

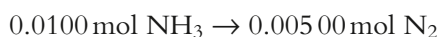
Answers to exam-style questions

Topic 1

- 1 C
2 B
3 D
4 A
5 B
6 B
7 C
8 A
9 B
10 D
- 11 volume of ammonia in $\text{dm}^3 = \frac{227}{1000} = 0.227 \text{ dm}^3$

$$\text{no. moles of ammonia} = \frac{\text{volume}}{\text{molar volume}} = \frac{0.227}{22.7} = 0.0100 \text{ mol}$$

From the equation, two moles of NH_3 produce one mole of N_2 .



0.00500 mol of N_2 has a volume of 0.00500×22.7 , i.e. 0.1135 dm^3 . This is $0.1135 \times 1000 = 113.5 \text{ cm}^3$, which is the theoretical yield of N_2 .

$$\text{percentage yield} = \frac{85}{113.5} \times 100 = 75\%$$

Alternative method: NH_3 and N_2 are both gases and so we do not have to convert to moles. From the equation, two moles of NH_3 react to give one mole of N_2 . Therefore two volumes of NH_3 react to give one volume of N_2 , so 227 cm^3 of NH_3 react to give $\frac{227}{2}$, i.e. 113.5 cm^3 of N_2 . This is the theoretical yield of N_2 .

The rest of the method is the same as above. [3]

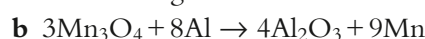
- 12 a Because the masses of the two substances are given, we must check to see if one of the substances is limiting.
- molar mass of $\text{Mn}_3\text{O}_4 = 228.82 \text{ g mol}^{-1}$
- no. moles of $\text{Mn}_3\text{O}_4 = \frac{100000}{228.82} = 437.0 \text{ mol}$
- molar mass of Al = 26.98 g mol^{-1}
- no. moles of Al = $\frac{100000}{26.98} = 3706 \text{ mol}$

437.0 mol of Mn_3O_4 will react with $437.0 \times 8/3$, i.e. 1165 mol. The number of moles of Al is greater than this, so Al is present in excess and Mn_3O_4 is the limiting reactant. So Mn_3O_4 must be used in all calculations.

3 mol Mn_3O_4 produces 9 mol Mn. Therefore 437.0 mol of Mn_3O_4 will produce 437.0×3 , i.e. 1311, mol of Mn.

$$\text{molar mass of Mn} = 54.94 \text{ g mol}^{-1}$$

mass of Mn = 1311×54.94 , i.e. 72 030 g, i.e. 72.03 kg. [4]



$$200.0 \text{ kg of Mn is } \frac{200000}{54.94}, \text{ i.e. } 3640 \text{ mol.}$$

This number of moles is produced from $\frac{3640}{3}$, i.e. 1213 mol Mn_3O_4 . The mass of 1213 mol Mn_3O_4 is $1213 \times 228.82 = 277\,661 \text{ g}$, i.e. 277.7 kg. To convert to tonnes, we divide by 1000 to get 0.2777 tonnes.

Therefore, the percentage Mn_3O_4 in the ore = $\frac{0.2777}{1.23} \times 100$, i.e. 22.6%. [3]

- 13 a A hydrocarbon contains carbon and hydrogen only. The percentage hydrogen in the hydrocarbon is $100 - 88.8$, i.e. 11.2%.

	C	H
	88.8	11.2
	<u>88.8</u>	<u>11.2</u>
divide by A_r	12.01	1.01
moles	7.39	11.09
	<u>7.39</u>	<u>11.09</u>
divide by smallest	7.39	7.39
ratio	1	1.5

Multiplying by 2 to get whole numbers, we get C_2H_3 , which is the empirical formula. [3]

- b To do this, we have to work out the relative molecular mass of the hydrocarbon.

Use $PV = nRT$ to calculate the number of moles. Convert volume in cm^3 to volume in m^3 :

$$\frac{98.9}{(1 \times 10^6)} = 9.89 \times 10^{-5} \text{ m}^3$$

$$P = 1.00 \times 10^5 \text{ Pa} \quad V = 9.89 \times 10^{-5} \text{ m}^3 \quad n = ?$$

$$R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1} \quad T = 320 \text{ K}$$

$$n = \frac{PV}{RT}$$

$$n = \frac{1.00 \times 10^5 \times 9.89 \times 10^{-5}}{8.31 \times 320}$$

$$n = 3.72 \times 10^{-3} \text{ mol}$$

$$\text{relative molecular mass} = \frac{\text{mass}}{\text{no. moles}} = \frac{0.201}{3.72 \times 10^{-3}}$$

$$= 54.0$$

The empirical formula mass
 $= (2 \times 12.01) + (3 \times 1.01) = 27.05$ and
 $\frac{54.0}{27.05} = 2$

Therefore the molecular formula is $(\text{C}_2\text{H}_3)_2$, i.e. C_4H_6 . [3]

14 a volume of CO_2 in $\text{m}^3 = \frac{258}{1\,000\,000}$
 $= 2.58 \times 10^{-4} \text{ m}^3$
 $P = 1.10 \times 10^5 \text{ Pa}$ $V = 2.58 \times 10^{-4} \text{ m}^3$ $n = ?$
 $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ $T = 300 \text{ K}$
 $n = \frac{PV}{RT}$
 $n = \frac{1.10 \times 10^5 \times 2.58 \times 10^{-4}}{8.31 \times 300}$
 $= 0.0114 \text{ mol}$ [3]

b The number of moles of CaCO_3 that must react to produce this number of moles of CO_2 is worked out from the chemical equation:
no. moles of $\text{CaCO}_3 = 0.0114 \text{ mol}$
molar mass of $\text{CaCO}_3 = 100.09 \text{ g mol}^{-1}$
mass of $\text{CaCO}_3 = 0.0114 \times 100.09 = 1.14 \text{ g}$
percentage CaCO_3 in the limestone $= \frac{1.14}{1.20} \times 100$
 $= 95.0\%$ [3]

15 a In this question the number of moles of copper(II) nitrate is equivalent to the number of moles of Cu^{2+} and the number of moles of potassium iodide is equivalent to the number of moles of I^- .

$$\text{no. moles of copper(II) nitrate} = \frac{25.0}{1000} \times 0.100$$

$$= 2.50 \times 10^{-3} \text{ mol}$$

$$\text{no. moles of potassium iodide} = \frac{15.0}{1000} \times 0.500$$

$$= 7.50 \times 10^{-3} \text{ mol}$$

From the ionic equation we can deduce that two moles of $\text{Cu}(\text{NO}_3)_2$ will react with four moles of KI . Therefore $2.50 \times 10^{-3} \text{ mol}$ of $\text{Cu}(\text{NO}_3)_2$ will react with $2 \times 2.50 \times 10^{-3}$, i.e. $5.00 \times 10^{-3} \text{ mol}$ of KI . The number of moles of potassium iodide present is greater than this, so the KI is present in excess. [3]

b We must use the number of moles of the limiting reactant ($\text{Cu}(\text{NO}_3)_2$) for subsequent calculations. From the chemical equation, 2 mol Cu^{2+} react to form 1 mol I_2 . Therefore $2.50 \times 10^{-3} \text{ mol}$ of $\text{Cu}(\text{NO}_3)_2$ will react to form $\frac{2.50 \times 10^{-3}}{2}$, i.e. $1.25 \times 10^{-3} \text{ mol I}_2$.

$$\text{molar mass of I}_2 = 253.80 \text{ g mol}^{-1}$$

$$\text{mass of I}_2 = 1.25 \times 10^{-3} \times 253.80, \text{ i.e. } 0.317 \text{ g} \quad [3]$$

16 a molar mass of $\text{PbI}_2 = 461.0 \text{ g mol}^{-1}$
moles of $\text{PbI}_2 = \frac{0.1270}{461.0} = 2.755 \times 10^{-4} \text{ mol}$ [2]
b $\text{Pb}(\text{NO}_3)_2(\text{aq}) + \text{MI}_2(\text{aq}) \rightarrow \text{PbI}_2(\text{s}) + \text{M}(\text{NO}_3)_2(\text{aq})$ [1]
c From the chemical equation, we can deduce that the number of moles of MI_2 is the same as the number of moles of PbI_2 . Therefore the number of moles of MI_2 is $2.755 \times 10^{-4} \text{ mol}$. [1]
d We know the mass of $2.755 \times 10^{-4} \text{ mol}$ of MI_2 is 0.0810 g. The molar mass of MI_2 is $\frac{0.0810}{2.755 \times 10^{-4}}$, i.e. 294.0 g mol^{-1} . Some of this mass is due to the two I^- ions in the formula – these contribute 2×126.90 to the mass, i.e. 253.8. The relative atomic mass of M is $294.0 - 253.8 = 40.20$. We know that this is a group 2 element, so from the periodic table we can see that it must be calcium. [3]

17 a molar mass of $\text{BaSO}_4 = 233.40 \text{ g mol}^{-1}$
no. moles of BaSO_4 formed $= \frac{3.739 \times 10^{-2}}{233.40}$
 $= 1.602 \times 10^{-4} \text{ mol}$ [2]
b $\text{CuSO}_4(\text{aq}) + \text{BaCl}_2(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + \text{CuCl}_2(\text{aq})$ [1]
c From the chemical equation we can deduce that the number of moles of CuSO_4 is the same as the number of moles of BaSO_4 . Therefore the number of moles of CuSO_4 is $1.602 \times 10^{-4} \text{ mol}$. [1]
d Only 10.00 cm^3 of the original solution (100.0 cm^3) was used in the reaction, so the number of moles of CuSO_4 that were dissolved in water was $10.00 \times 1.602 \times 10^{-4} \text{ mol}$, i.e. $1.602 \times 10^{-3} \text{ mol}$. [1]
e 0.4000 g of hydrated copper sulfate ($\text{CuSO}_4 \cdot x\text{H}_2\text{O}$) contains $1.602 \times 10^{-3} \text{ mol}$ of CuSO_4 . The molar mass of CuSO_4 is $159.62 \text{ g mol}^{-1}$. The mass of CuSO_4 present in the sample is $1.602 \times 10^{-3} \times 159.62$, i.e. 0.2557 g of CuSO_4 . The rest of the hydrated copper sulfate is water. Therefore the mass of water present in the sample is $0.4000 - 0.2557$, i.e. 0.1443 g.
no. moles of water $= \frac{0.1443}{18.02} = 8.008 \times 10^{-3} \text{ mol}$
ratio of no. moles of water to no. moles of CuSO_4
 $= \frac{8.008 \times 10^{-3}}{1.602 \times 10^{-3}} = 4.999$
This will be a whole number in the formula. Therefore the value of x is 5, and the formula is $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$. [3]