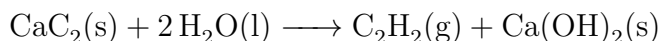
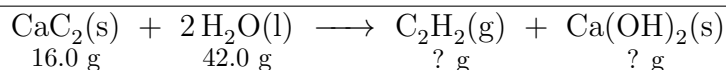


SHEET 6 ANSWERS

1. 16.0 g of CaC_2 reacts with 42.0 g of H_2O according to the following reaction:



- Determine which reactant is the limiting reagent.
- What mass of $\text{C}_2\text{H}_2(\text{g})$ and $\text{Ca}(\text{OH})_2(\text{s})$ is produced.
- Calculate the excess mass of the excess reagent



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 :

$$16.0 \text{ g CaC}_2 \times \frac{1 \text{ mol CaC}_2}{64.10 \text{ g CaC}_2} = 0.250 \text{ mol CaC}_2 \text{ available}$$

$$0.250 \text{ mol CaC}_2 \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol CaC}_2} = 0.499 \text{ mol H}_2\text{O required}$$

Consider H_2O :

$$42.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 2.33 \text{ mol H}_2\text{O available}$$

$$2.23 \text{ mol H}_2\text{O} \times \frac{1 \text{ mol CaC}_2}{2 \text{ mol H}_2\text{O}} = 1.17 \text{ mol CaC}_2 \text{ required}$$

therefore the limiting reagent is CaC_2

b) These are examples of typical final questions.

$$0.250 \text{ mol CaC}_2 \times \frac{1 \text{ mol C}_2\text{H}_2}{1 \text{ mol CaC}_2} \times \frac{26.04 \text{ g C}_2\text{H}_2}{1 \text{ mol C}_2\text{H}_2} = 6.50 \text{ g C}_2\text{H}_2 \text{ produced}$$

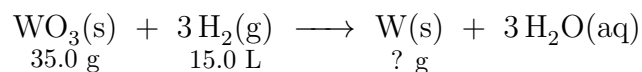
$$0.250 \text{ mol CaC}_2 \times \frac{1 \text{ mol Ca}(\text{OH})_2}{1 \text{ mol CaC}_2} \times \frac{74.10 \text{ g Ca}(\text{OH})_2}{1 \text{ mol Ca}(\text{OH})_2} = 18.5 \text{ g Ca}(\text{OH})_2 \text{ produced}$$

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

$$0.250 \text{ mol CaC}_2 \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol CaC}_2} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 9.00 \text{ g H}_2\text{O consumed}$$

$$\begin{aligned} \text{mass H}_2\text{O excess} &= (\text{mass H}_2\text{O available}) - (\text{mass H}_2\text{O consumed}) \\ &= (42.0 \text{ g H}_2\text{O}) - (9.00 \text{ g H}_2\text{O}) \\ &= 33.0 \text{ g H}_2\text{O remains after reaction} \end{aligned}$$

2. Consider the following reaction at S.T.P. If 35 g of tungsten trioxide reacts with 15 L of H₂ at S.T.P., what mass of tungsten is produced?



Consider WO₃ :

$$35.0 \text{ g WO}_3 \times \frac{1 \text{ mol WO}_3}{231.84 \text{ g WO}_3} = 0.151 \text{ mol WO}_3 \text{ available}$$

$$0.151 \text{ mol WO}_3 \times \frac{3 \text{ mol H}_2}{1 \text{ mol WO}_3} = 0.453 \text{ mol H}_2 \text{ required}$$

Consider H₂ :

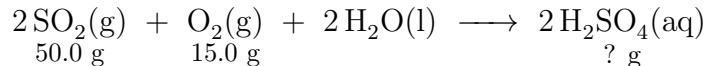
$$15.0 \text{ L H}_2 \times \frac{1 \text{ mol H}_2}{22.414 \text{ L H}_2} = 0.669 \text{ mol H}_2 \text{ available}$$

$$0.669 \text{ mol H}_2 \times \frac{1 \text{ mol WO}_3}{3 \text{ mol H}_2} = 0.223 \text{ mol WO}_3 \text{ required}$$

therefore the limiting reagent is WO₃

$$0.151 \text{ mol WO}_3 \times \frac{1 \text{ mol W}}{1 \text{ mol WO}_3} \times \frac{183.84 \text{ g W}}{1 \text{ mol W}} = 27.8 \text{ g W}$$

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:



Consider SO_2 :

$$50.0 \text{ g SO}_2 \times \frac{1 \text{ mol SO}_2}{64.07 \text{ g SO}_2} = 0.780 \text{ mol SO}_2 \text{ available}$$

$$0.780 \text{ mol SO}_2 \times \frac{1 \text{ mol O}_2}{2 \text{ mol SO}_2} = 0.390 \text{ mol O}_2 \text{ required}$$

Consider O_2 :

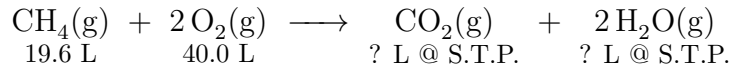
$$15.0 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 0.469 \text{ mol O}_2 \text{ available}$$

$$0.469 \text{ mol O}_2 \times \frac{2 \text{ mol SO}_2}{1 \text{ mol O}_2} = 0.938 \text{ mol SO}_2 \text{ required}$$

therefore the limiting reagent is SO_2

$$0.780 \text{ mol SO}_2 \times \frac{2 \text{ mol H}_2\text{SO}_4}{2 \text{ mol SO}_2} \times \frac{98.09 \text{ g H}_2\text{SO}_4}{1 \text{ mol H}_2\text{SO}_4} = 76.5 \text{ g H}_2\text{SO}_4$$

4. 40.0 L of O₂ react with 19.6 L of methane (CH₄) at S.T.P. according to the reaction shown below. What volume of water and carbon dioxide are produced at S.T.P.



Consider CH₄ :

$$19.6 \text{ L CH}_4 \times \frac{1 \text{ mol CH}_4}{22.414 \text{ L CH}_4} = 0.874 \text{ mol CH}_4 \text{ available}$$

$$0.874 \text{ mol CH}_4 \times \frac{2 \text{ mol O}_2}{1 \text{ mol CH}_4} = 1.75 \text{ mol O}_2 \text{ required}$$

Consider O₂ :

$$40.0 \text{ L O}_2 \times \frac{1 \text{ mol O}_2}{22.414 \text{ L O}_2} = 1.78 \text{ mol O}_2 \text{ available}$$

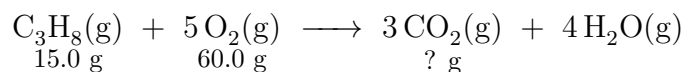
$$1.78 \text{ mol O}_2 \times \frac{1 \text{ mol CH}_4}{2 \text{ mol O}_2} = 0.892 \text{ mol CH}_4 \text{ required}$$

therefore the limiting reagent is CH₄

$$0.874 \text{ mol CH}_4 \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol CH}_4} \times \frac{22.414 \text{ L H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 39.2 \text{ L H}_2\text{O}$$

$$0.874 \text{ mol CH}_4 \times \frac{1 \text{ mol CO}_2}{1 \text{ mol CH}_4} \times \frac{22.414 \text{ L CO}_2}{1 \text{ mol CO}_2} = 19.6 \text{ L CO}_2$$

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?



Consider C_3H_8 :

$$15.0 \text{ g C}_3\text{H}_8 \times \frac{1 \text{ mol C}_3\text{H}_8}{44.11 \text{ g C}_3\text{H}_8} = 0.340 \text{ mol C}_3\text{H}_8 \text{ available}$$

$$0.340 \text{ mol C}_3\text{H}_8 \times \frac{5 \text{ mol O}_2}{1 \text{ mol C}_3\text{H}_8} = 1.70 \text{ mol O}_2 \text{ required}$$

Consider O_2 :

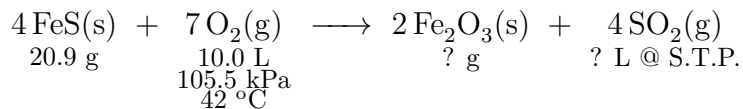
$$60.0 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 1.88 \text{ mol O}_2 \text{ available}$$

$$1.88 \text{ mol O}_2 \times \frac{1 \text{ mol C}_3\text{H}_8}{5 \text{ mol O}_2} = 0.375 \text{ mol C}_3\text{H}_8 \text{ required}$$

therefore the limiting reagent is C_3H_8

$$0.340 \text{ mol C}_3\text{H}_8 \times \frac{3 \text{ mol CO}_2}{1 \text{ mol C}_3\text{H}_8} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 44.9 \text{ g CO}_2$$

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.?



Consider O₂ :

$$\begin{array}{l}
 P = 105.5 \text{ kPa} \\
 V = 10.0 \text{ L} \\
 n = ? \\
 R = 8.314 \frac{\text{kPa} \cdot \text{L}}{\text{K} \cdot \text{mol}} \\
 T = 42 \text{ }^\circ\text{C} \rightarrow 315.15 \text{ K}
 \end{array}
 \qquad
 \begin{array}{l}
 n = \frac{PV}{RT} \\
 n = \frac{105.5 \text{ kPa} \times 10.0 \text{ L}}{8.314 \frac{\text{kPa} \cdot \text{L}}{\text{K} \cdot \text{mol}} \times 315.15 \text{ K}} \\
 n = 0.403 \text{ mol O}_2 \text{ available}
 \end{array}$$

$$0.403 \text{ mol O}_2 \times \frac{4 \text{ mol FeS}}{7 \text{ mol O}_2} = 0.230 \text{ mol FeS required}$$

Consider FeS :

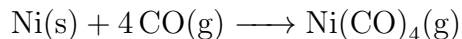
$$\begin{array}{l}
 20.9 \text{ g FeS} \times \frac{1 \text{ mol FeS}}{87.92 \text{ g FeS}} = 0.238 \text{ mol FeS available} \\
 0.238 \text{ mol FeS} \times \frac{7 \text{ mol O}_2}{4 \text{ mol FeS}} = 0.416 \text{ mol O}_2 \text{ required}
 \end{array}$$

therefore the limiting reagent is O₂

$$0.403 \text{ mol O}_2 \times \frac{2 \text{ mol Fe}_2\text{O}_3}{7 \text{ mol O}_2} \times \frac{159.70 \text{ g Fe}_2\text{O}_3}{1 \text{ mol Fe}_2\text{O}_3} = 18.4 \text{ g Fe}_2\text{O}_3$$

$$0.403 \text{ mol O}_2 \times \frac{4 \text{ mol SO}_2}{7 \text{ mol O}_2} \times \frac{22.414 \text{ L SO}_2}{1 \text{ mol SO}_2} = 5.16 \text{ L SO}_2$$

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



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 °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:

$$\begin{aligned} P &= 60.75 \text{ atm kPa} \times \frac{101.325 \text{ kPa}}{1 \text{ atm}} = 6155 \text{ kPa} & n &= \frac{PV}{RT} \\ V &= 5.00 \text{ L} & n &= \frac{6155 \text{ kPa} \times 5.00 \text{ L}}{8.314 \frac{\text{kPa} \cdot \text{L}}{\text{K} \cdot \text{mol}} \times 875 \text{ K}} \\ n &= ? \\ R &= 8.314 \frac{\text{kPa} \cdot \text{L}}{\text{K} \cdot \text{mol}} \\ T &= 875 \text{ K} & n &= 4.231 \text{ mol CO available} \end{aligned}$$

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

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

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:

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

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:

$$\begin{aligned} P &= ? & P &= \frac{nRT}{V} \\ V &= 5.00 \text{ L} \\ n &= 2.186 \text{ mol gas} \\ R &= 8.314 \frac{\text{kPa} \cdot \text{L}}{\text{K} \cdot \text{mol}} \\ T &= 25 \text{ }^\circ\text{C} \rightarrow 298.15 \text{ K} \end{aligned}$$
$$P = \frac{2.186 \text{ mol} \times 8.314 \frac{\text{kPa} \cdot \text{L}}{\text{K} \cdot \text{mol}} \times 298.15 \text{ K}}{5.00 \text{ L}}$$
$$P = 1084 \text{ kPa}$$

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