

Worked solutions

Chapter 1

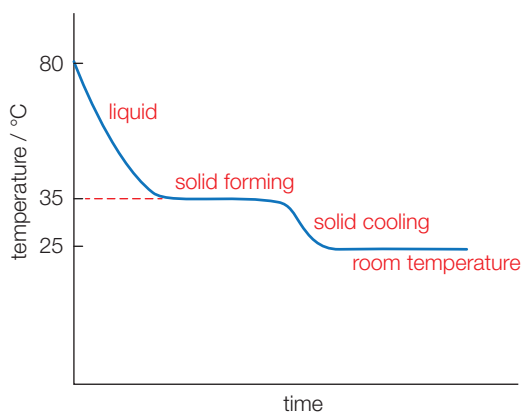
Exercises

- 1 For each of these questions (a) to (e):
- write the information from the question in the form of an equation
 - check the number of atoms on each side of the equation
 - introduce coefficients in front of the formulae in order to ensure that there are equal numbers of atoms on each side of the equation.
- (a) $\text{CuCO}_3 \rightarrow \text{CuO} + \text{CO}_2$
- (b) $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$
- (c) $\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$
- (d) $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$
- (e) $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$
- 2 For each of these questions (a) to (e):
- introduce coefficients in front of each formula to ensure that there are equal numbers of atoms on each side of the equation.
- (a) $2\text{K} + 2\text{H}_2\text{O} \rightarrow 2\text{KOH} + \text{H}_2$
- (b) $\text{C}_2\text{H}_5\text{OH} + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O}$
- (c) $\text{Cl}_2 + 2\text{KI} \rightarrow 2\text{KCl} + \text{I}_2$
- (d) $4\text{CrO}_3 \rightarrow 2\text{Cr}_2\text{O}_3 + 3\text{O}_2$
- (e) $\text{Fe}_2\text{O}_3 + 3\text{C} \rightarrow 3\text{CO} + 2\text{Fe}$
- 3 For each of these questions (a) to (e):
- introduce coefficients in front of each formula to ensure that there are equal numbers of atoms on each side of the equation.
- (a) $2\text{C}_4\text{H}_{10} + 13\text{O}_2 \rightarrow 8\text{CO}_2 + 10\text{H}_2\text{O}$
- (b) $4\text{NH}_3 + 5\text{O}_2 \rightarrow 4\text{NO} + 6\text{H}_2\text{O}$
- (c) $3\text{Cu} + 8\text{HNO}_3 \rightarrow 3\text{Cu}(\text{NO}_3)_2 + 2\text{NO} + 4\text{H}_2\text{O}$
- (d) $6\text{H}_2\text{O}_2 + 2\text{N}_2\text{H}_4 \rightarrow 2\text{N}_2 + 10\text{H}_2\text{O} + \text{O}_2$
- (e) $4\text{C}_2\text{H}_7\text{N} + 15\text{O}_2 \rightarrow 8\text{CO}_2 + 14\text{H}_2\text{O} + 2\text{N}_2$
- 4 (a) Sand is an insoluble solid and water a liquid – heterogeneous.
- (b) Smoke is made up of solid particles dispersed in air (a gas) – heterogeneous.
- (c) Sugar dissolves in water to give a clear solution – homogeneous. (If it is a saturated solution with excess sugar that cannot dissolve, the overall mixture is then heterogeneous.)
- (d) Salt and iron filings mix but don't interact with each other – heterogeneous.
- (e) Ethanol dissolves in water to give a clear solution – homogeneous.
- (f) Steel consists of an alloy of iron and carbon, it has the same properties throughout – homogeneous.
- 5 For each of questions (a) to (e):
- introduce coefficients in front of each formula to ensure that there are equal numbers of atoms on each side of the equation
 - remember when assigning state symbols that if there is water present and one of the products is soluble, then the symbol aq (aqueous) must be used (water itself as a liquid is always (l), never (aq)).
- (a) $2\text{KNO}_3(\text{s}) \rightarrow 2\text{KNO}_2(\text{s}) + \text{O}_2(\text{g})$
- (b) $\text{CaCO}_3(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{CaSO}_4(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
- (c) $2\text{Li}(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{LiOH}(\text{aq}) + \text{H}_2(\text{g})$
- (d) $\text{Pb}(\text{NO}_3)_2(\text{aq}) + 2\text{NaCl}(\text{aq}) \rightarrow \text{PbCl}_2(\text{s}) + 2\text{NaNO}_3(\text{aq})$
- (e) $2\text{C}_3\text{H}_6(\text{g}) + 9\text{O}_2(\text{g}) \rightarrow 6\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{l})$
- 6 X has diffused more quickly, so it must be a lighter gas. Its particles have greater velocity than the particles of Y at the same temperature. (They will, however, both have the same average kinetic energy.)
- 7 From the kinetic molecular theory we would expect a solid to be more dense than its liquid. We would expect that ice would sink in water. That ice floats is an indication that something

else is involved in the structure of ice (see page 152).

- 8 Bubbles will be present through the volume of the liquid. A brown gas is visible above the brown liquid. As the two states are at the same temperature, the particles have the same average kinetic energy and are moving at the same speed. The inter-particle distances in the gas are significantly larger than those in the liquid.
- 9 At certain conditions of low temperature and low humidity, snow changes directly to water vapour by sublimation, without going through the liquid phase.
- 10 Steam will condense on the skin, releasing energy as it forms liquid at the same temperature (e–d on Figure 1.4). This energy is in addition to the energy released when both the boiling water and the condensed steam cool on the surface of the skin.
- 11 B, as a change of state is taking place.

12



- 13 Use $L = 6.02 \times 10^{23} \text{ mol}^{-1}$.
- (a) 1 mole of $\text{C}_2\text{H}_5\text{OH}$ contains 6 moles of hydrogen atoms
 1 mole of $\text{C}_2\text{H}_5\text{OH}$ contains $6 \times (6.02 \times 10^{23})$ hydrogen atoms
 1 mole of $\text{C}_2\text{H}_5\text{OH}$ contains 3.61×10^{24} hydrogen atoms
 0.020 moles therefore contains $(3.61 \times 10^{24}) \times 0.020$ hydrogen atoms
 $= 7.2 \times 10^{22}$ hydrogen atoms
- (b) 1 mole of H_2O contains 2 moles of hydrogen atoms

1 mole of H_2O contains $2 \times (6.02 \times 10^{23})$ hydrogen atoms

1 mole of H_2O contains 1.20×10^{24} hydrogen atoms

2.50 moles therefore contains $(1.20 \times 10^{24}) \times 2.50$ hydrogen atoms
 $= 3.01 \times 10^{22}$ hydrogen atoms

- (c) 1 mole of $\text{Ca}(\text{HCO}_3)_2$ contains 2 moles of hydrogen atoms
 1 mole of $\text{Ca}(\text{HCO}_3)_2$ contains $2 \times (6.02 \times 10^{23})$ hydrogen atoms
 1 mole of $\text{Ca}(\text{HCO}_3)_2$ contains 1.20×10^{24} hydrogen atoms
 0.10 moles therefore contains $(1.20 \times 10^{24}) \times 0.10$ hydrogen atoms
 $= 1.2 \times 10^{23}$ hydrogen atoms

- 14 Propane contains three carbon atoms and eight hydrogen atoms. If the three carbon atoms are equivalent to 0.20 moles of carbon then one carbon atom would be equivalent to $0.20/3$ moles of carbon. So eight atoms of hydrogen would be equivalent to $(0.20/3) \times 8$ moles of hydrogen, i.e. 0.53 moles of H.

- 15 Sulfuric acid contains four oxygen atoms. If there are 6.02×10^{23} atoms of oxygen in total then there must be $(6.02 \times 10^{23})/4$ molecules of sulfuric acid, i.e. 1.51×10^{23} molecules of sulfuric acid (= 0.250 mol of sulfuric acid).

- 16 (a) Magnesium phosphate, $\text{Mg}_3(\text{PO}_4)_2$

Element	Relative atomic mass	Number of atoms of each element	Relative mass
Mg	24.31	3	72.93
P	30.97	2	61.94
O	16.00	8	128.00
Molar mass			262.87 g mol ⁻¹

(b) Ascorbic acid, $C_6H_8O_6$

Element	Relative atomic mass	Number of atoms of each element	Relative mass
C	12.01	6	72.06
H	1.01	8	8.08
O	16.00	6	96.00
Molar mass			176.14 g mol ⁻¹

(c) Calcium nitrate, $Ca(NO_3)_2$

Element	Relative atomic mass	Number of atoms of each element	Relative mass
Ca	40.08	1	40.08
N	14.01	2	28.02
O	16.00	6	96.00
Molar mass			164.10 g mol ⁻¹

(d) Hydrated sodium thiosulfate, $Na_2S_2O_3 \cdot 5H_2O$

Element	Relative atomic mass	Number of atoms of each element	Relative mass
Na	22.99	2	45.98
S	32.07	2	64.14
O	16.00	8	128.00
H	1.01	10	10.10
Molar mass			248.22 g mol ⁻¹

17 Calculate the molar mass of calcium arsenate, $Ca_3(AsO_4)_2$

Element	Relative atomic mass	Number of atoms of each element	Relative mass
Ca	40.08	3	120.24
As	74.92	2	149.84
O	16.00	8	128.00
Molar mass			398.08 g mol ⁻¹

$$\text{Mass} = nM = 0.475 \text{ mol} \times 398.08 \text{ g mol}^{-1} = 189.1 \text{ g}$$

18 (If not using a calculator, use rounded values for A_r .)

$$M \text{ of } CO_2 = (12 + (16 \times 2)) \text{ g mol}^{-1} = 44 \text{ g mol}^{-1}$$

$$\text{moles} = \frac{m}{M} = \frac{66 \text{ g}}{44 \text{ g mol}^{-1}} = 1.5 \text{ mol}$$

19 Copper(II) chloride, $CuCl_2$, has M of $(63.55 + (35.45 \times 2)) \text{ g mol}^{-1} = 134.45 \text{ g mol}^{-1}$

$$0.50 \text{ g is equivalent to } (0.50 \text{ g}/134.45 \text{ g mol}^{-1}) \text{ mol of copper chloride, i.e. } 3.7 \times 10^{-3} \text{ mol}$$

There are two chloride ions in copper chloride, $CuCl_2$

There must be $2 \times (3.7 \times 10^{-3}) \text{ mol}$ of chloride ions present, i.e. $7.4 \times 10^{-3} \text{ mol}$ (= 0.0074 mol)

20 36.55 g of carbon = $36.55 \text{ g}/12.01 \text{ g mol}^{-1} = 3.043 \text{ mol}$ of carbon

1 mole of carbon contains 6.02×10^{23} atoms of carbon

Therefore 3.043 moles of carbon contain $3.043 \times (6.02 \times 10^{23})$ atoms of carbon, i.e. 1.83×10^{24} atoms

21 (If not using a calculator, use rounded values for A_r .)

Calculate the M_r of sucrose, $C_{12}H_{22}O_{11}$

Element	Relative atomic mass	Number of atoms of each element	Relative mass
carbon	12	12	144
hydrogen	1	22	22
oxygen	16	11	176
M			342 g mol ⁻¹

$$\text{Mass} = nM = 0.500 \text{ mol} \times 342 \text{ g mol}^{-1} = 171 \text{ g}$$

22 (If not using a calculator, use rounded values for A_r .)

Water: the M_r of H_2O is 18 (= $16 + 2 \times 1$)

Therefore 10.0 g of water is equivalent to $(10 \text{ g}/18 \text{ g mol}^{-1}) \text{ mol}$ of water (= 0.55 mol)

Mercury: the relative atomic mass of mercury is 201

Therefore 10.0 g is equivalent to (10 g/201 g mol⁻¹) mol of mercury (≈ 0.05 mol)

10.0 g of water contains more particles than 10.0 g of mercury.

- 23** (If not using a calculator, use rounded values for A_r .)

M_r of N_2H_4 is $(2 \times 14) + (4 \times 1) = 32$, therefore 1.0 mol has a mass of 32 g

M_r of N_2 is $(2 \times 14) = 28$, therefore 2.0 mol has a mass of 56 g

M_r of NH_3 is $14 + (3 \times 1) = 17$, therefore 3.0 mol has a mass of 51 g

M_r of H_2 is $(2 \times 1) = 2$, therefore 25.0 mol has a mass of 50 g

So in order of decreasing order of mass:

2.0 mol nitrogen > 3.0 mol ammonia > 25.0 mol hydrogen > 1.0 mol hydrazine

- 24** (a) C_2H_2 : the ratio of carbon to hydrogen atoms can be simplified to CH
 (b) $C_6H_{12}O_6$: the ratio of atoms can be simplified to CH_2O
 (c) $C_{12}H_{22}O_{11}$: the ratio of atoms cannot be simplified – the empirical and molecular formula are the same.
 (d) C_8H_{18} : the ratio of atoms can be simplified to C_4H_9
 (e) C_8H_{14} : the ratio of atoms can be simplified to C_4H_7
 (f) CH_3COOH , i.e. $C_2H_4O_2$: the ratio of atoms can be simplified to CH_2O

	Sodium	Sulfur	Oxygen
mass / g	0.979	1.365	1.021
moles	$\frac{0.979}{22.99} = 0.0426$	$\frac{1.365}{32.07} = 0.0426$	$\frac{1.021}{16.00} = 0.06381$
divide by smallest	1.00	1.00	1.50
nearest whole number ratio	2	2	3

The empirical formula is $Na_2S_2O_3$

26

	Cobalt	Sulfur	Oxygen	Water (H ₂ O)
mass / g	2.10	1.14	2.28	4.50
moles	$\frac{2.10}{58.93} = 0.0356$	$\frac{1.14}{32.07} = 0.0355$	$\frac{2.28}{16.00} = 0.143$	$\frac{4.50}{18.02} = 0.250$
divide by smallest	1.00	1.00	4.03	7.04
nearest whole number ratio	1	1	4	7

The empirical formula is $CoSO_4 \cdot 7H_2O$

27

	Carbon	Hydrogen	Nitrogen
% by mass	83.89	10.35	5.76
moles	$\frac{83.89}{12.01} = 6.985$	$\frac{10.35}{1.01} = 10.2$	$\frac{5.76}{14.01} = 0.411$
divide by smallest	17.0	24.8	1.00
nearest whole number ratio	17	25	1

The empirical formula is $C_{17}H_{25}N$

28

M_r of $NH_3 = 14.01 + (3 \times 1.01) = 17.04$

% by mass of N is $\frac{14.01}{17.04} \times 100 = 82.22\%$

M_r of $CO(NH_2)_2 = 12.01 + 16.00 + 2 \times [14.01 + (2 \times 1.01)] = 62.07$

% by mass of N is $\frac{28.02}{62.07} \times 100 = 45.14\%$

(Note: 28.02 since there are two nitrogen atoms in the molecule)

M_r of $(NH_4)_2SO_4 = (2 \times 14.01) + (8 \times 1.01) + 32.07 + (4 \times 16.00) = 132.17$

% by mass of N is $\frac{28.02}{132.17} \times 100 = 21.20\%$

(Note: 28.02 since there are two nitrogen atoms in the molecule)

So overall, ammonia, NH_3 , has the highest % by mass of nitrogen.

29

moles of nitrogen = $0.673 \text{ g} / 14.01 \text{ g mol}^{-1} = 0.0480 \text{ mol}$

In the formula there are 3 moles of nitrogen associated with each mole of metal. Therefore moles of metal in the compound = $3 \times 0.0480 = 0.144$

$$\text{atomic mass} = \frac{\text{mass}}{\text{moles}} = \frac{1.00 \text{ g}}{0.144 \text{ g mol}^{-1}} = 6.94 \text{ g mol}^{-1}$$

The relative atomic mass of the element is 6.94. By looking at the periodic table (section 6 of the IB data booklet), it can be seen that the element is lithium.

$$30 \quad \text{percentage by mass} = \frac{\text{relative atomic mass of cadmium}}{M_r} \times 100$$

- For CdS, percentage by mass = $\frac{112.41}{112.41 + 32.07} \times 100 = 77.80\%$
- For CdSe, percentage by mass = $\frac{112.41}{112.41 + 78.96} \times 100 = 58.74\%$
- For CdTe, percentage by mass = $\frac{112.41}{112.41 + 127.60} \times 100 = 46.84\%$

Overall, CdS has the highest percentage by mass of cadmium.

You could also approach this question by considering the A_r of the other element in the compound. Sulfur has the lowest A_r and so CdS will have the highest percentage by mass of cadmium.

31	Carbon	Hydrogen
% by mass	$100 - 7.74 = 92.26$	7.74
moles	$\frac{92.26}{12.01} = 7.681$	$\frac{7.74}{1.01} = 7.66$
divide by smallest	1.00	1.00
nearest whole number ratio	1	1

Empirical formula is therefore CH. This has a mass of 13.02 g mol^{-1} . This number divides into the molar mass of the whole compound six times (i.e. $\frac{78.10 \text{ g mol}^{-1}}{13.02 \text{ g mol}^{-1}} = 6$).

The molecular formula is therefore six times the empirical formula, i.e. C_6H_6

32	Hydrogen	Phosphorus	Oxygen
mass / g	0.0220	0.3374	$0.8821 - (0.0220 + 0.3374) = 0.5227$
moles	$\frac{0.0220}{1.01} = 0.0218$	$\frac{0.3374}{30.97} = 0.01089$	$\frac{0.5227}{16.00} = 0.03267$
divide by smallest	2.00	1.00	3.00
nearest whole number ratio	2	1	3

Empirical formula is therefore H_2PO_3 . This has a mass of 80.99 g mol^{-1} . This number divides into the molar mass of the whole compound twice (i.e. $\frac{162 \text{ g mol}^{-1}}{80.99 \text{ g mol}^{-1}} = 2$).

The molecular formula is therefore twice the empirical formula, i.e. $\text{H}_4\text{P}_2\text{O}_6$

33	Carbon	Hydrogen	Nitrogen
mass / g	0.1927	0.02590	0.1124
moles	$\frac{0.1927}{12.01} = 0.01604$	$\frac{0.02590}{1.01} = 0.0256$	$\frac{0.1124}{14.01} = 0.008022$
divide by smallest	3.332	5.32	1.666
nearest whole number ratio	10	16	5

	Phosphorus	Oxygen
mass / g	0.1491	0.3337
moles	$\frac{0.1491}{30.97} = 0.004814$	$\frac{0.3337}{16.00} = 0.02086$
divide by smallest	1.000	4.333
nearest whole number ratio	3	13

Note:

- the mass for oxygen is obtained by subtracting all the masses of the other elements from 0.8138 g
- the nearest whole number ratio is obtained by multiplying by 3 to round everything up (numbers ending in '.33' and '.66' are the clue here).

The empirical formula is $C_{10}H_{16}N_5O_{13}P_3$. The formula mass of this is 507 g mol^{-1} so the empirical and molecular formulae are the same.

$$34 \quad \text{Moles of CO}_2 = \frac{0.66 \text{ g}}{(12.01 + (2 \times 16.00)) \text{ g mol}^{-1}} = 0.015 \text{ mol}$$

- This is the same as the number of moles of carbon atoms present.

$$\text{Moles of water} = \frac{0.36 \text{ g}}{(1.01 \times (2 + 16.00)) \text{ g mol}^{-1}} = 0.020 \text{ mol}$$

- Twice this number is the number of moles of hydrogen atoms present, i.e. 0.040 mol.

Convert these into masses in order to find the mass of oxygen in the original sample:

$$\text{mass of carbon} = 0.015 \text{ mol} \times 12.01 \text{ g mol}^{-1} = 0.18 \text{ g}$$

$$\text{mass of hydrogen} = 0.040 \text{ mol} \times 1.01 \text{ g mol}^{-1} = 0.040 \text{ g}$$

$$\text{therefore mass of oxygen} = 0.30 \text{ g} - 0.18 \text{ g} - 0.040 \text{ g} = 0.08 \text{ g}$$

Now the calculation can proceed as usual.

	Carbon	Hydrogen	Oxygen
mass / g	0.18	0.040	0.08
moles	$\frac{0.18}{12.01} = 0.015$	$\frac{0.040}{1.01} = 0.040$	$\frac{0.08}{16.00} = 0.005$
divide by smallest	3	8	1
nearest whole number ratio	3	8	1

Empirical formula is C_3H_8O

- 35
- Weigh the chalk before and after the name has been written.
 - Subtract the values to obtain the mass of chalk used.
 - Calculate the number of moles of chalk used.

Let y = the mass of chalk used in g

$$\begin{aligned} \text{moles of chalk used} &= \frac{\text{mass used}}{M_r(\text{CaCO}_3)} \\ &= \frac{y \text{ g}}{100.09 \text{ g mol}^{-1}} \end{aligned}$$

This is the same as the number of moles of carbon atoms used.

$$\text{Therefore the number of carbon atoms used} = \text{moles of chalk} \times (6.02 \times 10^{23} \text{ mol}^{-1}) = \frac{6.20 \times 10^{23} y}{100.09}$$

- 36 (a) From the stoichiometric equation 2 moles of iron can be made from 1 mole of iron oxide. Hence $2 \times 1.25 \text{ mol} = 2.50 \text{ mol}$ of iron can be made from 1.25 mol of iron oxide.

- (b) From the stoichiometric equation 2 moles of iron need 3 moles of hydrogen.

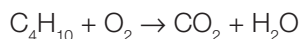
$$\text{Hence } 3.75 \text{ mol of iron need } \frac{3}{2} \times 3.75 \text{ mol} = 5.63 \text{ mol of hydrogen.}$$

- (c) From the stoichiometric equation 3 moles of water are produced from 1 mole of iron oxide.

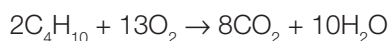
$$\text{Hence } 12.50 \text{ moles of water are produced from } \frac{1}{3} \times 12.50 \text{ mol} = 4.167 \text{ moles of iron oxide.}$$

$$4.167 \text{ moles of iron oxide have a mass of } 4.167 \times M_r(\text{Fe}_2\text{O}_3) = 4.167 \text{ mol} \times 159.70 \text{ g mol}^{-1} = 665.5 \text{ g} (= 665 \text{ g})$$

- 37 (a) Write the chemical equation:



Then balance the equation by deducing the appropriate numbers in front of the formulae:



- (b) From the equation 2 moles of butane produce 10 moles of water.

$$\frac{2.46 \text{ g}}{18.02 \text{ g mol}^{-1}} = 0.137 \text{ moles of water were produced.}$$

0.137 moles of water must have been made from $\frac{2}{10} \times 0.137 = 0.0274$ moles of butane.

0.0274 moles of butane has a mass of $0.0274 \times M_r(\text{C}_4\text{H}_{10}) = 0.0274 \text{ mol} \times 58.14 \text{ g mol}^{-1} = 1.59 \text{ g}$

38 From the equation, 1 mole of Al reacts with 1 mole of NH_4ClO_4

26.98 g of Al react with $14.01 + (4 \times 1.01) + 35.45 + (4 \times 16.00) = 117.50 \text{ g}$ of NH_4ClO_4

Therefore 1000 g of Al react with $\frac{117.50}{26.98} \times 1000 = 4355 \text{ g} = 4.355 \text{ kg}$ of NH_4ClO_4

39 (a) $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$

(b) $0.657 \text{ g of CO}_2 = \frac{n}{M} = \frac{0.657 \text{ g}}{44.01 \text{ g mol}^{-1}} = 0.0149 \text{ moles of CO}_2$

This was produced from 0.0149 moles of CaCO_3 .

0.0149 moles of CaCO_3 has a mass of $0.0149 \times M_r(\text{CaCO}_3) = 0.0149 \text{ mol} \times 100.09 \text{ g mol}^{-1} = 1.49 \text{ g}$

Therefore % of CaCO_3 in the impure limestone = $\frac{1.49 \text{ g}}{1.605 \text{ g}} \times 100 = 92.8\%$

(c) Assumptions are:

- CaCO_3 is the only source of carbon dioxide
- all the CaCO_3 undergoes complete decomposition
- all CO_2 released is captured
- heating does not cause any change in mass of any of the other minerals present.

40 (a) moles of $\text{H}_2 = \frac{12.0 \text{ g}}{2.02 \text{ g mol}^{-1}} = 5.94 \text{ mol}$

moles of $\text{CO} = \frac{74.5 \text{ g}}{28.01 \text{ g mol}^{-1}} = 2.66 \text{ mol}$

As the CO reacts with H_2 in a 1 : 2 ratio, this means that the H_2 is in excess ($2 \times 2.66 \text{ mol} = 5.32 \text{ mol}$).

2.66 mol of CO therefore produce 2.66 mol of CH_3OH .

This has a mass of $2.66 \times M_r(\text{CH}_3\text{OH}) = 2.66 \text{ mol} \times 32.05 \text{ g mol}^{-1} = 85.2 \text{ g}$

(b) moles of H_2 in excess = $5.94 \text{ mol} - (2 \times 2.66) \text{ mol} = 0.62 \text{ mol}$

This is equivalent to a mass of $0.62 \text{ mol} \times 2.02 \text{ g mol}^{-1} = 1.3 \text{ g}$

41 moles of $\text{C}_2\text{H}_4 = \frac{15.40 \text{ g}}{(2 \times 12.01) + (4 \times 1.01) \text{ g mol}^{-1}} = 0.5488 \text{ mol}$

moles of $\text{Cl}_2 = \frac{3.74 \text{ g}}{(2 \times 35.45) \text{ g mol}^{-1}} = 0.0528 \text{ mol}$

As the reactants react in the ratio of 1 : 1, C_2H_4 is in excess.

moles of product formed = 0.0528 mol

mass of product formed = $0.0528 \times M_r(\text{C}_2\text{H}_4\text{Cl}_2) = 0.0528 \text{ mol} \times 98.96 \text{ g mol}^{-1} = 5.23 \text{ g}$

42 moles of $\text{CaCO}_3 = \frac{255 \text{ g}}{(40.08 + 12.01 + (4 \times 16.00)) \text{ g mol}^{-1}} = 2.55 \text{ mol}$

moles of $\text{SO}_2 = \frac{135 \text{ g}}{(32.07 + (2 \times 16.00)) \text{ g mol}^{-1}} = 2.11 \text{ mol}$

Therefore as they react in a 1 : 1 ratio, the number of moles of CaSO_3 produced = 2.11 mol
mass of $\text{CaSO}_3 = 2.11 \text{ mol} \times (40.08 + 32.07 + (3 \times 16.00)) \text{ g mol}^{-1} = 254 \text{ g}$

Therefore percentage yield = $\frac{198 \text{ g}}{254 \text{ g mol}^{-1}} \times 100 = 77.9\%$

43 moles of CH_3COOH used = $\frac{3.58 \text{ g}}{((2 \times 12.01) + (4 \times 1.01) + (2 \times 16.00)) \text{ g mol}^{-1}} = 0.0596 \text{ mol}$

moles of $\text{C}_5\text{H}_{11}\text{OH} = \frac{4.75 \text{ g}}{((5 \times 12.01) + (12 \times 1.01) + 16.00) \text{ g mol}^{-1}} = 0.0539 \text{ mol}$

Therefore $\text{C}_5\text{H}_{11}\text{OH}$ is the limiting reagent, so 0.0539 mol of $\text{CH}_3\text{COOC}_5\text{H}_{11}$ is the maximum that can form.

This will have a mass of $0.0539 \times [(7 \times 12.01) + (14 \times 1.01) + (2 \times 16.00)] \text{ g mol}^{-1} = 7.01 \text{ g}$

This is the 100% yield, therefore 45% yield has a mass of $0.45 \times 7.01 \text{ g} = 3.16 \text{ g}$

- 44 100 g of $\text{C}_6\text{H}_5\text{Cl}$ is equivalent to

$$\frac{100 \text{ g}}{[(6 \times 12.01) + (5 \times 1.01) + 35.45] \text{ g mol}^{-1}} = 0.888 \text{ mol}$$

If this is 65% yield then 100% yield would be

$$0.888 \text{ mol} \times \frac{100}{65} = 1.37 \text{ mol}$$

1.37 moles of benzene has a mass of

$$1.37 \text{ mol} \times [(6 \times 12.01) + (6 \times 1.01)] \text{ g mol}^{-1} = 107 \text{ g}$$

- 45 (a) 1 mole of gas has a volume of 22.7 dm^3 at STP.

$$\text{Therefore } 54.5 \text{ dm}^3 \text{ is equivalent to } \frac{54.5}{22.7} \text{ mol} = 2.40 \text{ mol}$$

- (b) 1 mole of gas has a volume of 22.7 dm^3 at STP.

$$\text{This is equivalent to } 22.7 \times 1000 \text{ cm}^3 = 227\,000 \text{ cm}^3$$

Therefore 250.0 cm^3 of gas contains

$$\frac{250.0}{227\,000} \text{ mol} = 1.10 \times 10^{-3} \text{ mol} (= 0.0110 \text{ mol})$$

- (c) 1 mole of gas has a volume of 22.7 dm^3 at STP.

This is equivalent to 0.0227 m^3

$$1.0 \text{ m}^3 \text{ of gas therefore contains } \frac{1.0}{0.0227} \text{ mol} = 44 \text{ mol}$$

- 46 (a) 44.00 g of N_2 is equivalent to

$$\frac{44.00 \text{ g}}{(2 \times 14.01) \text{ g mol}^{-1}} \text{ of } \text{N}_2 \text{ gas} = 1.57 \text{ mol}$$

1 mole of gas has a volume of 22.7 dm^3 at STP.

Therefore 1.57 mol has a volume of 35.6 dm^3

- (b) 1 mole of gas has a volume of 22.7 dm^3 at STP.

Therefore 0.25 mol of ammonia has a volume of 5.7 dm^3

47 moles $\text{HgO} = \frac{12.45 \text{ g}}{(200.59 + 16.00) \text{ g mol}^{-1}} = 0.0575 \text{ mol}$

On decomposition this would produce 0.0287 mol of oxygen (since 2 mol of HgO produces 1 mol of O_2).

1 mole of gas has a volume of 22.7 dm^3 at STP.

Therefore 0.0287 mol have a volume of 0.652 dm^3

- 48 Assume all measurements are made at STP.

3.14 dm^3 of bromine is equivalent to

$$\frac{3.14 \text{ dm}^3}{22.7 \text{ mol dm}^3} = 0.138 \text{ mol Br}_2$$

11.07 g of chlorine is equivalent to

$$\frac{11.07 \text{ g}}{(2 \times 35.45) \text{ g mol}^{-1}} = 0.1561 \text{ mol Cl}_2$$

Therefore the sample of chlorine contains more molecules.

49 0.200 g calcium is $\frac{0.200 \text{ g}}{40.08 \text{ g mol}^{-1}} =$

$$4.99 \times 10^{-3} \text{ mol}$$

$4.99 \times 10^{-3} \text{ mol}$ of Ca will make $4.99 \times 10^{-3} \text{ mol}$ of hydrogen

$4.99 \times 10^{-3} \text{ mol}$ of hydrogen will occupy

$$4.99 \times 10^{-3} \text{ mol} \times 22.7 \text{ dm}^3 \text{ mol}^{-1} = 0.113 \text{ dm}^3 \text{ (or } 113 \text{ cm}^3) \text{ at STP.}$$

- 50 The first step, as usual, is to calculate how many moles of reactant we have.

$$1.0 \text{ g of ammonium nitrate is } \frac{1.0 \text{ g}}{80.06 \text{ g mol}^{-1}} = 0.012 \text{ mol}$$

According to the balanced chemical equation, 0.012 mol of ammonium nitrate will produce 0.012 mol of dinitrogen oxide.

At STP, 1 mole of gas occupies 22.7 dm^3 , so 0.012 mol will occupy $22.7 \text{ dm}^3 \text{ mol}^{-1} \times 0.012 \text{ mol} = 0.28 \text{ dm}^3$

51 Using $\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$

List all the data that is given in the question:

- $P_1 = 85 \text{ kPa}$

- $P_2 = ???$
- $V_1 = 2.50 \text{ dm}^3$
- $V_2 = 2.75 \text{ dm}^3$
- $T_1 = 25 \text{ }^\circ\text{C} (= 298 \text{ K})$
- $T_2 = 75 \text{ }^\circ\text{C} (= 348 \text{ K})$

$$\frac{85 \text{ kPa} \times 2.50 \text{ dm}^3}{298 \text{ K}} = \frac{P_2 \times 2.75 \text{ dm}^3}{348 \text{ K}}$$

Rearranging and solving for P_2 gives the final pressure = 90 kPa

52 Using $\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$

List all the data that is given in the question:

- $P_1 = 1.00 \times 10^5 \text{ Pa}$
- $V_1 = 675 \text{ cm}^3$
- $T_1 = ???$
- $P_2 = 2.00 \times 10^5 \text{ Pa}$
- $V_2 = 350 \text{ cm}^3$
- $T_2 = 27.0 \text{ }^\circ\text{C} (= 300 \text{ K})$

$$\frac{(1.00 \times 10^5) \text{ Pa} \times 675 \text{ cm}^3}{T_1} = \frac{(2.00 \times 10^5) \text{ Pa} \times 350 \text{ cm}^3}{300 \text{ K}}$$

Rearranging and solving for T_1 gives initial temperature = 289 K = 16 °C

53 Using $\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$

$$\frac{P_1 \times 4.0 \text{ dm}^3}{T_1} = \frac{4P_1 \times V_2}{3T_1}$$

Rearranging and solving for V_2 gives

$$V_2 = \frac{3 \times T_1 \times P_1 \times 4.0 \text{ dm}^3}{4 \times P_1 \times T_1} = 3.0 \text{ dm}^3$$

54 Using $M = \frac{mRT}{PV}$

List all the data that is given in the question:

- $m = \text{mass} = 4.40 \text{ g}$
- $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$
- $T = \text{temperature in K} = (273 + 27) = 300 \text{ K}$
- $P = \text{pressure} = 90 \text{ kPa} = 90 \times 10^3 \text{ Pa}$
- $M = \text{molar mass} = (12.01 + (2 \times 16.00)) \text{ g mol}^{-1} = 44.01 \text{ g mol}^{-1}$

Rearranging for V gives

$$V = \frac{mRT}{PM}$$

$$V = \frac{4.40 \text{ g} \times 8.31 \text{ J K}^{-1} \text{ mol}^{-1} \times 300 \text{ K}}{90 \times 10^3 \text{ Pa} \times 44.01 \text{ g mol}^{-1}}$$

$$V = 2.8 \times 10^{-3} \text{ m}^3 (= 2.8 \text{ dm}^3)$$

55 At STP, 1 mole of the gas would occupy 22.7 dm³

1 mole would have a molar mass of

$$5.84 \text{ g dm}^{-3} \times 22.7 \text{ dm}^3 \text{ mol}^{-1} = 133 \text{ g mol}^{-1}$$

From section 6 of the IB data booklet, xenon is the noble gas with the closest molar mass, 131.29 g mol⁻¹

56 Using $M = \frac{mRT}{PV}$

List all the data that is given in the question:

- $m = 12.1 \text{ mg} = 0.0121 \text{ g}$
- $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$
- $T = 25 \text{ }^\circ\text{C} = 298 \text{ K}$
- $P = 1300 \text{ Pa}$
- $V = 255 \text{ cm}^3 = 255 \times 10^{-6} \text{ m}^3$

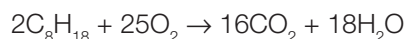
$$M = \frac{0.0121 \text{ g} \times 8.31 \text{ J K}^{-1} \text{ mol}^{-1} \times 298 \text{ K}}{1300 \text{ Pa} \times 255 \times 10^{-6} \text{ m}^3}$$

$$= 90.4 \text{ g mol}^{-1}$$

57 As density = $\frac{\text{mass}}{\text{volume}}$ for a fixed volume of gas,

the density will depend on the formula mass of the element. Hydrogen has a formula mass of 2.02 g mol⁻¹ and helium of 4.00 g mol⁻¹. Hence helium has the greater density.

58 The equation for the complete combustion of octane is:



1 mole of octane reacts with 12.5 moles of oxygen.

$$M(\text{C}_8\text{H}_{18}) = ((8 \times 12.01) + (18 \times 1.01)) \text{ g mol}^{-1} = 114.26 \text{ g mol}^{-1}$$

$$\text{moles of octane} = \frac{n}{M} = \frac{125 \text{ g}}{114.26 \text{ g mol}^{-1}}$$

$$= 1.09 \text{ mol}$$

Therefore 1.09 mol of octane react with $1.09 \times \frac{25}{2} = 13.7 \text{ mol}$ of oxygen

1 mol of gas occupies 22.7 dm³, hence 13.7 mol occupy 13.7 mol \times 22.7 dm³ mol⁻¹ = 311 dm³

59 Using $M = \frac{mRT}{PV}$

List all the data that is given in the question:

- $m = 3.620 \text{ g}$
- $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$
- $T = 25 \text{ }^\circ\text{C} = 298 \text{ K}$
- $P = 99 \text{ kPa} = 99 \times 10^3 \text{ Pa}$
- $V = 1120 \text{ cm}^3 = 1120 \times 10^{-6} \text{ m}^3$

$$M = \frac{3.620 \text{ g} \times 8.31 \text{ J K}^{-1} \text{ mol}^{-1} \times 298 \text{ K}}{99 \times 10^3 \text{ Pa} \times 1120 \times 10^{-6} \text{ m}^3}$$

$$= 80.8 \text{ g mol}^{-1}$$

	Oxygen	Sulfur
mass / g	2.172	1.448
moles	$\frac{2.172}{16.00} = 0.1357$	$\frac{1.448}{32.07} = 0.04515$
divide by smallest	3.01	1.00
nearest whole number ratio	3	1

Empirical formula = SO_3

Since this has an M_r of 80.07, the empirical and molecular formulae must be the same.

- 60 At higher altitude the external pressure is less. As the air in the tyre expands on heating, due to friction with the road surface, the internal pressure increases. (This can be a much greater problem during a descent when friction from the brakes on the wheel rim causes the tyre to heat up further.)
- 61 (a) • Particles are in constant random motion and collide with each other and with the walls of the container in perfectly elastic collisions.
- The kinetic energy of the particles increases with temperature.
 - There are no inter-particle forces.
 - The volume of the particles is negligible relative to the volume of the gas.
- (b) At low temperature, the particles have lower kinetic energy, which favours the formation of inter-particle forces and reduces gas pressure. $\frac{PV}{nRT} < 1$

62 NH_3 shows greater deviation than CH_4 due to stronger intermolecular attractions, especially at low temperature.

63 B Ideal gases are assumed to have no attractive forces between the particles; however, gases are not ideal.

64 number of moles required = concentration \times volume
 $= 0.200 \text{ mol dm}^{-3} \times 0.250 \text{ dm}^3$ ($1 \text{ dm}^3 = 1000 \text{ cm}^3$)
 $= 0.0500 \text{ mol}$

Formula of potassium hydroxide is KOH. (You should know that the ions are K^+ and OH^- .)

Molar mass of KOH = $(39.10 + 16.00 + 1.01) \text{ g mol}^{-1} = 56.11 \text{ g mol}^{-1}$

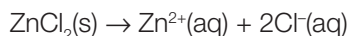
Mass of 0.0500 mol of KOH = $0.0500 \text{ mol} \times 56.11 \text{ g mol}^{-1} = 2.81 \text{ g}$

65 $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ has a molar mass of $(24.31 + 32.07 + 11 \times 16.00 + 14 \times 1.01) \text{ g mol}^{-1} = 246.52 \text{ g mol}^{-1}$

0.100 dm^3 of a $0.200 \text{ mol dm}^{-3}$ solution contains $0.100 \times 0.200 \text{ mol} = 0.0200 \text{ mol}$ of solute
 0.0200 mol of $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ has a mass of $0.0200 \text{ mol} \times 246.52 \text{ g mol}^{-1} = 4.93 \text{ g}$

66 In 0.250 dm^3 of $0.0200 \text{ mol dm}^{-3}$ of solution there are $0.250 \times 0.0200 = 0.00500 \text{ mol}$ of solute

For every mole of ZnCl_2 , two moles of chloride ions are released in solutions:



so, 0.00500 mol of ZnCl_2 will give 0.0100 mol of chloride ions in solution

67 250 cm^3 of solution contain 5.85 g of sodium chloride
 1 dm^3 of solution contains $5.85 \text{ g} = \frac{1000 \text{ cm}^3}{250 \text{ cm}^3} = 23.40 \text{ g}$ of sodium chloride
 23.40 g of NaCl is equivalent to $\frac{23.40 \text{ g}}{(22.99 + 35.45) \text{ g mol}^{-1}} = 0.400 \text{ mol}$
Hence concentration is $0.400 \text{ mol dm}^{-3}$

68 100 cm^3 of 0.50 mol dm^{-3} nitric acid contains 0.050 moles of acid

volume of 16.0 mol dm^{-3} acid to contain this number of moles = $\frac{n}{c} = \frac{0.050 \text{ mol}}{16.0 \text{ mol dm}^{-3}} = 3.1 \times 10^{-3} \text{ dm}^3 = 3.1 \text{ cm}^3$

- 69** $\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$
 moles of NaOH = $cV = 0.147 \text{ mol dm}^{-3} \times \frac{36.42}{1000} \text{ dm}^3 = 5.35 \times 10^{-3} \text{ mol}$
 As this reacts with the sulfuric acid in a 2 : 1 ratio, the number of moles of sulfuric acid present = $2.68 \times 10^{-3} \text{ mol}$

Hence concentration of sulfuric acid = $\frac{n}{V} = \frac{2.68 \times 10^{-3} \text{ mol}}{\left(\frac{15.00}{1000}\right) \text{ dm}^3} = 0.178 \text{ mol dm}^{-3}$

- 70** moles of KOH used in titration = $cV = 0.0100 \text{ mol dm}^{-3} \times \frac{11.00}{1000} \text{ dm}^3 = 1.10 \times 10^{-4} \text{ mol}$

Since KOH reacts with HCl in a 1:1 ratio, moles of HCl = $1.10 \times 10^{-4} \text{ mol}$

Concentration of HCl = $\frac{n}{V} = \frac{1.10 \times 10^{-4} \text{ mol}}{\left(\frac{5.00}{1000}\right) \text{ dm}^3} = 0.0220 \text{ mol dm}^{-3}$

Molar mass of HCl = $(1.01 + 35.45) \text{ g mol}^{-1} = 36.46 \text{ g mol}^{-1}$

Quantity of HCl is equivalent to $36.46 \times 0.0220 = 0.802 \text{ g}$ dissolved in 1 dm^3

Concentration of HCl in $\text{g dm}^{-3} = 0.0220 \text{ mol dm}^{-3} \times 36.46 \text{ g mol}^{-1} = 0.802 \text{ g dm}^{-3}$

Therefore in 1.00 cm^3 there would be 0.000802 g of HCl

Assuming a density of 1.00 g cm^{-3} a 1.00 cm^3 sample of solution has a mass of 1.00 g

$\% \text{HCl} = \frac{\text{mass HCl}}{\text{mass solution}} \times 100 = \frac{0.000802 \text{ g}}{1.00 \text{ g}} \times 100 = 0.0802\%$

- 71** $\text{Na}_2\text{SO}_4(\text{aq}) + \text{Pb}(\text{NO}_3)_2(\text{aq}) \rightarrow \text{PbSO}_4(\text{s}) + 2\text{NaNO}_3(\text{aq})$

First determine the moles of PbSO_4 formed in the reaction:

$$n(\text{PbSO}_4) = m/M(\text{PbSO}_4) = \frac{1.13 \text{ g}}{303.25 \text{ g mol}^{-1}} = 3.73 \times 10^{-3} \text{ mol}$$

From the balanced equation:

$n(\text{PbSO}_4)$ formed = $n(\text{Pb}(\text{NO}_3)_2)$ reacted = $n(\text{Na}_2\text{SO}_4)$ reacted = $3.73 \times 10^{-3} \text{ mol}$

$$[\text{Pb}(\text{NO}_3)_2] = \frac{n}{V} = \frac{3.73 \times 10^{-3} \text{ mol}}{(32.50/1000) \text{ dm}^3} = 0.115 \text{ mol dm}^{-3}$$

$$[\text{Na}_2\text{SO}_4] = \frac{n}{V} = \frac{3.73 \times 10^{-3} \text{ mol}}{(35.30/1000) \text{ dm}^3} = 0.106 \text{ mol dm}^{-3}$$

Two assumptions are:

- No side reactions occur that generate other products.
- All of the PbSO_4 formed precipitates out as a solid and can be weighed.

Practice questions

Questions 1–17 are multiple-choice questions similar to those in Paper 1 of the IB examinations. As calculators are not allowed in Paper 1 it is appropriate for whole-number values to be used for molar masses. Any relevant constants can also be rounded where appropriate to make the calculations simpler.

- One mole of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ will contain 9 moles of oxygen atoms.
 0.100 moles of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ will contain 0.900 moles of oxygen atoms.
 0.900 moles = $0.900 \times N_A$ atoms = $0.900 \times 6.02 \times 10^{23}$ atoms = 5.42×10^{23} atoms.
 Correct answer is D.
- The balanced equation is $\text{Fe}_2\text{O}_3(\text{s}) + 3\text{CO}(\text{g}) \rightarrow 2\text{Fe}(\text{s}) + 3\text{CO}_2(\text{g})$.
 Sum of coefficients = $1 + 3 + 2 + 3 = 9$.
 Correct answer is D.
- We can assume that the gases are all ideal gases. This means that under the same conditions of pressure, volume and temperature, they all contain the same amount of gas molecules.

The heaviest container will therefore contain the gas with the largest molar mass.

Nitrogen: $M(\text{N}_2) = 2 \times 14 \text{ g mol}^{-1} = 28 \text{ g mol}^{-1}$

Oxygen: $M(\text{O}_2) = 2 \times 16 \text{ g mol}^{-1} = 32 \text{ g mol}^{-1}$

Ethane: $M(\text{C}_2\text{H}_6) = (2 \times 12) + (6 \times 1) \text{ g mol}^{-1} = 30 \text{ g mol}^{-1}$

Neon: $M(\text{Ne}) = 20 \text{ g mol}^{-1}$

Oxygen has the largest molar mass so the heaviest container will be B.

Correct answer is B.

- 4 This can be solved by converting temperatures from degrees Celsius to Kelvin and applying Charles' Law, $\frac{V}{T} = \text{constant}$:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$V_2 = \frac{V_1 T_2}{T_1} = \frac{1.0 \text{ dm}^3 \times 323 \text{ K}}{298 \text{ K}} = 1.1 \text{ dm}^3$$

Correct answer is C.

- 5 $n(\text{FeSO}_4) = cV = 0.020 \text{ mol dm}^{-3} \times \frac{100}{1000} \text{ dm}^3 = 0.0020 \text{ mol} = 2.0 \times 10^{-3} \text{ mol}$

$$n(\text{SO}_4^{2-}) = n(\text{FeSO}_4) = 2.0 \times 10^{-3} \text{ mol}$$

Correct answer is A.

- 6 To calculate a concentration we first need to convert the mass of NaNO_3 to moles using the M_r of 85 provided for NaNO_3 :

$$n(\text{NaNO}_3) = \frac{1.7 \text{ g}}{85 \text{ g mol}^{-1}} = 0.020 \text{ mol}$$

$$[\text{NaNO}_3] = \frac{n}{V} = \frac{0.020 \text{ mol}}{0.20 \text{ dm}^3} = 0.10 \text{ mol dm}^{-3}$$

Correct answer is B.

- 7 From the balanced equation provided in the question:

$$n(\text{H}_2) \text{ produced} = \frac{3}{2}n(\text{Al}) \text{ reacted}$$

If 3 mol of Al reacts, the amount of H_2 produced $= \frac{3}{2} \times 3 \text{ mol} = 4.5 \text{ mol}$.

$$m(\text{H}_2) = nM(\text{H}_2) = 4.5 \text{ mol} \times 2 \text{ g mol}^{-1} = 9.0 \text{ g}$$

Correct answer is D.

- 8 Based on the empirical formula of the gas, CH_2 , the relative formula mass can be calculated:

relative formula mass (CH_2) = $12 + (2 \times 1) = 14$

Dividing the relative molar mass by the relative formula mass gives us the multiplier, x , that must be applied to the empirical formula to give the molecular formula:

$$x = \frac{\text{relative molecular mass}}{\text{relative formula mass}} = \frac{56}{14} = 4$$

Molecular formula = ' $4 \times \text{CH}_2$ ' = C_4H_8

Correct answer is D.

- 9 The balanced equation is $2\text{C}_2\text{H}_2 + 5\text{O}_2(\text{g}) \rightarrow 4\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}$.

Sum of coefficients = $2 + 5 + 4 + 2 = 13$

Correct answer is D.

- 10 1 mole of benzamide, $\text{C}_6\text{H}_5\text{CONH}_2$, will contain 7 moles of hydrogen atoms.

1.0 moles of $\text{C}_6\text{H}_5\text{CONH}_2$ will contain 7.0 moles of hydrogen atoms.

$$7.0 \text{ moles} = 7.0 \times (6.02 \times 10^{23}) \text{ atoms} = 4.2 \times 10^{24} \text{ atoms}$$

Correct answer is D.

- 11 O_2 is the limiting reagent as 10.0 mol of $\text{C}_2\text{H}_3\text{Cl}$ would require 25.0 mol of O_2 .

From the balanced equation, $n(\text{H}_2\text{O}) = \frac{2}{5}n(\text{O}_2)$.

If 10.0 mol of O_2 reacts the amount of H_2O produced $= \frac{2}{5} \times 10.0 \text{ mol} = 4.00 \text{ mol}$.

Correct answer is A.

- 12 We first calculate the total number of moles of NaCl present in the two solutions:

$$n(\text{total}) = n(\text{solution 1}) + n(\text{solution 2})$$

$$= c_1V_1 + c_2V_2 = (0.200 \text{ mol dm}^{-3} \times \frac{10.0}{1000} \text{ dm}^3) + (0.600 \text{ mol dm}^{-3} \times \frac{30.0}{1000} \text{ dm}^3) = 0.00200 \text{ mol} + 0.0180 \text{ mol} = 0.0200 \text{ mol}$$

$$[\text{NaCl}] = \frac{n}{(V_1 + V_2)} = \frac{0.0200 \text{ mol}}{\frac{(10.0 + 30.0)}{1000} \text{ dm}^3}$$

$$= \frac{0.0200 \text{ mol}}{0.0400 \text{ dm}^3} = 0.500 \text{ mol dm}^{-3}$$

Correct answer is C.

- 13 Determine the empirical formula of the compound.

	C	H	O
Mass (g)	12	2	16
Moles	$\frac{12 \text{ g}}{12 \text{ g mol}^{-1}}$	$\frac{2 \text{ g}}{1 \text{ g mol}^{-1}}$	$\frac{16 \text{ g}}{16 \text{ g mol}^{-1}}$
	= 1.0	= 2.0	= 1.0

Empirical formula = CH₂O

$$\begin{aligned} \text{Formula mass (CH}_2\text{O)} &= (12 + (2 \times 1) + 16) \text{ g mol}^{-1} \\ &= 30 \text{ g mol}^{-1} \end{aligned}$$

Dividing the molar mass by the formula mass gives us the multiplier, *x*, that must be applied to the empirical formula to give the molecular formula:

$$x = \frac{\text{relative molecular mass}}{\text{relative formula mass}} = \frac{60 \text{ g mol}^{-1}}{30 \text{ g mol}^{-1}} = 2$$

Molecular formula = '2 × CH₂O' = C₂H₄O₂

Correct answer is D.

- 14 We can calculate the final concentration, *c*₂, using the dilution formula *c*₁*V*₁ = *c*₂*V*₂.

$$\begin{aligned} c_2 &= \frac{c_1 V_1}{V_2} = \frac{0.5 \text{ mol dm}^{-3} \times 200 \text{ cm}^3}{(200 + 300) \text{ cm}^3} \\ &= \frac{0.5 \text{ mol dm}^{-3} \times 200 \text{ cm}^3}{500 \text{ cm}^3} = 0.2 \text{ mol dm}^{-3} \end{aligned}$$

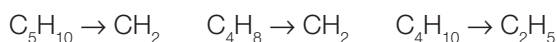
Correct answer is C.

- 15 The question asks for an approximate value so we can use whole numbers for the atomic masses of the constituent elements of MgSO₄·7H₂O:

$$\begin{aligned} M(\text{MgSO}_4 \cdot 7\text{H}_2\text{O}) &= (24 + 32 + (4 \times 16) + (14 \times 1) \\ &+ (7 \times 16)) \text{ g mol}^{-1} = 246 \text{ g mol}^{-1} \end{aligned}$$

Correct answer is D.

- 16 For a molecular formula to also be an empirical formula it cannot be converted to a simpler ratio. With the exception of C₅H₁₂, all of the formulas provided can be simplified.



Correct answer is A.

- 17 (a) Temperature is 25.00 °C. This has **four** significant figures.

Mass is 0.0650 kg. This has **three** significant figures.

Pressure is 1.08 atm. This has **three** significant figures.

$$\text{(b) } n(\text{NaN}_3) = \frac{m}{M(\text{NaN}_3)} = \frac{65.0 \text{ g}}{65.02 \text{ g mol}^{-1}} = 1.00 \text{ mol}$$

$$\text{(c) From the balanced equation, } n(\text{N}_2) = \frac{3}{2} n(\text{NaN}_3) = \frac{3}{2} \times 1.00 \text{ mol} = 1.50 \text{ mol}.$$

We can calculate the volume of N₂(g) produced using the ideal gas equation, *PV* = *nRT*, recognizing that temperature must be converted to K and pressure to Pa:

$$25.00 \text{ °C} = (25.00 + 273.15) \text{ K} = 298.15 \text{ K}$$

$$\begin{aligned} 1.08 \text{ atm} &= 1.08 \times (1.01 \times 10^5) \text{ Pa} \\ &= 1.09 \times 10^5 \text{ Pa} \end{aligned}$$

$$V = \frac{nRT}{P} =$$

$$\frac{1.50 \text{ mol} \times 8.314 \text{ J K}^{-1} \text{ mol}^{-1} \times 298.15 \text{ K}}{1.09 \times 10^5 \text{ Pa}} =$$

$$0.0341 \text{ m}^3 = (0.0341 \times 1000) \text{ dm}^3 = 34.1 \text{ dm}^3$$

- 18 (a) The amount of C present in **J** can be found from the mass of CO₂ formed in the combustion reaction:

$$\begin{aligned} n(\text{CO}_2) &= \frac{m}{M(\text{CO}_2)} = \frac{0.872 \text{ g}}{44.01 \text{ g mol}^{-1}} \\ &= 0.0198 \text{ mol} \end{aligned}$$

$$n(\text{C}) = n(\text{CO}_2) = 0.0198 \text{ mol}$$

$$\begin{aligned} \text{mass of C} &= nM = 0.0198 \text{ mol} \times 12.01 \text{ g mol}^{-1} \\ &= 0.238 \text{ g} \end{aligned}$$

$$\begin{aligned} \% \text{ mass of C in J} &= \frac{0.238 \text{ g}}{1.30 \text{ g}} \times 100\% \\ &= 18.3\% \end{aligned}$$

The amount of H present in **J** can be found from the mass of H₂O formed in the combustion reaction:

$$\begin{aligned} n(\text{H}_2\text{O}) &= \frac{m}{M(\text{H}_2\text{O})} = \frac{0.089 \text{ g}}{18.02 \text{ g mol}^{-1}} \\ &= 0.0049 \text{ mol} \end{aligned}$$

$$\begin{aligned} n(\text{H}) &= 2 \times n(\text{H}_2\text{O}) = 2 \times 0.0049 \text{ mol} \\ &= 0.0098 \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{mass of H} &= nM = 0.0098 \text{ mol} \times 1.01 \text{ g mol}^{-1} \\ &= 0.0099 \text{ g} \end{aligned}$$

$$\begin{aligned}\% \text{ mass of H in J} &= \frac{0.0099 \text{ g}}{1.30 \text{ g}} \times 100\% \\ &= 0.76\%\end{aligned}$$

(An answer of 0.77% is also acceptable. This value is obtained if the calculation is performed in a single step and no rounding error is introduced.)

- (b) The amount of Cl present in **J** can be found from the mass of AgCl formed in the precipitation reaction:

$$\begin{aligned}n(\text{AgCl}) &= \frac{m}{M(\text{AgCl})} = \frac{1.75 \text{ g}}{143.32 \text{ g mol}^{-1}} \\ &= 0.0122 \text{ mol}\end{aligned}$$

$$n(\text{Cl}) = n(\text{AgCl}) = 0.0122 \text{ mol}$$

$$\begin{aligned}\text{mass of Cl} &= nM = 0.0122 \text{ mol} \times 35.45 \text{ g mol}^{-1} \\ &= 0.432 \text{ g}\end{aligned}$$

$$\begin{aligned}\% \text{ mass of Cl in J} &= \frac{0.432 \text{ g}}{0.535 \text{ g}} \times 100\% \\ &= 80.7\%\end{aligned}$$

(An answer of 80.9% is also acceptable. This value is obtained if the calculation is performed in a single step and no rounding error is introduced.)

- (c) We can determine the empirical formula using the % compositions obtained in part (b) and the mass that would be in a 100 g sample of **J**.

	C	H	O
Mass/g	18.3	0.76	80.7
M/g mol ⁻¹	12.01	1.01	16.00
Number of moles/mol	1.52	0.75	2.28
Divide by smallest	2.03	1.0	3.04
Nearest whole number ratio	2	1	3

Empirical formula of **J** = C₂HCl₃

The formula mass of C₂HCl₃ = (2 × 12.01) + 1.01 + (3 × 35.45) g mol⁻¹ = 131.38 g mol⁻¹.

As the formula mass calculated for the empirical formula is the same as the molar mass given in the question the molecular formula of **J** is confirmed as C₂HCl₃.

- 19 We can determine which gas is in excess based on the assumption that they are ideal gases and therefore the volumetric ratios are the same as the molar ratios.

From the balanced equation, $n(\text{NO}) = 1.5n(\text{NH}_3)$

Therefore assuming ideal gases, $V(\text{NO}) = 1.5V(\text{NH}_3)$

30.0 dm³ of NH₃ would require 1.5 × 30 cm³ of NO = 45.0 dm³ of NO

As there are only 30.0 dm³ of NO it is the limiting reactant and NH₃ is in excess.

We can determine how much NH₃ reacts with 30.0 dm³ of NO:

$$V(\text{NH}_3) = \frac{V(\text{NO})}{1.5} = \frac{30.0 \text{ dm}^3}{1.5} = 20.0 \text{ dm}^3$$

The volume of excess NH₃ = 30.0 dm³ – 20.0 dm³ = 10.0 dm³

Using the balanced equation, the volume of N₂ produced can be determined from the volume of NO reacted:

$$V(\text{N}_2) = \frac{5}{6} V(\text{NO}) = \frac{5}{6} \times 30.0 \text{ dm}^3 = 25.0 \text{ dm}^3$$

- 20 (a) $n(\text{HCl}) = c \times V = 0.200 \text{ mol dm}^{-3} \times \frac{27.20}{1000} \text{ dm}^3 = 0.00544 \text{ mol}$

- (b) We need to write the balanced equation for the neutralization of HCl(aq) with NaOH(aq):
HCl(aq) + NaOH(aq) → NaCl(aq) + H₂O(l)

From the balanced equation, $n(\text{HCl}) = n(\text{NaOH})$:

$$\begin{aligned}n(\text{NaOH}) &= cV = 0.100 \text{ mol dm}^{-3} \times \frac{23.80}{1000} \text{ dm}^3 \\ &= 0.00238 \text{ mol}\end{aligned}$$

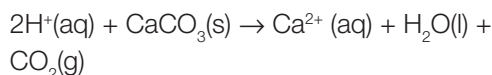
$n(\text{HCl})$ in excess = 0.00238 mol

- (c) The amount of HCl that reacted with the calcium carbonate in the eggshell is found from the difference between the original HCl added and the excess HCl that then reacted with the NaOH:

$$\begin{aligned}n(\text{HCl reacted}) &= 0.00544 \text{ mol} - 0.00238 \text{ mol} \\ &= 0.00306 \text{ mol}\end{aligned}$$

- (d) $2\text{HCl}(\text{aq}) + \text{CaCO}_3(\text{s}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$

An acceptable alternative that omits spectator ions is:



- (e) From the balanced equation in (d) the amount of CaCO_3 can be determined based on the amount of HCl that reacts:

$$\begin{aligned} n(\text{CaCO}_3) &= \frac{1}{2} n(\text{HCl reacted}) \\ &= \frac{1}{2} \times 0.00306 \text{ mol} = 0.00153 \text{ mol} \end{aligned}$$

- (f) $m(\text{CaCO}_3) = nM(\text{CaCO}_3) = 0.00153 \text{ mol} \times 100.09 \text{ g mol}^{-1} = 0.153 \text{ g}$

$$\begin{aligned} \% \text{CaCO}_3 \text{ in eggshell} &= \frac{\text{mass of CaCO}_3}{\text{mass of eggshell}} \times 100\% \\ &= \frac{0.153 \text{ g}}{0.188 \text{ g}} \times 100\% = 81.4\% \end{aligned}$$

- (g) The main assumption is that CaCO_3 is the only component of the eggshell that reacts with HCl , i.e. there are no basic impurities present in the eggshell that would also react with HCl .

- 21 We first need to determine the number of moles of AgCl solid that were precipitated:

$$\begin{aligned} n(\text{AgCl}) &= \frac{m}{M(\text{AgCl})} = \frac{6.127 \text{ g}}{143.32 \text{ g mol}^{-1}} \\ &= 0.04275 \text{ mol} \end{aligned}$$

$$n(\text{Cl}^-) = n(\text{AgCl}) = 0.04275 \text{ mol}$$

We can now solve this problem using simultaneous equations:

Let x = mass of NaCl in g

Let y = mass of CaCl_2 in g

$$x + y = 2.450 \quad (1)$$

We can determine a second equation based on the number of moles of Cl^- in the two solids along with the total number of moles of Cl^- in the precipitate:

$$n(\text{Cl}^- \text{ from NaCl}) + n(\text{Cl}^- \text{ from CaCl}_2) = 0.04275$$

$$\frac{x}{M(\text{NaCl})} + 2\left(\frac{y}{M(\text{CaCl}_2)}\right) = 0.04275$$

$$\frac{x}{58.44} + 2\left(\frac{y}{110.98}\right) = 0.04275 \quad (2)$$

Applying simultaneous equations:

$$x + y = 2.450 \quad (1)$$

$$\frac{x}{58.44} + 2\left(\frac{y}{110.98}\right) = 0.04275 \quad (2)$$

$$(2) \times 58.44 \quad x + 1.053y = 2.498 \quad (2')$$

$$(2') - (1) \quad 0.053y = 0.048$$

$$y = \frac{0.048}{0.053}$$

$$= 0.91$$

$$\text{Subbing } y \text{ into (1)} \quad x + 0.91 = 2.450$$

$$x = 1.54$$

The sample contains 1.54 g of NaCl and 0.91 g of CaCl_2 .

$$\begin{aligned} \% \text{NaCl} &= \frac{\text{mass NaCl}}{\text{total mass}} \times 100\% = \frac{1.54 \text{ g}}{2.450 \text{ g}} \times 100\% \\ &= 62.9\% \end{aligned}$$

$$\begin{aligned} \% \text{CaCl}_2 &= \frac{\text{mass CaCl}_2}{\text{total mass}} \times 100\% = \frac{0.91 \text{ g}}{2.450 \text{ g}} \times 100\% \\ &= 37.1\% \end{aligned}$$

- 22 (a) We first find the mass of water that was removed in drying the potassium carbonate:

$$m(\text{H}_2\text{O}) = m(\text{hydrated K}_2\text{CO}_3) - m(\text{dry K}_2\text{CO}_3)$$

$$= 10.00 \text{ g} - 7.93 \text{ g}$$

$$= 2.07 \text{ g}$$

$$\begin{aligned} n(\text{H}_2\text{O}) &= \frac{m}{M(\text{H}_2\text{O})} = \frac{2.07 \text{ g}}{18.02 \text{ g mol}^{-1}} \\ &= 0.115 \text{ mol} \end{aligned}$$

- (b) $n(\text{K}_2\text{CO}_3) = \frac{m}{M(\text{K}_2\text{CO}_3)} = \frac{7.93 \text{ g}}{138.21 \text{ g mol}^{-1}} = 0.0574 \text{ mol}$

- (c) We can determine the value of x in the hydrated potassium carbonate:

$$x = \frac{n(\text{H}_2\text{O removed})}{n(\text{dry K}_2\text{CO}_3)} = \frac{0.115 \text{ mol}}{0.0574 \text{ mol}} = 2.00$$

Therefore the formula of the hydrate is $\text{K}_2\text{CO}_3 \cdot 2\text{H}_2\text{O}$

- (d) By repeating the process of heating and weighing until a constant mass is obtained. (To ensure accurate results it will be necessary to cool the sample to room temperature before each weighing.)

- 23 (a) We can determine the moles of each reactant using the ideal gas law, $PV = nRT$, recognizing that temperature must be converted to K, pressure to Pa and volume to m^3 .

For ammonia gas:

$$T = (42 + 273) \text{ K} = 315 \text{ K}, P = 160 \times 10^3 \text{ Pa},$$

$$V = \frac{625}{1 \times 10^6} \text{ m}^3 = 6.25 \times 10^{-4} \text{ m}^3$$

$$n(\text{NH}_3) = \frac{PV}{RT} = \frac{160\,000 \text{ Pa} \times 6.25 \times 10^{-4} \text{ m}^3}{8.31 \text{ J K}^{-1} \text{ mol}^{-1} \times 315 \text{ K}} = 0.0382 \text{ mol}$$

For hydrogen chloride gas:

$$T = (57 + 273) \text{ K} = 330 \text{ K}, P = 113.3 \times 10^3 \text{ Pa}, V = \frac{740}{1 \times 10^6} \text{ m}^3 = 7.40 \times 10^{-4} \text{ m}^3$$

$$n(\text{HCl}) = \frac{PV}{RT} = \frac{113\,300 \text{ Pa} \times 7.40 \times 10^{-4} \text{ m}^3}{8.31 \text{ J K}^{-1} \text{ mol}^{-1} \times 330 \text{ K}} = 0.0306 \text{ mol}$$

Therefore ammonia gas (NH_3) is in excess.

- (b) If ammonia gas is in excess then the hydrogen chloride gas (HCl) is limiting.
- (c) The balanced equation for the reaction is $\text{NH}_3(\text{g}) + \text{HCl}(\text{g}) \rightarrow \text{NH}_4\text{Cl}(\text{s})$.

From this equation, $n(\text{NH}_4\text{Cl}) = n(\text{HCl}) = 0.0306 \text{ mol}$.

$$m(\text{NH}_4\text{Cl}) = nM(\text{NH}_4\text{Cl}) = 0.0306 \text{ mol} \times 53.50 \text{ g mol}^{-1} = 1.64 \text{ g}$$

- 24 (a) $2\text{PbS}(\text{s}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{PbO}(\text{s}) + 2\text{SO}_2(\text{g})$
 (b) $2\text{PbS}(\text{s}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{PbO}(\text{s}) + 2\text{SO}_2(\text{g})$

The terms in red indicate those of interest to the question.

Mole ratio $2\text{PbS}:2\text{SO}_2 = 1:1$

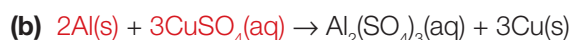
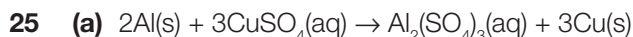
$$M(\text{PbS}) = 207.19 + 32.06 = 239.25 \text{ g mol}^{-1}$$

$$M(\text{SO}_2) = 32.06 + (16.00 \times 2) = 64.06 \text{ g mol}^{-1}$$

Therefore $239.25 \text{ g PbS} \rightarrow 64.06 \text{ g SO}_2$

$$\text{So } 1 \text{ tonne PbS} \rightarrow \frac{1 \text{ tonne} \times 64.06 \text{ g}}{239.25 \text{ g}} = 0.268 \text{ tonne} = 0.268 \times 10^3 \text{ kg} = 268 \text{ kg}$$

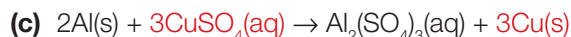
- (c) The calculation assumes that PbS is the limiting reactant and so is fully reacted in the presence of excess oxygen. It also assumes that the only reaction occurring is the production of PbO and SO_2 .



The question is about the reacting ratio of the terms in red.

$$\text{Mole ratio } 2\text{Al}:3\text{CuSO}_4 \quad \text{so} \quad \frac{x}{2} = \frac{10.38}{3}$$

$$x = 6.920 \text{ mol Al}$$



Mole ratio $3\text{CuSO}_4:3\text{Cu} = 1:1$

Therefore 3.95 mol Cu is produced from 3.95 mol CuSO_4 .

Note this question only asks for amounts in moles so it is not necessary to calculate the molar masses of the reactants and products.

- (d) The colour changes that occur when metals react with solutions of other metal ions will become familiar after you have studied redox reactions in Chapter 9.

Challenge yourself

- In cold climates, the temperature may approach or go below the boiling point of butane so the butane stays liquid even when it is released from the pressure it is under when stored in its canister. This makes it ineffective as a fuel in these heating devices in cold climates as they require gaseous fuels.
- We can determine which compounds are hydrated by comparing the molar masses provided in the photograph with the calculated molar masses for the unhydrated compounds.

Compound	Mass shown in photograph	Calculated molar mass	Hydrated or unhydrated
potassium iodide (KI)	166.0 g mol ⁻¹	166.0 g mol ⁻¹	unhydrated
sodium chloride (NaCl)	58.5 g mol ⁻¹	58.5 g mol ⁻¹	unhydrated

Compound	Mass shown in photograph	Calculated molar mass	Hydrated or unhydrated
potassium manganate(VII) (KMnO ₄)	158.0 g mol ⁻¹	158.0 g mol ⁻¹	unhydrated
iron(III) chloride (FeCl ₃)	270.3 g mol ⁻¹	162.2 g mol ⁻¹	hydrated
copper(II) sulfate (CuSO ₄)	249.7 g mol ⁻¹	159.6 g mol ⁻¹	hydrated
cobalt(II) nitrate (Co(NO ₃) ₂)	291.0 g mol ⁻¹	183.0 g mol ⁻¹	hydrated

Iron(III) chloride, copper(II) sulfate and cobalt(II) nitrate are hydrated.

Formulas of hydrated compounds

Mass of H₂O in one mole of hydrated FeCl₃ = 270.3 g – 162.2 g = 108.1 g

$$n(\text{H}_2\text{O}) = \frac{m}{M(\text{H}_2\text{O})} = \frac{108.1 \text{ g}}{18.02 \text{ g mol}^{-1}} \approx 6$$

In 1 mole of hydrated iron(III) chloride there are 6 moles of water, therefore the formula of hydrated iron(III) chloride is FeCl₃·6H₂O.

Applying the same working to the hydrated copper(II) sulfate and cobalt(II) nitrate gives the formulas CuSO₄·5H₂O and Co(NO₃)₂·6H₂O.

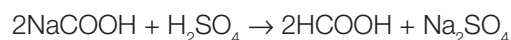
- 3 From the information box on page 24 we can see that a N-P-K rating of 18-51-20 means the fertilizer is 18% N, 51% P₂O₅ and 20% K₂O.

$$\begin{aligned} \% \text{ P} &= \frac{2 \times M_r(\text{P})}{M_r(\text{P}_2\text{O}_5)} \times \% \text{ P}_2\text{O}_5 \\ &= \frac{2 \times 30.07}{(2 \times 30.07 + 5 \times 16.00)} \times 51 = 0.436 \times 51 \\ &= 22\% \end{aligned}$$

$$\begin{aligned} \% \text{ K} &= \frac{2 \times M_r(\text{K})}{M_r(\text{K}_2\text{O})} \times \% \text{ K}_2\text{O} \\ &= \frac{2 \times 39.10}{(2 \times 39.10 \text{ g} + 16.00)} \times 20 = 0.830 \times 20 \\ &= 17\% \end{aligned}$$

% N = 18%, % P = 22%, % K = 17%

- 4 Many reactions with 'useless' by-products could have a high stoichiometric yield under optimum conditions, but low atom economy, for example the production of methanoic acid from the reaction of sodium methanoate and sulfuric acid:



For 100% conversion with stoichiometric reactants, the yield = 100%.

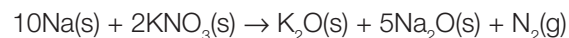
$$\% \text{ atom economy} = \frac{\text{mass of desired product}}{\text{total mass of reactants}} \times 100\%$$

$$\begin{aligned} &= \frac{2 \times M(\text{HCOOH})}{2 \times M(\text{NaHCOO}) + M(\text{H}_2\text{SO}_4)} \times 100\% \\ &= \frac{2 \times 46.03}{2 \times 68.01 + 98.08} \times 100\% \\ &= 39.33\% \end{aligned}$$

- 5 From the information on page 33, the main reaction in an airbag is the conversion of sodium azide, NaN₃, to nitrogen gas, N₂. Recognizing that solid sodium must be the other product gives the balanced equation:



Sodium metal is hazardous so it is converted to sodium oxide, Na₂O, through reaction with potassium nitrate, KNO₃:



The oxides formed in the above reaction are then reacted with silicon dioxide to form the harmless silicate Na₂K₂SiO₄ (alkaline silicate glass):



- 6 As NaOH dissolves, the separated Na⁺ and OH⁻ ions become hydrated due to the polar H₂O molecules being attracted to the charge on the ions and surrounding them. This attraction to the ions disrupts the hydrogen bonding between the H₂O molecules and allows for closer packing of the H₂O molecules in the NaOH solution, which reduces the volume of the solution.