

# Corrections to key .

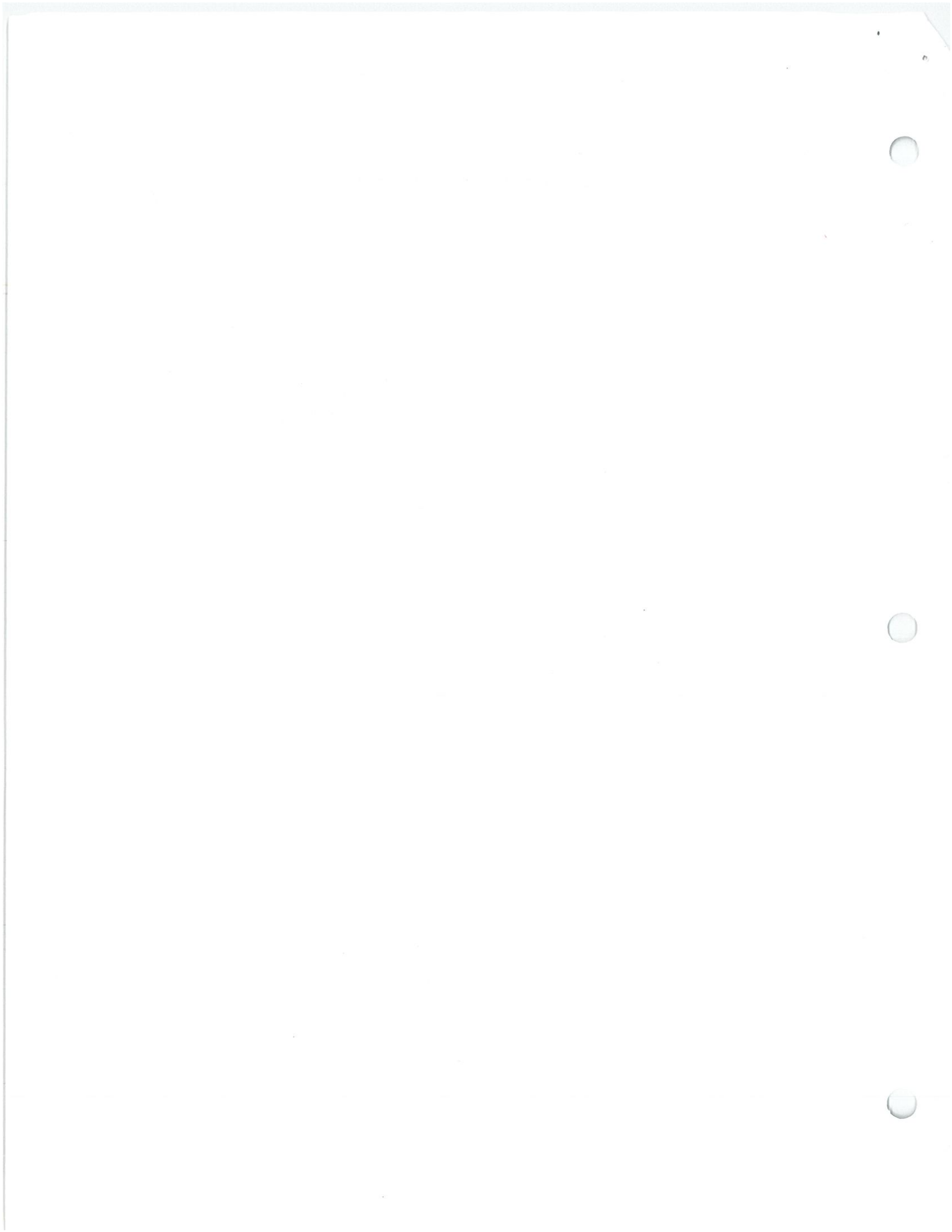
p. 10 1c should be Carbor

~~p. 9 4 should be 50g~~

p. 12 6c should ~~add~~ add ~~(.553g)~~

p. 16 4b 87% 0.553g

p 16 4c 230g



# Mass Relationships in Chemical Reactions

Key



$$\frac{6.23 \text{ mol Li} \mid 1 \text{ mol H}_2}{2 \text{ mol Li}} = 3.12 \text{ mol H}_2$$

$$\textcircled{2} \frac{80.57 \text{ g Li} \mid 1 \text{ mol Li} \mid 1 \text{ mol H}_2 \mid 2.02 \text{ g H}_2}{6.94 \text{ g Li} \mid 2 \text{ mol Li} \mid 1 \text{ mol H}_2} = 11.73 \text{ g H}_2$$



$$\frac{0.254 \text{ mol O}_2 \mid 2 \text{ mol NO}_2}{1 \text{ mol O}_2} = 0.508 \text{ mol NO}_2$$

$$\textcircled{4} \frac{1.44 \text{ g NO} \mid 1 \text{ mol NO} \mid 2 \text{ mol NO}_2 \mid 46.01 \text{ g NO}_2}{30.01 \text{ g NO} \mid 2 \text{ mol NO} \mid 1 \text{ mol NO}_2} = 2.21 \text{ g NO}_2$$

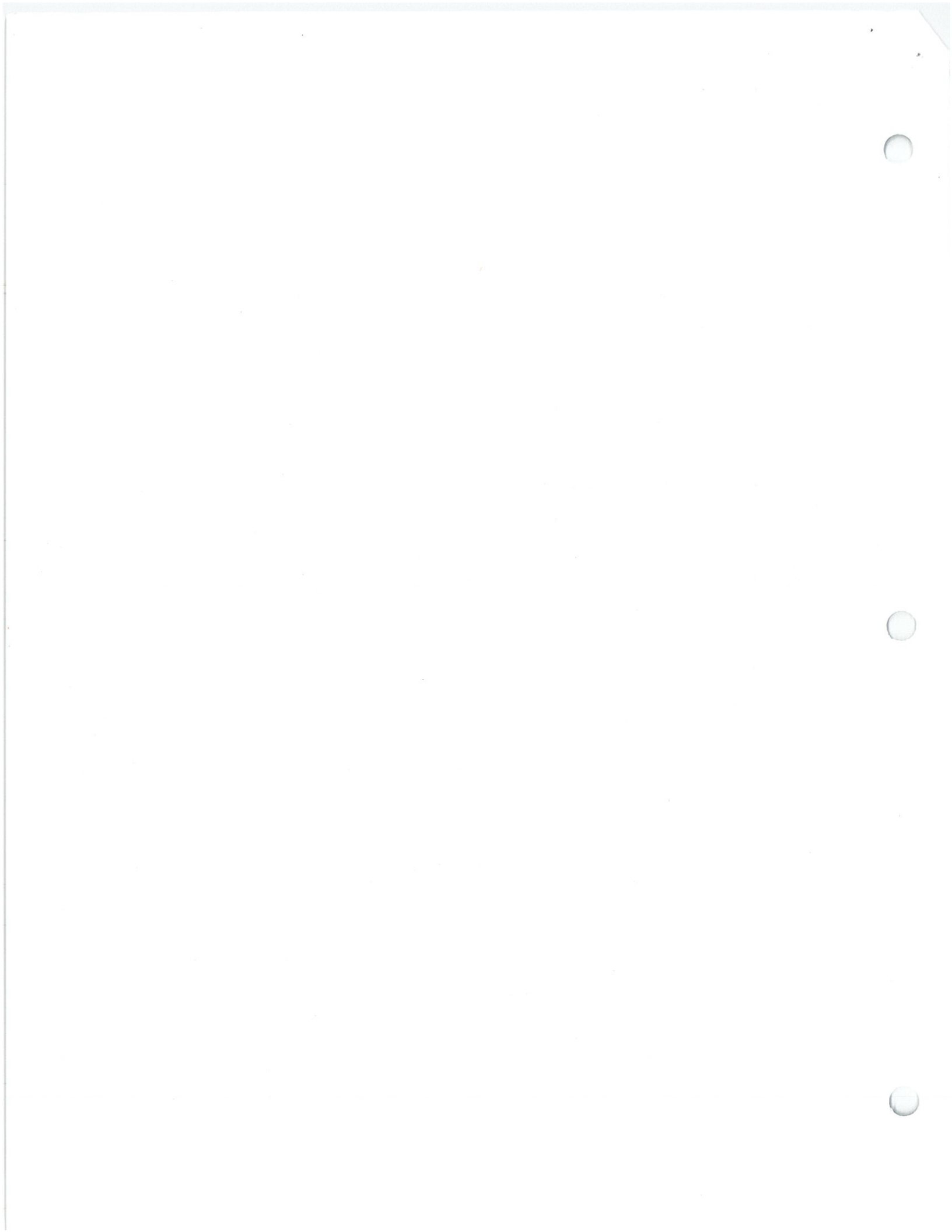


$$\frac{856 \text{ g C}_6\text{H}_{12}\text{O}_6 \mid 1 \text{ mol C}_6\text{H}_{12}\text{O}_6 \mid 6 \text{ mol CO}_2 \mid 44.01 \text{ g CO}_2}{116.06 \text{ g C}_6\text{H}_{12}\text{O}_6 \mid 1 \text{ mol C}_6\text{H}_{12}\text{O}_6 \mid 1 \text{ mol CO}_2} = \text{[scribble]}$$



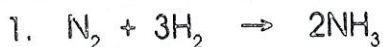
$$\frac{209 \text{ g CH}_3\text{OH} \mid 1 \text{ mol CH}_3\text{OH} \mid 3 \text{ mol O}_2 \mid 32.00 \text{ g O}_2}{32.01 \text{ g CH}_3\text{OH} \mid 2 \text{ mol CH}_3\text{OH} \mid 1 \text{ mol O}_2} = 313 \text{ g O}_2$$

$$\textcircled{7} \frac{209 \text{ g CH}_3\text{OH} \mid 1 \text{ mol CH}_3\text{OH} \mid 2 \text{ mol CO}_2 \mid 44.01 \text{ g CO}_2}{32.01 \text{ g CH}_3\text{OH} \mid 2 \text{ mol CH}_3\text{OH} \mid 1 \text{ mol CO}_2} = 287 \text{ g CO}_2$$

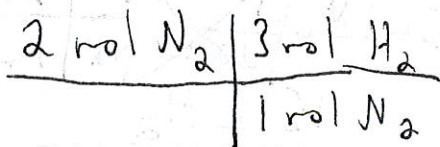


# STOICHIOMETRY: MOLE-MOLE PROBLEMS

Name \_\_\_\_\_



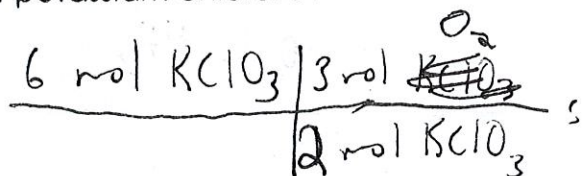
How many moles of hydrogen are needed to completely react with two moles of nitrogen?



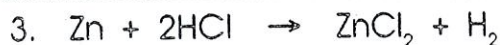
6 mol  $H_2$



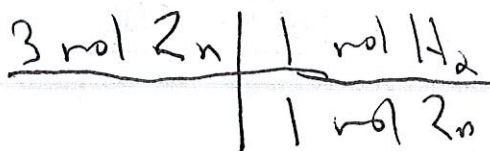
How many moles of oxygen are produced by the decomposition of six moles of potassium chlorate?



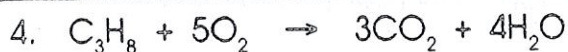
9 mol  $O_2$



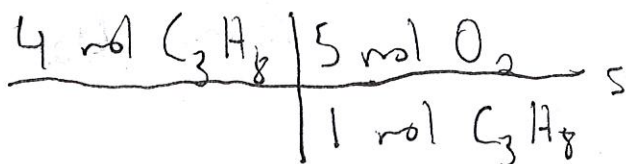
How many moles of hydrogen are produced from the reaction of three moles of zinc with an excess of hydrochloric acid?



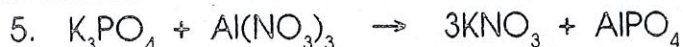
3 mol  $H_2$



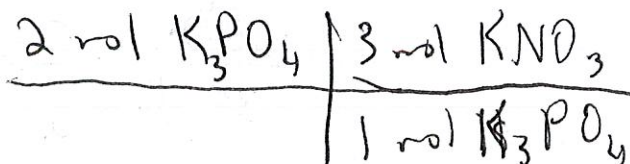
How many moles of oxygen are necessary to react completely with four moles of propane ( $C_3H_8$ )?



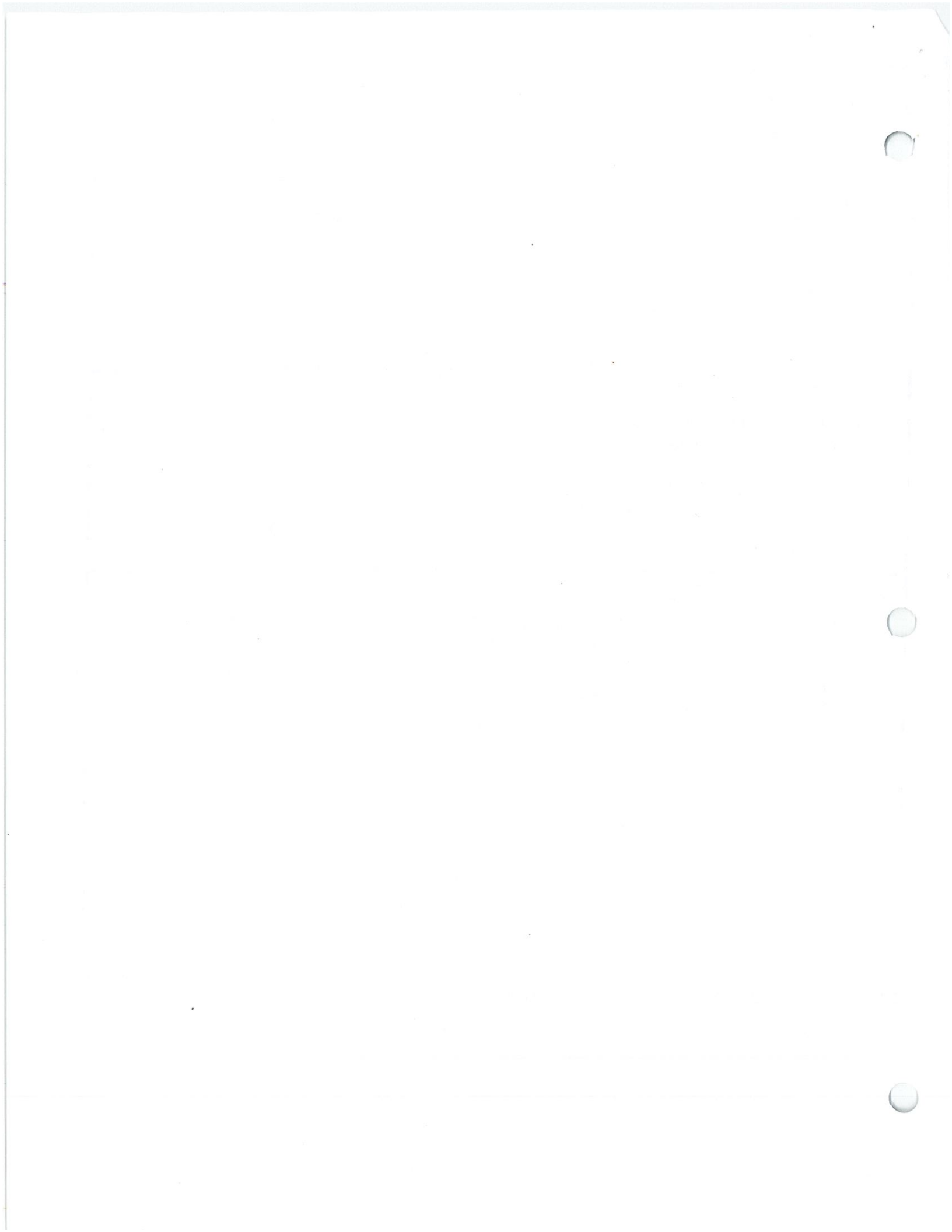
20 mol  $O_2$



How many moles of potassium nitrate are produced when two moles of potassium phosphate react with two moles of aluminum nitrate?

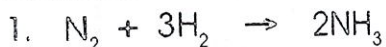


6 mol  $KNO_3$



# STOICHIOMETRY: GAS VOLUME PROBLEMS

Name \_\_\_\_\_



What volume of hydrogen is necessary to react with five liters of nitrogen to produce ammonia? (Assume constant temperature and pressure.)

$$\frac{5L N_2}{1 \text{ mol } N_2} \Bigg| \frac{3 \text{ mol } H_2}{1 \text{ mol } N_2} = 15L H_2$$

$$\frac{5L N_2}{2 \text{ mol } N_2} \Bigg| \frac{3 \text{ mol } H_2}{1 \text{ mol } N_2} = 15L H_2$$

2. What volume of ammonia is produced in the reaction in Problem 1?

$$\frac{5L N_2}{1 \text{ mol } N_2} \Bigg| \frac{2 \text{ mol } NH_3}{1 \text{ mol } N_2} =$$

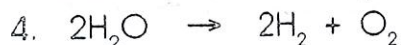
$$10L NH_3$$



If 20 liters of oxygen are consumed in the above reaction, how many liters of carbon dioxide are produced?

$$\frac{20L O_2}{5 \text{ mol } O_2} \Bigg| \frac{3 \text{ mol } CO_2}{5 \text{ mol } O_2} =$$

$$12L CO_2$$



If 30 mL of hydrogen are produced in the above reaction, how many milliliters of oxygen are produced?

$$\frac{30 \text{ mL } H_2}{2 \text{ mol } H_2} \Bigg| \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2} =$$

$$15 \text{ mL } O_2$$



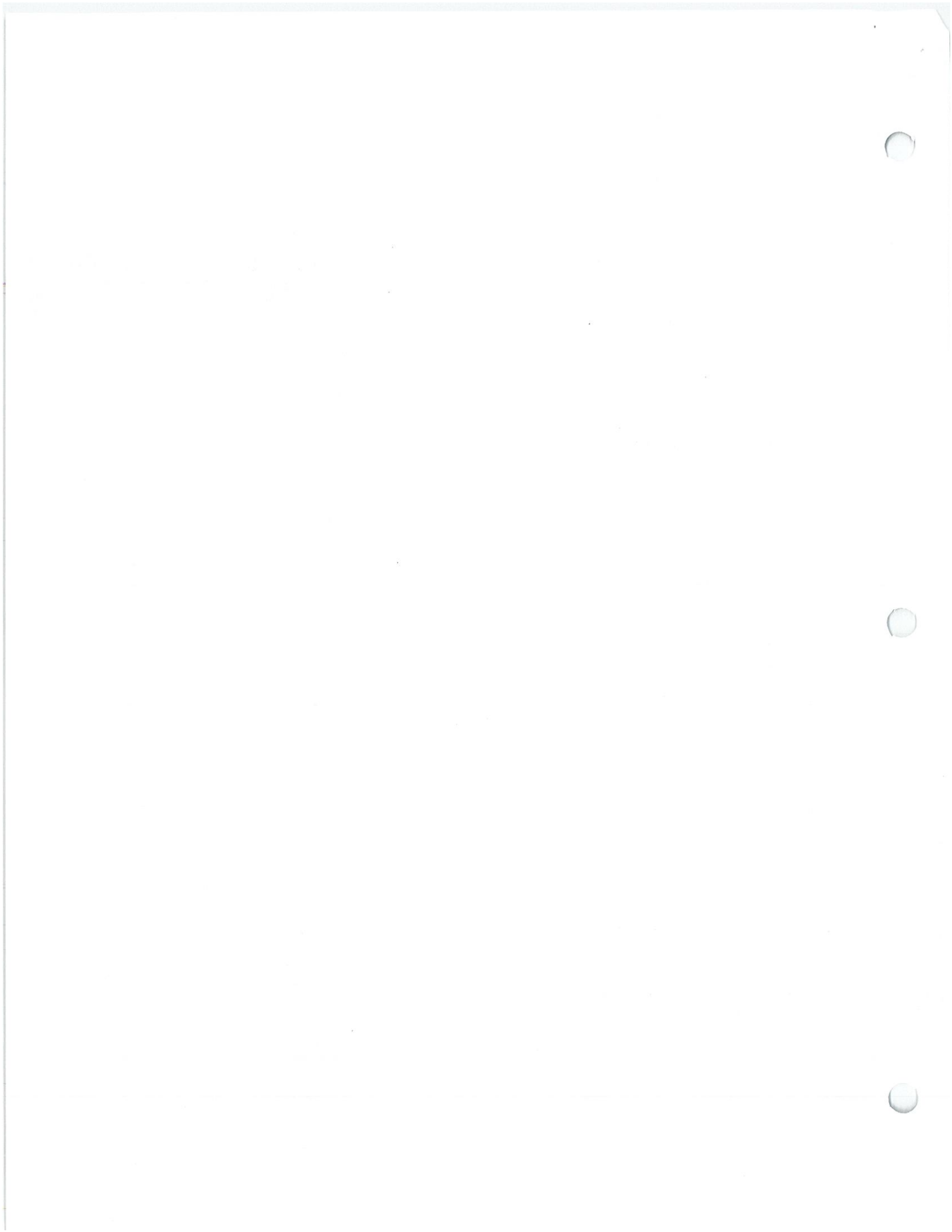
How many liters of carbon dioxide are produced if 75 liters of carbon monoxide are burned in oxygen? How many liters of oxygen are necessary?

$$\frac{75L CO}{2 \text{ mol } CO} \Bigg| \frac{2 \text{ mol } CO_2}{2 \text{ mol } CO} = 75L CO_2$$

$$\frac{75L CO}{2 \text{ mol } CO} \Bigg| \frac{1 \text{ mol } O_2}{2 \text{ mol } CO} = 38L O_2$$

$$75$$

$$\frac{75L CO_2}{38L O_2}$$





## Sec 3.9 Limiting Reactants



$$637.2\text{g} \qquad 1142\text{g} \qquad \times \text{g}$$

$$637.2\text{g NH}_3 \times \frac{1\text{mol NH}_3}{17.03\text{g}} = 37.42\text{ moles NH}_3$$

$$1142\text{g CO}_2 \times \frac{1\text{mol CO}_2}{44.01\text{g}} = 25.95\text{ moles CO}_2$$

$$25.95\text{mol CO}_2 \times \frac{2\text{mol NH}_3}{1\text{mol CO}_2} = 51.90\text{ mol NH}_3 \text{ needed to react with all of the CO}_2.$$

b) LR is  $\text{NH}_3$  because  $< 51.90$  moles is available

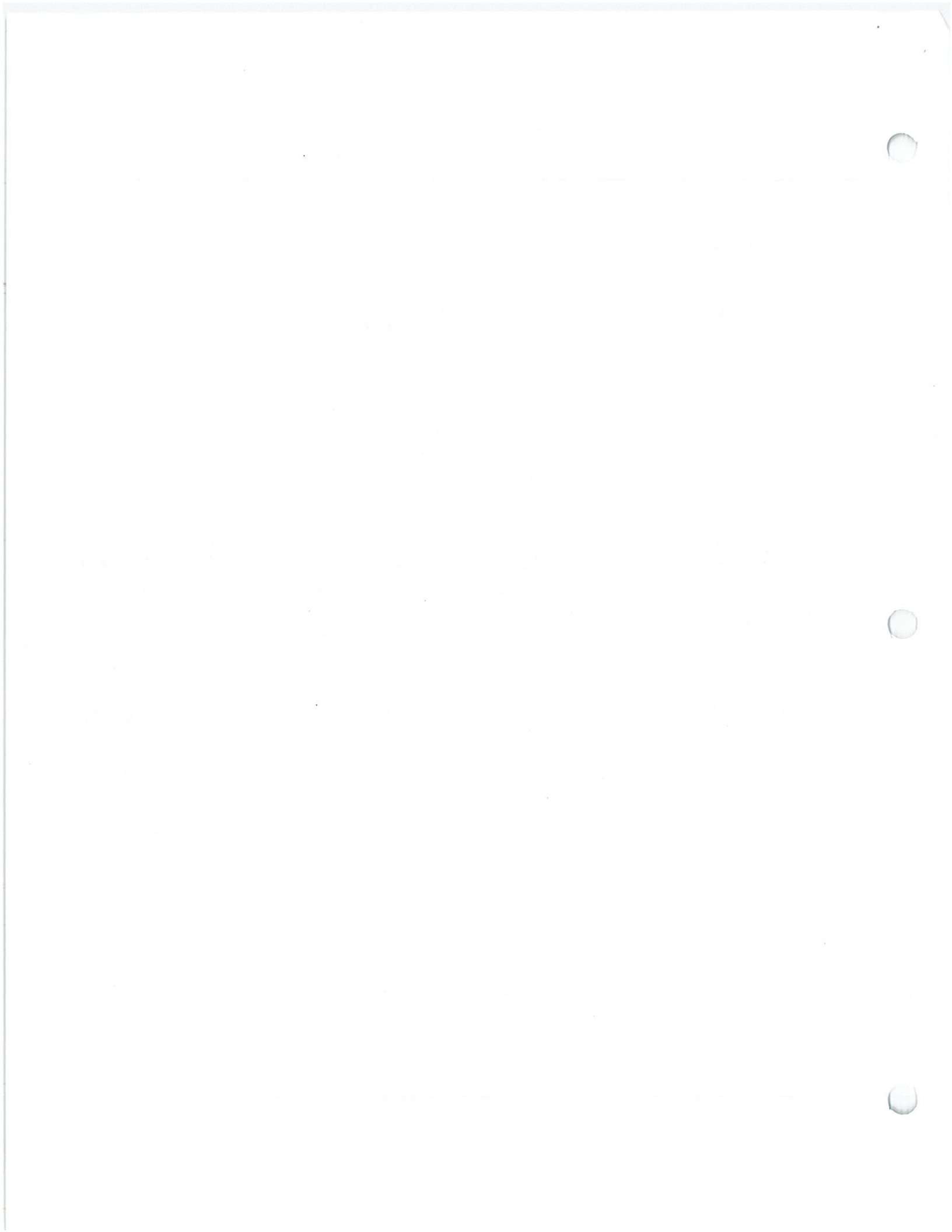
$$\text{a) } 37.42\text{ mol NH}_3 \times \frac{1\text{mol } (\text{NH}_2)_2\text{CO}}{2\text{mol NH}_3} \times \frac{60.06\text{g}}{1\text{mol } (\text{NH}_2)_2\text{CO}} = \boxed{1124\text{g } (\text{NH}_2)_2\text{CO}}$$

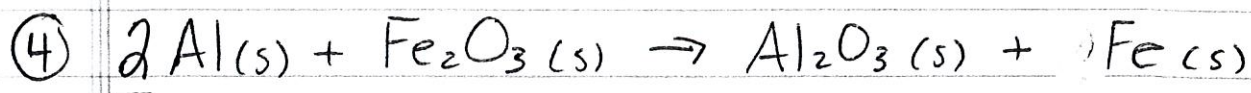
c) Excess reactant is  $\text{CO}_2$

$$37.42\text{ moles NH}_3 \times \frac{1\text{mole CO}_2}{2\text{mol NH}_3} = 18.71\text{ moles CO}_2 \text{ react}$$

$$25.95 - 18.71 = 7.24\text{ mol CO}_2 \times \frac{44.01\text{g CO}_2}{1\text{mole}}$$

$$= \boxed{318\text{g CO}_2 \text{ remain}}$$





$$124\text{g} \quad 601\text{g} \quad \quad \quad \times \text{g}$$

$$124\text{g Al} \times \frac{1\text{mol Al}}{26.98\text{g}} \times \frac{2\text{mol Fe}}{2\text{mol Al}} \times \frac{55.85\text{g Fe}}{1\text{mole Fe}} = 257\text{g Fe}$$

$$601\text{g Fe}_2\text{O}_3 \times \frac{1\text{mol Fe}_2\text{O}_3}{159.7\text{g}} \times \frac{2\text{mol Fe}}{1\text{mol Fe}_2\text{O}_3} \times \frac{55.85\text{g Fe}}{1\text{mole Fe}} = 420\text{g Fe}$$

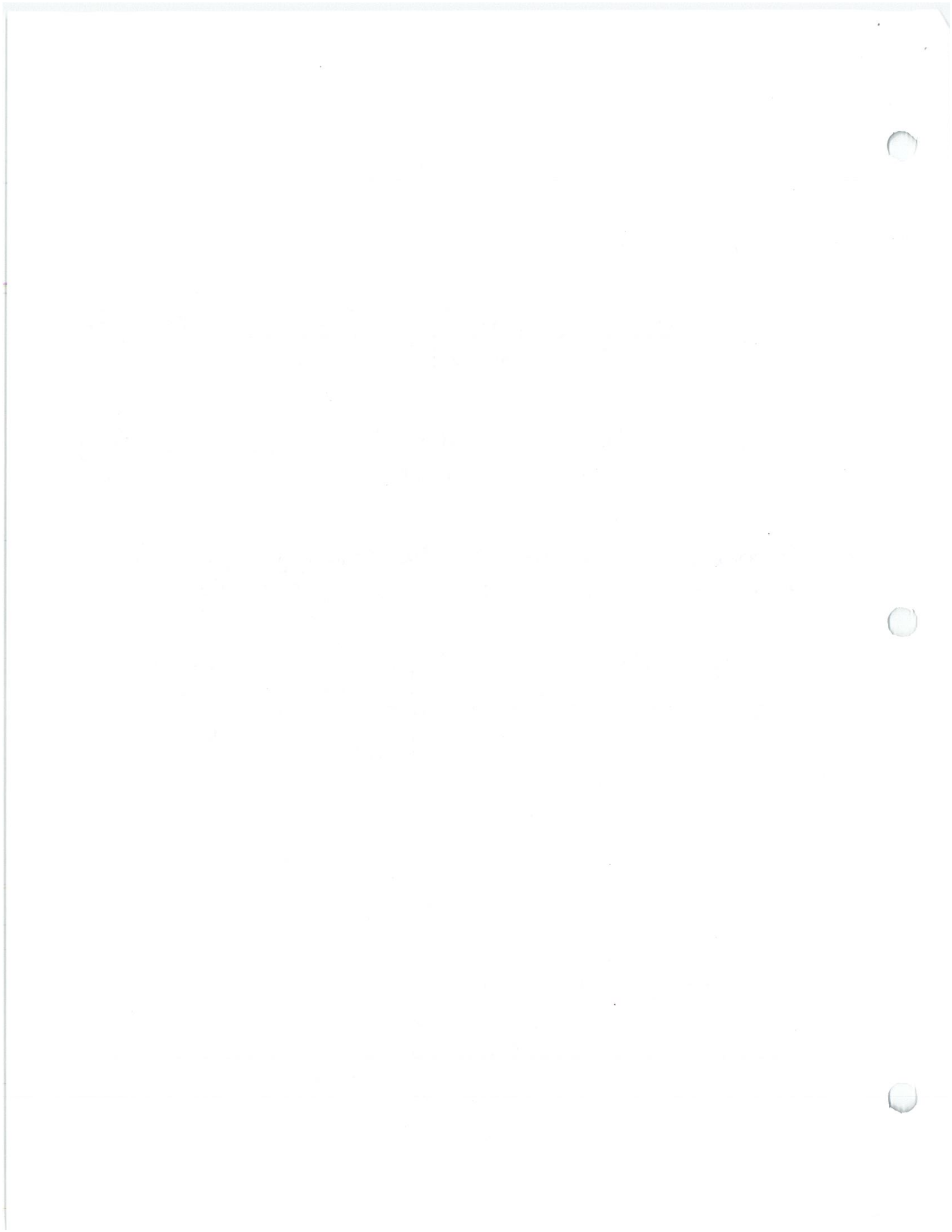
a) Smaller amount is the maximum amount that can be produced.  $\boxed{257\text{g Fe}}$

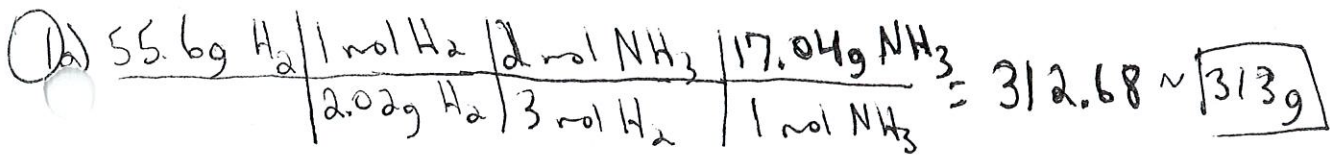
b) Limiting Reactant is Aluminum. When 257 g of iron have been produced, the Aluminum is used up and the reaction stops.

$$\text{c) } 124\text{g Al} \times \frac{1\text{mol Al}}{26.98\text{g}} \times \frac{1\text{mol Fe}_2\text{O}_3}{2\text{mol Al}} \times \frac{159.7\text{g Fe}_2\text{O}_3}{1\text{mol}}$$

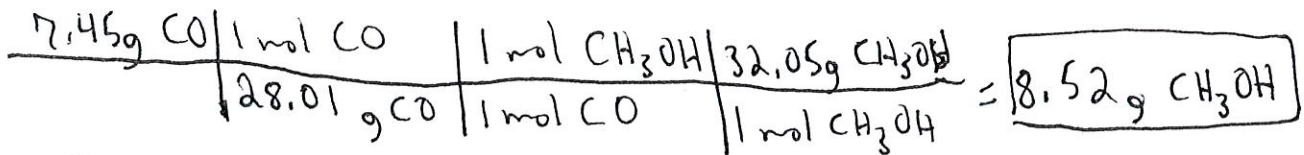
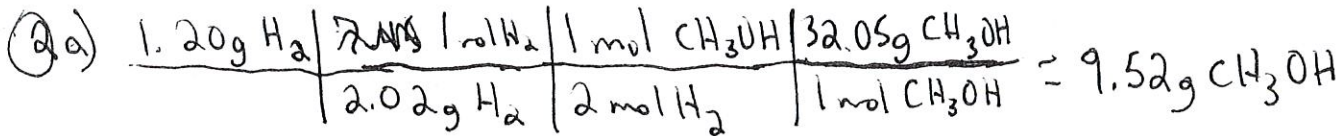
$$= 367\text{g Fe}_2\text{O}_3 \text{ reacts}$$

$$601 - 367 = \boxed{234\text{g of Fe}_2\text{O}_3 \text{ remain}}$$

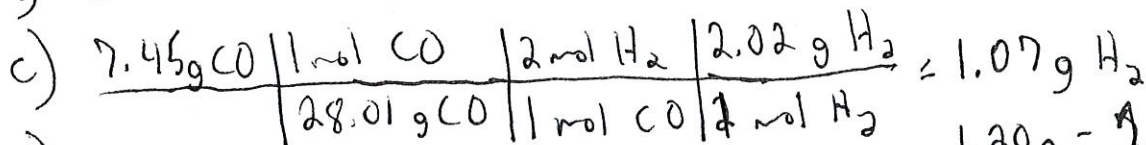




b) 
$$\frac{159 \text{ g}}{313 \text{ g}} \times 100 = \boxed{50.5\%}$$

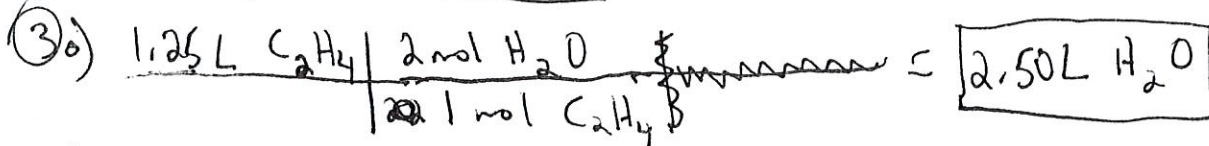


b) CO

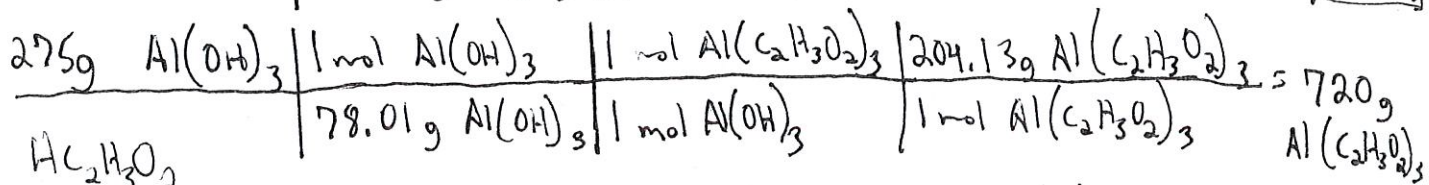
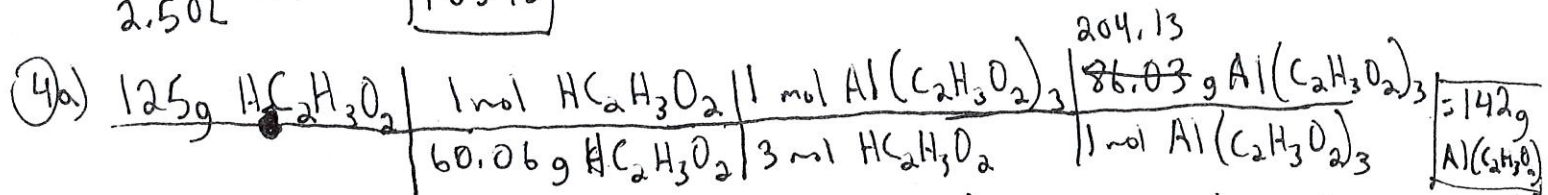


$$1.20 \text{ g} - 1.07 \text{ g} = \boxed{0.13 \text{ g H}_2}$$

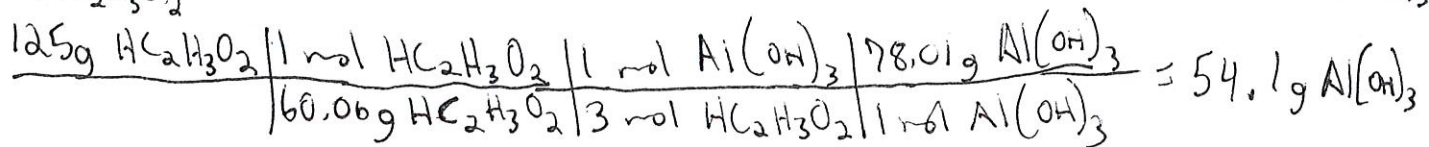
d) 
$$\frac{7.52 \text{ g H}_2}{8.52 \text{ g}} \times 100 = \boxed{88.3\%}$$



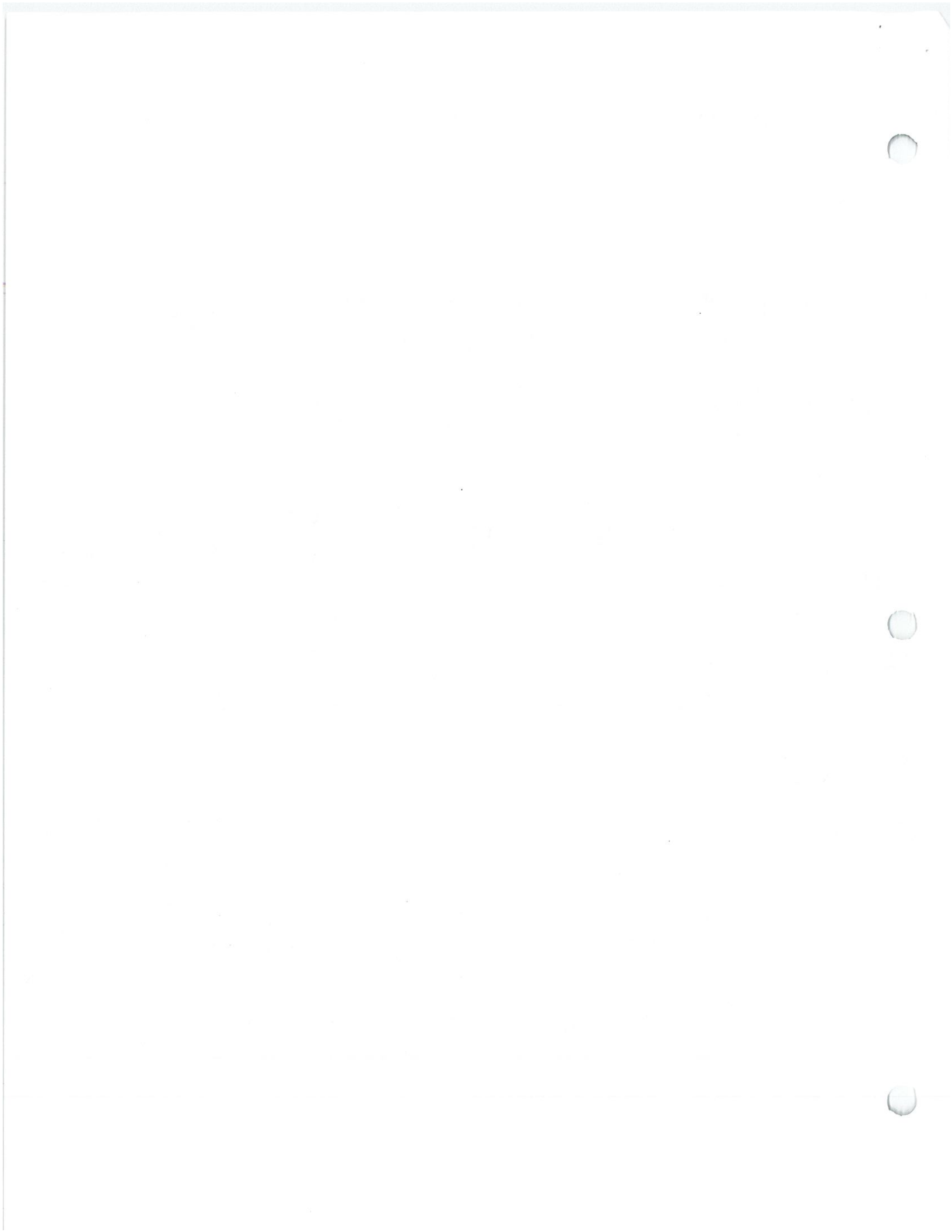
b) 
$$\frac{2.58 \text{ L}}{2.50 \text{ L}} \times 100 = \boxed{103\%}$$



b) HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub>



$$275 \text{ g} - 54.1 \text{ g} = \boxed{221 \text{ g Al}(\text{OH})_3}$$



3.10 % Yield



$$3.54 \times 10^7 \text{g} \quad 1.13 \times 10^7 \text{g} \quad \times \text{g}$$

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$$3.54 \times 10^7 \text{g TiCl}_4 \times \frac{1 \text{ mole TiCl}_4}{189.68 \text{g}} \times \frac{1 \text{ mol Ti}}{1 \text{ mol TiCl}_4} \times \frac{47.88 \text{g Ti}}{1 \text{ mole}}$$

$$= 8.94 \times 10^6 \text{g Ti possible}$$

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$$1.13 \times 10^7 \text{g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{g}} \times \frac{1 \text{ mol Ti}}{2 \text{ mol Mg}} \times \frac{47.88 \text{g Ti}}{1 \text{ mole}}$$

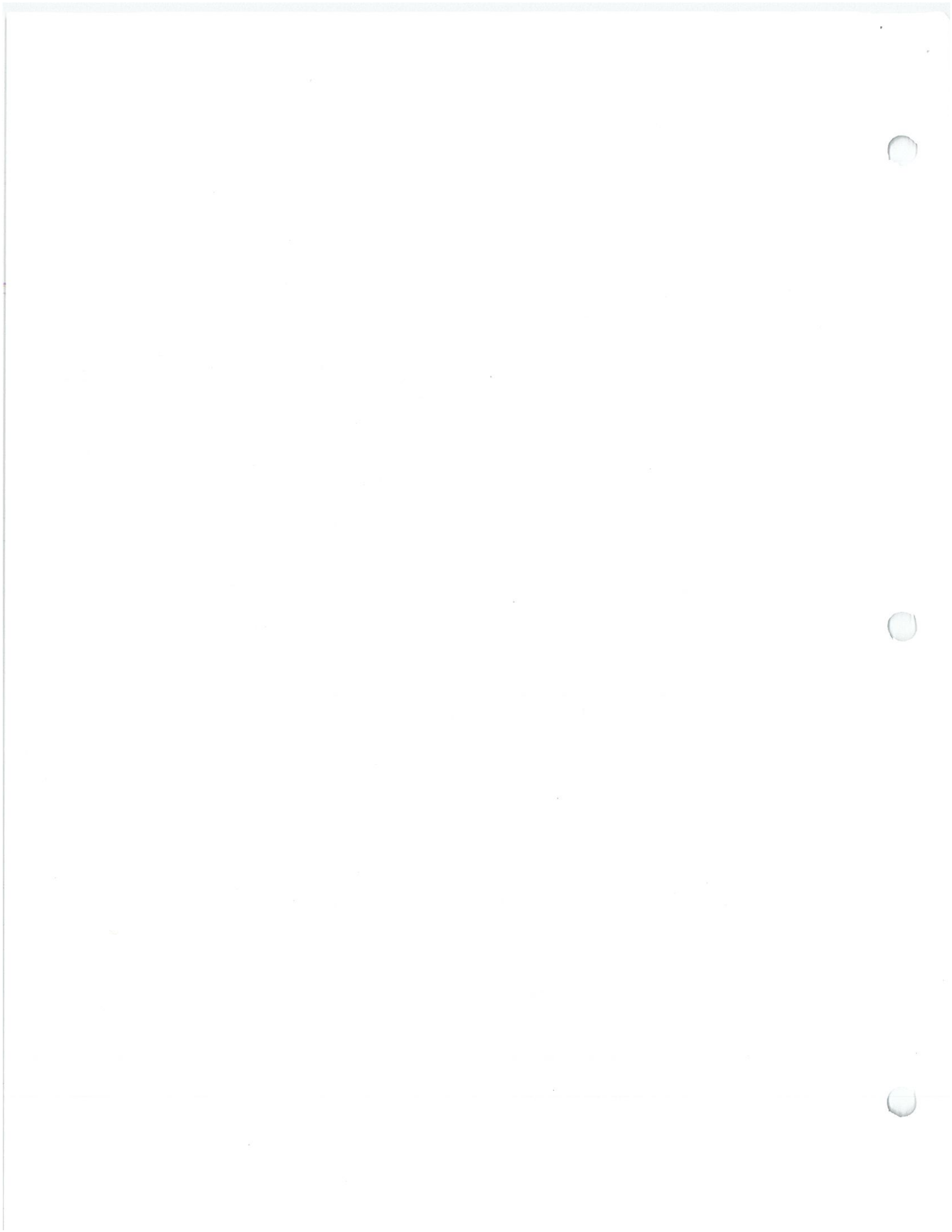
$$= 1.11 \times 10^7 \text{g Ti possible}$$

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a) Theoretical yield =  $8.94 \times 10^6 \text{g Ti}$  because when this is produced, the reaction will stop since all of the  $\text{TiCl}_4$  is used up.

$$\text{b. } \% \text{ Yield} = \frac{\text{Actual}}{\text{Theoretical}} \times 100 = \frac{7.91 \times 10^6 \text{g Ti}}{8.94 \times 10^6 \text{g possible}} \times 100$$

$$= \boxed{88.5\%}$$







$$1.96 \times 10^3 \text{g} \quad 1.54 \times 10^3 \text{g} \quad \text{x g}$$


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$$1.96 \times 10^3 \text{g Ca} \times \frac{1 \text{ mol Ca}}{40.08 \text{g}} \times \frac{2 \text{ mol V}}{5 \text{ mol Ca}} \times \frac{50.94 \text{g V}}{1 \text{ mole}}$$

$$= 9.96 \times 10^2 \text{g V possible}$$


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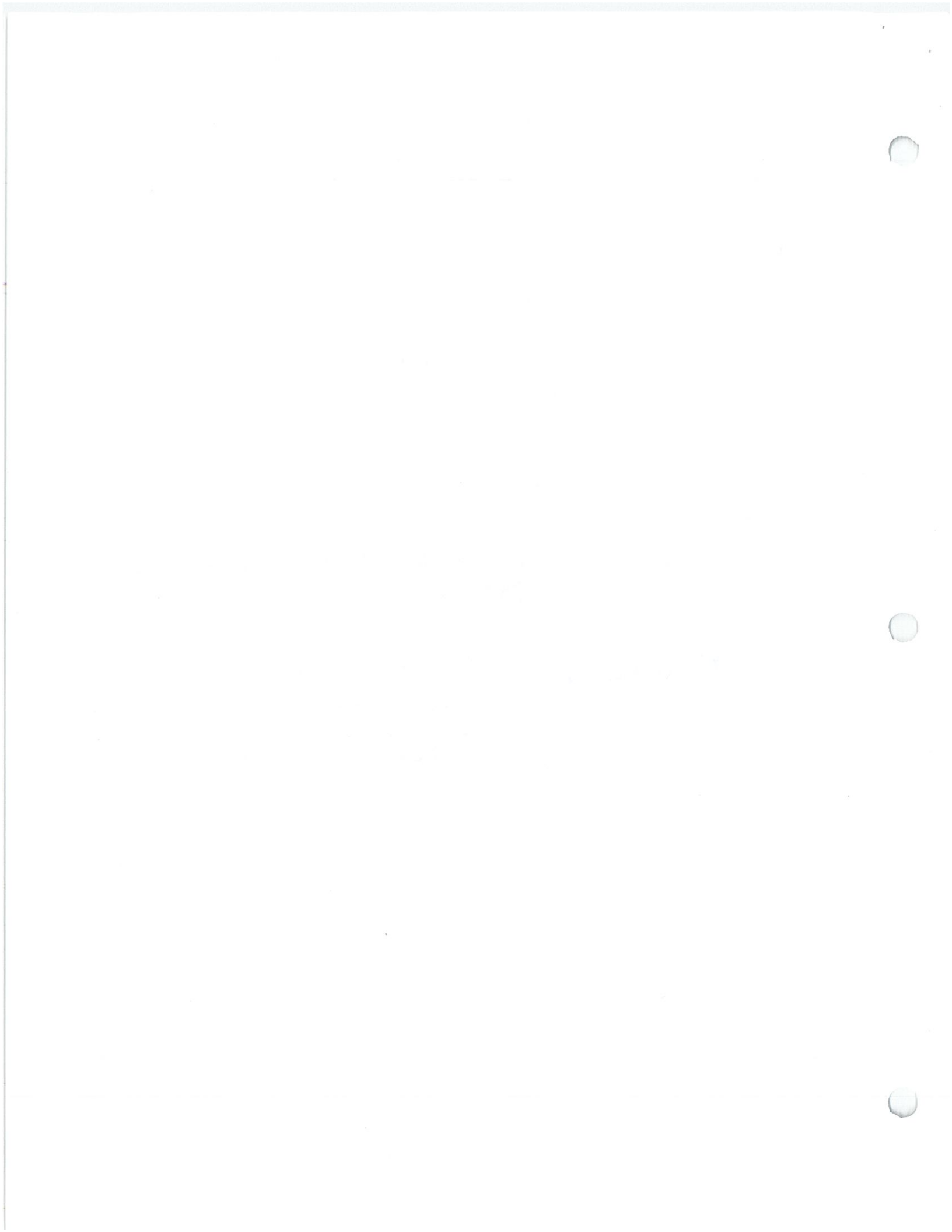
$$1.54 \times 10^3 \text{g V}_2\text{O}_5 \times \frac{1 \text{ mol V}_2\text{O}_5}{181.0 \text{g}} \times \frac{2 \text{ mol V}}{1 \text{ mol V}_2\text{O}_5} \times \frac{50.94 \text{g V}}{1 \text{ mol}}$$

$$= \textcircled{862} \text{g V possible}$$


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a) Theoretical yield =  $\frac{862}{509.96} \text{g V}$  because when this is produced, all of the  $\text{V}_2\text{O}_5$  is used up.

$$\text{b) \% Yield} = \frac{803 \text{g}}{862 \text{g}} \times 100 = \textcircled{93.2\%}$$



### Limiting Reactants

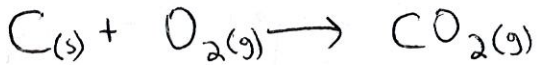
Limiting reactant -

Excess reactant -

Assume all gases are at STP.

1. Carbon reacts with oxygen to produce carbon dioxide.

a. Write the balanced equation showing physical states



b. How many grams of CO<sub>2</sub> are formed if 10.0 g of carbon are burned in

IF C is LR

$$\frac{10.0 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol C}} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2}}{20.0 \text{ L of O}_2?} = 36.6 \text{ g CO}_2$$

IF O<sub>2</sub> is LR

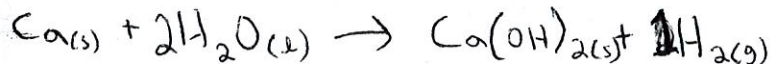
$$\frac{20.0 \text{ L O}_2 \times \frac{1 \text{ mol O}_2}{22.4 \text{ L O}_2} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol O}_2} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2}}{20.0 \text{ L of O}_2?} = 39.3 \text{ g}$$

c. What is the limiting reactant?

Carbon

2. Calcium reacts with water to produce solid calcium hydroxide, Ca(OH)<sub>2</sub>, and hydrogen gas.

a. Write the balanced equation showing physical states



b. What volume of hydrogen is formed when 15.0 g of calcium reacts with 10.0 g of water?

IF Ca is LR

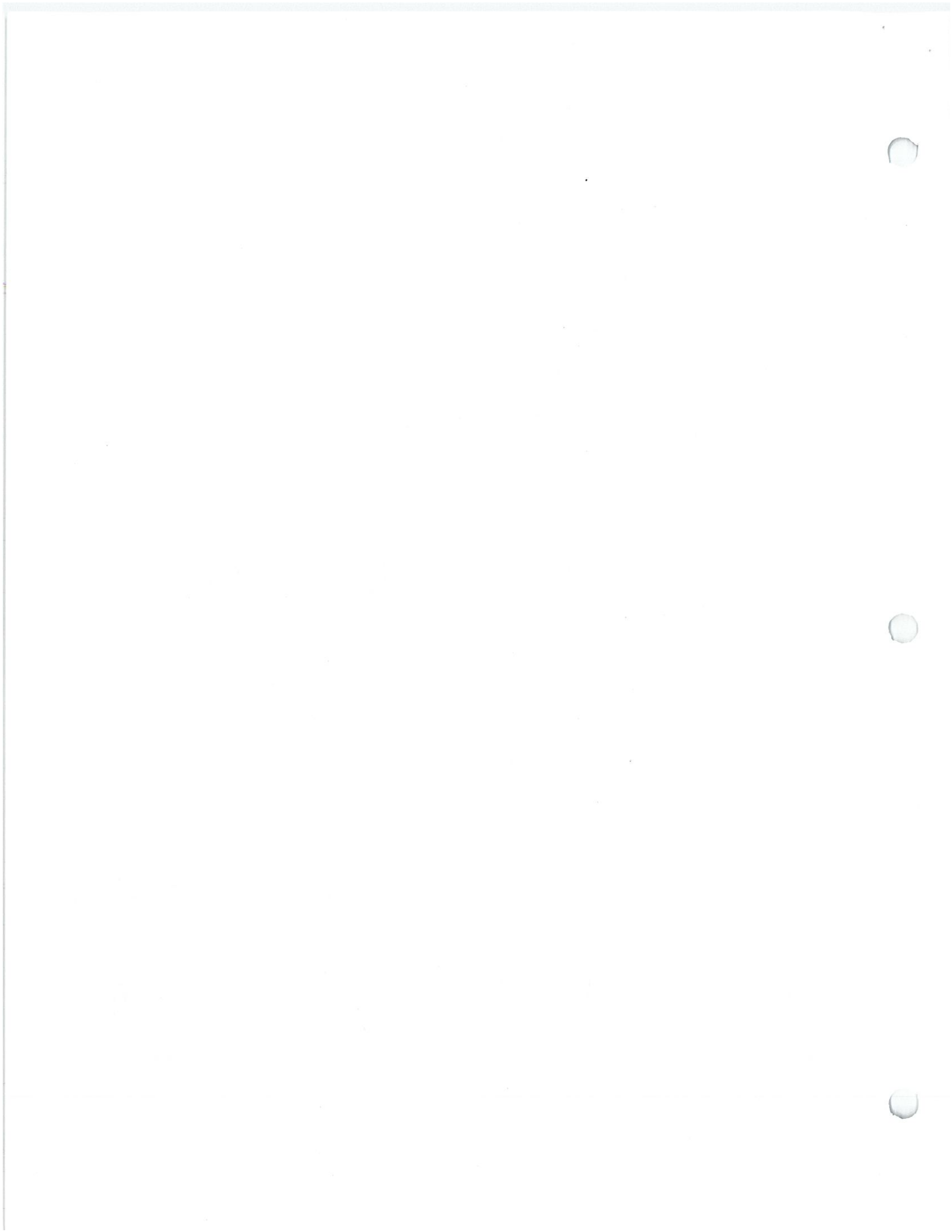
$$\frac{15.0 \text{ g Ca} \times \frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} \times \frac{1 \text{ mol H}_2}{1 \text{ mol Ca}} \times \frac{22.4 \text{ L H}_2}{1 \text{ mol H}_2}}{10.0 \text{ g of water?}} = 8.38 \text{ L H}_2$$

IF H<sub>2</sub>O is LR

$$\frac{10.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{1 \text{ mol H}_2}{2 \text{ mol H}_2\text{O}} \times \frac{22.4 \text{ L H}_2}{1 \text{ mol H}_2}}{10.0 \text{ g of water?}} = 6.22 \text{ L H}_2$$

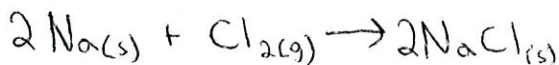
c. What is the limiting reactant?

Water



3. Sodium metal reacts with chlorine to produce crystalline sodium chloride, NaCl.

a. Write the balanced equation showing physical states



b. How much sodium chloride is produced when 2.50g of sodium reacts with 2.5 L of chlorine gas?

if Na is LR

$$\frac{2.50 \text{ g Na} \left| \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} \right| \frac{2 \text{ mol NaCl}}{2 \text{ mol Na}}}{}$$

$$\frac{0.109 \text{ mol}}{\cancel{0.22 \text{ mol NaCl}}}$$

if Cl<sub>2</sub> is LR

$$\frac{2.5 \text{ L Cl}_2 \left| \frac{1 \text{ mol Cl}_2}{22.4 \text{ L Cl}_2} \right| \frac{2 \text{ mol NaCl}}{1 \text{ mol Cl}_2}}{}$$

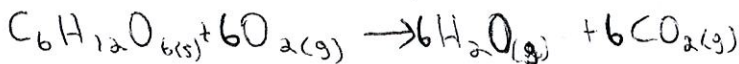
$$\frac{\cancel{0.22 \text{ mol NaCl}}}{0.22 \text{ mol NaCl}}$$

c. What is the limiting reactant?

~~Chlorine gas~~ Sodium

4. Glucose, C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>(s), undergoes combustion to form the same products as a hydrocarbon.

a. Write the balanced equation showing physical states.



b. What mass of carbon dioxide is formed when 85.0 g of glucose combines with 75.0 L of oxygen?

if glucose is LR

$$\frac{85.0 \text{ g C}_6\text{H}_{12}\text{O}_6 \left| \frac{1 \text{ mol C}_6\text{H}_{12}\text{O}_6}{180.2 \text{ g C}_6\text{H}_{12}\text{O}_6} \right| \frac{6 \text{ mol CO}_2}{1 \text{ mol C}_6\text{H}_{12}\text{O}_6} \left| \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \right|}{}$$

if O<sub>2</sub> is LR

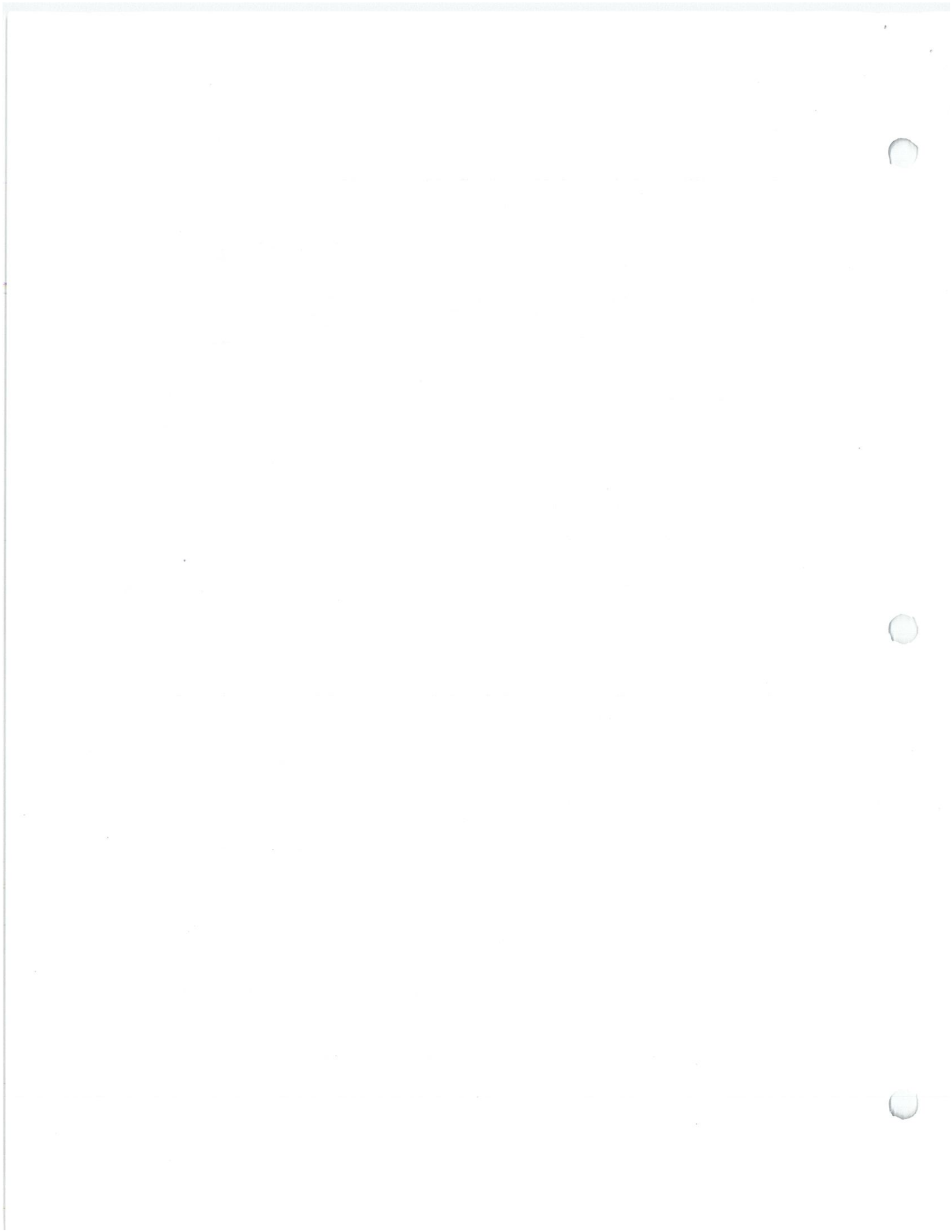
$$\frac{75.0 \text{ L O}_2 \left| \frac{1 \text{ mol O}_2}{22.4 \text{ L O}_2} \right| \frac{6 \text{ mol CO}_2}{1 \text{ mol O}_2} \left| \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \right|}{}$$

c. What volume does the carbon dioxide occupy?

$$\frac{125 \text{ g CO}_2 \left| \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \right| \frac{22.4 \text{ L CO}_2}{1 \text{ mol CO}_2}}{}$$

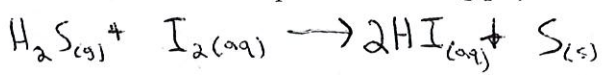
d. What is the limiting reactant?

glucose



5. Hydrogen sulfide reacts with an aqueous solution of iodine to produce hydroiodic acid, HI (aq) and sulfur.

a. Write the balanced equation showing physical states



b. What mass of sulfur is produced by 4.11 g of iodine and 317 mL of hydrogen sulfide?

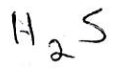
if I<sub>2</sub> is LR

$$\frac{4.11 \text{ g } I_2}{253.8 \text{ g } I_2} \times \frac{1 \text{ mol } I_2}{1 \text{ mol } I_2} \times \frac{1 \text{ mol } S}{1 \text{ mol } I_2} \times \frac{32.07 \text{ g } S}{1 \text{ mol } S} = 0.519 \text{ g } S$$

if H<sub>2</sub>S is LR

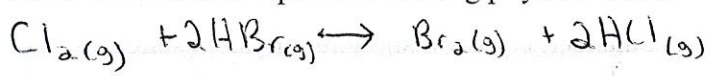
$$\frac{317 \text{ mL } H_2S}{1000 \text{ mL } H_2S} \times \frac{1 \text{ L } H_2S}{22.4 \text{ L } H_2S} \times \frac{1 \text{ mol } H_2S}{1 \text{ mol } H_2S} \times \frac{1 \text{ mol } S}{1 \text{ mol } H_2S} \times \frac{32.07 \text{ g } S}{1 \text{ mol } S} = \boxed{0.454 \text{ g } S}$$

c. What is the limiting reactant?



6. Chlorine combines with hydrogen bromide to produce bromine vapor and hydrogen chloride.

a. Write the balanced equation showing physical states



b. What volume of bromine gas is produced when 212 mL and chlorine reacts with 155 mL of hydrogen bromide?

if Cl<sub>2</sub> is LR

$$\frac{212 \text{ mL } Cl_2}{1000 \text{ mL } Cl_2} \times \frac{1 \text{ L } Cl_2}{22.4 \text{ L } Cl_2} \times \frac{1 \text{ mol } Cl_2}{1 \text{ mol } Cl_2} \times \frac{1 \text{ mol } Br_2}{1 \text{ mol } Cl_2} \times \frac{22.4 \text{ L } Br_2}{1 \text{ mol } Br_2} = 212 \text{ mL } Br_2$$

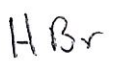
if HBr is LR

$$\frac{155 \text{ mL } HBr}{1000 \text{ mL } HBr} \times \frac{1 \text{ L } HBr}{22.4 \text{ L } HBr} \times \frac{1 \text{ mol } HBr}{2 \text{ mol } HBr} \times \frac{1 \text{ mol } Br_2}{1 \text{ mol } HBr} \times \frac{22.4 \text{ L } Br_2}{1 \text{ mol } Br_2} = \boxed{0.0775 \text{ L } Br_2}$$

c. What is the mass of this volume of bromine?

$$\frac{0.0775 \text{ L } Br_2}{22.4 \text{ L } Br_2} \times \frac{1 \text{ mol } Br_2}{1 \text{ mol } Br_2} \times \frac{159.8 \text{ g } Br_2}{1 \text{ mol } Br_2} = \boxed{0.553 \text{ g } Br_2}$$

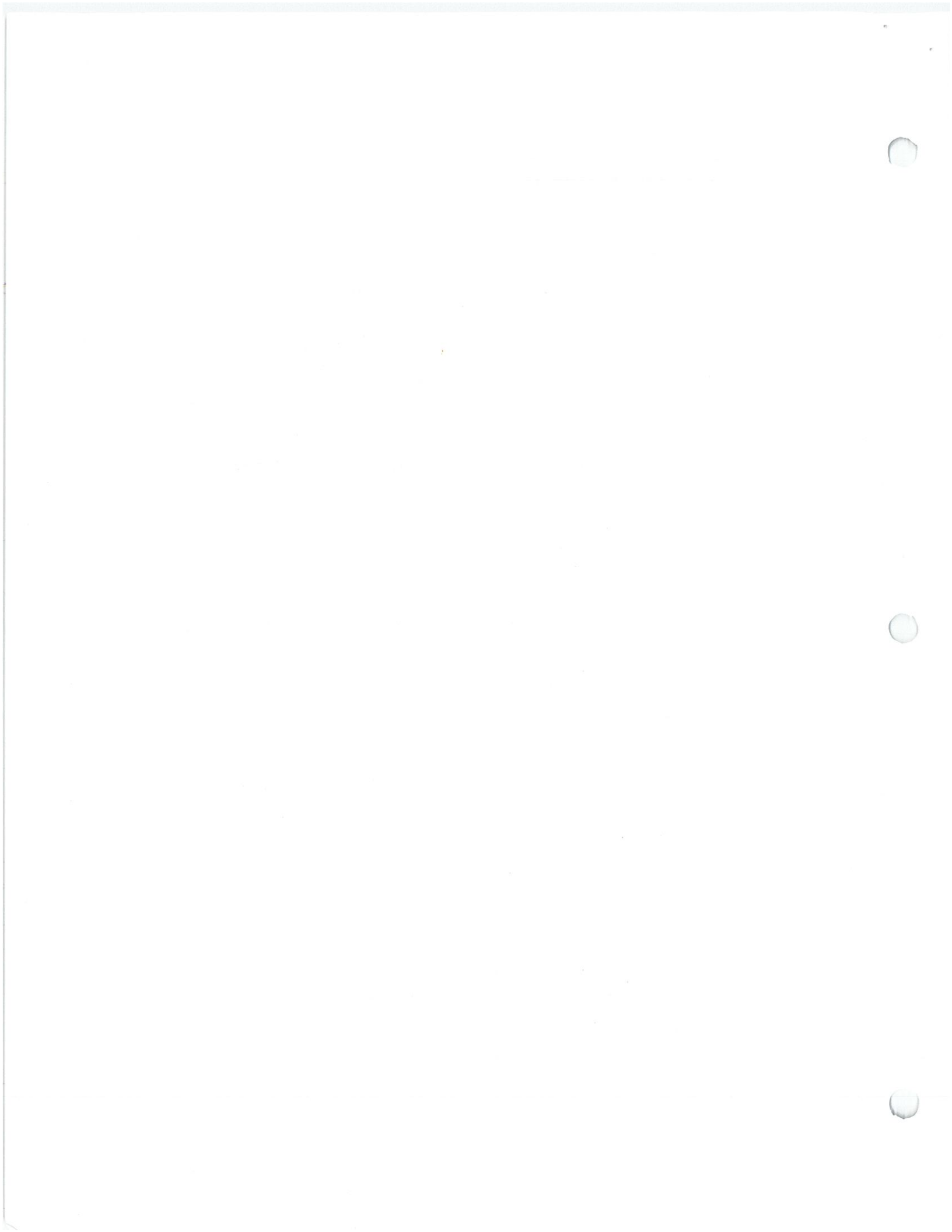
d. What is the limiting reactant?



e. What volume of the excess reactant remains?

$$\frac{155 \text{ mL } HBr}{1000 \text{ mL } HBr} \times \frac{1 \text{ L } HBr}{22.4 \text{ L } HBr} \times \frac{1 \text{ mol } HBr}{2 \text{ mol } HBr} \times \frac{1 \text{ mol } Cl_2}{1 \text{ mol } HBr} \times \frac{22.4 \text{ L } Cl_2}{1 \text{ mol } Cl_2} = 77.5 \text{ mL } Cl_2$$

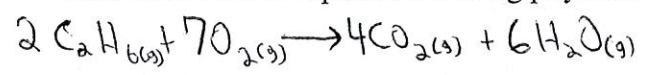
$$212 \text{ mL} - 77.5 \text{ mL} = 134.5 \text{ mL} \approx \boxed{135 \text{ mL } Cl_2}$$





7. 725 mL of ethane gas, C<sub>2</sub>H<sub>6</sub>, undergoes combustion in the presence of 3.5 L of oxygen.

a. Write the balanced equation showing physical states



b. What volume of carbon dioxide is formed if the reaction goes to completion?

if C<sub>2</sub>H<sub>6</sub> is LR

$$\frac{725 \text{ L C}_2\text{H}_6 \left| \frac{1 \text{ mol C}_2\text{H}_6}{22.4 \text{ L C}_2\text{H}_6} \right| \frac{4 \text{ mol CO}_2}{2 \text{ mol C}_2\text{H}_6} \left| \frac{22.4 \text{ L CO}_2}{1 \text{ mol CO}_2} \right|}{1} = 1.45 \text{ L CO}_2$$

if ~~C<sub>2</sub>H<sub>6</sub>~~ O<sub>2</sub> is LR

$$\frac{3.5 \text{ L O}_2 \left| \frac{1 \text{ mol O}_2}{22.4 \text{ L O}_2} \right| \frac{4 \text{ mol CO}_2}{7 \text{ mol O}_2} \left| \frac{22.4 \text{ L CO}_2}{1 \text{ mol CO}_2} \right|}{1} = 2.0 \text{ L CO}_2$$

c. What is the limiting reactant?

ethane

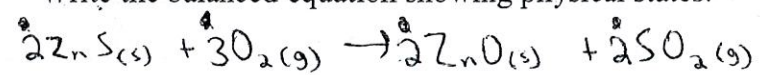
d. What volume of the excess reactant remains?

$$\frac{725 \text{ L C}_2\text{H}_6 \left| \frac{1 \text{ mol C}_2\text{H}_6}{22.4 \text{ L C}_2\text{H}_6} \right| \frac{7 \text{ mol O}_2}{2 \text{ mol C}_2\text{H}_6} \left| \frac{22.4 \text{ L O}_2}{1 \text{ mol O}_2} \right|}{1} = 2.54 \text{ L O}_2$$

$$3.5 \text{ L} - 2.54 \text{ L} = 0.96 \sim 1.0 \text{ L O}_2 \text{ remains}$$

8. Zinc sulfide, ZnS(s) reacts with oxygen to produce zinc oxide, ZnO(s) and sulfur dioxide.

a. Write the balanced equation showing physical states.



b. What mass of zinc oxide could be produced from 418 g of zinc sulfide and 185 L of oxygen?

if ZnS is LR

$$\frac{418 \text{ g ZnS} \left| \frac{1 \text{ mol ZnS}}{97.46 \text{ g ZnS}} \right| \frac{2 \text{ mol ZnO}}{2 \text{ mol ZnS}} \left| \frac{81.39 \text{ g ZnO}}{1 \text{ mol ZnO}} \right|}{1} = 349 \text{ g ZnO}$$

if O<sub>2</sub> is LR

$$\frac{185 \text{ L O}_2 \left| \frac{1 \text{ mol O}_2}{22.4 \text{ L O}_2} \right| \frac{2 \text{ mol ZnO}}{3 \text{ mol O}_2} \left| \frac{81.39 \text{ g ZnO}}{1 \text{ mol ZnO}} \right|}{1} = 448 \text{ g ZnO}$$

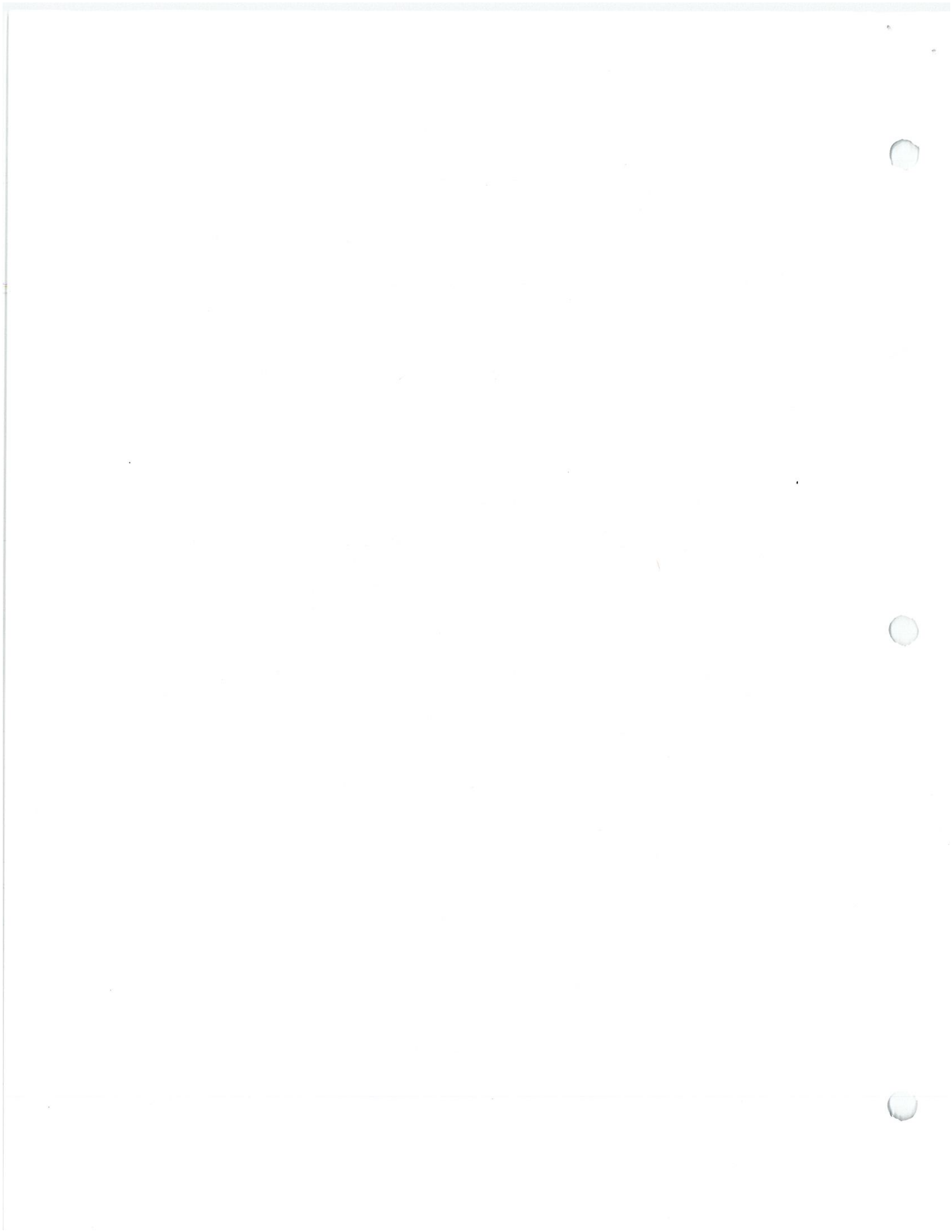
c. What is the limiting reactant?

ZnS

d. How much of the excess reactant remain?

$$\frac{418 \text{ g ZnS} \left| \frac{1 \text{ mol ZnS}}{97.46 \text{ g ZnS}} \right| \frac{3 \text{ mol O}_2}{2 \text{ mol ZnS}} \left| \frac{22.4 \text{ L O}_2}{1 \text{ mol O}_2} \right|}{1} = 144 \text{ L O}_2$$

$$185 \text{ L} - 144 \text{ L} = 41 \text{ L O}_2$$



9. Use the same equation in #8 above to answer the questions below.

a. What volume, at STP, of sulfur dioxide would be produced from 4.66g of zinc sulfide and 1250 mL of oxygen?

if ZnS is LR

$$\frac{4.66 \text{ g ZnS} \left| \frac{1 \text{ mol ZnS}}{97.46 \text{ g ZnS}} \right| \frac{2 \text{ mol SO}_2}{2 \text{ mol ZnS}} \left| \frac{22.4 \text{ L SO}_2}{1 \text{ mol SO}_2} \right.}{= 1.07 \text{ L SO}_2$$

if O<sub>2</sub> is LR

$$\frac{1.25 \text{ L O}_2 \left| \frac{1 \text{ mol O}_2}{22.4 \text{ L O}_2} \right| \frac{2 \text{ mol SO}_2}{3 \text{ mol O}_2} \left| \frac{22.4 \text{ L SO}_2}{1 \text{ mol SO}_2} \right.}{= \boxed{0.833 \text{ L SO}_2}$$

b. What is the limiting reactant?

Oxygen

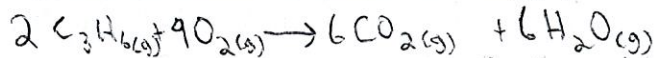
c. How much excess reactant remains?

$$\frac{1.25 \text{ L O}_2 \left| \frac{1 \text{ mol O}_2}{22.4 \text{ L O}_2} \right| \frac{2 \text{ mol ZnS}}{3 \text{ mol O}_2} \left| \frac{97.46 \text{ g ZnS}}{1 \text{ mol ZnS}} \right.}{= 3.63 \text{ g ZnS}}$$

$$4.66 \text{ g} - 3.63 \text{ g} = \boxed{1.03 \text{ g ZnS}}$$

10. 8.4 grams of cyclopropane gas, C<sub>3</sub>H<sub>6</sub>, undergoes combustion. If 5.5 L of oxygen is available for this reaction:

a. Write the balanced equation showing physical states.



b. What is the limiting reactant?

Oxygen

c. What is the excess reactant?

C<sub>3</sub>H<sub>6</sub>

d. What mass of carbon dioxide is formed?

if C<sub>3</sub>H<sub>6</sub> is LR

$$\frac{8.4 \text{ g C}_3\text{H}_6 \left| \frac{1 \text{ mol C}_3\text{H}_6}{42.08 \text{ g C}_3\text{H}_6} \right| \frac{6 \text{ mol CO}_2}{2 \text{ mol C}_3\text{H}_6} \left| \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \right.}{= 26 \text{ g CO}_2$$

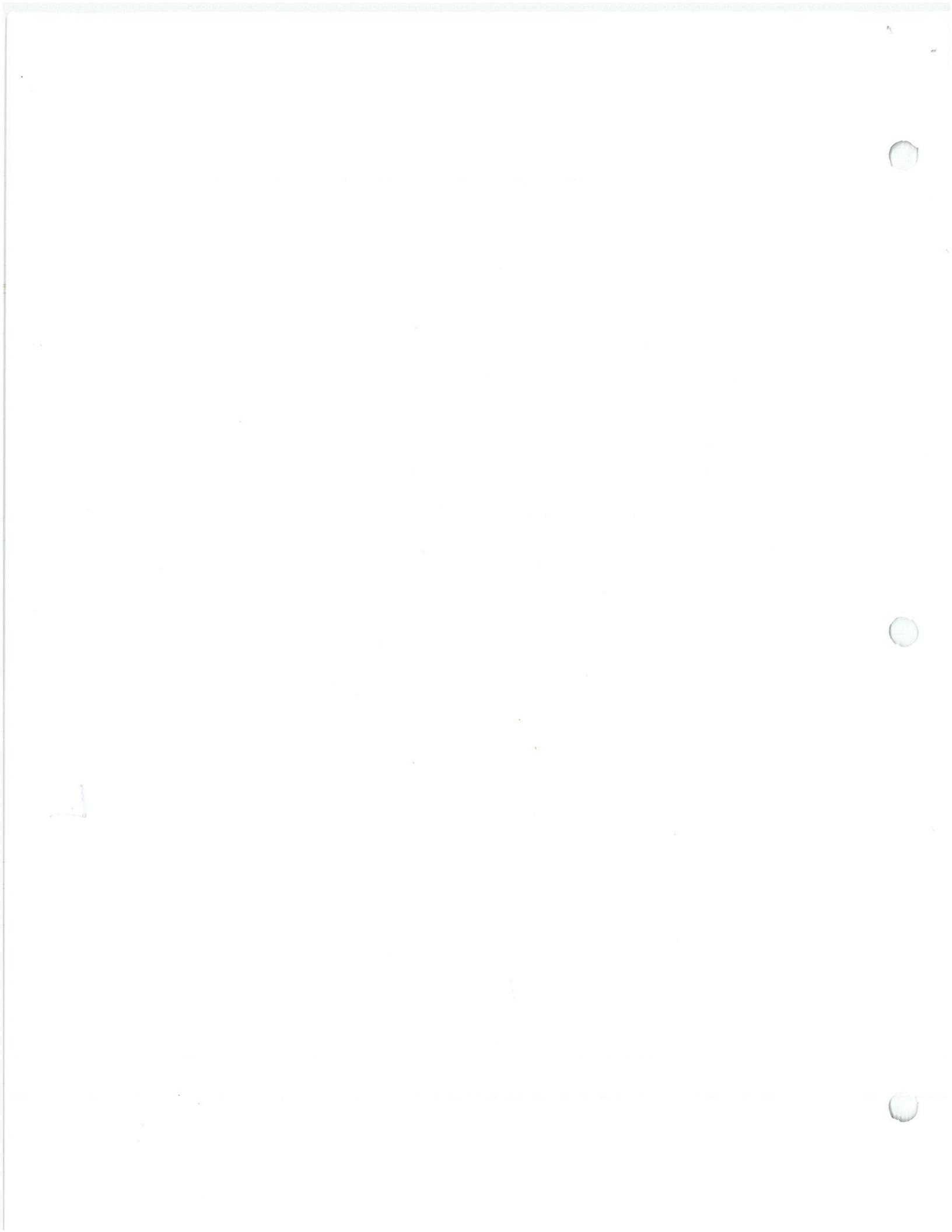
if O<sub>2</sub> is LR

$$\frac{5.5 \text{ L O}_2 \left| \frac{1 \text{ mol O}_2}{22.4 \text{ L O}_2} \right| \frac{6 \text{ mol CO}_2}{9 \text{ mol O}_2} \left| \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \right.}{= \boxed{17.2 \text{ g CO}_2}$$

d. How much of the excess reactant remains?

$$\frac{5.5 \text{ L O}_2 \left| \frac{1 \text{ mol O}_2}{22.4 \text{ L O}_2} \right| \frac{2 \text{ mol C}_3\text{H}_6}{9 \text{ mol O}_2} \left| \frac{42.08 \text{ g C}_3\text{H}_6}{1 \text{ mol C}_3\text{H}_6} \right.}{= 2.3 \text{ g C}_3\text{H}_6}$$

$$8.4 - 2.3 = \boxed{6.1 \text{ g C}_3\text{H}_6}$$



## Percent Yield

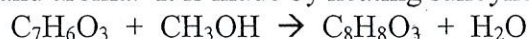
The percentage yield of a product is the actual amount of product expressed as a percentage of the calculated theoretical yield of that product.

Actual yield – the amount actually produced by the reaction

Theoretical yield – the calculated amount from the balanced equation. Assumes that the reaction goes to completion, that all of the reactants react, that all of the product is recovered and that there are no side reactions.

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

- Oil of wintergreen (methyl salicylate) is used in a variety of commercial products for its flavor and aroma. It is made by heating salicylic acid with methanol.



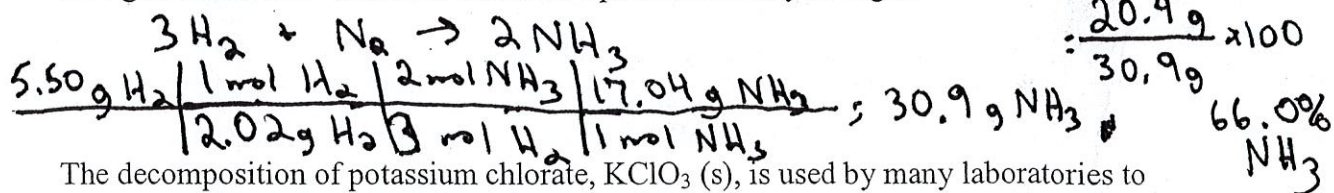
A chemist starts with 1.75g of salicylic acid and excess methanol and calculates the maximum possible yield to be 1.93 g. However, after the reaction is run, the chemist finds that the amount of methyl salicylate produced and isolated is only 1.42 g. What is the percentage yield of the process?

$$\frac{1.42\text{g}}{1.93\text{g}} \times 100 = 73.6\% \text{ methyl salicylate}$$

- The actual amount of product in a reaction is 39.7 g although a mass-mass calculation predicted 65.6 g. What is the percent yield of this product?

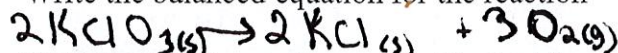
$$\frac{39.7\text{g}}{65.6\text{g}} \times 100 = 60.5\%$$

- What is the percentage yield if 5.50 g of hydrogen react with excess nitrogen to form 20.4 g of ammonia? Write the balanced equation before you begin.



- The decomposition of potassium chlorate, KClO<sub>3</sub> (s), is used by many laboratories to produce oxygen. The other product is potassium chloride, KCl(s).

a. Write the balanced equation for the reaction



b. If 75.6 g KClO<sub>3</sub> are decomposed in the lab producing 21 L O<sub>2</sub> what is the percent yield?

$$\frac{21}{20.7} = 101\%$$

75.6 g KClO <sub>3</sub>	1 mol KClO <sub>3</sub>	3 mol O <sub>2</sub>	22.4 L O <sub>2</sub>	= 20.7 L O <sub>2</sub>
122.55 g KClO <sub>3</sub>	2 mol KClO <sub>3</sub>	1 mol O <sub>2</sub>		

c. Using the percent yield you calculated above, how much KClO<sub>3</sub> must be decomposed to produce 54 L of oxygen?

$$\frac{101}{100} = \frac{54\text{L}}{x}$$

53.5 L O <sub>2</sub>	1 mol O <sub>2</sub>	2 mol KClO <sub>3</sub>	122.55 KClO <sub>3</sub>	= 195 g KClO <sub>3</sub>
22.4 L O <sub>2</sub>	3 mol O <sub>2</sub>	1 mol KClO <sub>3</sub>		

