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 $dm^3 = \frac{227}{1000} = 0.227\ dm^3$

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

$= 0.0100\ mol$

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

$0.0100\ mol\ NH_3 \rightarrow 0.00500\ mol\ N_2$

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

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 $Mn_3O_4 = 228.82\ g\ mol^{-1}$

no. moles of $Mn_3O_4 = \frac{100\,000}{228.82} = 437.0\ mol$

molar mass of $Al = 26.98\ g\ mol^{-1}$

no. moles of $Al = \frac{100\,000}{26.98} = 3706\ mol$

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.

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

mass of of $Mn = 1311 \times 54.94$, i.e. 72 030 g, i.e. 72.03 kg. **[4]**

b $3Mn_3O_4 + 8Al \rightarrow 4Al_2O_3 + 9Mn$

200.0 kg of Mn is $\frac{200\,000}{54.94}$, i.e. 3640 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 $1213 \times 228.82 = 277\,661$ 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
divide by A_r	$\frac{88.8}{12.01}$	$\frac{11.2}{1.01}$
moles	7.39	11.09
divide by smallest	$\frac{7.39}{7.39}$	$\frac{11.09}{7.39}$
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:

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

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

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

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

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

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

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 $(C_2H_3)_2$, i.e. C_4H_6. **[3]**

14 a volume of CO_2 in $m^3 = \frac{258}{1\,000\,000}$

$= 2.58 \times 10^{-4}\,m^3$

$P = 1.10 \times 10^5\,Pa \quad V = 2.58 \times 10^{-4}\,m^3 \quad n = ?$

$R = 8.31\,J\,K^{-1}\,mol^{-1} \quad T = 300\,K$

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

$n = \frac{1.10 \times 10^5 \times 2.58 \times 10^{-4}}{8.31 \times 300}$

$= 0.0114\,mol$ **[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 $CaCO_3 = 0.0114\,mol$

molar mass of $CaCO_3 = 100.09\,g\,mol^{-1}$

mass of $CaCO_3 = 0.0114 \times 100.09 = 1.14\,g$

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^-.

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

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

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

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

From the ionic equation we can deduce that two moles of $Cu(NO_3)_2$ will react with four moles of KI. Therefore $2.50 \times 10^{-3}\,mol$ of $Cu(NO_3)_2$ will react with $2 \times 2.50 \times 10^{-3}$, i.e. $5.00 \times 10^{-3}\,mol$ 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 ($Cu(NO_3)_2$) for subsequent calculations. From the chemical equation, 2 mol Cu^{2+} react to form 1 mol I_2. Therefore $2.50 \times 10^{-3}\,mol$ of $Cu(NO_3)_2$ will react to form $\frac{2.50 \times 10^{-3}}{2}$, i.e. $1.25 \times 10^{-3}\,mol\ I_2$.

molar mass of $I_2 = 253.80\,g\,mol^{-1}$

mass of $I_2 = 1.25 \times 10^{-3} \times 253.80$, i.e. $0.317\,g$ **[3]**

16 a molar mass of $PbI_2 = 461.0\,g\,mol^{-1}$

moles of $PbI_2 = \frac{0.1270}{461.0} = 2.755 \times 10^{-4}\,mol$ **[2]**

b $Pb(NO_3)_2(aq) + MI_2(aq) \rightarrow PbI_2(s) + M(NO_3)_2(aq)$ **[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 $2.755 \times 10^{-4}\,mol$. **[1]**

d We know the mass of $2.755 \times 10^{-4}\,mol$ 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 $BaSO_4 = 233.40\,g\,mol^{-1}$

no. moles of $BaSO_4$ formed $= \frac{3.739 \times 10^{-2}}{233.40}$

$= 1.602 \times 10^{-4}\,mol$ **[2]**

b $CuSO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + CuCl_2(aq)$ **[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 $1.602 \times 10^{-4}\,mol$. **[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 $10.00 \times 1.602 \times 10^{-4}\,mol$, i.e. $1.602 \times 10^{-3}\,mol$. **[1]**

e $0.4000\,g$ of hydrated copper sulfate ($CuSO_4{\cdot}xH_2O$) contains $1.602 \times 10^{-3}\,mol$ of $CuSO_4$. The molar mass of $CuSO_4$ is $159.62\,g\,mol^{-1}$. 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 $= \frac{0.1443}{18.02} = 8.008 \times 10^{-3}\,mol$

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 $CuSO_4{\cdot}5H_2O$. **[3]**