UNIT 1 ANSWERS

CHAPTER 1

1 ► a melting
   c subliming/sublimation
   d subliming/sublimation

b freezing

2 ► a
   ![Solid, Liquid, Gas](image)

Note: Solids should have regularly packed particles touching. Liquids should have most of the particles touching at least some of their neighbours, but with gaps here and there, and no regularity. Gases should have the particles well spaced.

b Solids: vibration around a fixed point. Liquids: particles can move around into vacant spaces, but with some difficulty because of the relatively close packing.

c Evaporation: Some faster moving particles break away from the surface of the liquid. Boiling: Attractive forces are broken throughout the liquid to produce bubbles of vapour.

3 ► a i A – gas; B – liquid; C – solid; D – liquid; E – solid
   b A, because it is a gas.
   c It sublimes and therefore is converted directly from a solid to a gas without going through the liquid stage.
   d D – because it has a lower boiling point the forces of attraction between particles will be weaker therefore it will also evaporate more easily than substance B (the only other substance that is a liquid at 25 °C).

4 ► a The ammonia and hydrogen chloride particles have to diffuse through the air in the tube, colliding with air particles all the way.
   b i Its particles will move faster.
   ii It would take slightly longer for the white ring to form, because the gas particles would be moving more slowly at the lower temperature.

   c Ammonia particles are lighter than hydrogen chloride particles and so move faster. The ammonia covers more distance than the hydrogen chloride in the same time.
   d i Ammonium bromide.
   ii The heavier hydrogen bromide particles would move more slowly than the hydrogen chloride particles, and so the ring would form even closer to the hydrobromic acid end than it was to the hydrochloric acid end. The ring will also take slightly longer to form because of the slower moving particles.

5 ► Sodium chloride dissolves in water to form a **solution**. The water is called the **solvent** and the sodium chloride is the **solute**. If the solution is heated to 50 °C some of the water evaporates until the solution becomes **saturated** and sodium chloride crystals start to form.

6 ► a
   ![Graph](image)

b 94 +/-1 g per 100 g

The values obtained in this question and in c depend on the line of best fit. In the exam there will always be some tolerance – a range of values will be accepted.

c From the graph, the solubility at 30 °C is 10 g per 100 g of water.

\[
40 \times \frac{10}{100} = 4 g
\]

Therefore 4 g of sodium chlorate will dissolve.

d i 53 +/-1 °C
   ii The solubility at 17 °C is 7 ±1 g per 100 g, therefore 20 – 7 = 13 g must precipitate out of the solution. Answers of 13 ±1 g are acceptable.

CHAPTER 2

1 ►

<table>
<thead>
<tr>
<th>Element</th>
<th>Compound</th>
<th>Mixture</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrogen</td>
<td>magnesium oxide</td>
<td>sea water</td>
</tr>
<tr>
<td>calcium</td>
<td>copper(II) sulfate</td>
<td>honey</td>
</tr>
<tr>
<td></td>
<td>blood</td>
<td></td>
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<tr>
<td></td>
<td>mud</td>
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<tr>
<td></td>
<td>potassium iodide</td>
<td>solution</td>
</tr>
</tbody>
</table>

2 ► a mixture   b mixture   c element
   d element    e compound  f compound

3 ► Substance X is the pure substance – it melts at a fixed temperature. Substance Y is impure – it melts over a range of temperatures.

4 ► a crystallisation   b (simple) distillation
   c fractional distillation   d chromatography
   e filtration

5 ► For example: Stir with a large enough volume of cold water to dissolve all the sugar. Filter to leave the diamonds on the filter paper. Wash on the filter paper with more water to remove any last traces of sugar solution. Allow to dry.

6 ► a M         b R
   c 0.45 ±0.01 (measure to the centre of the spot and remember to measure from the base line and not from the bottom of the paper)
   d G and T     e P
Palladium is a metal and so is likely to have any of the following properties:
- good conductor of electricity
- forms a basic oxide
- is shiny when polished or freshly cut
- is malleable
- is ductile
- is a good conductor of heat

The first two are mentioned specifically on the syllabus.

They have a full outer shell (energy level) and so they have no tendency to form compounds by losing/gaining electrons or sharing electrons.

Argon and potassium OR iodine and tellurium. The elements would then be in a different group in the Periodic Table. They would not have the same number of electrons in the outer shell as other members of the group and would react in a completely different way. For example, potassium would be in Group 0 with the noble gases, and argon, which is very unreactive, would be in Group 1, with the highly reactive alkali metals.
In each case, work out the M, by adding up the relative atomic masses (A_r values), and then attach the unit “g” to give the mass of 1 mole.

a 27 g
b 331 g
c 4.30 \times 16 = 68.8 g
d 0.70 \times 62 = 43.4 g
e 0.015 \times 85 = 1.275 g
f 0.24 \times 286 = 68.64 g Don’t forget the water of crystallisation

Strictly speaking the answers to d), e) and f) shouldn’t be quoted to more than 2 significant figures, because the number of moles is only quoted to that precision.

In each case, work out the mass of 1 mole as above, and then work out how many moles you’ve got in the stated mass. You can use the equation:

\[
\text{number of moles} = \frac{\text{mass}}{\text{mass of 1 mole}}
\]

a \frac{20}{40} = 0.5 mol
b \frac{3.20}{160} = 0.0200 mol
c \frac{2000}{79.5} = 25.2 \text{ mol; Don’t forget to convert kg to g!}
d \frac{50}{249.5} = 0.2 mol
e \frac{1000000}{56} = 17900 \text{ mol (or 17857, although this is precise to more significant figures than the A_r.)}
f \frac{0.032}{64} = 5.0 \times 10^{-4} \text{ mol (0.0005 mol)}

\[
\begin{array}{c|c|c}
\text{Combining mass} & \text{Na} & \text{S} \\
\hline
3.22 g & 4.48 g & 3.36 g \\
\end{array}
\]

\[
\begin{array}{c|c|c|c}
\text{No. of moles of atoms} & \text{Na} & \text{S} & \text{O} \\
\hline
\frac{3.22}{23} = 0.14 & \frac{4.48}{32} = 0.14 & \frac{3.36}{16} = 0.21 \\
\end{array}
\]

Empirical formula is \( \text{Na}_2\text{S}_2\text{O}_3 \)

\[
\begin{array}{c|c|c|c}
\text{Combining mass} & \text{C} & \text{H} & \text{Br} \\
\hline
22.0 & 4.6 & 73.4 \\
\end{array}
\]

\[
\begin{array}{c|c|c|c}
\text{No. of moles of atoms} & \text{C} & \text{H} & \text{Br} \\
\hline
\frac{22.0}{12} = 1.833 & 4.6 & \frac{73.4}{80} = 0.9175 \\
\end{array}
\]

Empirical formula is \( \text{C}_2\text{H}_5\text{Br} \)

\[
\begin{array}{c|c|c|c|c|c|c}
\text{Combining mass} & \text{P} & \text{O} \\
\hline
1.24 g & 1.60 g \\
\end{array}
\]

\[
\begin{array}{c|c|c|c|c|c|c}
\text{No. of moles of atoms} & \text{P} & \text{O} \\
\hline
\frac{1.24}{31} = 0.04 & \frac{1.60}{16} = 0.10 \\
\end{array}
\]

\[
\begin{array}{c|c|c|c|c|c|c}
\text{Ratio of moles (divide by smallest number)} & \text{P} & \text{O} \\
\hline
2 & 5 \\
\end{array}
\]

Empirical formula is \( \text{P}_2\text{O}_5 \)

\[
\begin{array}{c|c|c|c|c|c|c}
\text{Combining mass in 100g} & \text{C} & \text{H} & \text{O} \\
\hline
66.7 g & 11.1 g & 22.2 g \\
\end{array}
\]

\[
\begin{array}{c|c|c|c|c|c|c}
\text{No. of moles of atoms} & \text{C} & \text{H} & \text{O} \\
\hline
\frac{66.7}{12} = 5.558 & \frac{11.1}{1} = 11.1 & \frac{22.2}{16} = 1.3875 \\
\end{array}
\]

\[
\begin{array}{c|c|c|c|c|c|c}
\text{Ratio of moles (divide by smallest number)} & \text{C} & \text{H} & \text{O} \\
\hline
4 & 8 & 1 \\
\end{array}
\]

Empirical formula is \( \text{C}_4\text{H}_8\text{O} \)
b The mass of the empirical formula is $4 \times 12 + 8 \times 1 + 16 = 72$. Since this is equal to the relative formula mass, the molecular formula is the same as the empirical formula, that is $C_3H_6O$.

10 ▶ You know the mass of anhydrous sodium sulfate (1.42 g). You can work out the mass of water of crystallisation $(3.22 - 1.42 = 1.8)$ g. You can work out the mass of 1 mole of sodium sulfate, $Na_2SO_4 = 142$ g; and the mass of 1 mole of water = 18 g.

Number of moles of sodium sulfate = $\frac{1.42}{142} = 0.01$ mol.

Number of moles of water = $\frac{1.8}{18} = 0.1$ mol.

So for every 1 mole of sodium sulfate, there are 10 moles of water.

The value of $n$ is 10.

11 ▶ mass of anhydrous calcium sulfate = $44.14 - 37.34 = 6.80$ g

mass of water of crystallisation = $45.94 - 44.14 = 1.80$ g

mass of 1 mole of calcium sulfate, $CaSO_4 = 136$ g

number of moles of sodium sulfate = $\frac{6.80}{136} = 0.05$ mol

mass of 1 mole of water = 18 g

number of moles of water = $\frac{0.1}{0.1} = 2$

number of moles of calcium sulfate 0.05

the value of $n = 2$

12 ▶ a 0.36 mol

b From the chemical equation, the number 2 in front of the HCl indicates that 2 mol HCl react with 1 mol $CaCO_3$, therefore 0.4 mol $CaCO_3$ react with $2 \times 0.4 = 0.8$ mol HCl.

c 6 mol HCl react to form 3 mol $H_2S$

Therefore the number of moles of $H_2S$ is half the number of moles of HCl.

$0.4$ mol HCl form $0.2$ mol $H_2S$

d 3 mol CO form 2 mol Fe

The number of moles of Fe is $\frac{2}{3}$ the number of moles of CO.

$0.9 \times \frac{2}{3} = 0.6$ mol iron

e $0.8 \times \frac{2}{3} = 1.2$ mol hydrogen

13 ▶ a number of moles of iron = $\frac{10}{56} = 0.179$ mol

b From the chemical equation, the number of moles of bromine that reacted with this number of moles of iron is $\frac{3}{2} \times 0.179 = 0.277$ mol.

c From the chemical equation, the number of moles of FeBr$_3$ will be the same as the number of moles of iron = 0.179 mol.

d mass of 1 mol FeBr$_3$ = 56 + 3 \times 80 = 296

mass of FeBr$_3$ = 0.179 \times 296 = 53 g

(A common mistake here is to multiply the number of moles of FeBr$_3$ by the mass of 2FeBr$_3$. The 2 has already been used when you worked out that 2 mol Fe formed 2 mol FeBr$_3$ – do not use it again. The equation for working out the mass is mass = number of moles \times mass of 1 mole).

14 ▶ a $\frac{1.0}{190} = 0.0053$ mol

b 0.0053 mol

c $0.0053 \times 48 = 0.25$ g

d number of moles of $NaCl = 0.0053 \times 4 = 0.0212$ mol

mass of $NaCl = 0.0212 \times 58.5 = 1.2$ g

e You can carry out a moles calculation as above:

1 tonne = 1 000 000 g

moles of TiCl$_4$ = $\frac{1 000 000}{190} = 5300$ mol

moles of Ti = 5300 mol

mass of Ti = 5300 \times 48 = 254 400 g

Alternatively, you can reason that, if 1 g TiCl$_4$ forms 0.25 g Ti, 1 tonne TiCl$_4$ will form 0.25 tonne Ti.

15 ▶ mass of 1 mole of AlCl$_3$ = $27 + 3 \times 35.5 = 133.5$ g

moles of Aluminum chloride = $\frac{2.67}{133.5} = 0.0200$ mol

moles of AgCl = $3 \times 0.0200 = 0.0600$ mol

mass of 1 mole of AgCl = $108 + 35.5 = 143.5$ g

mass of AgCl = $0.0600 \times 143.5 = 8.61$ g

16 ▶ a mass of 1 mol $Cr_2O_3 = 2 \times 52 + 3 \times 16 = 152$ g

number of moles of $Cr_2O_3$ = $\frac{50}{152} = 0.33$ mol

number of moles of Al = $2 \times 0.33 = 0.66$ mol

mass of Al = $0.66 \times 27 = 17.8$ g

b number of moles of Cr = $2 \times 0.33 = 0.66$ mol

mass of Cr = $0.66 \times 52 = 34.3$ g

c 5 kg is 5000 g

number of moles of $Cr_2O_3 = \frac{5000}{152} = 33$ mol

number of moles of Cr = $2 \times 33 = 66$ mol

mass of Cr = $66 \times 52 = 3430$ g or 3.43 kg

Alternatively, we can reason that 5 kg is 100 times as much as 50 g. If we start with 100 times as much $Cr_2O_3$, we will make 100 times as much Cr.

d 5 tonnes is 5000000 g

number of moles of $Cr_2O_3 = \frac{5000000}{152} = 33000$ mol

number of moles of Cr = $2 \times 33000 = 66000$ mol

mass of Cr = $66000 \times 52 = 3430000$ g or 3.43 tonnes

Alternatively, we can reason that, if 5 kg of $Cr_2O_3$ produces 3.43 kg Cr, then 5 tonnes of $Cr_2O_3$ will produce 3.43 tonnes of Cr.

17 ▶ a mass of 1 mol $CuO = 63.5 + 16 = 79.5$ g

number of moles of $CuO = \frac{4.00}{79.5} = 0.0503$ mol

number of moles of $CuSO_4 = 0.0503$ mol

number of moles of $CuSO_4 \cdot 5H_2O = 0.0503$ mol

mass of 1 mol $CuSO_4 \cdot 5H_2O = 63.5 + 32 + 4 \times 16 + 5 \times (16 + 2) = 249.5$ g

mass of $CuSO_4 \cdot 5H_2O = 0.0503 \times 249.5 = 12.55$ g
18 a mass of 1 mol ethanol = 2 × 12 + 6 × 1 + 16 = 46 g
   number of moles of ethanol = \frac{20.0}{46} = 0.435 mol
   number of moles of ethyl ethanoate = 0.435 mol
   mass of 1 mol ethyl ethanoate = 4 × 12 + 2 × 16 + 8 × 1 = 88 g
   mass of ethyl ethanoate = 0.435 × 88 = 38.3 g
   b theoretical yield = 38.3 g
   actual yield = 30.0 g
   percentage yield = \frac{30.0}{38.3} × 100 = 78.3%

19 a 0.5 mol HCl would react with 0.25 mol Na₂CO₃.
     There is more than 0.25 mol Na₂CO₃, therefore Na₂CO₃
     is in excess.
     b 0.02 mol O₂ would react with \frac{0.02}{5} = 0.004 mol C₂H₆.
     There is more than 0.004 mol C₂H₆, therefore C₂H₆
     is in excess.
     c 28 g of CO is \frac{28}{28} = 1 mol
     1 mol CO would react with \frac{1}{3} = 0.33 mol Fe₂O₃.
     There is more than 0.33 mol Fe₂O₃, therefore Fe₂O₃
     is in excess.
     d 16 g O₂ is \frac{16}{32} = 0.5 mol
     16 g SO₂ is \frac{16}{64} = 0.25 mol
     0.25 mol SO₂ would react with 0.125 mol O₂.
     There is more than 0.125 mol O₂, therefore O₂ is in excess.

20 a 1.0 g of CaCO₃ is \frac{1.0}{100} = 0.010 mol
    0.010 mol CaCO₃ would react with 2 × 0.010 = 0.020 mol HCl
    There is less than 0.020 mol HCl, therefore there is
    not enough HCl to react with all the CaCO₃. Therefore
    CaCO₃ is in excess.
    b To calculate the number of moles of CO₂ you must use
    the number of moles of HCl because not all the CaCO₃
    reacted.
    moles of HCl = 0.015 mol
    moles of CO₂ = 0.5 × 0.015 = 0.0075 mol
    mass of CO₂ = 0.0075 × 44 = 0.33 g
    (If you got the answer 0.44 g you used the number of
    moles of CaCO₃. CaCO₃ was in excess; therefore not
    all of it will react.)

2 a 2.0 × 24 = 48 dm³ (48 000 cm³)
   b 0.10 × 24 = 2.4 dm³ (2400 cm³)
   c \frac{1}{2} × 0.024 = 0.012 dm³ (120 cm³)

3 a 200 cm³ of chlorine is \frac{200}{24 000} = 0.00833 mol
    (be careful with units here – if the volume is in cm³ you
    must use 24 000 as the molar volume)
    mass of 1 mol of Cl₂ = 2 × 35.5 = 71 g
    mass of 0.00833 mol Cl₂ = 0.00833 × 71 = 0.592 g
    b mass of 1 mol of O₂ = 2 × 16 = 32 g
    number of moles of O₂ = \frac{0.16}{32} = 0.005 mol
    0.005 mol O₂ has a volume of 0.005 × 24 000 = 120 cm³ (0.12 dm³)
    c 1 dm³ of the gas is \frac{1}{24} = 0.0417 mol
    the mass of 0.0417 mol is \frac{1.42}{0.0417} = 34.1 g

4 a \frac{0.240}{24} = 0.0100 mol
    From the chemical equation, the number of moles of H₂ is
    the same as the number of moles of Mg: 0.0100 mol.
    volume of hydrogen = 0.0100 × 24 = 0.24 dm³ (240 cm³)

5 \frac{1}{24} = 0.0417 mol O₂
    From the chemical equation, the number of moles of KNO₃ is twice the number of moles of O₂
    2 × 0.0417 = 0.0833 mol
    mass of 1 mol KNO₃ = 39 + 16 + 1 = 56 g
    mass of 0.0833 mol KNO₃ = 0.0833 × 56 = 4.67 g

6 a \frac{2.00}{87} = 0.0230 mol
    From the chemical equation: 1 mol MnO₂ produces
    1 mol Cl₂
    number of moles of Cl₂ produced = 0.0230 mol
    volume of Cl₂ produced = 0.0230 × 24 000 = 552 cm³ (0.552 dm³)

7 a mass of 1 mol H₂SO₄ = 98 g
    number of moles of H₂SO₄ \frac{4.90}{98} = 0.0500 mol
    this number of moles is in 1 dm³ of solution
    concentration = 0.0500 mol/dm³ (0.0500 to show that the
    answer is precise to 3 significant figures)

8 a mass of 1 mol KOH = 39 + 16 + 1 = 56 g
    mass of 0.200 mol = 0.200 × 56 = 11.2 g
    concentration = 11.2 g/dm³

9 a In each of these questions the number of moles is given by:
    number of moles = volume in dm³ × concentration in mol/dm³
    to convert a volume in dm³ to cm³ divide by 1000
    a \frac{25.0}{1000} × 0.100 = 0.00250 mol
    b \frac{200}{1000} × 0.200 = 0.0400 mol
mass of 1 mol CaCO₃ = 100 g
number of moles of CaCO₃ = \( \frac{10.0}{100} = 0.100 \) mol
from the chemical equation 0.100 mol CaCO₃ reacts with 2 × 0.100 = 0.200 mol HCl

volume (dm³) = \( \frac{\text{number of moles (mol)}}{\text{concentration (mol/dm³)}} \)

volume = \( \frac{0.200}{2.00} = 0.100 \) dm³ or 100 cm³

In each of these questions the number of moles is given by
number of moles = volume in dm³ × concentration in mol/dm³
to convert a volume in dm³ to cm³ divide by 1000

a no. of moles of NaOH solution = \( \frac{25.0}{1000} \times 0.100 = 0.00250 \) mol
the equation shows that you need the same number of moles of nitric acid = 0.00250 mol
concentration (mol/dm³) = \( \frac{\text{number of moles (mol)}}{\text{volume (dm³)}} \)
concentration = \( \frac{0.00250}{0.0200} = 0.125 \) mol/dm³

b no. of moles of nitric acid = \( \frac{30.0}{1000} \times 0.100 = 0.00300 \) mol
the equation shows that you need half the number of moles of sodium carbonate = 0.5 × 0.00300 = 0.00150 mol
correction (mol/dm³) = \( \frac{\text{number of moles (mol)}}{\text{volume (dm³)}} \)
concentration = \( \frac{0.00150}{0.0250} = 0.0600 \) mol/dm³

c no. of moles of potassium carbonate solution = \( \frac{25.0}{1000} \times 0.150 = 0.01125 \) mol
the equation shows that you need twice the number of moles of ethanoic acid = 2 × 0.00625 = 0.0125 mol
correction (mol/dm³) = \( \frac{\text{number of moles (mol)}}{\text{volume (dm³)}} \)
concentration = \( \frac{0.0125}{0.0125} = 1.00 \) mol/dm³

25.0 cm³ is \( \frac{25.0}{1000} = 0.0250 \) dm³
concentration = \( \frac{0.000376}{0.0250} = 0.0150 \) mol/dm³

75.0 cm³ is \( \frac{75.0}{1000} = 0.0750 \) dm³
concentration = \( \frac{0.000600}{0.0750} = 0.00800 \) mol/dm³

d 22.4 cm³ is \( \frac{22.4}{1000} = 0.0224 \) dm³
concentration = \( \frac{0.000400}{0.0224} = 0.0179 \) mol/dm³

concentration (mol/dm³) = \( \frac{\text{number of moles (mol)}}{\text{volume (dm³)}} \)
volume = \( \frac{0.000400}{0.0179} = 0.0224 \) dm³ or 22.4 cm³

The volume of oxygen is 0.000300 × 24 000 = 7.20 cm³

The volume of CO₂ = 0.0250 × 24 000 = 600 cm³

The number of moles of NO₂ = 0.0250 mol
correction (mol/dm³) = \( \frac{\text{number of moles (mol)}}{\text{volume (dm³)}} \)
concentration = \( \frac{0.000600}{0.0250} = 0.0240 \) mol/dm³

The number of moles of H₂O₂ is 0.000300 mol (3.00 × 10⁻⁶)

The volume of oxygen = 0.000300 × 24 000 = 7.20 cm³

From the chemical equation, the number of moles of O₂ produced is half the number of moles of O₂
number of moles of O₂ = 0.000300 mol (3.00 × 10⁻⁶)
volume of oxygen = 0.000300 × 24 000 = 7.20 cm³

The number of moles of NO₂ = 0.0250 mol
correction (mol/dm³) = \( \frac{\text{number of moles (mol)}}{\text{volume (dm³)}} \)
concentration = \( \frac{0.000600}{0.0250} = 0.0240 \) mol/dm³

The number of moles of H₂O₂ is 0.000300 mol (3.00 × 10⁻⁶)

The volume of oxygen = 0.000300 × 24 000 = 7.20 cm³

The number of moles of NO₂ = 0.0250 mol
correction (mol/dm³) = \( \frac{\text{number of moles (mol)}}{\text{volume (dm³)}} \)
concentration = \( \frac{0.000600}{0.0250} = 0.0240 \) mol/dm³

25.0 cm³ is \( \frac{25.0}{1000} = 0.0250 \) dm³
concentration = \( \frac{0.000376}{0.0250} = 0.0150 \) mol/dm³
mass of 1 mol Ca(OH)\textsubscript{2} = 40 + 2 \times (16+1) = 74 \text{ g} \\
there are 0.0150 \text{ mol in 1 dm}^3 \\
the mass of 0.0150 \text{ mol is } 0.0150 \times 74 = 1.11 \text{ g} \\
therefore the concentration is 1.11 \text{ g/dm}^3

18  a  0.1 mol HNO\textsubscript{3} reacts with 0.05 mol Na\textsubscript{2}CO\textsubscript{3}. There is more than this present, therefore Na\textsubscript{2}CO\textsubscript{3} is in excess.

b no. of moles of Na\textsubscript{2}CO\textsubscript{3} = \frac{20.0 \times 0.100 \text{ mol}}{1000} = 0.00200 \text{ mol} \\
0.0200 \text{ mol HNO\textsubscript{3} reacts with 0.0100 \text{ mol Na\textsubscript{2}CO\textsubscript{3}. There is less than this present, therefore HNO\textsubscript{3} is in excess.}

c no. of moles of Na\textsubscript{2}CO\textsubscript{3} = \frac{25.0 \times 0.300 \text{ mol}}{1000} = 0.00750 \text{ mol} \\
no of moles of HNO\textsubscript{3} = \frac{20.0 \times 0.400 \text{ mol}}{1000} = 0.00800 \text{ mol} \\
0.00800 \text{ mol HNO\textsubscript{3} reacts with 0.00400 \text{ mol Na\textsubscript{2}CO\textsubscript{3. There is more than this present, therefore Na\textsubscript{2}CO\textsubscript{3} is in excess.}

b Correct electronic structures for:

i  Na 2,8,1 and Cl 2,8,7 \\
ii  Li 2,1 and O 2,6 \\
iii  Mg 2,8,2 and F 2,7.

Diagrams (similar to those in the chapter) showing transfer of electrons, and the charges and electronic structures of the ions formed (or words to the same effect).

In (i), show 1 electron transferred from Na to Cl leaving Na\textsuperscript{+} [2,8]\textsuperscript{+} and Cl\textsuperscript{–} [2,8,8]

In (ii), show 2 lithium atoms each giving 1 electron to O leaving 2 × Li\textsuperscript{+} (2)+ and O\textsubscript{2–} [2,8]\textsuperscript{2–}

In (iii), show 1 Mg giving an electron each to 2 fluorines leaving Mg\textsuperscript{2+} [2,8]\textsuperscript{2+} and 2 × F\textsuperscript{–} [2,8]\textsuperscript{–}

3  a  formula  b  name

<p>| | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
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</thead>
<tbody>
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<tr>
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<td>strontium</td>
<td>Sr\textsuperscript{2+}</td>
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<tr>
<td>iii</td>
<td>potassium</td>
<td>K\textsuperscript{+}</td>
</tr>
<tr>
<td>iv</td>
<td>oxygen</td>
<td>O\textsuperscript{2–} oxide</td>
</tr>
<tr>
<td>v</td>
<td>sulfur</td>
<td>S\textsuperscript{2–} sulfide</td>
</tr>
<tr>
<td>vi</td>
<td>caesium</td>
<td>Cs\textsuperscript{+}</td>
</tr>
<tr>
<td>vii</td>
<td>chlorine</td>
<td>Cl\textsuperscript{–} chloride</td>
</tr>
<tr>
<td>viii</td>
<td>iodine</td>
<td>I\textsuperscript{–} iodide</td>
</tr>
<tr>
<td>ix</td>
<td>aluminium</td>
<td>Al\textsuperscript{3+}</td>
</tr>
<tr>
<td>x</td>
<td>calcium</td>
<td>Ca\textsuperscript{2+}</td>
</tr>
<tr>
<td>xi</td>
<td>nitrogen</td>
<td>N\textsuperscript{3–} nitride</td>
</tr>
</tbody>
</table>

4  PbO  NaBr \\
MgSO\textsubscript{4}  ZnCl\textsubscript{2} \\
K\textsubscript{2}CO\textsubscript{3}  (NH\textsubscript{4})\textsubscript{2}S \\
Ca(NO\textsubscript{3})\textsubscript{2}  Fe(OH)\textsubscript{3} \\
FeSO\textsubscript{4}  CuCO\textsubscript{3} \\
Al\textsubscript{2}(SO\textsubscript{4})\textsubscript{3}  Ca(OH)\textsubscript{2} \\
CoCl\textsubscript{2}  CaO \\
AgNO\textsubscript{3}  FeF\textsubscript{3} \\
NH\textsubscript{4}NO\textsubscript{3}  RbI \\
Na\textsubscript{2}SO\textsubscript{4}  Cr\textsubscript{2}O\textsubscript{3}

5  a  The electrostatic forces of attraction between oppositely charged ions are strong and require a lot of energy to break.

b  The ions are held tightly in place in the giant lattice structure and are not free to move.

c  The ions are free to move (it is important to use the word ions here; any mention of electrons will score 0 in an exam).

6  Potassium chloride will have a lower melting point than calcium oxide. The charges on the ions in KCl (K\textsuperscript{+} and Cl\textsuperscript{–}) are lower than in CaO (Ca\textsuperscript{2+} and O\textsuperscript{2–}). There are weaker electrostatic forces of attraction between oppositely charged ions in KCl than in CaO; these forces require less energy to break than the forces in CaO.

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FeSO\textsubscript{4}  CuCO\textsubscript{3} \\
Al\textsubscript{2}(SO\textsubscript{4})\textsubscript{3}  Ca(OH)\textsubscript{2} \\
CoCl\textsubscript{2}  CaO \\
AgNO\textsubscript{3}  FeF\textsubscript{3} \\
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The electrostatic forces of attraction between oppositely charged ions are strong and require a lot of energy to break.

The ions are held tightly in place in the giant lattice structure and are not free to move.

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Potassium chloride will have a lower melting point than calcium oxide. The charges on the ions in KCl (K\textsuperscript{+} and Cl\textsuperscript{–}) are lower than in CaO (Ca\textsuperscript{2+} and O\textsuperscript{2–}). There are weaker electrostatic forces of attraction between oppositely charged ions in KCl than in CaO; these forces require less energy to break than the forces in CaO.

CHAPTER 8

1  a  MgO  ionic \\
b  CH\textsubscript{3}Br  covalent \\
c  H\textsubscript{2}O\textsubscript{2}  covalent \\
d  FeCl\textsubscript{2}  ionic \\
e  NaF  ionic \\
f  HCN  covalent

2  a  A pair of electrons that is shared between two atoms. The atoms are held together because the nucleus of each is attracted to the shared pair.
352 ANSWERS

b (It doesn’t matter whether students use dots or crosses or just different colours - or what positions (N, S, E, W) the hydrogens occupy in the H₂S or PH₃.)

4 ▶ Carbon dioxide has a simple molecular structure; diamond has a giant covalent structure. When carbon dioxide sublimes, only the weak intermolecular forces of attraction must be broken – but when diamond sublimes, the strong covalent bonds must be broken. A lot more energy is required to break the strong covalent bonds in diamond than the weak intermolecular forces in carbon dioxide.

5 ▶ a Simple molecular because it is a liquid at room temperature. Only weak intermolecular forces of attraction must be broken to melt solid hexane. Compounds with giant structures have high melting points and boiling points and will be solids at room temperature.

b Pentane has a lower boiling point. The intermolecular forces of attraction are weaker in pentane because the relative formula mass is lower. Weaker intermolecular forces require less energy to break.

c It will not conduct electricity because there are no ions present and all the electrons are held tightly in atoms or covalent bonds.

6 ▶ a To break apart diamond, strong covalent bonds must be broken, which requires a large amount of energy. Much less energy is required to break the weak forces of attraction between the layers in graphite.

b C₆₀ fullerene has a molecular structure but graphite has a giant structure. To melt C₆₀ fullerene, only weak intermolecular forces must be broken, but to melt graphite strong covalent bonds must be broken. Much less energy is required to break the weak intermolecular forces in C₆₀ fullerene than the strong covalent bonds in graphite.

c Each C atom in graphite only forms 3 bonds so there is one electron left over on each on each atom. These delocalised electrons are free to move throughout the layers.

d All the outer shell electrons in diamond are held tightly in covalent bonds and unable to move around.

7 ▶ (Weakest intermolecular forces of attraction) hydrogen, phosphorus trifluoride, ammonia, ethanol, water, ethanamid (strongest intermolecular forces of attraction).

Higher intermolecular attractions produce higher boiling points – more energy has to be supplied to overcome stronger forces of attraction between molecules.

8 ▶ a Nitrogen usually forms 3 bonds because it has 5 electrons in its outer shell. Each N forms 1 bond to F, therefore there must be a double bond between the two N atoms. Formation of a double bond results in each N having 8 electrons in its outer shell.
(It doesn’t matter what variations of colours or dots and crosses are used. The F atoms could also be drawn in different positions.)

9 ▶ a

![Diagram of BCl3](image)

b The B atom does not have 8 electrons in its outer shell.

b ▶ a Mg$^{2+}$ + 2e$^-$ → Mg
   b Al$^{3+}$ + 3e$^-$ → Al
   c 2Br$^-$ → Br$_2$ + 2e$^-$
   d 2O$_2^-$ → O$_2$ + 4e$^-$
   e 2Cl$^-$ → Cl$_2$ + 2e$^-$
   f Ni$^{2+}$ + 2e$^-$ → Ni
   g 4OH$^-$ → O$_2$ + 2H$_2$O + 4e$^-$
   h 2H$_2$O + 2e$^-$ → H$_2$ + 2OH$^-$
   i 2H$_2$O + 2e$^-$ → H$_2$ + 2OH$^-$

b ▶ a Ions weren’t free to move.
   b anode
c iodine: 2I$^-$(l) → I$_2$(g) + 2e$^-$
   d K$^+(l)$ + e$^-$ → K(l)
   e anode: bromine
   f cathode: sodium
   f anode: 2Br$^-$ → Br$_2$(g) + 2e$^-$
   cathode: Na$^+$ + e$^-$ → Na

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   b anode
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   f cathode: sodium
   f anode: 2Br$^-$ → Br$_2$(g) + 2e$^-$
   cathode: Na$^+$ + e$^-$ → Na

2 ▶ a Pb$^{2+}$ + 2e$^-$ → Pb(l)
   b 2H$^+$ + 2e$^-$ → H$_2$(g)
   c Cu$^{2+}$(aq) + 2e$^-$ → Cu(s)
   d 2H$^+$ + 2e$^-$ → H$_2$(g)
   e 2H$^+$ + 2e$^-$ → H$_2$(g)
   f 2H$^+$ + 2e$^-$ → H$_2$(g)
   g 2H$^+$ + 2e$^-$ → H$_2$(g)
   h 2H$^+$ + 2e$^-$ → H$_2$(g)
   i 2H$^+$ + 2e$^-$ → H$_2$(g)

b ▶ a Mg 2, 8, 2. The two outer electrons will be lost from each Mg atom to form the sea of delocalised electrons. There will be a regular arrangement of Mg$^{2+}$ ions.
   b Na will form Na$^+$ ions, Mg will form Mg$^{2+}$ ions and Al will form Al$^{3+}$ ions. There will be stronger electrostatic attraction between the metal ions and the delocalised electrons when the charge on the ion is higher. Al$^{3+}$ has the strongest attraction between the metal ions and the delocalised electrons, therefore most energy has to be supplied to overcome the forces of attraction.
   c Delocalised electrons are free to move.
   d The layers of metal ions are able to slide over each other without changing the bonding.

c ▶ a giant covalent
   b molecular
   c molecular
   d giant ionic
   e giant metallic
   f molecular
   g giant metallic
   h giant metallic
   i giant metallic
   j giant metallic

**CHAPTER 9**

1 ▶ a

![Diagram of a lattice of metal ions](image)

b

![Diagram of a lattice of metal ions with sea of delocalised electrons](image)

(The diagrams should show at least 9 ions but can be of any size beyond that. Electrons could be shown as e$^-$)

2 ▶ a Mg 2, 8, 2. The two outer electrons will be lost from each Mg atom to form the sea of delocalised electrons. There will be a regular arrangement of Mg$^{2+}$ ions.
   b Na will form Na$^+$ ions, Mg will form Mg$^{2+}$ ions and Al will form Al$^{3+}$ ions. There will be stronger electrostatic attraction between the metal ions and the delocalised electrons when the charge on the ion is higher. Al$^{3+}$ has the strongest attraction between the metal ions and the delocalised electrons, therefore most energy has to be supplied to overcome the forces of attraction.
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3 ▶ a giant covalent
   b molecular
   c molecular
   d giant ionic
   e giant metallic
   f molecular
   g giant metallic
   h giant metallic
   i giant metallic
   j giant metallic
   k giant metallic

4 ▶ a Pb$^{2+}$(aq) + 2e$^-$ → Pb(l)
   b 2H$^+(aq) + 2e^-$ → H$_2(g)$
   c Cu$^{2+}$(aq) + 2e$^-$ → Cu(s)
   d 2H$^+(aq) + 2e^-$ → H$_2(g)$
   e 2H$^+(aq) + 2e^-$ → H$_2(g)$
   f 2H$^+(aq) + 2e^-$ → H$_2(g)$
   g 2H$^+(aq) + 2e^-$ → H$_2(g)$
   h 2H$^+(aq) + 2e^-$ → H$_2(g)$
   i 2H$^+(aq) + 2e^-$ → H$_2(g)$

b ▶ a lead bromine
   b zinc chlorine
   c hydrogen iodine
   d sodium iodine
   e copper chlorine
   f hydrogen chlorine
   g hydrogen oxygen
   h hydrogen oxygen
   i hydrogen oxygen

**CHAPTER 10**

1 ▶ cathode anode
   a lead bromine
   b zinc chlorine
   c hydrogen iodine
   d sodium iodine
   e copper chlorine
   f hydrogen chlorine
   g hydrogen oxygen
   h hydrogen oxygen
The melting point of S is too high to reach using a Bunsen, and so you would have to test a solution in water. On the other hand, T would melt easily, and won’t dissolve. Heat it until it melts.

![Diagram of electrolysis](image)

(You could also do the electrolysis of the solution using the more complicated apparatus in the chapter, but since there is no need to collect anything, there isn’t much point.)

If the substances are electrolytes, the bulbs will light up, and there will be signs of activity around the electrodes (gases given off, solids deposited, etc).

Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)

The Cu^{2+}(aq) ions are responsible for the blue colour and these are removed from the solution, therefore the blue colour fades.

Water dissociates

H_2O \rightleftharpoons H^+ + OH^-

Oxygen is produced at the anode. The half equation for this is:

4OH^- \rightarrow 2H_2O + O_2 + 4e^-

Because the OH^- is removed from solution, more water must dissociate to replace them. This produces an excess of H^+ ions in solution, so the solution is acidic.

There is no change in colour because the concentration of Cu^{2+}(aq) remains the same; when 1 Cu^{2+}(aq) ion is removed at the cathode another one replaces it at the anode.

There is no change in acidity because the OH^- ions are not removed from the solution to form oxygen.

END OF UNIT 1 QUESTIONS

1 a −255 °C (1) The temperature is between the melting point and boiling point.

b Particles randomly arranged (1) and mostly touching each other. (1)

c number of protons: 1 (1)

number of neutrons: 2 (1)

number of electrons: 1 (1)

d i 2, 5 (1)

ii NH_4SO_4 is incorrect (1) because the ammonium ion is NH_4^+ and the sulfate ion is SO_4^{2-}.

e fractional distillation (1)

The elements in the Periodic Table are arranged in order of atomic number (1). The vertical columns are called groups (1) and contain elements which have the same number of electrons (1) in their outer shell (1).

H (1) and Ar (1)

Metal oxides are basic (1) and would react with acids rather than with alkalis such as sodium hydroxide. (1)

ii The grey circles represent silicon atoms. (1) A silicon atom has 4 electrons in its outer shell and will form 4 covalent bonds. (1)

An oxygen atom has 6 electrons in its outer shell and will only form 2 covalent bonds. (1)

(Note: although some of the ‘grey atoms’ only appear to form 2 bonds this is because only part of the giant structure is shown. None of the ‘red atoms’ form 4 bonds.)

SiO_2 would be a solid at room temperature because it has a giant structure. (1)

A lot of energy is required to break all the strong covalent bonds. (1)

3 a idea of electron transfer (1)

electron transfer from Sr to Br (1)

the correct number of electrons transferred (1)

b high melting point (1)

strong electrostatic forces of attraction between oppositely-charged ions (1)

require a lot of energy to break (1)

c The relative atomic mass of an element is the weighted average mass of the isotopes of the element. (1)

It is measured on a scale on which a carbon-12 atom has a mass of exactly 12. (1)

d 50.69 \times 79 + 49.31 \times 81 \over 100 (1)

79.99 (1)

e Sr(NO_3)_2 (1)
4 a Diamond has a giant covalent structure (1), where all the atoms are held together by strong covalent bonds – a lot of energy is required to break these bonds. (1)

b i Graphite has the higher melting point. (1)
Graphite has a giant structure but C_{60} fullerene has a simple molecular structure. (1)
To melt graphite, all the strong covalent bonds between carbon atoms must be broken – this requires a lot of energy. (1)
To melt C_{60} fullerene, only the weak intermolecular forces of attraction must be broken – this requires less energy. (1)

ii Graphite has delocalised electrons. (1)
Delocalised electrons are free to move (throughout the structure). (1)

iii C_{60} has a molecular structure rather than a giant structure. (1)
Electrons cannot move between molecules. (1)

5 a O\O\O (1)
The correct diagram is above. (There is a double bond between the \O\ atoms.)

b K_2O is the correct formula (1)
c O\C\O (1)
The correct diagram is above. (Carbon forms 4 bonds and oxygen forms 2 bonds.)

6 a mass of lead = 24.16 − 17.95 = 6.21 g (1)
b mass of oxygen = 24.80 − 24.16 g = 0.64 g (1)
c Pb O

<table>
<thead>
<tr>
<th>combining masses</th>
<th>6.21 g</th>
<th>0.64 g</th>
</tr>
</thead>
<tbody>
<tr>
<td>no of moles of atoms</td>
<td>6.21/207</td>
<td>0.64/16</td>
</tr>
<tr>
<td>ratio of moles</td>
<td>3 : 4 (1)</td>
<td></td>
</tr>
<tr>
<td>empirical formula: Pb_3O_4 (1)</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

7 a number of moles of copper = \frac{2.00}{63.5} = 0.0315 mol (1)
From the chemical equation, the same number of moles of copper(II) nitrate will be produced.
number of moles of Cu(NO_3)_2 is 0.0315 mol (1)
M_\text{r} of Cu(NO_3)_2 = 63.5 + 2 \times (14 + 3 \times 16) = 187.5
mass of Cu(NO_3)_2 = 187.5 \times 0.0315 = 5.91 g (1)

b The student would use crystallisation. (1)
Heat the solution in an evaporating basin to drive off some of the water until the solution becomes saturated. (1)
Allow to cool then filter off the crystals that form. (1)

c percentage yield = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100
= \frac{5.23}{7.61} \times 100 (1)
= 68.7\% (1)

8 a M_\text{r} of FeS_2 = 56 + 2 \times 32 = 120
480 kg is 480 000 g
number of moles of FeS_2 = \frac{480 000}{120} = 4000 mol (1)

From the chemical equation, the number of moles of Fe_2O_3 is half the number of moles of FeS_2.
number of moles of Fe_2O_3 = \frac{1}{2} \times 4000 = 2000 mol (1)
M_\text{r} of Fe_2O_3 = 2 \times 56 + 3 \times 16 = 160
mass of Fe_2O_3 = 2000 \times 160 = 320 000 g or 320 kg (1)

b From the chemical equation 4000 mol FeS_2 produces 8000 mol SO_2. (1)
volume = 8000 \times 24 = 192 000 dm^3 (1)

c i 2SO_2 + O_2 \rightarrow 2SO_3
all formulae correct (1)
correct balancing (1)

ii From the equation 2 mol SO_2 react with 1 mol O_2. So the number of moles of O_2 will be half as many. Therefore the number of moles of O_2 will be 4000 mol. (1)
volume = 4000 \times 24 = 96 000 dm^3 (1)
For the second mark we could just reason that since half as many moles are required, then the volume will also be half.

9 a A saturated solution is a solution which contains as much dissolved solid as possible at a particular temperature. (1)
b number of moles = \frac{32.8}{1000} \times 0.100 = 3.28 \times 10^{-3} mol or 0.00328 mol (1)
c From the chemical equation, the number of moles of Sr(OH)_2 is half the number of moles of HCl.
number of moles of Sr(OH)_2 = 1.64 \times 10^{-3} or 0.00164 mol (1)
volume of Sr(OH)_2 in dm^3 is \frac{25}{100} = 0.025 dm^3 (1)
concentration of Sr(OH)_2 = \frac{0.00164}{0.025} = 0.0656 mol/dm^3 (1)

Mr of Sr(OH)_2 = 88 + 2 \times (16 + 1) = 122
mass of Sr(OH)_2 in 1 dm^3 of solution is 122 \times 0.0656 = 8.00 g (1)
1 dm^3 is 1000 cm^3, therefore mass of Sr(OH)_2 present in 100 cm^3 or 100 g of water is one tenth of this, that is 0.800 g (1)

e (The three points in bold are essential and then 1 mark each for 2 further points.)
an empty evaporating basin is weighed (1)
some of the strontium hydroxide solution is poured into the evaporating basin (1)
care must be taken to not pour in any solid (1)
weigh the evaporating basin and strontium hydroxide solution (1)
the evaporating basin is heated over a tripod and gauze using a Bunsen burner (1)
until all the water is driven off / heat to constant mass (1)
weigh the evaporating basin and solid strontium hydroxide (1)

10 a Number of moles of sodium hydroxide = \frac{25.0}{1000} \times 0.100 = 2.50 \times 10^{-3} mol or 0.0025 mol (1)
From the chemical equation, the number of moles of sulfuric acid is half of this.
Some gas could have escaped from the apparatus / there was a leak in the apparatus / they could have used less than 0.15 g limestone / they used a different concentration of hydrochloric acid that had too low a concentration / they used too low a volume of hydrochloric acid. (1)

e All the hydrochloric acid has been used up. (1)
There is no more hydrochloric acid to react with any additional limestone, so no more gas can be produced. (1)

f 0.42 g (1)
(This is the point at which the two lines cross because this is the point at which adding any more CaCO₃ does not produce any more gas.)

g i \[ \text{CaCO}_3 (s) + 2\text{HCl (aq)} \rightarrow \text{CaCl}_2 (aq) + \text{H}_2\text{O (l)} + \text{CO}_2 (g) \]

ii Number of moles of \( \text{CO}_2 \) = \( \frac{91}{24000} \) = 3.79 \times 10^{-3} \text{ mol or 0.00379 mol}. (1)
From the chemical equation, the number of moles of \( \text{CaCO}_3 \) is the same as the number of moles of \( \text{CO}_2 \).
Number of moles of \( \text{CaCO}_3 \) is 3.79 \times 10^{-3} \text{ mol or 0.00379 mol}. (1)
Mr of \( \text{CaCO}_3 \) is 40 + 12 + 3 \times 16 = 100
mass of \( \text{CaCO}_3 \) = 0.00379 \times 100 = 0.379 g (1)

iii \( \frac{0.379}{0.42} \times 100 = 90 \% \) (1)
You have worked out above the mass of limestone that reacts with the hydrochloric acid is 0.42 g. The volume of \( \text{CO}_2 \) produced by this mass of limestone is 91 cm³. 91 cm³ of \( \text{CO}_2 \) is produced by 0.379 g of \( \text{CaCO}_3 \), therefore this is the mass of \( \text{CaCO}_3 \) present in 0.42 g of limestone.

UNIT 2 ANSWERS

CHAPTER 11

1 ▶ They have the same number of electrons in the outer shell (1 electron) and all react in the same way.

2 ▶ a A = lithium
b B = potassium; C = hydrogen; D = potassium hydroxide
c \[ 2\text{K(s)} + 2\text{H}_2\text{O(l)} \rightarrow 2\text{KOH(aq)} + \text{H}_2(g) \]
d The paper goes blue/purple.
e Lots of heat evolved. Melting point of potassium is low.
f E is sodium and F is sodium oxide
sodium + oxygen \( \rightarrow \) sodium oxide
\[ 4\text{Na(s)} + \text{O}_2(g) \rightarrow 2\text{Na}_2\text{O(s)} \]

3 ▶ a FALSE  b TRUE  c TRUE  d FALSE, the formula of lithium chloride is LiCl

4 ▶ a More dense. Density increases down the group.  b 1
c Edexcelium hydroxide and hydrogen.
d More reactive than francium. Reactivity increases down the group.
Reactivity increases as the size of the atom increases.
a An edexcelium atom is larger than a francium atom.
The outer electron would be more easily lost from edexcelium – it is less strongly attracted by the nucleus.
e $2\text{Ed} + 2\text{H}_2\text{O} \rightarrow 2\text{EdOH} + \text{H}_2$
f Alkaline, because $\text{OH}^-$ ions are formed.
g $\text{Ed}_2\text{O}$

CHAPTER 12

1 ▶ a

![F]

b A fluorine atom has 7 electrons in its outer shell – the group number indicates the number of outer shell electrons.

c $\text{H}^+ + \text{F}^-$

d There are weaker intermolecular forces of attraction between fluorine molecules than between chlorine molecules. Weaker forces require less energy to break. The intermolecular forces are weaker because fluorine has a lower relative molecular mass.
e No reaction because chlorine is less reactive than fluorine and therefore cannot displace it.

2 ▶ a

![At]

Astatine atom

![At−]

Astatide ion

![At]![At]

Astatine molecule

b Astatine would be a solid – melting point increases down the group and iodine is a solid.
c Reactivity decreases down the group so astatine would be less reactive than iodine.
Astatine is a larger atom than iodine and so has a lower attraction for electrons.
d pH 1–2
e Caesium astatide would be a colourless/white solid because it is an ionic compound. It would probably be soluble in water because many ionic compounds are soluble in water and most of the alkali metal halides are soluble in water (e.g. sodium chloride, potassium iodide etc.).

f $\text{Cl}_2(\text{aq}) + 2\text{At}^- (\text{aq}) \rightarrow \text{At}_2(\text{s}) + 2\text{Cl}^- (\text{aq})$
The chlorine is reduced in this reaction because it gains electrons. The $\text{At}^-$ is oxidised in the reaction because it loses electrons. Oxidation and reduction both occur, therefore it is a redox reaction.

3 ▶ a

$\text{Cl}_2(\text{g}) + \text{F}_2(\text{g}) \rightarrow 2\text{ClF}(\text{g})$

b

![F] ![Cl]

c $\text{F}_2 < \text{ClF} < \text{Cl}_2$
The boiling point increases as the relative molecular mass increases and the strength of the intermolecular forces increases.
d i $\text{ClF}_3$
ii $\text{ClF}_3(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 3\text{HF}(\text{aq}) + \text{HCl}(\text{aq}) + \text{O}_2(\text{g})$

CHAPTER 13

1 ▶ a 78.1%  b 21.0 %  c 0.04%  d 0.9%
2 ▶ a $95 - 80 = 15 \text{cm}^3$ of oxygen $\frac{15}{95} \times 100 = 15.8\%$
b The answer is less than expected. The apparatus was not left long enough; there was not enough time for all the oxygen in the air to react. The experiment could be improved by leaving the apparatus set up for longer.
3 ▶ a The student should follow a procedure such as:
• Weigh a boiling tube.
• Pour 1 g of the metal carbonate into the boiling tube and weigh the boiling tube again.
• Connect a delivery tube to the boiling tube.
Connect the other end of the delivery tube to a gas syringe or put in to a measuring cylinder filled with water.
• Heat the boiling tube.
• Allow the apparatus to cool.
• Record the volume of gas collected.
• Re-weigh the boiling tube and contents.
b Carbon dioxide.
c $\text{PbCO}_3(\text{s}) \rightarrow \text{PbO}(\text{s}) + \text{CO}_2(\text{g})$
d Thermal decomposition.
e Sodium carbonate is more thermally stable than the other carbonates; it does not decompose when heated.
f The $\text{M}_r$ of $\text{PbCO}_3$ is 267, whereas the $\text{M}_r$ of $\text{CuCO}_3$ is 123.5.
1 g of $\text{CuCO}_3$ is more moles than 1 g of $\text{PbCO}_3$ – more than twice as many moles.
When more moles of solid are heated, more moles of gas will be produced, so a greater volume of gas will be produced.
g The mass goes down because $\text{CO}_2$ is given off. In the first experiment the mass goes down by 0.15 g, which must be the mass of $89\text{cm}^3$ of gas.
CHAPTER 14

1. a) sodium, aluminum, iron, copper
   b) i) magnesium oxide, copper
      ii) Mg(s) + CuO(s) → MgO(s) + Cu(s) (Include state symbols in all equations for preference.)
      iii) Copper(II) oxide (CuO) has been reduced to copper (Cu) because it has lost oxygen - reduction is the loss of oxygen.
      iv) Copper(II) oxide (CuO) is the oxidizing agent because it oxidizes the magnesium (gives oxygen to it) and is, in the process, reduced.
   c) i) Zinc is higher in the reactivity series because it takes the oxygen from the cobalt(II) oxide
      ii) A reducing agent is a substance which reduces something else. Zinc removes oxygen from the cobalt(II) oxide. Removal of oxygen is reduction.
      iii) Zinc because it gains oxygen - oxidation is gain of oxygen.
   d) aluminium, manganese, chromium (Statement 1: Al is above Cr. Statement 2: Mn is below Al. Statement 3: Mn is above Cr. Putting this together gives the final list.)

2. a) oxidised; gain of oxygen
   b) reduced; loss of oxygen
   c) oxidised; loss of electrons
   d) reduced; gain of electrons

3. Magnesium is above lead because it removes the oxygen from the lead(II) oxide.

4. a) Either: grey iron filings become coated with brown solid. Or: solution fades from blue to colourless (very pale green).
   b) iron (Fe) has been oxidised (to Fe_{2+}) because it has lost electrons - oxidation is loss of electrons.
   c) Fe(s) + CuSO_{4}(aq) → FeSO_{4}(aq) + Cu(s)

5. a) nickel, copper, silver
   b) i) Ni(s) + CuSO_{4}(aq) → NiSO_{4}(aq) + Cu(s)
      ii) Ni(s) + Cu^{2+}(aq) → Ni^{2+}(aq) + Cu(s)
      iii) Nickel has been oxidised by loss of electrons.

6. a) X is between iron and hydrogen in the reactivity series. It displaces hydrogen from dilute hydrochloric acid, and copper from copper(II) sulfate, and so must be above hydrogen and copper. It won’t displace iron from iron(II) sulfate, and so must be below iron.
   b) i) yes: X nitrate and silver
      ii) no reaction
      iii) no reaction (at least, not in the short term)
      iv) yes: X chloride and copper
     v) yes: X sulfate and hydrogen

7. a) hydrogen
   b) aluminium chloride
   c) 2Al(s) + 6HCl(aq) → 2AlCl_{3}(aq) + 3H_{2}(g)
   d) Aluminium is covered by a very thin, but very strong, layer of aluminium oxide which prevents the acid getting at the aluminium underneath. On heating, the acid reacts with the oxide and removes it. The aluminium then shows its true reactivity, and produces a vigorous reaction.

8. Drop a very small piece into cold water. If it reacts, judge its reactivity relative to K, Na, Ca or Mg.
   If it doesn’t react, add a small piece to dilute hydrochloric acid and warm if necessary. Rapid reaction in the cold would place it as ‘similar to magnesium’. A few bubbles of hydrogen in the cold, but more on heating would place it as ‘similar to iron or zinc’.
   If there is no reaction, then it is ‘below hydrogen’.

9. a) The iron/steel must be exposed to oxygen (air) and water.
   b) Painting prevents the oxygen and water from coming into contact with the iron/steel.
   c) It is iron/steel coated with zinc.
   d) It would take much longer before the car went rusty. Zinc is more reactive than iron and so corrodes in preference to the iron. The zinc prevents iron rusting even when the coating is scratched. When paint is scratched the steel underneath will rust – this will not happen when the steel is galvanised, therefore the life of the car is extended.

10. a) Nickel has been oxidised by loss of electrons.

CHAPTER 15

1. Iron is below carbon in the reactivity series and so can be extracted by heating with carbon.
Because lead is below iron in the reactivity series it is a reaction. Aluminium must be extracted by electrolysis. $\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$

2 a Sodium is high in the reactivity series, and so will need to be produced by electrolysis. It can be manufactured by electrolyzing molten sodium chloride using suitable electrodes (industrially, carbon anode, steel cathode), keeping the sodium and chlorine apart to prevent them from reacting back to give sodium chloride again. Sodium would be released at the cathode: $\text{Na}^+(l) + \text{e}^- \rightarrow \text{Na}(l)$

b It needed the discovery of electricity.

c Any other metals high in the reactivity series – e.g. potassium, lithium, calcium, magnesium.

3 Because lead is below iron in the reactivity series it is also below carbon. Lead can be extracted from its ore by heating with carbon.

4 a Chromium(III) oxide contains the $\text{Cr}^{3+}$ ion. The oxide ion has the formula O$^{2-}$, therefore the formula of chromium(III) oxide is $\text{Cr}_2\text{O}_3$.

$\text{Cr}_2\text{O}_3 + 2\text{Al} \rightarrow 2\text{Cr} + \text{Al}_2\text{O}_3$

b Al is the reducing agent because it takes the O away from the $\text{Cr}_2\text{O}_3$ – it reduces it.

c Chromium is less reactive than aluminium because the aluminium is able to take the oxygen away from the chromium oxide.

5 a $2\text{FeTiO}_3 + 7\text{Cl}_2 + 6\text{C} \rightarrow 2\text{TiCl}_4 + 2\text{FeCl}_3 + 6\text{CO}$

b $\text{TiCl}_4 + 2\text{Mg} \rightarrow \text{MgCl}_2 + \text{Ti}$

c Magnesium is more reactive than titanium as it is able to take the chlorine away from the TiCl$\text{}_4$; the magnesium is able to reduce the TiCl$\text{}_4$.

d Magnesium is more reactive than carbon, so it could be extracted from its ore by electrolysis or by heating with a more reactive metal such as sodium.

e Titanium is more expensive than magnesium. Although it is expensive to extract magnesium using electrolysis, the magnesium is used in the extraction of titanium, therefore titanium must be more expensive.

6 a i It has a low density. (Note: saying that it has a low density is usually better than just saying it is light)

ii Alloys are stronger than the pure metals. Alloys contain some atoms of different sizes; this makes it more difficult for the layers of positive ions to slide over each other.

iii Advantages:
An aluminium body is lighter than a steel body and so the car will use less fuel. An aluminium body will not corrode.
Disadvantage:
Aluminium is more expensive than steel.

b i a good conductor of electricity and ductile (can be drawn out into wires)

ii does not react with water and malleable (can be hammered into shape)

c High-carbon steel is brittle and not very malleable so it would be difficult to make it into the correct shapes for car bodies. It would also not bend/deform very well during a crash – the fact that mild steel deforms upon impact (because it is malleable) is important for absorbing some of the force upon impact (from a Physics point of view, the change in momentum occurs over a longer period of time).

### CHAPTER 16

<table>
<thead>
<tr>
<th>methyl orange</th>
<th>phenolphthalein</th>
<th>litmus</th>
</tr>
</thead>
<tbody>
<tr>
<td>yellow</td>
<td>pink</td>
<td>blue</td>
</tr>
<tr>
<td>red</td>
<td>colourless</td>
<td>red</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>solution</th>
<th>pH</th>
<th>strongly acidic</th>
<th>weakly acidic</th>
<th>neutral</th>
<th>weakly alkaline</th>
<th>strongly alkaline</th>
</tr>
</thead>
<tbody>
<tr>
<td>potassium iodide</td>
<td>7</td>
<td>✔</td>
<td></td>
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<td></td>
<td></td>
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<tr>
<td>propanoic acid</td>
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<td></td>
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<td>✔</td>
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<tr>
<td>potassium hydroxide</td>
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<td></td>
<td></td>
<td></td>
<td>✔</td>
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<tr>
<td>iron(III) chloride</td>
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<td></td>
<td></td>
<td></td>
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<tr>
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<td>1.3</td>
<td>✔</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

3 a $\text{Na}_2\text{O} + 2\text{HNO}_3 \rightarrow 2\text{NaNO}_3 + \text{H}_2\text{O}$

b $2\text{KOH} + \text{H}_2\text{SO}_4 \rightarrow \text{K}_2\text{SO}_4 + 2\text{H}_2\text{O}$

4 Neutralisation – reaction a) involves an acid reacting with a base (sodium oxide) and reaction b) involves an acid reacting with an alkali (KOH).

5 Measure out 25.0 cm$^3$ of potassium hydroxide solution using a pipette. Transfer the potassium hydroxide solution to a conical flask. Add a few drops of an indicator. Put the sulfuric acid into the burette. Add the acid to the alkali until the indicator changes colour.

### CHAPTER 17

1 a copper(II) oxide, copper(II) hydroxide, copper(II) carbonate

b $\text{CuO(s)} + \text{H}_2\text{SO}_4(aq) \rightarrow \text{CuSO}_4(aq) + \text{H}_2\text{O(l)}$ $\text{Cu(OH)}_2(s)$ $\text{H}_2\text{SO}_4(aq) \rightarrow \text{CuSO}_4(aq) + 2\text{H}_2\text{O(l)}$ $\text{CuCO}_3(s) + \text{H}_2\text{SO}_4(aq) + \text{CO}_2(g) + \text{H}_2\text{O(l)}$

2 a $\text{A} = \text{copper; B} = \text{copper(II) oxide}; \text{C} = \text{copper(II) sulfate}; \text{D} = \text{copper(II) carbonate}; \text{E} = \text{carbon dioxide}; \text{F} = \text{copper(II) nitrate}$

b i $\text{CuO(s)} + \text{H}_2\text{SO}_4(aq) \rightarrow \text{CuSO}_4(aq) + \text{H}_2\text{O(l)}$

ii $\text{CuCO}_3(s) + 2\text{HNO}_3(aq) \rightarrow \text{Cu(NO}_3)_2(aq) + \text{CO}_2(g) + \text{H}_2\text{O(l)}$

3 a i Description could include: Silvery metal. Colourless acid. Slow bubbles of gas on warming. Formation of green solution. Possible disappearance of nickel if very small quantities were used.
Pops with lighted splint held to mouth of tube.

ii Hydrogen

iii Na(s) + H₂SO₄(aq) → NaSO₄(aq) + H₂(g)

b i Description could include: Green solid. Colourless acid. Bubbles of gas produced in the cold. Formation of green solution. Green powder disappears if excess acid is used.

Testing for gas: Bubble through lime water to give a white precipitate (milky, cloudy).

ii NiCO₃(s) + 2HCl(aq) → NiCl₂(aq) + CO₂(g) + H₂O(l)

iii CO₃²⁻(s) + 2H⁺(aq) → CO₂(g) + H₂O(l)

Some teachers may prefer to write this as:

NiCO₃(s) + 2H⁺(aq) → Ni²⁺(aq) + CO₂(g) + H₂O(l)

4 ▶ soluble

sodium chloride
zinc nitrate
iron(III) sulfate
potassium sulfate
aluminum nitrate
ammonium chloride
magnesium nitrate
sodium phosphate
potassium dichromate(VI)

5 ▶ a Dilute sulfuric acid in a beaker/flask. Heat gently. Add copper(II) oxide a little at a time until no more reacts. Filter into an evaporating basin. Evaporate gently until a sample will crystallise on cooling. Leave to crystallise. Separate and dry crystals.

b i CuO(s) + H₂SO₄(aq) → CuSO₄(aq) + H₂O(l)

ii CuSO₄(aq) + 5H₂O(l) → CuSO₄·5H₂O(s)

6 ▶ a using a pipette: Need to be able to measure the volume of sodium carbonate solution accurately so that exactly the same volume can be used later without the indicator.

Few drops of methyl orange were added: An indicator to show when the solution becomes ‘neutral’. (In fact the colour change happens around pH 4. The pH changes very quickly around the end point for the indicator, and the difference in volume of acid added to take the pH from the desired pH 7 to the actual pH 4 for the indicator will be a fraction of a drop.)

Until the solution became orange: Exactly the right volume of acid has been added (the ‘neutral’ colour for methyl orange).

Without the methyl orange: Otherwise the crystals would be contaminated by methyl orange.

Evaporated until... Shows when the solution is sufficiently concentrated to crystallise on cooling. If a sample will crystallise, so will the bulk of the solution.

Left to cool... Crystals form on cooling because the solubility of the sodium sulfate is lower in the cold.

b i Na₂CO₃(aq) + H₂SO₄(aq) → Na₂SO₄(aq) + CO₂(g) + H₂O(l)

ii Na₂SO₄(aq) + 10H₂O(l) → Na₂SO₄·10H₂O(s)

7 ▶ a A solution of any soluble silver salt (silver nitrate is always used) plus a solution of any soluble carbonate including dilute hydrochloric acid

Ag⁺(aq) + Cl⁻(aq) → AgCl(s).

b A solution of any soluble calcium salt (calcium chloride or nitrate) plus a solution of a soluble carbonate (sodium, potassium or ammonium carbonate).

Ca²⁺(aq) + CO₃²⁻(aq) → CaCO₃(s)

c A solution of a soluble lead(II) salt (most commonly, lead(II) nitrate) plus a solution of any soluble sulfate including dilute sulfuric acid

Pb²⁺(aq) + SO₄²⁻(aq) → PbSO₄(s).

d A solution of a soluble lead(II) salt (most commonly, lead(II) nitrate) plus a solution of any soluble chloride including dilute hydrochloric acid.

Pb²⁺(aq) + Cl⁻(aq) → PbCl₂(s)

8 ▶ Mix solutions of barium chloride or barium nitrate and sodium carbonate, potassium carbonate or ammonium carbonate. Filter, wash and dry the precipitate.

Ba₂⁺(aq) + CO₃²⁻(aq) → BaCO₃(s)

9 ▶ a A dilute sulfuric acid + solid zinc (or zinc oxide, hydroxide or carbonate)

H₂SO₄(aq) + Zn(s) → ZnSO₄(aq) + H₂(g)

or H₂SO₄(aq) + ZnO(s) → ZnSO₄(aq) + H₂O(l)

or H₂SO₄(aq) + Zn(OH)₂(s) → ZnSO₄(aq) + 2H₂O(l)

or H₂SO₄(aq) + ZnCO₃(s) → ZnSO₄(aq) + H₂O(l) + CO₂(g)

b C: solutions of a soluble barium salt + soluble sulfate (including H₂SO₄)

Ba₂⁺(aq) + SO₄²⁻(aq) → BaSO₄(s)

There is no point in writing full equations for precipitation reactions like this. The ionic equation is always easier.

c B: dilute nitric acid + potassium hydroxide or carbonate solution

HNO₃(aq) + KOH(aq) → KNO₃(aq) + H₂O(l)

or 2HNO₃(aq) + K₂CO₃(aq) → 2KNO₃(aq) + H₂O(l) + CO₂(g)

d A: dilute nitric acid + solid copper(II) oxide or hydroxide or carbonate.

2HNO₃(aq) + CuO(s) → Cu(NO₃)₂(aq) + H₂O(l)

or 2HNO₃(aq) + Cu(OH)₂(s) → Cu(NO₃)₂(aq) + 2H₂O(l)

or 2HNO₃(aq) + CuCO₃(s) → Cu(NO₃)₂(aq) + H₂O(l) + CO₂(g)

(Unlike the other simple acids, nitric acid also reacts with copper metal, but it doesn’t produce hydrogen, and is not on the Edexcel International GCSE syllabus. Any equation using copper to produce hydrogen should be disallowed.)

e C: solutions of a soluble lead(II) salt (normally the nitrate) + a soluble chromate(VI). (Since you don’t know anything specifically about chromates, choose one of sodium, potassium or ammonium chromate (VI) – these are bound to be soluble because all sodium, potassium and ammonium salts are soluble.)

Pb²⁺(aq) + CrO₄²⁻(aq) → PbCrO₄(s)

10 ▶ a acid: H₂O⁺  base: CO₃²⁻

b acid: H₂SO₄  base: MgO

c acid: HNO₃  base: NH₃
**CHAPTER 18**

**1**
- a. chlorine
- b. ammonia
- c. carbon dioxide
- d. hydrogen
- e. oxygen

**2**
- a. Clean a nichrome or platinum wire by dipping it into concentrated hydrochloric acid and then into a flame until no colour shows. Moisten the wire with concentrated hydrochloric acid, dip it into the solid, and then back into flame. Lithium ions give a red flame colour.
- b. Warm the solid very gently with sodium hydroxide solution. Test any gases given off with a piece of damp litmus paper. If it turns blue, ammonia is being given off from an ammonium compound.
- c. Make a solution in pure water. Add dilute hydrochloric acid + barium chloride solution. A white precipitate shows the sulfate ions.
- d. Add dilute hydrochloric acid or dilute nitric acid. (Not sulfuric acid which forms an insoluble layer of calcium sulfate around the calcium carbonate, which stops the reaction.) Look for a colourless odourless gas turning limewater milky.
- e. Make a solution in pure water. Add dilute nitric acid + silver nitrate solution. A yellow precipitate shows the iodide ions.

**3**
- a. A = iron(III) chloride; B = iron(III) hydroxide; C = silver chloride
- b. B: Fe^{3+}(aq) + 3OH^−(aq) → Fe(OH)_3(s) or: FeCl_3(aq) + 3NaOH(aq) → Fe(OH)_3(s) + 3NaCl(aq)
- c. C: Ag^+(aq) + Cl^−(aq) → AgCl(s) or 3AgNO_3(aq) + FeCl_3(aq) → 3AgCl(s) + Fe(NO_3)_3(aq)

**4**
- a. D = iron(II) sulfate; E = iron(II) hydroxide; F = barium sulfate
- b. E: Fe^{2+}(aq) + 2OH^−(aq) → Fe(OH)_2(s) or: FeSO_4(aq) + 2NaOH(aq) → Fe(OH)_2(s) + Na_2SO_4(aq)
- c. F: Ba^{2+}(aq) + SO_4^{2−}(aq) → BaSO_4(s) or BaCl_2(aq) + MgSO_4(aq) → BaSO_4(s) + MgCl_2(aq)

**5**
- a. G = potassium carbonate; H = potassium nitrate; I = carbon dioxide (Note: g could also be potassium hydrogen carbonate, but, because hydrogen carbonates haven’t been covered in the course, a student is unlikely to come up with it.)
- b. CO_2^{2−}(s) + 2H^+(aq) → CO_2(g) + H_2O(l) or K_2CO_3(s) + 2HNO_3(aq) → 2KNO_3(aq) + CO_2(g) + H_2O(l)

**6**
- a. J = ammonium bromide; K = ammonia; L = silver bromide
- b. Ag^+(aq) + Br^−(aq) → AgBr(s)

**7**
- a. Dissolve the mixture in the minimum possible amount of hot water, and then allow it to cool again. The small amount of the more soluble potassium carbonate will stay in solution, but crystals of potassium nitrate will be formed on cooling. Filter these off, and wash them on the filter paper with a small amount of very cold water. Allow them to dry.
- b. Take a sample and add any named dilute acid. If the crystals are free of potassium carbonate, there won’t be any fizzing.

**8**
- a. Add a few drops of the liquid to some anhydrous copper(II) sulfate. If the colour changes from white to blue then water is present.
- b. The boiling point of the liquid must be measured. If the liquid boils at 100 °C it is pure water. The freezing/melting point can also be measured – pure water freezes/melts at 0°C.
- c. The test-tube becoming warm indicates that an exothermic reaction has occurred. An exothermic reaction with sodium hydroxide solution could mean that the colourless liquid is an acid – neutralisation reactions are exothermic. The fact that there is no other visual reaction indicates that there is no metal present that has an insoluble hydroxide.

**END OF UNIT 2 QUESTIONS**

**1**
- a. They all have 1 electron in the outer shell. (1)
- b. Potassium oxide. (1)
- c. The piece of lithium floats; hydrogen gas is formed; the final solution is alkaline.
- d. i. it forms a 1+ ion in compounds; it reacts with water and air
   ii. The density increases from lithium to sodium but decreases from sodium to potassium (1)
   - The trend is not clear for the whole group (1)

**2**
- a. Because there will be no reaction (1)

**ANSWERS**

<table>
<thead>
<tr>
<th>Compound</th>
<th>Reaction with Hydrochloric Acid</th>
<th>Reaction with Sodium Hydroxide</th>
<th>Reaction with Copper(II) Sulfate</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH_3COOH</td>
<td>Orange solution formed</td>
<td>No reaction</td>
<td>Brown solution formed</td>
</tr>
<tr>
<td>H_2SO_4</td>
<td>Orange solution formed</td>
<td>No reaction</td>
<td>Brown solution formed</td>
</tr>
<tr>
<td>HCl</td>
<td>No reaction</td>
<td>No reaction</td>
<td>No reaction</td>
</tr>
<tr>
<td>NH_3</td>
<td>No reaction</td>
<td>No reaction</td>
<td>No reaction</td>
</tr>
<tr>
<td>Fe^{2+}(aq) + 2OH^−(aq)</td>
<td>Fe(OH)_2(s)</td>
<td>Fe(OH)_2(s) + Na_2SO_4(aq)</td>
<td>Fe(OH)_2(s) + Na_2SO_4(aq)</td>
</tr>
<tr>
<td>Ba^{2+}(aq) + SO_4^{2−}(aq)</td>
<td>BaSO_4(s)</td>
<td>BaSO_4(s) + MgCl_2(aq)</td>
<td>BaSO_4(s) + MgCl_2(aq)</td>
</tr>
</tbody>
</table>
362 ANSWERS

4  ▶  a Put dilute sulfuric acid into a beaker and heat it on a tripod and gauze using a Bunsen burner. (1)
   Add excess copper(II) oxide (1)
   Filter off the excess copper(II) oxide (1)

b Test: add hydrochloric acid followed by barium chloride solution (1)
   Positive result: white precipitate (1)

c Copper(II) hydroxide (1)

4  ▶  d i Cl₂(aq) + 2Br⁻(aq) → 2Cl⁻(aq) + Br₂(aq)
   Cl⁻(aq) (1)
   Br₂(aq) (1)
   balancing (1)

   e Chlorine gains electrons to form chloride ions (1)
   Bromide ions lose electrons to form bromine (1)
   Gain of electrons is reduction and loss of electrons is oxidation (1)

5  ▶  a An ore is a sample of rock that contains enough of a mineral for it to be worthwhile to extract the metal. (1)

   b

<table>
<thead>
<tr>
<th>thermal decomposition</th>
<th>neutralisation</th>
<th>redox</th>
</tr>
</thead>
<tbody>
<tr>
<td>reaction 1</td>
<td></td>
<td>✓</td>
</tr>
<tr>
<td>reaction 3</td>
<td>✓</td>
<td></td>
</tr>
</tbody>
</table>

   c MoO₃ + 3C → Mo + 3CO
      products (1)
      balancing (1)

   d Carbon is the reducing agent (1)
      It reduces the MoO₃ by removing the oxygen (1)

   e Molybdenum is below carbon in the reactivity series (1)
      Carbon must be more reactive than molybdenum to be able to take the oxygen away from molybdenum oxide. (1)
6 ▶ a Mix together solutions (1) of lead(II) nitrate and sodium sulfate (1). Filter the mixture (1) Wash the residue with distilled/pure water (1); Leave the solid in a warm oven to dry. (1)
b Because lead(II) iodide is yellow and insoluble in water. (1)
The precipitate could just be lead(II) iodide. (1)

c i

![Image of a graph showing a line of best fit]

all point plotted correctly (2) −1 for each mistake line of best fit (1)
ii the point at 6 minutes (1); 10.8 ± 0.2 cm³ (1)
iii as the time increases the volume of gas increases (1)
directly proportional relationship (1)
just stating that the volume is directly proportional to the time scores (2)
iv test: burning splint (1)
positive result: squeaky pop (1)

7 ▶ a

\[ \text{H}_2\text{Cl} \]

*Dot and cross in the covalent bond.* (1)
*All other electrons.* (1)

b A proton/hydrogen ion/H⁺ is transferred from the HCl to the water. (1)
HCl is an acid because it donates a proton/hydrogen ion/H⁺ (1)
H₂O is a base because it accepts a proton/hydrogen ion/H⁺ (1)

c

<table>
<thead>
<tr>
<th>Reagent</th>
<th>Solid or solution?</th>
<th>Salt</th>
<th>Other product</th>
</tr>
</thead>
<tbody>
<tr>
<td>magnesium</td>
<td>solid</td>
<td>hydrochloric</td>
<td>magnesium</td>
</tr>
<tr>
<td>copper(II) oxide</td>
<td>solid</td>
<td>hydrochloric</td>
<td>copper(II)</td>
</tr>
<tr>
<td>sodium hydroxide</td>
<td>solution</td>
<td>hydrochloric</td>
<td>sodium</td>
</tr>
<tr>
<td>silver nitrate</td>
<td>solution</td>
<td>hydrochloric</td>
<td>silver</td>
</tr>
</tbody>
</table>

|                              |                    |               |               |
|                              | magnesium chloride | hydrochloric  | hydrogen       |
| copper(II) chlorine          | copper(II) chloride| sodium chloride| water         |
| sodium chloride              | sodium chloride    | water         |
| silver nitrate               | silver chloride    | nitric acid   |

d i a white precipitate is formed (1)
ii hydrochloric acid contains chloride ions (1)

Will always give a positive test for chloride ions. (1)

e \[ \text{CuCO}_3(s) + 2\text{HCl(aq)} \rightarrow \text{CuCl}_2(aq) + \text{H}_2\text{O(l)} + \text{CO}_2(g) \]
reactants (1) products (1) balance (1)

8 ▶ a iron(II) (1) chloride (1)
b \[ B = \text{iron(II) hydroxide (1) Fe(OH)}_2 (1) \]
c \[ D = \text{iron(III) hydroxide (1) Fe(OH)}_3 (1) \]
d redox (1)
The iron(II) is oxidised to iron(III) (1)

9 ▶ a \[ \text{Na}_2\text{SO}_4 (1) \]
b Sodium hydroxide solution/sulfuric acid is measured out using a pipette and put into conical flask/beaker. (1)
Sulfuric acid/sodium hydroxide placed in burette. (1)
c Energy is needed to break the bonds in heptane and in oxygen. Energy is released when new bonds are made to produce the carbon dioxide and water. More energy is released when the new bonds are made than was used to break the old ones.

2 ▶ a A reaction in which heat energy is absorbed from the surroundings.

\[
\text{CaO(s) + H}_2\text{O(l)} \rightarrow \text{Ca(OH)}_2(\text{s})
\]

b ▶

\[
\Delta H = +180 \text{kJ/mol}
\]

N\textsubscript{2} + O\textsubscript{2} \rightarrow 2\text{NO}

\[
\text{progress of the reaction}
\]

\[
\text{energy}
\]

\[
\Delta H = +180 \text{kJ/mol}
\]

3 ▶ a exothermic b exothermic c endothermic d exothermic e endothermic f exothermic

4 ▶ The cans have two chambers. The outer one contains the beverages or food to be heated. The inner chamber contains calcium oxide and water. The calcium oxide and water are kept separate by a seal. Pushing a button on the bottom of the can breaks the seal and allows the calcium oxide and water to come into contact. You get the reaction below which is exothermic.

CaO(s) + H\textsubscript{2}O(l) → Ca(OH\textsubscript{2})(s)

5 ▶ a Bonds broken: \(4 \times \text{C–H} = 4 \times (+413) = +1652 \text{kJ}\)
\(1 \times \text{Br–Br} = +193 \text{kJ}\)
\(\text{total} = +1845 \text{kJ}\)

Bonds made: \(3 \times \text{C–H} = 3 \times (–413) = -1239 \text{kJ}\)
\(1 \times \text{C–Br} = –290 \text{kJ}\)
\(1 \times \text{H–Br} = –366 \text{kJ}\)
\(\text{total} = –1895 \text{kJ}\)

Overall change = +1845 – 1895 = –50 kJ (exothermic)

b Bonds broken: \(1 \times \text{H–H} = +436 \text{kJ}\)
\(1 \times \text{Cl–Cl} = +243 \text{kJ}\)
\(\text{total} = +679 \text{kJ}\)

Bonds made: \(2 \times \text{H–Cl} = 2 \times (–432) = -864 \text{kJ}\)
\(\text{Overall change} = +679 – 864 = –185 \text{kJ} \) (exothermic)

c Bonds broken: \(2 \times \text{H–H} = 2 \times (+436) = +872 \text{kJ}\)
\(1 \times \text{O=O} = +498 \text{kJ}\)
\(\text{total} = +1370 \text{kJ}\)

Bonds made: \(4 \times \text{O–H} = 4 \times (–464) = –1856 \text{kJ}\)
\(\text{Overall change} = +1370 – 1856 = –486 \text{kJ} \) (exothermic)

d Bonds broken: \(1 \times \text{N≡N} = +944 \text{kJ}\)
\(3 \times \text{H–H} = 3 \times (+436) = +1308 \text{kJ}\)
\(\text{total} = +2252 \text{kJ}\)

Bonds made: \(6 \times \text{N–H} = 6 \times (–388) = –2328 \text{kJ}\)
\(\text{Overall change} = +2252 – 2328 = –76 \text{kJ} \) (exothermic)

6 ▶ a Her first two results weren’t reliable. There was too much difference between them.

b Any two from: Wear a lab coat; avoid touching the hot copper can; tie long hair back.

c Heat energy change = \(Q = mc\Delta T\)

Mass of water being heated = \(m = 100 \text{g}\)

c is the specific heat capacity of the water (it is the water that is being heated)
\(c = 4.18 \text{J/g°C}\)

Temperature change of water = \(\Delta T = 55.0 – 19.0 = 36.0 \text{°C}\)

Heat evolved = \(Q = mc\Delta T = 100 \times 4.18 \times 36.0 \text{J} = 15.048 \text{kJ} \) to 3 significant figures

d Mass of hexane burnt = \(35.62 – 35.23 = 0.39 \text{g}\)

Heat evolved per gram = \(15.048 \text{kJ} = 38.6 \text{kJ/g} \) to 3 significant figures

The combustion reaction is exothermic, therefore heat energy released per gram = –38.6 kJ/g

e Number of moles of hexane, \(\text{C}_6\text{H}_{14}\) burnt = \(\frac{\text{mass (m)}}{\text{relative molecular mass (Mr)}} = \frac{0.39}{86} = 0.00453 \text{mol}\)

The molar enthalpy change of combustion of hexane (\(\Delta H\))

heat energy change (Q) = \(\frac{\text{number of moles of combustion of hexane} (\Delta H)}{\text{mass of hexane burnt (m)}} = \frac{15.048}{0.00453} = 3320 \text{kJ/mol to 3 significant figures}\)

The combustion reaction is exothermic, therefore \(\Delta H = -3320 \text{kJ/mol}\)

f Any two from: Misreading one of the weighings of the spirit burner so that it looked as if less hexane had been burnt than was really the case; misreading the thermometer to give a final temperature higher than it should have been; adding less than 100 cm\(^3\) of water to the flask, so that the temperature went up more than it should because the heat was going into a smaller volume of water.

g Any two from: Heat loss to the surroundings; heat is lost to warm up the copper calorimeter or the thermometer; incomplete combustion of the fuel.

7 ▶ a Heat energy change = \(Q = mc\Delta T\)

Mass of solution being heated = \(m = 50 \text{g}\), the mass of the lithium chloride is relatively small and it is ignored in the calculation.

\(c\) is the specific heat capacity of the diluted solution of lithium chloride, which we assume to be the same as the heat capacity of water \(c = 4.18 \text{J/g°C}\)

Temperature change of water = \(\Delta T = 33.5 – 17.0 = 16.5 \text{°C}\)

Heat evolved = \(Q = mc\Delta T = 100 \times 4.18 \times 36.0 \text{J} = 3448.5 \text{kJ} = 34.5 \text{kJ} \) to 3 significant figures

Mass of lithium chloride = \(0.00453 \text{mol}\)

Relative molecular mass (Mr) = \(0.39 \text{kJ/mol}\)

Dissolving of lithium chloride is exothermic, \(\Delta H = -28.5 \text{kJ/mol}\)

b Number of moles of lithium chloride dissolved = \(\frac{\text{mass (m)}}{\text{relative formula mass (Mr)}} = \frac{0.121}{42.5} = 0.00453 \text{mol}\)

The enthalpy change of solution (\(\Delta H\))

heat energy change (Q) = \(\frac{\text{number of moles of lithium chloride dissolved (n)}}{\text{mass of lithium chloride dissolved (m)}} = \frac{0.121}{0.00453} = 26.5 \text{kJ/mol} \) to 3 significant figures

The dissolving of lithium chloride is exothermic, therefore \(\Delta H = -28.5 \text{kJ/mol}\)
At the very beginning of the reaction. The reaction can only happen when acid particles collide with the solid dolomite. Numbers of acid particles per unit volume are greatest at the beginning of the reaction before any get used up. Therefore the greatest number of collisions per second and the fastest reaction is at the beginning.

70 seconds (read this off the graph and allow some tolerance depending on the size of graph paper available.)

Volume produced within the first 80 seconds = 55 cm³
(The average rate = \( \frac{55}{80} \) = 0.688 cm³/s to 3 significant figures)

There would be a lower initial rate; same volume of gas.

There would be a lower initial rate; half the volume of gas (50 cm³).

The initial rate would be the same; half the volume of gas (50 cm³). (The initial rate depends on the original concentration of the acid that is still the same.)

The initial rate would be faster; same volume of gas.

Time taken for the reaction would increase. A reaction happens when acid particles collide with the magnesium. The concentration of acid is less, so there is less number of acid particles per unit volume. There will be fewer collisions per second, and therefore a slower reaction.

Time taken for the reaction would decrease. The acid particles are moving faster and so collide with the magnesium more often. Reaction only happens if the energy of the collision equals or exceeds activation energy. At higher temperatures more acid particles have energy greater than or equal to the activation energy, so there will be more successful collisions per unit time.

CHAPTER 20

Collection of gas over water into an inverted measuring cylinder is an acceptable alternative. Pieces of dolomite are put in a weighing bottle into the conical flask before the reaction starts. This is to make sure that the reaction can be started without losing any carbon dioxide.

Graph should be completely smooth with the axes properly labelled.

Reactions only happen if collisions have energies equaling or exceeding activation energy. Catalysts provide an alternative route for the reaction with a lower activation energy. More particles now have energy greater than or equal to the activation energy, so there will be more successful collisions per unit time.
To find out whether it speeds the reaction up: You could do this most simply by having two test-tubes with equal volumes of the same hydrogen peroxide solution side-by-side. Add some copper(II) oxide to one and look for the faster production of bubbles.

To show that it is unchanged: Use a known mass of copper(II) oxide. When the reaction stops, filter through previously weighed filter paper, allow to dry, and re-weigh. Show that the mass of copper(II) oxide is unchanged. (If it had changed, and you hadn’t lost any during the separation process, it must have reacted in some way.)

A high pressure favours the reaction producing fewer moles of gas, shifting the position of equilibrium to the right hand side.

Very high pressures are expensive to produce for two reasons: all the reaction vessels and pipework have to be built much more strongly, and it takes a lot more energy to drive the compressors producing the very high pressures. The extra ammonia produced isn’t worth the extra cost, so a compromise of 200 atmospheres pressure is used.

At a low temperature. Lower temperature favours the forward exothermic reaction, shifting the position of equilibrium to the right hand side to produce more ammonia.

At low temperatures the reaction is extremely slow even in the presence of a catalyst. 450 °C is chosen because it gives a compromise between the rate and the yield of the reaction.

The reversible symbol ‘⇌’ shows that the reaction can go both ways. The reactants can form products and the products can react to form the reactants.

White solid decomposes to form colourless gases.

‘Dynamic’ means that the reactions are still happening and the rate of the forward reaction is equal to the rate of the reverse reaction. ‘Equilibrium’ means that the concentrations of the reactants and the products remain constant.

The proportion of sulfur trioxide will increase. An increase in pressure shifts the position of equilibrium to the right because this side has fewer moles of moles of gas molecules.

The position of equilibrium shifts to the right. A decrease in temperature favours the forward exothermic reaction.

Less carbon monoxide and hydrogen would be produced. (A high pressure favours the reverse reaction, as the left hand side of the reaction has fewer number of moles of gas molecules.)

A high temperature favours the endothermic forward reaction and shifts the position of equilibrium to the right hand side of the reaction.

Using a catalyst increases the rate of the reaction, as it provides an alternative pathway for the reaction with lower activation energy. More particles will have energy greater than or equal to the activation energy so there will be more successful collisions per unit time. More products will be made quickly and relatively low temperature and pressure can be used in the presence of a catalyst. This saves energy and money. Adding a catalyst has no effect on the position of equilibrium, as the rate of the forward reaction and the rate of the reverse reaction are increased by an equal amount.

A high pressure favours the reaction producing fewer moles of gas, shifting the position of equilibrium to the right hand side.

Very high pressures are expensive to produce for two reasons: all the reaction vessels and pipework have to be built much more strongly, and it takes a lot more energy to drive the compressors producing the very high pressures. The extra ammonia produced isn’t worth the extra cost, so a compromise of 200 atmospheres pressure is used.

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2 a i 19.3 (1), 16.6 (1), 2.7 (1)
   ii A
b i $Q = mc \Delta T = 100.0 \times 4.2 \times (23.2 - 15.9) = 3.066 \text{kJ}$
calculation of temperature change (1)
ii $n = \frac{5}{111} = 0.0450 \text{mol to 3 significant figures.}$
calculation of $M_r$ (1)
iii $\Delta H = \frac{3.066}{0.0450} = 68.1 \text{kJ/mol to 3 significant figures.}$
   The molar enthalpy change is $-68.1 \text{kJ/mol}$ as the dissolving is exothermic.
d i The gas particles have more kinetic energy (1)
   so more particles have energy greater than or equal to the activation energy (1). There are more successful collisions per unit time (1).
ii $\frac{5}{111} \times \frac{0.353}{150} = 0.0450 \text{mol to 3 significant figures.}$
calculation of $n$ (1)
   correct answer to 2–4 significant figures (1)
iii $\frac{3.066}{0.0450} = 68.1 \text{kJ/mol}$ as the dissolving is exothermic.
   dividing $Q$ by $n$ (1)
   correct answer (must have a negative sign) (1)
c heat loss to the surrounding air through beaker (1)
some magnesium chloride is left on the weighing boat/did not dissolve in water completely (1)

3 a $2\text{H}_2\text{O}_2(\text{aq}) \rightarrow 2\text{H}_2\text{O}(l) + \text{O}_2(\text{g})$
correct formulae of reactants and products (1)
correct balancing (1)
b To prevent the loss of oxygen at the beginning of the reaction. (1)
c
correct labelling of axes with units (1)
correct points plotted (2)
a smooth curve of best fit, going through all of the points (1)
d i 125 seconds
ii 43 cm$^3$
iii $\frac{53}{150} = 0.353 \text{cm}^3/\text{sec}$
correct numerical answer (1)
correct unit (1)
e Reaction has stopped (1) because all the hydrogen peroxide is used up (1).
f i B (1)
ii Weigh a sample of manganese(IV) oxide and add to hydrogen peroxide (1). Oxygen is produced at a faster rate with manganese(IV) oxide than without (1). Filter the reaction mixture and dry the solid (1).

4 a i the mass (1) and the surface area / the size (1).
b To let the gas escape but keep the liquid inside the conical flask. (1)
c i B
ii $\frac{600}{0\times120} = \frac{5}{150} = 0.353 \text{cm}^3/\text{sec}$
c correct labelling of axes with units (1)
c correct points plotted (2)
a smooth curve of best fit, going through all but one point (1)
iii 150 seconds
iv C

5 a i The gas particles have more kinetic energy (1)
   so more particles have energy greater than or equal to the activation energy (1). There are more successful collisions per unit time (1).
ii The rate increases (1) as gas particles are closer together and there are more frequent collisions (or more successful collisions per unit time) (1).
iii Reactions happen on the surface of the catalyst (1) and gauzes have greater surface area (1).

b The use of a catalyst increases the rate of the reaction without itself being used up (1). It saves money which would be spent on increasing temperature or pressure (1).

6 a reversible reaction (1), enthalpy change (1)
b The forward reaction is endothermic. (1)
c

Re-weigh the solid and the mass should be the same as before if it acts as a catalyst (1).
g Shallower curve than the original (1), but the end volume remains the same (1).
h Shallower curve than the original (1) and only 30 cm$^3$ (half the volume) of gas is produced (1).
d i Increasing the temperature favours the forward endothermic reaction. (1)
The position of equilibrium shifts to the right. (1)
The reaction mixture becomes more colourless. (1)
i i Increasing the pressure favours the left hand side of the reaction as it has fewer moles of moles of gases. (1)
The position of equilibrium shifts to the left. (1)
The reaction mixtures becomes more brown. (1)
e No change in the position of equilibrium (1) The use of a catalyst increases the rate of both the forward and reverse reaction equally (1).

7 a i Increasing the pressure favours the right hand side of the reaction which has fewer moles of moles of gas molecules. (1)
The position of equilibrium shifts to the right. (1)
The yield of NH₃ increases. (1)
ii As temperature increases, the percentage of NH₃ at equilibrium decreases. (1)
iii \(\Delta H\) for the forward reaction is negative. (1)
Increasing temperature favours the reverse endothermic reaction. (1)
iv Increasing temperature increases the rate of reaction. (1)
More particles will have energy greater than or equal to the activation energy. (1)
The frequency of successful collisions increases. (1)

UNIT 4 ANSWERS

CHAPTER 22

1 a i methane
ii propane
iii hexane
iv propene
v ethene
vi but-1-ene
vii propan-1-ol
viii butan-2-ol
b i

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{C} \quad \text{C} \quad \text{H}
\end{align*}
\]

ii

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{C} \quad \text{C} \quad \text{C} \\
\text{OH} & \quad \text{H} \quad \text{H}
\end{align*}
\]

iii

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{C} \quad \text{C} \quad \text{C} \\
\text{C} & \quad \text{C} \quad \text{H} \\
\text{H} & \quad \text{H} \quad \text{H}
\end{align*}
\]

d i

\[
\begin{align*}
\text{H} & \quad \text{C} \quad \text{C} \quad \text{C} \\
\text{H} & \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H}
\end{align*}
\]

v

\[
\begin{align*}
\text{H} & \quad \text{C} \quad \text{O} \quad \text{H} \\
\text{H} & \quad \text{H}
\end{align*}
\]

vi

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{C} \quad \text{C} \\
\text{H} & \quad \text{C} \quad \text{H} \\
\text{H} & \quad \text{H}
\end{align*}
\]

vii

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{C} \quad \text{C} \\
\text{C} & \quad \text{C} \quad \text{H} \\
\text{H} & \quad \text{C} \quad \text{H} \\
\text{H} & \quad \text{H}
\end{align*}
\]

viii

\[
\begin{align*}
\text{H} & \quad \text{C} \quad \text{C} \\
\text{C} & \quad \text{C} \quad \text{H} \\
\text{H} & \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H}
\end{align*}
\]

2 a The existence of molecules with the same molecular formula but different structural formulae.
b

\[
\begin{align*}
\text{CH}₃ & \quad \text{CH}₂ \quad \text{CH}₂ \quad \text{CH}₂ \quad \text{CH}₂ \\
\text{CH}₃ & \quad \text{CH} \quad \text{CH} \quad \text{CH}₂ \quad \text{CH}_3
\end{align*}
\]

butane 2-methylpropane
(In these and subsequent formulae, if you aren’t asked specifically for displayed formulae, these quicker forms are acceptable.)
c

\[
\begin{align*}
\text{CH}_3 & \quad \text{CH}_2 \quad \text{CH}_2 \quad \text{CH}_2 \quad \text{CH}_3 \\
\text{CH}_3 & \quad \text{CHCH}_2 \quad \text{CH}_2 \quad \text{CH}_3 \\
\text{CH}_3 & \quad \text{CH}_2 \quad \text{CH}_2 \quad \text{CH}_2 \quad \text{CH}_3
\end{align*}
\]

hexane 2-methylpentane
3-methylpentane 2,3-dimethylbutane 2,2-dimethylbutane
d

\[
\begin{align*}
\text{CH}_3 & \quad \text{CH}_2 \quad \text{CH} \quad \text{CH} \quad \text{CH} \\
\text{CH}_3 & \quad \text{CH} \quad \text{CH} \quad \text{CH} \quad \text{CH} \\
\text{CH}_3 & \quad \text{CH}_2 \quad \text{CH} \quad \text{CH} \quad \text{CH} \\
\text{CH} & \quad \text{CH} \quad \text{CH} \quad \text{CH} \quad \text{CH}
\end{align*}
\]

but-1-ene but-2-ene (2-methyl(propene))
e

\[
\begin{align*}
\text{CH}_2 & \quad \text{CH}_2 \\
\text{CH}_2 & \quad \text{CH}_3 \\
\text{CH} & \quad \text{CH}
\end{align*}
\]

2,3-dimethylbutane
3 ▶ a  i (first answer is shown in the textbook as an example)
   ii \( \text{CH}_3\text{CH}_2\text{CH}_3 \)
   iii \( \text{CH}_3 \text{CH}(\text{CH}_3)_2 \)
   iv \( \text{CH}_3\text{CCH}_3 \)
   v \( \text{CH}_3\text{OCH}_2\text{CH}_2\text{CH}_3, \text{CH}_3\text{CH}_2\text{OCH}_2\text{CH}_3, \text{CH}_3\text{OCHCH}_3 \)

b These compounds are known as ethers.

4 ▶ a addition
   b combustion
   c addition
   d substitution
   e combustion
   f substitution

5 ▶ a C, it contains element other than carbon and hydrogen.
   b
   c
   d \( \text{C}_n\text{H}_{2n+2} \)
   e Same functional group / similar chemical properties;
     Shows a gradation in physical properties; Each member differs from the next by a \(-\text{CH}_2-\).
   f \( \text{C}_6\text{H}_{12} + 8\text{O}_2 \rightarrow 5\text{CO}_2 + 6\text{H}_2\text{O} \)

CHAPTER 24

1 ▶ a Contains only C-C single bonds and has no double or triple bonds.
   b i \( \text{C}_{11}\text{H}_{24} \)
      ii Liquid
   iii \( \text{C}_{11}\text{H}_{24}(l) + 17\text{O}_2(g) \rightarrow 11\text{CO}_2(g) + 12\text{H}_2\text{O}(l) \)
   iv \( 2\text{C}_{11}\text{H}_{24} + 23\text{O}_2 \rightarrow 22\text{CO} + 24\text{H}_2\text{O} \)
   Carbon monoxide is poisonous, as it reduces the ability of the blood to carry oxygen around the body.

2 ▶ a The sulfur (or sulfur compound) burns to make sulfur dioxide. The sulfur dioxide reacts with water and oxygen in the atmosphere to produce sulfuric acid that falls as acid rain.
   b The spark in the engine causes nitrogen to react with oxygen to give various oxides of nitrogen.
   c Sulfur dioxide reacts with water and oxygen in the atmosphere to produce sulfuric acid. The sulfuric acid in acid rain can react with calcium carbonate and corrode the buildings.

3 ▶ a Crude oil produces too many larger hydrocarbons / not enough of the more useful smaller ones. Smaller alkanes can be used as fuel for cars. Cracking also produces alkenes that can be used to make polymers.
   b Heat the vaporised fraction in the presence of a silicon dioxide or aluminium oxide catalyst at 600-700 °C.
   c \( \text{C}_{11}\text{H}_{24} \rightarrow 2\text{C}_2\text{H}_4 + \text{C}_7\text{H}_{16} \)
   d Any other valid cracking equation starting with \( \text{C}_{11}\text{H}_{24} \).
     For example:
     \( \text{C}_{11}\text{H}_{24} \rightarrow \text{C}_2\text{H}_4 + \text{C}_9\text{H}_{20} \)
     or \( \text{C}_{11}\text{H}_{24} \rightarrow \text{C}_2\text{H}_4 + \text{C}_3\text{H}_6 + \text{C}_6\text{H}_{14} \)
     or lots of other variants. In each case, at least one hydrocarbon should be an alkane (\( \text{C}_n\text{H}_{2n+2} \)), and at least one an alkene (\( \text{C}_n\text{H}_{2n} \)).
b Any 2 of the following:

4 ▶ a UV light
   b Substitution
   c

5 ▶ a C₄H₁₀ + Cl₂ → C₄H₉Cl + HCl
   b CH₃CH₂CH₂CH₂Cl  CH₃CH(OH)CH₃
      1-chlorobutane  2-chlorobutane

CHAPTER 25

1 ▶ a i Any two from: same general formula; same functional group or similar chemical properties; shows a gradation in physical properties; each member differs from the next by a –CH₂–.
   ii Compounds containing one or more carbon carbon double or triple bond.
   b i Starting: orange; Finishing: colourless
   ii

2 ▶ a a carbon-carbon double bond
   b
   c
   d i CH₃CH₃ + Br₂(g) → CH₃CH₂Br + HBr(g)
      ii In this (substitution) reaction, one of the hydrogens has been replaced by a bromine during the reaction. In the previous (addition) reaction, nothing was lost when the two molecules combined together.

3 ▶ a 1,2-dichloropropane CH₂ClCHClCH₃
   b ethane CH₃CH₃

4 ▶ a Cracking.
   Conditions: 1) The presence of a silicon dioxide (silica) or aluminium oxide (alumina) catalyst 2) heat the vaporized alkanes at high temperatures of 600 – 700 °C.
   b Test: shake with bromine water
   Result with propane: no colour change (solution remains orange)
   Result with propene: solution changes colour from orange to colourless

5 ▶ a CH₃CH₂CH₂CH₂OH or CH₃CH₂CH(OH)CH₃
      butan-1-ol  butan-2-ol
   b CH₃CH₂CH₂CH₂Cl or CH₃CH₂CHClCH₃
      1-chlorobutane  2-chlorobutane
   c CH₃CH₂CHBrCH₂Br
      1,2-dibromobutane

CHAPTER 26

1 ▶ a i CH₃OH
   ii CH₃CH₂CH₂OH
3 ▶ a

Orange to green.

b Propan-1-ol is very flammable and should not be heated with an open flame. The vapour can easily catch fire when heated with a Bunsen burner.

d i bubbles/fizzes/effervescence; Mg decreases in size

\[
\text{Mg(aq)} + 2\text{CH}_3\text{COOH(aq)} \rightarrow (\text{CH}_3\text{COO})_2\text{Mg(aq)} + \text{H}_2(\text{g})
\]

d ii bubbles/fizzes/effervescence

\[
\text{Na}_2\text{CO}_3(\text{aq}) + 2\text{CH}_3\text{COOH(}a\text{q)} \rightarrow 2\text{CH}_3\text{COONa(}a\text{q)} + \text{CO}_2(\text{g}) + \text{H}_2\text{O(l)}
\]

4 ▶ a \(\text{CH}_3\text{COOH} + \text{NaHCO}_3 \rightarrow \text{CH}_3\text{COONa} + \text{CO}_2 + \text{H}_2\text{O}\)

b Neutralisation.

c Bubbles will form.

### CHAPTER 28

1 ▶ a i \(\text{CH}_3\text{COOCH}_2\text{CH}_3\)

ii \(\text{CH}_3\text{CH}_2\text{CH}_2\text{COOCH}_2\text{CH}_3\text{CH}_3\)

b i Add 1 cm³ of ethanoic acid and 1 cm³ of ethanol into a boiling tube, then add a few drops of concentrated sulfuric acid.

• Warm at about 80 °C.

• Allow the contents of the tube to cool. When cool, pour the mixture into a beaker containing sodium carbonate solution.

• A layer of ester will separate and float on top of the water.

• Repeat with other combinations of acid and alcohol.

c Esters can be detected by smelling the product: gently waft the odour towards nose. For example, propyl ethanoate smells like pear.

d Food flavourings and perfumes.

3 ▶ a Alcohol Carboxylic acid

<table>
<thead>
<tr>
<th>Alcohol</th>
<th>Carboxylic acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>Methanol</td>
<td>Propanoic acid</td>
</tr>
<tr>
<td>(\text{CH}_3\text{OH})</td>
<td>(\text{CH}_3\text{CH}_2\text{COOH})</td>
</tr>
<tr>
<td>(\text{CH}_3\text{CH}_2\text{CH}_2\text{OH})</td>
<td>HCOOH</td>
</tr>
</tbody>
</table>

b Propan-1-ol Methanoic acid

\(\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}\) HCOOH
4. **a**

\[
\begin{array}{c}
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{C} \\
\text{C} \\
\text{O} \\
\text{C} \\
\text{H} \\
\text{H} \\
\end{array}
\]

**b** Ethanoic acid.

**c** By heating with potassium dichromate(VI) in dilute sulfuric acid.

**d** \( \text{CH}_3\text{COOH}(l) + \text{CH}_3\text{CH}_2\text{OH}(l) \rightarrow \text{CH}_3\text{COOCH}_2\text{CH}_3(l) + \text{H}_2\text{O}(l) \)

**e** Heat at about 80 °C with a few drops of concentrated sulfuric acid as a catalyst.

5. **a** Test 1: Take out 3 cm³ of each of the three samples and put them into three separate test-tubes. Add a few pieces of Mg or a spatula tip worth of Na₂CO₃ solid into each of the test-tubes. The one that bubbles contains ethanoic acid.

Test 2: To the remaining two samples, take out another 3 cm³ of each and put them into two separate boiling tubes. Add a few drops of potassium dichromate(VI) in dilute sulfuric acid. Heat the boiling tubes in a warm water bath. The one that changes colour from orange to green contains ethanol.

### CHAPTER 29

1. **a** Unsaturated: containing one or more carbon carbon double or triple bond.

   Hydrocarbon: compounds containing carbon and hydrogen only.

   **b** \( \text{C}_2\text{H}_4 + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 2\text{H}_2\text{O} \)

   **c** Ethene can be used to make addition polymer polyethene, which is used for making plastic bags, milk bottles etc.

2. **a** Joining up of lots of little molecules (the monomers) to make one big one (a polymer).

**b**

\[
\begin{array}{c}
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{C} \\
\text{C} \\
\end{array}
\]

**c** (The “continuation” bonds at each end are an important part of this structure. Marks will be lost in an exam if they are omitted.)

\[
\begin{array}{c}
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{C} \\
\text{C} \\
\text{C} \\
\text{C} \\
\text{C} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\end{array}
\]

**d** Joining two or more molecules together without anything being lost in the process.

**e**

\[
\begin{array}{c}
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{C} \\
\text{C} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\end{array}
\]

**f** Drawing the molecule to show its relationship with the structure of the polymer as drawn in the question:

\[
\begin{array}{c}
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{C} \\
\text{C} \\
\text{O} \\
\text{C} \\
\text{H} \\
\text{H} \\
\text{O} \\
\text{O} \\
\end{array}
\]

**g** Either: buried in landfill sites because there would be no emission of poisonous carbon monoxide or no carbon dioxide to contribute to global warming or no sulfur dioxide to contribute to acid rain. Or: incineration because no space is filled up for landfills or energy can be generated to provide heat for office buildings.

### Additional Content

**3.**

**a** Condensation polymer

**b**

\[
\begin{array}{c}
\text{H} \\
\text{O} \\
\text{C} \\
\text{C} \\
\text{O} \\
\text{H} \\
\end{array}
\]

and hexandioic acid as:

\[
\begin{array}{c}
\text{H} \\
\text{O} \\
\text{C} \\
\text{O} \\
\text{H} \\
\end{array}
\]

**4.**

\[
\begin{array}{c}
\text{H} \\
\text{O} \\
\text{C} \\
\text{C} \\
\text{O} \\
\text{C} \\
\text{C} \\
\text{C} \\
\text{O} \\
\text{C} \\
\text{C} \\
\text{C} \\
\end{array}
\]

**5.**

**a** Drawing 1,6-diaminohexane as:

\[
\begin{array}{c}
\text{H} \\
\text{N} \\
\text{N} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\end{array}
\]

and hexandioic acid as:

\[
\begin{array}{c}
\text{H} \\
\text{O} \\
\text{C} \\
\text{O} \\
\text{H} \\
\end{array}
\]

**i** In a condensation reaction, when two molecules join together a small molecule is lost in the process. When the two molecules above join together, a molecule of water is lost every time they come together. In a condensation polymerisation reaction, this happens repeatedly.

**ii**

\[
\begin{array}{c}
\text{H} \\
\text{O} \\
\text{C} \\
\text{N} \\
\text{N} \\
\text{H} \\
\text{H} \\
\text{H} \\
\end{array}
\]

\[
\begin{array}{c}
\text{H} \\
\text{O} \\
\text{C} \\
\text{O} \\
\text{H} \\
\end{array}
\]

etc
There will be four more CH₂ groups in the “box” in the part of the diagram which comes from the dioic acid.

The bonding between the two monomers will be identical, as will the size of the diamine.

**END OF UNIT 4 QUESTIONS**

1 **a** D (1)

- Used as liquefied petroleum gas / domestic heating or cooking. (1)

- Gasoline. (1)

- Crude oil is separate into fractions by fractional distillation. (1)

  - It is heated until it boils and the vapour enters a fractionating column (1), which is cooler at the top and hotter at the bottom (1). The fractions with a lower boiling point condense nearer to the top of the column and can be tapped off / the fractions with a higher boiling point condense nearer to the bottom of the column and can be tapped off (1).

- As the number of carbon atoms increase, the boiling point of a hydrocarbon increases. (1)

  - This is because the intermolecular forces increases so it takes more energy to break them during boiling. (1)

2 **a** E contains an element that is not carbon or hydrogen. (1)

- 1,3-dibromopropane  1 mark for dibromopropane

- Poly(chloroethene) or polyvinylchloride (1)

- They are inert. (1)

  - Or non-biodegradable. (1)

  - Or if they are burned, they release toxic gases including carbon monoxide or hydrogen chloride. (1)

- No colour change / bromine water remains orange. (1)

  - Polymer of compound A is saturated / contains no C=C bond. (1)

3 **a** C  H  O

<table>
<thead>
<tr>
<th>Mass (in 100g)</th>
<th>55</th>
<th>9</th>
<th>36</th>
</tr>
</thead>
<tbody>
<tr>
<td>Number of moles</td>
<td>4.58</td>
<td>9</td>
<td>2.25</td>
</tr>
<tr>
<td>Ratio</td>
<td>2.25</td>
<td>2.25</td>
<td>2.25</td>
</tr>
</tbody>
</table>

The empirical formula is C₃H₆O.
b  
\[2 \times 12 + 4 \times 1 + 1 \times 16 = 44\]  
\[
\frac{44}{2} = 2 
\]  
The molecular formula is \(C_4H_8O_2\).

c  
\[
\text{butanoic acid (1)} 
\]

d  
\[
\text{Methylbutanoate (l)} 
\]

4 a Test ‘Add bromine water’  
Result with compound 1 Solution changes colour from orange to colourless  
Result with compound 2 Solution stays orange  

b 1 mark for the correct structure; 1 mark for the continuation bonds

c Poly(but-2-ene) (1)

5 a i 1 mark for the ester linkage; 1 mark for the correct structure of butyl and ethanoate group

ii Potassium dichromate. (VI); Dilute sulfuric acid. (1); Heat. (1)

b

\[
\text{H}_2\text{CCH}_2\text{CH}_2\text{COOH} + \text{CH}_3\text{OH} \rightarrow \text{CH}_3\text{CH}_2\text{CH}_2\text{COOCH}_3 + \text{H}_2\text{O} 
\]

\[
\text{Methylbutanoate (l)} 
\]

6 a

\[
\begin{align*}
\text{H} & \quad \text{CH}_3 & \quad \text{CH}_3 & \quad \text{CH}_3 & \quad \text{b} & \quad \text{CH}_2=\text{CH} \\
\text{CH} & \quad \text{C} & \quad \text{C} & \quad \text{C} & \quad \text{C} & \quad \text{O} \\
\text{H} & \quad \text{H} & \quad \text{H} & \quad \text{O} & \quad \text{H} \\
\text{H} & \quad \text{H} & \quad \text{H} & \quad \text{C} & \quad \text{H} \\
\text{H} & \quad \text{H} & \quad \text{O} & \quad \text{C} & \quad \text{H} \\
\text{H} & \quad \text{H} & \quad \text{C} & \quad \text{C} & \quad \text{O} \\
\text{C} & \quad \text{H} & \quad \text{H} & \quad \text{C} & \quad \text{H} \\
\text{H} & \quad \text{H} & \quad \text{H} & \quad \text{H} \\
\text{b} & \quad \text{NH}_3 \\
\end{align*}
\]

b i A polyester that is biodegradable (1)

ii C

7 a i Phosphoric acid (1); 300 °C or 60–70 atm pressure. (1)

ii Bond breaking  
\[
4 \times \text{C}−\text{H} + 1 \times \text{C}=\text{C} + 2 \times \text{O−H} =
4 \times 412 + 612 + 2 \times 463 = 3186 \text{kJ/mol. (1)}
\]

Bond forming  
\[
5 \times \text{C}−\text{H} + 1 \times \text{C}−\text{C} + 1 \times \text{C}−\text{O} + 1 \times \text{O−H} =
5 \times 412 + 1 \times 348 + 1 \times 360 + 1 \times 463 = 3231 \text{kJ/mol. (1)}
\]

Overall  
\[
3186 – 3231 = −45 \text{kJ/mol. (1)}
\]

b i Fermentation (1)

ii To provide the enzyme (zymase) which speeds up the reaction. (1)

iii In the presence of oxygen, yeast would produce carbon dioxide and water rather than ethanol. (1)

iv Warm climate or plenty of land to grow sugar canes (1); Lack of oil resources to produce ethane. (1)

c i Dilute sulfuric acid (1); Heating. (1)

ii \(\text{CH}_3\text{CH}_2\text{OH} + 2\text{[O]} \rightarrow \text{CH}_3\text{COOH} + \text{H}_2\text{O}\)

iii Orange (1) to green. (1)

iv \(\text{Mg(s)} + 2\text{CH}_3\text{COOH(aq)} \rightarrow (\text{CH}_3\text{COO})_2\text{Mg(aq)} + \text{H}_2(\text{g})\)

1 mark for the correct formulae of products; 1 mark for correct balancing.

d i Water is formed in the reaction. (1)

ii Pour the reaction mixture into water (or sodium carbonate solution) (1); The alcohol and the acid will dissolve but the ester will float on the surface of water. (1)