bonding between the molecules, which is a stronger intermolecular force than dipole–dipole attractions. Finally, carboxylic acids have the highest boiling point of the four listed homologous series, as the molecules form dimers, so the dispersion forces between the dimers is equivalent to a molecule with double the molecular mass.

All the other alternatives do not show these homologous series in increasing boiling point order.

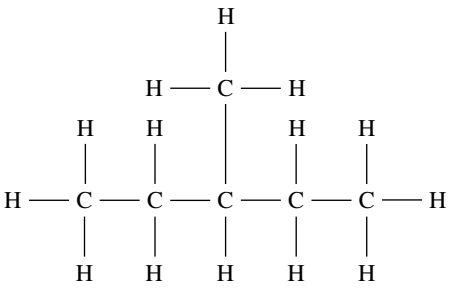
5 **a** ethane



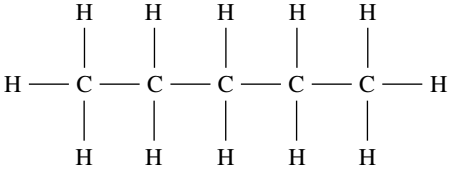
b 2-methylpropane



c 3-methylpentane



d pentane



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6 **a** Alkanes have the general formula C_nH_{2n+2}, and for this hydrocarbon $n = 16$, so $2n + 2 = 34$.
b $C_{17}H_{36}$. One more than 16 is 17, so for this hydrocarbon $n = 17$, and $2n + 2 = 36$.
c $C_{15}H_{32}$. One less than 16 is 15, so for this hydrocarbon $n = 15$, and $2n + 2 = 32$.
d $C_{16}H_{32}$. With one carbon–carbon double bond, the compound is an alkene, general formula C_nH_{2n}, so for this hydrocarbon $n = 16$, and $2n = 32$.

7 Remembering that the general formula for alkanes is C_nH_{2n+2} and for alkenes is C_nH_{2n}
a alkene
b alkane
c alkene
d alkane
e alkane

8 **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.

9 There are four structural isomers of $C_3H_6Br_2$: 1,1-dibromopropane, 1,2-dibromopropane, 2,2-dibromopropane and 1,3-dibromopropane.

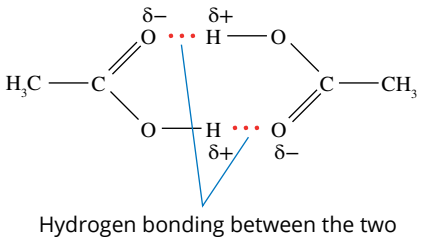
10 The following is an example of an answer. There are many possible answers to this question.

Homologous series	Particular compound	Use
alkanes	petrol (octane), diesel	fuel for transport
	methane	gas burnt in Bunsen burners
alkenes	ethene	monomer for making polymers
	but-1-ene	manufacture of polymers
haloalkanes	iodomethane	pesticide

11 **a** $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$
b $2C_6H_{14}(l) + 19O_2(g) \longrightarrow 12CO_2(g) + 14H_2O(l)$
c C_4H_{10}: $2C_4H_{10}(g) + 13O_2(g) \longrightarrow 8CO_2(g) + 10H_2O(l)$

12 Renewable means that there is a continuous source of the fuel, so as they are used up they can be replenished. There is only a limited supply of fossil fuels present in the Earth's crust as crude oil, because these were formed from the remains of prehistoric marine microorganisms, such as bacteria and plankton, which have been converted into hydrocarbons over millions of years. Because this process is no longer occurring, our supply of fossil fuels is not renewable.

13 While alcohols have hydrogen bonding between the molecules, carboxylic acid molecules pair up due to the hydrogen bonding between the carboxyl groups, forming dimers. As a result, the dispersion forces between the dimers are equivalent to a molecule with double the molecular mass, so the intermolecular forces are stronger than those between the equivalent alcohols.



Hydrogen bonding between the two carboxyl groups holds the two molecules together in a dimer.

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Application and analysis

14 C. If you track the molecule from left to right, structure C gives the correct order of carbons and hydrogens. Alternative A is incorrect because it combines the two CH_3 groups into C_2H_6. Alternative B is incorrect because it attributes the wrong number of hydrogens to some of the carbons in the structure and misses other carbons. Alternative D is almost correct, but it does not allow for the double bond between the last two carbons (shown on the right hand side), so it has too many hydrogens.

15 Using polyethene (polyethylene) made from sugarcane instead of from crude oil products means that the non-renewable crude oil products can be saved for use in making other products. Polyethene made from sugarcane is considered renewable and thus sustainable, because the sugarcane can be regrown quickly after it had been used.

16 **a** hex-2-ene. This is an alkene with 6 carbon atoms and the carbon–carbon double bond in the second position from the end.

b 2-methylbutane. The molecule has been drawn with the longest carbon chain bending downwards. The longest chain has 4 carbons and there is a methyl group on the second carbon from the end.

c 2,3-dimethylbut-1-ene. The longest carbon chain, containing the carbon–carbon double bond, which is in the first bonding position, has 4 carbons and there are two methyl groups on the second and third carbon from the end nearest the double bond.

d 2-methylpentan-3-ol. The molecule has been drawn with the longest carbon chain zig-zagging. The longest carbon chain has 5 carbons and there is a methyl group on the second carbon from the end. The hydroxyl group is on the third carbon from the end.

e 3-methylbutanoic acid. The longest carbon chain has 4 carbons and starts with the carboxyl group. There is a methyl group on the third carbon from the carboxyl group.

17 This compound is 3-methylpent-2-ene. The numbering of carbon chain must start at the end nearest the functional group (the carbon-carbon double bond), so the molecule is 3-methylpen-2-ene, rather than 3-methyl pent-3-ene.

18

	Alkanes	Alkenes	Haloalkanes	Alcohols
Solubility in water	does not dissolve	does not dissolve	small haloalkanes dissolve slightly	small alcohols dissolve well in water, due to formation of hydrogen bonds between the hydroxyl group and water molecules
Boiling point	low, increases with size of carbon chain	low, increases with size of carbon chain	higher than alkanes and alkenes, increases with size of carbon chain and with size of halogen atom or number of halogen atoms	higher than haloalkanes due to the strength of the hydrogen bonding between molecules
Bonding between molecules	dispersion forces	dispersion forces	dipole–dipole attraction and dispersion forces	hydrogen bonding and dispersion forces

19 The terms in bold have been corrected.

a **Alkanes** have two more hydrogen atoms per carbon atom than **alkenes**. / Alkenes have two **less** hydrogen atoms per carbon atom than alkanes.

b A haloalkane with five carbon atoms and one chlorine atom bonded to the end carbon could be called **1-chloropentane**. / A haloalkane with **three** carbon atoms and one chlorine atom bonded to the end carbon could be called 1-chloropropane.

c Pentane has **3** structural isomers.

d Alkanes are **saturated** hydrocarbons. / **Alkenes** are unsaturated hydrocarbons.

e The carboxylic acid with seven carbon atoms is called **heptanoic acid**.

f The **hydroxyl** functional group is found in alcohols and has the formula –OH.

g Compounds with the same molecular formula **can have molecules that are structural isomers of each other**.

20 **a** As the number of carbon atoms increases, the number of electrons in the molecules also increases, the attraction between temporary dipoles formed by the random movement of the electrons increases, and so the strength of the dispersion forces between the molecules increases. Because this intermolecular force of attraction has increased, more energy is needed to separate the molecules, and so the boiling point increases.

b Haloalkanes, alcohols or carboxylic acids would all have a higher boiling point than their corresponding alkane. The functional groups in these compounds enable the intermolecular forces to be stronger, so more energy is needed to separate the molecules than would be the case for alkanes, and so the boiling point is higher.

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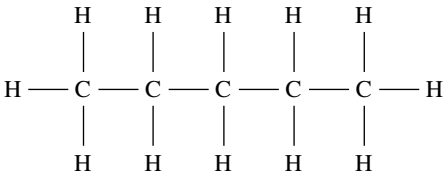
21 a 1-chloropropane is a molecule with three carbon atoms. While the two end (terminal) carbon atoms are equivalent to each other, the second carbon in the chain is not, so the number is required to distinguish whether the chlorine atom is bonded to an end carbon (designated as carbon number 1) or the second carbon in the chain. Chloroethane has only two carbon atoms which are equivalent to each other, so no distinction between the atoms is needed.

b 2,2-dibromopropane has a higher molecular mass, and more electrons than 2-bromopropane, so it has stronger dispersion forces between its molecules. As a result, more energy is needed to separate the molecules when the state changes from liquid to gas, and so it has a higher boiling point.

c While the hydroxyl group in ethanol is able to make hydrogen bonds with water molecules and enable the ethanol to dissolve in water, octan-1-ol has a much larger non-polar chain than ethanol (8 carbon atoms, compared to 2 carbon atoms), so the polar hydroxyl group has less influence over its solubility and so octan-1-ol cannot dissolve in water.

d In a carboxylic acid, the carboxyl group is always on carbon number 1 (by convention), so no number is needed for the carboxyl group. In comparison, the hydroxyl group can be anywhere along the carbon chain, so needs to be numbered in molecules larger than ethanol.

22 a



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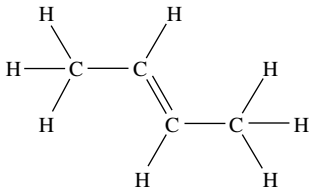
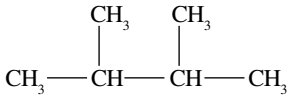


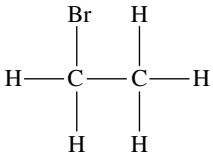
but–2–ene

c



2,3-dimethylbutane

d



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Chapter 9 Polymers and society

9.1 Polymer formation

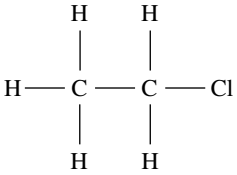
KEY QUESTIONS

Knowledge and understanding

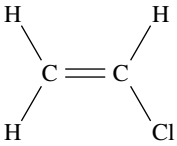
1 **a** An organic reaction where two or more molecules combine to form a larger molecule, without another product being formed.

b Monomers join when a covalent double bond (usually C=C) breaks. A very long molecule forms without the loss of another smaller molecule.

2 **a** chloroethane



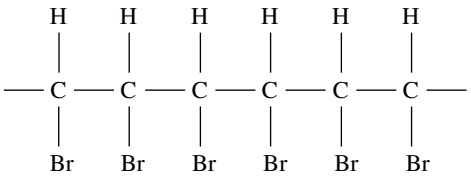
chloroethene



b Chloroethene can form a polymer due to its C=C double bond.

3 Polymers usually consist of thousands of monomer repeating units, so this would not be considered a polymer.

4

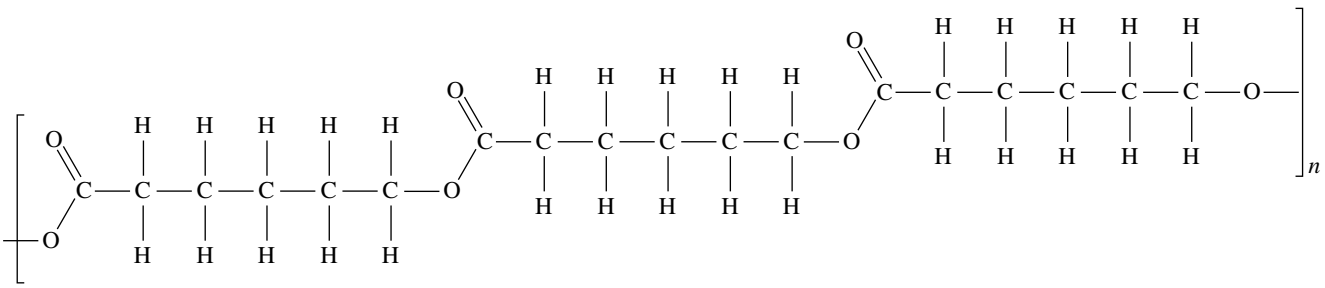


5 **a** Condensation reaction: Type of reaction in which two molecules combine to form a single molecule, usually with the loss of a small molecule such as water.

b Condensation polymerisation occurs between the functional groups on monomers, which react to form larger structural units while releasing smaller molecules (e.g. water) as a by-product.

Analysis

6



7 The repeating units from the monomer are $–CH_2CF_2–$ and $–CH_2CH_2–$ so there are 4 in this segment.

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8 **a** High-density polyethene (HDPE) is made of relatively unbranched chains of polyethene, which can pack together more closely than the chains of low-density polyethene (LDPE). LDPE contains branched chains of polyethene that cannot pack together as closely. HDPE is therefore stronger and slightly less flexible than LDPE. Both HDPE and LDPE are chemically unreactive, waterproof, non-conductors and only slightly permeable to gases.

b **i** LDPE
ii HDPE
iii LDPE

9 No. The melting point depends upon the length of the molecules and the degree of branching, which are related to manufacturing conditions These two factors are likely to vary between different manufacturers.

9.2 Thermoplastic and thermosetting polymers

KEY QUESTIONS

Knowledge and understanding

1 **a** thermosetting
b thermoplastic
c thermosetting
d thermoplastic

2 **a** covalent bonds within the chains and dispersion forces between chains
b dispersion forces between chains

Analysis

3 Thermosetting items are difficult to recycle as they do not melt. They are often disposed of in landfill. Thermoplastics that can be recycled offer better sustainability.

4 All the atoms in a thermosetting polymer are connected by strong covalent bonds, so it has a higher resistance to heat than a thermoplastic polymer. A thermoplastic polymer is made up of many individual polymer molecules held together by dispersion forces. The dispersion forces can be broken by heat, causing the polymer to melt.

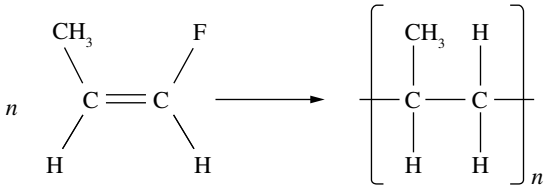
5 A thermosetting polymer is more likely to produce toxic gases when heated. If it is heated strongly, the covalent bonds will break and smaller, often toxic, compounds will form. Thermoplastic polymers initially melt when heated.

9.3 Designing polymers for a purpose

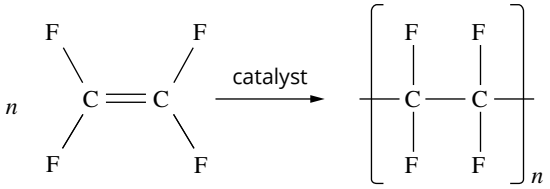
KEY QUESTIONS

Knowledge and understanding

1 **a** polypropene



b Teflon



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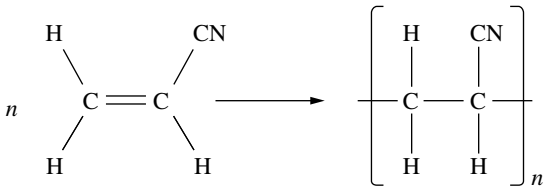
c polypropenenitrile



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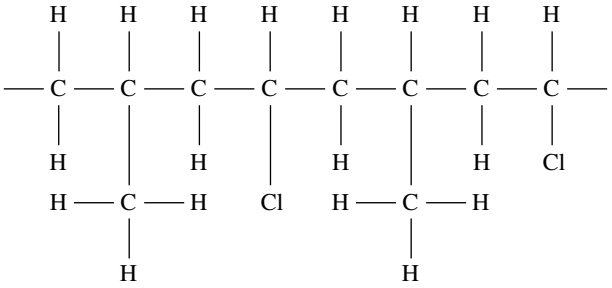

3 **a** 4

b The monomer used is 1,1-difluoroethene

4 –H (in polyethene)

–F (in Teflon)

–Cl (in polyvinyl chloride)

–C_6H_5 (in polystyrene)

–$NC_{12}H_8$ (in polyvinyl carbazole)

Analysis

5 **a** dispersion forces

b dipole–dipole attractions

c dispersion forces

d dispersion forces

e dipole–dipole attractions

6 **a** Lowest melting point to highest melting point will be polypropene, polychloroethene, polytetrafluoroethene (assuming the chains are arranged relative to each other in similar ways).

b The melting point of a polymer increases with the length of the molecules. A fair comparison of the impact of the monomer would require polymer chains of equivalent length, degree of branching and arrangement of the chains relative to each other.

7 **a** Pentane is added to polystyrene to make polystyrene foam. As the polystyrene is heated, the pentane turns to a gas, so the trapped gas in the polymer makes it a foam. Regular polystyrene does not contain any gas.

b Insulating, light weight, tough and inert.

8 **a** mass of polystyrene required to manufacture a cube of side length 10 cm = 950 g
($V = l^3 = 10^3 = 1000\,cm^3$: $m = d \times V = 1000 \times 0.95 = 950\,g$)
mass of polystyrene foam = $d \times V = 1000 \times 0.05 = 50\,g$

b The low density of a foam means that a small amount of polymer can be used to produce a large object.

9.4 Recycling plastics

CASE STUDY: ANALYSIS

Some choices are smarter than others

Analysis

1 coffee cup: convenient to use, but adds to landfill

starch-based container: can be made from waste and composts easily

biopolyethene cup: made from renewable waste, but adds to landfill

compostable spoon: drains resources, but does not add to landfill

PET drink bottle: resources used in manufacture, but does not contribute to waste or landfill

fork made from crude oil: uses resources in its manufacture and contributes to landfill

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2 **a** Factors to consider include: whether the source material is renewable or non-renewable, the energy required to manufacture the product, whether the product can be recycled or whether it is compostable. You would also need the packaging to serve its purpose of containing the food without having an impact on its taste.
b A ranking from most sustainable to least: starch-based container, compostable spoon, PET drink container, biopolyethene cup, biodegradable coffee cup, non-recyclable fork.

3 A class discussion is required here.

4 The recycling codes from 1 to 7 are used all across Australia. Container deposit schemes vary between states and territories.

KEY QUESTIONS

Knowledge and understanding

1 **a** Plastics need to have a code number from 1 to 7 on them, and their recycling status checked before they are added to the recycle bin. The numbers have the recycling triangle around them.
b thermoplastic

2 suitable for recycle bins: 2 L milk cartons, 600 mL plastic water bottles
not suitable for recycle bins: plastic wrap, polystyrene foam, Teflon tape

3 many possible answers: banning of plastic plates, replacement of foam take-away hamburger packaging, return of paper grocery bags, bring your own bags, bamboo cutlery, paper plates, keep cups

Analysis

4 PET drink bottles are easy to separate from waste and they are easy to remould into useful products. The majority of plastic water bottles are PET, so it is an obvious market to target.

5 Sorting machines are expensive, but they can do the job of many humans and often do the job more effectively. As the quality of the sorting technology improves, the potential gains increase. The sorting of garbage is not usually considered to be a pleasant job for humans. To progress to a circular economy, the product should be made from a renewable resource and the product needs to either be remoulded to a new product or compostable to turn into non-toxic substances.

6 Many recycling projects produce products of a lower grade, such as outside furniture and mats. The weaving of PET into clothing is producing a higher quality, more desirable product. The impact on scarce raw material supplies is alleviated through the use of recycled material.

9.5 Innovations in polymer manufacture

KEY QUESTIONS

Knowledge and understanding

1 **a** bio-plastic: a polymer produced from living organisms or biomass
b bio-monomer: a monomer produced from living organisms or biomass
c compostable: product that will degrade within 90 days
d microbial: a process or reaction caused by microorganisms

2 Biodegradable has proved to be a misleading term. Many biodegradable items do not actually degrade in landfill. They require elevated temperatures or added bacteria to degrade within a reasonable time period. The term compostable has been introduced to apply to plastics that generally degrade in normal landfill conditions in a reasonable time.

3 **a** sugar cane waste
b dairy waste
c potato-processing waste
d any plastic waste

Analysis

4 **a** For a polymer to be soluble in water, its side chain needs to have significant dipoles. A repeating –OH group is an example of the dipole necessary.
b casing around capsules, soluble stitches

5 Composting leads to breakdown of a molecule. Sometimes the products of breakdown are toxic and damaging to the environment. Composting might also produce significant levels of greenhouse gases such as methane and CO_2.

6 **a** The action of yeast on biomass can cause the carbohydrates to ferment to ethanol and CO_2.
b If polyethene uses ethanol made from biomass, it means less crude oil is being used. This preserves scarce resources and it uses up waste biomass.

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7 **a** A: example of a linear economy—drains scarce resources and creates landfill.
B: potato starch is a waste material so it is good to find a use for it. However, producing landfill is not helpful.
C: the addition of some lactic acid will lower the environmental impact of the polymer, especially if it leads to a compostable product. Less crude oil is used.
D: crude oil is used, but the final product does not end up in landfill.

b Perhaps option C offers the most benefit. The lactic acid used for the copolymer can be made from biomass waste so it does not drain resources and it composts to non-toxic substances.

8 **a** mechanical recycling: PET is heated and remoulded into a new shape. The polymer molecules have not been altered. New PET drink bottles or faux fur jackets are examples of mechanical recycling.

b organic recycling: microorganisms are able to break the ester bonds in PET to produce smaller and less toxic molecules. PETase is the organism used.

c chemical recycling: heat and catalysts are used to break the long polymer molecules and to form smaller molecules that are a form of synthetic oil. It is possible to re-form the same monomers to make further PET, but the synthetic oil can be used to make other monomers.

Chapter 9 Review

REVIEW QUESTIONS

Knowledge and understanding

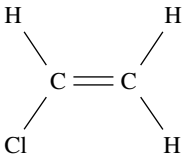
1 B. Increasing the length of the molecule chains leads to stronger forces between molecules. The other options lower the forces between molecules.

2 D. The presence of electronegative chlorine atoms adds to the intermolecular bonding. PVC does not have cross-links.

3 C. HDPE is denser, harder and less transparent. Options A and B are not correct as the branching in LDPE disrupts the packing of the molecules, lowering the density and melting point.

4 All except B. The monomers need to be unsaturated.

5 **a**



b 6

c Dipole–dipole attractions.

6 **a** Monomers are small molecules that are able to react to form long chains of repeating units, called polymers. They often contain a carbon–carbon double bond.

b When a thermoplastic is heated, the bonds between molecules are broken and the molecules become free to move, so the plastic melts.

c When a thermosetting is heated it does not melt, but at high temperatures covalent bonds are broken and the material decomposes or burns. It cannot be moulded into a different shape.

d A cross-link is one or more covalent bonds that connect neighbouring polymer chains.

7 **a** false

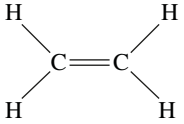
b false

c false

d true

8 **a** The ethene molecule has a carbon-to-carbon double bond.

b



c Ethane cannot undergo addition polymerisation because it is a saturated compound.

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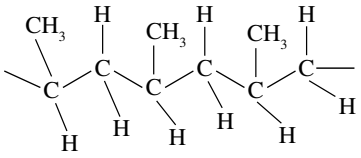
9 a



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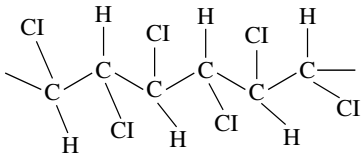

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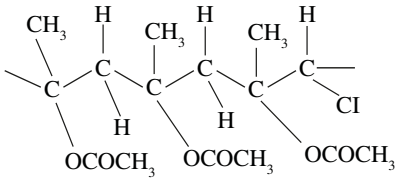
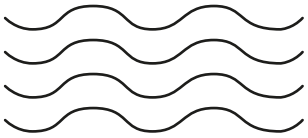


10 a i



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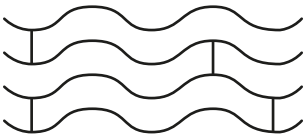

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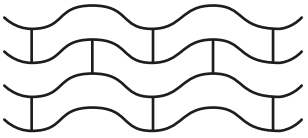
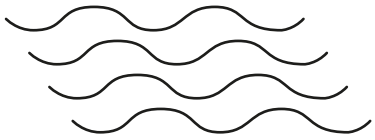
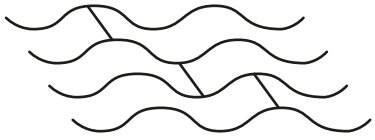


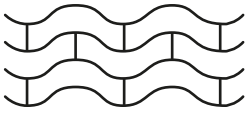
b



polymer easily stretched



Some stretching but polymer tends to return to its original shape.



no stretching possible

11 a Thermosetting polymers have much stronger inter-chain bonds than thermoplastic polymers.

b The inter-chain bonds in thermosetting polymers are about the same strength as the covalent bonds within chains. When heated strongly both the inter-chain and within-chain bonds break.

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12 **a** thermosetting
b thermoplastic
c thermosetting
d thermoplastic
e thermoplastic

13 **a** The polymer used in saucepan handles
b Saucepan handles are much harder and more resistant to the effect of heat than elastic bands. The polymer on the outside of the golf ball is also very hard, but, unlike the polymer in saucepan handles, is quite elastic.

14 Look for 1 to 7 recycle codes on the product itself, not just the packaging. Expanded polystyrene foam should not be recycled.

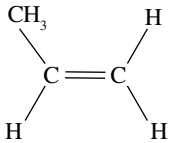
15 **a** A polymer formed when more than one type of monomer is used.
b A copolymer can display a blend of the best properties of each monomer. Sometimes only a small amount of a second monomer significantly improves the performance of the polymer.

16 Thermoplastic, as it can be melted and remoulded.

17 **a** biobased—a material made from substances derived from living organisms
b biodegradable—potentially will degrade under the right conditions
c compostable—can degrade in a useful time frame, usually less than 90 days.

Application and analysis

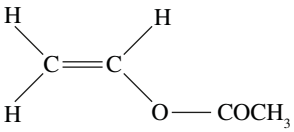
18 **a**



propene

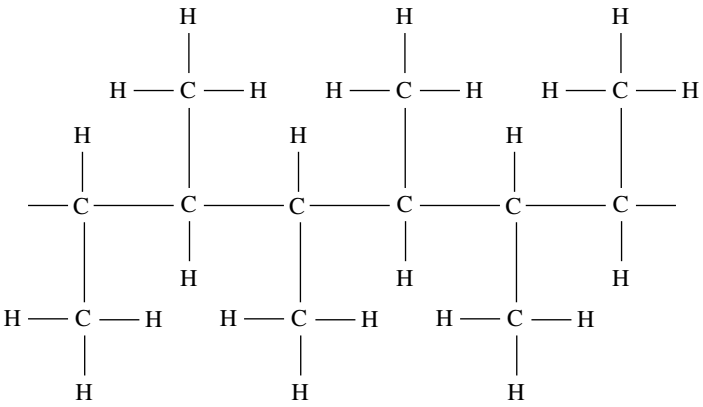
b 3
c The melting point is likely to be higher than polyethene due to the larger side group (assuming similar chain lengths and degrees of branching).
d This polymer is recyclable – separate the items made from this polymer and remould them.

19 **a**



b PVA contains oxygen atoms in its side-group. They add a significant dipole to the molecule and allow hydrogen bonds to form with water.
c Being water soluble means the polymer might fail in a moist environment. It will degrade more easily, however, than most polymers.

20



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21 Left: C_8H_8 (styrene); middle: C_3H_3N (propenenitrile); right: C_4H_6 (buta-1,3-diene)

22 **a** The melting point and the toughness will increase with longer polymer molecules

b Incorporation of a large side group will make the polymer tougher and harder.

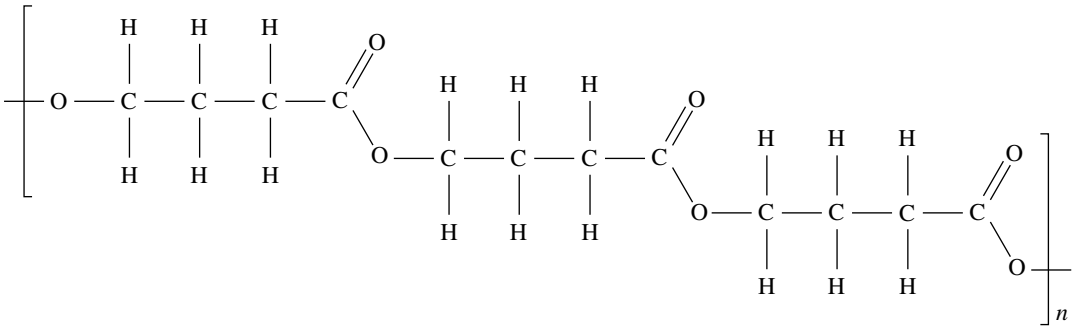
c A foaming agent will expand the polymer material significantly. It will have a much larger volume and a lower density.

23 The relative molecular mass would have increased. The melting point of the polymer would have increased. The overall strength of inter-chain forces would have increased. The electrical conductivity of the polymer would have remained the same.

24 **a** Answer to include—polyethene: crude oil fractional distillation to isolate ethane. Ethane cracked to ethene. Ethene polymerised to polyethene. Biopolyethene—carbohydrate ferments to ethanol. Ethanol cracked to ethene. Ethene polymerised.

b World consumption of polyethene is extremely high—there is not enough waste biomass to support the large-scale, rapid production of bioethanol. Fermentation is a slow reaction.

25 **a**



b Condensation and addition polymers both have long molecules. The characteristic structure of the chain, however, will differ. Condensation polymers are likely to have some oxygen or nitrogen atoms in the chain, while the chains of addition polymers contain carbon atoms only. Addition polymers form when unsaturated monomers react, while condensation polymers form when the functional groups on each end of the monomers react.

c The chain of a condensation polymer often contains oxygen and/or nitrogen atoms. These atoms are susceptible to attack from microorganisms.

26 PLA can be manufactured from biomass and it is fully compostable to non-toxic substances. Its production is not draining scarce resources and it is not adding to landfill. Its use leads to little impact on the environment.

27 **a** Most polyethene is made from non-renewable crude oil. Once used, it is added to landfill. Therefore, it is a drain on resources and a problem in landfill. This is an example of the linear economy in action, often referred to as 'take-make-dispose'.

b A biomonomer is made from biomass of some form. The most common source is bioethanol made by fermentation. This removes the problem of polyethene being a drain on non-renewable resources.

c In mechanical recycling, the polyethene is shredded and remoulded into a new product. The fraction of polyethene recycled has been increased since supermarkets have started collecting soft plastic wrap.

d Licella heats plastic material under pressure. The addition of a catalyst leads to the polymer being broken down to smaller, synthetic oil. This is an example of chemical recycling, where the polymer structure is changed in the recycling process.

e Addition polymers have only carbon atoms in their chains. Microorganisms cannot break the bonds in the chains, so the polymer cannot undergo organic recycling.

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Unit 1 Area of Study 2

How are materials quantified and classified?

Multiple-choice questions

1 C. By definition, the empirical formula shows the simplest mole ratio of each element in the compound.

2 A. It is an alkane with 3 carbons, so propane. There is only one way to arrange 3 carbon atoms, so there are no isomers. Therefore, no number prefix is used.

3 D. The molecular formula for the compound is $C_5H_{11}OH$.

3.5 mol of $C_5H_{11}OH = 3.5 \times 6.02 \times 10^{23} = 2.1 \times 10^{24}$ molecules

As each molecule of pentan-1-ol has 18 atoms:

total number of atoms $= 18 \times 2.1 \times 10^{24} = 3.8 \times 10^{25}$

4 B. Crude oil is mainly a mixture of alkanes.

5 C. If 0.235 mol has a mass of 16.0 g, then 1 mol will have a mass of $\frac{16}{0.235} = 68\,\text{g}$. So the molar mass of the compound is 68 g mol^{-1}. The only compound with this molar mass is C_5H_8.

6 D. $n(C) : n(H)$ in the compound $= \frac{80.0}{12.0} : \frac{20.0}{1.0} = 6.667 : 20 = 1 : 3$

So the empirical formula is CH_3. This cannot be the molecular formula as carbon needs to form 4 bonds. C_2H_6 is the only formula to represent a molecule with empirical formula CH_3.

7 C. There are 3 compounds: but-1-ene, but-2-ene and 2-methylprop-1-ene.

8 A. Alcohols contain the polar –OH group, so molecules are attracted to each other by hydrogen bonds, as well as by dispersion forces. Therefore, alcohols have a higher boiling point than similar-sized alkanes.

Therefore, ethanol has a higher boiling point than propane.

The strength of dispersion forces between molecules depends on the size of the molecules. The forces are weaker between ethane molecules than between propane molecules, so ethane has a lower boiling point than propane. Similarly, ethanol has a lower boiling point than propan-1-ol.

9 B. Molybdenum has a relative atomic mass of 96.0, compared to the value of 12 for carbon-12. This is a ratio of 8 : 1. Each atom of molybdenum will weigh eight times more than each carbon-12 atom. The mass ratio will be the same no matter what quantity of atoms we are weighing.

10 A. The double bond breaks and the monomers join together.

11 C. A molar mass of 56 g mol^{-1} corresponds to a molecular formula of C_4H_8. This is likely to be a form of butene. The product after reaction with bromine is a symmetrical molecule, so the original molecule must have the C=C double bond in the middle, making it but-2-ene.

12 B. Higher density is the result of closer packing of polymer chains. Polymer chains can pack more closely if they have little or no branching. The closer the chains are packed, the more significant the dispersion forces between the chains and therefore the higher the softening temperature.

Short-answer questions

13 **a** Crude oil is fossilised organic material, mostly of plant and microbial origin.

b **i** It is a mixture of hydrocarbons, mostly alkanes from C1 to about C70.

ii Any correct formulas of two alkanes with general formula C_nH_{2n+2}; e.g. C_5H_{12}, C_6H_{14}.

14 **a** Magnesium has three different isotopes, ^{24}Mg, ^{25}Mg and ^{26}Mg. The most abundant of these isotopes is ^{12}Mg.

b $A_r = \frac{(24 \times 0.79) + (25 \times 0.10) + (26 \times 0.11)}{100} = 24.3$

15 **a** Each molecule has four phosphorus atoms, therefore number of mol $= 5 \times 4 = 20$ mol

b Each molecule has 14 atoms, therefore $5 \times 14 = 70$ mol

c Number of atoms = number of mole $\times 6.02 \times 10^{23} = 70 \times 6.02 \times 10^{23} = 4.21 \times 10^{25}$

16 **a** but-2-ene

b heptane

c 2,2,3-trimethylpentane

d butanoic acid

e 2-methylpropan-1-ol

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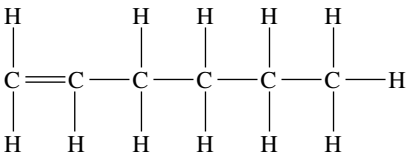
17 a



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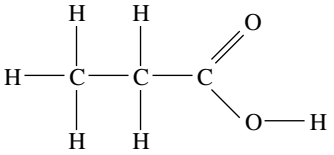

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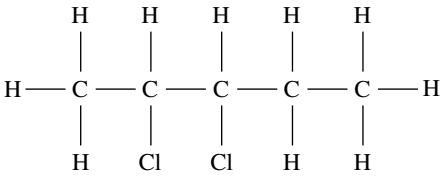

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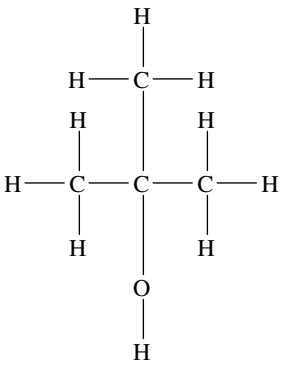


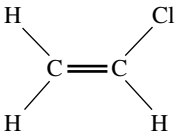
18 a $CH_3(CH_2)_4COOH$ or $CH_3CH_2CH_2CH_2CH_2COOH$

b $CH_3C(CH_3)_2CH_2CH(C_2H_5)$ $CH_2CH_2CH_3$ or $(CH_3)_3CCH_2CH(C_2H_5)CH_2CH_2CH_3$

c $CH_3CH(OH)CH_2CH_3$

d $CH_2C(CH_3)CH_2CH(CH_3)CH_2CH_3$

19 a



b i The plastic will char and blacken.

ii Thermosetting polymers have strong cross-links between polymer chains so the layers cannot slide past each other and melt.

20 a A monomer is a small molecule that is able to bond with other monomers to form a long chain molecule called a polymer; e.g. ethene molecules are the monomers that join together to form the polyethene polymer.

b A thermoplastic polymer is one that softens on heating, but becomes hard again when cooled, e.g. polyethene.
A thermosetting polymer is one that doesn't soften on gentle heating, but if heated sufficiently it will char, e.g. urea-formaldehyde.

c A branched polymer is a linear polymer that has some of the atoms forming branches attached to the polymer backbone, e.g. low density polyethene. Relatively weak intermolecular forces exist between chains.
A cross-linked polymer has covalent bonds linking polymer chains, resulting in a rigid polymer that does not soften on heating, e.g. urea-formaldehyde.

d Addition polymer: long molecule formed when the C=C double bonds in the monomer break, allowing the monomers to join to each other. Polyethene is an example.
Condensation polymer: long molecule formed when the functional groups on the ends of the monomers react, allowing the monomers to join and a small molecule such as water is also formed. Polyester is an example.

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21 Percentage by mass of oxygen in the compound = 100 – 59.7 = 40.3%

In 100 g of compound:

$m(O) = 40.3\,g$ and $m(X) = 59.7\,g$

So, $n(O) = \frac{40.3}{16.0} = 2.52$ mol

As ratio of X:O = 2 : 3, $n(X) = \frac{2}{3} \times 2.52 = 1.68$

So 59.7 g of X is equivalent to 1.68 mol.

And then $M(X) = \frac{m(X)}{n(X)} = \frac{59.7}{1.68} = 35.5\,g\,mol^{-1}$

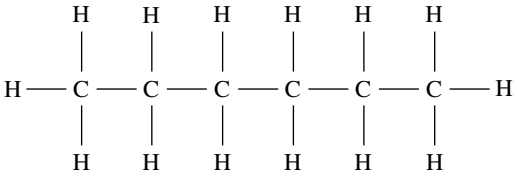
The element must be Cl.

22 **a** N_2O will have the highest percentage of nitrogen. $(\% = \frac{2 \times 14}{(2 \times 14 + 16)} \times 100 = 63.6\%)$

b NO_2 and N_2O_4 will have the same nitrogen content as they have the same empirical formula.

c Each molecule contains 7 atoms => number of atoms = $7 \times 6.02 \times 10^{23} = 4.21 \times 10^{24}$ atoms

23 **a** Hexane molecules have no significant dipoles as the structure only contains carbon and hydrogen atoms. The very small dipoles will cancel each other out as well.



b To be used in the food industry, the molecule must have no long-term health impact on humans. Some hydrocarbons are toxic to humans. Also, hexane is a liquid; some smaller alkanes would also be non-polar, but not suitable as a solvent in the gas state.

24 **a** In 100 g of compound, the mass of hydrogen = 100 – 54.54 – 36.36 = 9.10 g

$= n(C) : n(H) : n(O)$

$= \frac{54.54}{12.0} : \frac{9.10}{1.0} : \frac{36.36}{16.0}$

$= 4.55 : 9.10 : 2.27$

$= 2 : 4 : 1$

So empirical formula is C_2H_4O.

b $n = \frac{m}{M}$, so $M = \frac{m}{n} = \frac{30.8}{0.350} = 88\,g\,mol^{-1}$

c Relative formula mass of $C_2H_4O = 44$

Molar mass = 2 × formula mass

So molecular formula is $C_4H_8O_2$.

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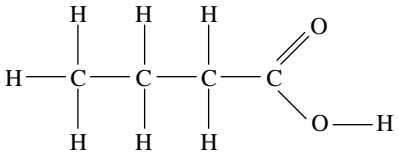
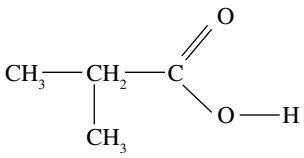
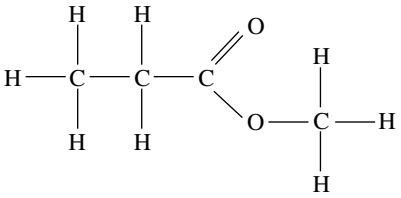
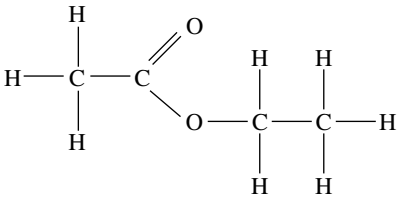
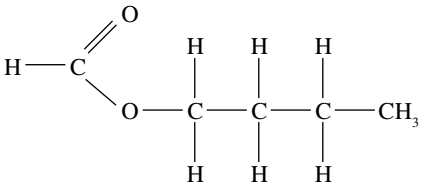
d (Students should be able to draw the two structures shown under the heading Compound 1. Some students might also be aware of another group of molecules shown under the heading Compound 2, which are not studied until Unit 4.)

Compound 1	**Compound 2**
butanoic acid  or 2-methylpropanoic acid 	methyl propanoate  or ethyl ethanoate  or propyl methanoate 

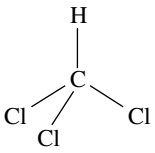
25 a m(H) in the compound = 4.738 – (0.476 + 4.22) = 0.042 g

so, n(C) : n(H) : n(Cl) = $\frac{0.476}{12.0} : \frac{0.042}{1.0} : \frac{4.22}{35.35}$ = 0.040 : 0.042 : 0.119 = 1 : 1 : 3

So, the empirical formula is $CHCl_3$.

b $n = \frac{m}{M}$, so molar mass of compound = $\frac{m}{n} = \frac{39.0}{0.326}$ = 120 g mol^{-1}

Formula mass of $CHCl_3$ is 119.5 g mol^{-1}, so the molecular formula is the same as the empirical formula.

c 

d There are three polar C–Cl bonds in the molecule; the molecule is not symmetrical, so it is polar overall. The interactions between the molecules are dipole–dipole interactions (as well as dispersion forces).

e n(Cl) in the sample = $\frac{7.2 \times 10^{23}}{N_A} = \frac{7.2 \times 10^{23}}{6.02 \times 10^{23}}$ = 1.2 mol

$n(CHCl_3) = \frac{1}{3} \times 1.2$ = 0.40 mol

So $m(CHCl_3) = n \times M$ = 0.40 × 120 = 48 g

26 a $\%N = \frac{2 \times N}{(2 \times N) + (4 \times H) + (3 \times O)} = \frac{2 \times 14}{(2 \times 14) + 4 + (3 \times 16)} \times 100$ = 35%

b mass in 5 kg bag = $\frac{85 \times 5}{100}$ = 4.25 kg

c mass of N = $\frac{4.25 \times 35}{100}$ = 1.49 kg

d n(N) = $\frac{1490}{14}$ = 106 mol

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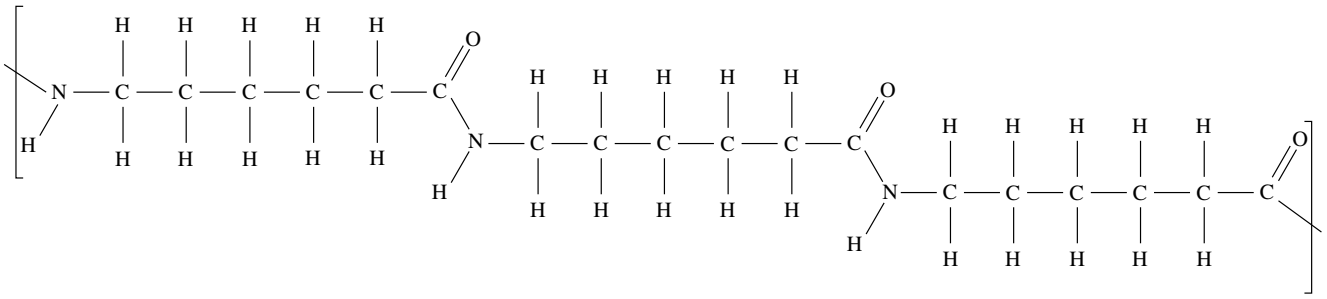
27 a



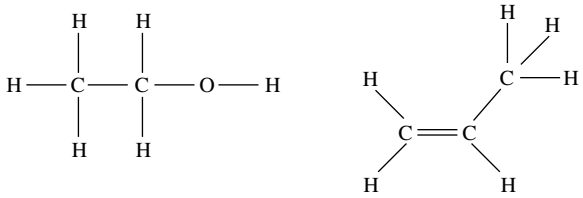
b The number of water molecules = number of monomers – 1 = 3000 – 1 = 2999

c The polymer will have nitrogen and oxygen atoms in the polymer chains. These atoms are susceptible to the action of microorganisms. Addition polymers like polyethene have only carbon atoms in the chains.

28 a Casein will be a condensation polymer. Natural polymers form from reactions between functional groups.

b Casein is a good example of a circular economy. It can be made from waste milk, so non-renewable crude oil is not required. As it degrades readily in water, it composts easily to form harmless products. Casein is therefore a far more sustainable option, both in source material and waste created.

29 a



b i ethanol: $\%O = \frac{16}{46} \times 100 = 34.8\ \%$

ii ethanoic acid $\%O = \frac{32}{60} \times 100 = 53.3\%$

c Ethanoic acid has a higher boiling point. Both molecules can form hydrogen bonds, but two hydrogen bonds can form between each pair of ethanoic acid molecules, whereas each pair of ethanol molecules have only one hydrogen bond between them.

d Many microorganisms function best around 35°C. If the temperature is too high, the organism might be killed. Reactions with chemicals derived from crude oil are not limited by the impact on living organisms.

e The production of either ethanol or ethanoic acid from biomass offers many advantages. Biomass is often a waste material, so it is good to find a use for the waste. If biomass is not used, scarce crude oil resources have to be used instead.

30 a i independent variable: mass of magnesium

ii dependent variable: mass of ash

b The precision of most data, except the data of student D, is high. The results of the other six students have produced a relatively linear graph. Student D's results can be viewed as an outlier, where the mass obtained is lower than expected.

c i Very close to 0.800 g

ii The mass of oxygen is the ash mass – magnesium mass = 0.800 – 0.500 = 0.300 g. Therefore the empirical formula: $n(\text{Mg}):n(\text{O})\ \frac{0.5}{24.3} : \frac{0.3}{16} = 0.021 : 0.019$. This is closer to 1 : 1 than any other ratio, so formula is MgO.

iii The mole ratio is not close enough to give 1 : 1. It is often found in this experiment that some magnesium oxide escapes, leaving the mass of oxide a little lower than it should be. This means the mass of oxygen will be low and the ratio drifts towards 1.1 : 1 or 1.2 : 1

d MgO, as magnesium has an electrovalency of 2+ (2 electrons in the outer shell of its atoms), while oxygen has an electrovalency of 2– (6 electrons in the outer shell), and forms O^{2-} ions.

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