SHEET 6 ANSWERS

1. 16.0 g of Ca_2C reacts with 42.0 g of H_2O according to the following reaction:

$$CaC_2(s) + 2\,H_2O(l) \longrightarrow C_2H_2(g) + Ca(OH)_2(s)$$

a) Determine which reactant in the limiting reagent.
b) What mass of $C_2H_2(g)$ and $Ca(OH)_2(s)$ is produced.
c) Calculate the excess mass of the excess reagent

$$\underset{16.0\text{ g}}{CaC_2(s)} \;+\; \underset{42.0\text{ g}}{2\,H_2O(l)} \;\longrightarrow\; \underset{?\text{ g}}{C_2H_2(g)} \;+\; \underset{?\text{ g}}{Ca(OH)_2(s)}$$

a) This question is usually not asked directly. If there is information given about two or more reactants, this step MUST BE TAKEN.

Consider CaC_2 :

$$16.0\text{ g }CaC_2 \text{ x } \frac{1\text{ mol }CaC_2}{64.10\text{ g }CaC_2} = 0.250\text{ mol }CaC_2\text{ available}$$

$$0.250\text{ mol }CaC_2 \text{ x } \frac{2\text{ mol }H_2O}{1\text{ mol }CaC_2} = 0.499\text{ mol }H_2O\text{ required}$$

Consider H_2O :

$$42.0\text{ g }H_2O \text{ x } \frac{1\text{ mol }H_2O}{18.02\text{ g }H_2O} = 2.33\text{ mol }H_2O\text{ available}$$

$$2.23\text{ mol }H_2O \text{ x } \frac{1\text{ mol }CaC_2}{2\text{ mol }H_2O} = 1.17\text{ mol }CaC_2\text{ required}$$

therefore the limiting reagent is CaC_2

b) These are examples of typical final questions.

$$0.250\text{ mol }CaC_2 \text{ x } \frac{1\text{ mol }C_2H_2}{1\text{ mol }CaC_2} \text{ x } \frac{26.04\text{ g }C_2H_2}{1\text{ mol }C_2H_2} = 6.50\text{ g }C_2H_2\text{ produced}$$

$$0.250\text{ mol }CaC_2 \text{ x } \frac{1\text{ mol }Ca(OH)_2}{1\text{ mol }CaC_2} \text{ x } \frac{74.10\text{ g }Ca(OH)_2}{1\text{ mol }Ca(OH)_2} = 18.5\text{ g }Ca(OH)_2\text{ produced}$$

c) This is not a typical question but it helps to point out that there will be left overs for the excess reagent.

$$0.250\text{ mol }CaC_2 \text{ x } \frac{2\text{ mol }H_2O}{1\text{ mol }CaC_2} \text{ x } \frac{18.02\text{ g }H_2O}{1\text{ mol }H_2O} = 9.00\text{ g }H_2O\text{ consumed}$$

$$\begin{aligned}\text{mass }H_2O\text{ excess} &= (\text{mass }H_2O\text{ available}) - (\text{mass }H_2O\text{ consumed})\\ &= (42.0\text{ g }H_2O) - (9.00\text{ g }H_2O)\\ &= 33.0\text{ g }H_2O\text{ remains after reaction}\end{aligned}$$

2. Consider the following reaction at S.T.P. If 35 g of tungsten trioxide reacts with 15 L of H_2 at S.T.P., what mass of tungsten is produced?

$$\underset{35.0\text{ g}}{WO_3(s)} + \underset{15.0\text{ L}}{3\,H_2(g)} \longrightarrow \underset{?\text{ g}}{W(s)} + 3\,H_2O(aq)$$

Consider WO_3 :

$$35.0\text{ g }WO_3 \text{ x } \frac{1\text{ mol }WO_3}{231.84\text{ g }WO_3} = 0.151\text{ mol }WO_3\text{ available}$$

$$0.151\text{ mol }WO_3 \text{ x } \frac{3\text{ mol }H_2}{1\text{ mol }WO_3} = 0.453\text{ mol }H_2\text{ required}$$

Consider H_2 :

$$15.0\text{ L }H_2 \text{ x } \frac{1\text{ mol }H_2}{22.414\text{ L }H_2} = 0.669\text{ mol }H_2\text{ available}$$

$$0.669\text{ mol }H_2 \text{ x } \frac{1\text{ mol }WO_3}{3\text{ mol }H_2} = 0.223\text{ mol }WO_3\text{ required}$$

therefore the limiting reagent is WO_3

$$0.151\text{ mol }WO_3 \text{ x } \frac{1\text{ mol W}}{1\text{ mol }WO_3} \text{ x } \frac{183.84\text{ g W}}{1\text{ mol W}} = 27.8\text{ g W}$$

3. What mass of H_2SO_4 can be produced from 50.0 g of SO_2 , 15.0 g O_2 and an unlimited amount of H_2O ? The equation is:

$$\underset{50.0\text{ g}}{2\,SO_2(g)} + \underset{15.0\text{ g}}{O_2(g)} + 2\,H_2O(l) \longrightarrow \underset{?\text{ g}}{2\,H_2SO_4(aq)}$$

Consider SO_2 :

$$50.0 \text{ g } SO_2 \text{ x } \frac{1 \text{ mol } SO_2}{64.07 \text{ g } SO_2} = 0.780 \text{ mol } SO_2 \text{ available}$$

$$0.780 \text{ mol } SO_2 \text{ x } \frac{1 \text{ mol } O_2}{2 \text{ mol } SO_2} = 0.390 \text{ mol } O_2 \text{ required}$$

Consider O_2 :

$$15.0 \text{ g } O_2 \text{ x } \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} = 0.469 \text{ mol } O_2 \text{ available}$$

$$0.469 \text{ mol } O_2 \text{ x } \frac{2 \text{ mol } SO_2}{1 \text{ mol } O_2} = 0.938 \text{ mol } SO_2 \text{ required}$$

therefore the limiting reagent is SO_2

$$0.780 \text{ mol } SO_2 \text{ x } \frac{2 \text{ mol } H_2SO_4}{2 \text{ mol } SO_2} \text{ x } \frac{98.09 \text{ g } H_2SO_4}{1 \text{ mol } H_2SO_4} = 76.5 \text{ g } H_2SO_4$$

4. 40.0 L of O_2 react with 19.6 L of methane (CH_4) at S.T.P. according to the reaction shown below. What volume of water and carbon dioxide are produced at S.T.P.

$$\underset{19.6\text{ L}}{CH_4(g)} + \underset{40.0\text{ L}}{2\,O_2(g)} \longrightarrow \underset{?\text{ L @ S.T.P.}}{CO_2(g)} + \underset{?\text{ L @ S.T.P.}}{2\,H_2O(g)}$$

Consider CH_4 :

$$19.6 \text{ L } CH_4 \text{ x } \frac{1 \text{ mol } CH_4}{22.414 \text{ L } CH_4} = 0.874 \text{ mol } CH_4 \text{ available}$$

$$0.874 \text{ mol } CH_4 \text{ x } \frac{2 \text{ mol } O_2}{1 \text{ mol } CH_4} = 1.75 \text{ mol } O_2 \text{ required}$$

Consider O_2 :

$$40.0 \text{ L } O_2 \text{ x } \frac{1 \text{ mol } O_2}{22.414 \text{ L } O_2} = 1.78 \text{ mol } O_2 \text{ available}$$

$$1.78 \text{ mol } O_2 \text{ x } \frac{1 \text{ mol } CH_4}{2 \text{ mol } O_2} = 0.892 \text{ mol } CH_4 \text{ required}$$

therefore the limiting reagent is CH_4

$$0.874 \text{ mol } CH_4 \text{ x } \frac{2 \text{ mol } H_2O}{1 \text{ mol } CH_4} \text{ x } \frac{22.414 \text{ L } H_2O}{1 \text{ mol } H_2O} = 39.2 \text{ L } H_2O$$

$$0.874 \text{ mol } CH_4 \text{ x } \frac{1 \text{ mol } CO_2}{1 \text{ mol } CH_4} \text{ x } \frac{22.414 \text{ L } CO_2}{1 \text{ mol } CO_2} = 19.6 \text{ L } CO_2$$

5. What is the maximum mass of carbon dioxide that can be produced by the reaction between 15.0 g of propane (C_3H_8) with 60.0 g of oxygen gas?

$$\underset{15.0\text{ g}}{C_3H_8(g)} + \underset{60.0\text{ g}}{5\,O_2(g)} \longrightarrow \underset{?\text{ g}}{3\,CO_2(g)} + 4\,H_2O(g)$$

Consider C_3H_8 :

$$15.0 \text{ g } C_3H_8 \text{ x } \frac{1 \text{ mol } C_3H_8}{44.11 \text{ g } C_3H_8} = 0.340 \text{ mol } C_3H_8 \text{ available}$$

$$0.340 \text{ mol } C_3H_8 \text{ x } \frac{5 \text{ mol } O_2}{1 \text{ mol } C_3H_8} = 1.70 \text{ mol } O_2 \text{ required}$$

Consider O_2 :

$$60.0 \text{ g } O_2 \text{ x } \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} = 1.88 \text{ mol } O_2 \text{ available}$$

$$1.88 \text{ mol } O_2 \text{ x } \frac{1 \text{ mol } C_3H_8}{5 \text{ mol } O_2} = 0.375 \text{ mol } C_3H_8 \text{ required}$$

therefore the limiting reagent is C_3H_8

$$0.340 \text{ mol } C_3H_8 \text{ x } \frac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8} \text{ x } \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 44.9 \text{ g } CO_2$$

6. What mass of iron(III) oxide is produced when 20.9 g of iron(II) sulphide reacts with 10.0 L of oxygen gas at 105.5 kPa and a temperature of 42 °C? What volume of sulphur dioxide is produced at S.T.P.?

$$\underset{20.9\ \text{g}}{4\,\mathrm{FeS(s)}} \;+\; \underset{\substack{10.0\ \text{L}\\105.5\ \text{kPa}\\42\ ^\circ\text{C}}}{7\,\mathrm{O_2(g)}} \longrightarrow \underset{?\ \text{g}}{2\,\mathrm{Fe_2O_3(s)}} \;+\; \underset{?\ \text{L @ S.T.P.}}{4\,\mathrm{SO_2(g)}}$$

Consider O_2 :

$$\begin{aligned} P &= 105.5\ \text{kPa} \\ V &= 10.0\ \text{L} \\ n &= ? \\ R &= 8.314\frac{\text{kPa}\cdot\text{L}}{\text{K}\cdot\text{mol}} \\ T &= 42\ ^\circ\text{C} \rightarrow 315.15\ \text{K} \end{aligned}$$

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

$$n = \frac{105.5\ \text{kPa} \times 10.0\ \text{L}}{8.314\frac{\text{kPa}\cdot\text{L}}{\text{K}\cdot\text{mol}} \times 315.15\ \text{K}}$$

$$n = 0.403\ \text{mol } O_2 \text{ available}$$

$$0.403\ \text{mol } O_2 \text{ x } \frac{4\ \text{mol FeS}}{7\ \text{mol } O_2} = 0.230\ \text{mol FeS required}$$

Consider FeS :

$$20.9\ \text{g FeS x } \frac{1\ \text{mol FeS}}{87.92\ \text{g FeS}} = 0.238\ \text{mol FeS available}$$

$$0.238\ \text{mol FeS x } \frac{7\ \text{mol } O_2}{4\ \text{mol FeS}} = 0.416\ \text{mol } O_2 \text{ required}$$

---

therefore the limiting reagent is O_2

$$0.403\ \text{mol } O_2 \text{ x } \frac{2\ \text{mol } Fe_2O_3}{7\ \text{mol } O_2} \text{ x } \frac{159.70\ \text{g } Fe_2O_3}{1\ \text{mol } Fe_2O_3} = 18.4\ \text{g } Fe_2O_3$$

$$0.403\ \text{mol } O_2 \text{ x } \frac{4\ \text{mol } SO_2}{7\ \text{mol } O_2} \text{ x } \frac{22.414\ \text{L } SO_2}{1\ \text{mol } SO_2} = 5.16\ \text{L } SO_2$$

---

7. Nickel metal can be highly purified using the Mond Process:

$$\mathrm{Ni(s)} + 4\,\mathrm{CO(g)} \longrightarrow \mathrm{Ni(CO)_4(g)}$$

In the first step of this process nickel metal is reacted with carbon monoxide under high pressure and heat to produce a gas product known as nickel carbonyl ($Ni(CO)_4$). If 40.0 g of nickel metal is reacted with 5.00 L of carbon monoxide at 60.75 atm. pressure and a temperature of 875 K, calculate the resulting total pressure of all gases at 25 ^{o}C and total volume 5.00 L. Hints: nickel is the limiting reagent, Dalton's Law of Partial Pressures could be used to solve this problem

---

Calculate the amount of CO(g) available:

$$\mathrm{P} = 60.75 \text{ atm kPa x } \frac{101.325 \text{ kPa}}{1 \text{ atm}} = 6155 \text{ kPa}$$

$$\mathrm{V} = 5.00 \text{ L}$$

$$\mathrm{n} = ?$$

$$\mathrm{R} = 8.314 \frac{\mathrm{kPa \cdot L}}{\mathrm{K \cdot mol}}$$

$$\mathrm{T} = 875 \text{ K}$$

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

$$\mathrm{n} = \frac{6155 \text{ kPa} \times 5.00 \text{ L}}{8.314 \frac{\mathrm{kPa \cdot L}}{\mathrm{K \cdot mol}} \times 875 \text{ K}}$$

$$\mathrm{n} = 4.231 \text{ mol CO available}$$

Calculate amount of CO(g) consumed in the reaction:

$$40.0 \text{ g Ni x } \frac{1 \text{ mol Ni}}{58.69 \text{ g Ni}} \text{ x } \frac{4 \text{ mol CO}}{1 \text{ mol Ni}} = 2.726 \text{ mol CO consumed}$$

Calculate the amount of CO remaining (unreacted):

$$\begin{aligned} \text{amount CO remaining} &= (\text{amount CO available}) - (\text{amount CO consumed}) \\ &= (4.231 \text{ mol CO}) - (2.726 \text{ mol CO}) \\ &= 1.505 \text{ mol CO remains after reaction} \end{aligned}$$

Calculate the amount of $Ni(CO)_4$ formed:

$$40.0 \text{ g Ni x } \frac{1 \text{ mol Ni}}{58.69 \text{ g Ni}} \text{ x } \frac{1 \text{ mol Ni(CO)}_4}{1 \text{ mol Ni}} = 0.6815 \text{ mol Ni(CO)}_4 \text{ formed}$$

Calculate the total amount of gases after reaction:

$$\begin{aligned} \text{total amount of gases} &= (\text{amount CO remaining}) - (\text{amount Ni(CO)}_4 \text{ formed}) \\ &= (1.505 \text{ mol CO}) + (0.6815 \text{ mol Ni(CO)}_4) \\ &= 2.186 \text{ mol of gas remains after reaction} \end{aligned}$$

Calculate the pressure of remaining gas:

$$
\begin{aligned}
\mathrm{P} &= ? \\
\mathrm{V} &= 5.00\ \mathrm{L} \\
\mathrm{n} &= 2.186\ \mathrm{mol\ gas} \\
\mathrm{R} &= 8.314\frac{\mathrm{kPa\cdot L}}{\mathrm{K\cdot mol}} \\
\mathrm{T} &= 25\ ^{\circ}\mathrm{C} \rightarrow 298.15\ \mathrm{K}
\end{aligned}
$$

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

$$\mathrm{P} = \frac{2.186\ \mathrm{mol} \times 8.314\frac{\mathrm{kPa\cdot L}}{\mathrm{K\cdot mol}} \times 298.15\ \mathrm{K}}{5.00\ \mathrm{L}}$$

$$\mathrm{P} = 1084\ \mathrm{kPa}$$

$$1084\ \mathrm{kPa} \times \frac{1\ \mathrm{atm}}{101.325\ \mathrm{kPa}} = 10.696\ \mathrm{atm}$$