Unit 1

Area of study 1 How can knowledge of elements explain the properties of matter?

Multiple-choice questions

1 D. In an atom, there are forces of attraction between the positively charged nucleus and the negatively charged electrons. Also there are forces of repulsion between positively charged protons in the nucleus and between negatively charged electrons surrounding the nucleus.

2 C. All atoms of the same element have the same number of protons in their nuclei.

3 B. All atoms of copper have the same atomic number. However, isotopes differ in the number of neutrons in the nucleus. The number of neutrons is the mass number minus the atomic number, so an atom of copper could have either 34 neutrons (63 – 29) or 36 neutrons (65 – 29).

4 C. Elements in the same group of the periodic table have similar properties as they have the same number of valence electrons. Only N and P belong to the same group.

5 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.

6 D. For main group elements, the group number is equal to the number of valence electrons, i.e. number of electrons in the outer electron shell.

7 D. In general, for the first series of transition metals, the 4s-subshell fills before the 3d-subshell. However, Cu and Cr are exceptions. Cr is more stable with a half-filled 4s-subshell and 5 electrons in its 3d-subshell.

8 C. A sulfur atom has 16 electrons so a S^{2-} ion has 18 electrons. D is not correct because the 3p-subshell contains 6 electrons in the ground state.

9 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.

10 B. Isotopes have the same atomic number but different mass numbers.

11 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.

12 B. The reactivity of metals increases down a group and decreases across a period.

13 A. Transition metals are dense and hard.

14 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.

15 C. Ionic bonding involves the transfer of electrons, not the sharing of electrons, from a metal to a non-metal.

16 B. 1 mol of NaOH contains 2 mol of ions (1 mol of Na^+ ions and 1 mol OH^- ions).

1 mol of K_3PO_4 contains 4 mol of ions (3 mol of K^+ ions and 1 mol of PO_4^{3-} ions).

So 0.5 mol of K_3PO_4 contains the same number of ions as 1 mol of NaOH.

17 A. Emission spectra involve the release of energy. When electrons at higher energy levels fall back to a lower energy level, the excess energy is emitted and can be observed as lines of light of particular frequencies on an emission spectrum.

18 A. 1 mole of CO_2 has a mass of 44 g (12 + 2 × 16 = 44 g)

So 6.02×10^{23} molecules have a mass of 44 g

And 1 molecule has a mass of $\frac{44}{6.02 \times 10^{23}} = 7.3 \times 10^{-23}$ g The contents of this box has changed.

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19 A. $m(\text{urea})$ in 125 g fertiliser $= \frac{58.5}{100} \times 125 = 73.1$ g

% of N in urea $= \frac{M(\text{N})}{M(\text{urea})} \times 100 = \frac{28.0}{60.0} \times 100 = 46.7\%$

So $m(\text{N})$ in 125 g fertiliser $= \frac{46.7}{100} \times 73.1 = 34.1$ g

20 B. 6.02×10^{23} is equivalent to 1 mol of Cl^- ions. The formula for calcium chloride is $CaCl_2$, so 0.5 mol $CaCl_2$ contains 1 mol of Cl^- ions.

The molar mass of $CaCl_2$ is 111.1 g mol^{-1}. So 0.5 mol will have a mass of 55.6 g.

Short-answer questions

* Represents 1 mark

21 **a** K_3PO_4
b Al_2O_3
c $NaNO_3$
d Fe_2S_3

22 **a** Calcium nitrate
b Copper(II) carbonate
c Magnesium hydroxide
d Chromium(III) oxide

23 **a** One mark for correct formula, one mark for appropriate use; e.g. NaCl preserving food, $NaHCO_3$ as baking soda, NaF in toothpaste to harden tooth enamel.

b **i** Although charged particles are present, they are held in fixed positions in the lattice and so cannot move to carry a current.

ii If a strong force is applied to a crystal of the compound, the layers of ions will move relative to one another, causing ions of like charge to be adjacent and hence repel. The crystal thus shatters.

iii A large amount of energy is needed to overcome the electrostatic attraction between oppositely charged ions.

24 **a** Mass spectrometer

b $A_r = \frac{(37.30 \times 190.97) + (62.70 \times 192.97)}{100} = 192.2$*

(Answer to 4 sig figs consistent with data to the smallest number of sig figs.)

25 **a** For example: aluminium is extracted by electrolysis of molten alumina. Huge amounts of electricity are required.

b **i** Harder and more brittle
ii Harder
iii Softer and more ductile

c Quenching is heating the metal to a moderate temperature and then cooling very quickly. Tempering involves first quenching a metal and then heating it again to a lower temperature.

d In annealing, larger crystals forms and in quenching tiny crystals form.

26 **a** $1s^22s^22p^63s^2$

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 there are fewer protons in the nucleus and so the attraction for the outer shell electron would be weaker, pulling it 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 electrons and cations.

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

iii The model cannot explain the differences in melting points/densities/electrical conductivities/magnetism between metals.

d **i** Observations: bubbles of gas evolved or increase in temperature

$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

or $Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

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

e Transition elements have an incomplete d-subshell of electrons.

f An alloy is a substance formed when other materials, usually carbon or other metals, are mixed with a metal. Alloying increases the hardness of a metal.

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27 **a** When atoms absorb or emit energy, a spectrum is obtained.

Emission and absorption spectra are not continuous but contain discrete lines corresponding to discrete amounts of energy emitted or absorbed as electrons move from one energy level to another.

Bohr proposed that electrons can only move in fixed orbits or shells of particular energy and could only jump from one orbit to another without stopping in between; thus, only certain energies can be emitted or absorbed.

b It does not explain why electrons move in circular orbits or why shells have particular energies. The mathematical model does not hold well for larger atoms.

c Electrons behave as waves around the nucleus. There are no definite orbits, instead electrons are thought to move in regions of space called orbitals. Electron shells can be further divided into subshells.

28 **a** Sodium has 11 electrons, which are distributed in 3 main shells or energy levels: 2 electrons in the first shell closest to the nucleus, 8 electrons in the second shell and one in the third shell.

Shells are divided into subshells labelled s, p, d and f. Two of sodium's second-shell electrons are in the s-subshell and the remaining 6 in the p-subshell.

s-, p-, d- and f-subshells are further divided into orbitals, which are regions of space in which electrons may be found. The s-subshell consists of only one orbital, the p-subshell has 3. The 6 electrons of sodium in the 2p-subshell are distributed in pairs in the 3p-orbitals.

b An orbital can hold a maximum of two electrons.

c An excited atom because the 3d-subshell is a higher energy subshell than 4s, which should be filled first.

29 **a** **i** Ca
ii Ar
iii C
iv Na or Mg
v Li
vi N
vii F

b 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.

30 **a** The total number of protons and neutrons in an atom

b The total number of protons in an atom

c The mass of an atom of the isotope relative to the mass of an atom of ^{12}C taken as 12 units exactly

d 6.02×10^{23}, which is the number of particles in a mole

e An atom, or a group of atoms, that has either gained or lost one or more electrons and so carries a charge

31 **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 core 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 core charge increases from left to right across the periodic table. As core 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.

32 **a** $n(Al(NO_3)_3) = \frac{m}{M} = \frac{30.5}{213} = 0.143$ mol

b $n(N) = 3 \times 0.143 = 0.430$ mol

c Total number of mol of atoms = $13 \times 0.143 = 1.86$

Total number of atoms = $1.86 \times 6.02 \times 10^{23} = 1.12 \times 10^{24}$

d $\%\ O = 9 \times \frac{16}{213} \times 100 = 67.6\%$

33 **a** $m(O)$ in compound = $2.127 - 1.704 = 0.423$

$n(Cu) : n(O) = \frac{1.704}{63.4} : \frac{0.423}{16} = 0.0268 : 0.0264 = 1:1$

So empirical formula is CuO.

b If the sample was a mixture of CuO and Cu_2O, the ratio of Cu to O would be greater and so the mass of Cu extracted would be greater than 1.704 g.

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34 **a** $m(\text{O})$ in the compound = 1.368 – 0.753 = 0.615 g

$n(\text{O})$ in the compound = $\frac{0.615}{16}$ = 0.03844 mol

As the empirical formula is XO_2:

$n(\text{M})$ in the compound = $\frac{1}{2}$ × 0.03844 = 0.01922 mol

As $n = \frac{m}{M}$, $M(\text{X}) = \frac{m(\text{X})}{n(\text{X})} = \frac{0.752}{0.01922} = 39.1\ \text{g mol}^{-1}$

b The metal must be potassium.

35 % N in KNO_3 = $\frac{14}{101.1}$ × 100 = 13.8 %

So $m(\text{N})$ in 65.0 g KNO_3 = 9.00 g

% N in $(NH_4)_2SO_4$ = $\frac{28.0}{132.0}$ × 100 = 21.2%

So $m((NH_4)_2SO_4)$ that contains 9.00 g N = $\frac{9.00}{21.2}$ × 100 = 42.4 g

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Unit 1

Area of Study 2 How can the versatility of non-metals be explained?

Multiple-choice questions

1 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. Hence one X atom would bond with 2 Y atoms.

2 C. The structure of HCN is H–C≡N:. There is a triple bond between C and N. Therefore 6 electrons are shared between these two atoms.

3 C. By definition, the empirical formula shows the simplest mole ratio of each element in the compound.

4 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 $= 18 \times 2.1 \times 10^{24} = 3.8 \times 10^{25}$

5 B. Crude oil is a mixture of alkanes.

6 C. If 0.235 mol has a mass of 16.0 g, then 1 mol will have a mass of $\frac{16}{0.235} = 68$ g. So the molar mass of the compound is 68 g mol^{-1}. The only compound with this molar mass is C_5H_8.

7 D. $n(\text{C}) : n(\text{H})$ 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.

8 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.

9 C. All alkynes have names ending in -yne. All alkynes have a triple bond between two carbon atoms, so the smallest alkyne must have two carbons. A two-carbon hydrocarbon has a name with the prefix eth-.

10 B. The general formula for alkynes is C_nH_{2n-2}. B is an alkane.

11 C. There are 3 compounds: but-1-ene, but-2-ene and methylpropene.

12 A. Diatomic molecules, e.g. HF, can only be linear. CO_2 is also linear because it has 2 bonding regions (each a double bond). These repel and so are oriented as far from one another as possible, i.e. at 180°.

H_2S is V-shaped. The sulfur atom has two bonding pairs of electron and two non-bonding pairs. The four pairs of electrons take on a distorted tetrahedral arrangement that results in a V-shaped structure. CH_4 has 4 C–H bonds in a tetrahedral structure.

13 D. HF, OF_2 and NF_3 are all polar molecules due to the polar nature of the bonds with fluorine. So there are dipole–dipole attractions between molecules. In CF_4 the polarising effects the individual C–F bonds are cancelled out due to the symmetrical tetrahedral structure of the molecule, so it is a non-polar molecule so there are only dispersion forces between molecules.

14 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. 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.

15 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.

So 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.

16 A. The double bond breaks and the monomers join together.

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17 D. In general, ionic compounds have a high melting temperature and covalent molecular compounds have a lower melting temperature. However, there are exceptions; e.g. in diamond, the C atoms are held together by covalent bonds and diamond has a very high melting temperature. This is also the case for hardness and brittleness.

In ionic compounds in the solid state, the ions are held in fixed positions and so cannot carry a charge. However, in the molten state, the ions are free to move and so the compound is an electrical conductor in that state. Covalently bonded compounds have no charged particles and so are non-conductors.

18 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 so the higher the softening temperature.

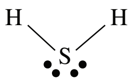
19 B. Crystalline regions are formed when polymer chains line up parallel to one another. This allows closer packing of the chains and more significant dispersion forces between them. So, the greater the percentage of crystalline regions, the stronger the polymer and the higher its softening temperature. Crystalline regions scatter light passing through them. So the greater the percentage of crystalline regions, the more light is scattered and the less transparent the polymer becomes.

20 B. 1 nm = 10^{-9} m, so IV is correct.

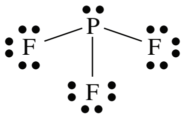
1 mm = 10^{-3} m, so 10^{-6} mm = 10^{-9} m so I is also correct.

Short-answer questions

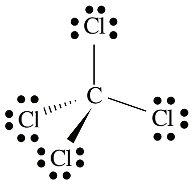
21 **a** V-shaped/bent and polar



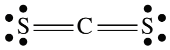
b Triangular pyramid and polar



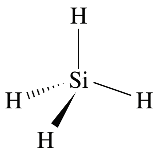
c Tetrahedral and non-polar There are no non-bonding pairs shown on Cl atoms



d Linear and non-polar



e Tetrahedral and non-polar



22 **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 hydrocarbons; e.g. C_5H_{12}, C_6H_{14}.

23 **a** But-2-yne

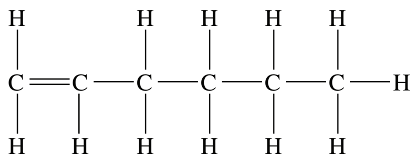
b Heptane

c 2,2,3-Trimethylpentane

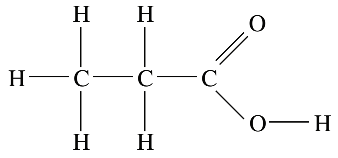
d Butanoic acid

e 2-Methylpropan-1-ol

24 **a**



b



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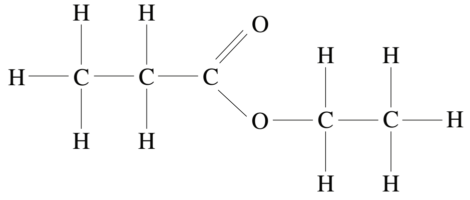

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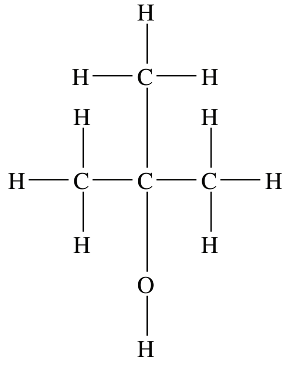
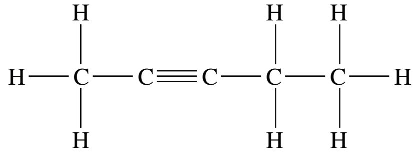


e



25 **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 $CHCCH_3$

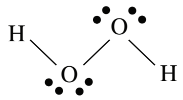
e $CH_2C(CH_3)CH_2C(CH_3)CH_2CH_3$

26 **a** The bonds between the molecules of ice are hydrogen bonds. The bonds between the 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 In ice, each water molecule forms hydrogen bonds with four other water molecules. This arrangement is very open, causing ice to have a lower density than liquid water.

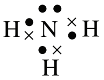
c Ethyne has the structure CHCH. The C atoms have a triple bond between them, each using three of their four valence electrons to form the triple bond. The fourth valence electron of each carbon atom forms a covalent bond with a hydrogen atom. There are no lone pairs and only two bonding regions on the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

Hydrogen peroxide has the following structure:

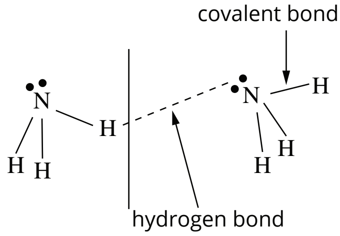


Each oxygen has 6 valence electrons. Two are involved in bonding leaving 2 pairs of non-bonding electrons. So there are 4 regions where electrons are present. These assume a tetrahedral arrangement to minimise repulsion. So the molecule is not linear.

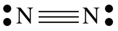
27 **a**



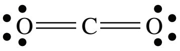
b The four electron pairs form a tetrahedral arrangement around the atom due to the repulsion of the electron pairs. The result is a pyramidal shaped molecule.



c **i**



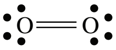
ii

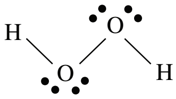


d **i** N_2 has no polar bonds so the intermolecular bonds are weak dispersion forces.
Although the bonds within CO_2 are polar, the molecule overall is symmetrical so the dipoles cancel and the overall molecule is non-polar. So the only intermolecular forces are dispersion forces.

ii Nitrogen is a highly electronegative element and 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.

28 **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.

29 **a** HCl, HF, H_2O, H_2S

b HCl, H_2S

c HCl, HF, F_2, H_2, O_2

d H_2O, HF

e O_2

f H_2 as it is the smallest non-polar molecule in the list.

30 **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 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.

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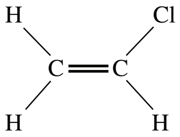
31 **a** 

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.

c **i** They are polymers designed for a specific purpose.

ii Use two different monomers

32 **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. melamine.

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. melamine.

d Crystalline regions occur when atoms in chains are arranged in a regular pattern thus strengthening the polymer. Non-crystalline regions occur when the atoms in the chain are randomly arranged.

33 % by mass of oxygen in the compound = 100 – 59.7 = 40.3

In 100 g of compound:

$m(\text{O}) = 40.3$ g and $m(\text{X}) = 59.7$ g

So, $n(\text{O}) = \frac{40.3}{16.0} = 2.52$

As ratio of X:O = 2:3, $n(\text{X}) = \frac{2}{3} \times 2.52 = 1.68$

So 59.7 g of X is equivalent to 1.68 mol.

And then $M(\text{X}) = \frac{m(\text{X})}{n(\text{X})} = \frac{59.7}{1.68} = 35.5 \text{ g mol}^{-1}$

The element must be Cl.

34 **a** In 100 g of compound, the mass of hydrogen = 100 – 54.54 – 36.36 = 9.10 g

$= n(\text{C}) : n(\text{H}) : n(\text{O})$

$= \frac{54.54}{12.0} : \frac{9.10}{1.0} : \frac{36.36}{16.0}$

$= 4.55 : 9.10 : 2.27$

$= 2 : 4 : 1$

So empirical formula is C_2H_4O.

b $n = \frac{m}{M}$, so $M = \frac{m}{n} = \frac{30.8}{0.350} = 88 \text{ g mol}^{-1}$

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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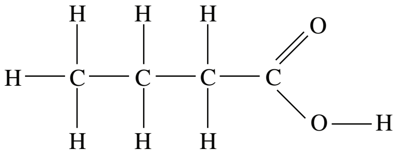
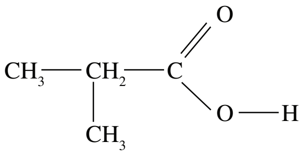
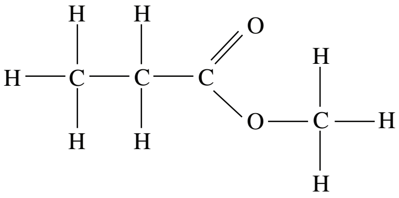
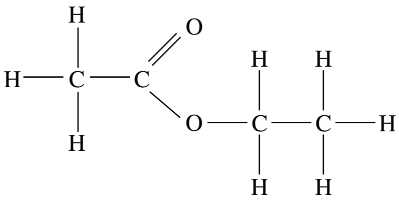
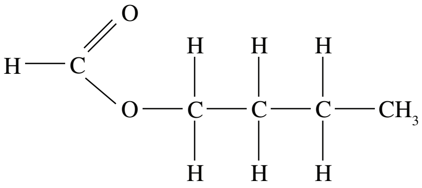
d

Compound 1	**Compound 2**
butanoic acid  or methylpropanoic acid 	methylpropanoate  or ethylethanoate  or propylmethanoate 

35 a m(H) in the compound = 4.738 – (0.476 + 4.22) = 0.042 g

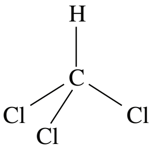
so, $n(\text{C}) : n(\text{H}) : n(\text{Cl}) = \frac{0.476}{12.0} : \frac{0.042}{1.0} : \frac{4.22}{35.35} = 0.040 : 0.042 : 0.119 = 1 : 1 : 3$

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 $= \frac{7.2 \times 10^{23}}{N_A} = \frac{7.2 \times 10^{23}}{6.02 \times 10^{23}} = 1.2$

$n(CHCl_3) = \frac{1}{3} \times 1.2 = 0.40$

So $m(CHCl_3) = n \times M = 0.40 \times 120 = 48$ g

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Unit 2

Area of Study 1 How do substances interact with water?

Multiple-choice questions

1 D. Water has a high specific heat capacity and so is able to store heat energy effectively. The temperature of large bodies of water does not change as much as that of surrounding land, thus moderating extremes of temperature.

2 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.

3 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.

4 A. Reaction A would form a precipitate of $BaSO_4$.

Reaction C would form a precipitate of $CuCO_3$.

In reaction D, Cu metal would precipitate out of solution as Cu^{2+} ions are reduced to Cu metal by Zn.

5 C. Lemon juice, hydrochloric acid and ammonium ions are all acidic so have a pH of less than 7. Sodium hydroxide is a base and has a pH greater than 7.

6 A. CH_3COOH is a weak acid and so is only partially ionised in solution, unlike HCl. 6.0 M indicates a more concentrated solution than 0.01 M.

7 B. When OH^- acts as an acid, it donates a proton and forms its conjugate base O^{2-}.

8 A. To act as either a Brønsted–Lowry acid or base, a species must be able to accept and to donate a $H^+(aq)$. NH_4^+ has a positive charge and so is unlikely to attract a $H^+(aq)$. $CO_3^{2-}(aq)$ has no $H^+(aq)$ to donate and CH_4 has no acidic hydrogens.

9 B. Polyprotic refers to an acid that can donate more than one proton.

CH_3COOH donates only one proton to form its conjugate base CH_3COO^-.

NH_4^+ also only donates one proton to form its conjugate base NH_3.

H_2SO_3 can donate one or two protons to form HSO_3^- or SO_3^{2-} respectively.

10 D. Both solutions have the same concentration and volume but nitric acid is a strong acid whereas ethanoic acid is a weak acid. So statement I is not correct.

As nitric acid is a stronger acid, it is ionised more fully in water releasing more $H^+(aq)$. So, the pH of the nitric acid will be lower than that of ethanoic acid and it will have greater electrical conductivity. So statement II is incorrect but statement III is correct.

Statement IV is also correct as both solutions contain the same number of moles of acid, which will react fully with a strong base such as NaOH.

11 D. For a change of 1 unit in pH, the concentration of $H^+(aq)$ changes by a factor of 10. So a change of pH of 2 units means a change of $H^+(aq)$ concentration of a factor of 100. The pH of beaker A is higher, so the $H^+(aq)$ concentration will be 100 times lower.

12 C. pH = 13 so $[H^+] = 10^{-13}$ M and $[OH^-] = 10^{-1}$ M = 0.10 M = [NaOH]

So $n(\text{NaOH}) = c \times V = 0.10 \times 0.500 = 0.050$ mol

And $m(\text{NaOH}) = n \times M = 0.050 \times 40.0 = 2.0$ g

13 A. An increase in the level of atmospheric carbon dioxide since the Industrial Revolution has resulted in an increased amount of dissolved carbon dioxide in the oceans. CO_2 is an acidic gas that forms carbonic acid with water.

$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$

and $H_2CO_3(aq) \rightarrow HCO_3^-(aq) + H^+(aq)$

so causing the pH of the oceans to drop.

14 C. A reductant donates electrons to another species causing them to be reduced. Mg^{2+} and Ag^+ are positively charged and so are unlikely to donate electrons. Mg is more reactive than Ag; it donates its electrons more readily and so is the stronger reductant.

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15 D. There is no exchange of electrons between Ag^+ and Cl^-. This is a precipitation reaction.

16 A. Cu is less reactive than Mg so will not displace magnesium from a solution of its ions.

Fe is also less reactive that Al so there will be no reaction here either. Substitute Pb^{2+}(aq) for Al^{3+}(aq)

17 C. There is no proton transfer so the reaction is not an acid–base one. There is a transfer of electrons from Mg to H^+ to form Mg^{2+} and H_2. The Mg loses electrons so is oxidised by the H^+, which is therefore the oxidising agent.

18 B. The reaction tells us that metal X is more reactive than Pb as it displaces Pb from a solution of its salt. So A and D are not correct. As Pb is sufficiently reactive to react with HCl to produce hydrogen gas, X, being more reactive than Pb, will also react similarly with HCl, so B is correct. We don't know how much more reactive X is than Pb or whether it is more or less reactive than Mg, so C *could* be true, but one cannot deduce from the data that it *must* be true.

19 B. A reducing agent causes reduction in another substance by donating electrons to it.

20 A. Each Ni atom loses two electrons to form a Ni^{2+}(aq) ion, so Ni is oxidised. The electrons are accepted by a Cu^{2+}(aq), so Cu^{2+}(aq) causes the oxidation and hence is the oxidising agent.

Short-answer questions

21 The stronger the intermolecular bonds, the higher the boiling point of a substance.

a Water molecules are highly polar, since hydrogen is bonded to oxygen. So between water molecules, there are hydrogen bonds as well as weak dispersion forces.

H_2S is much less polar, so between H_2S molecules there are dipole–dipole interactions and weak dispersion forces. As H_2S molecules are larger than water molecules, the dispersion forces between H_2S molecules are more significant than those between water molecules. However, hydrogen bonds are much stronger than dipole–dipole interactions, so, overall, the intermolecular bonds between H_2S molecules are weaker than those between water molecules and the boiling point of H_2S is consequently lower. This statement is correct but the polarity of the O-H bond is the reason hydrogen bonds form. Perhaps this should be worked into the first two sentences.

b The polarity of H_2Se is about the same as that of H_2S. However, H_2Se consists of much larger molecules so the dispersion forces between them are more significant. Therefore, overall, the intermolecular forces between H_2S molecules are weaker than those between H_2Se molecules and the boiling point of H_2S is consequently lower.

22 **a** Heat energy = $C_{water} \times m \times \Delta T = 4.18 \times 250.0 \times (100 - 18) = 8.6 \times 10^4$ J

b Heat energy = $C_{copper} \times m \times \Delta T = 0.39 \times 250.0 \times (100 - 18) = 8.0 \times 10^3$ J

c $8.0 \times 10^3 = 2.4 \times 250 \times \Delta T$; $\Delta T = \frac{8 \times 10^3}{2.4 \times 250} = 13°C$

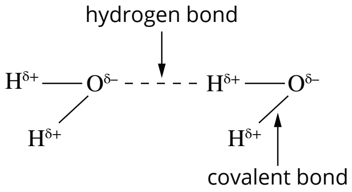
d The high heat capacity of water is due to the relatively strong hydrogen bonds between water molecules.

e e.g. as a coolant in factories

23 **a** The specific heat capacity of a substance is a measure of the energy required to raise the temperature of a certain amount of the substance (usually 1 g) by 1°C.

b The bonds between water molecules are relatively strong hydrogen bonds whereas those between hexane molecules are weaker dispersion forces.

Water has a higher specific heat capacity than ethanol because of the greater strength of its hydrogen bonds.



c Heat absorbed by water = $mC\Delta T = 100 \times 4.18 \times 12.0 = 5016$ J

The same amount of heat is provided to the hexane.

So, for hexane, $5016 = mC\Delta T = 100 \times 2.29 \times \Delta T$

So $\Delta T = 21.9°C$ and final temperature of hexane is $(20.0 + 21.9) = 41.9°C$.

24 **a** The water molecule is very polar. It is an asymmetrical molecule with a large difference in electronegativity between the H atoms and the O atom. It forms ion–dipole interactions with the positive and negative ions of ionic compounds, thus allowing them to dissolve in water. Polar covalent compounds in which H atoms are bonded to N, O or F atoms are able to form hydrogen bonds with water. Some polar covalent compounds are dissolved by ionising in water.

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b Necessary substances for metabolic processes include ionic compounds such as NaCl and molecular compounds such as glucose. These readily dissolve in water and are carried to the various parts of the body dissolved in the blood stream. Are the parts of the body dissolved in the blood stream? I did not try to correct this sentence because I was not sure to what depth you wanted to go. Waste products such as urea are also water soluble and can be excreted as a solution.

c Because of its excellent solvent properties, water in nature contains dissolved mineral salts and some soluble organic matter. Many salts are leached from rocks or result from the decay of plant and animal matter. In nature, gases dissolve in water which explains why rain water is slightly acidic due to the dissolution of CO_2.

25 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 lots of 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.

26 a $2HCl(aq) + Zn(s) \rightarrow ZnCl_2(aq) + H_2(g)$
b $2HNO_3(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_3)_2(aq) + 2H_2O(l)$
c $H_2SO_4(aq) + Na_2CO_3(aq) \rightarrow Na_2SO_4(aq) + CO_2(g) + H_2O(l)$
d $Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

27 a $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$
b $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
c $2H^+(aq) + CO_3^{2-}(aq) \rightarrow CO_2(g) + H_2O(l)$
d $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

28 a A strong acid is an acid that readily donates protons. Or, an acid that completely ionises in water.
b H_2CO_3 or H_2SO_4 or H_3PO_4 (or other correct polyprotic acids)
c i $HCO_3^-(aq) + H_2O(l) \rightarrow H_2CO_3(aq) + OH^-(aq)$
ii $HCO_3^-(aq) + H_2O(l) \rightarrow CO_3^{2-}(aq) + H_3O^+(aq)$
d H_3O^+

29 a i $[OH^-] = 2 \times [Ca(OH)_2] = 2 \times 0.050 = 0.10\ M$

$$[H_3O^+] = \frac{1.00 \times 10^{-14}}{[OH^-]} = \frac{1.00 \times 10^{-14}}{0.10} = 1.0 \times 10^{-13}\ M$$

$pH = -\log_{10}[H_3O^+] = -\log_{10}(1.0 \times 10^{-13}) = 13.0$ if you must but not 13.00

ii $n(HNO_3) = cV = 0.125 \times 0.400 = 0.0500$

Total volume of solution after dilution = 500 mL

so concentration of diluted $HNO_3 = \frac{n}{V} = \frac{0.0500}{0.500} = 0.100\ M = [H_3O^+]$

$pH = -\log_{10}[H_3O^+] = -\log_{10}(0.100) = 1.00$

iii $n(HCl) = \frac{m}{M} = \frac{2.45}{36.5} = 0.0671\ mol$

$c(HCl) = \frac{n}{V} = \frac{0.0671}{0.065} = 1.03\ M = [H_3O^+]$

$pH = -\log_{10}[H_3O^+] = -\log_{10}(1.03) = -0.014$

b $[H_3O^+] = 10^{-pH} = 10^{-0.050} = 0.891\ M$
$c(HNO_3) = 0.891\ M$
$n(HNO_3) = cV = 0.891 \times 0.300 = 0.267\ mol$
$m(HNO_3) = nM = 0.267 \times 63.0 = 16.8\ g$

30 a C. It is the only solution with a pH above 7.00 and sodium hydroxide is a base.

b B. As it is an acid, the pH must be less than 7. So it cannot be C.
A and B must both be acids (pH less than 7 or turns litmus red). B has the higher conductivity and, as they all have the same concentration, B must ionise the most and so must be the stronger acid.

c The pH of A must be greater than 1.5 but less than 7. It is an acid as it turns litmus paper red, but it will have a higher pH than solution B because it has a lower conductivity.

d i $[H_3O^+] = 10^{-pH} = 10^{-1.5} = 0.03\ M$
ii $[H_3O^+] = 10^{-pH} = 10^{-10.5} = 3 \times 10^{-11}\ M$

31 a Propanoic acid is a weak acid so it is partially ionised. The concentration of $H^+(aq)$ ions in a propanoic acid solution will therefore be lower than in solution of nitric acid or sulfuric acid, strong acids, of equal concentration. As the concentration of $H^+(aq)$ ions is lowest, the pH of the propanoic acid solution will be highest.

b Sulfuric acid is diprotic. In water, one hydrogen ion is completely donated to water. In addition, there is partial ionisation of the $HSO_4^-(aq)$ ion to release a second hydrogen ion. So the overall concentration of $H^+(aq)$ ions is higher than in a solution of the monoprotic nitric acid of the same concentration, and the pH of the solution will be lower.

32 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

d i $n(Pb(NO_3)_2) = \frac{m}{M} = \frac{3.15}{207.2 + 2 \times 62} = 0.00951$ mol

ii $n(PbCl_2)$ formed = $n(Pb(NO_3)_2)$ reacted = 0.00951 mol

So $m(PbCl_2)$ formed = $n \times M = 0.00951 \times (207.2 + 2 \times 35.5) = 2.65$ g

33 a $O_2(g) + 2H_2O(l) + 2Fe(s) \rightarrow 4OH^-(aq) + 2Fe^{2+}(aq)$

b The oxidising agent is the reactant that causes the other reactant to be oxidised. Fe loses electrons so it is oxidised. Oxygen is therefore the oxidising agent and Fe is the reducing agent.

c i Zn(s) is more reactive than iron and so will lose electrons in preference to Fe. So when there is some Zn attached to the Fe, the Zn is oxidised and the Fe is protected.

ii Sn is less reactive than Fe so if Fe were to be in contact with oxygen, it is the Fe that will preferentially corrode. If the Sn completely covers the Fe, the oxygen is prevented from reaching the Fe and the Fe is thus protected.

34 a i $2Br^-(aq) \rightarrow Br_2(aq) + 2e^-$

ii $H_2O_2(aq) + 2H^+(aq) + 2e^- \rightarrow 2H_2O(l)$

b Oxidation half-equation is $Ag(s) \rightarrow Ag^+(aq) + e^-$

The reduction half-equation is the same as in part a (ii).

In order to balance the charge, the oxidation half-equation needs to be multiplied by 2.

The overall equation is:

$H_2O_2(aq) + 2H^+(aq) + 2Ag(s) \rightarrow 2H_2O(l) + 2Ag^+(aq)$

35 a i V and VII

ii In V, there is a proton donation from HCOOH to $H_2O(l)$.

In VII, there is a proton donation from HCl to NH_3.

iii Any of $HCOOH/HCOO^-$, H_3O^+/H_2O, NH_4^+/NH_3, HCl/Cl^-

b i Any two of II, III and VI

ii In II, electrons are transferred from Ca to H^+.

In III, electrons are transferred from Mg to O_2.

In VI, electrons are transferred from Br^- to Cl_2.

iii Any one of Ca^{2+}/Ca, H^+/H_2, MgO/Mg, O_2/MgO, Cl_2/Cl^-, Br_2/Br^-

c i I and IV

ii In I, a precipitate of $PbCl_2(s)$ forms when two soluble compounds react

In IV, a precipitate of $Fe(OH)_3(s)$ when two soluble compounds react.

iii In IV, Na^+ ions and NO_3^- ions are spectator ions.

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Unit 2

Area of Study 2 How are substances in water measured and analysed?

Multiple-choice questions

1 C. 2.5% of the water on Earth is drinkable. Most of this water is not accessible as it is locked in the icecaps, glaciers or in the soil. Only 0.5 % of all the water on Earth is both drinkable and accessible.

2 A. Heavy metals are usually described as metals with high density which have a toxic effect on living things. Cadmium, lead, arsenic are included in lists of heavy metals. Potassium is not considered to be a heavy metal.

3 C. The solubility of gases decreases as the temperature increases but that of solids increases

4 B. At 0°C, 28 g dissolves in 100 g water.

So, at 0°C, 7.0 g dissolves in 25 g water.

As there is a total of 8.0 g solute, 8.0 – 7.0 = 1.0 g crystallises out.

5 C. $n(\text{Al}) = \frac{0.540}{27} = 0.0200$ mol

$n(\text{H}_2\text{SO}_4) = \frac{3}{2} \times 0.0200 = 0.0300$ mol

$V(\text{H}_2\text{SO}_4) = \frac{0.0300}{1.10} = 0.0273$ L = 27.3 mL

6 A. $n(\text{Na}_2\text{CO}_3)$ in the dilute solution = 0.150 × 0.750 = 0.113 mol

$V(3.0\text{ M Na}_2\text{CO}_3)$ needed $= \frac{0.113}{3.0} = 0.0375$ L = 37.5 mL

7 A. $n(\text{NaCl})$ to be diluted = 0.50 × 0.020 = 0.010 = $n(\text{NaCl})$ in the diluted solution

$V(\text{diluted NaCl}) = \frac{n}{c} = \frac{0.0100}{0.200} = 0.0500$ L = 50.0 mL

So 30.0 mL of water needs to be added to the 20 mL of the more concentrated solution.

8 A. If the acid had been a strong acid and therefore completely ionised, the initial pH would have been –log 0.1; i.e. 1.0. However, the initial pH is approximately 4, an indication of only partial ionisation and hence a weak acid.

The final pH of the mixture is close to 13 and the pH at equivalence point is higher than 7. Both these observations point to the base being strong.

9 C. The indicator must change colour at a pH range within the sharply changing pH of the reaction mixture.

10 B. The average titre is calculated from the three concordant results. That is, the three titres that only differ by 0.10 mL from smallest to largest are averaged.

So average titre is $= \frac{20.75 + 20.65 + 20.75}{3} = 20.72$ mL

The first result is not concordant and so is not included in the average titre.

11 D. Electrical conductivity in solution is due to the presence of ions. The higher the concentration of ions, the higher the conductivity.

H_2CO_3 is a weak acid and so the concentration of ions in 0.01 M H_2CO_3 is less than 0.01 M.

NH_4Cl dissociates into NH_4^+ ions and Cl^- ions, a total of 2 ions, so the solution will have an ion concentration of 0.02 M.

NaCl dissociates into Na^+ ions and Cl^- ions, so the solution will also have an ion concentration of 0.02 M.

Na_2SO_4 dissociates into 2Na^+ ions and one SO_4^{2-} ions, a total of 3 ions, so the solution will have an ion concentration of 0.03 M and thus have the highest electrical conductivity.

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12 B. Option A is correct, as the average speed of ions is lower at lower temperature and so the conductivity is lower.

B is incorrect because if some of the water evaporated from Y, the concentration of ions would be higher and so would the conductivity.

C and D are correct as the closer to the ocean the greater the salinity, and acidic wastes increase the concentration of ions in solution.

13 D. Dissolved carbon dioxide reacts with water to form the weak acid carbonic acid.

14 A. ppb is equivalent to mg in 1000 L

1.75 kg in 500 ML is equivalent to 1.75 mg in 500 L or 3.50 mg in 1000 L, i.e. 3.5 ppb.

15 C. $n(\text{AgCl}) = \frac{20.37}{143.3} = 0.1421 = n(\text{Cl}^-)$

$m(\text{Cl}^-)$ in 250 mL = 0.1421 × 35.5 = 5.05 g

$m(\text{Cl}^-)$ in 1.0 L = $\frac{5.05}{0.250}$ = 20.2 g

16 D. AAS is used for measuring metal ion concentrations.

17 B. AAS is used for measuring metal ion concentrations; UV–visible spectroscopy is applicable for coloured compounds and gravimetric analysis relies on the formation of precipitates.

18 C. Very little of the green light is absorbed, so the compound will appear green.

19 B. Analysis is best performed at the wavelength of greatest absorption, so at about 460 nm.

20 D. An absorbance of 3.0 is equivalent to 16 ppm from the graph. As the dam water had been diluted by a factor of 4, the concentration of the water in the dam is 4 × 16 = 64 ppm.

Short-answer questions

21 **a** A chemical contaminant is an element or compound that may be harmful if consumed. It can occur in nature or be synthetic (i.e. man-made).

b Examples: lead or cadmium from mining, fungicides or herbicides from agricultural runoff

22 **a** The solubility of gases increases as temperature decreases. More carbon dioxide is therefore dissolved in cold lemonade than in lemonade at room temperature.

b When the temperature of waterways is increased, gases come out of solution because their solubility has decreased. The decrease in oxygen concentration, in particular, may make the waterways unsuitable for some forms of aquatic life.

c Dissolved ionic compounds, principally NaCl, are the cause of salinity. In water, soluble ionic compounds dissociate into ions. The greater the salinity, the more ions there are in solution. As ions are charged and free to move in solution, they conduct an electric current. So, the electrical conductivity is related to the concentration of ions in solution and hence the salinity.

23 **a** **i** 110 g

ii Crystals of NaNO_3 would precipitate out of the solution.

iii Solubility at 10°C is 80 g/100 g water, so to produce a saturated solution containing 110 g, $\frac{110}{80} \times 100 = 138$ g water at 10°C is needed.

b At 20°C, the solubility of sucrose is 200 g in 100 g water, so in 200 g water, 400 g sucrose can dissolve.

c **i** A solution containing 30 g KNO_3 in 100 g water is unsaturated because up to 80 g can dissolve in that quantity of water at 50°C. The solution containing 87 g of KNO_3 is saturated; it contains the maximum amount of KNO_3 that can dissolve at that temperature. The one containing 90 g of KNO_3 is supersaturated; it contains more than the maximum amount of KNO_3 that can dissolve at that temperature and would be an unstable solution.

ii The three solutions can be identified by observing what happens when a small amount of solute is added to each solution. A small amount of solute added to an unsaturated solution will dissolve. A small amount of solute added to a saturated solution will not dissolve. A small amount of solute added to a supersaturated solution or even knocking the solution will cause the excess solute to crystallise out of the solution.

iii On cooling the solutions to 30°C, crystals will precipitate out of the saturated and supersaturated solutions but not from the unsaturated one, because even at 30°C the solution can hold more than 30 g KNO_3 in 100 g water.

d The solubility of $\text{CuSO}_2.5\text{H}_2\text{O}$ at 20°C is 18 g in 100 g water. So to make a saturated solution containing 35 g $\text{CuSO}_2.5\text{H}_2\text{O}$, $\frac{35}{18} \times 100 = 194$ g water is needed.

24 **a** 16 g sodium chloride dissolving in 50 g water is the same as 32 g dissolving in 100 g water. This less than the solubility of sodium chloride therefore all 16 g would dissolve.

b **i** The solubility of potassium sulfate is 12 g in 100 g water at 20°C. This equates to 6 g in 50 g water. If 9 g potassium sulfate is added to 50 g water, only 6 g will dissolve (and 3 g will remain undissolved).

ii At 80°C the solubility of potassium sulfate is 10.5 g in 50 g water. All the 9 g will dissolve 80°C.

c Since graphite is insoluble none will dissolve.

d Molar mass of K_2SO_4 = 174.3 g mol^{-1}
Assuming the density of water is 1.00 g mL^{-1}
Volume of 100 g water = 100 mL
Mass of K_2SO_4 that dissolves in 100 mL water = 12.0 g
So:

$$n(K_2SO_4) \text{ in 100 mL water} = \frac{12.0}{174.3} = 0.0688 \text{ mol}$$

$$c(K_2SO_4) = \frac{n}{V} = \frac{0.0688}{0.100} = 0.688 \text{ mol L}^{-1}$$

e **i** Add the mixture to 50 g water and heat to 80°C. Both salts are soluble at this temperature but graphite is insoluble. The graphite can be separated by filtration of the hot solution, leaving both K_2SO_4 and NaCl dissolved in the hot solution.

Allow the filtrate to cool to 20°C. Some of the K_2SO_4 will precipitate as the solution is saturated with respect to this salt at 20°C. The precipitated potassium sulfate can be recovered by filtration. The filtrate should be dried in an oven to obtain a pure sample of potassium sulfate.

ii At 20°C, as calculated in part a, 6.0 g will remain dissolved and 3.0 g will therefore precipitate out of solution. So the maximum amount of K_2SO_4 that can be obtained is 3.0 g.

f Sodium chloride is soluble at both 20°C. and 80°C and the solution is unsaturated at both temperatures, so NaCl will not precipitate out of solution. The filtrate solution contains both sodium chloride and potassium sulfate, so a pure sample could not be obtained even if it was evaporated to dryness.

25 **a** **i** A colourless, odourless gas is produced.
ii No change
iii A white precipitate is formed.
iv A green precipitate is produced.
v No change
vi No change
vii No change
viii No change
ix No change
x No change

b **i** $CO_3^{2-}(aq) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$
iii $Mg^{2+}(aq) + CO_3^{2-}(aq) \rightarrow MgCO_3(s)$
iv $Cu^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CuCO_3(s)$

c Sodium carbonate solution will react with three of the other solutions. It will form a gas with one solution, a white precipitate with another solution, and a green precipitate with the third. No reaction will be observed with the fourth solution.
The sulfuric acid solution will react with only one other solution producing gas bubbles.
Potassium nitrate solution will not react with any of the other four 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 **a** pH = 4 so $[H_3O^+] = 10^{-4}$ mol L^{-1}

$$[OH^-] = \frac{10^{-14}}{10^{-4}} = 10^{-10} \text{ mol L}^{-1}$$

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b $n(\text{HCl}) = 0.100 \times 0.0100 = 0.00100 \text{ mol} = n(\text{H}^+)$
$n(\text{NaOH}) = 0.100 \times 0.0120 = 0.00120 \text{ mol} = n(\text{OH}^-)$
The OH^- ions are in excess by 0.00120 – 0.00100 = 0.00020 mol
The volume of the mixture is 200 mL = 0.200 L

$$c(\text{OH}^-) = \frac{0.00020}{0.200} = 0.0010 = 1.0 \times 10^{-3} \text{ mol L}^{-1}$$

$$[\text{H}^+] = \frac{1.00 \times 10^{-14}}{[\text{OH}^-]} = \frac{1.00 \times 10^{-14}}{1.0 \times 10^{-3}} = 1.0 \times 10^{-11} \text{ mol L}^{-1}$$

$\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}(1.0 \times 10^{-11}) = 11.00$

27 a $n(\text{Pb(NO}_3)_2) = \frac{m}{M} = \frac{9.80}{331.2} = 0.0296 \text{ mol}$

$c(\text{Pb(NO}_3)_2) = \frac{0.296}{0.0500} = 0.592 \text{ M}$

b New volume of solution is 80.0 mL.
So $c(\text{Pb(NO}_3)_2) = \ = 0.370 \text{ M}$

c i $\text{Pb}^{2+}(aq) + 2\text{I}^-(aq) \rightarrow \text{PbI}_2(s)$

ii $n(\text{Pb}^{2+}) = 0.0296 \text{ mol}$
From the equation, this would require:
$2 \times 0.0296 = 0.0592$ mol I^- to react
$n(\text{NaI}) = cV = 0.650 \times 0.0500 = 0.0325 = n(\text{I}^-)$
So there is not sufficient I^- to react with all of the Pb^{2+}.
I^- is the limiting reagent and will limit the amount of product that forms.

$n(\text{I}^-)$ reacting = 0.0325 and $n(\text{PbI}_2)$ formed $= \frac{1}{2} \times 0.0325 = 0.0163$ mol

$m(\text{PbI}_2)$ formed $= m \times M = 0.0163 \times 461.0 = 7.49$ g

28 a $\text{Fe}^{3+}(aq) + 3\text{OH}^-(aq) \rightarrow \text{Fe(OH)}_3(s)$

b $2\text{Fe(OH)}_3(s) \rightarrow \text{Fe}_2\text{O}_3(s) + 3\text{H}_2\text{O}(g)$

c i $n(\text{Fe}_2\text{O}_3) = \frac{m}{M} = \frac{1.095}{159.6} = 0.006861 \text{ mol}$

So $n(\text{Fe}^{3+}) = 2 \times n(\text{Fe}_2\text{O}_3) = 0.01372$ mol
and m(iron ions) $= 0.01372 \times 55.8 = 0.766$ g

ii $c(\text{iron ions}) = \frac{n}{V} = \frac{0.01372}{1.30} = 0.0106 \text{ M}$

29 a A sample of a solution containing Ni^{2+} ions is scanned in a UV-visible spectrophotometer to determine its absorbance across a range of wavelengths. The wavelength that gives the highest absorbance is then chosen.

b Prepare a series of standard solutions containing nickel ions. Measure their absorbance at an appropriate wavelength and plot a calibration curve. Measure the absorbance of the unknown solution and determine its concentration from the calibration curve.

30 a $n(\text{NaHCO}_3) = cV = 0.0500 \times \frac{200}{1000} = 0.0100 \text{ mol}$

$m(\text{NaHCO}_3) = nM = 0.0100 \times (23.0 + 1.0 + 12 + 3 \times 16.0)$
$= 0.0100 \times 84.0 = 0.840$ g

b i $\text{NaHCO}_3(aq) + \text{HCl}(aq) \rightarrow \text{NaCl}(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(l)$

ii $n(\text{NaHCO}_3) = cV = 0.0500 \times 0.02000 = 0.00100 = n(\text{HCl})$

$c(\text{HCl}) = \frac{n}{V}$ (V = 35.05 mL = 0.03505 L)

$\frac{0.00100}{0.3505} = 0.0285 \text{ M}$

31 a $\text{NH}_3(aq) + \text{HCl}(aq) \rightarrow \text{NH}_4\text{Cl}(aq)$ or $\text{NH}_3(aq) + \text{H}^+(aq) \rightarrow \text{NH}_4^+(aq)$

b $n(\text{HCl}) = 0.0950 \times 0.01740 = 0.001653 \text{ mol} = n(\text{NH}_3)$ in the 20.00 mL titre
$n(\text{NH}_3)$ in 100 mL diluted cleaner $= 5 \times 0.001653 = 0.008265$ mol $= n(\text{NH}_3)$ in 10.00 mL original cleaner

$c(\text{NH}_3)$ in cleaner $= \frac{0.008265}{0.01000} = 0.8265 = 0.827 \text{ mol L}^{-1}$

c $m(\text{NH}_3)$ in 1 L cleaner $= 0.8265 \times 17.0 = 14.05$ g
$m(\text{NH}_3)$ in 750 mL bottle $= 14.05 \times 0.75 = 10.5$ g

d The observed colour change will be from yellow at the start of the titration when the indicator is added to the ammonia solution, to orange at equivalence point and pink when excess acid is present.

e **i** The result would be the same because the amount, in mole, of ammonia in the flask is unchanged by the addition of water, so the same titre of HCl would be needed for the reaction.

ii It would be lower because phenolphthalein indicator would change colour before equivalence point is reached. So the titre of HCl would be smaller than it should be. Hence, the calculated concentration of ammonia would be lower than the actual one.

iii Higher because the flask would contain some drops of diluted cleaner as well as the 20.00 mL of it measured out. More HCl would be needed to reach equivalence. Hence the calculated concentration of ammonia would be higher than the actual one.

32 **a** $n(\text{NaOH}) = c \times V = 0.790 \times 0.0234 = 0.01849$ mol

Succinic acid is diprotic, so:

n(succinic acid) in the 20.00 mL sample that reacted with the NaOH

$= \frac{1}{2} \times 0.01849 = 0.00924$ mol

b **i** Concentration in mol $L^{-1} = \frac{n}{V} = \frac{0.00924}{0.0200} = 0.462$ M

ii Molar mass of succinic acid is 118.1 g mol^{-1}

Concentration of succinic acid in g $L^{-1} = 0.462 \times 118.1 = 54.6$ g L^{-1}

iii %(m/v) is equivalent to grams of solute in 100 mL of solution

As there are 54.6 g in 1000 mL solution

%(m/v) of succinic acid = 5.46 %(m/v)

c The most likely reason would be that the solution was in fact not saturated. Another possibility is that the solution cooled during handling and so the solubility decreased.

33 **a** $n(S_2O_3^{2-}) = 0.0155 \times 0.01725 = 2.67 \times 10^{-4}$ mol

So from the equation, $n(O_2) = \frac{1}{4} \times 2.67 \times 10^{-4} = 6.68 \times 10^{-5}$ mol

i This was present in 300 mL water, so concentration

$= \frac{6.68 \times 10^{-5}}{0.300} = 2.23 \times 10^{-4}$ mol L^{-1}

ii ppm is equivalent to mg in 1 kg solution

1 L of water has a mass of 1 kg

So $m(O_2)$ in 1 L $= 2.23 \times 10^{-4} \times 32.0 = 0.00713$ g = 7.13 mg

So concentration is 7.13 ppm.

b % saturation $= \frac{7.13}{9.2} \times 100 = 78\%$

c It would be less than 9.2 as gases become less soluble as the temperature rises.

d **i** The concentration of oxygen may vary at different locations and depths. So samples from different depths and locations are analysed to obtain a reliable estimate of the overall oxygen level.

ii They were stoppered immediately to prevent dissolved oxygen from escaping into the atmosphere.

34 **a** From the calibration curve, an absorbance reading of 0.25 is equivalent to a concentration of mercury of 20 mg L^{-1}.

So, the 3.0 g of fish that was ground up into 100 mL water must contain

$0.100 \times 20 = 2.0$ mg

As a % by mass of fish, this represents $\frac{2.0 \times 10^{-3}}{3.0} \times 100 = 0.067\%$

b ppm is equivalent to mass in mg per kg

So 2.0 mg in 3.0 g fish is $\frac{2.0}{3.0} \times 1000 = 6.7 \times 10^2$ mg Hg in 1 kg fish, i.e. 6.7×10^2 ppm.

35 **a** Insecticides, pesticides, oil, dioxins

b **i** X travelled fastest as it had the smallest retention time.

ii Y is present in the highest concentration because it has the highest peak area.

c **i** A calibration curve is necessary to relate the concentration of a substance to the area of the peak.

ii A solution of X of precise concentration is made up. This solution is used to prepare a set of standard diluted solutions of various concentrations. Analyse these diluted solutions by HPLC under the same conditions as the analysis of the pond water.

Plot the area under the peak against the concentration of each standard.

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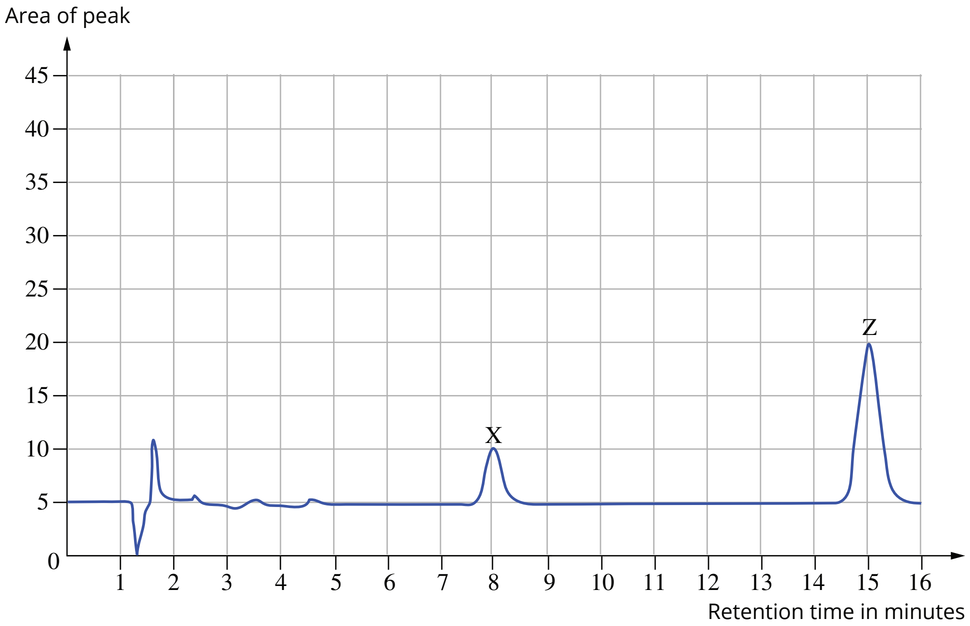
d



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Chapter 1 The atomic nature of matter

Section 1.1 Nanomaterials and nanoparticles

Worked example: Try yourself 1.1.1

CONVERTING KILOMETRES TO NANOMETRES

Convert 2.43 km into nanometres.	
Thinking	**Working**
Write the length in metres.	2.43 km = 2430 m
Write the length in scientific notation.	2430 m = 2.43×10^3 m
Multiply by 10^9 to convert to nanometres. (Hint: to do this, you just need to add 9 to the index on the number 10.)	$(2.43 \times 10^3) \times 10^9 = 2.4310^{(3 + 9)}$ $= 2.43 \times 10^{12}$ nm

1.1 Key questions

1 8.35 cm = 0.0835 m

0.0835 m = 8.35×10^{-2} m

$(8.35 \times 10^{-2}) \times 10^9 = 8.35 \times 10^{-2+9}$

$= 8.35 \times 10^7$ nm

2 **a** 1.35 cm = 1.35×10^{-2} m

1.35×10^{-2} m $\times 10^9 = 1.35 \times 10^7$ nm

b 4.2 mm = 4.2×10^{-3} m

$4.2 \times 10^{-3} \times 10^9 = 4.2 \times 10^6$ nm

3 0.050 mm = 5×10^{-5} m

$5 \times 10^{-5} \times 10^9 = 5.0 \times 10^4$ nm

$5.0 \times 10^4 \div 20 = 2500$ times larger

4 **a**

Shape	**Surface area**	**Volume**	$\frac{\textbf{Surface area}}{\textbf{Volume}}$
Cube 2 cm × 2 cm × 2 cm	24 cm^2	8 cm^3	3
Sphere of radius 1.38 cm	23.9 cm^2	11 cm^3	2.17
Tube (cylinder) Radius of 1 cm Height of 2 cm	19 cm^2	6.3 cm^3	3

b Sphere

c Sphere

d A large SA : *V* ratio is required for rapid cooling (e.g. on a hot day people spread out). A small SA : *V* ratio is required to minimise loss of heat when cold. Hence, people huddle or wrap their arms around themselves to keep warm.

5 There are two reasons why their size makes nanoparticles useful for transporting medicine. First, their small size makes it easy for the nanoparticles to get into the body by inhalation, injection or absorption through the skin. Second, their small size gives nanoparticles a large surface area compared to their volume. This means that a large number of molecules can be adsorbed onto the surface of a relatively small volume of nanoparticles.

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Section 1.2 The atomic world

1.2 Key questions

1. Noble gases
2. **a** Element
   b Element
   c Compound
   d Compound
   e Element
   f Compound
   g Element
   h Compound
3. An element is made up of just one type of atom and these atoms are identical. A compound is a molecule that contains at least two different types of atoms in a definite proportion.
4. **a** Iron
   b Potassium
   c Tungsten
   d Lead

Section 1.3 Inside atoms

1.3 Key questions

1. 10 000–100 000 larger depending on the element
2. Protons and neutrons found in the nucleus
3. The electrostatic attraction of the protons; the negative electrons are attracted to the positive protons and pulled towards them.

Section 1.4 Classifying atoms

Worked example: Try yourself 1.4.1

CALCULATING THE NUMBER OF SUBATOMIC PARTICLES

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

1.4 Key questions

1. Mass number
2. $Z = 15$. Protons = 15, electrons = 15. Neutrons = 31 – 15 = 16
3. 7 is its atomic number as it has 7 protons. The element that contains 7 protons is nitrogen.
4. Isotopes of the same element have the same atomic number and therefore the same number of protons and electrons. The number of neutrons is different between isotopes of the same element; therefore they have different mass numbers.
5. Carbon-14 has a mass number $A = 14$. The number of neutrons is:
   mass number – the atomic number = 14 – 6 = 8
   Carbon-12 has a mass number of 12 and therefore has 6 neutrons.
   Therefore the difference in the number of neutrons is 8 – 6, which is 2 neutrons.

Section 1.5 Electronic structure of atoms

1.5 Key questions

1. Each line in an emission spectrum corresponds to a specific amount of energy. This energy is emitted when electrons from higher-energy electron shells transition to a lower-energy shell. Different lines indicate that there are differences in energy between shells. This is evidence that electrons are found in shells with discrete energy levels.
2. Electrons revolve around the nucleus in fixed, circular orbits.
   Electrons' orbits correspond to specific energy levels in the atom.
   Electrons can only occupy fixed energy levels and cannot exist between two energy levels.
   Orbits of larger radii correspond to energy levels of higher energy.
3. Energy is emitted as light or electromagnetic radiation.

Section 1.6 Electronic configuration and the shell model

Worked example: Try yourself 1.6.1

ELECTRONIC CONFIGURATION FOR UP TO 36 ELECTRONS

<table>
<tr><td colspan="2">Apply the rules of the shell model to determine the electronic configuration of an atom with 34 electrons.</td></tr>
<tr><td>Thinking</td><td>Working</td></tr>
<tr><td>Recall the maximum number of electrons that each shell can hold.</td><td><table><tr><th>Shell (<i>n</i>)</th><th>Maximum number of electrons</th></tr><tr><td>1</td><td>2</td></tr><tr><td>2</td><td>8</td></tr><tr><td>3</td><td>18</td></tr><tr><td>4</td><td>32</td></tr></table></td></tr>
<tr><td>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.</td><td><table><tr><th>Shell (<i>n</i>)</th><th>Electrons in atom</th></tr><tr><td>1</td><td>2</td></tr><tr><td>2</td><td>8</td></tr><tr><td>3</td><td>8</td></tr><tr><td>4</td><td></td></tr></table></td></tr>
</table>

<table>
<tr><td>Place the next two electrons in the fourth shell.</td><td><table><tr><th>Shell (n)</th><th>Electrons in atom</th></tr><tr><td>1</td><td>2</td></tr><tr><td>2</td><td>8</td></tr><tr><td>3</td><td>8</td></tr><tr><td>4</td><td>2</td></tr></table></td></tr>
<tr><td>Continue filling the third shell until it holds up to 18 electrons. Put any remaining electrons in the fourth shell.</td><td><table><tr><th>Shell (n)</th><th>Electrons in atom</th></tr><tr><td>1</td><td>2</td></tr><tr><td>2</td><td>8</td></tr><tr><td>3</td><td>18</td></tr><tr><td>4</td><td>6</td></tr></table>The 10 remaining electrons from the previous step have gone into the third shell.</td></tr>
<tr><td>Write the electronic configuration by listing the number of electrons in each shell separated by commas.</td><td>2,8,18,6</td></tr>
</table>

Worked example: Try yourself 1.6.2

CALCULATING THE NUMBER OF VALENCE ELECTRONS IN AN ATOM

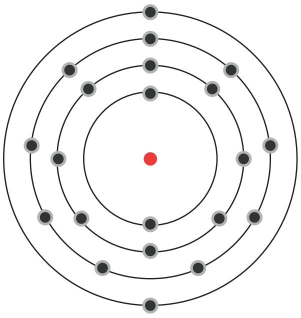
<table>
<tr><td colspan="2">How many valence electrons are present in an atom of sulfur, which has an atomic number of 16?</td></tr>
<tr><td>Thinking</td><td>Working</td></tr>
<tr><td>Recall the maximum number of electrons that each shell can hold.</td><td><table><tr><th>Shell (n)</th><th>Maximum number of electrons</th></tr><tr><td>1</td><td>2</td></tr><tr><td>2</td><td>8</td></tr><tr><td>3</td><td>18</td></tr><tr><td>4</td><td>32</td></tr></table></td></tr>
<tr><td>Place 16 electrons in the shells from the lowest energy to the highest energy. Do not exceed the maximum number of electrons allowed.</td><td><table><tr><th>Shell (n)</th><th>Electrons in atom</th></tr><tr><td>1</td><td>2</td></tr><tr><td>2</td><td>8</td></tr><tr><td>3</td><td>6</td></tr></table></td></tr>
<tr><td>Write the electronic configuration by listing the number of electrons in each shell separated by commas.</td><td>The electronic configuration is: 2,8,6.</td></tr>
<tr><td>Determine the number of electrons in the outer shell or valence shell.</td><td>The number of electrons in the valence shell is 6.</td></tr>
</table>

1.6 Key questions

1 Each shell (n) can hold up to $2n^2$ electrons. The fourth shell always accepts 2 electrons after the third electron has received 8 electrons. Once the fourth shell has 2 electrons, the third shell will continue to fill.
  - **a** 2,3
  - **b** 2,8,2
  - **c** 2,8,8,2
  - **d** 2,8,18,7

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

3



Sc 2,8,9,2

4 **a** Helium, He
  b Fluorine, F
  c Aluminium, Al
  d Nitrogen, N
  e Chlorine, Cl

5 The electronic configuration is 2,8,18,7. Therefore the number of valence electrons is 7.

Section 1.7 The Schrödinger model of the atom

Worked example: Try yourself 1.7.1

WRITING ELECTRONIC CONFIGURATIONS USING THE SCHRÖDINGER MODEL

Write the Schrödinger 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.	1s, 1 orbital 2s, 1 orbital 2p, 3 orbitals 3s, 1 orbital 3p, 3 orbitals 4s, 1 orbital 3d, 5 orbitals 4p, 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.	see table below
Write the electronic configuration by writing each subshell with the number of electrons as a superscript. Remember to group subshells from the same shell.	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^3 4s^2$

Subshell	**Electrons in subshell**	**Progressive total of electrons**
1s	2	2
2s	2	4
2p	6	10
3s	2	12
3p	6	18
4s	2	20
3d	3	23

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1.7 Key questions

1

Element (atomic number)	**Electronic configuration using the shell model**	**Electronic configuration using the subshell model**
Boron (5)	2,3	$1s^22s^22p^1$
Lithium (3)	2,1	$1s^22s^1$
Chlorine (17)	2,8,7	$1s^22s^22p^63s^23p^5$
Sodium (11)	2,8,1	$1s^22s^22p^63s^1$
Neon (10)	2,8	$1s^22s^22p^6$
Potassium (19)	2,8,8,1	$1s^22s^22p^63s^23p^64s^1$
Scandium (21)	2,8,9,2	$1s^22s^22p^63s^23p^63d^14s^2$
Iron (26)	2,8,16	$1s^22s^22p^63s^23p^63d^64s^2$
Bromine (35)	2,8,18,7	$1s^22s^22p^63s^23p^63d^{10}4s^24p^5$

2 The subshell model is a refinement of the shell model. The shell 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.

CHAPTER 1 REVIEW

1.1 NANOMATERIALS AND NANOPARTICLES

2 **a** 5 cm = 0.05 m = 5.0×10^{-2} m = $5.0 \times 10^{-2} \times 10^9 = 5.0 \times 10^7$ nm
b 12 mm = 0.012 m = 1.2×10^{-2} m = $1.2 \times 10^{-2} \times 10^9 = 1.2 \times 10^7$ nm
c 2 km = 2000 m = 2.0×10^3 m = $2.0 \times 10^3 \times 10^9 = 2.0 \times 10^{12}$ nm

3 Zinc oxide nanoparticles are colourless.

1.2 THE ATOMIC WORLD

4 Atoms are hard, indivisible structures.

5 Monoatomic = helium, neon, krypton.
Molecules = sulfur, oxygen, nitrogen.
Large network = copper, gold, diamond, tin

1.3 INSIDE ATOMS

6 The nucleus

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

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

1.4 CLASSIFYING ATOMS

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

10 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.

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

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

1.5 ELECTRONIC STRUCTURE OF ATOMS

13 $n = 1$

1.6 ELECTRONIC CONFIGURATION AND THE SHELL MODEL

14 $n = 3$. 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.

15 Magnesium. Total of 12 electrons means the element has 12 protons. The element with 12 protons and atomic number 12 is magnesium.

1.7 THE SCHRÖDINGER MODEL OF THE ATOM

16 **a** $1s^2$
- **b** $1s^22s^22p^2$
- **c** $1s^22s^22p^5$
- **d** $1s^22s^22p^63s^23p^1$
- **e** $1s^22s^22p^63s^23p^6$
- **f** $1s^22s^22p^63s^23p^63d^84s^2$
- **g** $1s^22s^22p^63s^23p^63d^{10}4s^24p^5$

17 A fluorine atom contains nine electrons. The electrons are arranged in energy levels called shells; two electrons are in the first shell and seven electrons are in the second shell, which has higher energy. The electron arrangement in the shells can be written as 2,7.

Shells are regarded as being made up of energy levels called subshells. The first shell contains an s-type subshell, which is labelled ‘1s’. The second shell contains both s- and p-type subshells, labelled ‘2s’ and ‘2p’ respectively.

Within subshells, electrons occupy regions of space known as orbitals. An orbital can hold up to two electrons. Subshells of an s-type contain one orbital, whereas p-type subshells contain three orbitals. The electron arrangement in the subshells of a fluorine atom can be represented as $1s^22s^22p^5$.

18 $1s^22s^22p^63s^23p^4$

19 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 4s-subshell is lower in energy than the 3d-subshell. Therefore, the 4s-subshell begins filling after the 3s- and 3p-subshells but before the 3d-subshell.

20 There is very little difference between the energy levels of the 3d- and 4s-subshells. As orbitals fill in sequence, one electron at a time, it is more stable for chromium to have all d-orbitals exactly half-filled than to have one empty d-orbital. Likewise for copper, it is more stable to have all d-orbitals completely filled than to have one half-filled and a full 4s-orbital.

CONNECTING THE MAIN IDEAS

21 Nothing

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

b Although Rutherford’s atomic model accounted for a number of atomic properties, it was not able to account for the characteristic emission spectrum of each element. The model was also in conflict with the principles of classical physics, which suggested that electrons moving in circular orbits should continuously lose energy and spiral into the nucleus.

c The Bohr model of the atom did not adequately explain why electrons adopted some energy levels but not others. In addition, calculated frequencies for lines in the emission spectra of atoms with more than one electron gave poor agreement with measured values.

Chapter 2 Electron arrangements and the periodic table

Section 2.1 The periodic table

2.1 Key questions

1 Groups 1, 2 and 13–18

2
- **a** 1
- **b** 15 – 10 = 5
- **c** 17 – 10 = 7
- **d** 2

3 The element is in period 3 and therefore has three occupied shells. As the element is in group 2, it will have two valence electrons. This gives an electron configuration of 2,8,2.

4
- **a**
  - **i** Group 13
  - **ii** Group 17
  - **iii** Group 1
  - **iv** Group 18
  - **v** Group 14
  - **vi** Group 14
- **b**
  - **i** 4
  - **ii** 2
  - **iii** 1
  - **iv** 1
  - **v** 7
  - **vi** 3
- **c**
  - **i** Silicon, Si 2,8,4 or $1s^22s^22p^63s^23p^2$
  - **ii** Beryllium, Be, 2,2 or $1s^22s^2$
  - **iii** Argon, Ar, 2,8,8 or $1s^22s^22p^63s^23p^6$

5 Atoms are listed in the periodic table in order of atomic number because the atomic number determines the number of electrons, and this in turn sets the properties that make each element unique. Relative atomic mass does not relate directly to either the atomic number or the electron number.

Section 2.2 Trends in the periodic table—Part 1

Worked example: Try yourself 2.2.1

CORE CHARGE

Determine the core 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 $1s^22s^22p^{5.}$
Determine the core charge. Core charge = number of protons – number of inner-shell electrons	There are 2 inner-shell electrons. Core charge = 9 – 2 = +7

2.2 Key questions

1 Carbon has 6 protons and 6 electrons. Its electronic configuration is $1s^22s^22p^2$. Carbon has two inner-shell electrons. Core charge = 6 – 2 = +4.

2 As core charge increases, electronegativity increases.

3 **a** **i** F
- **ii** Fr

b **i** Group 17
- **ii** Group 1

c Elements in group 18, the noble gases, have a very stable electronic configuration and so are unreactive.

Section 2.3 Trends in the periodic table—Part 2

2.3 Key questions

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

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

2 A metalloid is an element that exhibits both metallic and non-metallic properties.

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

CHAPTER 2 REVIEW

2.1 THE PERIODIC TABLE

1 **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

2 **a** Period 2, group 2
b Period 3, group 14
c Period 4, group 13
d Period 1, group 18

3 **a** Elements in the s-block are filling an s-subshell. As an s-subshell accommodates a maximum of two electrons, there are only two groups of elements in the s-block. They have outer-shell configurations of s^1 and s^2.

b Elements in the p-block are filling a p-subshell. As a p-subshell accommodates a maximum of six electrons, there are six groups of elements in the p-block. They have outer-shell configurations of s^2p^1 to s^2p^6.

c Elements in the d-block are filling a d-subshell. As a d-subshell accommodates a maximum of 10 electrons, there are 10 elements in each transition series.

d The lanthanides and actinides are filling an f-subshell. As an f-subshell accommodates a maximum of 14 electrons, there are 14 elements in each of these series.

4 **a** Silicon
b Potassium, caesium
c Bromine
d Nitrogen, arsenic

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2.2 TRENDS IN THE PERIODIC TABLE—PART 1

5 As one moves 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.

6 **a** Magnesium and phosphorus, with outer electrons in the third shell, are in the same period. Magnesium has a nuclear charge of +12 but, with completed inner shells of $1s^2 2s^2 2p^6$, the outer electrons experience the attraction of a core charge of +2. The outer-shell electrons of phosphorus, which has a nuclear charge of +15 and the same number of inner shells as magnesium, are attracted by a core charge of +5. The stronger attraction of the phosphorus electrons to the core means that more energy is required to remove an electron from a phosphorus atom than from a magnesium atom.

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

7 **a** $1s^2 2s^2 2p^3$

b Period 2 and group 15

c 5

d +5

2.3 TRENDS IN THE PERIODIC TABLE—PART 2

8 As you move across period 2 from lithium to fluorine:

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

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

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

9 **a** Lithium, sodium, potassium, rubidium

b The reactivity of metals increases down a group. This is because the number of electron shells increases down a group, so the valence electrons are further from the nucleus and more easily lost. Since metals lose electrons in their reactions, those that lose electrons most easily will be most reactive.

10 **a** Fluorine

b Aluminium and magnesium are metals. Fluorine has the highest core charge (it is located furthest right on the periodic table) while also containing the least number of electron shells. This means that fluorine attracts an electron more strongly than other elements and is therefore the most reactive non-metal.

11 **a** Nitrogen

b Chlorine

c Chlorine

12 The reactivity decreases then increases. Metal atoms lose electrons in chemical reactions. It is easier for atoms with a low core charge to do this. These atoms are found on the left-hand side of the periodic table. Moving across the period from left to right, core charge increases and reactivity decreases. For the non-metals, which are located on the right-hand side of the period, electrons are gained or shared by their atoms in chemical reactions. So as the core charge increases from left to right across the periodic table, the reactivity of non-metals increases.

CONNECTING THE MAIN IDEAS

13 Students' own answers.

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Chapter 3 **Metals**

Section 3.1 **Properties of metals**

Worked example: Try yourself 3.1.1

DETERMINING CHARGES

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

3.1 Key questions

1 **a** Li atoms have 3 electrons. The electronic structure is $1s^2 2s^1$. There is one electron in the outer shell. The charge of the cation will therefore be 1+.

b Mg atoms have 12 electrons. The electronic structure is $1s^2 2s^2 2p^6 3s^2$. There are two electrons in the outer shell. The charge of the cation will therefore be 2+.

c Ga atoms have 31 electrons. The electronic structure is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^1$. There are three electrons in the outer shell. The charge of the cation will therefore be 3+.

d Ba atoms have 56 electrons. The electronic structure is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 6s^2$. There are two electrons in the outer shell. The charge of the cation will therefore be 2+. Another way to approach this is to notice that barium is in the same group as magnesium. All alkali earth metals form 2+ cations by losing their two outer-shell electrons.

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

b Sodium

c Silver

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

3 **a** Silver, copper, gold, aluminium

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

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

5 Tensile strength, cost, availability

Section 3.2 Metallic bonding

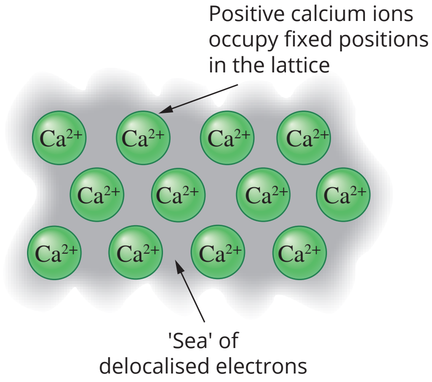
Worked example: Try yourself 3.2.1

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 a periodic table.)	Mg has an atomic number of 12. This means a neutral atom of magnesium has 12 electrons. The electronic configuration is $1s^22s^22p^63s^2$.
From the electronic configuration, find how many outer-shell electrons are lost to form cations that have a stable, noble gas electron configuration. These electrons become delocalised.	Mg has 2 electrons in its outer shell (the $3s^2$ electrons). 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.

3.2 Key questions

1 **a**



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

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

3 **a** Both graphite and metals are lustrous and conduct heat and electricity.

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

Section 3.3 Reactivity of metals

3.3 Key questions

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

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

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

3 Zn > Fe > Au

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

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Section 3.4 Extraction of iron from its ore

3.4 Key questions

1 **a** False. Slag is less dense and floats on top of the molten iron.
  b False. Iron ore contains impurities such as silica, alumina and manganese oxides.
  c True
  d True
  e False. The reaction between carbon in the coke and oxygen to produce carbon dioxide releases a considerable amount of heat energy to maintain the high temperatures required in the furnace.

2 $Fe_3O_4(l) + 4CO(g) \rightarrow 3Fe(l) + 4CO_2(g)$ magnetite + carbon monoxide → iron + carbon dioxide

3 Carbon dioxide is the main environmental pollutant from the production of iron in a blast furnace. The carbon dioxide is formed by the reaction of carbon from the coke with oxygen at high temperatures.
$C(s) + O_2(g) \rightarrow CO_2(g)$

4 The production of iron requires large amounts of coke, which is usually produced on site from coal. Proximity to a coal mine is required.

The production of iron may generate local environmental pollution. Iron furnaces should be located away from residential populations, agriculture and conservation areas.

The production of iron requires many workers. A site must have suitable accommodation and support for workers.

The iron produced must be able to be processed and shipped from the production site. Access is required to transport networks such as ports or railways.

Section 3.5 Modifying metals

3.5 Key questions

1 **a** A 20-cent coin contains copper and nickel. High-carbon steel contains iron and carbon.
  b The 20-cent coin is a substitutional alloy similar to Figure 3.5.3. High-carbon steel is an interstitial alloy, similar to Figure 3.5.2.

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

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

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

Section 3.6 Metallic nanomaterials

3.6 Key questions

1 Gold nanoparticles are not large enough to have a ‘sea’ of delocalised electrons or a consistent surface to reflect light.

2 A is a nanoparticle. B is a nanowire, as the length is more than 5 times the diameter. C is a nanorod, as the length is between 3 and 5 times the diameter. D is not a nanoparticle as all of its dimensions are above 100 nm.

3 iron + oxygen → iron oxide

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CHAPTER 3 REVIEW

3.1 PROPERTIES OF METALS

1 Mg, Ca, Sr

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

3 Electrical conductivity

4 **a** Low density
  b High electrical conductivity
  c High tensile strength

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

6 Al: $1s^22s^22p^63s^23p^1$; Al^{3+}: $1s^22s^22p^6$

7 It is able to be drawn into a wire.

3.2 METALLIC BONDING

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

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

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

11 In a metal lattice, metal cations are in a regular three-dimensional arrangement and have a positive charge. The positive cations are surrounded by a mobile sea of delocalised electrons. This is shown in Figure 3.2.1.

3.3 REACTIVITY OF METALS

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

13 Aluminium

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

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

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

17 **a** False
  b True
  c False
  d False
  e True

18 magnesium + oxygen → magnesium oxide

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3.4 EXTRACTION OF IRON FROM ITS ORE

19 Iron reacts with water and oxygen to produce iron oxides found in ores.

20 **a** Rocks and minerals from which iron can be economically extracted.
- **b** Naturally occurring solid substances with a definite chemical composition, structure and chemical and physical properties
- **c** The type of iron oxide that is most often found in iron ore. It has the formula Fe_2O_3.
- **d** A compound containing iron and oxygen

21 **a** Substance B and F are carbon dioxide.
- **b** Substance A is coke.
- **c** Solid D is iron ore, which is a mixture of iron oxide and other minerals.
- **d** Limestone is needed to remove unwanted contaminants, such as silica.

22 Iron ore is the source of iron. Limestone is the source of calcium oxide, which reacts with unwanted contaminants to form molten slag. Coke is used to produce heat and carbon monoxide, which reacts with the iron oxides to form iron metal. Air is pumped in as a source of oxygen for the combustion of the coke.

23 calcium oxide + silica → calcium silicate

3.5 MODIFYING METALS

24 **a** Copper and nickel; harder, more corrosion resistant and a silver colour
- **b** Tin and lead; lower melting temperature
- **c** Gold, silver and copper; harder
- **d** Iron, nickel and chromium; resists rusting, stronger
- **e** Mercury and zinc (sometimes a little silver is added); harder, non-toxic

25 Needle 2 < needle 3 < needle 1

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

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

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

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

3.6 METALLIC NANOMATERIALS

27 Nanowire

28 8.34×10^{-7} m is 834 nm. This particle would be too big to be classified as a nanoparticle.

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

CONNECTING THE MAIN IDEAS

30 **a** **i** Iron (steel) or aluminium
- **ii** Iron and steel are strong. Aluminium has a low density (light) and can be easily coloured.
- **iii** Iron rusts easily. Aluminium is soft and lacks strength.

b **i** Copper
- **ii** It is a good conductor of electricity and is ductile.

c **i** Gold, silver and platinum
- **ii** They are non-reactive, malleable and ductile, lustrous.

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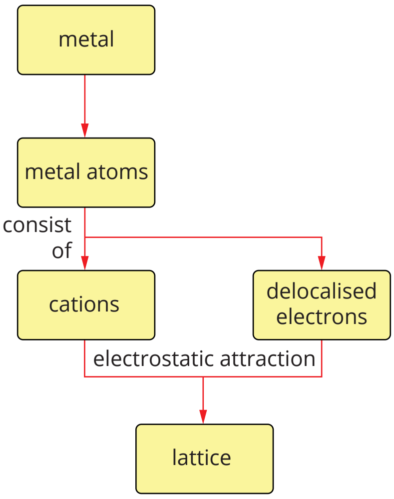
31 A variety of answers is possible. An example of a possible answer is shown.



32 **a** Aluminium Al, copper Cu, gold Au, iron Fe, silver Ag

b Aluminium: period 3, group 13, p-block
Copper: period 4, 1st transition series, d-block
Gold: period 6, 3rd transition series, d-block
Iron: period 4, 1st transition series, d-block
Silver: period 5, 2nd transition series, d-block

c Gold and silver

d Copper, gold, iron and silver

e Gold

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

b Stress corrosion cracking can occur between the crystal grains.

34 **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. So the electrostatic forces of attraction between delocalised electrons and cations are stronger in Ca and so it requires more energy to overcome the metallic bonding to boil the metal.

35 Aluminium is extracted from its ore by electrolysis. There was no source of electricity available for this process until 1886.

Chapter 4 Ionic bonding

Section 4.1 Properties and structures of ionic compounds

4.1 Key questions

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

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

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

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

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

Section 4.2 Using the ionic bonding model to explain properties

4.2 Key questions

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

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

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

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

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

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Section 4.3 Formation of ionic compounds

Worked example: Try yourself 4.3.1

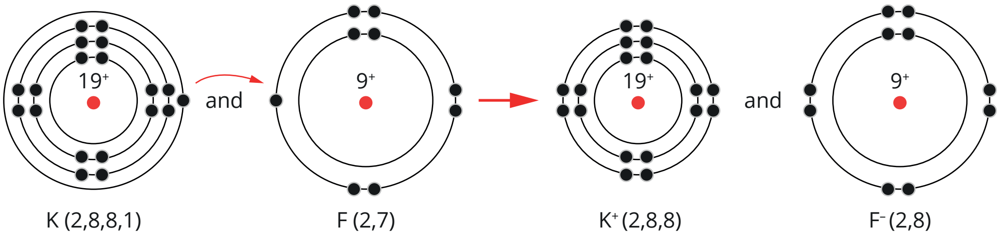
WRITING EQUATIONS FOR REACTIONS BETWEEN METAL AND NON-METAL ATOMS

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)
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) → 3Ca^{2+} (2,8,8) + 2P^{3-} (2,8,8)

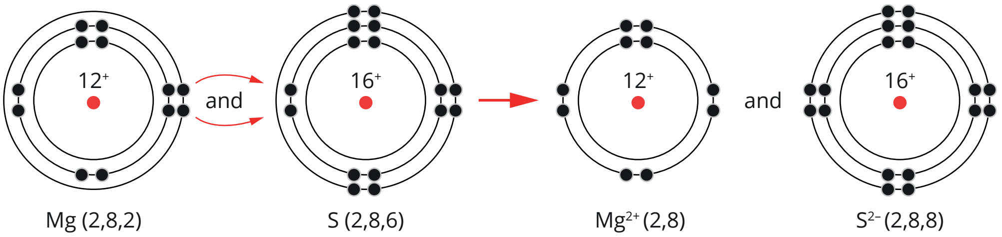
4.3 Key questions

1 Cations: calcium, aluminium. Anions: nitrogen, fluorine and phosphorus. 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.

2 **a**



b



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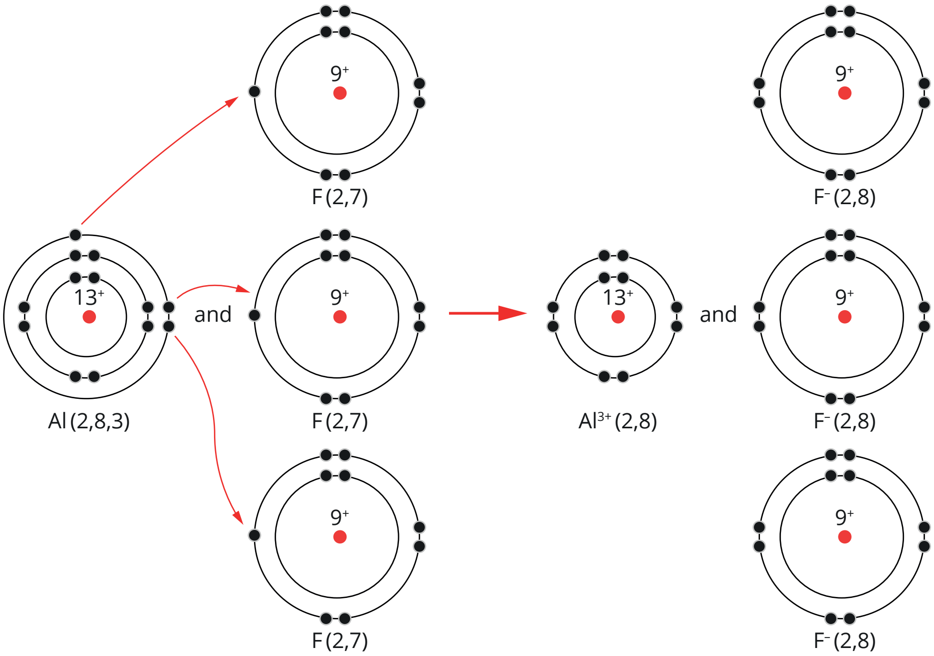
c



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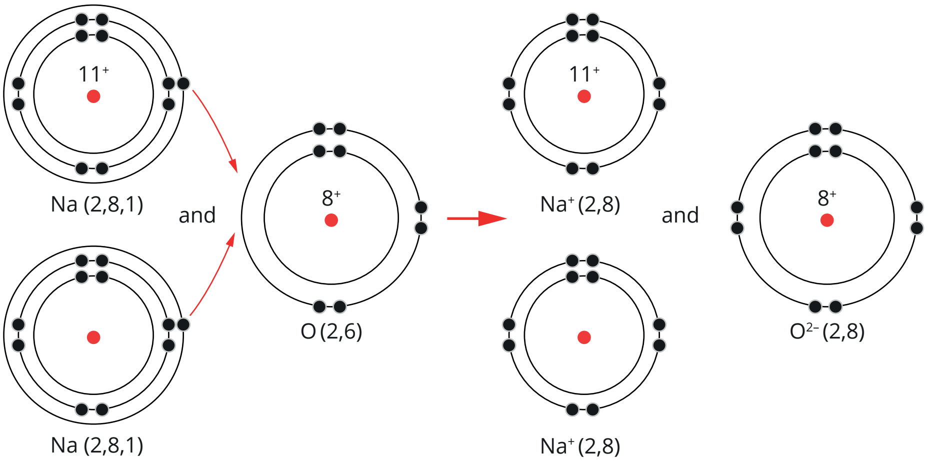

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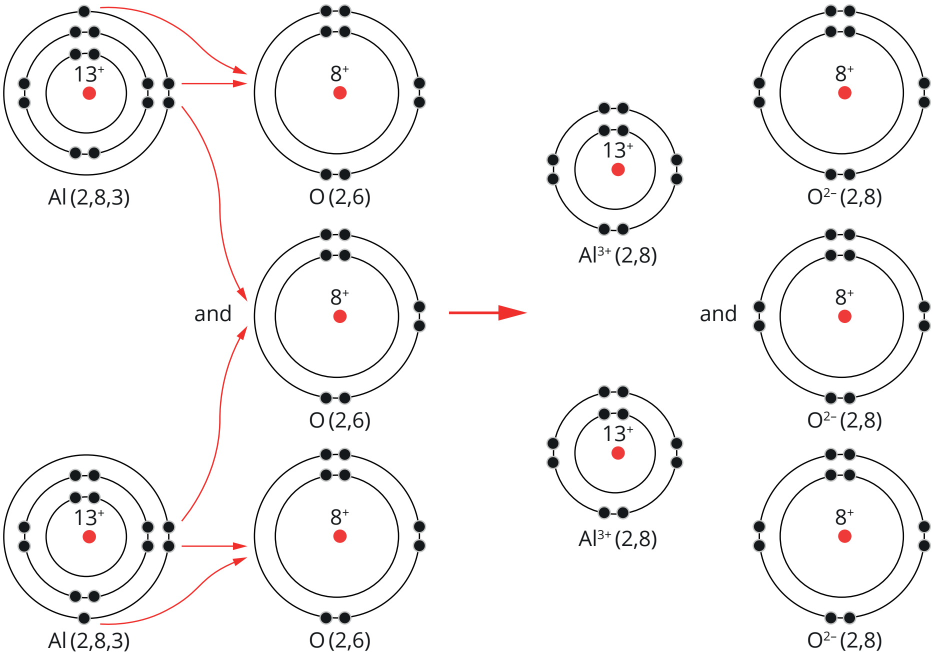

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3 Group 2 metals have two electrons in their valence shell. They lose these two electrons and therefore become positively charged, as they still have the original number of protons but have lost two electrons. Cations are ions with a positive charge.

4 The electronic configuration of an atom of potassium is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ and that of an atom of chlorine is $1s^2 2s^2 2p^6 3s^2 3p^5$. Because an atom of K has one more electron than an atom of a noble gas and an atom of Cl has one less, in a reaction one K atom can donate one electron to one Cl atom to give K^+ ($1s^2 2s^2 2p^6 3s^2 3p^6$) and Cl^- ($1s^2 2s^2 2p^6 3s^2 3p^6$), both of which have the electronic configuration of the noble gas argon. The formula of this compound is, therefore, KCl.

Calcium, however, has the electronic configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$. A Ca atom will lose two electrons to gain a noble gas configuration. Because each Cl atom will gain only one electron, there will be two chlorine atoms for each Ca atom. This reaction will therefore produce Ca^{2+} ($1s^2 2s^2 2p^6 3s^2 3p^6$) ions and Cl^- ($1s^2 2s^2 2p^6 3s^2 3p^6$) ions. The formula of the compound is, therefore, $CaCl_2$.

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

Section 4.4 Chemical formulas of simple ionic compounds

Worked example: Try yourself 4.4.1

STEPS IN WRITING A CHEMICAL FORMULA

Determine the chemical formula of the ionic compound formed between barium and fluoride ions. You may need to refer to Tables 4.4.1 and 4.4.2 on page 99.

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

4.4 Key questions

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

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

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3 **a** $NaCl$
- **b** KBr
- **c** $ZnCl_2$
- **d** K_2O
- **e** $BaBr_2$
- **f** Al_2I_3
- **g** $AgBr$
- **h** ZnO
- **i** BaO
- **j** Al_2S_3

4 **a** Potassium chloride
- **b** Calcium oxide
- **c** Magnesium sulfide
- **d** Potassium oxide
- **e** Sodium fluoride

Section 4.5 Writing formulas of more complex ionic compounds

4.5 Key questions

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

2 **a** $CuCl$
- **b** Fe_2O_3
- **c** CuO
- **d** $Cr_2(SO_4)_3$
- **e** FeO
- **f** $Pb(NO_3)_2$
- **g** PbO_2
- **h** $Sn(NO_3)_2$

3 **a** Magnesium hydroxide
- **b** Sodium carbonate
- **c** Iron(II) sulfate
- **d** Copper(II) sulfate
- **e** Barium nitrate
- **f** Copper(I) sulfate
- **g** Iron(III) sulfate
- **h** Ammonium nitrate
- **i** Disodium hydrogen phosphate

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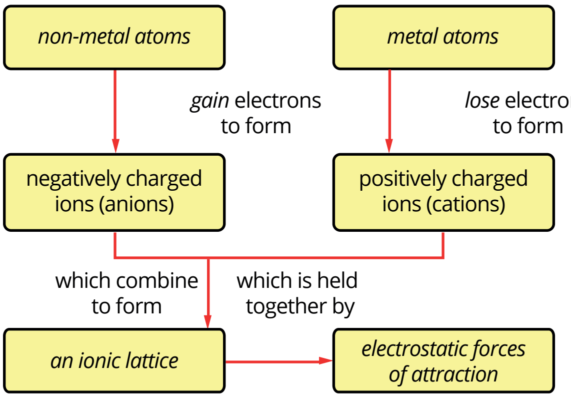
CHAPTER 4 REVIEW

4.1 PROPERTIES AND STRUCTURES OF IONIC COMPOUNDS

1 **a** Assemble equipment to test conductivity. Add a globe to the circuit. When the electrodes are touching the solid magnesium chloride, the globe will not light up.
  b Using the same equipment with molten sodium chloride, the globe will glow. Care is needed, as sodium chloride melts at 801°C.
  c If a crystal of sodium chloride was hit firmly with a hammer, it would shatter. Again, care is needed—safety glasses must be worn.

2 **a** Both metallic and ionic lattices
  b Both metallic and ionic lattices
  c Both metallic and ionic lattices
  d Ionic lattices only
  e Both metallic and ionic lattices

3



4.2 USING THE IONIC BONDING MODEL TO EXPLAIN PROPERTIES

4 **a** The electrostatic forces of attraction between the positive and negative ions are strong and will be overcome only at high temperatures.
  b The strong electrostatic forces of attraction between the ions mean that a strong force is needed to break up the lattice, giving the ionic crystals the property of hardness. However, the crystal lattice will shatter when a strong force is applied, suddenly causing ions of like charge to become adjacent to each other and be repelled.
  c In the solid state, the ions are not free to move. However, when the solid melts or dissolves in water, the ions are free to move and conduct electricity.

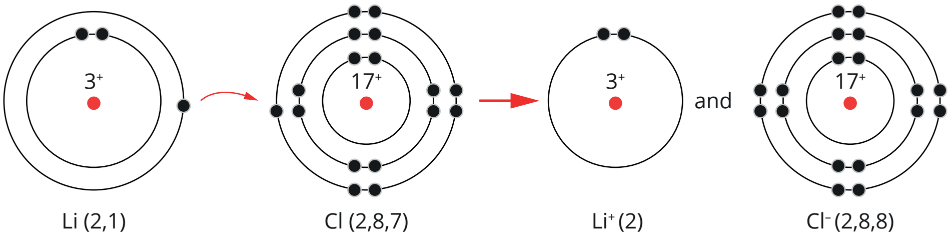
5 **a** Na^+, Cl^-; Mg^{2+}, O^{2-}
  b MgO. More energy is required to overcome the stronger forces. The higher melting temperature therefore reflects the solid with stronger forces between particles.
  c The strength of electrostatic attraction between ions will depend on the size of the ions and on their charge. The Mg^{2+} ion is slightly smaller than the Na^+ ion, and the O^{2-} ion is much smaller than the Cl^- ion. More importantly, the Mg^{2+} ion and the O^{2-} ions each have twice the charge of the Na^+ ion and the Cl^- ion. The attraction between the ions in MgO is therefore much stronger than in NaCl. Magnesium oxide therefore has a much higher melting temperature.

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

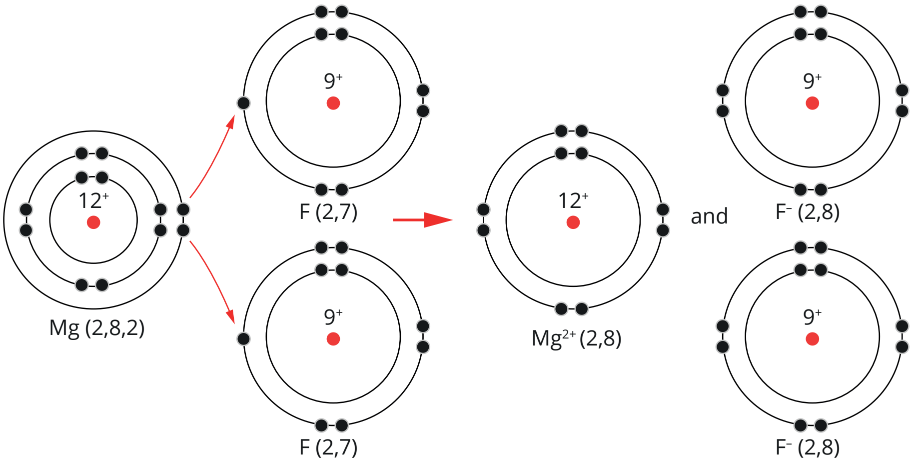
7 **a** Solid ionic compounds do not conduct electricity.
  b Ionic compounds are hard.
  c In solution ionic compounds conduct electricity.

4.3 FORMATION OF IONIC COMPOUNDS

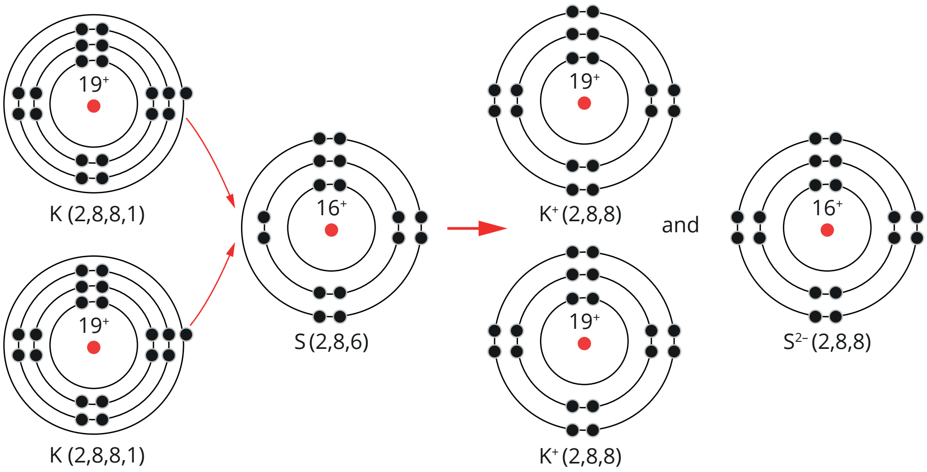
8 a



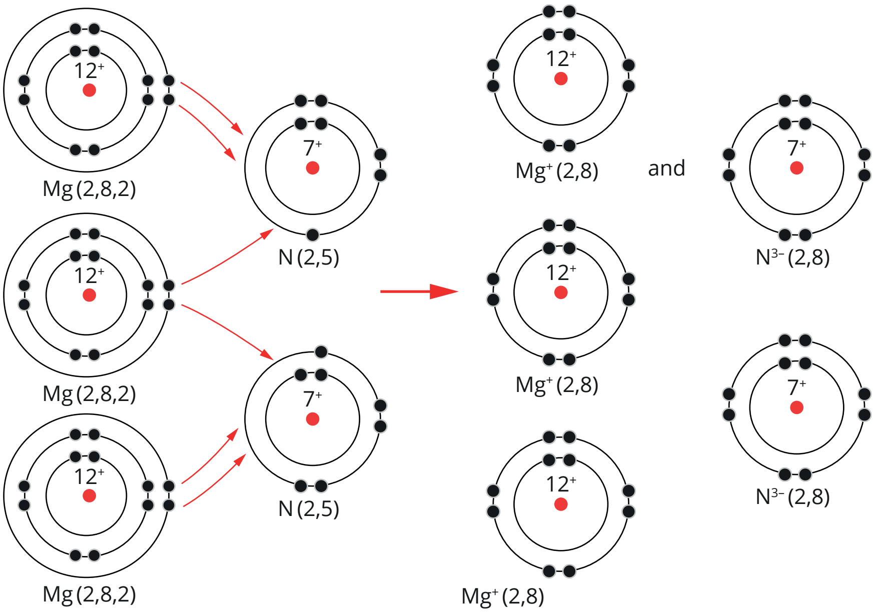
b



c



d



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

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

11 **a** $MgCl_2$ or MgF_2
  b NaCl or CaS
  c Na_2O or K_2S
  d Na_3N or Li_3N
  e $AlCl_3$ or AF_3
  f Mg_3N_2

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

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

4.4 CHEMICAL FORMULAS OF SIMPLE IONIC COMPOUNDS

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

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

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

4.5 WRITING FORMULAS OF MORE COMPLEX IONIC COMPOUNDS

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

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

19 **a** Ammonium carbonate
  b Copper(II) nitrate
  c Chromium(III) bromide

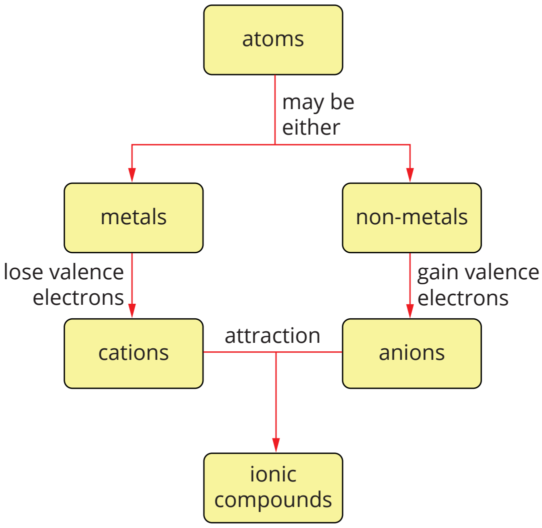
CONNECTING THE MAIN IDEAS

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

- **b** Agree. In a metallic lattice, each positive ion attracts the delocalised electrons in its region, and each delocalised electron is attracted to all neighbouring positive ions. In an ionic lattice, each positive ion is attracted to the negative ions that surround it, and vice versa.
- **c** Agree. In a metallic solid, there will be repulsion between the positive ions, and between the delocalised electrons. The particles are arranged to minimise these repulsions.

  In an ionic lattice, the arrangement of alternating positive and negative ions also minimises repulsion between like charges.
- **d** Agree. The energy required to remove the outer electron(s) is known as the ionisation energy, with each electron removed having a specific ionisation energy. Metals with low ionisation energy, such as sodium and potassium, form positive ions more readily than metals with high ionisation energies.
- **e** Disagree. Although this is true, it is not the reason ionic solids do not conduct. In a metal, delocalised electrons are free to move so it conducts electricity; in an ionic solid, the ions are not free to move so it does not conduct electricity.

21 A possible answer is shown.

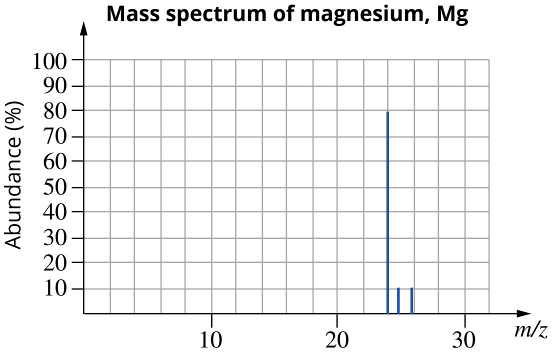


Chapter 5 Quantifying atoms and molecules

Section 5.1 Masses of particles

Worked example: Try yourself 5.1.1

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. Give your answer to two decimal places.	$A_r = \frac{199.36 + 881.71}{100} = 10.811$
Express the answer to two decimal places.	$A_r(B) = 10.81$

Worked example: Try yourself 5.1.2

CALCULATING PERCENTAGE ABUNDANCE OF EACH ISOTOPE FROM THE MASS SPECTRUM

Figure 5.1.8 shows a simplified mass spectrum of lead. From this mass spectrum, calculate the percentage abundances of each of its three isotopes.

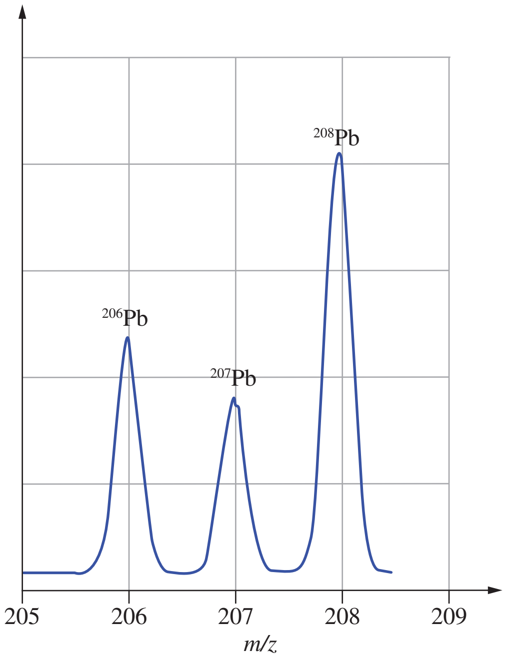


Figure 5.1.8 Mass spectrum of lead.

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

Worked example: Try yourself 5.1.3

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 $= 100 - x$ $= 100 - 70.50$ $= 29.50\%$

Worked example: Try yourself 5.1.4

CALCULATING THE RELATIVE MOLECULAR MASS OF MOLECULES

Calculate the relative molecular mass of nitric acid (HNO_3).	
Thinking	**Working**
Use the periodic table to find the relative atomic mass for the elements represented in the formula.	$A_r(\text{H}) = 1.0$ $A_r(\text{N}) = 14.0$ $A_r(\text{O}) = 16.0$
Determine the number of atoms of each element present, taking into consideration any brackets in the formula.	1 × H atom 1 × N atom 3 × O atoms
Determine the relative molecular mass by adding the appropriate relative atomic masses.	$M_r = 1 \times A_r(\text{H}) + 1 \times A_r(\text{N}) + 3 \times A_r(\text{O})$ $= 1 \times 1.0 + 1 \times 14.0 + 3 \times 16.0$ $= 63.0$

Worked example: Try yourself 5.1.5

CALCULATING THE RELATIVE FORMULA MASS OF IONIC COMPOUNDS.

Calculate the relative formula mass of copper(II) nitrate ($Cu(NO_3)_2$).	
Thinking	**Working**
Use the periodic table to find the relative atomic mass for the elements represented in the formula.	$A_r(\text{Cu}) = 63.5$ $A_r(\text{N}) = 14.0$ $A_r(\text{O}) = 16.0$
Determine the number of atoms of each element present, taking into consideration any brackets in the formula.	1 × Cu atom 1 × 2 = 2 N atoms 3 × 2 = 6 O atoms
Determine the relative formula mass by adding the appropriate relative atomic masses.	Relative formula mass $= 1 \times A_r(\text{Cu}) + 2 \times A_r(\text{N}) + 6 \times A_r(\text{O})$ $= 63.5 + 2 \times 14.0 + 6 \times 16.0$ $= 187.5$

5.1 Key questions

1 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.

B is the correct answer.

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

$= 15.999$

b $A_r(\mathrm{O}) = \frac{(106.9 \times 51.8) + (108.9 \times 48.2)}{100}$

$= 107.9$

c $A_r(\mathrm{O}) = \frac{(1.008 \times 99.986) + (2.014 \times 0.014) + (3.016 \times 0.0001)}{100}$

$= 1.008$

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

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

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

$694 = 6.02x + 702 - 7.02x$

$x = 8\%$

Percentage abundance of the lighter isotope is 8%.

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

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

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

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

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

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

5 **a** $M_r = 2 \times A_r(\mathrm{H}) + A_r(\mathrm{S}) + 4 \times A_r(\mathrm{O})$
$= 2 \times 1.0 + 32.1 + 4 \times 16.0$
$= 98.1$

b $M_r = A_r(\mathrm{N}) + 3 \times A_r(\mathrm{H})$
$= 14.0 + 3 \times 1.0$
$= 17.0$

c $M_r = 2 \times A_r(\mathrm{C}) + 6 \times A_r(\mathrm{H})$
$= 2 \times 12.0 + 6 \times 1.0$
$= 30.0$

6 **a** $M_r = A_r(\mathrm{K}) + A_r(\mathrm{Cl})$
$= 39.1 + 35.5$
$= 74.6$

b $M_r = 2 \times A_r(\mathrm{Na}) + A_r(\mathrm{C}) + 3 \times A_r(\mathrm{O})$
$= 2 \times 23.0 + 12.0 + 3 \times 16.0$
$= 106.0$

c $M_r = 2 \times A_r(\mathrm{Al}) + 3 \times A_r(\mathrm{S}) + 12 \times A_r(\mathrm{O})$
$= 2 \times 27.0 + 3 \times 32.1 + 12 \times 16.0$
$= 342.3$

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Section 5.2 Introducing the mole

Worked example: Try yourself 5.2.1

CALCULATING THE NUMBER OF MOLECULES

Calculate the number of molecules in 1.6 moles of carbon dioxide (CO_2).	
Thinking	**Working**
List the data given in the question next to the appropriate symbol. Include units.	The number of carbon dioxide molecules is the unknown, so: $N(CO_2) = ?$ $n(CO_2) = 1.6$ 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 5.2.2

CALCULATING THE NUMBER OF ATOMS

Calculate the number of hydrogen atoms in 0.35 mol of methane (CH_4).	
Thinking	**Working**
List the data given in the question next to the appropriate symbol. Include units.	The number of hydrogen atoms is the unknown so: $N(H) = ?$ $n(CH_4) = 0.35$ 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

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Worked example: Try yourself 5.2.3

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

Worked example: Try yourself 5.2.4

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

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

5.2 Key questions

1 Number of particles = amount (mol) × N_A

a $N(\text{Na}) = 2.0 \times 6.02 \times 10^{23}$
$= 1.2 \times 10^{24}$ atoms

b $N(N_2) = 0.10 \times 6.02 \times 10^{23}$
$= 6.02 \times 10^{22}$ molecules

c $N(\text{C}) = 20.0 \times 6.02 \times 10^{23}$
$= 1.20 \times 10^{25}$ atoms

d $N(H_2O) = 4.2 \times 6.02 \times 10^{23}$
$= 2.5 \times 10^{24}$ molecules

e $N(\text{Fe}) = 1.0 \times 10^{-2} \times 6.02 \times 10^{23}$
$= 6.02 \times 10^{21}$ atoms

f $N(CO_2) = 4.62 \times 10^{-5} \times 6.02 \times 10^{23}$
$= 2.78 \times 10^{19}$ molecules

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

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

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

$= 0.5$ mol

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

$= 0.25$ mol

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

$= 70$ mol

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

$= 70$ mol

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

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

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

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

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

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

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

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

4 **a** Each Cl_2 molecule has 2 Cl atoms.
$\therefore n(\text{Cl atoms}) = 0.4 \times 2$
$= 0.8$ mol

b Each methane molecule has 4 H atoms.
$\therefore n(\text{H atoms}) = 4 \times 1.2$
$= 4.8$ mol

c Each ethane molecule has 6 H atoms.
$\therefore n(\text{H atoms}) = 6 \times 0.12$
$= 0.72$ mol

d Each sulfate ion has 4 O atoms.
$\therefore n(\text{O atoms}) = 4 \times 1.5$
$= 6.0$ mol

Section 5.3 **Molar mass**

Worked example: Try yourself 5.3.1

CALCULATING THE MASS OF A SUBSTANCE

Calculate the mass of 4.68 mol of sodium carbonate (Na_2CO_3).	
Thinking	**Working**
List the data given to you in the question. Remember that whenever you are given a formula, you can calculate the molar mass.	$m(Na_2CO_3) = ?$ g $n(Na_2CO_3) = 4.68$ mol $M(Na_2CO_3) = 2 \times 23.0 + 12.0 + (3 \times 16.0) = 106.0 \text{ g mol}^{-1}$
Calculate the mass of sodium carbonate using: $n = \frac{m}{M}$	$n = \frac{m}{M}$, so $m = n \times M$ $m(Na_2CO_3) = 4.68 \times 106.0$ $= 496$ g

Worked example: Try yourself 5.3.2

CALCULATING THE NUMBER OF MOLECULES

Calculate the number of sucrose molecules in a teaspoon (4.2 g) of sucrose ($C_{12}H_{22}O_{11}$).	
Thinking	**Working**
List the data given to you in the question. Convert mass to grams if required. Remember that whenever you are given a formula you can calculate the molar mass.	$N(C_{12}H_{22}O_{11}) = ?$ $M(C_{12}H_{22}O_{11}) = (12 \times 12.0) + (22 \times 1.0) + (11 \times 16.0)$ $= 342.0\ \text{g mol}^{-1}$ $m(C_{12}H_{22}O_{11}) = 4.2\ \text{g}$
Calculate the amount, in mol, of $C_{12}H_{22}O_{11}$, using: $n = \frac{m}{M}$	$n(C_{12}H_{22}O_{11}) = \frac{m}{M}$ $= \frac{4.2}{342.0}$ $= 0.012\ \text{mol}$
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

5.3 Key questions

1 **a** $M_r = 2 \times A_r(N)$
$= 2 \times 14.0$
$= 28.0$

b $M_r = A_r(N) + 3 \times A_r(H)$
$= 14.0 + 3 \times 1.0$
$= 17.0$

c $M_r = 2 \times A_r(H) + A_r(S) + 4 \times A_r(O)$
$= 2 \times 1.0 + 32.1 + 4 \times 16.0$
$= 98.1$

d $M_r = A_r(Fe) + 3 \times A_r(N) + 9 \times A_r(O)$
$= 55.8 + 3 \times 14.0 + 9 \times 16.0$
$= 241.8$

e $M_r = 2 \times A_r(C) + 4 \times A_r(H) + 2 \times A_r(O)$
$= 2 \times 12.0 + 4 \times 1.0 + 2 \times 16.0$
$= 60.0$

f $M_r = A_r(S)$
$= 32.1$

g $M_r = 6 \times A_r(C) + 8 \times A_r(H) + 6 \times A_r(O)$
$= 6 \times 12.0 + 8 \times 1.0 + 6 \times 16.0$
$= 176.0$

h $M_r = A_r(Cu) + A_r(S) + 4 \times A_r(O) + 5 \times M_r(H_2O)$
$= 63.5 + 32.1 + 4 \times 16.0 + 5 \times 18.0$
$= 249.6$

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

a $m(\text{Na atoms}) = 1.0\ \text{mol} \times 23\ \text{g mol}^{-1}$
$= 23.0\ \text{g}$

b $m(O_2) = 2.0\ \text{mol} \times 32.0\ \text{g mol}^{-1}$
$= 64.0\ \text{g}$

c $m(CH_4) = 0.10\ \text{mol} \times 16.0\ \text{g mol}^{-1}$
$= 1.60\ \text{g}$

d $m(Al_2O_3) = 0.25\ \text{mol} \times 102\ \text{g mol}^{-1}$
$= 25.5\ \text{g}$

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

a $n(\text{H atoms}) = \frac{5}{1} = 5$ mol

b $n(H_2) = \frac{5}{2} = 2.5$ mol

c $n(\text{Al atoms}) = \frac{2.7}{27} = 0.10$ mol

d $n(CH_4) = \frac{0.4}{16} = 0.025$ mol

e $n(O_2) = \frac{0.10}{32} = 0.0031$ mol

f $n(\text{O atoms}) = \frac{0.10}{16} = 0.0063$ mol

g $n(P_4) = \frac{1.2 \times 10^{-3}}{124} = 9.7 \times 10^{-6}$ mol

h $n(\text{P atoms}) = \frac{1.2 \times 10^{-3}}{31} = 3.9 \times 10^{-5}$ mol

4 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 $n(\text{Na}) = \frac{23}{23} = 1.0$ mol

Number of Na atoms = $1.0 \times 6.0 \times 10^{23} = 6.0 \times 10^{23}$ atoms

b $n(\text{Ar}) = \frac{4.0}{39.95} = 0.10$ mol

Number of Ar atoms = $0.10 \times 6.0 \times 10^{23} = 6.0 \times 10^{22}$ atoms

c $n(\text{Mg}) = \frac{0.243}{24.3} = 0.01$ mol

Number of Mg atoms = $0.01 \times 6.0 \times 10^{23} = 6.0 \times 10^{21}$ atoms

d $n(\text{Au}) = \frac{10.0}{196.97} = 0.051$ mol

Number of Mg atoms = $0.051 \times 6.0 \times 10^{23} = 3.1 \times 10^{22}$ atoms

5 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 **i** $n(O_2) = \frac{16}{32} = 0.50$ mol

Number of O_2 molecules = $0.5 \times 6.0 \times 10^{23} = 3.0 \times 10^{23}$ molecules

ii $n(N_2) = \frac{2.8}{28} = 0.10$ mol

Number of N_2 molecules = $0.1 \times 6.0 \times 10^{23} = 6.0 \times 10^{22}$ molecules

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

Number of SO_2 molecules = $0.05 \times 6.0 \times 10^{23} = 3.0 \times 10^{22}$ molecules
Each molecule contains 2 oxygen atoms.
So, number of oxygen atoms = 6.0×10^{22} atoms.

c $n(NH_3) = \frac{288}{17} = 16.9$ mol

Number of NH_3 molecules = $16.9 \times 6.02 \times 10^{23} = 1.017 \times 10^{25}$ molecules
Each molecule contains 4 atoms (1 of N and 3 of H).
So, total number of atoms = $4.0 \times 1.017 \times 10^{25} = 4.07 \times 10^{25}$ atoms.

Section 5.4 Percentage composition and empirical formulas

Worked example: Try yourself 5.4.1

CALCULATING PERCENTAGE COMPOSITION

Calculate the percentage by mass of nitrogen in ammonium nitrate (NH_4NO_3).	
Thinking	**Working**
Find the molar mass of the compound.	$M(NH_4NO_3) = (2 \times 14.0) + (4 \times 1.0) + (3 \times 16.0)$ $= 80.0\ g\ mol^{-1}$
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 $= \frac{\text{mass of N in 1 mol of } NH_4NO_3}{\text{molar mass of } NH_4NO_3} \times 100$ $= \frac{28.0}{80.0} \times 100$ $= 35.0\%$

Worked example: Try yourself 5.4.2

DETERMINING THE EMPIRICAL FORMULA

0.50 g of magnesium is heated and allowed to completely react with chlorine. 1.96 g of white powder is formed. Determine the empirical formula of the compound.		
Thinking	**Working**	
Write down the mass, in g, of all elements present in the compound. If masses are given as percentages, assume that the sample weighs 100 g, then the percentages become masses in grams.	$m(Mg) = 0.50\ g$	$m(Cl) = 1.96 - 0.50 = 1.46\ g$
Calculate the amount, in mol, of each element in the compound using: $n = \frac{m}{M}$	$n(Mg) = \frac{0.50}{24.3}$ $= 0.0206\ mol$	$n(Cl) = \frac{1.46}{35.5}$ $= 0.0411\ mol$
Simplify the ratio by dividing each number of moles by the smallest number of moles calculated in the previous step. This gives you a ratio of the number of atoms of each element.	$\frac{0.0206}{0.0206}$ $= 1$	$\frac{0.0411}{0.0206}$ $= 2$
Find the simplest whole number ratio.	1:2	
Write the empirical formula.	$MgCl_2$	

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Worked example: Try yourself 5.4.3

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 × 12.0 + 5 × 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 × C_2H_5 = C_4H_{10}

5.4 Key questions

1 **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. For example, $M(Fe_2O_3)$ = 159.6 g mol^{-1}.

$$\therefore \%(Fe) = \frac{2 \times 55.8 \times 100}{159.6}$$

$$= 69.9\%$$

b $$\%(U) = \frac{3 \times 238.03 \times 842.09}{159.6}$$

$$= 84.8\%$$

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

$$= 26.2\%$$

d $$\%(O) = \frac{6 \times 16 \times 100}{187.5}$$

$$= 51.2\%$$

2 **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	$\frac{70}{14}$ = 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	$\frac{3.57}{3.57}$ = 1
Round off to whole numbers	1	1

∴ empirical formula is CO

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c When 6.03 g is the mass of magnesium in 10.0 g of the compound, the mass of oxygen is (10.0 – 6.03) = 3.97 g.

	Mg	**O**
Mass	6.03 g	3.97 g
Molar mass	24.3 g mol^{-1}	16 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{6.03}{24.3} = 0.248$ mol	$n = \frac{3.97}{16} = 0.248$ mol
Divide all by the smallest amount	$\frac{0.248}{0.248} = 1$	$\frac{0.248}{0.248} = 1$
Round off to whole numbers	1	1

∴ empirical formula is MgO

d When 2.4 g is the mass of carbon in 3.2 g of the hydrocarbon, the mass of hydrogen is (3.2 – 2.4) = 0.8 g.

	C	**H**
Mass	2.4 g	0.8 g
Molar mass	12 g mol^{-1}	1 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{2.4}{12} = 0.2$ mol	$n = \frac{0.8}{1} = 0.8$ mol
Divide all by the smallest amount	$\frac{0.2}{0.2} = 1$	$\frac{0.8}{0.2} = 4$
Round off to whole numbers	1	1

∴ empirical formula is CH_4

3 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.

Molar mass of a CH unit (empirical formula) = 12 + 1 = 13 g mol^{-1}
Molar mass of the compound (molecular formula) = 78 g mol^{-1}

∴ number of CH units in one molecule = $\frac{78}{13} = 6$

∴ molecular formula is C_6H_6 (which is benzene)

b Molar mass of an HO unit (empirical formula) = 1 + 16 = 17 g mol^{-1}
Molar mass of the compound (molecular formula) = 34 g mol^{-1}

∴ number of OH units in one molecule = $\frac{34}{17} = 2$

∴ molecular formula is H_2O_2

c Molar mass of a CH_2O unit (empirical formula) = 12 + (2 × 1) + 16 = 30 g mol^{-1}
Molar mass of the compound (molecular formula) = 90 g mol^{-1}

∴ number of CH_2O units in one molecule = $\frac{90}{30} = 3$

∴ molecular formula is $C_3H_6O_3$

d Molar mass of an NO_2 unit (empirical formula) = 14 + (16 × 2) = 46 g mol^{-1}
Molar mass of the compound (molecular formula) = 46 g mol^{-1}

∴ number of NO_2 units in one molecule = $\frac{46}{46} = 1$

∴ molecular formula is NO_2

e Molar mass of a CH_2 unit (empirical formula) = 12 + (1 × 2) = 14 g mol^{-1}
Molar mass of the compound (molecular formula) = 154 g mol^{-1}

∴ number of CH_2 units in one molecule = $\frac{154}{14} = 11$

∴ molecular formula $C_{11}H_{22}$

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4 **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}.

A hydrocarbon contains only carbon and hydrogen.

Hence, %H = 100 – 85.7 = 14.3%

	H	**Cl**
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

b Molar mass of a CH_2 unit (empirical formula) = 12 + 2 = 14 g mol^{-1}
Molar mass of the compound (molecular formula) = 70 g mol^{-1}

∴ number of CH_2 units in one molecule = $\frac{70}{14} = 5$

∴ molecular formula is C_5H_{10}

5 **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 g mol^{-1}	1 g mol^{-1}	16 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{1.8}{12} = 0.15$ mol	$n = \frac{0.3}{1} = 0.30$ mol	$n = \frac{2.4}{16} = 0.15$ 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

b Molar mass of a CH_2O unit (empirical formula) = 12 + (2 × 1) + 16 = 30 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 5 REVIEW

5.1 MASSES OF PARTICLES

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

2 $$A_r(\text{Pd}) = \frac{(101.9049 \times 0.96) + (103.9036 \times 10.97) + (104.9046 \times 22.23) + (105.9032 \times 27.33) + (107.9039 \times 26.71) + (109.9044 \times 11.800)}{100}$$

= 106.4

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

$= 39.96$

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

$= 39.11$

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.

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

$= 52$

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

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

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

$\therefore x = 48.0$

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

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

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

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

$\therefore x = 61.5$

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

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

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

$\therefore M_r(\text{H}_2\text{O}) = 2 + 16$

$= 18$

b $M_r(\text{P}_4) = 4 \times 31$

$= 124$

c $M_r(\text{CO}) = 12 + 16$

$= 28$

8 **a** Relative formula mass zinc bromide = 65.4 + 2 × 79.9 = 225.2

b Relative formula mass barium hydroxide = 137.3 + 2 × 16.0 + 2 × 1.0 = 171.3

c Relative formula mass iron(III) carbonate = 2 × 55.8 + 3 × 12.0 + 9 × 16.0 = 297.6

5.2 INTRODUCING THE MOLE

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

a $n(\text{H}_2\text{O}) = \frac{4.50 \times 10^{23}}{6.02 \times 10^{23}} = 0.75$ mol

b $n(\text{CH}_4) = 15.0$ mol

c $n(\text{Cl}_2) = 3.8 \times 10^4$ mol

d $n(\text{C}_{12}\text{H}_{22}\text{O}_{11}) = \frac{1}{6.02 \times 10^{23}} = 1.7 \times 10^{-24}$ mol

10 **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

5.3 MOLAR MASS

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

12 **a** 55.8 g mol^{-1}

b $M_r(H_2SO_4) = 2 + 32 + 64 = 98$

$\therefore M_r(H_2SO_4) = 98$ g mol^{-1}

c $M_r(Na_2O) = (2 \times 23) + 16 = 62$

$\therefore M_r(Na_2O) = 62$ g mol^{-1}

d $M_r(Zn(NO_3)_2) = 65.4 + (2 \times 14) + (6 \times 16) = 189.4$

$\therefore M_r(Zn(NO_3)_2) = 189.4$ g mol^{-1}

e $M_r(H_2NCH_2COOH) = (5 \times 1) + 14 + (2 \times 12) + (2 \times 16) = 75$

$\therefore M_r(H_2NCH_2COOH) = 75.0$ g mol^{-1}

f $M_r(Al_2(SO_4)_3) = (2 \times 27) + (3 \times 32) + (12 \times 16) = 342$

$\therefore M_r(Al_2(SO_4)_3) = 342$ g mol^{-1}

g $M_r(FeCl_3.6H_2O) = 55.8 + (3 \times 35.5) + (12 \times 1) + (6 \times 16) = 270$

$\therefore M_r(FeCl_3.6H_2O) = 270$ g mol^{-1}

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

a $m(C_2H_6) = 0.060 \times (24 + 6) = 1.8$ g

b $m(C_6H_{12}O_6) = 0.32 \times ((6 \times 12) + (12 \times 1) + (6 \times 16)) = 58$ g

c $m((NH_2)_2CO_2) = 6.8 \times 10^{-3} \times ((2 \times 14) + (4 \times 1) + 12 + 16) = 0.41$ g

d $m(Cu) = 6.12 \times 63.5 = 389$ g

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

a $n(C) = \frac{1.201}{12} = 0.10$ mol

b $n(S_8) = \frac{10.0}{257} = 0.0389$ mol

c $n(CH_4) = \frac{20}{16} = 1.25$ mol

d $n(C_6H_4(OCOCH_3)COOH) = \frac{0.300}{180} = 0.00167$ mol

e $n(Al_2O_3) = \frac{3\,500\,000}{102} = 3.4 \times 10^4$ mol

15 **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 $= \frac{18}{6.02 \times 10^{23}} = 3.0 \times 10^{-23}$ g

c Mass of one CO_2 molecule $= \frac{44}{6.02 \times 10^{23}} = 7.3 \times 10^{-23}$ g

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16 **a** **i** $n(P_4) = \frac{m}{M} = \frac{4.2}{4 \times 31.0} = 0.034$ mol

ii $N(P_4) = n \times 6.02 \times 10^{23} = 2.04 \times 10^{22}$ molecules

iii Total number of atoms = $4 \times 2.04 \times 10^{22} = 8.2 \times 10^{22}$ atoms

b **i** $n(S_8) = \frac{m}{M} = \frac{75.0}{8 \times 32.1} = 0.292$ mol

ii $N(S_8) = n \times 6.02 \times 10^{23} = 1.75 \times 10^{23}$ molecules

iii Total number of atoms = $0.292 \times 8 \times 6.02 \times 10^{23} = 1.41 \times 10^{24}$ atoms

c **i** $n(HCl) = \frac{m}{M} = \frac{0.32}{1.01 \times 35.5} = 0.0088$ mol

ii $N(HCl) = n \times 6.02 \times 10^{23} = 5.3 \times 10^{21}$ molecules

iii Total number of atoms = $0.0088 \times 2 \times 6.02 \times 10^{23} = 1.1 \times 10^{22}$ atoms

d **i** $n(C_4H_{12}O_6) = \frac{m}{M} = \frac{2.2 \times 10^{-2}}{(6 \times 12.01) + (12 \times 1.01) + (6 \times 16.0)} = 1.22 \times 10^{-4}$ mol

ii $n(C_6H_{12}O_6) = n \times 6.02 \times 10^{23} = 7.3 \times 10^{19}$ molecules

iii Total number of atoms = $1.22 \times 10^{-4} \times 24 \times 6.02 \times 10^{23} = 1.8 \times 10^{21}$ atoms

17 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}

$\therefore n(H_2O) = \frac{20}{18}$

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

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

18 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** $n(NaCl) = \frac{20}{18} = 0.100$ mol

ii $n(Na^+) = n(NaCl) = 0.100$ mol
$n(Cl^-) = n(NaCl) = 0.100$ mol

b **i** $n(CaCl_2) = \frac{45.0}{111} = 0.405$ mol

ii $n(Ca^{2+}) = n(CaCl_2) = 0.405$ mol
$n(Cl^-) = 2 \times n(CaCl_2) = 2 \times 0.405 = 0.81$ mol

c **i** $n(Fe_2(SO_4)_3) = \frac{1.68}{399.6} = 0.00420$ 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

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

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

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

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

a $M(\text{substance}) = \frac{80}{2} = 40$ g mol^{-1}

All other parts follow the same process.

b 98 g mol^{-1}

c 44 g mol^{-1}

d 106 g mol^{-1}

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

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

$m(\text{Fe}) = 4.0 \times 55.8 = 223$ g

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

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

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

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

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

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

c Number of molecules $= nN_A$
$= 1.6 \times 10^{-7} \times 6.02 \times 10^{23} = 9.6 \times 10^{16}$ molecules

5.4 PERCENTAGE COMPOSITION AND EMPIRICAL FORMULAS

23 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 useful check of these answers is provided by seeing that they add up to 100%, or somewhere close to that value.

a $M(\text{Al}) = 27$ g mol^{-1}, $M(\text{O}) = 16$ g mol^{-1}, $M(\text{Al}_2\text{O}_3) = 102$ g mol^{-1}

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

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

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

c $M(\text{Mg}) = 24.3$ g mol^{-1}, $M(\text{Cl}) = 35.5$ g mol^{-1}, $M(\text{MgCl}_2.6\text{H}_2\text{O}) = 203.3$ g mol^{-1}

$\%(\text{Mg}) = \frac{24.3}{203.3} \times 100 = 12.0\%$

$\%(\text{Cl}) = \frac{2 \times 35.5}{203.3} \times 100 = 34.9\%$

$\%(\text{H}) = \frac{12 \times 1}{203.3} \times 100 = 5.9\%$

$\%(\text{O}) = \frac{6 \times 16}{203.3} \times 100 = 47.2\%$

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

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

24 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 useful check of these answers is provided by seeing that they add up to 100%, or close to that value.

a $M(C_{10}H_8) = 128$ g mol^{-1}, $M(H) = 1$ g mol^{-1}, $M(C) = 12$ g mol^{-1}

$\%(C) = \frac{10 \times 12}{108} \times 100 = 93.8\%$

b %C = 40%

c %C = 19.9%

d %C = 60.0%

25 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 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	$\frac{3.575}{3.569} = 1$	$\frac{3.569}{3.569} = 1$
Round off to whole numbers	1	1

∴ empirical formula is CO

b

	C	**O**
Mass	27.2 g	72.8 g
Molar mass	12 g mol^{-1}	16 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{27.2}{12} = 2.267$ mol	$n = \frac{72.8}{16} = 4.55$ mol
Divide all by the smallest amount	$\frac{2.267}{2.267} = 1$	$\frac{4.55}{2.267} = 2$
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 g mol^{-1}	16 g mol^{-1}	1 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n(O) = \frac{34.8}{16} = 4.54$ mol	$n = \frac{364}{16} = 2.275$ mol	$n = \frac{9.1}{1} = 9.1$ mol
Divide all by the smallest amount	$\frac{4.54}{2275} = 1.99$	$\frac{2.275}{2.275} = 2$	$\frac{9.1}{2.275} = 4$
Round off to whole numbers	2	2	4

∴ empirical formula is C_2H_4O

d

	Fe	**S**	**O**
Mass	1.72 g	1.48 g	3.02 g
Molar mass	55.8 g mol^{-1}	32 g mol^{-1}	16 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{1.72}{55.8} = 4.54$ mol	$n = \frac{1.48}{32} = 2.275$ mol	$n = \frac{3.02}{16} = 0.189$ mol
Divide all by the smallest amount	$\frac{0.031}{0.031} = 1$	$\frac{0.046}{0.031} = 1.5$	$\frac{0.189}{0.031} = 6$
Round off to whole numbers	2	3	12

∴ empirical formula is $Fe_2S_3O_{12}$, which is $Fe_2(SO_4)_3$

e

	C	**Cl**	**H**
Mass	9.6 g	4.7 g	0.67 g
Molar mass	12 g mol^{-1}	35.5 g mol^{-1}	1 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{9.6}{12}$ = 0.800 mol	$n = \frac{4.7}{35.5}$ = 0.132 mol	$n = \frac{0.67}{1}$ = 0.67 mol
Divide all by the smallest amount	$\frac{0.800}{0.132} = 6$	$\frac{0.132}{0.132} = 1$	$\frac{0.67}{0.132} = 5$
Round off to whole numbers	6	1	5

∴ empirical formula is C_6H_5Cl

f

	C	**H**
Mass	4.42 g	0.842 g
Molar mass	12 g mol^{-1}	1 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{4.42}{12}$ = 0.368 mol	$n = \frac{0.842}{16}$ = 0.842 mol
Divide all by the smallest amount	$\frac{0.368}{0.368} = 1$	$\frac{0.842}{0.368} = 2.3$
Round off to whole numbers	1 × 7 = 7	16

∴ empirical formula is C_7H_{16}

26 The empirical formula provides the simplest whole-number ratio of atoms in a compound. The amount of each atom is found by using $n = \frac{m}{M}$, 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 g mol^{-1}	1 g mol^{-1}	14 g mol^{-1}	16 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{32}{12}$ = 2.67 mol	$n = \frac{6.7}{1}$ = 6.7 mol	$n = \frac{18.7}{14}$ = 1.34 mol	$n = \frac{42.6}{16}$ = 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 $C_2H_5NO_2$

27 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 W and S.

	W	**S**
Mass	1.84 g	0.64 g
Molar mass	184 g mol^{-1}	32 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{1.84}{184}$ = 0.01 mol	$n = \frac{0.64}{32}$ = 0.02 mol
Divide all by the smallest amount	$\frac{0.01}{0.01} = 1$	$\frac{0.02}{0.01} = 2$
Round off to whole numbers	1	2

∴ empirical formula is WS_2

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28 **a** Step 1: Calculate mass of each element present in 100 g.

$m(\text{C}) = 52.2$ g

$m(\text{O}) = 34.8$ g

$m(\text{H}) = 100 - 52.2 - 34.8 = 13$ g

Step 2: Calculate amount, in mol, of each element present.

$$n(\text{C}) = \frac{m}{M} = \frac{52.2}{12} = 4.35 \text{ mol}$$

$$n(\text{O}) = \frac{34.8}{16} = 2.175 \text{ mol}$$

$$n(\text{H}) = \frac{13}{1} = 13 \text{ mol}$$

Step 3: Convert to whole-number ratios.

$$n(\text{C}) = \frac{4.35}{2.175} = 2$$

$$n(\text{O}) = \frac{2.175}{2.175} = 1$$

$$n(\text{H}) = \frac{13}{2.175} = 6$$

Step 4: Write as empirical formula: C_2H_6O

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

29 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 P and O.

	O	**P**
Mass	0.456 – 0.200 = 0.256 g	0.200 g
Molar mass	16 g mol^{-1}	31 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{0.256}{16} = 0.016$ mol	$n = \frac{0.200}{31} = 0.00645$ mol
Divide all by the smallest amount	$\frac{0.016}{0.00645} = 2.5$	$\frac{0.00645}{0.00645} = 1$
Round off to whole numbers	5	2

∴ empirical formula is P_2O_5

30 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 and H.

	C	**H**
Mass	90.0 g	100 – 90.0 = 10.0 g
Molar mass	12 g mol^{-1}	1 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{90}{12} = 7.5$ mol	$n = \frac{10.0}{1} = 10$ mol
Divide all by the smallest amount	$\frac{7.5}{7.5} = 1$	$\frac{10}{7.5} = 1.3$
Round off to whole numbers	1 × 3 = 3	1.3 × 3 = 4

∴ empirical formula is C_3H_4

31 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 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 $n = \frac{m}{M}$	$n = \frac{3.370}{M(\text{Ni})}$	$n = \frac{0.916}{16} = 0.0573$ mol

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

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

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

32 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 mass of Cl.

	W	**Cl**
Mass	4.150 g	8.950 – 4.150 = 4.800 g
Molar mass	? g mol^{-1}	35.5 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{4.150}{M(\text{W})}$	$n = \frac{4.800}{35.5} = 0.0135$ mol

$$\frac{n(\text{W})}{n(\text{Cl})} = \frac{4.150}{M(\text{W}) \times 0.135}$$

As the empirical formula is WCl_6, $\frac{4.150}{M(\text{W}) \times 0.135} = \frac{1}{6}$

$$\therefore M(\text{W}) = 4.150 \times \frac{6}{0.135} = 184.2 \text{ g mol}^{-1}$$

33 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 mass of O.

	X	**O**
Mass	3.72 g	4.80 g
Molar mass	? g mol^{-1}	16 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{3.72}{M(\text{X})}$	$n = \frac{4.80}{16} = 0.3$ mol

$$\frac{n(\text{X})}{n(\text{O})} = \frac{3.72}{M(\text{X}) \times 0.3}$$

As the empirical formula is X_4O_{10}, $\frac{3.72}{M(\text{X}) \times 0.3} = \frac{4}{10}$

$$\therefore M(\text{X}) = \frac{3.72 \times 10}{0.3 \times 4} = 31.0 \text{ g mol}^{-1}$$

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34 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 mass of C and H.

a

	C	**H**
Mass	42.9 g	17.25 g
Molar mass	12 g mol^{-1}	1 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{82.75}{12} = 6.89$ mol	$n = \frac{17.25}{1} = 17.25$ mol
Divide all by the smallest amount	$\frac{6.89}{6.89} = 1$	$\frac{17.25}{6.89} = 2.5$
Round off to whole numbers	1 × 2 = 2	2.5 × 2 = 5

∴ empirical formula is C_2H_5

Molar mass of a C_2H_5 unit (empirical formula) = 24 + 5 = 29 g mol^{-1}.
Molar mass of the compound (molecular formula) = 58 g mol^{-1}.

∴ number of C_2H_5 units in one molecule = $\frac{58}{29} = 2$

∴ molecular formula is C_4H_{10}

b P_4O_{10}

c $C_6H_{12}O_6$

d 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 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 O, S and H.

	C	**S**	**O**
Mass	0.164 g	5.25 g	9.18 g
Molar mass	1 g mol^{-1}	32 g mol^{-1}	16 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{0.164}{1} = 0.164$ mol	$n = \frac{5.25}{32} = 0.164$ mol	$n = \frac{9.18}{16} = 0.574$ mol
Divide all by the smallest amount	$\frac{0.164}{0.164} = 1$	$\frac{0.164}{0.164} = 1$	$\frac{0.574}{0.164} = 3.5$
Round off to whole numbers	1 × 2 = 2	1 × 2 = 2	3.5 × 2 = 7

∴ empirical formula is $H_2S_2O_7$

Molar mass of a $H_2S_2O_7$ unit (empirical formula) = (2 × 1) + (2 × 32) + (7 × 16) = 178 g mol^{-1}.
Molar mass of the compound (molecular formula) = 178 g mol^{-1}.

∴ number of $H_2S_2O_7$ units in one molecule = $\frac{178}{178} = 1$

∴ molecular formula is $H_2S_2O_7$

CONNECTING THE MAIN IDEAS

35 **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 (see Table 5.1.3, page 112) 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 (see Worked example: Try yourself 5.1.1, on page 113) is 10.81.

c The relative molecular mass (M_r) of a compound is the mass of one molecule of that substance relative to the mass of a ^{12}C atom, which is 12 exactly. For example, the relative molecular mass of carbon dioxide is 44.0.

d Relative formula mass is calculated by taking the sum of the relative atomic masses of the elements in the formula. Relative formula mass (rather than relative molecular mass) is the appropriate term to use for ionic compounds, as these do not contain molecules. For example, the relative formula mass of sodium chloride is 58.5.

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e The molar mass of an element is the mass of one mole 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 24.3 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. The molar mass of a compound is the mass of one mole of the compound. It is equal to the relative molecular or relative formula mass of the compound expressed in grams. For example, the molar mass of sodium chloride is 58.5 g mol^{-1}.

36 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 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.

Take 100 g of caffeine. Mass of oxygen can be found by subtraction
= 100 – (49.48 + 5.15 + 28.87)
= 16.5 g

	C	**H**	**N**	**O**
Mass	49.48 g	5.15 g	28.87 g	16.5 g
Molar mass	12.0 g mol^{-1}	1.01 g mol^{-1}	32 g mol^{-1}	16 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{49.48}{12.0}$ = 0.4.17 mol	$n = \frac{5.15}{1.01}$ = 5.10mol	$n = \frac{28.87}{14.0}$ = 2.06 mol	$n = \frac{16.5}{16}$ = 1.03 mol
Divide all by the smallest amount	$\frac{4.17}{1.03} = 4.04$	$\frac{5.10}{1.03} = 5.0$	$\frac{2.06}{1.03} = 2.0$	$\frac{1.03}{1.03} = 1.0$
Round off to whole numbers	4	5	2	1

∴ empirical formula is $C_4H_5N_2O$

b Molar mass of caffeine = $38.8 \times \frac{1}{0.2}$ = 194 g mol^{-1}

c The molecule must contain a whole number of ($C_4H_5N_2O$) units.
Molar mass of a $C_4H_5N_2O$ unit is
(4 × 12.0) + (5 × 1.01) + (2 × 14.0) + (16.0) = 97 g mol^{-1}.
If the compound has a molar mass of 194 g mol^{-1}, then the number of ($C_4H_5N_2O$) units in a molecule

$$= \frac{\text{molar mass of compound}}{\text{molar mass of one unit}}$$

$$= \frac{194\ \text{g mol}^{-1}}{97\ \text{g mol}^{-1}} = 2$$

The molecular formula of caffeine is therefore 2 × ($C_4H_5N_2O$); that is, $C_8H_{10}N_4O_2$.

d $n(\text{caffeine}) = \frac{m}{M_r} = \frac{1.00}{194}$
$= 5.15 \times 10^{-3}$ mol

e $N(\text{caffeine})$ molecules = $n(\text{caffeine}) \times N_A$
$= 5.15 \times 10^{-3} \times 6.02 \times 10^{23} = 3.10 \times 10^{21}$ molecules of caffeine

f $N(\text{atoms in caffeine}) = N(\text{molecules}) \times 6.02 \times 10^{23} = 7.44 \times 10^{24}$ atoms

37 **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.

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

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

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.

$$M = \frac{m}{n} = \frac{0.542}{0.135} = 40.1$$

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 periodic table to identify metal as calcium.

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Chapter 6 Materials made of molecules

Section 6.1 Properties of non-metallic substances

6.1 Key questions

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

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

3 **a** Non-metals do not conduct electricity because they do not contain free-moving charged particles (neither delocalised electrons nor ions).
  b Non-metals have low melting and boiling temperatures because they have weak intermolecular bonds between molecules.

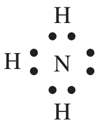
4 **a** True. Methane is formed from two types of non-metal atoms (carbon and hydrogen).
  b False. Methane contains 5 atoms (1 carbon and 4 hydrogen atoms).
  c True. There are 5 atoms per molecule and so 4 molecules will have 20 atoms.
  d False. Bonds in the molecule are intramolecular bonds.

5 When sugar turns to a liquid, it is melting; the intermolecular bonds are broken. When the liquid turns black and a gas is produced, the intramolecular bonds are broken, allowing new substances to be produced.

Section 6.2 Covalent bonding

Worked example: Try yourself 6.2.1

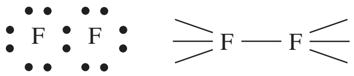
ELECTRON DOT DIAGRAMS

Draw an electron dot diagram of the molecule formed from nitrogen and hydrogen atoms.	
Thinking	**Working**
Write the electronic configuration of the atoms in the molecule.	N electronic configuration: 2,5 H electronic configuration: 1
Determine how many electrons each atom requires for a stable outer shell.	N requires 3 electrons H requires 1 electrons
Draw an electron dot diagram of the likely molecule, ensuring that each atom has a stable outer shell. Electrons not involved in bonding will be in non-bonding (lone) pairs.	Draw an electron dot diagram of the molecule. 

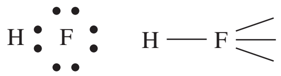
6.2 Key questions

1 **a** 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.

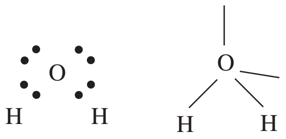
2 fluorine (F_2)



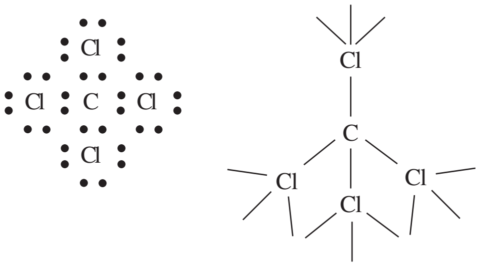
hydrogen fluoride (HF)



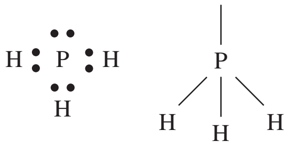
water (H_2O)



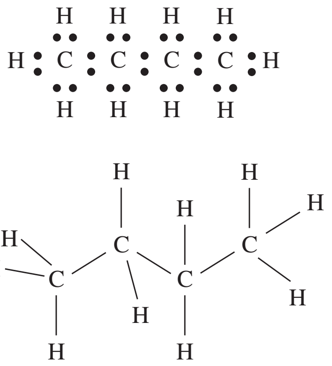
tetrachloromethane (CCl_4)



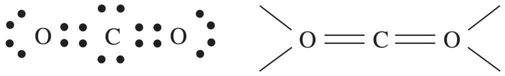
phosphine (PH_3)



butane (C_4H_{10})



carbon dioxide(CO_2)



3
- **a** 1
- **b** 2
- **c** 3
- **d** 4
- **e** 1
- **f** 0

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

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

6
- **a** Ball-and-stick model
- **b** Electron dot diagram
- **c** Space-filling model

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CHAPTER 6 REVIEW

6.1 PROPERTIES OF NON-METALLIC SUBSTANCES

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

2 Non-metallic substances do not conduct electricity.

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

4 **a** Space-filling model
  b Valence structure
  c Ball-and-stick model
  d Electron dot diagram

6.2 COVALENT BONDING

5 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.

6 **a** 6. Each nitrogen atom shares three electrons with the other nitrogen atom.
  b 3. Each nitrogen atom shares three electrons with the other nitrogen atom, so three bonding pairs are formed, which is equal to a triple covalent bond.
  c Each nitrogen atom has two non-bonding electrons, making a total of four when both atoms are counted.

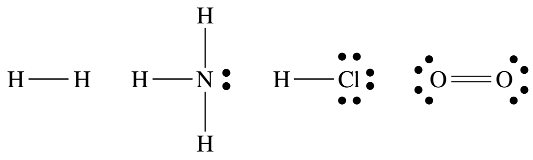
7 B. To have a formula XY_4, atom X must require four more outer-shell electrons and atom Y must require one. The likely molecule is CH_4 as carbon has four outer-shell electrons and hydrogen requires one more.

Oxygen has 6 valence electrons and needs to form 2 covalent bonds. Hydrogen has one valence electron and needs to form one bond. For OH_4, each hydrogen would need to form a single covalent bond, giving a total of 4, but oxygen only needs to form 2, so this option is incorrect. Bromine has 7 valence electrons. Each bromine needs to form a single covalent bond, giving a total of 4, but hydrogen only needs to form one bond, so this option is incorrect. Carbon has 4 outer-shell electrons and needs to form 4 bonds. Oxygen has 6 valence electrons and needs to form 2 covalent bonds, to have a total of 8, but carbon only needs to form 4, so this option is incorrect.

8 Atom X requires 2 electrons to form a stable outer shell. Each atom of element Y can share 1 electron. Therefore the molecular formula will be XY_2.

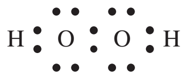
9 Neon will not form bonds to other atoms as it has a stable outer shell.

10 **a** 0
  b 1
  c 3
  d 4

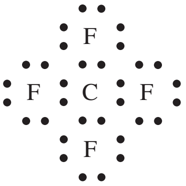


11 The diagrams of each molecule below show the number of electron pairs.
  a 2 bonding electrons, 6 non-bonding electrons

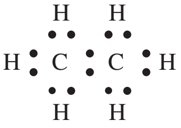
  b 4 bonding electrons, 4 non-bonding electrons



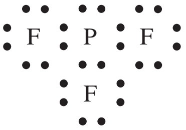
c 8 bonding electrons, 24 non-bonding electrons



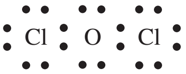
d 14 bonding electrons, 0 non-bonding electrons



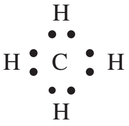
e 6 bonding electrons, 20 non-bonding electrons



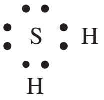
f 4 bonding electrons, 16 non-bonding electrons



g 8 bonding electrons



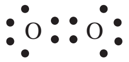
h 4 bonding electrons, 4 non-bonding electrons

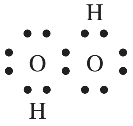


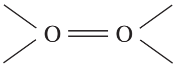
12 a The bonds are similar in that they all involve the sharing of electron pairs between two atoms; that is, they are covalent bonds.

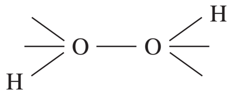
b They differ in the number of electron pairs shared: one pair (fluorine), two pairs (oxygen) and three pairs (nitrogen).

13 a









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

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

CONNECTING THE MAIN IDEAS

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

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

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

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

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

b True. Carbon dioxide contains 3 atoms (1 carbon and 2 oxygen atoms).

c False. There are 2 oxygen atoms per molecule and so 33 molecules will have 66 oxygen atoms.

d True. Bonds in the molecule are intramolecular bonds.

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

f True. Each oxygen atom has 2 lone pairs of electrons. As there are 2 oxygen atoms, there are 4 lone pairs of electrons in a molecule of carbon dioxide.

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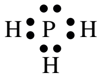
Chapter 7 Intermolecular forces

Section 7.1 Shapes of molecules

Worked example: Try yourself 7.1.1

PREDICTING THE SHAPE OF MOLECULES

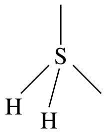
Predict the shape of a molecule of hydrogen sulfide (H_2S).	
Thinking	**Working**
Draw the electron dot diagram for the molecule.	
Count the number of bonds and lone pairs on the central atom.	There are two bonds and two lone pairs.
Determine how the groups of electrons will be arranged to get maximum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

7.1 Key questions

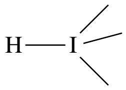
1 The VSEPR theory is based on the principle that negatively charged electron pairs in the outer shell of an atom repel each other. As a consequence, these electron pairs are arranged as far away from each other as possible.

2 4. One pair is in the single bond between the fluorine and hydrogen atoms; the other three pairs exist as lone pairs of electrons.

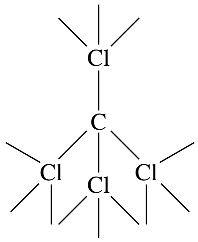
3 **a**



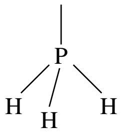
b



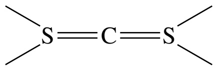
c



d



e



4 **a** V-shaped
b Linear
c Tetrahedral
d Pyramidal
e Linear

5 **a** Tetrahedral
b Pyramidal
c V-shaped

Section 7.2 Properties of covalent molecular substances

Worked example: Try yourself 7.2.1

COMPARING THE POLARITY OF COVALENT BONDS

Compare the polarity of the covalent bonds in nitrogen monoxide (NO) and hydrogen chloride (HCl).	
Thinking	**Working**
Use the table of electronegativity values in Figure 7.2.3 on page 163 to find the electronegativities of the atoms in each molecule.	NO: nitrogen 3.04; oxygen 3.44 HCl: hydrogen 2.20; chlorine 3.16
For each molecule, subtract the lowest electronegativity value from the highest value.	NO: 3.44 – 3.04 = 0.40 HCl: 3.16 – 2.20 = 0.96
Determine which molecule has the biggest difference in electronegativity in order to determine the more polar molecule.	HCl is more polar than NO.

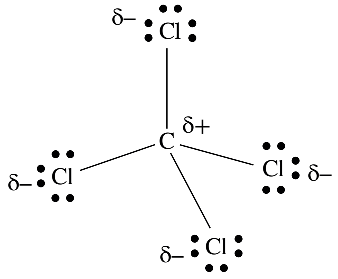
7.2 Key questions

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

2 **a** P–F
b C–H

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

4



5 **a** Polar
b Polar
c Polar
d Polar
e Non-polar

Section 7.3 Types of intermolecular forces

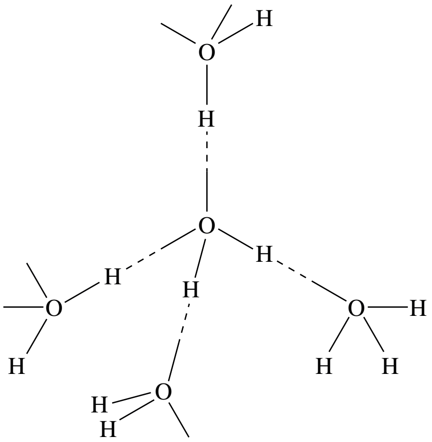
7.3 Key questions

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

2 C. Hydrogen chloride

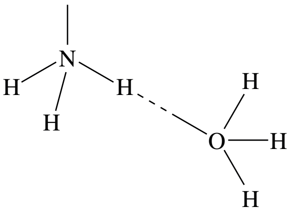
3 Dipole–dipole forces: a to h; hydrogen bonds: a, g, h

4



Because of hydrogen bonding, ice is less dense than liquid water, and so ice floats on water. (For most liquids, the solid is denser than the liquid.) This is good news for fish, but not good news for travellers on the *Titanic*!

5



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

CHAPTER 7 REVIEW

7.1 SHAPES OF MOLECULES

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

2 V-shaped

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

4 **a** Tetrahedral
  b Pyramidal
  c Tetrahedral
  d V-shaped
  e Pyramidal

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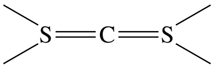
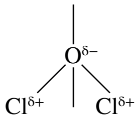
7.2 PROPERTIES OF COVALENT MOLECULAR SUBSTANCES

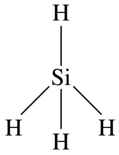
5 **a** Non-polar



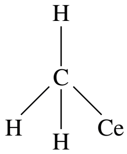
b Polar



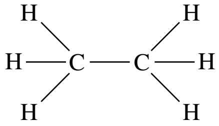
c Non-polar



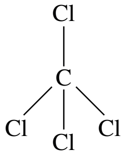
d Polar



e Non-polar



f Non-polar



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

7 **a** Non-polar
b Polar
c Non-polar
d Polar
e Polar

8 Si–O, H–Br, N–O, O–Cl, F–F

7.3 TYPES OF INTERMOLECULAR FORCES

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

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

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

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

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

14 **a** CCl_4

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

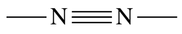
15 Iodine is a much larger molecule, so it has more electrons and therefore the dispersion forces are much stronger in iodine than in fluorine.

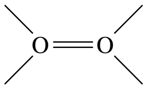
16 A permanent dipole is formed if there is a difference in electronegativity between the two atoms that form a bond. The more electronegative atom has a partial negative charge and the less electronegative atom has a partial positive charge. In larger molecules, individual bonds may be polar but if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

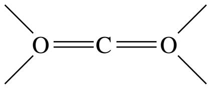
A temporary dipole is caused by random fluctuations in the electron distributions around the molecule. The electrons are constantly moving and can occasionally concentrate at one end of the molecule, causing that end to develop a temporary negative charge while the other end develops a temporary positive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

CONNECTING THE MAIN IDEAS

17







a **i** N_2
  ii O_2
  iii CO_2

b **i** NH_3, HCl, H_2O, $CHCl_3$
  ii N_2, Cl_2, O_2, CH_4, CO_2, CCl_4
  iii NH_3, H_2O

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

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

Chapter 8 Carbon lattices and carbon nanomaterials

Section 8.1 Carbon lattices

8.1 Key questions

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

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

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

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

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

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

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

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

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

Section 8.2 Carbon nanomaterials

8.2 Key questions

1 Fullerenes are similar to graphite in that they are allotropes of carbon in which each carbon atom has bonds to 3 other carbon atoms. Fullerenes conduct electricity and heat, similar to graphite. Fullerenes differ from graphite in that they are nanomaterials, which give them different physical properties.

2 Each carbon atom in a buckyball is covalently bonded to 3 other carbon atoms. Each carbon atom has one free electron, which is shared between the other carbon atoms in the sphere. The structure consists of hexagonal and pentagonal rings of atoms.

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

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

CHAPTER 8 REVIEW

8.1 CARBON LATTICES

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

2 Refer to Figure 8.1.8. Carbon atoms bond covalently to three other carbon atoms in graphite to form layers. These layers consist of hexagonal rings connected to each other. The fourth electron in each carbon atom is delocalised, which explains its electrical conductivity. There are weak dispersion forces between layers of graphite, allowing the layers to slide over each other and enabling it to act as a lubricant.

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

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

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

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

8.2 CARBON NANOMATERIALS

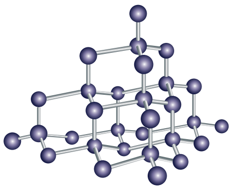
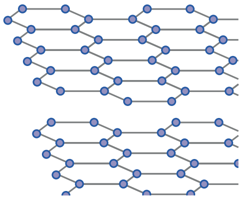
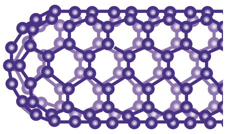
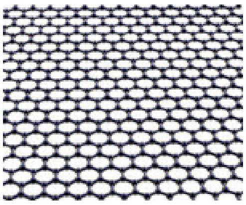
6 Carbon atoms in diamond have a tetrahedral bond geometry; carbon atoms in graphene have a trigonal planar bond geometry.

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

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

CONNECTING THE MAIN IDEAS

9

				
Allotrope name	Diamond	Graphite	Carbon nanotube	Graphene
Bonding feature	Each carbon is bonded to 4 others in a network lattice structure	Each carbon is bonded to 3 others in a layered structure	Each carbon is bonded to 3 others in a cylindrical shape	Each carbon is bonded to 3 others in a layer one atom thick
Properties	Hard, brittle	Conducts electricity, soft	Very strong, conducts electricity	Very strong, conducts electricity

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

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

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

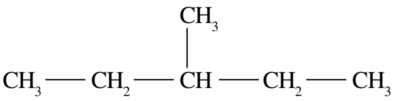
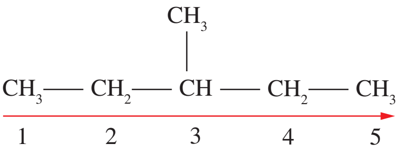
b Graphite, graphene and nanotubes are all conductive of electricity. Diamond is non-conductive of electricity.

Chapter 9 **Organic compounds**

Section 9.1 **Crude oil as a source of raw materials**

Worked example: Try yourself 9.1.1

IUPAC NAMING SYSTEM FOR ALKANES

Write the systematic name for the following molecule.  $CH_3—CH_2—CH(CH_3)—CH_2—CH_3$	
Thinking	**Working**
Identify the longest carbon chain in the molecule. The name of the molecule is based on this longest chain.	The longest chain has 5 carbons in the longest chain. The stem name is based on pentane.
Number the carbons starting from the end closest to the branch.	 Note: In this molecule, it does not matter which end the numbering begins at as the side branch comes off central carbon which is the third carbon.
Identify the branch.	The side chain is a methyl group.
Combine all components.	The name of the molecule is 3-methylpentane.

9.1 Key questions

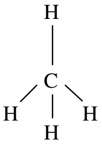
1 A, D

2 **a** CH_4

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

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

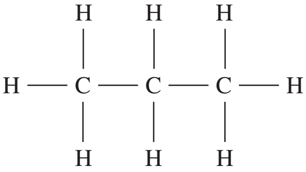
d This tetrahedral arrangement gives minimum electrostatic repulsion between the four pairs of bonding electrons. The hydrogen atoms are arranged around the central carbon atom in a tetrahedral configuration.



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

b $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$

c

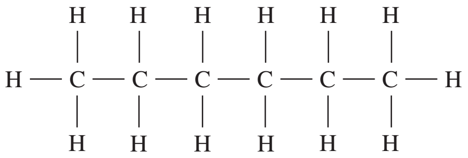


d $CH_3CH_2CH_3$

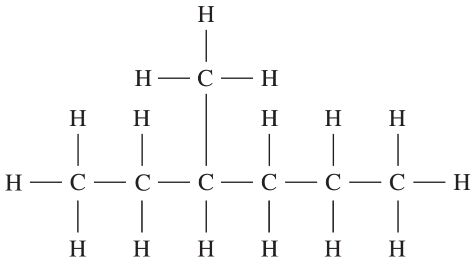
4
- **a** Butane. The longest chain of carbon atoms is 4 and there are no branches.
- **b** Heptane. The longest chain of carbon atoms is 7 and there are no branches. Note that the prefix hept- rather than sept- is used to denote 7.
- **c** 2-Methylpentane. The longest chain of carbon atoms is 5, giving the stem name of pentane, and there is one methyl, $–CH_3$, branch on carbon number 2 when numbered from the end closest to the branch.
- **d** 2-Methylbutane. The longest chain of carbon atoms is 4, giving the stem name of butane, and there is one methyl, $–CH_3$, branch on carbon number 2 when numbered from the end closest to the branch.
- **e** 2,4-Dimethylpentane. The longest chain of carbon atoms is 5, giving the stem name of pentane, and there are two methyl, $–CH_3$, branches: one on carbon number 2 and the other on carbon number 4 when numbered in either direction. Dimethyl is used in addition to the two numbers to indicate the presence of two separate methyl groups.

5
- **a** Butane
- **b** Methylpropane
- **c** 2,4-Dimethylpentane
- **d** 3-Methylhexane

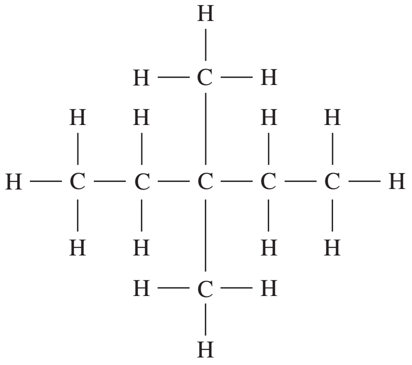
6
- **a** Hexane



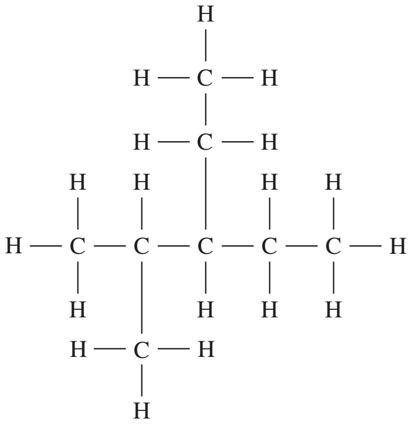
- **b** 3-Methylhexane



- **c** 3,3-Dimethylpentane



- **d** 3-Ethyl-2-methylpentane



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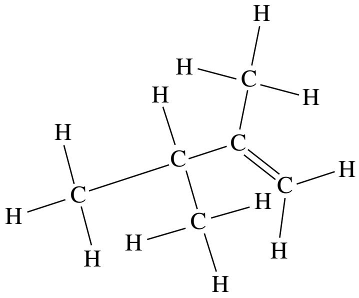
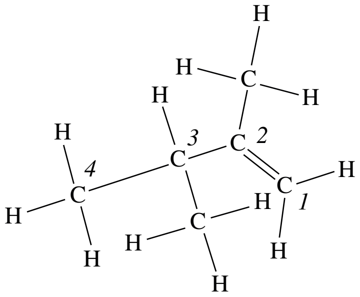
Section 9.2 Unsaturated hydrocarbons

Worked example: Try yourself 9.2.1

IUPAC NAMING SYSTEM FOR UNSATURATED HYDROCARBONS

<table>
<tr><td colspan="2">Write the systematic name for the following molecule.
</td></tr>
<tr><td>Thinking</td><td>Working</td></tr>
<tr><td>Identify the longest carbon chain in the molecule that contains the double or triple bond.
The name of the molecule is based on this longest chain.</td><td>There are 4 carbons in the longest chain.
The name is based on butene.</td></tr>
<tr><td>Number the carbon atoms starting from the end closest to the double or triple bond.
Note the position of any double or triple bond.</td><td>Number each carbon atom.

There is a double bond on carbon number 1 so the longest chain is but-1-ene.</td></tr>
<tr><td>Identify any branches and the number carbon that it is on.</td><td>The side chains are methyl groups and they are on carbons numbered (from smallest to largest) 2 and 3.</td></tr>
<tr><td>Combine all components.</td><td>The name of the molecule is 2,3-dimethylbut-1-ene.</td></tr>
</table>

9.2 Key questions

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

2 a 4. The structure of methylpropene is based on propene with a methyl group attached.
  b 8
  c Methylpropene is an isomer of butene (C_4H_8).

3 B. The condensed structural formula shows the arrangement of the atoms in a molecule, but does not show the bonds. A is the molecular formula of pentyne. D is the structural formula of pentyne and C is the condensed structural formula of pentene.

4 a 5. The structure of 3-methylbut-1-yne is based on but-1-yne with a methyl group attached.
  b 8. Methylpropene is an isomer of pentyne (C_5H_8).
  c Pentyne

5 a Hept-3-ene
  b Hex-3-yne
  c 2,3-Dimethylbut-2-ene.

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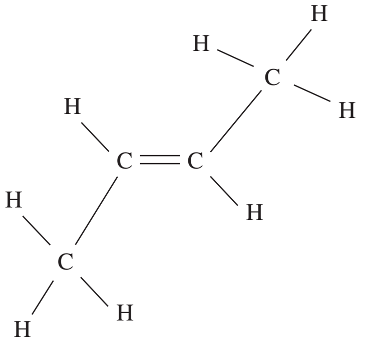
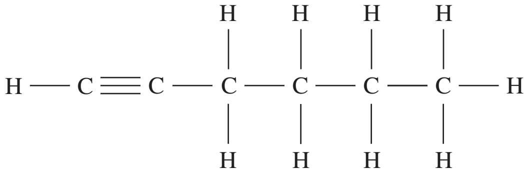
6 **a** But-2-ene

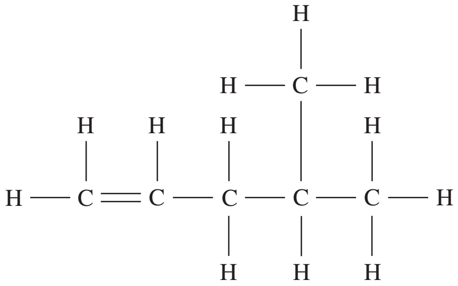


b Hex-1-yne

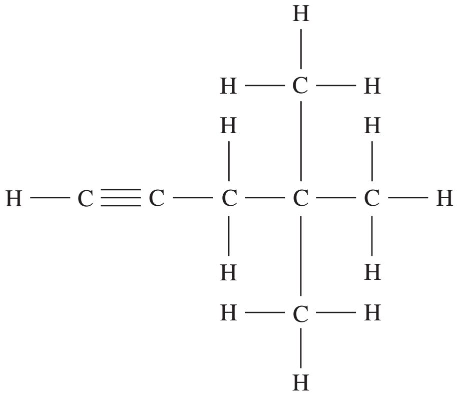


$$\begin{array}{ccccccccccccccc} & & & & & & H & & H & & H & & H & & \\ & & & & & & | & & | & & | & & | & & \\ H & - & C & \equiv & C & - & C & - & C & - & C & - & C & - & H \\ & & & & & & | & & | & & | & & | & & \\ & & & & & & H & & H & & H & & H & & \end{array}$$

c 4-Methylpent-1-ene



d 4,4-Dimethylpent-1-yne



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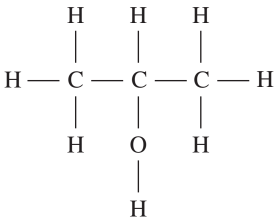
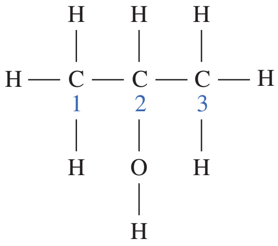
Section 9.3 Alcohols

Worked example: Try yourself 9.3.1

IUPAC NAMING SYSTEM FOR ALCOHOLS

Write the systematic name for the following molecule.

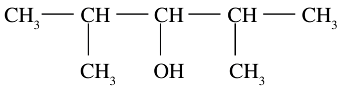


Thinking	Working
Identify the longest carbon chain in the molecule. The name of the molecule is based on this longest chain.	There are three carbons in the longest chain. The name is based on propane.
Identify the functional group that is present.	There is a hydroxyl group present.
Number the carbon atoms starting from the end closest to the functional group.	
Identify the position(s) and the type(s) of branch.	There are no branches in this molecule.
Combine all components. Place the number for the position of the branch 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 9.3.2

IUPAC NAMING SYSTEM FOR ALCOHOLS

Write the systematic name for the following molecule.



$CH_3-CH(CH_3)-CH(OH)-CH(CH_3)-CH_3$

Thinking	Working
Identify the longest carbon chain in the molecule. The name of the molecule is based on this longest chain.	There are 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 branches.	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 branch 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-dimethylpent-3-ol.

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9.3 Key questions

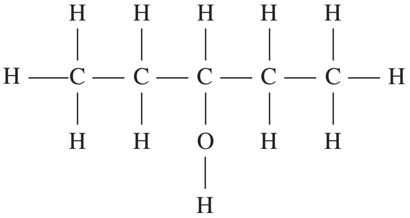
1 C. CH_3OH is the smallest alcohol with the shortest carbon chain.

2 C

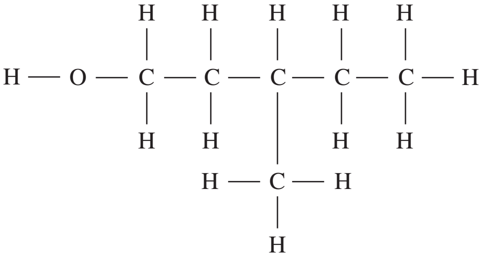
3 CH_3OH, C_4H_9OH, $C_5H_{11}OH$

4
- **a** Butan-2-ol
- **b** 3-Methylheptan-2-ol
- **c** 2,3-Dimethylhexan-1-ol

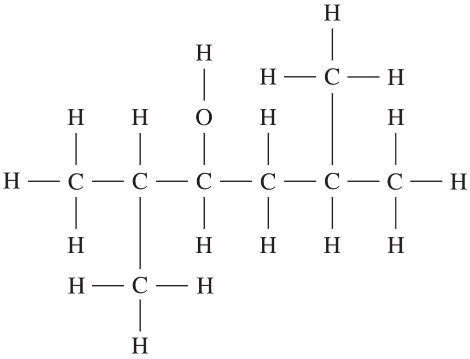
5 **a** Pentan-3-ol



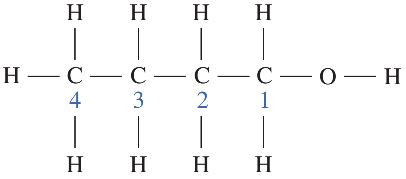
b 3-Methylpentan-1-ol



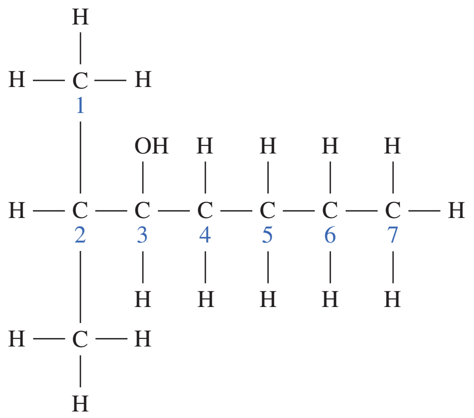
c 2,5-Dimethylhexan-3-ol



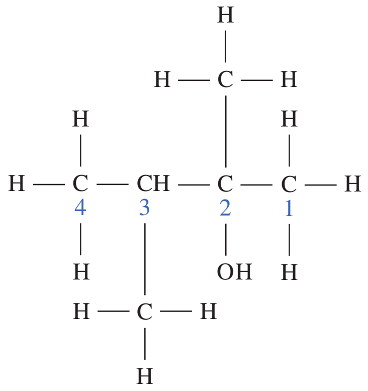
6 **a** The correct name is butan-1-ol. Carbon number 4 should be labelled carbon number 1 as that is the end closest to the hydroxyl group.



b The correct name is 2-methylheptan-3-ol. One of the mislabelled methyl branches is actually part of the longest carbon chain. Consequently, the hydroxyl group is on carbon number 3.



c 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.

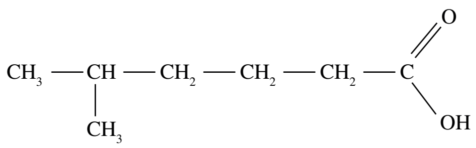


Section 9.4 Carboxylic acids

Worked example: Try yourself 9.4.1

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 six 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 branches.	There is a methyl group on C5.
Combine all components. Place the number for the position of the branch in front of the prefix and using the *–oic acid* ending.	The name of the molecule is 5-methylhexanoic acid.

9.4 Key questions

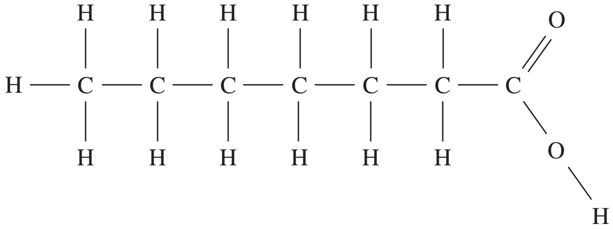
1 D

2 A. Option B is incorrect because alcohols can also form hydrogen bonds. Options C and D are unrelated.

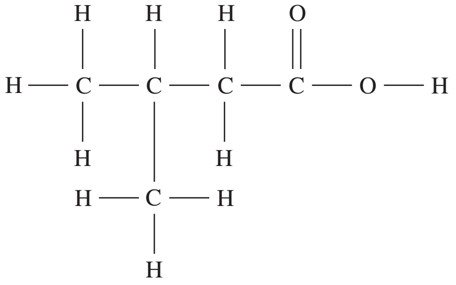
3 **a** $CH_3CH_2CH_2COOH$
b $(CH_3)_2CHCOOH$
c $CH_3CH_2CH(CH_3)CH(CH_3)COOH$
d $(CH_3)_3CHCH_2COOH$

4 **a** Propanoic acid
b 2-Methylbutanoic acid
c 2-Ethylhexanoic acid
d 3,4-Dimethylpentanoic acid

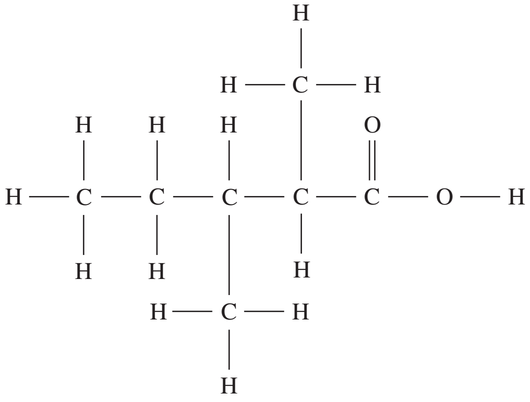
5 **a**



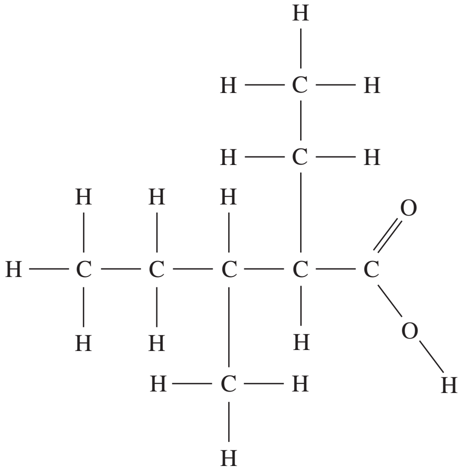
b



c



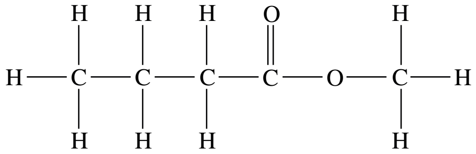
d



Section 9.5 **Esters**

Worked example: Try yourself 9.5.1

IUPAC NAMING SYSTEM FOR ESTERS

Write the systematic name for the following molecule. 	
Thinking	**Working**
Locate the ester functional group. Starting from the carbonyl carbon, determine the number of carbon atoms in the chain attached to it and name according to the stem name, adding the suffix *'-oate'*.	There are four carbons in the chain, which includes the carbonyl carbon, so the last part of the molecule's name is butanoate.
Count how many carbons are bonded to the singly bonded oxygen atom. The alkyl group bonded to the singly bonded oxygen atom is then added in front of the name.	There is one carbon bonded to the single bonded oxygen atom, so the alkyl group is methyl.
Combine the two components, placing the alkyl part of the name first.	The name of the molecule is methyl butanoate.

9.5 Key questions

1 C

2 C. Methyl = 1 carbon atom; propanoate = 3 carbon atoms. 3 + 1 = 4 carbon atoms in total. Option C is correct because ethyl = 2 carbon atoms; ethanoate = 2 carbon atoms: 2 + 2 = 4 carbon atoms

3
- **a** 2. Methyl = 1 carbon atom; methanoate = 1 carbon atom: 1 + 1 = 2 carbon atoms
- **b** 5. Ethyl = 2 carbon atoms; propanoate = 3 carbon atoms: 2 + 3 = 5 carbon atoms
- **c** 4. Propyl = 3 carbon atoms; methanoate = 1 carbon atom: 3 + 1 = 4 carbon atoms
- **d** 9. Pentyl = 5 carbon atoms; butanoate = 4 carbon atoms: 5 + 4 = 9 carbon atoms
- **e** 7. Butyl = 4 carbon atoms; propanoate = 3 carbon atoms: 4 + 3 = 7 carbon atoms

4
- **a** CH_3COOCH_3
- **b** $CH_3CH_2COOCH_2CH_3$
- **c** $CH_3COOCH_2CH_3$
- **d** $CH_3CH_2CH_2COOCH_2CH_2CH_3$

5
- **a** Methylpropanoate
- **b** Ethylpropanoate
- **c** Butylethanoate
- **d** Butylbutanoate

Section 9.6 Determining formulas of organic compounds

Worked example: Try yourself 9.6.1

CALCULATING PERCENTAGE COMPOSITION OF ORGANIC COMPOUNDS

8.38 g of an organic compound contains 5.44 g carbon, 1.13 g hydrogen and 1.81 g oxygen. Determine the percentage composition of the compound. Give all answers to three significant figures.

Thinking	**Working**
Divide the mass of the first element by the total mass of the compound and multiply by 100.	$\%\ \text{carbon} = \frac{\text{mass of the element present}}{\text{total mass of the compound}} \times 100$ $= \frac{5.44}{8.38} \times 100$ $= 64.9\%$
Divide the mass of the second element by the total mass of the compound and multiply by 100.	$\%\ \text{hydrogen} = \frac{\text{mass of the element present}}{\text{total mass of the compound}} \times 100$ $= \frac{1.13}{8.38} \times 100$ $= 13.5\%$
Repeat the previous step for any further elements in the compound.	$\%\ \text{oxygen} = \frac{\text{mass of the element present}}{\text{total mass of the compound}} \times 100$ $= \frac{1.81}{8.38} \times 100$ $= 21.6\%$
Add up all the elemental percentages to check that they equal 100%. (If they don't, then check your calculations and rounding off.)	If your calculations above are correct, then Total % elements = % carbon + % hydrogen + % oxygen = 64.9 + 13.5 + 21.6 = 100%

Worked example: Try yourself 9.6.2

CALCULATING 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: $n = \frac{m}{M}$	$n(\text{C}) = \frac{40.0}{12.0}$ $= 3.33$ $n(\text{H}) = \frac{6.7}{1.0}$ $= 6.7$ $n(\text{O}) = \frac{53.3}{16.0}$ $= 3.33$

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Simplify the ratio by dividing all of the numbers of moles by the smallest number of moles calculated above. This gives you a ratio by number of atoms.	$C = \frac{3.33}{3.33}$ $= 1$ $H = \frac{6.7}{3.33}$ $= 2$ $O = \frac{3.33}{3.33}$ $= 1$
Find the simplest whole number ratio.	C : H : O 1 2 1
Write the empirical formula.	CH_2O

Worked example: Try yourself 9.6.3

CALCULATING THE MOLECULAR FORMULA OF A COMPOUND

A compound has the empirical formula CH. The molar mass of this compound is 78 g mol^{-1}. What is the molecular formula of the compound?	
Thinking	**Working**
Calculate the molar mass, in g mol^{-1}, of one unit of the empirical formula.	$M(CH) = 12.0 + 1.0$ $= 13.0$
Determine the number of empirical formula units in the molecular formula.	Number of empirical formula units $= \frac{\text{molar mass}}{\text{molar mass of empirical formula unit}}$ $= \frac{78}{13.0}$ $= 6$
Multiply the empirical formula by this factor to determine the molecular formula of the compound.	The molecular formula is 6 × CH i.e. C_6H_6

9.6 Key questions

1 D. $\frac{4.66}{6.84} \times 100 = 68.1\%$

2 $\% \text{ of element} = \frac{\text{mass of element present}}{\text{total mass of the compound}} \times 100$

$\% \text{ of carbon} = \frac{6.13}{8.00} \times 100 = 76.6\%$

$\% \text{ of hydrogen} = \frac{0.51}{8.00} \times 100 = 6.4\%$

Mass of oxygen = mass of compound – mass of carbon – mass of hydrogen
= 8.00 – 6.13 – 0.51 = 1.36 g

$\% \text{ of oxygen} = \frac{1.36}{8.00} \times 100 = 17\%$

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3 **a** Carbon = 60.0 g; oxygen = 26.7 g; hydrogen = 100.0 – 60.0 – 26.7 = 13.3 g

b $n(\text{C}) = \frac{60.0}{12.0} = 5.00$ mol

$n(\text{O}) = \frac{26.7}{16.0} = 1.67$ mol

$n(\text{H}) = \frac{13.3}{1} = 13.3$ mol

c $\text{C} = \frac{5.00}{1.67} = 2.99$

$\text{O} = \frac{1.67}{1.67} = 1$

$\text{H} = \frac{13.3}{1.67} = 7.96$

d C : O : H = 3 : 1 : 8

e C_3H_8O

4 $M(\text{mannitol}) = \frac{m}{M} = \frac{152.7}{0.8390} = 182.0 \text{ g mol}^{-1}$

$M(\text{EF}) = (3 \times 12.0) + (7 \times 1.0) + (3 \times 16.0) = 91.0 \text{ g mol}^{-1}$

$\text{number of empirical formula units} = \frac{\text{molar mass}}{\text{molar mass of empirical formula unit}}$

$= \frac{182.0}{91} = 2$

Molecular formula = 2 × $C_3H_7O_3$ = $C_6H_{14}O_6$

CHAPTER 9 REVIEW

9.1 CRUDE OIL AS A SOURCE OF RAW MATERIALS

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

2 A

3 **a** Alkanes

b $C_{17}H_{36}$

c $C_{15}H_{32}$

d Percentage by mass of an element

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

$M(C_{16}H_{34}) = 226 \text{ g mol}^{-1}$

$\therefore \%(\text{C}) = \frac{16 \times 12.0}{226} \times 100$

= 85.0%

So, $M(\text{C})$ in 275 g of compound $= \frac{85.0 \times 275 \text{ g}}{100}$

= 234 g

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

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

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

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

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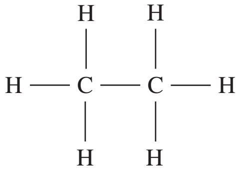
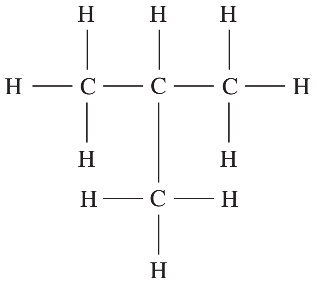
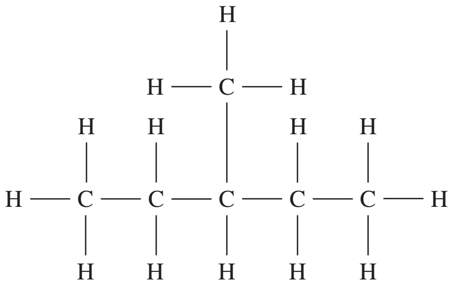
5 **a** Ethane



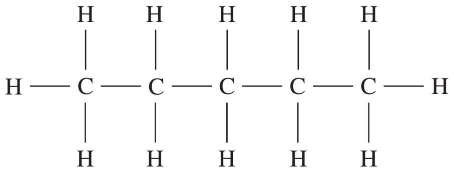
b Methylpropane



c 3-Methylpentane



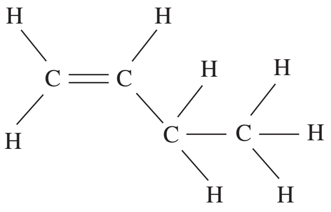
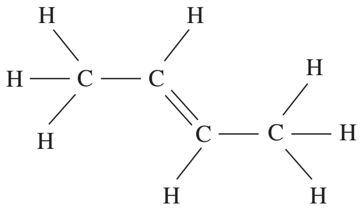
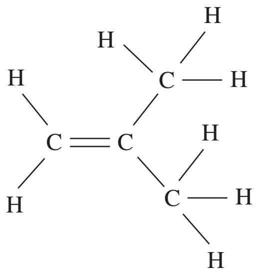
d Pentane



9.2 UNSATURATED HYDROCARBONS

6 **a** Alkane
b Alkene
c Alkane
d Alkyne
e Alkene
f Alkane
g Alkyne

7

But-1-ene	But-2-ene	Methylpropene
		

8 **a** Hex-2-ene
b 3,4-Dimethylpent-1-yne
c 4-Methylhex-2-yne

9 **a** $2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l)$
b $C_3H_6(g) + Br_2(org) \rightarrow C_3H_6Br_2(org)$
c $C_4H_8(g) + H_2O(l) \rightarrow C_4H_9OH(l)$
d $C_2H_4(g) + HCl(g) \rightarrow C_2H_5Cl(l)$

9.3 ALCOHOLS

10 hydroxyl; polar; hydrogen; intermolecular; small

11 **a** True. A functional group is an atom or group of atoms that changes the chemical and physical properties of a molecule.
b True.
c False. Alcohols contain the hydroxyl functional group.

12 Alcohols contain a hydroxyl group so hydrogen bonds can form between alcohol molecules. Hydrogen bonds are stronger than the dispersion forces found between alkane molecules. This results in alcohols having a higher boiling point than their corresponding alkane.

13 **a** Hydrogen bonds and dispersion forces are the intermolecular forces in both methanol and butan-1-ol, but butan-1-ol has a higher boiling point because it is a larger molecule, so the dispersion forces between its molecules are stronger.
b In butanol, the hydroxyl group could be placed at either carbon number 1 or 2, hence a number is required to indicate its position. In methanol, there is only one carbon atom so the position of the hydroxyl group does not need to be indicated.

9.4 CARBOXYLIC ACIDS

14 D. Only Option D contains the carboxyl functional group (–COOH).

15 A.

16 As the length of the carbon chain increases, the proportion of the carboxylic acid molecule that is non-polar increases, decreasing its solubility in water.

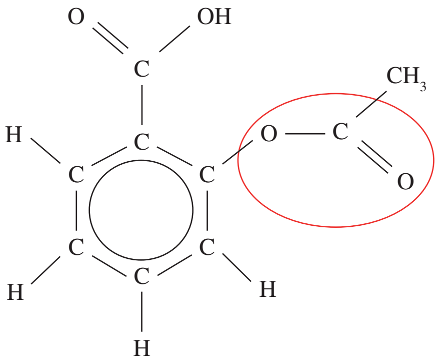
17 **a** Pentanoic acid
b 3-Methylbutanoic acid
c 4-Methylhexanoic acid
d 3-Ethylpentanoic acid

9.5 ESTERS

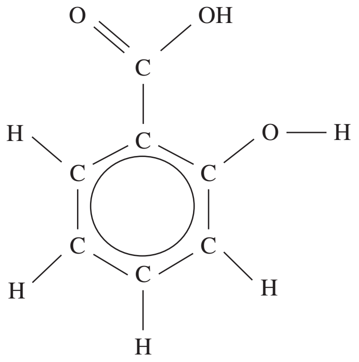
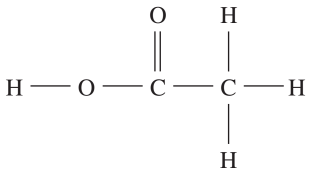
18 **a** Methylethanoate
b Ethylethanoate
c Ethylpropanoate

19 **a** $CH_3COOH(l) + CH_3CH_2CH_2OH(l) \rightarrow CH_3COOCH_2CH_2CH_3(l) + H_2O(l)$
b Esterification reaction
c Sulfuric acid acts as a catalyst

20 **a** Ester functional group is circled in red.



b

The alcohol	The carboxylic acid
	 $\mathrm{H{-}O{-}C({=}O){-}CH_3}$

9.6 DETERMINING THE FORMULAS OF ORGANIC COMPOUNDS

21 $m(\text{C}) = 92.3\% \times \text{mass of hydrocarbon} = 1.15\text{ g}$

$\text{mass of hydrocarbon} = \frac{1.15}{92.3\%} = \frac{1.15}{0.923} = 1.25\text{ g}$

$m(\text{H}) = \text{mass of hydrocarbon} - m(\text{C}) = 1.25 - 1.15\text{ g} = 0.10\text{ g}$

22

Empirical formula	**Relative molecular mass**	**Molecular formula**
CH	78.0	C_6H_6
CH_3	30.0	C_2H_6
CH_2O	90.0	$C_3H_6O_3$
C_3H_6O	116.0	$C_6H_{12}O_2$
CH_2	98.0	C_7H_{14}

23 **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 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}.

	C	**H**
Mass	85.7 g	14.3 g
Molar mass	12 g mol^{-1}	1 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{85.7}{12.0} = 7.14$ mol	$n = \frac{14.3}{1.0} = 14.3$ mol
Divide all by the smallest amount	$\frac{7.14}{7.14} = 1$	$\frac{14.3}{7.14} = 2$

∴ empirical formula is CH_2

b Molar mass of a CH_2 unit (empirical formula) = 12 + 2 = 14 g mol^{-1}
Molar mass of the compound (molecular formula) = between 40 and 50 g mol^{-1}
∴ Number of CH_2 units in one molecule = 3
∴ Molecular formula is C_3H_6

c Alkene

24 **a**

	C	**O**	**H**
Mass	40.0 g	6.7 g	9.1 g
Molar mass	$n = \frac{10.0}{12.0} = 3.33$ mol	$n = \frac{6.7}{1.0} = 6.7$ mol	$n = \frac{53.3}{16.0} = 3.33$ mol
Simplest mole ratio	$\frac{3.33}{3.33} = 1$	$\frac{6.7}{3.33} = 2$	$\frac{3.33}{3.33} = 1$

∴ EF = CH_2O

b $M(\text{compound}) = \frac{m}{n} = \frac{15.0}{0.250} = 60.0 \text{ g mol}^{-1}$

$M(\text{EF}) = 12.0 + (2 \times 1.0) + 16.0 = 30.0 \text{ g mol}^{-1}$

number of empirical formula units $= \frac{\text{molar mass}}{\text{molar mass of empirical formula unit}}$

$\frac{60.0}{30.0} = 2$

Molecular formula = 2 × CH_2O = $C_2H_4O_2$

CONNECTING THE MAIN IDEAS

25 **a** Single covalent bond
  b Hydroxyl group
  c Methyl group
  d Carboxyl group

26 Student poster

27 **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.

28 **a** Polyunsaturated: contains many double bonds; mono-unsaturated: contains one double bond; saturated: contains only single bonds between carbon atoms.
  b Individual student response required.

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Chapter 10 Polymers

Section 10.1 Types of polymers

10.1 Key questions

1 **a** CH_2
  b CH_2
  c CH
  d C_2H_3Cl

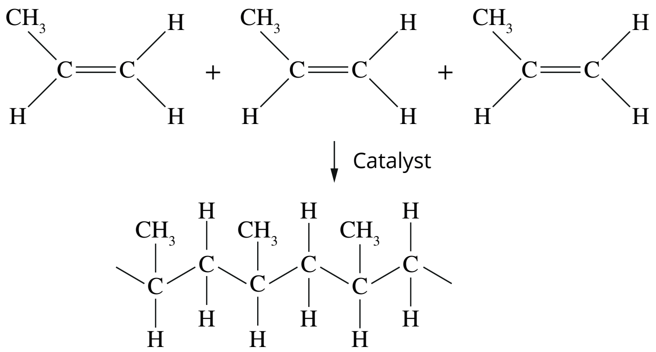
2 **a** Monomer formula C_3H_6, polymer formula has three monomers: C_9H_{18}
  b Monomer formula C_2H_3Cl, polymer formal has 6 monomers: $C_{12}H_{18}Cl_6$
  c Monomer formula C_2H_4, polymer formal has 65 monomers: $C_{130}H_{260}$

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

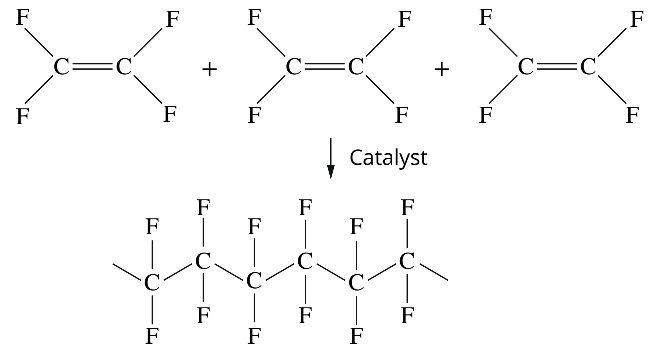
4 Polymers usually consist of thousands to millions of monomer repeating units, so this would not be considered a polymer.

5 **a** High-density polyethene (HDPE) is made of relatively unbranched chains of polyethene, which can pack more closely together 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

6 **a**



b



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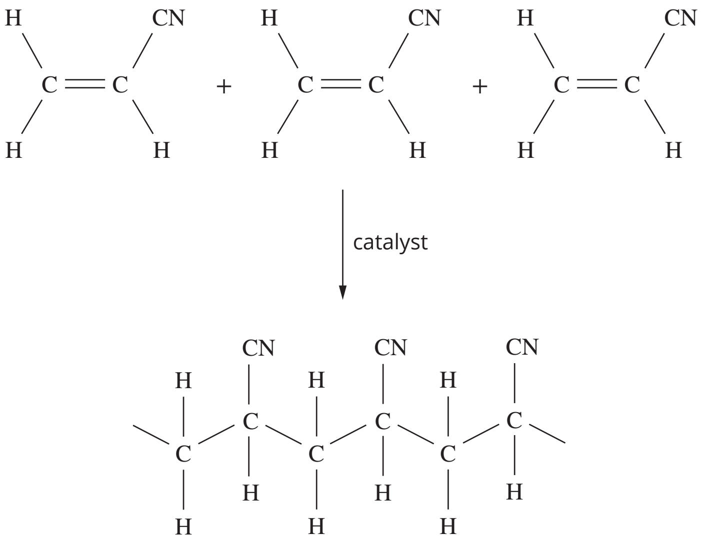
c



Section 10.2 Designing polymers for a purpose

10.2 Key questions

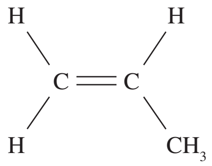
1 **a** Weak dispersion forces
   b Dipole–dipole attractions
   c Weak dispersion forces
   d Weak dispersion forces
   e Dipole–dipole attractions

2 Monomers from left to right: ethene, 1,1,2,2-tetrafluoroethene, ethene, 1,1,2,2-tetrafluoroethene

3 –H (in polyethene)
   –F (in Teflon)
   –Cl (in polyvinyl chloride)
   $–C_6H_5$ (in polystyrene)
   $–NC_{12}H_8$ (in polyvinyl carbazole)

4 **a** Increased crystallinity
   b Increased crystallinity
   c Decreased crystallinity

5 **a** Isotactic
   b



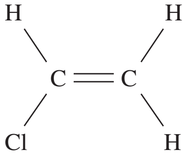
   c 3 units are shown.
   d This section is likely to be part of a crystalline region as the small side groups and isotactic arrangement will allow for close packing and strong interchain bonding.

6 **a** Increase flexibility
   b Improved heat insulating properties or mechanical properties
   c Increased crystallinity resulting in a harder, more brittle polymer

CHAPTER 10 REVIEW

10.1 TYPES OF POLYMERS

1 **a**



b 6

c Dipole–dipole bonds

2 **a** Monomers are small molecules that are able to react to form long chains of repeating units, called polymers. Often contains a carbon–carbon double bond.

b When the material is heated, the bonds between molecules are broken and the molecules become free to move.

c When the material 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 One or more covalent bonds that connect neighbouring polymer chains.

e A chemical added to plastics to make them more flexible.

3 **a** False

b False

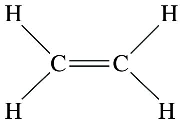
c False

d True

4 A, C, and D can act as monomers in addition polymerisation; B, propane, cannot as it does not contain a double bond.

5 **a** The ethene molecule has a carbon-to-carbon double bond.

b



Ethene has this structural formula because it ensures each atom has a stable outer-electron shell, and this arrangement of atoms gives the minimum repulsion between electron pairs.

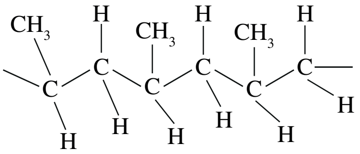
c Ethane cannot undergo addition polymerisation because it is a saturated compound.

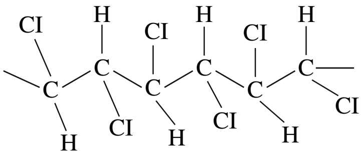
6 **a** Several thousand ethene monomers react to make one molecule of polyethene.

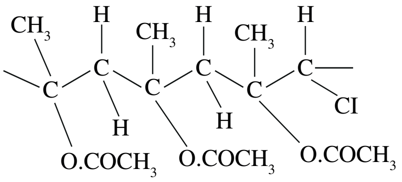
b Ethene ($CH_2=CH_2$) is unsaturated because it contains a double bond. Polyethene is saturated because it contains only single bonds between carbon atoms.

c The empirical formula is the simplest whole-number ratio of elements in a compound. The empirical formula of ethene ($CH_2=CH_2$) is CH_2.

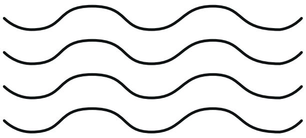
7







8 a i



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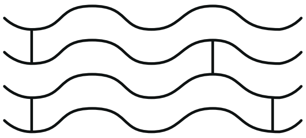

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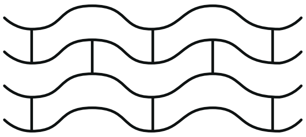
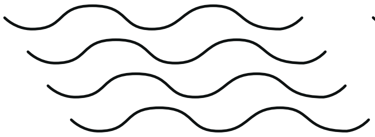
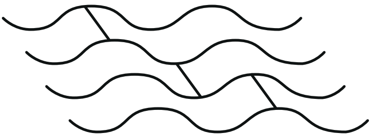


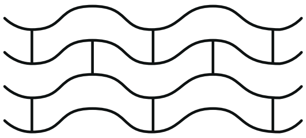
b



polymer easily stretched



Some stretching but polymer tends to return to its original shape.



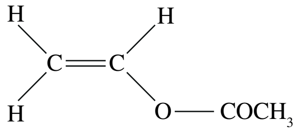
no stretching possible

9 An elastomer has occasional crosslinking that prevent the chains from slipping past each other too far and pull the chains back to their original positions when the force is removed. See Figure 10.1.15.

10 Thermosetting polymers have much stronger inter-chain bonds than thermoplastic polymers.

11 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.

12



10.2 DESIGNING POLYMERS FOR A PURPOSE

13 Distinguishing between thermoplastic and thermosetting polymers is important. The former get soft when heated, whereas the latter remain rigid and will char if heating continues.

- **a** Thermosetting
- **b** Thermoplastic
- **c** Thermosetting
- **d** Thermoplastic
- **e** Thermoplastic

14 As the polymer chains are held further apart by the plasticisers, it makes it easier for the polymer chains to slide over each other. As a result, these materials have a lower softening temperature, and are more malleable and soft.

15 Polymers without extensive branching and regular chain structures are most likely to form crystalline regions.

16 a PVC is flexible, non-conducting, strong, has a high melting point and low flammability.

b PVC has no free charged particles to conduct electricity. It has strong inter-chain bonding due to dipoles, which increase its strength and raise its melting point. PVC contains strong bonds that do not burn easily and it produces chemicals that extinguish flames when burned. PVC is plasticised to make it flexible.

17 Ultra-high molecular weight polyethene has very long polymer chains that cause stronger inter-chain forces than for HDPE.

18 Left: C_8H_8; middle: C_3H_3N; right: C_4H_6

19 a Thermoplastic with high melting temperature

b Lower density/more transparent polymer

20 Thermoplastic

CONNECTING THE MAIN IDEAS

21 **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.

22 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.

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Chapter 11 Research investigation

Section 11.1 Planning a research investigation

11.1 Key questions

1 **a** Evaluate
  b Compare
  c Contrast
  d Justify

2 D. The nuclear accident in Fukushima was due to radioactive material being released from a nuclear power plant.

3 **a** Secondary
  b Primary
  c Primary
  d Secondary

4 A and C. The company and the research scientist are more likely to be biased because they have vested interests in the outcomes. For example, if a research scientist has received financial support from a pharmaceutical company, there is the possibility that the scientist might try to influence the findings in favour of the company.

5 B. Repeating experiments and presenting results support the *reliability* of an experiment.

6 A. In APA style, the in-text citation should list the first author's last name and the year of publication.

Section 11.2 Communicating scientific ideas

11.2 Key questions

1 C. It is best to simplify or explain difficult words. 'Anthropogenic' describes pollutants originating from human activity.

2

Model	**Advantages**	**Limitations**
a Space-filling models	Useful for representing the sharing of space between 2 atoms where a pair of shared electron orbitals overlap. Useful for representing the overall 3D structure of molecules.	Sizes of atoms may not be accurate.
b Ball-and-stick models	Useful for representing the three-dimensional structure of molecules. Show each covalent bond.	Covalent bonds between atoms are often shown as 'lines' that do not represent actual bond length.
c Structural formula	Easy to draw.	Molecules are actually three-dimensional.

3 A. The first C is part of a methyl group, CH_3, the second C has 2 methyl groups attached to it, i.e. $C(CH_3)_2$, the third C in the chain has 2 H atoms attached to it, the fourth C has both a H atom and a methyl group attached, $CH(CH_3)$, and the final carbon is part of a methyl group.

4 **a** Economic
  b Social
  c Social
  d Environmental

5 B. The symbol of the unit gram is g.

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Section 11.3 Writing a scientific report

11.3 Key questions

1 B. Scientific writing should not use biased or absolute language.

2 D. First-person narrative uses the pronoun 'I'.

3 The statement 'Many repeats of the procedure were conducted' is unquantified. 'Thirty repeats of the procedure were conducted' is better because the number of trials is quantified.

4 B. A scientific article should always be written using objective language. A concluding paragraph must summarise the information presented in the article and connect it with the title. It should also include limitations, possible applications of the research and potential future research.

CHAPTER 11 REVIEW

11.1 PLANNING A RESEARCH INVESTIGATION

1 C

2 **a** Reflect
  b Create
  c Analyse
  d Investigate
  e Apply
  f Identify
  g Describe

3 C. All of these sources are secondary sources.

11.2 COMMUNICATING SCIENTIFIC IDEAS

4 C

5 C. The charge on the chloride ion is –1, therefore the charge on the copper ion is +2. Transition metals often can form more than one ion. To distinguish Cu^{2+} from Cu^{+}, Roman numerals are needed, thus copper(II).

6 C. *Environmental* factors are not mentioned in this excerpt. Medical implications fall into the social category, and costs and availability to certain groups and not others are economical and ethical factors.

11.3 WRITING A SCIENTIFIC REPORT

7 **a** Accuracy
  b Critical thinking
  c Precision
  d Validity of data
  e Reliability of data
  f Authority of data

8 **a** B. The article (advertisement) uses emotive language to imply that butter is healthier than margarine.
  b What's in it for you? It tastes better, naturally. Dairy Good. Good food. Good taste.

CONNECTING THE MAIN IDEAS

9 **a** **i** Social
    ii Bombarding
  b 'Through an electron microscope, Stevens' material looks like a bed of nails, each nail one thousandth the width of a human hair and one thousandth the length of a grain of rice.'

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c Possible answers:

1 This article is written in a scientific way because:
  - Evidence is provided to support arguments; e.g. ‘After one week there was a six-fold increase …’
  - Credible sources are cited, e.g. Molly Stevens and her team from Imperial College, London.

2 This article is written in a non-scientific way because:
  - Colloquial language is occasionally used such as ‘But placing DNA inside cells is tricky’.

d B. Further details would enable the reader to easily find the original article to read further, or confirm the information summarised.

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Chapter 12 Properties of water

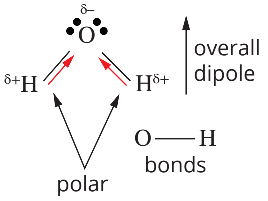
Section 12.1 Essential water

12.1 Key questions

1 **a** Physical properties that are unusual to water include:
- relatively high melting and boiling temperatures for its molecular size
- decrease in density on freezing
- high heat capacity
- high latent heat of fusion and evaporation for a substance of its molecular size

b The bond between H and O atoms in water is highly polar. As a result, hydrogen bonds exist between water molecules. Hydrogen bonds are stronger than other intermolecular bonds (although still weaker than the covalent intramolecular bonds) and so require more energy to break. Thus, water has a relatively high melting and boiling temperatures. Hydrogen bonding between water molecules in ice results in a very open arrangement of molecules so ice is less dense than liquid water.

2



3 Each water molecule has two hydrogen atoms and one oxygen atom. The oxygen atom has two pairs of non-bonding electrons, each of which can form one hydrogen bond.

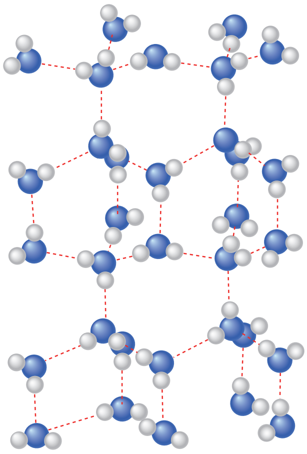
So the maximum number of water molecules with which one water molecule can form hydrogen bonds is four: up to two hydrogen bonds involving the two hydrogen atoms and up to two hydrogen bonds involving the two pairs of non-bonding electrons on the oxygen atom.

4 **a** H_2S, H_2Se, H_2Te, H_2Po, H_2O

b With the exception of water, the boiling points of the group 16 hydrides increase going down the group. This is due to increasing strength of dispersion forces as the molecules increase in mass. Water has a significantly higher boiling point than the other compounds due to the hydrogen bonds between its molecules.

5 **a** In ice, the distribution of hydrogen bonds in water is almost tetrahedral in shape. In order for these hydrogen bonds to form as the water solidifies, the water molecules must move into these positions, so occupying more volume than in liquid water.

b



c As it freezes, water expands, unlike most liquids. This is because of hydrogen bonding. Each water molecule is surrounded by four others in what is almost a crystal-type situation. (See Figure 12.1.7, which shows the variation in density of water with temperature.) Therefore, ice is less dense than liquid water, and it floats on liquid water. (For most liquids, the solid is denser than the liquid.) This is good news for fish, but not good news for travellers on the *Titanic*!

Section 12.2 Heat capacity

Worked example: Try yourself 12.2.1

CALCULATING THE AMOUNT OF ENERGY REQUIRED TO HEAT A SPECIFIED MASS OF A SUBSTANCE USING SPECIFIC HEAT CAPACITY

Calculate the heat energy, in kJ, needed to increase the temperature of 375 g of water by 45.0°C.

Thinking	Working
Find the specific heat capacity (*C*) of the substance from the data in Table 12.2.1.	The specific heat capacity of water is 4.18 J g^{-1} $°C^{-1}$.
To calculate the quantity of heat energy in joules, use the formula: $q = C \times m \times \Delta T$	$q = 4.18 \times 375 \times 45$ $= 7.05 \times 10^4$ J
Express the quantity of energy in kJ. Remember that to convert from J to kJ, you multiply 10^{-3}.	$q = 7.05 \times 10^4 \times 10^{-3}$ $= 70.5$ kJ

12.2 Key questions

1 D

2 Copper, iron, chlorofluorocarbon, concrete, glass, aluminium, wood, ethanediol, water

3 $q = C \times m \times \Delta T$

$= 0.48 \times 5.0 \times 12 = 29$ J

4 Convert the mass of water in kg to a mass in g (1000 g = 1 kg)

$1.5 \times 1000 = 1500$ g

$q = C \times m \times \Delta T$

$= 4.18 \times 1500 \times 15$

$= 94\,050$ J

The answer asks you to give your answer in kJ. To convert from joules to kilojoules divide by 1000.

94 050 J = 94 kJ

5 **a** 198 kJ

$q = C \times m \times \Delta T$

$= 4.18 \times 600 \times (100 - 21)$

$= 4.18 \times 600 \times 79$

$= 198\,132$ J

$= 198$ kJ

b $q = C \times m \times \Delta T$

$= 0.13 \times 600 \times 79$

$= 6162$ J

$= 6.2$ kJ

6 32°C

$q = C \times m \times \Delta T$°C

So rearranging the equation $\Delta T = \frac{q}{C \times m}$

$\Delta T = \frac{10\,000}{4.18 \times 250}$

$= \frac{10\,000}{1045}$

$= 9.57$°C

This is the change in temperature, so to work out the final temperature:

Temperature reached = 22°C + 9.57°C = 32°C

(Only to 2 sig figs as the energy is given as 10 kJ. When changing the unit you can't change the no. of sig figs.)

7 $C = 0.39\ \text{J g}^{-1}\,°\text{C}^{-1}$, therefore the unknown is copper.

$q = C \times m \times \Delta T$

Rearranging the equation $C = \frac{q}{m \times \Delta T}$

$C = \frac{9750}{500 \times 50} = 0.39\ \text{J g}^{-1}\,°\text{C}^{-1}$

From Table 12.2.1: $C = 0.39\ \text{J g}^{-1}\,°\text{C}^{-1}$ is copper.

Section 12.3 **Latent heat**

Worked example: Try yourself 12.3.1

CALCULATING THE HEAT ENERGY REQUIRED TO EVAPORATE A GIVEN MASS OF WATER AT ITS BOILING TEMPERATURE

Calculate the heat energy, in kJ, required to evaporate 75.0 g of water at 100°C.	
Thinking	**Working**
Determine the amount, in mol, of the substance using the formula: $n = \frac{m}{M}$	$n = \frac{75.0}{18.0}$ = 4.17 mol
Find the relevant latent heat value of the substance.	Water is being evaporated so the latent heat of vaporisation of water is required. $L = 44.0\ \text{kJ mol}^{-1}$
Calculate the heat energy, in kJ, using the formula: $q = n \times L$	$q = 4.17 \times 44.0$ = 70.5 kJ

Section 12.3 Key questions

1 At the melting point the flat region of the graph represents the substance *changing from a solid to a liquid*. The energy change is equal to *the latent heat of fusion*. At the boiling point the flat region of the graph represents the substance *changing from a liquid to a gas*. The energy change is equal to *the latent heat of vaporisation.*

2 Hydrogen bonding

3 The latent heat of fusion is the energy required to change the state of water from a solid to a liquid. This only requires the breaking of some of the hydrogen bonds between water molecules. The latent heat of vaporisation is much higher as the phase change from liquid to gas requires all the hydrogen bonds between water molecules to be broken.

4 110 kJ

Water is evaporating, so latent heat of fusion is required ($44.0\ \text{kJ mol}^{-1}$).

$q = n \times L$

$= 2.50 \times 44.0$

= 110 kJ

5 1 mL of water has a mass of 1 g so 100 mL has a mass of 100 g

$n = \frac{100}{18.0} = 5.56$ mol

$q = n \times L$

$= 5.56 \times 44.0$

= 245 kJ

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6 100 kJ

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

$= \frac{300}{18.0}$

= 16.7 mol

The ice is melting so the latent heat of fusion of water is required.

$q = n \times L$

$= 16.7 \times 6.0$

= 100 kJ

CHAPTER 12 **REVIEW**

12.1 ESSENTIAL WATER

1 Water is a *polar* molecule. Within a single molecule, hydrogen and oxygen atoms are held together by strong *covalent bonds*. Between different molecules, the most significant forces are *hydrogen bonds.*

It is the relatively *high* strength of the intermolecular forces that give water its unique properties of:

- relatively *high* boiling point, 100°C
- relatively *high* latent heat values 6.0 kJ mol^{-1} and 44.0 kJ mol^{-1}
- relatively *high* specific heat capacity 4.18 J g^{-1} °C^{-1}.

2 **a** Water molecule

b Hydrogen atom

c Hydrogen bond (and dispersion forces)

d Covalent bond

3 **a** Intermolecular forces are those between one molecule and other molecules. For water, these are hydrogen bonds. Intramolecular forces are those holding the atoms together within a molecule. For water, these are covalent bonds.

b Covalent bonds are stronger. Evidence for this is the high temperatures required to break the bonds between the oxygen and hydrogen atoms inside the water molecule and so decompose it into its constituent gases. Changing liquid water into gaseous water involves breaking hydrogen bonds to separate one molecule from another. The lower temperatures needed to do so indicate that hydrogen bonds are weaker.

4 Water has a significantly higher melting point than hydrogen sulfide due to the hydrogen bonds between water molecules. Hydrogen sulfide cannot form hydrogen bonds.

5 It is the high polarity of the water molecule that allows relatively strong hydrogen bonding to occur between molecules. As a consequence, a relatively large quantity of energy is required to break the hydrogen bonds between water molecules when water changes from a liquid to a gas. This gives water a high boiling point.

6 Water has a high latent heat of vaporisation. Water is effective as a cooler because it absorbs a relatively large amount of energy when it evaporates, giving it a high latent heat of vaporisation.

12.2 HEAT CAPACITY

7 D

$\Delta T = \frac{q}{C \times m}$

So the change in temperature of a substance (ΔT) is dependent upon the amount of heat energy transferred to the substance (q, option C), the mass of the substance (m, option A) and the heat capacity of the substance (C), which is related to the types of bonds present in the substance (option B).

8 85.7 kJ

250 mL of water has a mass of 250 g

$q = C \times m \times \Delta T$

$= 4.18 \times 250 \times (100 - 18)$

$= 4.18 \times 250 \times 82$

= 85 690 J

= 85.7 kJ

9 420 kJ

For water, 1 mL weighs 1 g, so 1.500 L of water (1500 mL) weighs 1500 g .

$q = C \times m \times \Delta T$

$= 4.18 \times 1500 \times (90 - 23)$

$= 4.18 \times 1500 \times 67$

= 420 090 J

= 420 kJ

10 9.6 kJ

1.0 kg = 1000 g

$q = C \times m \times \Delta T$

$= 2.4 \times 1000 \times 4.0$

= 9600 J

= 9.6 kJ

11 67.7°C

$q = C \times m \times \Delta T$

Rearrange for $\Delta T = \frac{q}{C \times m}$

$= \frac{12\,500}{0.39 \times 750} = 42.7°\text{C}$

Therefore final temperature = 25 + 42.7 = 67.7°C = 68°C

12 0.42 J g^{-1} $°C^{-1}$

Rearrange for $C = \frac{q}{m \times \Delta T}$

$= \frac{50}{15 \times 8.0}$

= 0.42 J g^{-1} $°C^{-1}$

12.3 LATENT HEAT

13

Definition	Property
a The temperature at which a liquid evaporates to form a gas	Boiling point
b The heat energy required to melt a solid to a liquid at its melting point	Latent heat of fusion
c The amount of heat energy required to increase a specific mass of a substance by a certain amount, e.g. 1 g by 1°C	Specific heat capacity
d The heat energy required to evaporate a liquid to a gas at its boiling point	Latent heat of vaporisation

14 B. Latent heat results in the change of state of a substance.

15 **a** The crystal lattice of ice is disrupted and molecules have greater freedom of movement. During this time, all solid ice is being converted to liquid water.

b The added energy is taken up in overcoming the hydrogen bonds between molecules, separating the molecules to form a gas.

16 611 kJ

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

$= \frac{250}{18.0}$

= 13.9 mol

The water is evaporating, so the latent heat of vaporisation is required.

$q = n \times L$

$= 13.9 \times 44.0$

= 611 kJ

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17 409 g

$q = n \times L$ can be rearranged to make the formula $n = \frac{q}{L}$

so $n = \frac{1000}{44.0}$

= 22.7 mol

$m = n \times M$

= 22.7 × 18.0

= 409 g

18 First calculate the amount, in mol, of water.

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

$= \frac{500}{18.0}$

= 27.8 mol

The heat energy required for melting the ice is then calculated:

$q = 6.0 \times 27.8$

= 167 kJ

The heat energy required for evaporating the water is calculated:

$q = 44.0 \times 27.8$

= 1223 kJ

Difference in heat energies = 1223 – 167 = 1056 kJ

19 Water's high latent heat of vaporisation means that living organisms need only release small amounts of water onto their skin to efficiently cool themselves. If the value were lower, the water would evaporate more easily, so more sweat would need to be produced to cool the body down. Living organisms would be at a greater risk of dehydration or overheating.

CONNECTING THE MAIN IDEAS

20 **a** The specific heat capacity of ethanol is almost half that of water. For the same amount of energy the change in temperature of ethanol will be almost twice that of water.

b Water has a higher specific heat capacity than ethanol due to the greater strength of its hydrogen bonds. Also for the same mass there are more molecules per gram of water than ethanol meaning that there will be a greater number of hydrogen bonds per gram of water. Heating 100 g of ethanol from 0°C to 20°C will require less energy than the energy needed to heat 100 g of water over the same temperature range.

21 0.80 $\text{J g}^{-1}\ {}^{\circ}\text{C}^{-1}$

$q = n \times L$

water is melting so need to use latent heat of fusion 6.0 kJ mol^{-1}

$n(\text{water}) = \frac{9.0}{18.0} = 0.50$ mol

$q = 0.50 \times 6.0$

= 3.0 kJ of energy absorbed by the ice

$q = C \times m \times \Delta T$

Rearrange for C

$C = \frac{q}{m \times \Delta T}$

$\Delta T = 150 - 0$

= 150 (not all of the ice has melted so the final temperature of the metal rod must still be at the melting point of water, 0°C)

$C = \frac{3000}{25 \times 150}$

= 0.80 $\text{J g}^{-1}\ {}^{\circ}\text{C}^{-1}$

Chapter 13 Water, the universal solvent

Section 13.1 Water as a solvent

13.1 Key questions

1 **a** Water – solvent
  b Wine – solution
  c Ethanol – solute
  d Sugar – solute

2 They all have water as the solvent.

3 **a** Ammonia (NH_3^-) – soluble
  b Oxygen gas (O_2) – likely to be insoluble
  c Hydrogen chloride (HCl) – soluble
  d Methanol (CH_3OH) – soluble
  e Methane (CH_4) – likely to be insoluble
  f Hydrogen fluoride (HF) – soluble
  g Carbon dioxide (CO_2) – likely to be insoluble

4 The solute–solvent forces are stronger than the solute–solute and solvent–solvent forces.

5

Shape	**Surface area**
Dissolution	The process by which a substance is dissolved in a solvent.
Solution	A liquid mixture in which a solute is dispersed throughout a solvent.
Solvent	The major component of a solution.
Solute	The minor component of a solution.
Soluble	Capable of dissolving
Insoluble	Incapable of dissolving

6 CH_4 (methane). Methane is a non-polar molecule. According to the principle of 'like dissolves like, a non-polar substance will not dissolve in a polar solvent such as water.

7 When sugar dissolves in water, (1) the solute–solute bonds between sugar molecules must be broken, (2) the solvent–solvent bonds between water molecules must also be broken and (3) new solute–solvent bonds form.

Section 13.2 Water as a solvent of molecular substances

13.2 Key questions

1 B, methanol, CH_3OH, and E, ethylene glycol, $HOCH_2CH_2OH$

An –OH group is polar and can form hydrogen bonds with water. As a consequence, methanol and ethylene glycol dissolve readily in water.

Hydrogen chloride contains polar covalent molecules; however, its molecules are not able to form hydrogen bonds with water molecules.

Ammonium sulfate and sodium hydroxide are ionic compounds containing polyatomic ions.

Diamond has a covalent lattice structure and is insoluble in water.

Hydrogen gas consists of non-polar molecules and will not dissolve in water.

2 $CH_3OH(l) \xrightarrow{H_2O(l)} CH_3OH(aq)$

$C_6H_{12}O_6(s) \xrightarrow{H_2O(l)} C_6H_{12}O_6(aq)$

3
- Hydrogen bonds between water molecules are broken.
- Hydrogen bonds between ethanol molecules are broken.
- New hydrogen bonds form between ethanol molecules and water molecules.

4 $HI(s) + H_2O(l) \rightarrow H_3O^+(aq) + I^-(aq)$

5 *Covalent bonds* within hydrogen chloride molecules are broken.

Hydrogen bonds between water molecules are broken. The HCl *ionises* and produces Cl^- and H^+ ions. *Covalent bonds* form between H^+ ions and water to produce *hydronium* ions. *Ion–dipole bonds* form between the Cl^- and H_3O^+ ions and polar water molecules.

Section 13.3 Water as a solvent of ionic compounds

Worked example: Try yourself 13.3.1

DETERMINING IF IONIC COMPOUNDS ARE SOLUBLE OR INSOLUBLE IN WATER

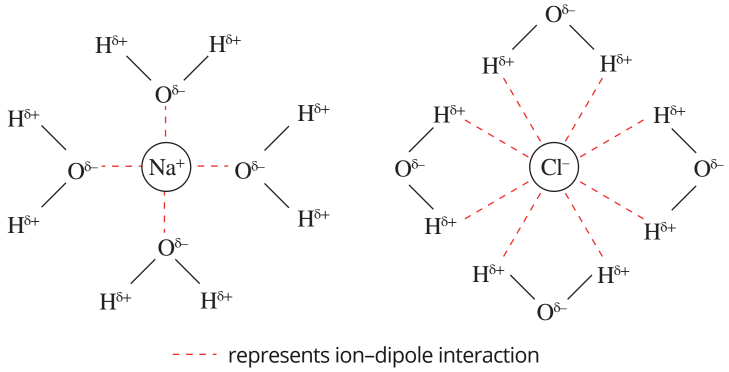
Is ammonium phosphate ($(NH_4)_3PO_4$) soluble or insoluble in water? You will need to refer to the solubility table to complete this question.	
Thinking	**Working**
Identify the ions that are present in the ionic compound.	Ammonium (NH_4^+) and phosphate (PO_4^{3-})
Check the solubility tables to see if compounds containing the cation are usually soluble or insoluble in water.	Compounds containing ammonium ions are soluble.
Check the solubility tables to see if compounds containing the anion are usually soluble or insoluble in water.	Ammonium phosphate is a noted water soluble phosphate ion containing compound.

13.3 Key questions

1 $NaNO_3(s) \xrightarrow{H_2O(l)} Na^+(aq) + NO_3^-(aq)$

$Ca(OH)_2(s) \xrightarrow{H_2O(l)} Ca^{2+}(aq) + 2OH^-(aq)$

2



The positive sodium ion attracts the partial negative charges on the oxygen atoms in the water molecule. The negative chloride ion attracts the partial positive charges on the hydrogen atoms in the water molecule.

3 A, B, D, E, H

4 A, C, D, E, F, H

5
- **a** Na^+/CO_3^{2-}
- **b** Ca^{2+}/NO_3^-
- **c** K^+/Br^-
- **d** Fe^{3+}/SO_4^{2-}
- **e** Cu^{2+}/Cl^-

6 **a** Nitrates are highly soluble in water. If found on Earth, they would dissolve in rainwater and wash into the oceans. Therefore, they are found only in areas of low rainfall.

b The high solubility of sodium, chloride and sulfate ions results in them dissolving and flowing into the world's oceans.

Section 13.4 Precipitation reactions

Worked example: Try yourself 13.4.1

PREDICTING THE PRODUCTS OF A PRECIPITATION REACTION

What precipitate, if any, will be produced when solutions of sodium sulfide, and copper(II) nitrate are added together? You will need to refer to the solubility tables (Tables 13.3.1 and 13.3.2) to complete this question.	
Thinking	**Working**
Identify which ions are produced by each of the ionic compounds in the mixture.	Ions produced are 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 13.4.2

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.	$CuSO_4(aq) + NaOH(aq) \rightarrow Cu(OH)_2(s)$
Add to the equation above the formula of the other compound formed in the reaction.	$CuSO_4(aq) + NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$
Balance the equation.	$CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$
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.

13.4 Key questions

1 **a** A precipitate of silver carbonate

b A precipitate of lead(II) hydroxide

c A precipitate of magnesium sulfide

d No precipitate

Silver, lead and magnesium ions are not found in the solubility table. However, the anions that each of these are combined with in these questions all form compounds that are usually insoluble. It is also worth noting that compounds containing sodium and nitrate ions are usually soluble.

The iron(II) ion is also not found in the solubility table. However, iron(II) does not appear in the 'exceptions' column of the table for either nitrates or sulfates, both of which generally form soluble compounds.

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2 Chemicals containing the ions Na^+, K+, NH_4^+ and NO_3^- almost never form a precipitate.

a i magnesium sulfide
ii silver chloride
iii aluminium hydroxide
iv magnesium hydroxide

b i $K_2S(aq) + MgCl_2(aq) \rightarrow MgS(s) + 2KCl(aq)$
ii $CuCl_2(aq) + 2AgNO_3(aq) \rightarrow 2AgCl(s) + Cu(NO_3)_2(aq)$
iii $AlCl_3(aq) + 3KOH(aq) \rightarrow Al(OH)_3(s) + 3KCl(aq)$
iv $MgSO_4(aq) + 2NaOH(aq) \rightarrow Mg(OH)_2(s) + Na_2SO_4(aq)$

3 a i $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
ii $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
b i $CuSO_4(aq) + Na_2CO_3(aq) \rightarrow CuCO_3(s) + Na_2SO_4(aq)$
ii $Cu^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CuCO_3(s)$
c i $(NH_4)_2SO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(aq) + 2NH_4Cl(aq)$
ii $SO_4^{2-}(aq) + Ba^{2+}(aq) \rightarrow BaSO_4(aq)$
d i $K_2S(aq) + Pb(NO_3)_2(aq) \rightarrow PbS(s) + 2KNO_3(aq)$
ii $S^{2-}(aq) + Pb^{2+}(aq) \rightarrow PbS(s)$
e i $3CaCl_2(aq) + 2Na_3PO_4(aq) \rightarrow 6NaCl(aq) + Ca_3(PO_4)_2(s)$
ii $3Ca^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$
f i $2NaOH(aq) + Pb(NO_3)_2(aq) \rightarrow 2NaNO_3(aq) + Pb(OH)_2(s)$
ii $2OH^-(aq) + Pb^{2+}(aq) \rightarrow Pb(OH)_2(s)$

4 a Na^+/NO_3^-
b Na^+/SO_4^{2-}
c NH_4^+/Cl^-
d K^+/NO_3^-
e Na^+/Cl^-
f Na^+/NO_3^-

CHAPTER 13 REVIEW

13.1 WATER AS A SOLVENT

1 Dispersion forces, hydrogen bonds, covalent bonds.

2 A solution is most likely to form when the polarity of bonding of the solute is similar to that of the solvent. The bonds formed between solute and solvent are then similar to those that existed between solute particles and between solvent particles. Water, being polar, is therefore a good solvent for ionic and polar substances.

3 Using the 'like dissolves like' rule, only polar substances will dissolve in water. Nitrogen is a non-polar molecule so will not dissolve well. Ethene is a non-polar hydrocarbon that will not be expected to dissolve in water. Ethanol, however, is a small polar molecule that can be expected to dissolve in water.

Full explanation:

- Nitrogen, N_2, is a covalent molecular substance. It is composed of non-polar molecules, so nitrogen would not be expected to dissolve well in water.
- Ethane, C_2H_6, is also a molecular compound. Ethane molecules contain slightly polar covalent bonds that are distributed symmetrically across the molecule. This makes the entire molecule non-polar, so ethane would be expected to not dissolve in water.
- Ethanol, C_2H_5O, is a molecular compound that dissolves in water by forming hydrogen bonds with water molecules.

4 Carbon is slightly more electronegative than hydrogen, so each C–H bond is slightly polar. However, the resulting partial charges are distributed symmetrically across the octane molecule, making the molecule non-polar overall.

The energy released in the formation of solute–solvent bonds is not enough to overcome the intermolecular bonds between solute molecules and the intermolecular bonds between solvent molecules, so octane will not dissolve in water.

5 **a** Biological context

Water dissolves wastes in the body so they can be safely removed from cells and the body.

Water dissolves glucose and other nutrients so they can be transported around the body for use in energy production.

Water dissolves nutrients so they can be transported around plants.

b Domestic context

Aqueous solutions are the basis for many drinks including tea, coffee, cordial, wine and soft drinks.

Most paints used for residential purposes are water-based, meaning they contain solutes dissolved in water.

Much cleaning in the home relies on water being able to dissolve detergents.

c Industrial context

Farmers add water-soluble fertilisers to increase crop growth.

Many medicines are water-soluble.

Water is used as a solvent in the production of paints, foods, paper and other products.

13.2 WATER AS A SOLVENT OF MOLECULAR SUBSTANCES

6 $C_6H_{12}O_6$ and C_3H_7OH are polar molecules. They contain the polar –OH group and so are able to form hydrogen bonds with water.

HI and HNO_3 contain polar molecules but they are unable to form hydrogen bonds with water. They dissolve by ionising.

I_2, CH_4 and C_2H_4 are non-polar covalent molecules. They do not dissolve well in polar water.

7 Y, Z, X

As the CH_3OH molecule is polar, it will dissolve readily in polar solvents. Because CH_3OH dissolves in Y, it can be ascertained that Y is a polar solvent. The non-polar nature of a CH_4 molecule means CH_4 will dissolve readily in non-polar solvents. It can be concluded that X is a non-polar solvent because CH_4 completely dissolves in it. CH_3OH molecules and CH_4 molecules partially dissolve in Z, which indicates it is more polar than X. No information is given about methanol dissolving in Z, although it can be implied.

8 DDT is most likely non-polar because it is soluble in fats, which are non-polar, and insoluble in polar water.

13.3 WATER AS A SOLVENT OF IONIC COMPOUNDS

9 **a** Dissociation

b **i** $Cu^{2+}(aq)$, $NO_3^-(aq)$

ii $Zn^{2+}(aq)$, $SO_4^{2-}(aq)$

iii $NH_4^+(aq)$, $PO_4^{3-}(aq)$

10 **a** K^+/CO_3^{2-}

b Pb^{2+}/NO_3^-

c Na^+/OH^-

d Na^+/SO_4^{2-}

e Mg^{2+}/Cl^-

f Zn^{2+}/NO_3^-

g K^+/S^{2-}

h Fe^{3+}/NO_3^-

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

12 Hydrated hydronium ions and chloride ions. Hydrogen chloride is a polar molecule that cannot form hydrogen bonds. When added to water, the molecule ionises to form a hydronium ion, H_3O^+, and a chloride ion, Cl^-. These ions become hydrated by ion–dipole attractions. There will also be water molecules and a very, very few HCl molecules.

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13 Magnesium ions are cations; they have a positive charge. This means the negatively charged non-bonding electron pairs on the oxygen atoms in the water molecules are attracted to them.

The water molecules arrange around the magnesium ion with their oxygen atoms, rather than the hydrogen atoms, closest to the chloride ion.

Chloride ions are anions; they have a negative charge. This means the hydrogen atoms in the water molecules, which have a partial positive charge, are attracted to them. The water molecules arrange around the chloride ion with their hydrogen atoms, rather than the oxygen atoms, closest to the chloride ion.

14 There are a number of possible answers to this question. Use the information in Tables 13.3.1 and 13.3.2. For example, Na_2CO_3, Li_2CO_3 and K_2CO_3 are soluble, whereas $CaCO_3$, $MgCO_3$ and Ag_2CO_3 are insoluble.

15 There are a number of possible answers to this question. Use the information in Tables 13.3.1 and 13.3.2. For example, Na_2SO_4, K_2SO_4 and $(NH_4)_2SO_4$ are soluble, whereas $CaSO_4$, $BaSO_4$ and $PbSO_4$ are insoluble.

16 Potassium ions and bromide ions are held in the ionic lattice by ionic bonds that are based on electrostatic attraction. These ionic bonds in the solute break when it dissolves in water.

Water, the solvent, has hydrogen bonds between water molecules.

When potassium bromide is added to water, the hydrogen atoms of the water molecules are attracted to the negative bromide ions, and the oxygen atoms of the water molecules are attracted to the positive potassium ions.

Ion–dipole bonds form between the ions and water molecules and the surface ions are pulled into solution. Gradually the ionic lattice dissociates and a solution is formed.

13.4 PRECIPITATION REACTIONS

17 **a** True
b True
c False
d False
e True
f False

18

	NaOH	**KBr**	**NaI**	**MgSO4**	**$BaCl_2$**
$Pb(NO_3)_2$	$Pb(OH)_2$	$PbBr_2$	PbI_2	$PbSO_4$	$PbCl_2$
KI					BaI_2
$CaCl_2$	$Ca(OH)_2$			$CaSO_4$	
Na_2CO_3				$MgCO_3$	$BaCO_3$
Na_2S				MgS	BaS

19 According to Tables 13.3.1 and 13.3.2, Fe^{2+} would precipitate in the presence of PO_4^{3-} and S^{2-} to produce $Fe_3(PO_4)_2$ and FeS.

20 **a** barium sulfate
b none
c lead(II) sulfate
d None

21 **a** $NH_4Cl(aq) + AgNO_3(aq) \rightarrow NH_4NO_3(aq) + AgCl(s)$
$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
b $Fe\ Cl2(aq) + Na_2S(aq) \rightarrow FeS(s) + 2NaCl(aq)$
$Fe^{2+}(aq) + S^{2-}(aq) \rightarrow FeS(s)$
c $Fe(NO_3)3(aq) + 3KOH(aq) \rightarrow 3KNO_3(aq) + Fe(OH)_3(s)$
$Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$
d $CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$
$Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$
e $Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$
$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

22 **a** $CuSO_4(aq) + Na_2CO_3(aq) \rightarrow CuCO_3(s) + Na_2SO_4(aq)$
$Cu^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CuCO_3(s)$
Spectator ions: Na^+, SO_4^{2-}

b $AgNO_3(aq) + KCl(aq) \rightarrow AgCl(s) + KNO_3(aq)$
$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
Spectator ions: K^+, NO_3^-

c $Na_2S(aq) + Pb(NO_3)_2(aq) \rightarrow PbS(s) + 2NaNO_3(aq)$
$Pb^{2+}(aq) + S^{2-}(aq) \rightarrow PbS(s)$
Spectator ions: Na^+, NO_3^-

d $FeCl_3(aq) + 3NaOH(aq) \rightarrow Fe(OH)_3(s) + 3NaCl(aq)$
$Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$
Spectator ions: Na^+, Cl^-

e $Fe_2(SO_4)_3(aq) + 6KOH(aq) \rightarrow 2Fe(OH)_3(s) + 3K_2SO_4(aq)$
$Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$
Spectator ions: K^+, SO_4^{2-}

CONNECTING THE MAIN IDEAS

23

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	**d**
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

24 **a** Ammonia is a highly polar molecule and forms hydrogen bonds with water. It is therefore very soluble in water. Methane, however, is non-polar. Weak (dispersion) forces would occur between methane and water, and these are unable to disrupt the stronger hydrogen bonds between water molecules. Therefore, methane does not dissolve in water.

b Glucose dissolves in water because it has very polar –OH groups that can form hydrogen bonds with water molecules. Sodium chloride is ionic; hence, there are ion–dipole attraction between the ions and water. This is strong enough to overcome the attraction between the sodium ions and chloride ions in the solid NaCl lattice.

25 Adding aqueous silver nitrate to a sample of each solution would distinguish between the two solutions. The sample that contained NaCl would form a white precipitate of AgCl(s); the glucose solution would not form a precipitate. Note electrical conductivity could also be used as the solution of dissociated ions (NaCl) would conduct an electrical current but glucose is a neutral covalent compound so would not conduct an electrical current.

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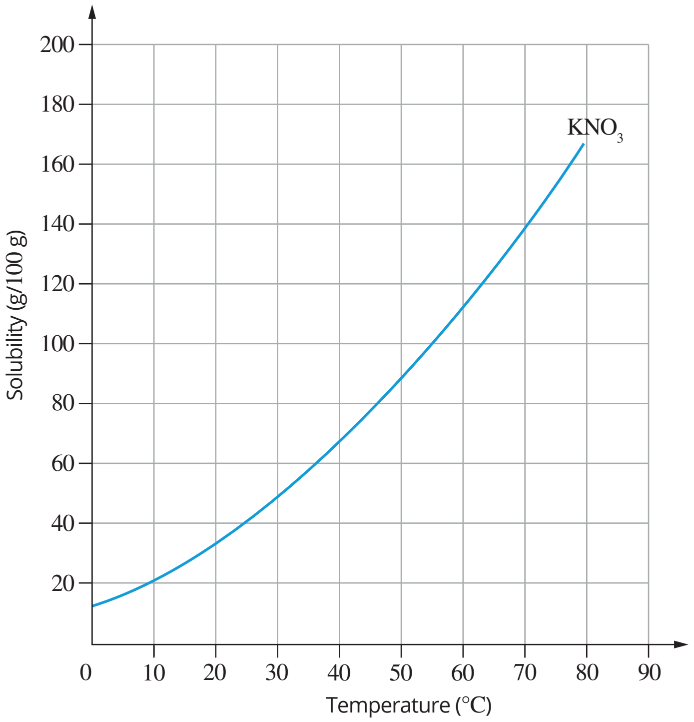
Chapter 14 Aqueous solutions

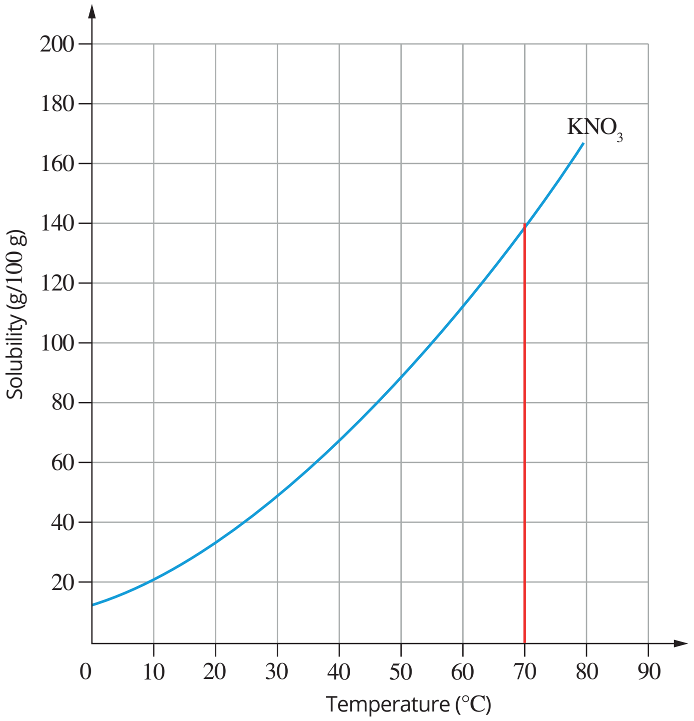
Section 14.1 Solubility

Worked example: Try yourself 14.1.1

SOLUBILITY CURVE CALCULATIONS

Use the solubility curve below to find how many grams of potassium nitrate (KNO_3) will dissolve in 100 g of water at 70°C.



Thinking	Working
Draw a vertical line from the required temperature on the horizontal axis to intersect with the appropriate solubility curve.	

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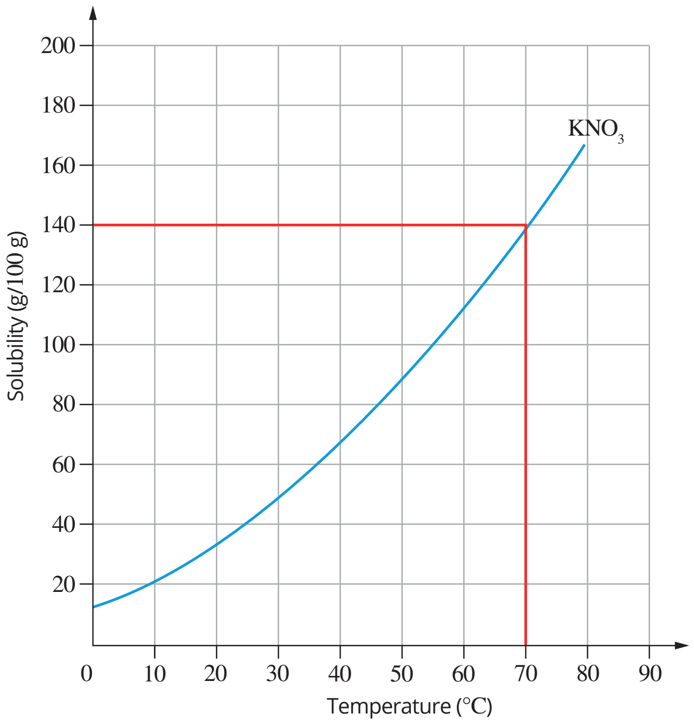
Draw a horizontal line from the intersection point of the vertical line drawn in the previous step. The point where this horizontal line intersects with the vertical axis will give the mass of dissolved substance.	
Read the solubility from the graph.	The horizontal line intersects with the vertical axis at 140 g. Therefore, the mass of that will dissolve in 100 g of water at 70°C is 140 g.

Worked example: Try yourself 14.1.2

SOLUBILITY CURVES

A 120 g sample of sodium nitrate ($NaNO_3$) is added to 300 mL of H_2O at 40°C. Use Figure 14.1.2 on page 337 to calculate how much more $NaNO_3$ needs to be added to make the solution saturated with $NaNO_3$ at 40°C.	
Thinking	**Working**
Use the solubility curve to find the mass of solute in a saturated solution of 100 g of H_2O at the required temperature.	Imagine a line drawn from 40°C on the horizontal axis to the solubility curve for $NaNO_3$ and find the corresponding value on the vertical axis. The value is 110 g.
Use the amount of solute that will dissolve in 100 g of H_2O to find the mass of solute to make a saturated solution in the mass of H_2O for this question.	The density of water is 1.0 g mL^{-1}, so 300 mL of water will weigh 300 g. So three times the mass of solute can dissolve in 300 g of water as in 100 g. $m(NaNO_3) = 3 \times 110$ g = 330 g
To find out how much extra solute you need to add, find the difference between the mass of solute needed to make a saturated solution and how much has already been added.	120 g of $NaNO_3$ has already been added to 300 g H_2O. So the extra mass of $NaNO_3$ needed: = 330 g – 120 g = 210 g

14.1 Key questions

1 **a** Saturated. According to the solubility table, this solution contains exactly the amount of sodium chloride necessary to saturate 100 g of water at 18°C.

b Unsaturated. A saturated solution of calcium carbonate in 100 g of water at 18°C contains 0.0013 g of solute. This solution contains less than 0.0013 g of dissolved calcium carbonate in 100 g of water, so the solution must be unsaturated.

   c Supersaturated. A saturated solution of sugar in 100 g of water at 18°C contains 200 g of dissolved sugar. Therefore a solution in 50 g contains 100 g of sugar. This solution contains more than 100 g of dissolved sugar, so the solution must be supersaturated.

2 C. A supersaturated solution is one that contains more solute than would normally be dissolved in a saturated solution. At 18°C, 35 g of sodium chloride would form a saturated solution in 100 g of water, so 17.5 g dissolves in 50 g of water. If 50 g of water at 18°C contained 20 g dissolved sodium chloride, it would therefore be supersaturated.

3 The solubility curve represents the maximum mass of solute that can be dissolved at a set temperature. The green dot represents a supersaturated solution as it is above the solubility curve. The orange dot represents a saturated solution as it lies on the solubility curve. The blue dot represents an unsaturated solution as it lies below the solubility curve.

4 **a** 40 g
   b 160 g
   c 20 g

5 **a** 48 g
   b 200 g
   c 70 g

   The mass that will dissolve in each case is the mass found from the solubility curve multiplied by 2 as the question asks for the amount in 200 g of solvent, not 100 g.

6 **a** Supersaturated. At 50°C 120 g of $NaNO_3$ will dissolve in 100 g of water. As the solution has 150 g of water the mass for a saturated solution will be $120 \times \frac{150}{100} = 180$ g, which is less than 200 g.

   b Unsaturated. At 20°C, 200 g of sucrose will dissolve in 100 g of water. As the solution has 200 g of water the mass for a saturated solution will be $200 \times \frac{200}{100} = 400$ g. The temperature of the solution is far greater than 20°C, therefore the solution could hold more solute.

   c Unsaturated. At 80°C, 58 g of $CuSO_4.5H_2O$ will dissolve in 100 g of water. As the solution has 20 g of water the mass for a soluble solution will be $58 \times \frac{20}{100} = 11.6$ g, which is greater than 8 g.

Section 14.2 Crystallisation

Worked example: Try yourself 14.2.1

CRYSTALLISATION

200 g of sucrose is dissolved in 100 g water at 20°C. What mass of sucrose crystals will form if the temperature is reduced to 10°C?

Thinking	Working
Identify the mass of solute dissolved in the original solution.	Mass of sucrose in original solution = 200 g
Find the maximum mass of solute hat will remain dissolved in 100 g water at the final temperature.	The solubility curve of sucrose shows that the maximum mass that will dissolve at 10°C is 190 g.
Calculate the mass of solute crystals that will form in the solution at the final temperature.	Mass of crystals formed = original mass – remaining = 200 – 190 = 10 g

14.2 Key questions

1 D. The maximum mass of oxygen that will dissolve in 1 kg water at 60°C is 0.023 g. The mass of oxygen that will leave the solution is 0.033 – 0.023 = 0.010 g.

2 A. The table shows the solubility of carbon dioxide in 1 kg water at 20°C and at 60°C. The solubility of carbon dioxide in 500 g water at these temperatures will be halved. At 20°C, 0.85 g of carbon dioxide will dissolve. At 60°C, 0.29 g of carbon dioxide will dissolve. The mass of carbon dioxide that will dissolve in 500 g water at 40°C would be expected to be between these two values.

3 B. At 55°C the maximum amount of KNO_3 that will dissolve in 100 g of water is 100 g. Therefore the mass of KNO_3 that will crystallise out of the solution at this temperature is 140 – 100 = 40 g.

4 20 g. The solubility curve for $CuSO_4.5H_2O$ shows that the maximum amount of solute that will dissolve in 100 g of solution at 60°C is 40 g. Therefore the mass of crystals will be 60 – 40 = 20 g.

5 55 g. The solubility curve for $NaNO_3$ shows that the maximum amount of solute that will dissolve in 100 g of water at 50°C is 120 g. Therefore the mass that may dissolve in 50 g of water will be 60 g. As this is greater than the amount present it will all remain dissolved.

6 Granite is a rock that is formed from magma cooling beneath the Earth's surface. This means the crystals are formed slowly. Crystals formed from slow cooling are larger than those formed from cooling quickly. Magma that cools above the surface forms basalt, the cooling process is faster, and therefore the crystals in basalt are smaller than those in granite.

Section 14.3 Concentration of solutions

Worked example: Try yourself 14.3.1

CALCULATING CONCENTRATION IN g L^{-1} (GRAMS PER LITRE)

What is the concentration, in g L^{-1}, of a solution containing 5.00 g of glucose in 250 mL of solution?	
Thinking	**Working**
Change the volume of solution so it is expressed in litres.	250 mL = $\frac{250}{1000}$ = 0.250 L
Calculate the concentration in g L^{-1}.	$c = \frac{\text{mass of solute (in g)}}{\text{volume of solution (in L)}}$ $= \frac{5.00}{0.250}$ = 20.0 g L^{-1}

Worked example: Try yourself 14.3.2

CALCULATING CONCENTRATION IN PARTS PER MILLION (ppm)

A sample of tap water was found to contain 0.0537 g of NaCl per 250.0 g of solution. Calculate the concentration of NaCl in parts per million (ppm).	
Thinking	**Working**
Calculate the mass of solute in mg. Remember: mass (in mg) = mass (in g) × 1000	Mass of solute (NaCl) in mg: = 0.0537 × 1000 = 53.7 mg
Calculate the mass of solution in kg. Remember: mass (in kg) = $\frac{\text{mass (in g)}}{1000}$	Mass of solution in kg = $\frac{250.0}{1000}$ = 0.2500 kg
Calculate the concentration of the solution in mg kg^{-1}. This is the same as concentration in ppm.	Concentration of NaCl in ppm $= \frac{\text{mass of solute (in g)}}{\text{mass of solution (in kg)}}$ $= \frac{53.7}{0.2500}$ = 215 mg kg^{-1} = 215 ppm

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Section 14.3 Key questions

1 D. Volume in litres = $\frac{60\text{ mL}}{1000}$ = 0.060 L. Concentration of sugar = $\frac{5.0\text{ g}}{0.060}$ = 83 g L^{-1}

2 **a** $c(CaCl_2) = \frac{m(CaCl_2)\text{ in mg}}{m(\text{solution})\text{ in kg}} = \frac{25}{5.0}$ = 5.0 ppm

b $c(Pb(NO_3)_2) = \frac{m(Pb(NO_3)_2)\text{ in mg}}{m(\text{solution})\text{ in kg}}$

$m(Pb(NO_3)_2)$ = 1.25 g × 1000 = 1250 mg

$m(\text{solution}) = \frac{2000\text{ g}}{1000}$ = 2.000 kg

$c(Pb(NO_3)_2) = \frac{1250}{2.000}$ = 625 ppm

c $c(MgSO_4) = \frac{m(MgSO_4)\text{ in mg}}{m(\text{solution})\text{ in kg}}$

$m(MgSO_4) = 4.0 \times 10^{-3}\text{ g} \times 1000$ = 4.0 mg

$m(\text{solution}) = \frac{150\text{ g}}{1000}$ = 0.150 kg

$c(MgSO_4) = \frac{4.0}{0.150}$ = 27 ppm

3 To find the % (m/v) the mass of each nutrient in grams is divided by the volume in millilitres and multiplied by 100.

Sugar = $\frac{35.0}{250} \times 100$ = 14.0% (m/v)

Fat = $\frac{7.5}{250} \times 100$ = 3.0% (m/v)

4 Parts per billion is the same as micrograms per kilogram.

Mass of ciguatera in µg = 15 mg × 1000 = 15 000 µg

The lethal concentration = $\frac{15\,000}{70}$ = 214 ppb

Section 14.4 **Molar concentration**

Worked example: Try yourself 14.4.1

CALCULATING MOLAR CONCENTRATIONS (MOLARITY)

Calculate the molar concentration of a solution that contains 0.24 mol of glucose dissolved in 500 mL of solution.	
Thinking	**Working**
Convert the given volume to litres.	$V(\text{glucose}) = \frac{500}{1000}$ = 0.500 L
Calculate the molar concentration using the formula: $c = \frac{n}{V}$	$c(\text{glucose}) = \frac{n}{V}$ $= \frac{0.24}{0.500}$ = 0.48 mol L^{-1} or 0.48 M

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Worked example: Try yourself 14.4.2

CALCULATING MOLARITY GIVEN THE MASS OF SOLUTE

Calculate the concentration, mol L^{-1}, of a solution that contains 4000 mg of ethanoic acid (CH_3COOH) dissolved in 100 mL of solution.

Thinking	**Working**
Convert the volume to litres.	$V(CH_3COOH) = \frac{100}{1000}$ $= 0.100$ L
Convert the mass to grams.	$m(CH_3COOH) = \frac{4000}{1000}$ $= 4.000$ g
Calculate the molar mass of the solute.	$M(CH_3COOH) = (12.0 \times 2) + (1.0 \times 4) + (16.0 \times 2)$ $= 600$ g mol^{-1}
Calculate the number of mol of solute using the formula: $n = \frac{m}{M}$	$n(CH_3COOH) = \frac{m}{M}$ $= \frac{4.000}{60.0}$ $= 0.0667$ mol
Calculate the molar concentration using the formula: $c = \frac{n}{V}$	$c(CH_3COOH) = \frac{n}{V}$ $= \frac{0.0667}{0.100}$ $= 0.0667$ mol L^{-1} or 0.667 M

Worked example: Try yourself 14.4.3

CALCULATING THE NUMBER OF MOLES OF SOLUTE IN A SOLUTION

Calculate the amount, in moles, of potassium permanganate ($KMnO_4$) in 100 mL of a 0.0250 M solution of the compound.

Thinking	**Working**
Convert the given volume to litres.	$V(KMnO_4) = \frac{100}{1000}$ $= 0.100$ L
Calculate the amount of compound, in moles, using the formula: $n = c \times V$	$n(KMnO_4) = c \times V$ $= 0.0250 \times 0.100$ $= 0.00250$ moL

14.4 Key questions

1 B. Convert the mass of H_2O_2 to moles:

$M(H_2O_2) = (1.0 \times 2) + (16.0 \times 2) = 34.0$ g mol^{-1}

$n(H_2O_2) = \frac{m(H_2O_2)}{M(H_2O_2)} = \frac{5.00}{34.0} = 0.147$ mol

Volume $= \frac{120 \text{ mL}}{1000} = 0.120$ L

Concentration $= \frac{0.147}{0.120} = 1.23$ M

2 **a** Volume $= \frac{25 \text{ mL}}{1000} = 0.025$ L

Concentration $= \frac{2 \times 10^{-3}}{0.025} = 0.08$ M

b Concentration $= \frac{1.23}{4.1} = 0.3$ M

c Concentration $= \frac{1.8 \times 10^{3}}{9.3 \times 10^{3}} = 0.19$ M

3 **a** Concentration $= \frac{2.0}{8.0} = 0.25$ M

b Volume $= \frac{500 \text{ mL}}{1000} = 0.5$ L, concentration $= \frac{0.25}{0.5} = 0.50$ M

c Volume $= \frac{200 \text{ mL}}{1000} = 0.2$ L, concentration $= \frac{0.0876}{0.2} = 0.438$ M

4 Convert the mass of $AgNO_3$ to moles:

$M(AgNO_3) = 107.9 + 14.0 + (16.0 \times 3) = 169.9 \text{ g mol}^{-1}$

$n(AgNO_3) = \frac{m(AgNO_3)}{M(AgNO_3)} = \frac{5.09}{169.9} = 0.0300$ mol

Volume $= \frac{250 \text{ mL}}{1000} = 0.250$ L, concentration $= \frac{0.0300}{0.250} = 0.120$ M

5 Convert the mass of $CaCl_2$ to moles:

$M(CaCl_2) = 40.1 + (35.5 \times 2) = 111.1 \text{ g mol}^{-1}$

$n(CaCl_2) = \frac{m(CaCl_2)}{M(CaCl_2)} = \frac{1.223}{111.1} = 1.101 \times 10^{-2}$ mol

Concentration $= \frac{1.101 \times 10^{-2}}{1.55} = 7.10 \times 10^{-3}$ M

6 **a** Amount $= c \times V = 0.1 \times 0.22 = 2.2 \times 10^{-2}$ mol

b Volume $= \frac{10 \text{ mL}}{1000} = 0.01$ L, amount $= c \times V = 0.64 \times 0.01 = 6.4 \times 10^{-3}$ mol

c Volume $= \frac{15.6 \text{ mL}}{1000} = 0.0156$ L, amount $= c \times V = 0.0150 \times 0.0156 = 2.34 \times 10^{-4}$ mol

d Volume $= \frac{1.5 \times 10^{-1} \text{ mL}}{1000} = 1.5 \times 10^{-4}$ L, amount $= c \times V = 5.2 \times 1.5 \times 10^{-1} = 0.78$ mol

Section 14.5 **Dilution**

Worked example: Try yourself 14.5.1

QUESTIONS INVOLVING DILUTION

Calculate the concentration of the solution formed when 95.0 mL of water is added to 5.00 mL of 0.500 M HCl.

Thinking	Working
Write down the value of c_1 and V_1. Note: c_1 and V_1 refer to the original solution, before water was added.	$c_1 = 0.500$ M $V_1 = 5.00$ mL
Write down the value of V_2. Note: V_2 is the total volume of the original solution plus the added water.	$V_2 = 5.00$ mL + 95.0 mL = 100.0 mL
Transpose the equation $c_1V_1 = c_2V_2$ to allow calculation of the concentration, c_2, of the new solution.	$c_1V_1 = c_2V_2$ $c_2 = \frac{c_1V_1}{V_2}$
Calculate the concentration of the diluted solution.	$c_2 = \frac{0.500 \times 5.00}{100}$ = 0.0250 M

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Worked example: Try yourself 14.5.2

CONCENTRATION UNIT CONVERSIONS

What is the concentration, in ppm, of a 0.0100 M solution of NaOH? Remember that concentration in ppm is the same as mg L^{-1}.	
Thinking	**Working**
Calculate the number of moles of solute in 1.00 L of the solution.	$n(\text{NaOH}) = c \times V$ $= 0.0100 \times 1.00$ $= 0.0100 \text{ mol}$
Calculate the mass, in grams, of solute in 1.00 L of the solution.	$M(\text{NaOH}) = 23.0 + 16.0 + 1.0$ $= 040.0 \text{ g mol}^{-1}$ $m(\text{NaOH}) = n \times M$ $= 0.0100 \times 40.0$ $= 0.400 \text{ g}$
Calculate the mass, in mg, of solute in 1.00 L of the solution.	$m(\text{NaOH}) = 0.400 \text{ g} \times 1000$ $= 400 \text{ mg}$
Express the concentration of the solute in ppm.	$c(\text{NaOH}) = 400 \text{ ppm}$

14.5 Key questions

1 **a** $V_2 = 10.0 \text{ mL} + 5.0 \text{ mL} = 15.0 \text{ mL}, c_2 = \frac{c_1V_1}{V_2} = \frac{1.2 \times 5.0}{15.0} = 0.4 \text{ M}$

b $V_2 = 1.0 \text{ L} + 3.0 \text{ L} = 4.0 \text{ L}, c_2 = \frac{c_1V_1}{V_2} = \frac{0.1 \times 3.0}{4.0} = 0.075 \text{ M}$

c $V_2 = 5.0 \text{ mL} + 95.0 \text{ mL} = 100.0 \text{ mL}, c_2 = \frac{c_1V_1}{V_2} = \frac{0.5 \times 5.0}{100.0} = 0.025 \text{ M}$

2 D. $V_2 = \frac{250 \text{ mL}}{1000} = 0.250 \text{ L}, V_1 = \frac{c_2V_2}{c_1} = \frac{0.30 \times 0.250}{10} = 0.0075 \text{ L} = 7.5 \times 10^{-3} \text{ L}$

3 Mass of ammonia in 1 L of original solution = 15 g

$M(\text{NH}_3) = 14 + (1.0 \times 3) = 17.0 \text{ g mol}^{-1}$

$n(\text{NH}_3) = \frac{m(\text{NH}_3)}{M(\text{NH}_3)} = \frac{15}{17} = 0.88 \text{ M}$

$c_1 = \frac{0.88}{1.0} = 0.88 \text{ M}, V_2 = 25 \text{ mL} + 250 \text{ mL} = 275 \text{ mL}$

$c_2 = \frac{c_1V_1}{V_2} = \frac{0.88 \times 25.0}{275.0} = 8.0 \times 10^{-2} \text{ M}$

4 In 1 L of solution there will be 170 g of NaOH.

$M(\text{NaOH}) = 23.0 + 1.0 + 16.0 = 40.0 \text{ g mol}^{-1}$

$n(\text{NaOH}) = \frac{m(\text{NaOH})}{M(\text{NaOH})} = \frac{170}{40.0} = 4.25 \text{ mol}, c(\text{NaOH}) = \frac{4.25}{1.0} = 4.25 \text{ mol L}^{-1}$

CHAPTER 14 **REVIEW**

14.1 SOLUBILITY

1 **a** 50 g
  b 10.5 g
  c 25 g

2 **a** $\frac{200\text{ g}}{100\text{ g}} \times 500\text{ g} = 1000\text{ g}$

b $\frac{40\text{ g}}{100\text{ g}} \times 50\text{ g} = 20\text{ g}$

3 $\frac{40\text{ g}}{25\text{ g}} \times 100\text{ g} = 160\text{ g}$ to be dissolved, therefore temperature must be 10°C.

14.2 CRYSTALLISATION

4 The maximum amount of potassium nitrate that will dissolve in 100 g water at 30°C is 50 g. Therefore the maximum amount that will dissolve in 200 g water is 100 g. There was initially 160 potassium nitrate dissolved, so the amount that will crystallise out is 160 – 100 = 60 g.

5 As the temperature increases, the solubility of most gases in water decreases. This means that less oxygen and other gases will be available to aquatic life.

6 0.5 mmol in 500 mL gives 1.0 mmol in 1 L. According to Figure 14.2.8, 23°C and 1.0 mmol L^{-1} corresponds to carbon monoxide, CO.

7 Table 14.2.1 shows that at 20°C, 1 kg water can dissolve 529 g ammonia.

The mass in 2 kg of water is $529\text{ g} \times \frac{2\text{ kg}}{1\text{ kg}} = 1058\text{ g}$.

The mass of ammonia that will leave the solution at 20°C is 1200 – 1058 = 142 g.

14.3 CONCENTRATION OF SOLUTIONS

8 **a** Concentration of lead (ppm) $= \frac{m(\text{Pb})\text{ in mg}}{m(\text{solution})\text{ in kg}} = \frac{12}{6} = 2\text{ ppm}$

b Percentage by mass $= \frac{m(\text{Pb})\text{ in mg}}{m(\text{solution})\text{ in kg}} \times 100$, mass of solution = 6 kg × 1 000 000 = 6 000 000 mg

$\%\ (\text{m/m}) = \frac{12}{6\,000\,000} \times 100 = 2 \times 10^{-4}\%$

9 $\%\ (\text{m/m}) = \frac{15}{300} \times 100 = 5.0\%$

10 Concentration of active ingredient (mg L^{-1}) $= \frac{\text{mass (in mg)}}{\text{volume of solution (in L)}}$

Volume of solution $= \frac{5.0\text{ mL}}{1000} = 0.0050\text{ L}$, concentration $= \frac{2.00}{0.0050} = 400\text{ mg L}^{-1}$

Concentration of active ingredient (% (m/v)) $= \frac{\text{mass (in g)}}{\text{volume (in mL)}} \times 100$

Mass of active ingredient $= \frac{2.00\text{ mg}}{1000} = 2.00 \times 10^{-3}\text{ g}$

$\%\ (\text{m/v}) = \frac{2.00 \times 10^{-3}}{5.0} \times 100 = 4.0 \times 10^{-2}$

14.4 MOLAR CONCENTRATION

11 $C = \frac{n}{V}$, volume $= \frac{160\text{ mL}}{1000} = 0.160\text{ L}$, $c(\text{NaBr}) = \frac{0.380}{0.160} = 2.38\text{ M}$

12 $c = \frac{n}{V}$, $M(\text{NaOH}) = 23.0 + 16.0 + 1.0 = 40.0\text{ g mol}^{-1}$,

$n(\text{NaOH}) = \frac{m(\text{NaOH})}{M(\text{NaOH})} = \frac{30}{40} = 0.75\text{ mol}$,

$c(\text{NaOH}) = \frac{0.75}{2.0} = 0.375\text{ M} = 0.38\text{ M}$

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13 $c = \frac{n}{V}$, in 1 L of solution there will be 42.6 g of solute,

$M(Na_2SO_4) = (23.0 \times 2) + 32.1 + (16.0 \times 4) = 142.1$ g mol^{-1}

$n(Na_2SO_4) = \frac{m(Na_2SO_4)}{M(Na_2SO_4)} = \frac{42.6}{142.1} = 0.300$ mol

$c(Na_2SO_4) = \frac{0.300}{1} = 0.300$ M

14 **a** Volume = $\frac{12 \text{ mL}}{1000} = 0.012$ L, $n = cV = 0.22 \times 0.012 = 2.6 \times 10^{-3}$ mol

b Volume = $\frac{150 \text{ mL}}{1000} = 0.150$ L, $n = cV = 0.0250 \times 0.150 = 3.75 \times 10^{-3}$ mol

c Concentration = $\frac{3.15 \text{ mL}}{1000} = 0.00315$ M, $n = cV = 0.00315 \times 7.2 = 2.3 \times 10^{-2}$ mol

15 **a** Volume = $\frac{100 \text{ mL}}{1000} = 0.100$ L

$n(NH_3) = cV = 1.2 \times 0.100 = 0.120$ mol
$M(NH_3) = 14.0 + (1.0 \times 3) = 17.0$ g mol^{-1}
$m(NH_3) = n(NH_3) \times M(NH_3)$
$= 0.120 \times 17.0 = 2.04$ g

b Volume = $\frac{20 \text{ mL}}{1000} = 0.020$ L

$n(AgNO_3) = cV = 0.50 \times 0.020 = 0.010$ mol
$M(AgNO_3) = 107.9 + 14.0 + (16.0 \times 3) = 169.9$ g mol^{-1}
$m(AgNO_3) = n(AgNO_3) \times M(AgNO_3) = 0.010 \times 169.9 = 1.7$ g

14.5 DILUTION

16 $c_2 = \frac{c_1V_1}{V_2} = \frac{5.00 \times 50.0}{250} = 1.00$ M

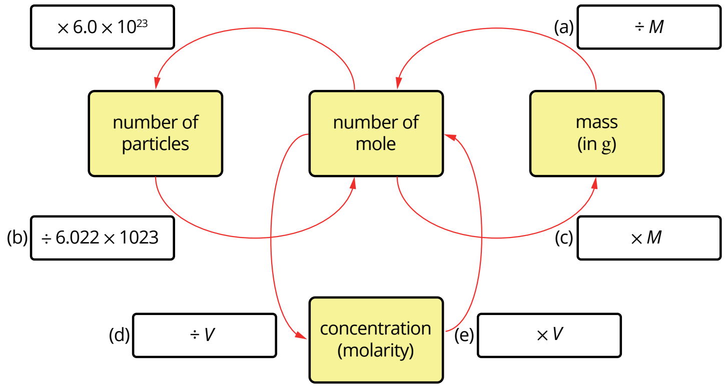
17 $V_2 = \frac{c_1V_1}{c_2} = \frac{4.0 \times 25}{1.6} = 62.5$ mL, so volume to be added = 62.5 – 25 = 37.5 mL

18 $n(CaCl_2)$ in the original solution = $c \times V$, volume = $\frac{1.5 \text{ mL}}{1000} = 0.00150$ L

$n(CaCl_2) = 0.0500 \times 0.00150 = 7.50 \times 10^{-5}$, and therefore 1.50×10^{-4} mol of Cl^- ions
$M(Cl^-) = 35.5$ g mol^{-1}
$m(Cl^-) = n(Cl^-) \times M(Cl^-) = 1.50 \times 10^{-4} \times 35.5 = 5.33 \times 10^{-3}$ g
Concentration of final solution (ppm) = $\frac{m(Cl^-)}{V(Cl^-) \text{ in L}}$
$m(Cl^-) = 5.33 \times 10^{-3}$ g $\times 1000 = 5.33$ mg
$c(Cl^-) = \frac{5.33}{10.0} = 0.533$ ppm

CONNECTING THE MAIN IDEAS

19



20 **a** The solubility curve for potassium nitrate shows that 140 g of solute will dissolve in 100 g of water at 70°C. The mass in 500 g of water is 140 g × $\frac{500}{100}$ = 700 g.

b The solubility curve for potassium nitrate shows that 140 g of solute will dissolve in 100 g of water at 70°C. The mass in 50 g of water is 140 g × $\frac{50}{100}$ = 70 g. The mass required to be added is 70 – 25 = 45 g.

c The solubility curve for potassium nitrate shows that 68 g of solute will dissolve in 100 g of water at 40°C. The mass in 50 g of water is 68 g × $\frac{50}{100}$ = 34 g. The mass that will remain undissolved is 100 – 34 = 66 g.

d $c = \frac{n}{V}$, $n(\text{KNO}_3) = \frac{m(\text{KNO}_3)}{M(\text{KNO}_3)}$

$M(\text{KNO}_3) = 39.1 + 14.0 + (3 \times 16.0) = 101.1\ \text{g mol}^{-1}$

$n(\text{KNO}_3) = \frac{34}{101.1} = 0.34\ \text{mol}$

$\text{Volume} = \frac{50\ \text{mL}}{1000} = 0.050\ \text{L}$

$c(\text{KNO}_3) = \frac{0.34}{0.050} = 6.7\ \text{M}$

21 **a** Figure 14.1.3 shows that 120 g of solute per 100 g water will make a saturated solution at 50°C. In 150 g of water, the mass of solute is 120 g × $\frac{150}{100}$ = 180 g.

b At 30°C only 100 g of solute will dissolve per 100 g of water. In 150 g of water, the mass of solute is 100 g × $\frac{150}{100}$ = 150 g. The mass of crystals that form will be = 180 – 150 = 30 g.

c $c = \frac{n}{V}$, $n(\text{NaNO}_3) = \frac{m(\text{NaNO}_3)}{M(\text{NanO}_3)}$

$M(\text{NaNO}_3) = 23.0 + 14.0 + (16.0 \times 3) = 85\ \text{g mol}^{-1}$

$n(\text{NaNO}_3) = \frac{30}{85} = 0.35\ \text{mol}$

$\text{Volume} = \frac{250\ \text{mL}}{1000} = 0.250\ \text{L}$

$c(\text{NaNO}_3) = \frac{0.35}{250} = 1.4\ \text{M}$

d $c_2 = \frac{c_1V_1}{V_2}$, $V_2 = 700 + 50 = 750\ \text{mL}$, $c_2 = \frac{1.4 \times 50}{750} = 9.4 \times 10^{-2}\ \text{M}$

In 1 L of solution there would be 0.094 mol of solute.
$m(\text{NaNO}_3) = n(\text{NaNO}_3) \times M(\text{NaNO}_3)$
$M(\text{NaNO}_3) = 23.0 + 14.0 + (16.0 \times 3) = 85\ \text{g mol}^{-1}$
$m(\text{NaNO}_3) = 0.094 \times 85.0 = 8.0\ \text{g}$
Concentration (in g L^{-1}) = $\frac{m\ (\text{in g})}{V\ (\text{in L})} = \frac{8.0}{1.0} = 8.0\ \text{g L}^{-1}$

22 Figure 14.1.3 shows that 20 g of solute will make a saturated solution at 20°C. The mass of solute in 1 L of solution is 20 g × $\frac{1000}{100}$ = 200 g.

$c = \frac{n}{V}$, $n(\text{CuSO}_4.5\text{H}_2\text{O}) = \frac{m(\text{CuSO}_4.5\text{H}_2\text{O})}{M(\text{CuSO}_4.5\text{H}_2\text{O})}$

$M(\text{CuSO}_4.5\text{H}_2\text{O}) = 63.5 + 32.1 + (16.0 \times 4) + 5 \times ((1.0 \times 2) + 16.0) = 249.6\ \text{g mol}^{-1}$

$n(\text{CuSO}_4.5\text{H}_2\text{O}) = \frac{200}{249.6} = 0.80\ \text{mol}$

$c(\text{CuSO}_4.5\text{H}_2\text{O}) = \frac{0.80}{1.0} = 0.80\ \text{M}$

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Chapter 15 Acid–base reactions in water

Section 15.1 Introducing acids and bases

15.1 Key questions

1 $HBr(g) + H_2O(l) \rightarrow H_3O^+(aq) + Br^-(aq)$

2 H_2SO_4/HSO_4^- and $H_2NO_3^+/HNO_3$

3 **a** H_2O
  b H_3O^+
  c CH_3NH_2

4 **a** NH_4^+
  b CH_3COOH
  c $H_2PO_4^-$
  d HCO_3^-

5 Brønsted–Lowry acid–base reactions are those involving the exchange of a proton (H^+ ion). The acid donates the proton to the base. In the reaction below, the HCl loses a proton to the base.

$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

The ionic equation provides a clearer way (by eliminating spectator ions) of noting the reaction between the H^+ and OH^- ions neutralising to form water.

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

6 Acting as an acid, whereby the reactant donates one proton:
  a $HCO_3^- + H_2O(l) \rightarrow CO_3^{2-}(aq) + H_3O^+(aq)$
  b $HPO_4^{2-} + H_2O(l) \rightarrow PO_4^{3-}(aq) + H_3O^+(aq)$
  c $HSO_4^- + H_2O(l) \rightarrow SO_4^-(aq) + H_3O^+(aq)$
  d $H_2O(l) + H_2O(l) \rightarrow OH^-(aq) + H_3O^+(aq)$

Acting as a base, whereby the reactant accepts one proton:
  a $HCO_3^- + H_2O(l) \rightarrow H_2CO_3(aq) + OH^-(aq)$
  b $HPO_4^{2-} + H_2O(l) \rightarrow H_2PO_4^-(aq) + OH^-(aq)$
  c $HSO_4^- + H_2O(l) \rightarrow H_2SO_4(aq) + OH^-(aq)$
  d $H_2O + H_2O(l) \rightarrow H_3O^+(aq) + OH^-(aq)$

Section 15.2 Strength of acids and bases

15.2 Key questions

1 **a** $HClO_4(aq) + H_2O(l) \rightarrow H_3O^+(aq) + ClO_4^-(aq)$
  b $HCN(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + CN^-(aq)$
  c $CH_3NH_2(aq) + H_2O(l) \leftrightarrow CH_3NH_3^+ (aq) + OH^-(aq)$

2 (1) $H_3AsO_4(aq) + H_2O(l) \rightarrow H_3O^+(aq) + H_2AsO_4^-(aq)$
  (2) $H_2AsO_4^-(aq) + H_2O(l) \rightarrow H_3O^+(aq) + HAsO_4^{2-}(aq)$
  (3) $HAsO_4^{2-}(aq) + H_2O(l) \rightarrow H_3O^+(aq) + AsO_4^{3-}(aq)$

3 A. A strong acid readily donates a proton to a water molecule to form the hydronium ion. The complete ionisation of the acid molecule is indicated by the use of a single arrow, →.

4 Stronger acids more readily ionise, forming ions in solution. As perchloric acid is stronger, more hydronium ions would be present in solution than in a solution of ethanoic acid, making it a better conductor of electricity.

Section 15.3 Acidity of solutions

Worked example: Try yourself 15.3.1

CALCULATING THE CONCENTRATION OF HYDRONIUM AND HYDROXIDE IONS IN AN AQUEOUS SOLUTION

For a 5.6 × 10^{-6} M HNO_3 solution at 25°C, calculate the concentration of: • H_3O^+ • OH^-	
Thinking	**Working**
Find the concentration of hydronium ions (H_3O^+).	HNO_3 is a strong acid, so it will ionise completely in solution. Each molecule of HNO_3 donates one proton to water to form one H_3O^+ ion: $HNO_3(aq) + H_2O(l) \rightarrow H_3O^+(aq) + NO_3^-(aq)$ Because HNO_3 is completely ionised in water, 5.6 × 10^{-6} M will produce a solution with a concentration of H_3O^+ ions of 5.6 × 10^{-6} M: i.e $[H_3O^+] = 5.6 \times 10^{-6}$ M
Use the expression for the ionisation constant of water to calculate the concentration of OH^- ions.	$K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$ M^2 $[OH^-] = \frac{1.00 \times 10^{-14}}{[H_3O^+]}$ Since $[H_3O^+] = 5.6 \times 10^{-6}$ M $[OH^-] = \frac{1.00 \times 10^{-14}}{5.6 \times 10^{-6}}$ $[OH^-] = 1.8 \times 10^{-9}$ M

Worked example: Try yourself 15.3.2

CALCULATING THE pH OF AN AQUEOUS SOLUTION FROM $[H_3O^+]$

What is the pH of a solution in which the concentration of $[H_3O^+]$ is 6 × 10^{-9} M? Express your answer to two significant figures.	
Thinking	**Working**
Note down the concentration of $[H_3O^+]$ ions in the solution.	$[H_3O^+] = 6 \times 10^{-9}$ M
Substitute the value of $[H_3O^+]$ into: $pH = -\log_{10}[H_3O^+]$ Use the logarithm function on your calculator to calculate the answer.	$pH = -\log_{10}[H_3O^+]$ $= -\log_{10}[6 \times 10^{-9}]$ (use your calculator) = 8.2

Worked example: Try yourself 15.3.3

CALCULATING pH IN A SOLUTION OF A BASE

What is the pH of a 0.01 M solution of $Ba(OH)_2$ at 25°C?	
Thinking	**Working**
Write down the reaction in which $Ba(OH)_2$ dissociates.	In water, each mole of $Ba(OH)_2$ completely dissociates to release 2 moles of OH^- ions. $Ba(OH)_2(aq) \rightarrow Ba^{2+}(aq) + 2OH^-(aq)$
Determine the concentration of $[OH^-]$ ions.	$[OH^-] = 2 \times [Ba(OH)_2]$ = 2 × 0.01 M = 0.02 M

Determine the $[H_3O^+]$ in the diluted solution by substituting the $[OH^-]$ into the ionic product of water: $K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$ M^2	$K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$ M^2 $[H_3O^+] = \frac{K_w}{[OH^-]}$ $= \frac{1.00 \times 10^{-14}}{0.02}$ $= 5 \times 10^{-13}$
Substitute the value of $[H_3O^+]$ into: pH $= -\log_{10}[H_3O^+]$ Use the logarithm function on your calculator to calculate the answer.	pH $= -\log_{10}[H_3O^+]$ $= -\log_{10}[5 \times 10^{-13}]$ (use your calculator) $= 12.3$

Worked example: Try yourself 15.3.4

CALCULATING pH IN A SOLUTION WHERE THE SOLUTE CONCENTRATION IS NOT GIVEN

What is the pH of a solution at 25°C that contains 0.50 g KOH in 500 mL of solution?

Thinking	**Working**
Determine the number of moles of KOH.	$n(\text{KOH}) = \frac{m}{M}$ $= \frac{0.50}{56.0}$ $= 8.9 \times 10^{-3}$ mol
Write the equation for dissociation of KOH.	$KOH(aq) \rightarrow K^+(aq) + OH^-(aq)$ KOH is completely dissociated in water.
Determine the number of moles of OH^- based on the dissociation equation.	$n(OH^-) = n(\text{KOH})$ $= 8.9 \times 10^{-3}$ mol
Use the formula for determining concentration given number of mole and volume: $c = \frac{n}{V}$	$n = 8.9 \times 10^{-3}$ mol $V = 0.500$ L $c = \frac{8.9 \times 10^{-3}}{0.500}$ $= 0.0179$ M
Determine the in the diluted solution by substituting the into the ionic product of water: $K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$ M^2	$K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$ M^2 $[H_3O^+] = \frac{K_w}{[OH^-]}$ $= \frac{1.00 \times 10^{-14}}{0.0179}$ $= 5.6 \times 10^{-13}$
Substitute the value of into: pH $= -\log_{10}[H_3O^+]$ Use the logarithm function on your calculator to calculate the answer.	pH $= -\log_{10}[H_3O^+]$ $= -\log_{10}[5.6 \times 10^{-14}]$ (use your calculator) $= 12.25$

Worked example: Try yourself 15.3.5

CALCULATING [H_3O^+] IN A SOLUTION OF A GIVEN pH

Calculate the [H_3O^+] in a solution of pH 10.4 at 25°C.	
Thinking	**Working**
Decide which form of the relationship between pH and [H_3O^+] should be used: pH = $-\log_{10}[H_3O^+]$ or $[H_3O^+] = 10^{-pH}$	As you have the pH and are calculating [H_3O^+], use: $[H_3O^+] = 10^{-pH}$
Substitute the value of pH into the relationship expression and use a calculator to determine the answer.	$[H_3O^+] = 10^{-pH}$ $= 10^{-10.4}$ $= 4 \times 10^{-11}$ M

15.3 Key questions

1 $K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$

$[OH^-] = \frac{1.00 \times 10^{-14}}{[H_3O^+]}$

$[OH^-] = \frac{1.00 \times 10^{-14}}{0.001} = 1.00 \times 10^{-11}$

2 $K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$

$[OH^-] = \frac{1.00 \times 10^{-14}}{[H_3O^+]}$

$[OH^-] = \frac{1.00 \times 10^{-14}}{5.70 \times 10^{-9}} = 1.75 \times 10^{-6}$

3 $K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$

$[OH^+] = \frac{1.00 \times 10^{-14}}{[OH^-]}$

$[OH^+] = \frac{1.00 \times 10^{-14}}{1.0 \times 10^{-5}} = 1.0 \times 10^{-9}$

4 pH = $-\log_{10}[H_3O^+]$

$= -\log_{10}0.01$

$= 2$

5 As nitric acid is a strong acid, the concentration of H_3O^+ ions is 0.0010 M.

pH = $-\log_{10}[H_3O^+]$

$= -\log_{10}0.0010$

$= 3$

6 $[H_3O^+] = 10^{-pH}$

$= 10^{-6}$

$= 1.0 \times 10^{-6}$ (or 0.000001) M

7 $M(HCl) = M(H) + M(Cl)$

$= 1.0 + 35.5$

$= 36.5$

As HCl completely dissociates to H_3O^+ and Cl^-, $n(HCl) = n(H_3O^+)$, therefore $n(H_3O^+) = 0.01$ mol.

$c = \frac{n}{V} = \frac{0.01}{0.20} = 0.050$ M

$[H_3O^+] = 0.050$ M

pH = $-\log_{10}[H_3O^+]$

$= -\log_{10}0.050 = 1.3$

$= 1.3$

Section 15.4 Dilution of acids and bases

Worked example: Try yourself 15.4.1

CALCULATING MOLAR CONCENTRATION AFTER DILUTION

Calculate the molar concentration of nitric acid when 80.0 mL of water is added to 20.0 mL of 5.00 M HNO_3.	
Thinking	**Working**
The number of moles of solute does not change during a dilution. So $c_1V_1 = c_2V_2$, where c is the concentration in M and V is the volume of the solution. (Each of the volume units must be the same, although not necessarily litres.)	$c_1V_1 = c_2V_2$
Identify given values for concentrations and volumes before and after dilution. Identify the unknown.	Remember that 80.0 mL was added to 20.0 mL, so the final volume is 100.0 mL. $c_1 = 5.00$ M $V_1 = 20.0$ mL $V_2 = 100.0$ mL You are required to calculate c_2, the concentration after dilution.
Transpose the equation and substitute the known values into the equation to find the required value.	$c_2 = \frac{c_1 \times V_1}{V_2}$ $= \frac{5.00 \times 20.0}{100.0}$ $= 1.00$ M

Worked example: Try yourself 15.4.2

CALCULATING THE VOLUME OF WATER TO BE ADDED IN A DILUTION

How much water must be added to 15.0 mL of 10.0 M NaOH to dilute the solution to 2.00 M?	
Thinking	**Working**
The number of moles of solute does not change during a dilution. So $c_1V_1 = c_2V_2$, where c is the concentration in M and V is the volume of the solution. (Each of the volume units must be the same, although not necessarily litres.)	$c_1V_1 = c_2V_2$
Identify given values for concentrations and volumes before and after dilution. Identify the unknown.	$c_1 = 10.0$ M $V_1 = 15.0$ mL $c_2 = 2.00$ M You are required to calculate V_2, the volume of the diluted solution.
Transpose the equation and substitute the known values into the equation to find the required value.	$V_2 = \frac{c_1 \times V_1}{c_2}$ $= \frac{10.0 \times 15.0}{2.00}$ $= 75.0$ mL
Calculate the volume of water to be added.	Volume of dilute solution = 75.0 mL Initial volume of NaOH = 15.0 mL So 75.0 – 15.0 = 60.0 mL of water must be added.

Worked example: Try yourself 15.4.3

CALCULATING THE pH OF A DILUTED ACID

10.0 mL of 0.1 M HCl is diluted to 30.0 mL. Calculate the pH of the diluted solution.

Thinking	**Working**
Identify given values for concentrations and volumes before and after dilution.	$c_1 = 0.1$ M $V_1 = 10.0$ mL $V_2 = 30.0$ mL $c_2 = ?$
Calculate c_2, which is the concentration of H_3O^+ after dilution, by transposing the formula: $c_1V_1 = c_2V_2$	$c_2 = \frac{c_1 \times V_1}{V_2}$ $= \frac{0.1 \times 10.0}{30.0}$ $= 0.03$ M
Calculate pH using: $pH = -\log_{10}[H_3O^+]$ Use the logarithm function on your calculator to calculate pH.	$pH = -\log_{10}[H_3O^+]$ $= -\log_{10}(0.03)$ $= 1.5$

Worked example: Try yourself 15.4.4

CALCULATING THE pH OF A DILUTED BASE

15.0 mL of 0.02 M KOH is diluted to 60.0 mL. Calculate the pH of the diluted solution.

Thinking	**Working**
Identify given values for concentrations and volumes before and after dilution.	$c_1 = 0.02$ M $V_1 = 15.0$ mL $V_2 = 60.0$ mL $c_2 = ?$
Calculate c_2, which is the concentration of $[OH^-]$ after dilution, by transposing the formula: $c_1V_1 = c_2V_2$	$c_2 = \frac{c_1 \times V_1}{V_2}$ $= \frac{0.02 \times 15.0}{60.0}$ $= 0.005$ M
Determine the $[H_3O^+]$ in the diluted solution by substituting the $[OH^-]$ into the ionic product of water: $K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$ M^2	$K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$ M^2 $[H_3O^+] = \frac{K_w}{[OH^-]}$ $= \frac{1.00 \times 10^{-14}}{0.005}$ $= 2 \times 10^{-12}$ M
Calculate pH using: $pH = -\log_{10}[H_3O^+]$ Use the logarithm function on your calculator to calculate pH.	$pH = -\log_{10}[H_3O^+]$ $= -\log_{10}(2 \times 10^{-12})$ $= 11.7$

15.4 Key questions

1 $c_1V_1 = c_2V_2$, $c_1 = 0.01$ M, $V_1 = 3.0$ L, $V_2 = 4.0$ L, $c_2 = ?$

$0.01 \times 3.0 = c_2 \times 4.0$

$c_2 = \frac{0.10 \times 3.0}{4.0} = 0.075$ M

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2 $c_1V_1 = c_2V_2$, $c_1 = 2.0$ M, $V_1 = 10$ mL, $V_2 = ?$, $c_2 = 0.50$ M

$2.0 \times 10 = 0.50 \times V_2$

$V_2 = \frac{2.0 \times 10}{0.50} = 40$ mL

Final volume – initial volume = 40 – 10 = 30 mL so 30 mL of water needs to be added.

3 $c_1V_1 = c_2V_2$, $c_1 = 0.60$ M, $V_1 = 20$ mL, $V_2 = ?$, $c_2 = 0.10$ M

$0.60 \times 20 = 0.10 \times V_2$

$V_2 = \frac{0.60 \times 20}{0.10} = 120$ mL

Final volume – initial volume = 120 – 20 = 100 mL, so 100 mL of water needs to be added.

4 This increases the pH by one unit.

5 $c_1V_1 = c_2V_2$, $c_1 = 0.100$ M, $V_1 = 20.0$ mL, $V_2 = 50.0$ mL, $c_2 = ?$

$0.100 \times 20.0 = c_2 \times 50.0$

$c_2 = \frac{0.100 \times 20.0}{50.0} = 0.040$ M

As NaOH dissociates completely to Na^+ and OH^-, $[NaOH] = [OH^-]$, therefore $[OH^-] = 0.040$ M

$K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$

$1.00 \times 10^{-14} = [H_3O^+] \times 0.040$

$[H_3O^+] = \frac{1.0 \times 10^{-14}}{0.040} = 2.5 \times 10^{-13}$

$pH = -\log_{10}[H_3O^+]$

$pH = -\log_{10}2.5 \times 10^{-13} = 12.6$

6 **a** **i** The concentration of H_3O^+ ions equals the concentration of a monoprotic acid = 0.0010 M = 10^{-3} M.

ii $[OH^-] = \frac{1.00 \times 10^{-14}\ M^2}{[H_3O^+]} = \frac{1.00 \times 10^{-14}\ M^2}{0.0010\ M} = 1 \times 10^{-11}$ M

iii $pH = -\log_{10}[H_3O^+] = -\log(10^{-3}\ M) = 3$

b **i** $[H_3O^+] = 0.030$ M

ii $[OH^-] = \frac{1.00 \times 10^{-14}\ M^2}{[H_3O^+]} = \frac{1.00 \times 10^{-14}\ M^2}{0.030\ M} = 3.33 \times 10^{-13}$ M

iii $pH = -\log_{10}[H_3O^+] = -\log(0.030\ M) = 1.52$

c **i** NaOH dissociates completely to Na^+ and OH^-, $[NaOH] = [OH^-]$, therefore $[OH^-] = 0.010$ M.

$[H_3O^+] = \frac{1.00 \times 10^{-14}\ M^2}{0.010\ M} = 1.0 \times 10^{-12}$ M

ii From part a, $[OH^-] = 0.010$ M

iii $pH = -\log_{10}[H_3O^+] = -\log(10^{-12}\ M) = 12$

d **i** $10^{-4.5}$ M HCl = $[H_3O^+] = 3.16 \times 10^{-5}$ M

ii $[OH^-] = \frac{1.00 \times 10^{-14}\ M^2}{[H_3O^+]} = \frac{1.00 \times 10^{-14}\ M^2}{3.16 \times 10^{-5}\ M} = 3.16 \times 10^{-10}$ M

iii $pH = -\log_{10}[H_3O^+] = -\log(3.16 \times 10^{-5}\ M) = 4.5$

e **i** $[OH^-] = 2 \times [Ba(OH)_2] = 2 \times 0.0050\ M = 0.010$ M

$[H_3O^+] = \frac{1.00 \times 10^{-14}\ M^2}{0.010\ M} = 1.0 \times 10^{-12}$ M

ii From part a, $[OH^-] = 0.010$ M

iii $pH = -\log_{10}[H_3O^+] = -\log(1.0 \times 10^{-12}\ M) = 12$

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Section 15.5 Reactions of acids and bases

Worked example: Try yourself 15.5.1

WRITING AN IONIC EQUATION FOR AN ACID–BASE REACTION

Write an ionic equation for the reaction that occurs when sulfuric acid is added to potassium hydroxide solution.	
Thinking	**Working**
What is the general reaction? Identify the products formed.	Acid + metal hydroxide → salt + water A solution of potassium sulfate and water is formed.
Identify the reactants and products. Indicate the state of each, i.e. (aq), (s), (l) or (g).	**Reactants:** H_2SO_4 is ionised in solution, forming H^+(aq) and SO_4^{2+}(aq). KOH is dissociated in solution, forming K^+(aq) and OH^-(aq). **Products:** Potassium sulfate is dissociated and exists as K^+(aq) and SO_4^{2-}(aq). Water is a molecular compound and its molecular formula is H_2O(l).
Write the equation showing all reactants and products. (There is no need to balance the equation yet.)	$H^+(aq) + SO_4^{2-}(aq) + K^+(aq) + OH^-(aq) \rightarrow K^+(aq) + SO_4^{2-}(aq) + H_2O(l)$
Identify the spectator ions: the ions that have an (aq) state both as a reactant and as a product.	K^+(aq) and SO_4^{2-}(aq)
Rewrite equation without the spectator ions. Balance the equation with respect to number of atoms of each element and charge.	$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$ Note that if hydronium ions are represented as H_3O^+(aq), rather than as H^+(aq), this reaction would be written as: $H_3O^+(aq) + OH^-(aq) \rightarrow 2H_2O(l)$

Worked example: Try yourself 15.5.2

WRITING IONIC EQUATIONS FOR REACTIONS BETWEEN ACIDS AND METAL CARBONATES

What products are formed when a solution of hydrochloric acid is added to a solution of sodium hydrogen carbonate? Write an ionic equation for this reaction.	
Thinking	**Working**
What is the general reaction? Identify the products.	Acid + metal hydrogen carbonate → salt + water + carbon dioxide Products of this reaction are sodium chloride in solution, water and carbon dioxide gas
Identify the reactants and products. Indicate the state of each, i.e. (aq), (s), (l) or (g).	**Reactants:** Hydrochloric acid is ionised in solution, forming H^+(aq) and Cl^-(aq) ions. Sodium hydrogen carbonate dissociates in solution and exists as Na^+(aq) and HCO_3^-(aq) ions. **Products:** Sodium chloride is dissociated into Na^+(aq) and Cl^-(aq) ions. Water has the formula H_2O(l). Carbon dioxide has the formula CO_2(g).
Write the equation showing all reactants and products. (There is no need to balance the equation yet.)	$H^+(aq) + Cl^-(aq) + Na^+(aq) + HCO_3^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l) + CO_2(g)$

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Identify the spectator ions.	$Na^+(aq) + Cl^-(aq)$
Rewrite the equation without the spectator ions. Balance the equation with respect to number of atoms of each element and charge.	The balanced equation with spectator ions omitted is: $H^+(aq) + HCO_3^-(aq) \rightarrow H_2O(l) + CO_2(g)$ Note that if hydronium ions are represented as $H_3O^+(aq)$, rather than as $H^+(aq)$, this reaction would be written as: $H_3O^+(aq) + HCO_3^-(aq) \rightarrow 2H_2O(l) + CO_2(g)$

Worked example: Try yourself 15.5.3

WRITING IONIC EQUATIONS FOR REACTIONS BETWEEN ACIDS AND REACTIVE METALS

Write an ionic equation for the reaction that occurs when aluminium is added to a dilute solution of hydrochloric acid.	
Thinking	**Working**
What is the general reaction? Identify the products formed.	Acid + reactive metal → salt + hydrogen Hydrogen gas and aluminium chloride solution are produced.
Identify the reactants and products. Indicate the state of each, i.e. (aq), (s), (l), (g).	**Reactants:** Aluminium is a solid, $Al(s)$. Hydrochloric acid is ionised, forming $H^+(aq)$ and $Cl^-(aq)$ ions. **Products:** Hydrogen gas, H_2. aluminium chloride is dissociated into $Al^{3+}(aq)$ and $Cl^-(aq)$ ions.
Write the equation showing all reactants and products. (There is no need to balance the equation yet.)	$H^+(aq) + Cl^-(aq) + Al(s) \rightarrow Al^{3+}(aq) + Cl^-(aq) + H_2(g)$
Identify the spectator ions.	$Cl^-(aq)$
Rewrite equation without the spectator ions. Balance the equation with respect to number of atoms of each element and charge.	$6H^+(aq) + 2Al(s) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

15.5 Key questions

1 a $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$
  $Mg(s) + 2H^+ (aq) \rightarrow Mg^{2+}(aq) + H_2(g)$
 b $Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$
  $Ca(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2(g)$
 c $Zn(s) + 2CH_3COOH(aq) \rightarrow Zn(CH_3COO)_2 + H_2(g)$
  $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$
 d $2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$
  $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

2 a Magnesium sulfate
 b Calcium chloride
 c Zinc acetate
 d Aluminium chloride

3 a i $ZnO(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2O(l)$
  ii $ZnO(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2O(l)$
 b i $Ca(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2(g)$
  ii $Ca(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2(g)$
 c i $Cu(OH)_2(s) + 2HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2H_2O(l)$
  ii $Cu(OH)_2(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + 2H_2O(l)$

d **i** $Mg(HCO_3)_2(s) + 2HCl(aq) \rightarrow MgCl_2(s) + 2H_2O(l) + 2CO_2(g)$
  ii $Mg(HCO_3)_2(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2O(l)$
e **i** $SnCO_3(s) + H_2SO_4(aq) \rightarrow SnSO_4(s) + H_2O(l) + CO_2(g)$
  ii $SnCO_3(s) + 2H^+(aq) \rightarrow Sn^{2+}(s) + H_2O(l) + CO_2(g)$

4 The first step, as in the development of any equation, is to write the correct chemical formulas for each of the chemicals involved.

a $2KOH(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$
$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$
b $NaOH(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l)$
$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$
c $Mg(OH)_2(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$
$Mg(OH)_2(s) + 2H^+(aq) \rightarrow Mg2^+(aq) + 2H_2O(l)$
d $CuCO_3(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l) + CO_2(g)$
$CuCO_3(s) + 2H^+(aq) \rightarrow Cu2^+(aq) + H_2O(l) + CO_2(g)$
e $KHCO_3(aq) + HF(aq) \rightarrow KF(aq) + H_2O(l) + CO_2(g)$
$HCO_3^-(aq) + H^+(aq) \rightarrow H_2O(l) + CO_2(g)$
f $Zn(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2(g)$
$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$
g $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$
$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$
h $NaHCO_3(s) + CH_3COOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l) + CO_2(g)$
$NaHCO_3(s) + H^+(aq) \rightarrow Na^+(aq) + H_2O(l) + CO_2(g)$

Section 15.6 Ocean acidity

15.6 Key questions

1 D. Carbon dioxide is sparingly soluble in water. It reacts to a limited extent with water, forming carbonic acid. Carbonic is a weak acid and is partially ionised, forming hydrogen carbonate and hydronium ions. The hydrogen carbonate ions are also ionised and form further hydronium ions and carbonate ions. There will also be a very low concentration of $OH^-(aq)$ ions because of the ionisation of water molecules.

2 Step 1. $H_2CO_3(aq) + H_2O(l) \rightarrow H_3O^+(aq) + HCO_3^-(aq)$
Step 2. $HCO_3^-(aq) + H_2O(l) \rightarrow H_3O^+(aq) + CO_3^{2-}(aq)$

3 Calcium carbonate formed through calcification.

4 $CO_3^{2-}(aq) + H_3O^+(aq) \rightarrow HCO_3^-(aq) + H_2O(l)$

CHAPTER 15 REVIEW

15.1 INTRODUCING ACIDS AND BASES

1 **a** NH_4^+
b HCl
c HCO_3^-
d H_3O^+
e $CH_3COOH(aq)$

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2 Remember that if a chemical acts as a base, it will accept a proton. In this question, the proton comes from a water molecule, but this is not always the case. If a chemical acts as an acid, it must be able to donate one or more protons.

a $PO_4^{3-}(aq) + H_2O(l) \rightarrow HPO_4^{2-}(aq) + OH^-(aq)$

b The $H_2PO_4^-$ accepts a proton from water, and acts as a base:
$H_2PO_4^-(aq) + H_2O(l) \rightarrow H_3PO_4(aq) + OH^-(aq)$
The $H_2PO_4^-$ donates a proton to the water, and acts as an acid:
$H_2PO_4^-(aq) + H_2O(l) \rightarrow HPO_4^{2-}(aq) + H_3O^+(aq)$

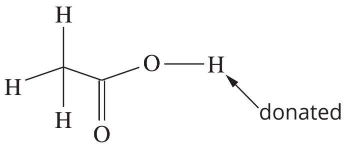
c $H_2S(aq) + H_2O(l) \rightarrow HS^-(aq) + H_3O^+(aq)$

3 **a** Cl^-
b H_3O^+
c OH^-
d SO_4^{2-}

4 **a** Sulfuric acid (H_2SO_4) is a diprotic acid because each molecule can donate two protons to a base:
i.e. $H_2SO_4(aq) + H_2O(l) \rightarrow H_3O^+(aq) + HSO_4^-(aq)$
$HSO_4^-(aq) + H_2O(l) \rightarrow H_3O^+(aq) + SO_4^{2-}(aq)$
The HSO_4^- ion, however, is amphiprotic because it can act as either an acid or a base, depending on the environment. In water it will undergo both acid and base reactions. For example:
As an acid: $HSO_4^-(aq) + H_2O(l) \rightarrow H_3O^+(aq) + SO_4^{2-}(aq)$
As a base: $HSO_4^-(aq) + H_2O(l) \rightarrow OH^-(aq) + H_2SO_4(aq)$

b A strong acid is one that ionises completely in solution (e.g. HCl). A concentrated acid is one in which there is a large amount of acid dissolved in a given volume of solution; for example, 5 M HCl and 5 M CH_3COOH are concentrated acids.

5



6 The first step, as in the development of any equation, is to write the correct chemical formulas for each of the chemicals involved.

a $HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(l)$
$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$
This is a Brønsted–Lowry acid–base reaction. H^+ donated to OH^-.

b $2HNO_3(aq) + Mg(s) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$
This is an acid reacting with a metal, not a Brønsted–Lowry acid–base reaction.

c $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
This a precipitation reaction.

d $CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$
$CuO(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l)$
This is a Brønsted–Lowry acid–base reaction. H^+ donated to O^{2-}.

15.2 STRENGTHS OF ACIDS AND BASES

7 $HClO_4(aq) + H_2O \rightarrow H_3O^+(aq) + ClO_4^-(aq)$
8 $HClO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + ClO_3^-(aq)$
9 $NH_3(aq) + H_2O(l) \rightleftharpoons OH^-(aq) + NH_4^+(aq)$
10 $H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons H_3PO_4\ (aq) + OH^-(aq)$

15.3 ACIDITY OF SOLUTIONS

11 At 25°C, $[OH^-] \times [H_3O^+] = 1 \times 10^{-14}\ M^2$

(Note: [ ] denotes concentration in molar units.)

$\therefore [OH^-] = \frac{1.00 \times 10^{-14}\ M^2}{[H_3O^+]}$

a 10^{-11} M

b 10^{-9} M

c 1.8×10^{-6} M

d 2.9×10^{-3} M

e 1.5×10^{-13} M

f 4.5×10^{-2} M

12 a i $pH = -\log_{10}[H_3O^+]$

$[H_3O^+] = 10^{-pH}$ (or 10 to the power of pH)

$= 10^{-1}$

$= 0.1$ M

ii At 25°C, $[OH^-] \times [H_3O^+] = 1 \times 10^{-14}\ M^2$

$\therefore [OH^-] = \frac{1.00 \times 10^{-14}\ M^2}{[H_3O^+]}$

$= \frac{1.00 \times 10^{-14}\ M^2}{0.10\ M} = 1.0 \times 10^{-13}$ M

b i 10^{-3} M

ii 10^{-11} M

c i 10^{-7} M

ii 10^{-7} M

d i 2×10^{-12} M

ii 5×10^{-3} M

13 A solution with a pH of 7.4 is basic.

At 25°C, $[H_3O^+] \times [OH^-] = 1 \times 10^{-14}\ M^2$ and $pH = -\log_{10}[H_3O^+]$

The assumption made is that at body temperature (approximately 38°C), the ionisation of water is not significantly different from that at 25°C.

14 Remember: $pH = -\log_{10}[H_3O^+]$

pH = 3

$\therefore [H_3O^+] = 10^{-3}$ M

pH = 5

$\therefore [H_3O^+] = 10^{-5}$ M

∴ the difference is a factor of 100.

15 a $[H^+] = 10^{-pH} = 0.001$ M

$[OH^-] = \frac{1.00 \times 10^{-14}\ M^2}{[H^+]} = \frac{1.00 \times 10^{-14}\ M^2}{10^{-3}\ M} = 1.0 \times 10^{-11}$ M

b $[H^+] = 10^{-10}$ M, $[OH^-] = 10^{-4}$

c $[H^+] = 3.16 \times 10^{-9}$ M, $[OH^-] = 3.16 \times 10^{-6}$ M

d $[H^+] = 1.58 \times 10^{-6}$ M, $[OH^-] = 6.3 \times 10^{-9}$ M

e $[H^+] = 2.5 \times 10^{-10}$ M, $[OH^-] = 4.0 \times 10^{-5}$ M

f $[H^+] = 3.16 \times 10^{-14}$ M, $[OH^-] = 0.316$ M

16 $pH = -\log_{10}[H_3O^+] = -\log_{10}5.3$

$[H_3O^+] = 10^{-5.3}$, $[OH^-] = \frac{1.00 \times 10^{-14}\ M^2}{10^{-5.3}\ M} = 2.0 \times 10^{-9}$ M

15.4 DILUTION OF ACIDS AND BASES

17 Remember to use $[OH^-] \times [H_3O^+] = 1 \times 10^{-14}$ M^2 at 25°C and pH = $-\log_{10}[H_3O^+]$.

a pH = 2, $\therefore$ $[H_3O^+] = 10^{-2}$ or 0.01 M

b The formula for amounts of substance in solution is $n = c \times V$ (where n is the amount in moles, c the concentration in mol L^{-1} and V the volume of the solution in litres).

$n = 0.01 \times 0.500 = 0.005$ mol

18 These three questions have two steps.

Step 1. Dilution questions are best answered using the formula $c_1V_1 = c_2V_2$, where c is the concentration in mol L^{-1} and V is the volume of the solution. Each of the volume units needs to be the same, although not necessarily litres.

Step 2. Also remember, pH = pH = $-\log_{10}[H_3O^+]$

a $c_2 = 0.25 \times \frac{10}{50} = 0.05$ M

$\therefore [H_3O^+] = 0.05$ M

pH = $-\log_{10}(0.005\text{ M}) = 1.3$

b $c_2 = 0.0050 \times \frac{20}{500} = 2 \times 10^{-4}$ M

$\therefore [OH^-] = 2.0 \times 10^{-4}$ M

$$[H_3O^+] = \frac{1.00 \times 10^{-14}\text{ M}^2}{[OH^-]} = \frac{1.00 \times 10^{-14}\text{ M}^2}{2 \times 10^{-4}\text{ M}} = 5.0 \times 10^{-11}\text{ M}$$

pH = $-\log_{10}(5.0 \times 10^{-11}\text{ M}) = 10.3$

c $c_2 = 0.15 \times \frac{10}{1500}$ (1.5 L has been converted to mL to maintain identical units)

= 0.001 M

$\therefore [H_3O^+] = 0.001\text{ M} = 10^{-3}$ M

$\therefore$ pH = 3

19 $c_1V_1 = c_2V_2$, $c_1 = 18.0$ M, $V_1 = ?$, $c_2 = 2.00$ M, $V_2 = 1.00$ L

$$V_1 = \frac{2.00 \times 1.00}{18.0} = 0.111\text{ L} = 111\text{ mL}$$

20 Initial pH = 2. $[H_3O^+] = 10^{-pH} = 10^{-2} = 0.010$ M

Final pH = 4. $[H_3O^+] = 10^{-pH} = 10^{-4} = 1.00 \times 10^{-4}$ M

$c_1V_1 = c_2V_2$, $c_1 = 0.010$ M, $V_1 = 10.0$ mL, $c_2 = 1.00 \times 10^{-4}$ M, $V_2 = ?$

$$V_2 = \frac{0.010 \times 10.0}{1.00 \times 10^{-4}} = 1000\text{ mL}$$

Volume of water added = final volume – initial volume = 1000 – 10.0 = 990 mL

21 pH(initial) = $-\log_{10}[H_3O^+] = 1.0$

$c_1V_1 = c_2V_2$, $c_1 = 0.10$ M, $V_1 = 40.0$ mL, $c_2 = ?$, $V_2 = 500$ mL

$$c_2 = \frac{0.10 \times 40.0}{500} = 8 \times 10^{-3}\text{ M}$$

pH(final) = $-\log_{10}(8 \times 10^{-3}) = 2.10$

Therefore the pH will increase.

15.5 REACTIONS OF ACIDS AND BASES

22 **a** $HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)$

b $H_2SO_4(aq) + K_2CO_3(aq) \rightarrow K_2SO_4(aq) + H_2O(l) + CO_2(g)$

c $2H_3PO_4(aq) + 3Ca(HCO_3)_2(s) \rightarrow Ca_3(PO_4)_2(s) + 6H_2O(l) + 6CO_2(g)$

d $2HF(aq) + Zn(OH)_2(s) \rightarrow ZnF_2(aq) + 2H_2O(l)$

23 E. An acid reacts with a base to produce a salt plus water. In this case, the salt produced is magnesium chloride, which is soluble in water and exists as magnesium ions and chloride ions.

24 $2H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)$

25 A carbonate

26 **a** $2Al(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$

b $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

15.6 OCEAN ACIDITY

27 $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$

28 Photosynthesis

29 $H_2CO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$

$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CO_3^{2-}(aq)$

30 **a** Krill eggs will not hatch at lower pH, therefore there will be a lower number of krill.

b Krill are a food source for many marine animals. If the number of krill cannot sustain organisms that feed directly on them, such as seals, penguins and small fish, then organisms such as killer whales and larger fish also would not survive. This will lead to an overall decrease in the number of large species that live within specific ecosystems.

CONNECTING THE MAIN IDEAS

31 Solution A: weaker base, few freely moving charged particles—ammonia

Solution B: neutral, no freely moving charged particles—glucose

Solution C: strong base, many freely moving charged particles—sodium hydroxide

Solution D: strong acid, many freely moving charged particles—hydrochloric acid

Solution E: weaker acid, few freely moving charged particles—ethanoic acid

32 **a** **i** A Brønsted–Lowry acid is a proton donor.

ii A strong base is a substance that ionises completely in water.

iii Molarity is a measure of concentration of a solution expressed in mol L^{-1}.

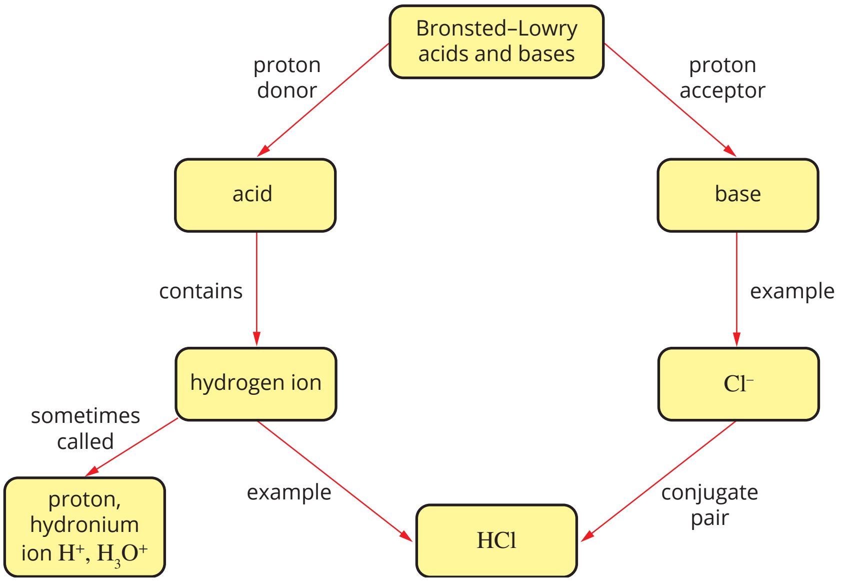
iv The conjugate acid of a base contains one more hydrogen ion (proton) than the base.

b An amphiprotic substance can act as an acid (proton donor) or a base (proton acceptor).

Acting as an acid: $HCO_3^-(aq) + H_2O(l) \rightarrow CO_3^{2-}(aq) + H_3O^+(aq)$

Acting as a base: $HCO_3^-(aq) + H_2O(l) \rightarrow H_2CO_3(aq) + OH^-(aq)$

33



Chapter 16 Redox reactions in water

Section 16.1 Introducing redox reactions

Worked example: Try yourself 16.1.1

IDENTIFYING OXIDATION AND REDUCTION

Write the oxidation and reduction half-equations for the reaction with the overall equation:
$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

Thinking	**Working**
Identify the ions in the product.	NaCl is made up of Na^+ and Cl^- ions.
Write the half-equation for the oxidation of the reactant that forms positive ions and balance the equation with electrons.	$Na(s) \rightarrow Na^+(s) + e^-$
Write the half-equation for the reduction of the reactant that forms negative ions and balance the equation with electrons.	$Cl_2(g) + 2e^- \rightarrow 2Cl^-(s)$

Worked example: Try yourself 16.1.2

WRITING SIMPLE HALF-EQUATIONS

When a piece of copper metal is placed into a silver nitrate solution, silver metal is formed and the solution gradually turns blue, indicating the presence of copper(II) ions in solution. The oxidation and reduction reactions can be represented by two half-equations.

Write these half-equations and identify the substances that are oxidised and reduced.

Thinking	**Working**
Identify one reactant and the product it forms and write them on each side of the equation. Balance the equation for the element.	$Ag^+(aq) \rightarrow Ag(s)$
Add electrons to balance the equation for charge.	$Ag^+(aq) + e^- \rightarrow Ag(s)$
To decide whether the reactant is oxidised or reduced, remember that oxidation is loss of electrons and reduction is gain of electrons.	Electrons are gained, so this is reduction. The $Ag^+(aq)$ is being reduced.
Identify the second reactant and the product it forms, and write them on each side of the equation. Balance the equation for the element.	$Cu(s) \rightarrow Cu^{2+}(aq)$
Add electrons to balance the equation for charge.	$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
To decide whether the reactant is oxidised or reduced, remember that oxidation is loss of electrons and reduction is gain of electrons.	Electrons are lost, so this is oxidation. The Cu(s) is being oxidised.

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Worked example: Try yourself 16.1.3

WRITING OVERALL REDOX EQUATIONS FROM HALF-EQUATIONS

Potassium metal is oxidised by oxygen gas in air to form solid potassium oxide. Write the half-equations for the reaction and hence write the balanced overall equation.	
Thinking	**Working**
Identify one reactant and the product it forms, and write them on each side of the equation. Balance the equation for the element.	$O_2(g) \rightarrow 2O^{2-}(s)$
Add electrons to balance the equation for charge.	$O_2(g) + 4e^- \rightarrow 2O^{2-}(s)$
Identify the second reactant and the product it forms, and write them on each side of the equation. Balance the equation for the element.	$K(s) \rightarrow K^+(s)$
Add electrons to balance the equation for charge.	$K(s) \rightarrow K^+(s) + e^-$
Multiply one equation by a suitable factor to ensure that the number of electrons is balanced.	$(K(s) \rightarrow K^+(s) + e^-) \times 4$ $4K(s) \rightarrow 4K^+(s) + 4e^-$
Add the oxidation and the reduction half-equations together, cancelling out electrons so that none are in the final equation. Combine ions to create the formula of the product.	$O_2(g) + 4e^- \rightarrow 2O^{2-}(s)$ $4K(s) \rightarrow 4K^+(s) + 4e^-$ When the electrons have been cancelled, the overall equation is: $4K(s) + O_2(g) \rightarrow 2K_2O(s)$

16.1 Key questions

1 a Oxidation
  b Reduction
  c Reduction
  d Oxidation

2 a $Fe(s) \rightarrow Fe^{3+}(aq) + 3e^-$ – oxidation
  b $K(s) \rightarrow K^+(aq) + e^-$ – oxidation
  c $F_2(g) + 2e– \rightarrow 2F^-(aq)$ – reduction
  d $O_2(g) + 4e– \rightarrow 2O^{2-}(aq)$ – reduction

3 a $Fe(s)$ has been oxidised to $Fe^{2+}(aq)$.
  b $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
  c $H^+(aq)$
  d $H^+(aq)$ has been reduced to H_2
  e $2H^+(aq) + 2e^- \rightarrow H_2(g)$
  f $Fe(s)$
  g $Fe^{2+}(aq)/Fe(s)$ and $H^+(aq)/H_2(g)$

4 a Magnesium is oxidised, copper ions are reduced.
  b $Mg(s) \rightarrow Mg^{2+}(aq) + e^-$
  c $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
  d $Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$
  e oxidant Cu^{2+}; reductant Mg
  f The solution loses some of its blue colour due to the loss of $Cu^{2+}(aq)$, which reacts to form Cu(s).

5 a Cu^{2+}
  b Cu

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6 **a** CaO

b Ca(s)

c $Ca(s) \rightarrow Ca^{2+}(s) + 2e^-$

d $O_2(g)$

e $O_2(g) + 4e^- \rightarrow 2O^{2-}(s)$

f $2Ca(s) + O_2(g) \rightarrow 2CaO(s)$

g Calcium has been *oxidised* by *oxygen* to calcium ions. The *oxygen* has gained electrons from the *calcium*. The oxygen has been *reduced* by *calcium* to oxide ions. The *calcium* has lost electrons to the *oxygen*.

16.1 Extension

Worked example: Try yourself 16.1.4

WRITING HALF-EQUATIONS FOR COMPLEX REDOX REACTIONS

Potassium dichromate ($K_2Cr_2O_7$) reacts with potassium iodide (KI) in acidified solution. The dichromate ion ($Cr_2O_7^{2-}$) is reduced to form Cr^{3+} , and the iodide ion (I^-) is oxidised to I_2.

Write:

a the half equation for the oxidation of the I^- to I_2

b the half equation for the reduction of $Cr_2O_7^{2-}$ to Cr^{3+}

c an overall ionic equation for the reaction.

(The potassium ions are spectators and do not appear in the ionic equation.)

Thinking	Working
The oxidation of I^- to I_2 is represented by a simple half-equation:	$2I^-(aq) \rightarrow I_2(aq) + 2e^-$
The half-equation to represent the reduction of $Cr_2O_7^{2-}$ to Cr^{3+} is more complex and will require the use of H^+ (aq) and $H_2O(l)$.	
Balance all atoms in the half equation except oxygen and hydrogen.	$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$
Balance the oxygen atoms by adding water. (Oxygen atoms react to form water in acidic solution.)	$Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$
Balance the hydrogen atoms by adding H^+ ions (which are present in the acidic solution).	$Cr_2O_7^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$
Balance the charges on both sides of the equation by adding electrons to the more positive side. Add states.	$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$
Add the oxidation half equation to the reduction half equation to get the overall equation, making sure that the number of electrons used in reduction equals the number of electrons released during oxidation.	$3 \times [2I^-(aq) \rightarrow I_2(aq) + 2e^-]$ $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$ $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6I^-(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(l) + 3I_2(aq)$

1 **a** Step 1: Balance all atoms in the half-equation except oxygen.

$SO_4^{2-} \rightarrow SO_2$

Step 2: Balance the oxygen atoms by adding water molecules.

$SO_4^{2-} \rightarrow SO_2 + 2H_2O$

Step 3: Balance the hydrogen atoms by adding H^+ ions (which are present in acidic solution).

$SO_4^{2-} + 4H^+ \rightarrow SO_2 + 2H_2O$

Step 4: Balance the charge by adding electrons to the more positive side.

$SO_4^{2-} + 4H^+ + 2e^- \rightarrow SO_2 + 2H_2O$

Step 5: Add symbols of state.

$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow SO_2(g) + 2H_2O(l)$

Use the same method to balance the other half-equations.

b $H_2O_2(aq) \rightarrow O_2(g) + 2H^+(aq) + 2e^-$

c $H_2S(g) \rightarrow S(s) + 2H^+(aq) + 2e^-$

d $MnO_4^-(aq) + 4H^+(aq) + 3e^- \rightarrow MnO_2(s) + 2H_2O(l)$

e $Ta_2O_5(s) + 10H^+(aq) + 10e^- \rightarrow 2Ta(s) + 5H_2O(l)$

f $SO_3^{2-}(aq) + H_2O(l) \rightarrow SO_4^-(aq) + 2H^+(aq) + 2e^-$

g $IO_3^-(aq)+ 6H^+(aq) + 6e^- \rightarrow I^-(aq) + 3H_2O(l)$

2 **a** $H_2O_2(aq) + 2H^+(aq) + 2e^- \rightarrow 2H_2O(l) \times 4$
$PbS(s) + 4H_2O(l) \rightarrow PbSO_4(s) + 8H^+(aq) + 8e^- \times 1$
$4H_2O_2(aq) + PbS(s) \rightarrow PbSO_4(s) + 4H_2O(l)$

b $I_2(aq) + 2e^- \rightarrow 2I^-(aq)$
$H_2S(g) \rightarrow S(s) + 2H^+(aq) + 2e^-$
$I_2(aq) + H_2S(g) \rightarrow 2I^- (aq) + S(s) + 2H^+(aq)$

c $SO_3^{2-}(aq) + H_2O(l) \rightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^- \times 5$
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l) \times 2$
$5SO_3^{2-}(aq) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 3H_2O(l)$

d $NO(g) + 2H_2O(l) \rightarrow NO_3^-(aq) + 4H^+(aq) + 3e^- \times 2$
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$
$2NO(g) + Cr_2O_7^{2-}(aq) + 6H^+(aq) \rightarrow 2NO_3^-(aq) + 2Cr^{3+}(aq) + 3H_2O(l)$

Section 16.2 The reactivity series of metals

Worked example: Try yourself 16.2.1

PREDICTING METAL DISPLACEMENT REACTIONS

Predict whether copper will displace gold from a solution containing gold ions and, if appropriate, write the overall equation for the reaction.

Thinking	Working
Locate the metal and the metal ions in the reactivity series.	Metals (reducing agents) are found on the right-hand side of the reactivity series and the metal ions (oxidising agents) are on the left-hand side of the reactivity series. $Au^+(aq) + e^- \rightarrow Au(s)$ $Ag^+(aq) + e^- \rightarrow Ag(s)$ $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
Determine whether the metal is below (and to the right of) the metal ion in the table. If this is the case, there will be a reaction.	We can see from the reactivity series that Cu is on the right-hand side because it is a reducing agent and it is below Au^+, so there will be a reaction.
Write the reduction reaction for the metal ions directly as it is written in the reactivity series, including states.	$Au^+(aq) + e^- \rightarrow Au(s)$
Write the oxidation reaction for the metal from the reactivity series, writing the metal on the left-hand side of the arrow (as a reactant). Include states.	$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
Combine the two half-equations, balancing electrons, to give the overall equation for the reaction.	$(Au^+(aq) + e^- \rightarrow Au(s)) \times 2$ $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ $2Au^+(aq) + Cu(s) \rightarrow 2Au(s) + Cu^{2+}(aq)$

16.2 Key questions

1 For reactions to occur spontaneously, the aqueous cation in the solution must be a stronger oxidant than the cation of the metal added.
- **a** No
- **b** No
- **c** Yes
- **d** Yes
- **e** Yes
- **f** Yes
- **g** No

2
- **a** $3Cu^{2+}(aq) + 2Al(s) \rightarrow 3Cu(s) + 2Al^{3+}(aq)$
- **b** No reaction
- **c** $2Ag^+(aq) + Zn(s) \rightarrow 2Ag(s) + Zn^{2+}(aq)$

3 Use iron and lead to test the solutions.

The copper(II) nitrate will react with both iron and lead.

The tin nitrate will react only with the iron.

The zinc nitrate will not react with either of the metals.

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Section 16.3 **Corrosion**

16.3 Key questions

1 $2Fe(s) + O_2(aq) + 2H_2O(l) \rightarrow 2Fe^{2+}(aq) + 4OH^-(aq)$
$4Fe(OH)_2(s) + O_2(aq) + 2H_2O(l) \rightarrow 4Fe(OH)_3(s)$

2 Near the coast, moisture in the air contains a higher amount of dissolved salts (e.g. NaCl(aq)). Salts are good electrolytes and accelerate the rusting process. They do this because they facilitate the production of ions at the oxidation and reduction sites of a corroding metal.

3 **a** Magnesium and zinc
b Only metals more reactive (i.e. more easily oxidised) than iron will act as sacrificial anodes. These more reactive metals will be oxidised in preference to the iron. The iron will become the cathode (the site of reduction) and no oxidation of the iron will take place.

4 Reduction of oxygen occurs when there is plenty of moisture and the oxygen concentration is high; that is, at or just above the surface of the water. A cathodic region forms here. The anodic region, then, is close to the cathode and where the oxygen concentration is less; that is, just below the surface. Oxidation, or corrosion, therefore occurs more rapidly below the surface.

5 As zinc is more reactive than steel, it will be the anode and steel will be the cathode. Oxidation, and so corrosion, of the nails will thus be prevented.

CHAPTER 16 **REVIEW**

16.1 INTRODUCING REDOX REACTIONS

1 D. The first half-equation shows that sodium metal, Na(s), is losing electrons, so is being oxidised. It must therefore be the reducing agent. The second half-equation shows that $Ag^+(aq)$ is gaining electrons, so it is being reduced. The $Ag^+(aq)$ must therefore be the oxidising agent.

2 A.

3 B. In this reaction, the half-equation for sodium is $Na(s) \rightarrow Na^+(s) + e^-$. Each sodium atom loses one electron.

4 C. The equation is not balanced correctly as it is written. Because the silver ion has a +1 charge, Ag^+ ions and Cu atoms will react in a 2:1 ratio in the reaction. The equation is balanced when it is written as:
$2Ag^+(aq) + Cu(s) \rightarrow 2Ag(s) + Cu^+(aq)$

5 **a** Oxidation is the gain of oxygen atoms; reduction is the loss of oxygen atoms.
b Oxidation is the loss of electrons; reduction is the gain of electrons.

6 The conjugate redox pairs are $Zn^{2+}(aq)/Zn(s)$ and $Co^{2+}(aq)/Co(s)$.

7 **a** Oxidation
b Oxidation
c Reduction
d Reduction

8 **a** $Ag^+(aq) + e^- \rightarrow Ag(s)$
b $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
c $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

9 **a** Mg(s) oxidised, $Cu^{2+}(aq)$ reduced
b Zn(s) oxidised, $Ag^+(aq)$ reduced
c CO(g) oxidised, PbO(s) reduced
d Ca(s) oxidised, $H^+(aq)$ reduced
e Fe(s) oxidised, $Pb^{2+}(aq)$ reduced

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10 **a** Zinc
  b $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
  c $H^+(aq)$
  d $2H^+(aq) + 2e^- \rightarrow H_2(g)$
  e $H^+(aq)$
  f $Zn(s)$
  g $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

16.2 THE REACTIVITY SERIES OF METALS

11 Ag, Cu, Sn, Ni, Mg, Li
12 **a** $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$
  b No reaction
  c No reaction
  d $Mg(s) + Pb^{2+}(aq) \rightarrow Mg^{2+}(aq) + Pb(s)$
  e No reaction
  f No reaction
  g No reaction
  h $Pb(s) + 2Ag^{2+}(aq) \rightarrow Pb^{2+}(aq) + 2Ag(s)$
13 **a** No
  b $Zn(s) + 2AgNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + 2Ag(s)$
  c $Zn(s) + SnCl_2(aq) \rightarrow ZnCl_2(aq) + Sn(s)$
  d $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$
14 Place 10 mL of each solution into a clean test-tube. Add a small piece of magnesium ribbon to each solution. The magnesium should be coated by displaced metal in the silver nitrate and lead nitrate solutions, but not in the sodium nitrate. The lead coating will be black. Over time, the silver coating will change from black to silver as more metal is deposited. To confirm the identity of the silver nitrate solution, take a fresh sample and add a small piece of copper. Copper will displace silver from the solution, giving silver deposit and a blue solution. Copper will not displace lead from the solution.
15 Zinc
16 Coatings of metals other than iron would be expected on the nails placed in 1 M solutions of $CuSO_4$ and $Pb(NO_3)_2$. These solutions contain oxidants that are strong enough to react with iron metal. The iron, in turn, reduces the metal ions in the solution, forming a coating on the nail.

16.3 CORROSION

17 D. A coating of paint will prevent contact between the metal and oxygen and water, so corrosion will not occur as long as the coating of paint is complete.
18 C. Water is very cold deep in the sea, so any reaction that occurs will do so very slowly. Rusting requires there to be water and oxygen for it to occur. The distance from the surface of the ocean, hence the low concentration of oxygen, and the cold temperature of the deep-sea would retard the corrosion of the iron.
19 $Fe^{2+}(aq) + OH^-(aq)$
  $Fe(OH)_2(s)$
  $Fe(OH)_3(s)$
  $Fe_2O_3.xH_2O$
20 Cathodic protection involves protection of a metal, usually iron, by making it the cathode of a cell. The metal can either be joined to a more reactive metal, such as zinc, or be connected to the negative terminal of a DC power source.
21 Copper is not a very reactive metal and so is not easily oxidised.
22 As the less reactive coating of tin on the nail was hammered, it would probably be damaged exposing the iron to water and oxygen. The more reactive iron will now donate electrons to the tin accelerating its rate of corrosion.
23 Near the seaside, moisture in the air contains a higher amount of dissolved salts (e.g. NaCl(aq)). Salts are good electrolytes and accelerate the rusting process. They do this because they facilitate the production of ions at the oxidation and reduction sites of a rusting metal.

CONNECTING THE MAIN IDEAS

24 Individual students' responses required.

25 **a** $Ag^+(aq)$; $Mg^{2+}(aq)$

b Mg(s); Ag(s)

c A coating of silver will form on the lead when it is placed in silver nitrate solution because Ag^+ ions are stronger oxidants than Pb^{2+} ions.

d Zinc and magnesium

26 Experiment 1 gives you evidence that T is a more reactive metal than R.

Experiment 2 suggests that S is a less reactive metal than T as it does not displace T from a solution of its salt, $T(NO_3)_2$.

Experiment 3 indicates that S is a less reactive metal than R as it does not displace R from a solution of its salt, RNO_3.

The information above tells you that S is less reactive than both T and R and that R is less reactive than T, so the correct order of increasing reactivity is $S < R < T$.

27 Group A will be oxidised because metals tend to lose electrons to non-metals when they react. Group B will be reduced because non-metals tend to gain electrons when they react.

28 **a** Tin is not very reactive and so does not corrode readily. It protects the steel layer under it from corroding by preventing its contact with water and oxygen.

b Zinc is too reactive to be used inside food cans. It would be oxidised easily by the contents of the can and so contaminate the food.

c If a can is dented, the tin layer may be cracked, exposing the steel underneath to the contents of the can. The steel forms an anode and the tin, which is less reactive, forms the cathode. The steel rapidly corrodes and contaminates the food. Rust that forms may flake off, leaving a hole in the can.

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Chapter 17 Water sample analysis

Section 17.1 The availability of drinking water

17.1 Key questions

1 Surface water has the highest risk of contamination since it can be easily polluted. It is not a protected water source like mains water where the water is tested and treated if needed.

2 Reservoirs provide water for reticulated water supplies through a system of pipes to individual households.

3 Most of the freshwater on Earth is locked in ice caps, glaciers or the soil.

4 Australia is the driest inhabited continent. Rainfall in Australia is extremely variable. Most of the rain that falls evaporates before it can enter rivers and reservoirs.

Section 17.2 Water quality

17.2 Key questions

1 A. Use of an amber glass container is not a requirement for sampling for mercury and no temperature is specified for the preservation of the sample.

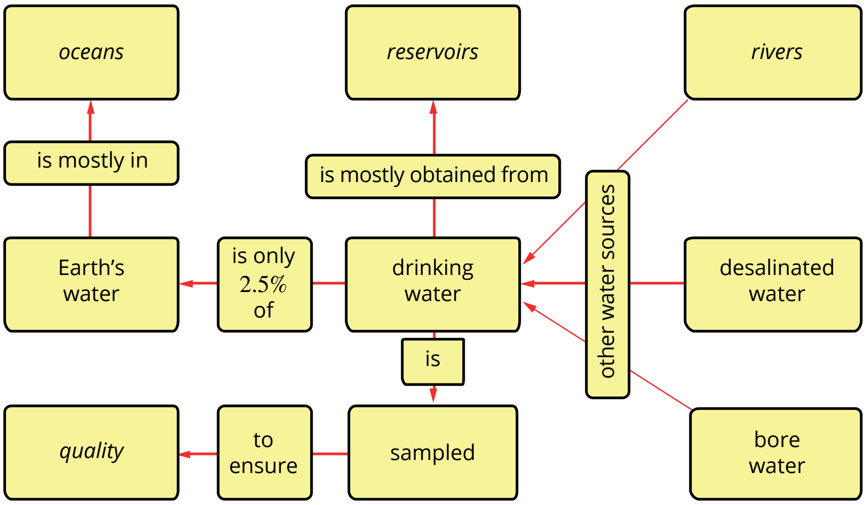
2 Sample should be taken upstream from where the person taking the sample stands.

3 To determine if the water is mixed and therefore what depth to take the samples. If the temperature is consistent, it can be assumed that the water is thoroughly mixed and samples can be taken half way down. If there is a temperature variation, then accurate information can only be obtained if a sample is taken in the middle of each temperature region.

CHAPTER 17 REVIEW

17.1 THE AVAILABILITY OF DRINKING WATER

1



2 A. Ice. Most of the drinkable water is found in the ice caps, although it is not readily accessible.

3 Bore water is water collected in aquifers (underground water-bearing rock) below the Earth's surface.

4 Most Australians live in the capital cities and their drinking water is supplied from reservoirs, as cities in Australia obtain their water from protected sources.

5 Oceans, ice caps and glaciers, groundwater, ground ice and permafrost, lakes, soil moisture, atmosphere as water vapour, rivers

17.2 WATER QUALITY

6 Considerations include:
- which chemical is to be analysed
- why the testing is to take place
- what are the health risks associated with the sampling
- which equipment is required
- what is the sample size required for the selected test
- what method is to be used to record the measurements
- where and at what depths should the samples should be taken
- how to obtain a representative sample
- what are the labelling, storage and transport requirements for the sample

7 D. Humidity. Temperature, depth and location must be recorded and future measurements made at the same depth and location. The humidity has no significant effect on the result.

8 Arsenic, cadmium and copper are below the Australian Drinking Water Guideline levels and therefore acceptable. Lead and mercury are above the Guideline levels and therefore unacceptable.

9 **a** Unsafe position
b Sample taken from surface of groundwater
c Container not sterile
d Container not protected from light
e Water sample not secured safely

10 Water samples are collected for analysis and comparison against water quality standards to determine its suitability for intended use and, if necessary, treatment required to purify the water. Protocols exist for obtaining these water samples to ensure they are collected in a consistent manner with the correct equipment. Without protocols, accurate data cannot be guaranteed.

CONNECTING THE MAIN IDEAS

11 Water can be used as a coolant in energy production, such as in power plants. It is used to remove wastes or as a cleaning agent in industries ranging from food production to car manufacture. Farming activities use large volumes of water for irrigation.

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Chapter 18 Analysis for salts in water

Section 18.1 The source of salts in water

18.1 Key questions

1 Heavy metals have a relatively high density. Examples of metals with a relatively high density that are toxic to living organisms include lead, mercury, copper, cadmium, chromium and the metalloid arsenic.

2 Toxic heavy metal levels are increased in our waterways by combustion of fuels and wastes (indirect), improper disposal of batteries, natural deposits in the Earth, and leaching from landfill.

3 A: Organometallic substances contain metal and non-metal atoms. They must have at least one metal–carbon bond to be defined as organometallic. Each of the other options consist of at least one compound that does not contain carbon.

4 **a** calcium ions, iron ions, magnesium ions

b $2C_{17}H_{35}COO^{-}(aq) + Ca^{2+}(aq) \rightarrow Ca(C_{17}H_{35}COO)_2(s)$

5 Carbonate ions from the washing soda precipitate calcium and magnesium ions from solution. This reaction takes place before the calcium and magnesium ions can react with the stearate ions from the soap. The soap is, therefore, much more effective in its washing action.

Ionic equation: $CO_3^{2-}(aq) + Ca^{2+}(aq) \rightarrow CaCO_3(s)$ or $CO_3^{2-}(aq) + Mg^{2+}(aq) \rightarrow MgCO_3(s)$

6 Total dissolved solids (TDS)

7

Electrical conductivity, EC (mS cm^{-1})	**a TDS (mg L^{-1})**	**b Fit or unfit to drink**
240	144	Fit
3200	1920	Fit
12 000	7200	Unfit

8 Group I and group II metal ions are usually safe to drink at low concentrations, but heavy metal ions are not.

Very toxic: Cu^{2+}, Pb^{2+}, Cd^{2+}, Hg^{2+}.

Relatively safe: Na^{+}, Mg^{2+}, Li^{2+}, K^{+}

Section 18.2 Gravimetric analysis

Worked example: Try yourself 18.2.1

USING MOLE RATIOS

Solutions of lead(II) nitrate and potassium iodide are mixed together, forming 0.30 mol of solid lead(II) iodide. How many moles of potassium iodide reacted in this reaction?	
Thinking	**Working**
Write a balanced equation for the reaction.	$2KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$
Determine the number of moles of the 'known substance'. The known substance is the one you are provided information about in the question.	$n(PbI_2) = 0.30$ mol
Find the mole ratio: mole ratio = $\frac{\text{coefficient of unknown}}{\text{coefficient of known}}$ The 'unknown substance' is the one whose moles you are required to calculate.	$\frac{n(KI)}{n(PbI_2)} = \frac{2}{1}$
Calculate the number of moles of the unknown substance from: $n(\text{unknown}) = n(\text{known}) \times (\text{mole ratio})$	$n(KI) = 0.30 \times \frac{2}{1}$ $= 0.60$ mol

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Worked example: Try yourself 18.2.2

MASS–MASS STOICHIOMETRIC PROBLEMS

A reaction between solutions of sodium sulfate and barium nitrate produces a precipitate of barium sulfate with a mass of 2.440 g. Calculate the mass of sodium sulfate required to produce this precipitate.

Thinking	**Working**
Write a balanced equation for the reaction.	$Na_2SO_4(aq) + Ba(NO_3)_2(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$
Calculate the number of moles of the known substance (the precipitate): $n = \frac{\text{mass } (m)}{\text{molar mass } (M)}$	$n(BaSO_4) = \frac{2.440}{233.4}$ = 0.01045 mol
Calculate the mole ratio: $\text{mole ratio} = \frac{\text{coefficient of unknown}}{\text{coefficient of known}}$	$\text{mole ratio} = \frac{n(Na_2SO_4)}{n(BaSO_4)} = \frac{1}{1}$
Calculate the number of moles of the known substance: $n(\text{unknown}) = n(\text{known}) \times \text{mole ratio}$	$n(Na_2SO_4) = 0.01045 \times \frac{1}{1}$ = 0.01045 mol
Calculate the mass of the unknown substance: $m = n(\text{unknown}) \times \text{molar mass}$	$m(Na_2SO_4) = 0.01045 \times 142.1$ = 1.486 g

Worked example: Try yourself 18.2.3

GRAVIMETRIC ANALYSIS

Water discharged from a mining plant contains silver ions present as silver nitrate ($AgNO_3$). Excess potassium chromate (K_2CrO_4) solution is added to a 50.0 g sample of water to precipitate the silver as silver chromate. The precipitate is heated to produce 1.32 g of silver chromate (Ag_2CrO_4).

Calculate the percentage mass of silver in the water supply. (The molar mass of Ag_2CrO_4 is 331.7 g mol^{-1}.)

Thinking	**Working**
Write a balanced equation for the precipitation reaction.	$2Ag_2NO_3(aq) + K_2CrO_4(aq) \rightarrow Ag_2CrO_4(s) + 2KNO_3(aq)$
Calculate the number of moles of precipitate using: $n = \frac{m}{M}$	$n(Ag_2CrO_4) = \frac{m}{M}$ $= \frac{1.32}{331.7}$ $= 3.98 \times 10^{-3}$ mol
Use the balanced equation to find the mole ratio of the 'known' and 'unknown' substances. The 'known substance' is the one you are provided about in the question; the 'unknown substance' is the one whose mass you are required to calculate.	$\text{mole ratio} = \frac{n(AgNO_3)}{n(Ag_2CrO_4)} = \frac{2}{1}$
Calculate the moles of unknown substance.	$n(Ag_2NO_3) = 3.98 \times 10^{-3} \times \frac{2}{1}$ $= 7.96 \times 10^{-3}$ mol
Calculate the mass of unknown substance in the sample.	The question asks for % of silver not silver nitrate, so using mole ratios: $n(Ag) = n(AgNO_3) = 7.96 \times 10^{-3}$ mol $m(Ag) = n \times M$ $= 7.96 \times 10^{-3} \times 107.9$ = 0.859 g
Calculate the percentage mass of the unknown substance in the 50.0 g of sample.	$\% \text{ Ag} = \frac{0.859}{50.0} \times 100$ = 1.72%

18.2 Key questions

1 C. $AgNO_3(aq) + LiBr(aq) \rightarrow AgBr(s) + LiNO_3(aq)$

This equation shows silver bromide precipitating from the mixture of solutions of silver nitrate and lithium bromide. To be used for gravimetric analysis the chemical reaction must form a solid (precipitation reaction).

2
- **a** False
- **b** True
- **c** True
- **d** True
- **e** True

Gravimetric analysis cannot be used if a salt is insoluble. Even if the salt is soluble, the technique still requires at least one of the ions in the salt to form insoluble substances when other reactants are added. For accurate results, the precipitate should be washed carefully and then dried thoroughly. The precipitate must be relatively stable so that its composition does not change during heating.

3

Poor practice initially displayed	**Impact**
Precipitate is not dried to constant mass.	Mass of the precipitate is high due to the presence of water.
Precipitate is not washed with deionised water.	Mass of the precipitate is high due to the presence of impurities.
Precipitate is left on the sides of the flask.	Precipitate is pure but its measured mass is lower than it should be.
Precipitate decomposes when heated.	The composition of the precipitate is not consistent.

4 $m(Ca(NO_3)_2) = 8.95$ g

Working:

$$n(Ca(PO_4)_2) = \frac{5.64}{310.2} = 0.0182 \text{ mol}$$

$n(Ca(NO_3)_2) = 0.0182 \times 3 = 0.0545$ mol

$m(Ca(NO_3)_2) = 0.0545 \times 164 = 8.95$ g

5
- **a** Silver nitrate. All nitrate salts are highly soluble but AgCl is insoluble.
  $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
- **b** The mass of the precipitate will be used to initially calculate the number of moles of silver chloride.
- **c** Mole ratio $= \frac{n(NaCl)}{n(AgCl)} = \frac{1}{1}$

6
- **a** $n(Al_2(CO_3)_3 = \frac{4.68}{234} = 0.020$ mol
- **b** Mole ratio $= \frac{n(Al(NO_3)_3)}{n(Al_2(CO_3)_3)} = \frac{2}{1}$
- **c** $m(Al(NO_3)_3) = 0.04 \times 213 = 8.52$ g

7 $m(Hg(CH_3COO)_2) = 3.38$ g

Working:

$$n(HgI_2) = \frac{4.82}{454.4} = 0.0106 \text{ mol}$$

Mole ratio $= \frac{n(Hg(CH_3COO)_2)}{n(HgI_2)} = \frac{1}{1}$

$m(Hg(CH_3COO)_2) = 0.0106 \times 318.7 = 3.38$ g

8 50.00 mL sample of hard water measured.

Excess of sodium carbonate is added.

Reaction mixture is filtered.

Precipitate is dried.

Precipitate is weighed.

Number of mole of precipitate is calculated.

Number of mole of calcium is calculated.

Mass of calcium ions is calculated.

Section 18.3 Calculations involving excess reactants

Worked example: Try yourself 18.3.1

MASS–MASS STOICHIOMETRY CALCULATION WITH ONE REACTANT ‘IN EXCESS’

A solution containing 0.600 g of lead(II) nitrate is added to a solution containing 2.60 g of potassium iodide. An equation for the reaction is: $2KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$	
(a) Which reactant is the limiting reactant?	
Thinking	**Working**
Calculate the number of moles of each of the reactants in the equation using $n = \frac{m}{M}$ or $n = c \times V$ as appropriate.	Use the equation $n = \frac{m}{M}$. For $Pb(NO_3)_2$: $n(Pb(NO_3)_2) = \frac{0.600}{331.2} = 1.81 \times 10^{-3}$ mol For KI: $n(KI) = \frac{2.60}{166.0} = 1.57 \times 10^{-2}$ mol
Use the coefficients of the equation to find the limiting reactant.	The equation shows that 2 mol of KI reacts with 1 mol of $Pb(NO_3)_2$ So to react all of the $Pb(NO_3)_2$ you will require $\frac{2}{1} \times n(Pb(NO_3))$ of KI $\frac{2}{1} \times 1.81 \times 10^{-3} = 3.62 \times 10^{-3}$ mol As there is 1.57×10^{-2} mol available, the KI is in excess. So, $Pb(NO_3)_2$ is the limiting reactant (it will be completely consumed).
(b) What will be the mass of the precipitate?	
Thinking	**Working**
Find the mole ratio of the unknown substance to the limiting reactant from the equation coefficients: mole ratio = $\frac{\text{coefficient of unknown}}{\text{coefficient of limiting reactant}}$	From the equation coefficients: $\frac{\text{coefficient of } (PbI_2)}{\text{coefficient of } (Pb(NO_3)_2)} \times \frac{1}{1}$
Calculate the number of moles of the unknown substance using the number of moles of the limiting reactant: n(unknown) = n(limiting reactant) × mole ratio	$n(PbI_2) = n(Pb(NO_3)_2) \times \frac{1}{1}$ $= 1.81 \times 10^{-3} \times \frac{1}{1}$ $= 181 \times 10^{-3}$ mol
Calculate the mass of the unknown substance using: m(unknown) = n(unknown) × molar mass	Molar mass of $PbI_2 = 461.0$ g mol^{-1} $m(PbI_2) = 1.81 \times 10^{-3} \times 461.0$ $= 0.835$ g

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18.3 Key questions

1 **B** Calculate the number of moles of aluminium and of oxygen.
D Refer to the balanced equation.
A Use mole ratios to determine which reactant is limiting.
E Calculate the number of moles of aluminium oxide that forms.
C Calculate the mass of aluminium oxide that forms.

2

Nitrogen molecules available	**Hydrogen molecules available**	**Ammonia molecules produced**	**Nitrogen molecules in excess**	**Hydrogen molecules in excess**
2	10	4	0	4
879	477	318	720	0
9 mol	6 mol	4 mol	7 mol	0 mol

Explanation

- Both nitrogen molecules will react with 2 × 3 = 6 hydrogen molecules to produce 2 × 2 = 4 ammonia molecules, so 10 – 6 = 4 hydrogen molecules will be in excess.
- All 477 hydrogen molecules will react with $477 \times \frac{1}{3} = 159$ nitrogen molecules to produce $477 \times \frac{2}{3} = 318$ ammonia molecules. 879 – 159 = 720 nitrogen molecules will be in excess.
- All 6 mol hydrogen molecules will react with $6 \times \frac{1}{3} = 2$ mol nitrogen molecules to produce $6 \times \frac{2}{3} = 4$ mol ammonia molecules. 9 – 2 = 7 nitrogen molecules will be in excess.

3 **a** $2Na(s) + Cl_2 \rightarrow 2NaCl(s)$

b $n(Na) = \frac{m}{M} = \frac{25.0}{23.0} = 1.09$ mol

$n(Cl_2) = \frac{m}{M} = \frac{50.0}{71.0} = 0.704$ mol

$\frac{n(Na)}{n(Cl_2)} = \frac{2}{1}$

Sodium is the limiting reagent; use the moles of sodium to calculate $n(NaCl)$:

$\frac{n(NaCl)}{n(Na)} = \frac{1}{1}$

$n(NaCl) = n(Na) = 1.09$ mol

$m(NaCl) = n \times M = 1.09 \times 58.5 = 63.8$ g

4 **a** In the equation for the reaction:

$\frac{\text{coefficient of KI}}{\text{coefficient of Pb(NO}_3)_2} = \frac{2}{1} = 2$

2 mol of KI react with 1 mol of $Pb(NO_3)_2$

In this case, not all the $Pb(NO_3)_2$ can react and it will be in excess.

From the equation:

$n(Pb(NO_3)_2) \text{ reacted} = \frac{1}{2} \times n(KI) = \frac{1}{2} \times 1.0 = 0.50$ mol

So, $Pb(NO_3)_2$ is in excess by (1.0 – 0.50) = 0.50 mol.

b In the equation for the reaction:

$\frac{\text{coefficient of KI}}{\text{coefficient of Pb(NO}_3)_2} = \frac{2}{1} = 2$

2 mol of KI react with 1 mol of $Pb(NO_3)_2$.

In this case, not all the $Pb(NO_3)_2$ can react and it will be in excess.

From the equation:

$n(Pb(NO_3)_2) \text{ reacted} = \frac{1}{2} \times n(KI)$

$= \frac{1}{2} \times 0.5 = 0.25$ mol

So, $Pb(NO_3)_2$ is in excess by (2.0 – 0.25) = 1.75 mol

c $n(Pb(NO_3)_2) = \frac{1.00}{331} = 0.00302$ mol

$n(KI) = \frac{1.50}{166} = 0.00904$ mol

In the equation for the reaction:

$\frac{\text{coefficient of KI}}{\text{coefficient of Pb(NO}_3)_2} = \frac{2}{1} = 2$

2 mol of KI react with 1 mol of $Pb(NO_3)_2$.

In this case, not all of the KI can react and some of it will be in excess. This means that $Pb(NO_3)_2$ is the limiting reactant.

The equation shows that 1 mol of PbI_2 precipitate will form for every 1 mol of $Pb(NO_3)_2$ that reacts.

$n(PbI_2) = n(Pb(NO_3)_2) = 0.00302$ mol

$m(PbI_2) = 0.00302 \times 461 = 1.39$ g

d $n(Pb(NO_3)_2) = c \times V = 1.00 \times 0.0500 = 0.0500$ mol

$n(KI) = c \times V = 0.500 \times 0.0750 = 0.0375$ mol

$\frac{\text{coefficient of KI}}{\text{coefficient of Pb(NO}_3)_2} = \frac{2}{1} = 2$

2 mol of KI react with 1 mol of $Pb(NO_3)_2$.

In this case, not all of the $Pb(NO_3)_2$ can react and some of it will be in excess. This means KI is the limiting reactant.

The equation shows that 1 mol of PbI_2 precipitate will form for every 2 mol of KI that reacts.

$n(PbI_2) = \frac{1}{2} \times n(KI) = \frac{1}{2} \times 0.0375 = 0.00188$ mol

$m(PbI_2) = 0.00188 \times 461 = 8.64$ g

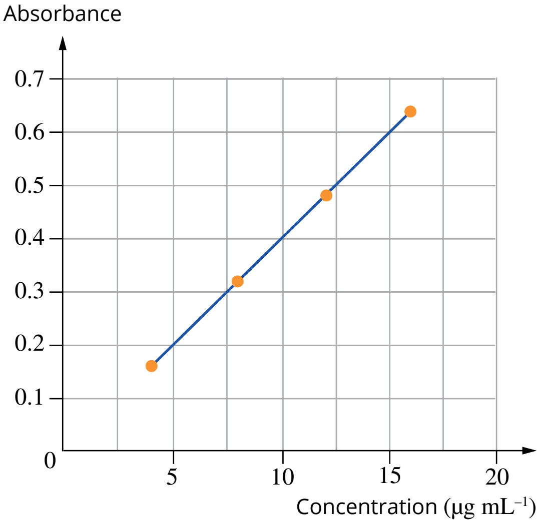
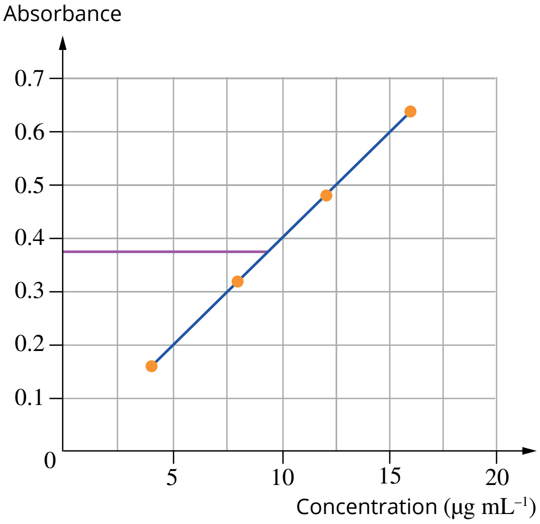
Section 18.4 Determining salt concentration by colorimetry and UV–visible spectroscopy

Worked example: Try yourself 18.4.1

USING A CALIBRATION CURVE

Determine the lead level in a solution using the following colorimetry data.

Concentration of Pb^{2+} (mg L^{-1})	**Absorbance**
2.5	0.18
5.0	0.35
7.5	0.51
10.0	0.68
Sample	0.60

Thinking	**Working**
Construct a calibration curve from the data above. Concentration will be on the horizontal axis.	**Calibration curve** 
Mark where the absorption of the sample lies on the calibration curve by tracing a horizontal line to the curve.	**Calibration curve** 

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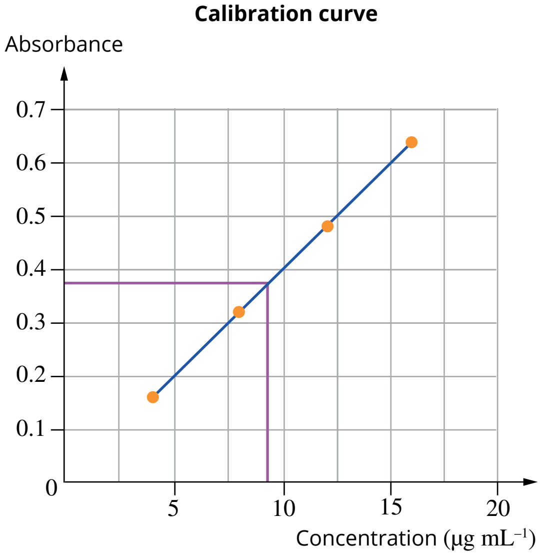
Draw a vertical line from the absorbance value to the *x*-axis.	
Determine the concentration of the sample by reading the concentration value for the unknown solution from the *x*-axis.	The concentration of Pb^{2+} in the sample can be determined from the graph as 8.8 mg L^{-1}.

18.4 Key questions

1 **a** 0.080 M solution
  b 0.30 M solution

2 Copper sulfate is blue because it transmits blue light and absorbs light of other frequencies. Since a colorimeter measures the amount of light absorbed by a sample, light of a colour other than blue must be used when measuring the concentration of a copper sulfate solution.

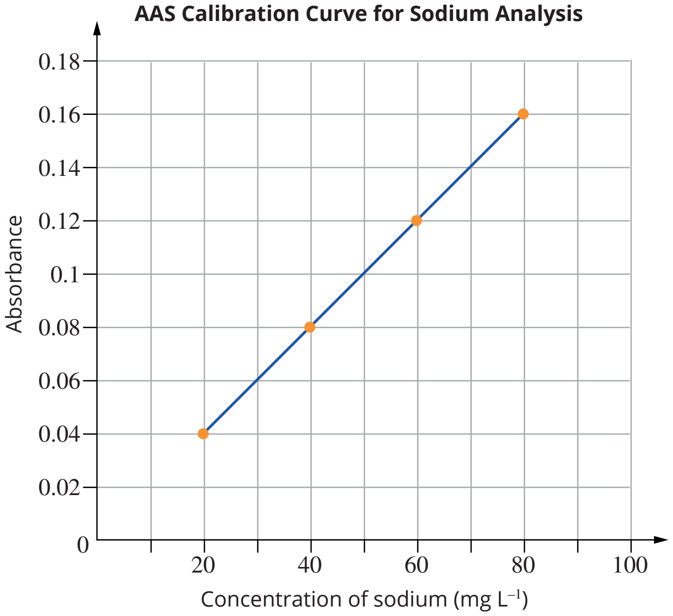
3 Blue. The sample is absorbing light with wavelengths of 600–700 nm. This means that the sample is absorbing red and yellow light but not blue light. If the sample is not absorbing blue light, the dye will observed to be blue. Blue is the complement of red and yellow light.

4 **a** An absorbance of 0.24 corresponds to a concentration of 0.15 M Fe^{2+} from the calibration graph.
  b Allowing for the dilution of the sample, the original solution was diluted by a factor of 4 so the concentration of Fe^{2+} ions in the tank water was 0.15 × 4 = 0.6 M.
  c Fe^{2+} ions are oxidised to Fe^{3+} with acidified permanganate (MnO_4^-/H^+) before the addition of KSCN. Reaction of KSCN with Fe^{3+} forms the highly coloured red metal complex FeSCN^{2+}. The concentration of the complex can then be determined using spectrophotometry.
  $Fe^{3+}(aq) + SCN^-(aq) \rightarrow FeSCN^{2+}(aq)$

5

Solution	**Complementary colour**
Green	Purple
Purple	Green
Orange	Blue
Green–yellow	Violet
Blue-green	Red
Green	Purple

6 **a**



b Concentration = 175 mg L^{-1}

Mark where the absorption of 0.42 of the sample lies on the calibration curve by tracing a horizontal line to the curve. Then drawing a vertical line from the curve to the horizontal axis. Read the concentration from the *x*-axis.

7 Blue is the complementary colour of orange. If potassium dichromate solutions are analysed in a colorimeter, a blue filter should be used. The sample will be placed in a glass or plastic cell and light directed at it. The higher the concentration of the solution, the more light is absorbed and the lower the intensity of the light arriving at the detector.

8 **a** There are two strong absorption peaks at 420 nm and 660 nm.

b Either 660 nm or 420 nm wavelengths could be used as chlorophyll absorbs strongly at both.

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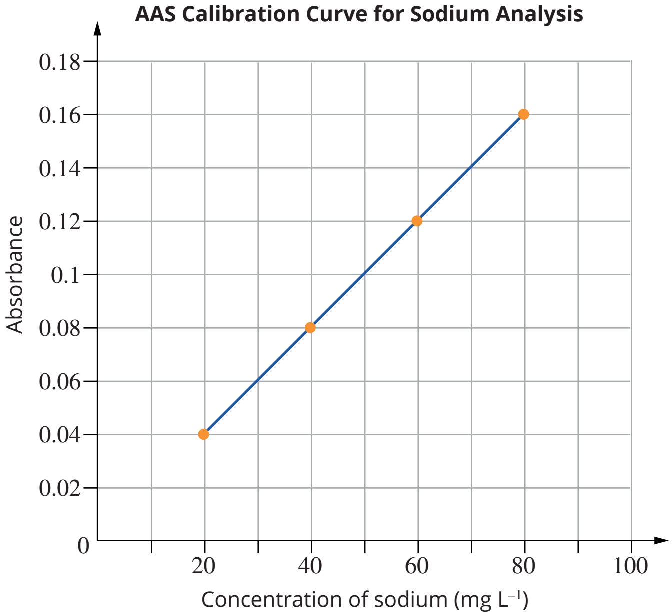
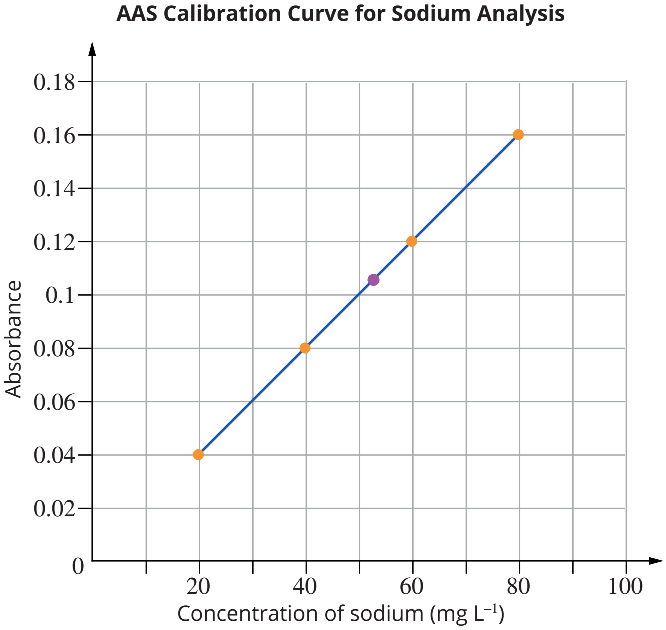
Section 18.5 Determining concentration by atomic absorption spectroscopy

Worked example: Try yourself 18.5.1

USING A CALIBRATION CURVE TO DETERMINE CONCENTRATION

Determine the concentration of sodium in a sample given the data in the following table.

Concentration of sodium (mg L^{-1})	20	40	60	80	Sample
Absorbance	0.041	0.080	0.121	0.159	0.104

Thinking	**Working**
Construct a calibration curve using the concentrations of the standard solutions and their absorbance.	
Mark where the absorption of the sample lies on the calibration curve. This can be done by finding the absorption of the sample on the *y*-axis and moving horizontally right until you reach the calibration curve and marking that point.	 The sample's absorption is 0.104.
Determine the concentration of the sample by reading off the graph on the *x*-axis at the point where the sample's absorption lies on the calibration curve.	From the graph it can be determined that the concentration of sodium in the sample is 52 mg L^{-1}.

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18.5 Key questions

1 When a sample containing copper is heated in the flame of a Bunsen burner, the flame turns a green colour. This is because the **electrons** in the copper atoms absorb energy and move to **higher** energy shells and then **emit** light that corresponds to a green colour as they return to **lower** energy shells.

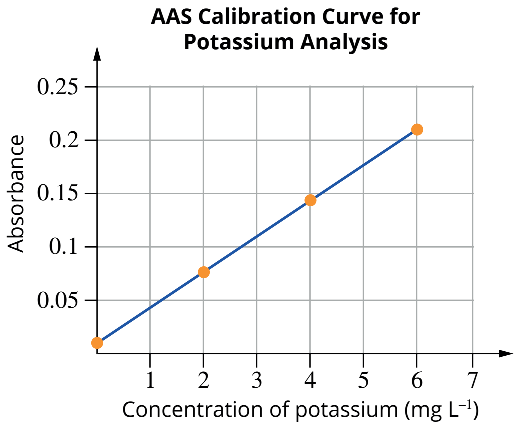
2 An emission spectrum is the spectrum of the light emitted by an element when it is heated. Atoms must be heated so that electrons move to higher energy levels, before returning to lower levels and emitting light.

3

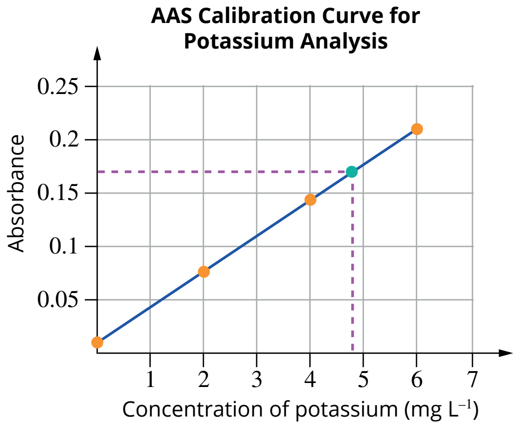
Component	**Description**
Monochromator	Selects a specific wavelength of light
Detector	Measures the amount of light
Hollow source cathode lamp	Produces light with wavelengths that are absorbed by the metal being analysed
Flame	Where the sample is sprayed and light is absorbed
Computer	Converts the amount of light detected into the amount of light absorbed by the sample

4 C. AAS does not heat a sample to a sufficient temperature to analyse all elements. It can, however, be used to analyse most metals but not non-metals.

5 **a**



b From the graph the concentration of potassium in the sample is 4.8 mg L^{-1}.



CHAPTER 18 REVIEW

18.1 THE SOURCE OF SALTS IN WATER

1 **a** Any three of the following human activities lead to an increase in salts in our waterways: mining and industry lead to an increase of mercury and other heavy metals in the water system, the burning of fuels containing trace amounts of mercury, leaching from landfill, incorrect disposal of batteries and fluorescent light bulbs.

Run off from agriculture increases ammonium and phosphate levels due to fertilisers. Sewage treatment and domestic grey water increase levels of phosphate.

b They can be removed by precipitation reactions. Addition of calcium hydroxide will result in the precipitation of insoluble heavy metal hydroxides.

Methylmercury (CH_3Hg^+) is on organometallic compound produced in industry and from the burning of mercury-containing substances.

Methylmercury and tetraethyl lead are organometallic compounds because they contain at least one bond between a carbon atom and the metal.

2 Salt solutions will conduct electricity due to the presence of *ions*. The higher the concentration of the ions, the *higher* the electrical conductivity. Conductivity readings are taken at a standard temperature of *25°C*. Electrical conductivity readings *cannot* be used to identify the actual metal ions present in a natural water sample.

3 B. A 0.1 M solution of $LiNO_3$ will be more conductive than a 0.001 M solution of KCl. Ethanoic acid and ammonia do not ionise completely, so the number of ions in 0.1 M solutions will be less than in a 0.1 M solution of $LiNO_3$.

4 Another example of an organometallic compound is methylmercury. Organometallic compounds are substances with at least one carbon–metal bond.

18.2 GRAVIMETRIC ANALYSIS

5

$Ca(NO_3)_2$	Na_3PO_4	$Ca_3(PO_4)_2$	$NaNO_3$
27 mol	18	9.0	54
0.72	0.48 mol	0.24	0.44
0.54	0.36	0.18 mol	1.08
1.2	0.8	0.4	2.4 mol

6 **a** lead(II) chloride, lead(II) hydroxide, copper(II) hydroxide

b $Pb(NO_3)_2(aq) + CuCl_2(aq) \rightarrow PbCl_2(s) + Cu(NO_3)_2(aq)$

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

$Pb(NO_3)_2(aq) + Ba(OH)_2(aq) \rightarrow Pb(OH)_2(s) + Ba(NO_3)_2(aq)$

$Pb^{2+}(aq) + 2OH^-(aq) \rightarrow Pb(OH)_2(s)$

$CuCl_2(aq) + Ba(OH)_2(aq) \rightarrow Cu(OH)_2(s) + BaCl_2(aq)$

$Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$

7 $n(KOH) = \frac{2}{1} \times n(Fe(OH)_2)$

$n(FeSO_4) = \frac{1}{2} \times n(KOH)$

$n(KOH) = \frac{2}{1} \times n(K_2SO_4)$

$n(Fe(OH)_2) = \frac{1}{1} \times n(FeSO_4)$

8 $\%(Ag) = 18.3\%$

$n(AgCl) = \frac{0.169}{143.4} = 0.00118$ mol

$n(Ag) = n(AgCl) = 0.00118$ mol

$m(Ag) = n \times M = 0.00118 \times 107.9 = 0.127$ g

$\%(Ag) = \frac{0.127}{0.693} \times 100 = 18.3\%$

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9 Step 1: Write a balanced equation.

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

Step 2: Calculate the amount of AgCl.

$$n(AgCl) = \frac{0.693\ g}{143.32\ g\ mol^{-1}} = 0.004438\ mol$$

Step 3: From the equation, 1 mol of NaCl produces 1 mol of AgCl.

$$\frac{n(NaCl)}{n(AgCl)} = \frac{1}{1}$$

$n(NaCl) = 0.004438$ mol

Step 4: Calculate the mass of salt.

$m(NaCl) = 0.004438 \times 58.44 = 0.2594$ g

Step 5: Convert to percentage.

$$\%(NaCl) = \frac{0.2594}{8.45} \times 100 = 3.0698\%$$

Step 6: Express the answer with the correct number of significant figures.

$\%(NaCl) = 3.07\%$

10 **a** Step 1: Find the percentage of P in $Mg_2P_2O_7$.

$$\%(P) = \frac{2 \times 30.97}{222.56} \times 100 = 27.83\%$$

Step 2: Using the %(P) in $Mg_2P_2O_7$, calculate the mass of P in precipitate.

$m(P) = 27.83\%$ of 0.085

$$\frac{27.83}{100} \times 0.085 = 0.02366\ g$$

Step 3: Calculate the percentage (P) in 2.0 g of washing powder.

$$\%(P) = \frac{0.02366\ g}{2.0\ g} = 1.183\%$$

Step 4: Express the answer with the correct number of significant figures.

$\%(P) = 1.2\%$

b You would need to consider the fact that the percentage of phosphate in the washing powder is greater than the percentage of the phosphorus.

11 The mass of the precipitate is higher than it should be. The experimental errors that could lead to a high result include the precipitate was not dried completely, the precipitate was not washed, the presence in the solution of other metal ions that form precipitates.

18.3 CALCULATIONS INVOLVING EXCESS REACTANTS

12

Carbon atoms available	Oxygen molecules available	Carbon dioxide molecules produced	Carbon atoms in excess	Oxygen molecules in excess
8	20	8	0	12
1000	3000	1000	0	2000
9 mol	6 mol	6 mol	3 mol	0 mol

13 **a** Oxygen is in excess.

b 1.5 mol of sodium oxide is produced.

$$\frac{n(O_2)}{n(Na)} = \frac{1}{4}$$

Since the $n(O_2)$ required is , oxygen is in excess and sodium is the limiting reagent.

Next, calculate the number of mole of sodium oxide from the number of mole of sodium.

$$\frac{n(Na_2O)}{n(Na)} = \frac{2}{4} = \frac{1}{2}$$

$$n(Na_2O) = 3 \times \frac{1}{2} = 1.5\ mol$$

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14 **a** $m(S) = 22.6$ g

$n(H_2S) = \frac{16}{34} = 0.471$ mol

$n(SO_2) = \frac{20}{64} = 0.313$ mol

From the balanced equation the number of mole of H_2S needs to be double that of SO_2. It is not; therefore, H_2S is the limiting reactant.

$n(S) = \frac{3}{2} \times 0.471 = 0.707$ mol

$\therefore m(S) = n \times M = 0.707 \times 32 = 22.6$ g

b $m(SO_2) = 4.93$ g

This is calculated by following the steps below.

$n(SO_2 \text{ reacting}) = \frac{1}{2} \times 0.471 = 0.236$ mol

$n(SO_2 \text{ remaining}) = 0.313 - 0.236 = 0.077$ mol

$\therefore m(SO_2) = n \times M = 0.077 \times 64 = 4.93$ g

15 **a** $n(P_4O_6) = \frac{4.40}{220.0} = 0.0200$ mol

$n(I_2) = \frac{3.00}{253.8} = 0.01182$ mol

Use $n(P_4O_6)$ provided to find $n(I_2)$ needed to react completely with the P_4O_6. From the equation, 8 mol of I_2 reacts with 5 mol of P_4O_6.

$\frac{n(I_2)}{n(P_4O_6)} = \frac{8}{5}$

$n(I_2) \text{ required} = \frac{8}{5} \times n(P_4O_6) = \frac{8}{5} \times 0.0200 = 0.0320$ mol

$n(I_2)$ available = 0.01182 mol

So, there is insufficient I_2 to react with all the P_4O_6, so I_2 is the limiting reagent and P_4O_6 is in excess.

Use $n(I_2)$ provided to find $n(P_4O_6)$ needed to react completely with the equation, 8 mol of I_2 reacts with 5 mol of P_4O_6.

$n(P_4O_6) \text{ required} = \frac{8}{5} \times n(I_2) = \frac{8}{5} \times 0.01182 = 0.00738$ mol

$n(P_4O_6)$ available = 0.0200 mol

So, P_4O_6 is in excess by $0.0200 - 0.00738 = 0.0126$ mol

$\therefore m(P_4O_6) \text{ in excess} = 0.0126 \times 220 = 2.77$ g

b Use the amount of the limiting reagent, I_2, to determine the amount of P_2I_4 formed. From the equation, 8 mol of I_2 reacts to form 4 mol of P_2I_4.

$n(P_2O_4) = 4$

$n(I_2) = 8$

So, $n(P_2I_4) = \frac{4 \times 0.01182}{8} = 0.00591$ mol

$\therefore m(P_2I_4) = 0.00591 \times 569.6 = 3.37$ g

c Use the amount of the limiting reagent, I_2, to determine the amount of P_4O_{10} formed. From the equation, 8 mol of I_2 reacts to form 3 mol of P_4O_{10}.

$n(P_4O_{10}) = 3$

$n(I_2) = 8$

So, $n(P_4O_{10}) = \frac{3 \times 0.01182}{8} = 0.00443$ mol

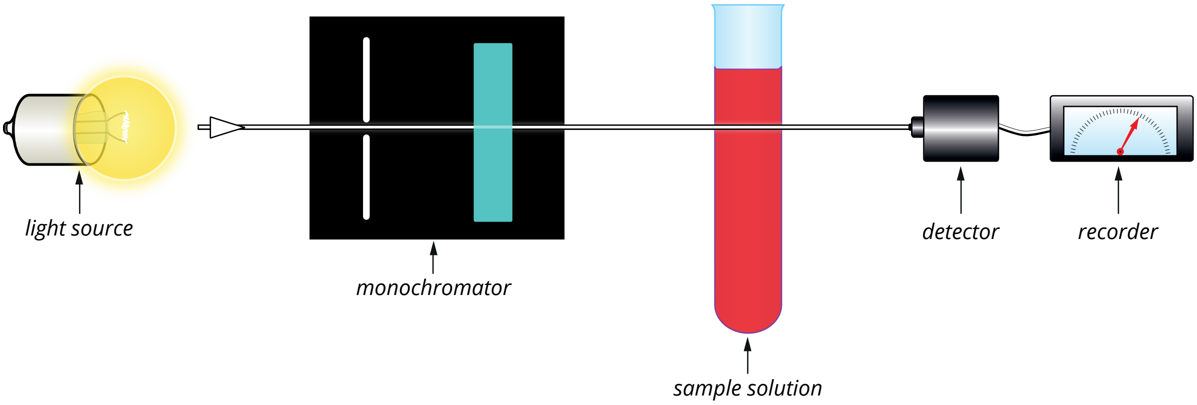
$\therefore m(P_4O_{10}) = 0.00443 \times 284.0 = 1.26$ g

d Total mass of products = 3.37 + 1.26 = 4.63 g

4.63 g of products formed plus 2.78 g unreacted P_4O_6 = 7.40 g, which is consistent with the total mass of reactants used initially.

18.4 DETERMINING SALT CONCENTRATION BY COLORIMETRY AND UV-VISIBLE SPECTROSCOPY

16



17 **a** 800 nm

b Light is absorbed in the red region.

c The solution would be blue.

18

Component	Function
Detector	Measures the intensity of light.
Filter	Selects a range of wavelengths of light.
Light source	Provides ultraviolet and visible light of all wavelengths.
Monochromator	Allows selection of light of a particular wavelength.
Glass cell	Transparent container that holds the sample.
Recorder	Displays the absorbance measurement.

19 **a** 39 mg L^{-1}

b Step 1: Using the answer from part a, calculate the mass of P in the 0.250 g sample of detergent powder, assuming that the density of the solution is 1 g mL^{-1}.

32 mg L^{-1} is 32 mg per litre (1000 mL) = ? mg per 250 mL solution

$$m(\mathrm{P}) = 32 \times \frac{250}{1000} = 8.0 \text{ mg}$$

Step 2: Calculate the percentage of P in the 0.250 g sample of detergent powder.

$$\%(\mathrm{P}) = \frac{8 \text{ mg}}{250 \text{ mg}} \times 100 = 3.2\% \text{ (to 2 significant figures)}$$

c Orange light of wavelength 600 nm is strongly absorbed by a blue solution.

20 **a** False

b False

c True

d False

18.5 DETERMINING CONCENTRATION BY ATOMIC ABSORPTION SPECTROSCOPY

21 **a** Copper gives a green flame.

b It is necessary vaporise traces of other substances that could lead to a false result.

c Copper wire would give the flame a green colour, and so interfere with the test. The wire must be made of an element that does not change the colour of the Bunsen burner flame.

d Flame tests are rarely used for analysis by today's chemists because relatively few elements may be analysed in this way and the presence of one element (sodium, for example) may mask the presence of another. Also colours emitted by some elements are similar and difficult and distinguish, and because a relatively large amount of sample is needed to give a clear colour.

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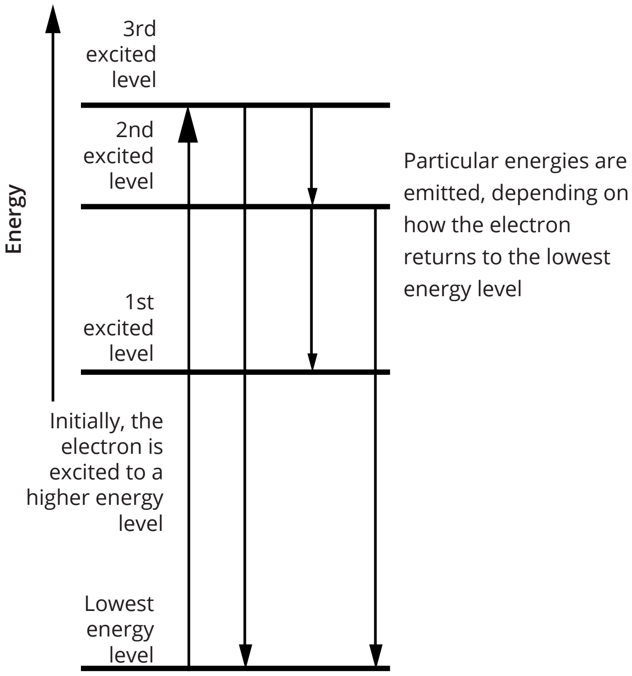
22 When atoms absorb energy, it is often possible for electrons to be promoted to various higher energy levels. Electrons can return to the ground state from these excited states by undergoing a number of transitions of different energy. (This is shown in the diagram.) Each transition results in a line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



23 Absorption occurs in the flame.

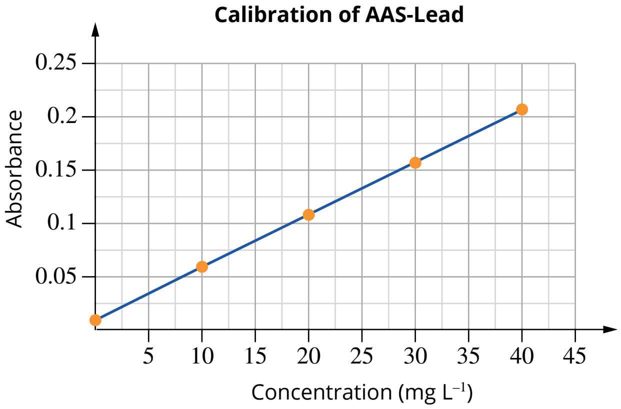
The sample is sprayed into the flame and the light from the lamp is passed through the flame. The metal atoms in the sample absorb some of the light passing through the flame. The higher the concentration of the metal, the more light is absorbed in the flame.

24 **a** Both involve the absorption of electromagnetic radiation of particular wavelengths characteristic of the substance under investigation. The amount of radiation absorbed is measured by a detector. The amount of light absorbed is proportional to the amount of the light-absorbing substance in the sample.

b In AAS, the sample is sprayed into a flame and light of a particular wavelength is passed through the flame. Atoms of the element being analysed absorb some of the radiation and the amount of light absorbed indicates the amount of the element present in the sample. AAS can be used to detect most metals.

In UV–visible spectroscopy, two lights are used to produce radiation covering the visible and UV spectrum. The ‘light’ is passed through a prism to produce the desired wavelength; for UV analysis a silica sample holder is used. UV spectroscopy can be used to analyse many colourless compounds as well as coloured ones. It can give information about the structure of the substance. UV–visible spectroscopy can be used to determine the concentration of atom ion and molecules include complex organic substances.

25 **a**



b From the graph, the concentration of lead in the sample is 26 mg L^{-1}.

CONNECTING THE MAIN IDEAS

26

Technique	**Feature**
Electrical conductivity	Ions flow between electrodes in a solution.
Atomic emission spectroscopy	A sample is injected into a flame.
Colorimeter	A filter is used to select a range of wavelengths of visible light.
UV–visible spectrophotometer	A monochromator allows complementary light of an exact wavelength to be used.
Gravimetric analysis	The mass of a precipitate is used to determine the concentration of a salt.

27 **a** $Al_2(SO_4)_2(aq) + 3Ba(NO_3)_2(aq) \rightarrow 3BaSO_4(s) + 2Al(NO_3)_2(aq)$

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

b n = 0.0161 mol

$n(BaSO_4) = \frac{3.76}{233.4} = 0.0161$ mol

$n(SO_4{}^{2-}) = n(BaSO_4) = 0.0161$ mol

c $n(Al_2(SO_4) = 0.00403$ mol

$n(Na_2SO_4) = 0.00403$ mol

The total number of moles of sulfate is 0.0161. Since $Al_2(SO_4)_3$ and Na_2SO_4 are present in equimolar amounts, the number of moles of sulfate from the $Al_2(SO_4)_3$ is three times that of the Na_2SO_4.

$n(Na_2SO_4) = \frac{0.0161}{4} = 0.00403$ mol

$n(Al_2(SO_4) = 3 \times 0.00403 = 0.0121$ mol

28 **a** Fe_2O_3

b $n(Fe) = 2\ n(Fe_2O_3) = 2 \times 2.4 = 4.8$ mol

c No. All compounds containing sodium ions are soluble in water.

29 **a** Use a green filter. (A complementary colour to red needs to be used in the calorimeter.)

b Mole ratio will be $\frac{1}{1}$.

c The observance obtained from a sample needs to be plotted on a calibration curve constructed from a series of standard solutions of the $FeSCN^-$ complex.

d The lamp used must match the element being tested. A hollow cathode lamp with an iron (Fe) filament should be used.

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Chapter 19 Analysis for organic compounds in water

Section 19.1 Organic contaminants in water

19.1 Key questions

1 A and B. Persistent organic pollutants (POPs) are organic contaminants that are resistant to environmental degradation. Heavy metals, minerals and organometallic compounds are not regarded as organic contaminants. Crude oil is not regarded as a POP because the alkanes in crude oil break down more rapidly in the environment than DDT or dioxins.

2 A, D and E. POPs can remain in the body for a long time. They impact upon the nervous system and reproductive organs. They are also linked to cancer.

3 B. PVC is a polymer called polyvinylchloride. It contains chlorine atoms, one of the elements present in most dioxins.

4 D. Dioxins are not soluble in water but they are soluble in fat. Blood and urine are both water-based solutions.

5 B. Oil and water do not mix. This leads to oil forming slicks when it spills into the ocean.

6 Dioxins can be formed from the *burning* of organic chemicals containing *chlorine* atoms. The dioxins can end up in *water supplies*. They are usually *insoluble* in water, so they become attached to plants or sediments. When consumed by an animal they are usually stored in *fat* They can have *adverse* effects on your health.

7 C, D and F. Ethanol and ethanoic acid are regularly consumed by humans, so small quantities in waste water are acceptable as they dissolve in water and are easily degraded. Methanol is also water soluble and decomposes relatively quickly in the environment.

Section 19.2 Principles of chromatography

Worked example: Try yourself 19.2.1

CALCULATING R_f VALUES

Calculate the R_f value of the blue component in Figure 19.2.5 on page 499.	
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

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19.2 Key questions

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

2

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

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

Band	**a Distance from origin (mm)**	**b R_f**	**c Compound**
Light green	20	0.33	Chlorophyll b
Dark green	27	0.45	Chlorophyll a
Orange	40	0.67	Xanthophyll
Yellow	50	0.83	ß-Carotene
Solvent front	60	–	–

d The chromatogram would be likely to be different because separation of components depends on their solubility in the mobile phase (as well as strength of adsorption to the stationary phase). The polarity of the solvent used in TLC and paper chromatography will affect the Rf of the sample components. A polar solvent will dissolve polar samples readily; a non-polar solvent will dissolve nonpolar 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. If a spot from the sample appears at the same distance from the origin as the spot from the standard solution, the tablet is likely to contain phenacetin.

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Section 19.3 High-performance liquid chromatography

Worked example: Try yourself 19.3.1

CONCENTRATION OF A PESTICIDE

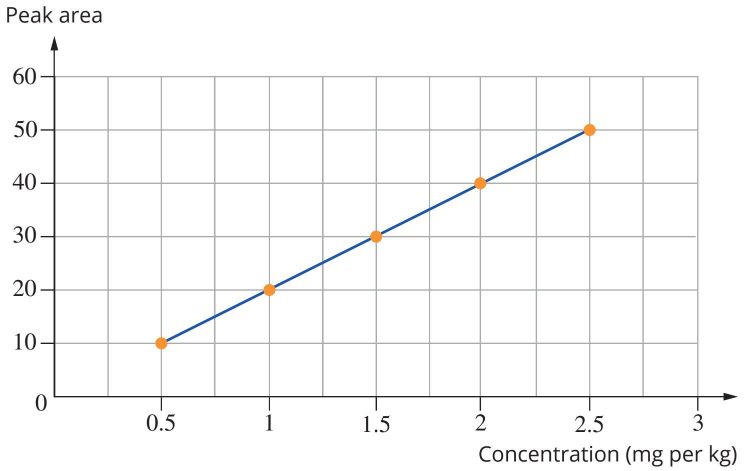
Procymidone is a pesticide that is used to prevent disease in oranges. Tests have been performed using HPLC to determine the concentration of this pesticide in a sample of orange juice.

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

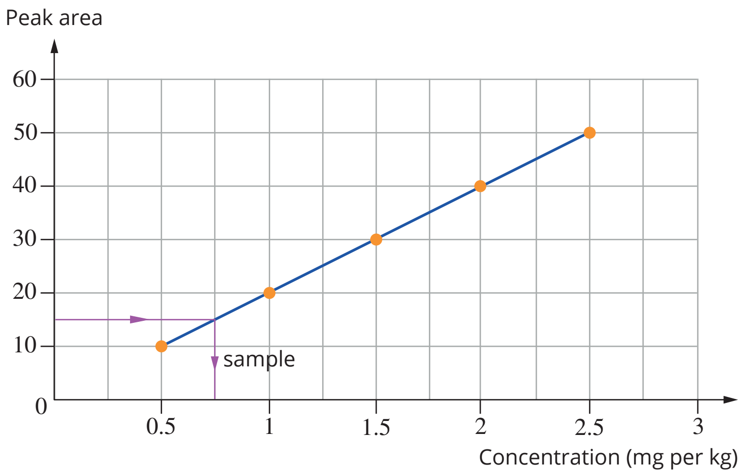
The results from the chromatograms of the sample and the standards are shown in the following table.

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

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

<table>
<tr><th>Thinking</th><th>Working</th></tr>
<tr><td>Construct a calibration graph by plotting peak area versus concentration.</td><td>


Calibration graph for concentration of procymidone in orange juice sample.
</td></tr>
<tr><td>Mark where the peak area of the sample lies on the calibration curve by tracing a horizontal line to the curve and then a vertical line down to the <i>x</i>-axis.</td><td>
Calibration graph for procymidone in orange juice



Calibration graph for procymidone sample.
</td></tr>
</table>

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Determine the concentration of the sample by reading the concentration value for the unknown solution from the *x*-axis.	The concentration of procymidone in the orange juice sample is 0.75 mg kg^{-1}.

19.3 Key questions

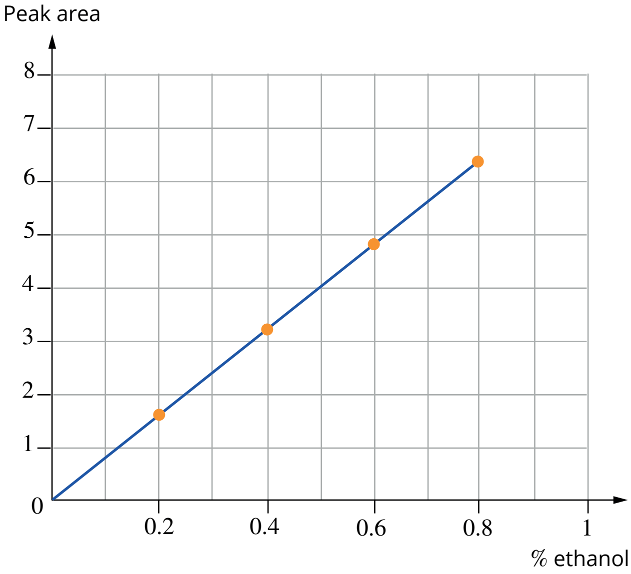
1 A

2 A

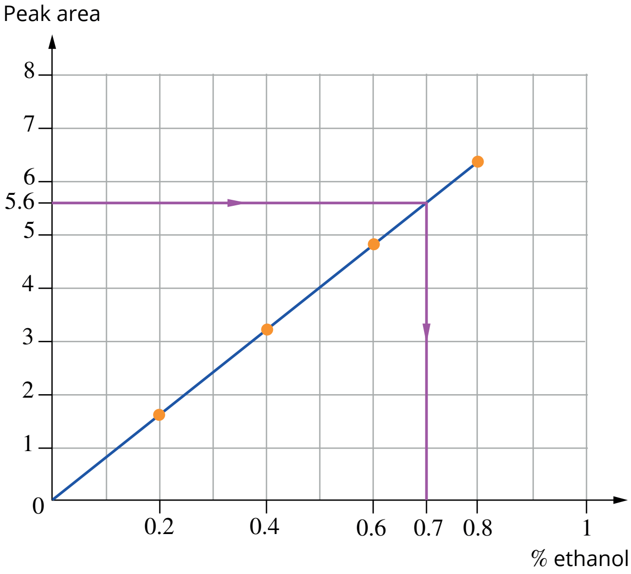
3 During column chromatography, the components of the sample *adsorb* onto the stationary phase and *desorb* into the liquid mobile phase. A component that adsorbs most strongly to the *stationary* phase and is least soluble in the *mobile* phase would be expected to take the *longest* time to pass through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

4 B, A, C

5 **a** Calibration graph for ethanol sample



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



CHAPTER 19 REVIEW

19.1 ORGANIC CONTAMINANTS IN WATER

1 A. Organometallic compounds.

Organometallic compounds contain metal ions bonded to the carbon atom of an organic compound. As it is the presence of the metal that is largely responsible for their action as a contaminant, they are not usually described as organic contaminants.

2 B. Non-polar contaminants will deposit in the fatty regions of the steak

3 H, D, G, B, E, A, C, F

4

DDT kills mosquitoes.	True
DDT degrades quickly in the environment.	False
DDT is toxic to humans.	True
It is illegal to use DDT in Australia.	True
DDT is an ionic compound.	False

DDT is now banned in most Western countries. Its toxicity is high and it remains active in the environment for a long time.

5 Whale, large fish, small fish, zooplankton, aquatic plants

The whale is likely to have the highest level of contaminants due to bioaccumulation. The levels will decrease with the size of the fish. The aquatic plants are likely to have the lowest levels of contaminants.

6 B. DDT – Dioxins and many other POPs contain chlorine.

7 To kill fungi growing on a tree, the tree can be sprayed with a *fungicide*. The chemical remains active even after the fungi have been killed. The term used to describe the time taken for 50% of the fungicide to degrade is known as the *half-life* of the chemical. Some of the chemical may wash into streams and lakes. It will adhere to sediment or plants if it is *insoluble* in water. When animals swallow the chemical, it tends to accumulate in *fat*.

8 Beneficial: produces energy; reduces landfill volume

Adverse: produces dioxins; produces carbon dioxide

Incinerators are used to remove unwanted household and industrial waste; however, the process produces carbon dioxide (a greenhouse gas) and dioxins. The dioxins are toxic and can accumulate in the fatty tissue of animals, where they take a long time to break down to less harmful molecules.

19.2 PRINCIPLES OF CHROMATOGRAPHY

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

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

10 **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.

11 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.

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

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

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

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

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

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

13 **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 3

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

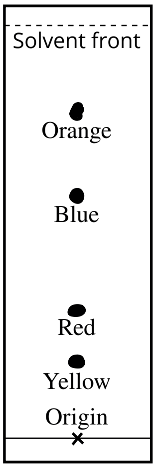
e blue

f 0.63; 0.13

14 **a** 3.2 cm

b 15 cm

c



15 **a** Taurine, glycine and an unknown. Two of the three spots produced by the medicine match those produced by taurine and glycine. The third spot does not match any of the amino acid standards and represents an unknown substance.

b There are several ways of visualising the spots:

- Viewing the chromatogram under UV light
- Spraying the finished chromatogram with a compound that causes the amino acids to fluoresce
- Spraying the finished chromatogram with sulfuric acid that causes a brown spot for organic compounds.

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

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

19.3 HIGH-PERFORMANCE LIQUID CHROMATOGRAPHY

16 The component most strongly adsorbed to the stationary phase is band A.

As band B begins to emerge from the column, it appears to separate into 2 bands. These two components would be more readily separated if you *increased the column length.*

HPLC differs from column chromatography because the particles in the stationary phase are *smaller* and *high* pressure is applied to the mobile phase.

17 A. HPLC does not directly produce measures on concentration. Standard solutions are used to find the relationship between peak area and concentration.

18 A. All forms of chromatography are based on the adsorption of the component particles to the stationary phase and desorption back into the mobile phase.

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19 B Obtain chromatograms of the stream water and of standard solutions of the herbicide.

D Use R_t values to identify the herbicide peak on the chromatogram of stream water.

A Measure the herbicide peak areas.

C Construct a calibration graph and mark the herbicide peak area on it.

E Determine the herbicide concentration in the stream water sample.

20 **a** Chemists can determine the number of components and the absorption of components, from which concentration can be determined.

b 4

c Solid samples are dissolved in a suitable solvent. The liquid sample is injected into the top of an HPLC column. The stationary and mobile phases are chosen to achieve a good separation of the components in the sample. The sample components alternately adsorb onto the stationary phase and then desorb into the solvent as they are swept forward. The time taken to exit the column increases if the component strongly absorbs onto the stationary phase and has a low solubility in the mobile phase.

CONNECTING THE MAIN IDEAS

21 B. The results obtained from HPLC would provide identification and accurate concentrations of the dioxins present.

22 **a** 1 mL: mass of dioxin = $3.0 \times 10^{-8} \times 0.001 = 3.0 \times 10^{-11}$ mg

b 600 mL water bottle: mass of dioxin = $3.0 \times 10^{-8} \times 0.600 = 1.8 \times 10^{-8}$ mg

c 750 ML reservoir: mass of dioxin = $3.0 \times 10^{-8} \times 750\,000\,000 = 22.5$ mg

23

Technique	Problem
Paper or thin-layer chromatography	B Difficult or impossible to obtain quantitative data D Samples must be able to be dissolved in solvent
High performance chromatography (HPLC)	A Requires large amounts of solvents to operate C Expensive equipment needed D Samples must be able to be dissolved in solvent

24 Most components of crude oil are insoluble in water and have a lower density than water. Therefore they form a slick on top of the water that spreads from the source.

Booms are used to contain the slick wand to protect the coastline. Some of the more volatile components will evaporate. Other components can be ‘skimmed’ from the water surface.

Dispersants are used on larger spills to try and break up the slick. Oil reaching the coast often has to be collected manually using shovels and buckets.

Damage is done to marine and coastal environments and to animals that are in contact with the oil.

25 **a**

Standard	Peak height (cm)
Standard 0 ppm parathion	0.2
Standard 10 ppm parathion	1.5
Standard 20 ppm parathion	3.1
Standard 30 ppm parathion	4.4
Suspected contaminated water	2.0

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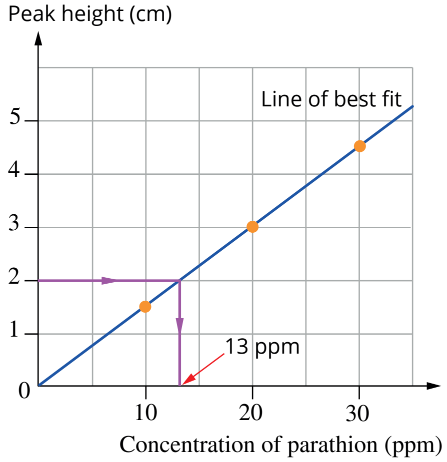
b



c Concentration of parathion in water sample = 13 ppm

d 13 ppm ≡ 13 μg/mL ≡ 13 mg/L

No, this is over 1000 times the allowable concentration.

e LD_{50} of 8 mg/kg ≡ 8 mg/1000 g × 150 g = 1.2 mg. A mouse of 150 g has a 50% chance of being killed by a 1.2 mg dose of parathion.

Volume of water that contains 1.2 mg parathion = $\frac{1.2}{13}$ L = 0.092 L = 92 mL

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Chapter 20 Volumetric analysis

Section 20.1 Sources of acids and bases in water

20.1 Key questions

1 Calcium carbonate raises the pH of river water.

2 Nitrogen oxides produced by the reaction of nitrogen and oxygen from the air in internal combustion engines contribute to the formation of acid rain. Acid rain is more acidic than pure water (pH of 7); therefore, nitrogen oxides lower the pH.

3 $4FeS_2(s) + O_2(g) + 15H_2O(l) \rightarrow Fe(OH)_3(s) + 8H_2SO_4(aq)$

4 $CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + CO_2(g) + H_2O(l)$

Section 20.2 Standard solutions

Worked example: Try yourself 20.2.1

CALCULATING THE CONCENTRATION OF A STANDARD SOLUTION PREPARED FROM A PRIMARY STANDARD

<table>
<tr><td colspan="2">Calculate the concentration of a standard solution prepared from 117.0 g of NaCl dissolved in a 500.0 mL volumetric flask.</td></tr>
<tr><th>Thinking</th><th>Working</th></tr>
<tr><td>Use the chemical formula to determine the molar mass (<i>M</i>) of the compound.</td><td>The molar mass (<i>M</i>) of NaCl is
$23.0 + 35.5 = 58.5$ g mol^{-1}</td></tr>
<tr><td>Use the mass (<i>m</i>) and molar mass (<i>M</i>) of the compound and the formula $n = \frac{m}{M}$ to determine the amount, in mol.</td><td>$n = \frac{117.0}{58.5}$
$= 2.00$ mol</td></tr>
<tr><td>Use the amount, in mol, to determine the concentration of the solution using the formula $c = \frac{n}{V}$</td><td>$c = \frac{n}{V}$
$= \frac{2.00}{0.5000}$
$= 4.00$ M</td></tr>
</table>

20.2 Key questions

1 Weigh the solid primary standard on an electronic balance.
Transfer the solid into the volumetric flask using a clean, dry funnel.
Rinse any remaining solid particles into the flask using deionised water.
Half fill the flask with deionised water and swirl vigorously to dissolve the solid.
Fill the flask with deionised water to just below the calibration mark.
Add deionised water drop by drop up to the calibration line on the flask until the bottom of the meniscus touches the line.
Stopper and shake the solution to ensure an even concentration throughout.

2 $n = \frac{m}{M} = \frac{2.042}{204.2} = 0.01000$ mol

$c = \frac{n}{V} = \frac{0.01000}{0.050000} = 0.2000$ M

3 $c = \frac{n}{V} = 0.500 = \frac{n}{0.250}$

$n = 0.500 \times 0.2500 = 0.125$ mol

$m = n \times M = 0.125 \times 105.99 = 13.25$ g $= 13.3$ g

Section 20.3 Calculations involving acids and bases

Worked example: Try yourself 20.3.1

A SOLUTION VOLUME–VOLUME CALCULATION

What volume of 0.500 M hydrochloric acid (HCl) reacts completely with 25.0 mL of 0.100 M calcium hydroxide ($Ca(OH)_2$) solution? The salt formed in this acid–base reaction is calcium chloride.	
Thinking	**Working**
Write a balanced full equation for the reaction.	$Ca(OH)_2(aq) + 2HCl(aq) \rightarrow CaCl_2(aq) + 2H_2O(l)$
Calculate the amount, in mol, of the substance with known volume and concentration.	The volume and concentration of calcium hydroxide solution are given, so you use $n = cV$ (remember that volume must be expressed in litres). $n(Ca(OH)_2) = cV$ $= 0.100 \times 0.0250$ $= 0.00250$ mol
Use the mole ratio from the equation to calculate the amount, in mol, of the required substances.	The balanced equation shows that 2 mol of hydrochloric acid reacts with 1 mol of calcium hydroxide. $\frac{n(HCl)}{n(Ca(OH)_2)} = \frac{2}{1}$ $n(HCl) = \frac{2}{1} \times n(Ca(OH)_2)$ $= \frac{2}{1} \times 0.00250$ $= 0.00500$ mol
Calculate the volume or concentration required.	The volume of HCl is found by using $n = cV$. $V(HCl) = \frac{n}{c}$ $= \frac{0.00500}{0.500}$ $= 0.0100$ L $= 10.0$ mL

Worked example: Try yourself 20.3.2

SOLUTION STOICHIOMETRY: A LIMITING REACTANT PROBLEM

30.0 mL of a 0.100M H_2SO_4 solution is mixed with 40.0 mL of a 0.200M KOH solution. The equation for the reaction that occurs is: $H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$	
(a) Which reactant is the limiting reactant?	
Thinking	**Working**
Calculate the number of moles of each of the reactants using $n = cV$.	$n(H_2SO_4) = 0.100 \times 0.0300$ $= 0.00300$ mol $n(KOH) = 0.200 \times 0.0400$ $= 0.00800$ mol

Use the coefficients of the equation to find the limiting reactant.	The equation shows that 1 mole of H_2SO_4 reacts with 2 mole of KOH. So H_2SO_4 is the limiting reactant (it will be completely consumed).
(b) What will be the mass of K_2SO_4 produced by this reaction?	
Thinking	**Working**
Find the mole ratio of the unknown substance to the limiting reactant from the equation coefficients: $\frac{\text{coefficient of unknown}}{\text{coefficient of limiting reactant}}$	From the equation coefficients, $\frac{n(K_2SO_4)}{n(H_2SO_4)} = \frac{\text{coefficient of } K_2SO_4}{\text{coefficient of } H_2SO_4} = \frac{1}{1}$
Calculate the number of moles of the unknown substance using the limiting reactant: n(unknown) = n(limiting reactant) × mole ratio	$n(K_2SO_4) = n(H_2SO_4) \times \frac{1}{1}$ $= 0.00300 \times \frac{1}{1}$ $= 0.00300$ mol
Calculate the mass of the unknown substance using: m(unknown) = n(unknown) × molar mass	Molar mass of $K_2SO_4 = 174.3$ g mol^{-1} $m(K_2SO_4) = 0.00300 \times 174.3$ $= 0.523$ g

20.3 Key questions

1 **a** $H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$

b $n(KOH) = cV = 0.300 \times 0.0100 = 0.00300$ mol

$\frac{n(H_2SO_4)}{n(KOH)} = \frac{1}{2} n(H_2SO_4) = 0.00300 \times \frac{1}{2} = 0.00150$

$V(H_2SO_4) = \frac{0.00150}{0.100} = 0.0150$ L = 15.0 mL

2 **a** $2HNO_3(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_3)_2(aq) + 2H_2O(l)$

b $n(Ca(OH)_2) = cV = 0.100 \times 0.010 = 0.00100$ mol

$n(HNO_3) = 2 \times 0.00100$ mol $= 0.00200$ mol

$c(HNO_3) = \frac{n}{V} = \frac{0.00200}{0.0150} = 0.133$ M

3 **a** $n(H_2SO_4) = cV = 0.200 \times 0.0100 = 0.00200$ mol

b $n(Na_2CO_3) = cV = 0.100 \times 0.0160 = 0.00160$ mol

c According to the reaction, $n(H_2SO_4) = n(Na_2CO_3)$, therefore Na_2CO_3 is the limiting reactant and H_2SO_4 is in excess.

d $n(H_2SO_4)$ reacts $= n(Na_2CO_3) = 0.00160 = 0.00040$ mol

$n(H_2SO_4)$ excess $= n(H_2SO_4)$ total $- n(H_2SO_4)$ reacts

$= 0.00200 - 0.00160 = 0.00040$mol in excess

4 **a** $HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)$

b $n(KOH) = cV = 0.09927 \times 0.02000 = 0.001985$ mol

c $n(HNO_3) = n(KOH) = 0.001985$ mol

d $c(HNO_3) = \frac{n}{V} = \frac{0.001985}{0.01826} = 0.1087$ M

Section 20.4 Volumetric analysis

20.4 Key questions

1 A. The concentration of NaOH doubled. Consequently, the volume of NaOH solution required to react will be halved. A strong acid is reacting with a strong base so the pH will be 7.

2

Indicator	**Colour after adding indicator**
Alizarin yellow	Yellow
Bromothymol blue	Yellow
Methyl red	Orange
Methyl orange	Yellow
Phenolphthalein	Colourless

3 Trials 2, 3 and 5 provide concordant titres.

Average titre = $\frac{25.46 + 25.38 + 25.42}{3}$ = 25.42 mL

Section 20.5 Calculations in volumetric analysis

Worked example: Try yourself 20.5.1

SIMPLE TITRATION

The concentration of a solution of barium hydroxide ($Ba(OH)_2$) was determined by titration with a standard solution of hydrochloric acid.

A 10.00 mL aliquot of $Ba(OH)_2$ solution was titrated with a 0.125 M solution of HCl. Titres of 17.23 mL, 17.28 mL and 17.21 mL of HCl were required to reach the end point. What is the concentration of the barium hydroxide solution?

Thinking	**Working**
Write a balanced chemical equation for the reaction.	An acid is reacting with a metal hydroxide, or base, so the products will be a salt and water. $2HCl(aq) + Ba(OH)_2(aq) \rightarrow BaCl_2(aq) + 2H_2O(l)$
Determine the volume of the average titre.	Average titre = $\frac{17.21 + 17.23 + 17.28}{3}$ = 17.24 mL
Calculate the amount, in mol, of the standard solution that was required to reach the end point.	$n(HCl) = c \times V$ $= 0.125 \times 0.01724$ $= 0.002155$ mol
Use the mole ratio in the equation to calculate the amount, in mol, of the unknown substance that would have reacted with the given amount, in mol, of the standard solution.	mole ratio = $\frac{n(Ba(OH)_2)}{n(HCl)} = \frac{1}{2}$ $n(Ba(OH)_2) = \frac{1}{2} \times n(HCl)$ $= \frac{1}{2} \times 0.002155$ $= 0.001078$ mol
Determine the concentration of the unknown substance.	$c(Ba(OH)_2) = \frac{n}{V}$ $= \frac{0.001078}{0.01000}$ = 0.108 M
Express your answer to the appropriate number of significant figures.	The final result is rounded off to three significant figures corresponding to the smallest number of significant figures in the original data. Four significant figures are used in the earlier steps of the calculation to avoid rounding off errors. Concentration of the $Ba(OH)_2$ = 0.108 M

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Worked example: Try yourself 20.5.2

TITRATION THAT INVOLVES DILUTION

A commercial concrete cleaner contains hydrochloric acid. A 10.00 mL volume of cleaner was diluted to 250.0 mL in a volumetric flask. A 20.00 mL aliquot of 0.2406 M sodium carbonate solution was placed in a conical flask. Methyl orange indicator was added and the solution was titrated with the diluted cleaner. The methyl orange indicator changed permanently from yellow to pink when 18.68 mL of the cleaner was added. Calculate the concentration of hydrochloric acid in the concrete cleaner.	
Thinking	**Working**
Write a balanced chemical equation.	A dilute acid is reacting with a metal carbonate so the products will be a salt, water and carbon dioxide gas. $2HCl(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$
Using the concentration of the standard solution, calculate the amount, in mol, of the known substance that reacted in the titration. Remember that volume must be expressed in litres.	$n(Na_2CO_3) = cV$ $= 0.2406 \times 0.02000$ $= 0.004812$ mol
Use the mole ratio in the equation to calculate the amount, in mol, of the diluted unknown solution that reacted in the titration.	$\frac{n(HCl)}{n(Na_2CO_3)} = \frac{2}{1}$ $n(HCl) = \frac{2}{1} \times n(Na_2CO_3)$ $= \frac{2}{1} \times 0.004812$ $= 0.009624$ mol
Calculate the concentration of diluted unknown solution reacting in the titration.	V(diluted HCl) = 0.01868 L $c(HCl) = \frac{n}{V}$ $= \frac{0.009624}{0.01868}$ $= 0.5152$ M
Multiply by the dilution factor to determine the concentration of undiluted concrete cleaner.	Dilution factor $= \frac{250.0}{10.00} = 25.00$ So undiluted $c(HCl)$ = diluted $c(HCl) \times 25.00$ $= 0.5152 \times 25.00$ $= 12.88$ M

20.5 Key questions

1 $n(KOH) = cV = 1.34 \times 0.02298 = 0.03079$ mol

$\frac{n(H_2SO_4)}{n(KOH)} = \frac{1}{2}$, therefore $n(KOH) \times \frac{1}{2} = n(H_2SO_4) = 0.01540$ mol

$\frac{n}{V} = \frac{0.01540}{0.02000} = 0.770$ M

2 **a** $CH_3COOH(aq) + NaOH(aq) \rightarrow NaCH_3COO(aq) + H_2O(l)$

b $n(NaOH) = cV = 0.995 \times 0.02156 = 0.02145$ mol

c $n(NaOH) = n(CH_3COOH) = 0.02145$ mol

d $c(CH_3COOH) = \frac{n}{V} = \frac{0.02145}{0.02500} = 0.858$ M

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3 $n(K_2CO_3) = cV = 0.4480 \times 0.02000 = 0.008960$ mol

$\frac{n(HCl)}{n(K_2CO_3)} = \frac{2}{1}$, therefore $n(HCl) = n(K_2CO_3) \times \frac{2}{1} = 0.01792$ mol

$c(HCl) = \frac{n}{V} = \frac{0.01792}{0.01984} = 0.9032$ M of diluted sample

Dilution factor $= \frac{250.0}{25.00} = 10$

Concentration of undiluted $= 0.9032 \times 10 = 9.032$ M

CHAPTER 20 REVIEW

20.1 SOURCES OF ACIDS AND BASES IN WATER

1 B. Iron pyrite. Acid mine drainage is caused by the oxidation of sulfur, found in some minerals and coal, to sulfuric acid, which leaches toxic metal ions from soil and rocks.

2 The oxidation of iron pyrites produces a solution containing sulfuric acid and iron(III) hydroxide. The acid leaches heavy metals from the soils and rocks. As a consequence, the water discharged from the mine contains iron(III) hydroxide, sulfuric acid and heavy metal ions.

3 Acid rain is rainfall made acidic when pollutant gases generated by smelters, the internal combustion engine and the burning of fossil fuels dissolve in rain. These gases contain SO_2, NO and NO_2 and dissolve in rain to form sulfuric, nitrous and nitric acids respectively.

4 D. Both reactions occur in river water. Carbon dioxide is absorbed from the atmosphere and is also produced by decaying plants. A small amount of calcium carbonate reacts with water when a river flows through limestone. Because the pH > 7, there must be an excess of OH^- ions, so the second equation is occurring to a greater extent.

20.2 STANDARD SOLUTIONS

5 C. The colour of a compound is not usually considered when selecting a primary standard. Primary standards must be soluble because they react with the substance being analysed in solution. A high level of purity minimises errors in calculating the results of the analyses. A high molar mass minimises the percentage error in calculations.

6 $n = \frac{m}{M} = \frac{25.21}{126.0} = 0.20008$ mol

$c = \frac{n}{V} = \frac{0.20008}{0.2500} = 0.8003$ M

7 $c = \frac{n}{V}$, $0.400 = \frac{n}{0.500}$, $n = 0.400 \times 0.500 = 0.200$ mol

$n = \frac{m}{M}$, $m = n \times M$, $m = 0.200 \times 106.0 = 21.2$ g

20.3 CALCULATIONS INVOLVING ACIDS AND BASES

8 $HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)$

$n(HNO_3) = cV = 0.100 \times 0.0300 = 0.00300$ mol

$n(HNO_3) = n(KOH)$, therefore $n(KOH) = 0.00300$ mol

$V(KOH) = \frac{n}{c} = \frac{0.00300}{0.200} = 0.0150$ L = 15.0 mL

9 **a** $n(HCl) = cV = 1.00 \times 0.0200 = 0.0200$ mol

b $n(K_2CO_3) = cV = 1.00 \times 0.0160 = 0.0160$ mol

c $2HCl(aq) + K_2CO_3(aq) \rightarrow 2KCl(aq) + CO_2(g) + H_2O(l)$

$n(K_2CO_3 \text{ reacts}) = n(HCl) \div 2 = 0.0200 \div 2 = 0.0100$ mol, $n(K_2CO_3) = 0.0160$ mol

Therefore K_2CO_3 is in excess, so HCl is the limiting reactant.

d $n(K_2CO_3) - n(K_2CO_3 \text{ reacts})$ total $= 0.0160 - 0.0100 = 0.0060$ mol in excess

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10 $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

$n(HCl) = cV = 0.200 \times 0.02000 = 0.00400$ mol

$n(HCl) = n(NaOH)$, so $n(NaOH) = 0.00400$ mol

$V(NaOH) = \frac{n}{c} = \frac{0.00400}{0.100} = 0.0400\text{ L} = 40\text{ mL}$

11 **a** $n(LiOH) = cV = 2.00 \times 0.0150 = 0.0300$ mol

$n(HNO_3) = cV = 0.400 \times 0.0250 = 0.0100$ mol

According to the reaction equation, $n(LiOH) = n(HNO_3)$, therefore HNO_3 is the limiting reactant.

b $n(HNO_3) = n(LiNO_3) = 0.0100$ mol

$M(LiNO_3) = 68.9\text{ g mol}^{-1}$

$n = \frac{m}{M}$, therefore $m = n \times M = 0.0100 \times 68.98 = 0.690$ g

20.4 VOLUMETRIC ANALYSIS

12 **a** A burette delivers a precise variable volume of solution.

b A pipette delivers a precise known volume of solution.

c A standard flask is used to prepare a solution of known concentration.

13 The equivalence point is the point in the reaction at which equivalent amounts of acid and base have been mixed according to the ratio given in the chemical equation. The end point is the point in the titration at which the indicator changes colour.

14 Near the equivalence point, a very small addition of either the acid or the base in the burette can cause a large change in pH. A sharp end point is one where the indicator changes colour (due to a large change in pH) with just one additional drop of the solution being added from the burette. Selecting the indicator based on the expected equivalence point enables for a sharp end point to be seen.

20.5 CALCULATIONS IN VOLUMETRIC ANALYSIS

15 **a** $H_2SO_4(aq) + K_2CO_3(aq) \rightarrow K_2SO_4(aq) + H_2O(l) + CO_2(g) + K_2SO_4(aq)$

b $c(K_2CO_3) = \frac{m}{MV} = \frac{1.227\text{ g}}{138.0\text{ g mol}^{-1} \times 0.2500\text{ L}}$

$= 0.03557$ M

c From the equation, 1 mol of H_2SO_4 reacts with 1 mol of K_2CO_3 in the titration.

$\frac{n(H_2SO_4)}{n(K_2CO_3)} = \frac{1}{1}$

$n(K_2CO_3) = cV = 0.03556\text{ M} \times 0.02000\text{ L} = 0.0007113$ mol

$n(H_2SO_4) = 0.03557\text{ M} \times 0.02000\text{ L} = 0.0007113$ mol

$\therefore c(H_2SO_4) = \frac{0.0007113\text{ M}}{0.02256\text{ L}} = 0.0315$ M

16 **a** $2HCl(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$

b $c(Na_2CO_3) = \frac{m}{MV} = \frac{1.358\text{ g}}{106.0\text{ g mol}^{-1} \times 0.2500\text{ L}}$

$= 0.05125$ M

c From the equation, 2 mol of HCl reacts with 1 mol of Na_2CO_3 in the titration.

$\frac{n(HCl)}{n(Na_2CO_3)} = \frac{2}{1}$

$n(Na_2CO_3) = cV = 0.05125\text{ M} \times 0.02000\text{ L} = 0.001025$ mol

$n(HCl) = 2 \times 0.05125\text{ M} \times 0.02000\text{ L} = 0.002050$ mol

$\therefore c(HCl) = \frac{n}{V} = \frac{0.002050}{0.02024} = 0.1013$ M

d Sulfuric acid is a diprotic acid so only half as much acid would be required to neutralise an equivalent amount of sodium carbonate. Since the two acids have the same concentration, half the volume of sulfuric acid is required.

17 $2HNO_3(aq) + Na_2CO_3(aq) \rightarrow 2NaNO_3(aq) + H_2O(l) + CO_2(g)$

a $c(Na_2CO_3) = \frac{m}{MV} = \frac{1.104\ \text{g}}{106.0\ \text{g mol}^{-1} \times 0.2500\ \text{L}}$

$= 0.04166\ \text{M}$

b From the equation, 2 mol of HNO_3 reacts with 1 mol of Na_2CO_3 in the titration.

$\frac{n(HNO_3)}{n(Na_2CO_3)} = \frac{2}{1}$

$n(Na_2CO_3) = cV = 0.04166\ \text{M} \times 0.02000\ \text{L} = 0.0008332\ \text{mol}$

$n(HNO_3) = 2n(Na_2CO_3) = 2 \times 0.0008332 = 0.001\ 666\ \text{mol}$

$c(HNO_3) = \frac{n}{V} = \frac{0.001666\ \text{mol}}{0.02347\ \text{L}} = 0.007100\ \text{M}$

CONNECTING THE MAIN IDEAS

18 **a** Sodium carbonate or potassium carbonate could be used as they are both good primary standards. They are also both bases, so they will neutralise the hydrochloric acid.

b To prepare a standard solution:

1. Weigh the pure solid on an electronic balance.
2. Transfer the solid into the volumetric flask using a clean, dry funnel.
3. Rinse any remaining solid particles into the flask using deionised water.
4. Half fill the flask with deionised water and swirl vigorously to dissolve the solid.
5. Add deionised water up to the calibration line on the neck of the flask. The bottom of the meniscus of the solution should be on the mark when viewing at eye level.
6. Stopper and shake the solution to ensure an even concentration throughout.

c Rinse the burette with the acid solution, rinse the pipette with the base and rinse the conical flask with the deionised water.

d Reaction equation: $K_2CO_3(aq) + 2HCl(aq) \rightarrow 2KCl(aq) + CO_2(g) + 2H_2O(l)$

$n(K_2CO_3) = cV = 0.5000 \times 0.02500 = 0.01250\ \text{mol}$

$\frac{n(HCl)}{n(K_2CO_3)} = \frac{2}{1}$, therefore $n(HCl) = 0.01250 \times \frac{2}{1} = 0.02500\ \text{mol}$

$c(HCl) = \frac{n}{V} = \frac{0.02500}{0.01844} = 1.356\ \text{M}$

Dilution factor $= \frac{250.00}{25.00} = 10$

Concentration of undiluted = 1.356 × 10 = 13.56 M

e The concentration calculated shouldn't be different if using a different base in the titration. The volume of base titrated and the concentration of the base may be different and therefore the mole ratio in the equation might be different, but the concentration of HCl in the concrete cleaner would be the same. If there were errors evident in either titration then there would be slight inaccuracies in the results and potentially a different concentration calculated.

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Chapter 21 Practical investigation

Section 21.1 Designing and planning investigations

21.1 Key questions

1 **a** Electrical conductivity
  b Concentration of lead
  c Electrical conductivity
  d pH

2 A hypothesis is a suggested outcome, based on evidence and prior knowledge, to answer the research question. A hypothesis often takes the form of a proposed relationship between two variables.

3 **a** The pH of the water
  b The mussel shells dissolving
  c The difference in mass of the mussel shells before and after the experiment
  d The temperature of the water, the mass of mussel shells, the length of time the mussel shells are left in the water

Section 21.2 Conducting investigations and recording and presenting data

21.2 Key questions

1 17.34 mL, 17.38 mL, 17.44 mL

2 B. Using a control with distilled water would help to correctly calibrate the instrument. This would make subsequent measurements more reliable.

3 **a** Systematic error
  b Mistake
  c Random error

4 C. Using a calibrated pH meter at a particular temperature

5 The mean is most affected by an outlier. The median value is the middle data value (when values are ordered). The mode is the value that occurs most frequently. Consequently, these will not be affected by an outlier.

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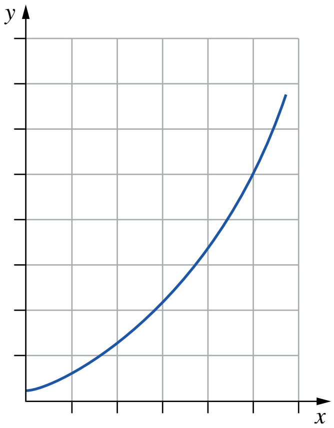
Section 21.3 Discussing investigations and drawing evidence-based conclusions

21.3 Key questions

1 

2 The trend in this graph is inversely non-linear. As the temperature of the water increases, the solubility of oxygen decreases markedly. The graph suggests that to increase the solubility of oxygen gas in water, the temperature of the water should decrease.

3 **a** Reliability
  b Validity
  c Accuracy
  d Precision

4 The conclusion stated is not directly related to the hypothesis—the conclusion should have referred to the temperature of the water and its electrical conductivity.

CHAPTER 21 REVIEW

21.1 DESIGNING AND PLANNING INVESTIGATIONS

1 A Concentration of lead – dependent variable
  B Analytical technique, temperature of water sample, type of sampling container – controlled variable
  C Source and location of water – independent variable

2 Independent variable = source of the water, dependent variable = phosphate concentration, controlled variables = temperature, time of testing, method of testing

3 Accuracy refers to the ability of the method to obtain the correct measurement close to a true or accepted value. Validity refers to whether an experiment or investigation is in fact testing the set hypothesis and aims

4 **a** It can dissolve or eat away at substances including tissues such as your skin or lungs.
  b It Is toxic (poisonous) if inhaled.
  c it is a highly combustible liquid that could catch on fire.

21.2 CONDUCTING INVESTIGATIONS AND RECORDING AND PRESENTING DATA

5 **a** Mistake
  b Random error
  c Systematic error

6 **a** Bar graph
  b Line graph
  c Scatter graph (with line of best fit)
  d Pie chart

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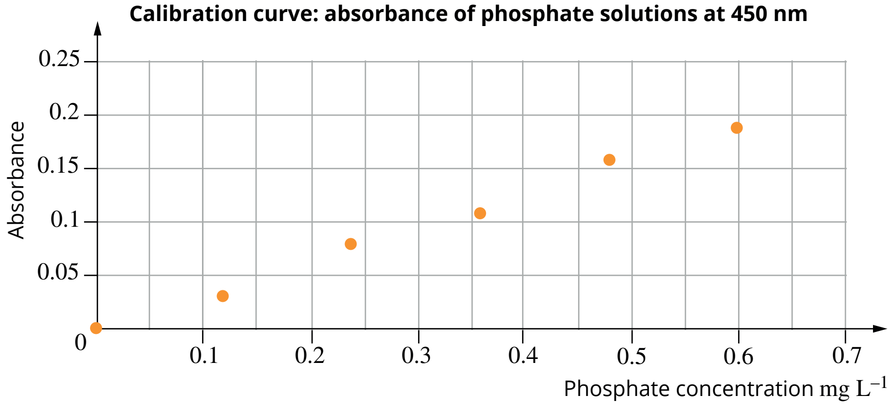
7 **a**



b Data point 4 is outlier.

c An outlier is a point in the data that does not fit the trend.

8 Temperature because dissolved oxygen varies with temperature

9 Add a title, label the *x*- and *y*-axes, add units to the *x*-axis variable, and exclude the obvious outliers when drawing the line of best fit.

21.3 DISCUSSING INVESTIGATIONS AND DRAWING EVIDENCE-BASED CONCLUSIONS

10 Evaluate the method; identify issues that could affect validity, accuracy, precision and reliability of data; state systematic sources of error and uncertainty; and recommend improvements to the investigation if it is to be repeated

11 A trend is a pattern or relationship that can be seen between the dependent and independent variables. It may be linear, in which the variables change in direct proportion to each other to produce a straight trend line. The relationship may be in proportion but non-linear, giving a curved trend line. The relationship may also be inverse, in which one variable decreases in response to the other variable increasing. This could be linear or non-linear.

12 The purpose of referencing and acknowledgments is to avoid plagiarism and ensure creators and sources are properly credited for their work.

CONNECTING THE MAIN IDEAS

13 **a** Aim: To determine the effect of increasing water temperature on the electrical conductivity of water.

b Independent variable – water temperature; dependent variable – electrical conductivity of water; controlled variables – pH, water source, sampling container

c The data collected would be conductivity using a probe and, therefore, it would be quantitative.

d 10 mL measuring cylinder ±0.1 mL, pH probe ±0.02 alcohol-filled glass thermometer ±0.1°C, electrical conductivity probe ±2

e Raw data is data collected in the field and recorded as measurements are taken. Processed data tabulates this into a form in which the reader can clearly see the temperature of the water and the conductivity at each separate temperature value. This can also then be processed and graphed with the independent variable on the *x*-axis (temperature) and dependent variable on the *y*-axis (conductivity). If the hypothesis is correct, the graph for this experiment would look like that below: