1. Classify the type of each of the following reactions:

a) \((\text{NH}_4)_2\text{SO}_4 \rightarrow 2 \text{NH}_3 + \text{SO}_2 + \text{H}_2\) __decomposition__

b) \(\text{Br}_2 + 2 \text{KI} \rightarrow 2 \text{KBr} + \text{I}_2\) __single replacement__

c) \(2 \text{Na} + \text{Cl}_2 \rightarrow 2 \text{NaCl}\) __synthesis__

d) \(\text{Al} (\text{OH})_3 + 3 \text{HCl} \rightarrow \text{AlCl}_3 + 3 \text{H}_2\text{O}\) __double replacement__

2. Balance each of the equations shown below:

a) \(\textcolor{red}{5} \text{C} + \textcolor{red}{2} \text{SO}_2 \rightarrow \textcolor{red}{1} \text{CS}_2 + \textcolor{red}{4} \text{CO}\)

b) \(\textcolor{red}{2} \text{Na}_3\text{N} \rightarrow \textcolor{red}{6} \text{Na} + \textcolor{red}{1} \text{N}_2\)

c) \(\textcolor{red}{1} \text{C}_3\text{H}_8 + \textcolor{red}{5} \text{O}_2 \rightarrow \textcolor{red}{3} \text{CO}_2 + \textcolor{red}{4} \text{H}_2\text{O}\)

d) \(\textcolor{red}{2} \text{Al} + \textcolor{red}{3} \text{H}_2\text{SO}_4 \rightarrow \textcolor{red}{1} \text{Al}_2(\text{SO}_4)_3 + \textcolor{red}{3} \text{H}_2\)

3. Predict the products for each reaction and balance the equation. If no reaction occurs, write “No Reaction” after the arrow:

a) \(\textcolor{red}{2} \text{Al} (\text{s}) + \textcolor{red}{6} \text{HCl} (\text{aq}) \rightarrow \textcolor{red}{2} \text{AlCl}_3 (\text{aq}) + \textcolor{red}{3} \text{H}_2 (\text{g})\)

b) \(\textcolor{red}{3} \text{Mg} (\text{s}) + \textcolor{red}{2} \text{FeCl}_3 (\text{aq}) \rightarrow \textcolor{red}{3} \text{MgCl}_2 (\text{aq}) + \textcolor{red}{2} \text{Fe} (\text{s})\)

c) \(\textcolor{red}{1} \text{F}_2 (\text{g}) + \textcolor{red}{2} \text{KCl} (\text{aq}) \rightarrow \textcolor{red}{2} \text{KF} (\text{aq}) + \text{Cl}_2 (\text{g})\)

d) \(\text{Cu} (\text{s}) + \text{Ca(NO}_3)_2 \rightarrow \text{No Reaction} (\text{Ca is more reactive than Cu})\)

e) \(\textcolor{red}{1} \text{Pb} (\text{s}) + \textcolor{red}{2} \text{AgNO}_3 (\text{aq}) \rightarrow \text{Pb(NO}_3)_2 (\text{aq}) + \textcolor{red}{2} \text{Ag} (\text{s})\)
4. Identify each reaction below as oxidation or reduction:

a) \[
Pb \rightarrow Pb^{2+} + 2 e^- \quad \text{oxidation} (\text{loss of electrons})
\]

b) \[
C_2H_2 + H_2 \rightarrow C_2H_4 \quad \text{reduction} (\text{gain of hydrogen})
\]

c) \[
Cr^{3+} + 3 e^- \rightarrow Cr \quad \text{reduction} (\text{gain of electrons})
\]

d) \[
C_2H_5OH \rightarrow C_2H_4O + H_2 \quad \text{oxidation} (\text{loss of hydrogen})
\]

5. In the following reactions, identify which reactant is oxidized and which is reduced:

a) \[
2 \text{Li (s)} + \text{F}_2 (g) \rightarrow 2 \text{LiF (s)}
\]

oxidized: \textbf{Li} \quad \text{reduced: \textbf{F}_2}

b) \[
\text{Cl}_2 (g) + 2 \text{KI (aq)} \rightarrow 2 \text{KCl (aq)} + \text{I}_2 (g)
\]

oxidized: \textbf{I} \quad \text{reduced: \textbf{Cl}_2}

c) \[
\text{Zn (s)} + \text{CuSO}_4 (aq) \rightarrow \text{ZnSO}_4 (aq) + \text{Cu (s)}
\]

oxidized: \textbf{Zn} \quad \text{reduced: \textbf{Cu}^{2+}}

d) \[
2 \text{PbO (s)} \rightarrow 2 \text{Pb (s)} + \text{O}_2 (g)
\]

oxidized: \textbf{O}^{2-} \quad \text{reduced: \textbf{Pb}^{2+}}
6. Calculate each of the following quantities:

a) Number of moles in 112 g of aspirin, C₉H₈O₄

Molar mass = [9(12.0) + 8(1.01) + 4(16.0)] = 180.1 g/mol

Moles = \( \frac{112 \text{ g}}{180.1 \text{ g/mol}} \times \frac{1 \text{ mol}}{180.1 \text{ g}} = 0.622 \text{ mol} \) (3 sig figs)

b) Mass of 3.82 moles of silver acetate, AgC₂H₃O₂

Molar mass = [107.9 + 2(12.0) + 3(1.01) + 2(16.0)] = 166.9 g/mol

Mass = \( 3.82 \text{ mol} \times \frac{166.9 \text{ g}}{1 \text{ mol}} = 638 \text{ g} \) (3 sig figs)

c) Number of molecules in 1.75 moles of CO₂

Molar mass = 12.0 + 2(16.2) = 44.0 g/mol

# of CO₂ molecules = \( 1.75 \text{ mol} \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 1.05 \times 10^{24} \text{ molecules} \) (3 sig figs)

d) Number of mole of H atoms in 20.0 g of CH₄

Molar mass = 12.0 + 4(1.01) = 16.0 g/mol

# of CH₄ molecules = \( 20.0 \text{ g} \times \frac{1 \text{ mol}}{16.0 \text{ g}} \times \frac{4 \text{ mol H atoms}}{1 \text{ mol CH}_4} = 5.00 \text{ mol H atoms} \) (3 sig figs)
7. Use the equation below to determine the mole ratios below:

\[ 2 \text{CH}_4 + 3 \text{O}_2 + 2 \text{NH}_3 \rightarrow 2 \text{HCN} + 6 \text{H}_2\text{O} \]

A) \[ \frac{\text{mol NH}_3}{\text{mol H}_2\text{O}} = \frac{2}{6} = \frac{1}{3} \]

B) \[ \frac{\text{mol HCN}}{\text{mol O}_2} = \frac{2}{3} \]

C) \[ \frac{\text{mol H}_2\text{O}}{\text{mol CH}_4} = \frac{6}{2} = \frac{3}{1} \]

D) \[ \frac{\text{mol O}_2}{\text{mol H}_2\text{O}} = \frac{3}{6} = \frac{1}{2} \]

Use the reaction shown below to answer the next 3 questions:

\[ 2 \text{C}_2\text{H}_6 (g) + 7 \text{O}_2 (g) \rightarrow 4 \text{CO}_2 (g) + 6 \text{H}_2\text{O} (g) \]

8. How many moles of water can be produced when 1.8 moles of \( \text{C}_2\text{H}_6 \) are used?

\[
1.8 \text{ mol } \text{C}_2\text{H}_6 \times \frac{6 \text{ mol H}_2\text{O}}{2 \text{ mol } \text{C}_2\text{H}_6} = 5.4 \text{ mol H}_2\text{O}
\]

9. How many moles of \( \text{CO}_2 \) are produced when 25.0 g of oxygen are consumed?

\[
25.0 \text{ g } \text{O}_2 \times \frac{1 \text{ mol } \text{O}_2}{32.0 \text{ g } \text{O}_2} \times \frac{4 \text{ mol CO}_2}{7 \text{ mol } \text{O}_2} = 0.446 \text{ mol CO}_2
\]

10. How many grams of water is produced when 78.0 g of \( \text{C}_2\text{H}_6 \) are burned?

\[
78.0 \text{ g } \text{C}_2\text{H}_6 \times \frac{1 \text{ mol } \text{C}_2\text{H}_6}{30.06 \text{ g } \text{C}_2\text{H}_6} \times \frac{6 \text{ mol H}_2\text{O}}{2 \text{ mol } \text{C}_2\text{H}_6} \times \frac{18.0 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 140. \text{ g H}_2\text{O}
\]
11. In the reaction shown below, if 10.0 g of CH₄ is combined with 30.0 g of O₂, what is maximum amount of CO₂ that can be produced?

\[ \text{CH}_4 (g) + 2 \text{O}_2 (g) \rightarrow \text{CO}_2 (g) + 2 \text{H}_2\text{O} (g) \]

Assume CH₄ is LR:

\[
10.0 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol CH}_4} = 0.623 \text{ mol CO}_2
\]

Assume O₂ is LR:

\[
30.0 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.0 \text{ g O}_2} \times \frac{1 \text{ mol CO}_2}{2 \text{ mol O}_2} = 0.469 \text{ mol CO}_2
\]

The second assumption is correct. Therefore, oxygen is the limiting reactant

\[
0.469 \text{ mol CO}_2 \times \frac{44.0 \text{ g}}{1 \text{ mol}} = 20.6 \text{ g CO}_2
\]

12. In an experiment with Zn and S, it was found that 30.7 g of ZnS was produced. If the percent yield of the reaction was 93.7%, what is the theoretical yield of this reaction?

Actual yield = 30.7 g
Theoretical yield = ???
Percent yield = 93.7

\[
\% \text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}}
\]

Theoretical yield = \[
\frac{\text{Actual yield}}{\% \text{ yield}} = \frac{30.7 \text{ g}}{0.937} = 32.8 \text{ g}
\]
13. How many grams of SO\textsubscript{2} can be produced from reaction of 10.0 g of H\textsubscript{2}S and 10.0 g of O\textsubscript{2}, as shown below:

\[ 2 \text{H}_2\text{S} + 3 \text{O}_2 \rightarrow 2 \text{SO}_2 + 2 \text{H}_2\text{O} \]

Assume H\textsubscript{2}S is LR:

\[
10.0 \text{ g H}_2\text{S} \times \frac{1 \text{ mol H}_2\text{S}}{34.1 \text{ g H}_2\text{S}} \times \frac{2 \text{ mol SO}_2}{2 \text{ mol H}_2\text{S}} = 0.293 \text{ mol SO}_2
\]

Assume O\textsubscript{2} is LR:

\[
10.0 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.0 \text{ g O}_2} \times \frac{2 \text{ mol SO}_2}{3 \text{ mol O}_2} = 0.208 \text{ mol SO}_2
\]

The second assumption is correct. Therefore, oxygen is the limiting reactant.

\[
0.208 \text{ mol SO}_2 \times \frac{64.1 \text{ g}}{1 \text{ mol}} = 13.3 \text{ g SO}_2
\]

14. When 50.0 g of N\textsubscript{2} is reacted with an excess of other reactants as shown below, 20.0 g of NaCN was produced. What is the percent yield of this reaction?

\[ \text{Na}_2\text{CO}_3 + 4 \text{C} + \text{N}_2 \rightarrow 2 \text{NaCN} + 3 \text{CO} \]

Calculate the theoretical yield:

\[
50.0 \text{ g N}_2 \times \frac{1 \text{ mol N}_2}{28.0 \text{ g N}_2} \times \frac{2 \text{ mol NaCN}}{1 \text{ mol N}_2} \times \frac{49.0 \text{ g NaCN}}{1 \text{ mol NaCN}} = 175 \text{ g NaCN}
\]

Calculate the percent yield:

\[
\% \text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 = \frac{20.0 \text{ g}}{175 \text{ g}} \times 100 = 11.4 \%
\]
15. The formation of Fe$_2$O$_3$ from iron and oxygen gas releases $1.7 \times 10^3$ kJ of heat, as shown below:

$$4 \text{Fe} \ (s) \ + \ 3 \text{O}_2 \ (g) \ \rightarrow \ 2 \text{Fe}_2\text{O}_3 \ (s) \quad \Delta H = -1.7 \times 10^3 \text{kJ}$$

a) How many kJ are released when 2.00 g of Fe react?

$$\text{2.00 g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \times \frac{1.7 \times 10^3 \text{ kJ}}{4 \text{ mol Fe}} = 15.2 \text{ kJ}$$

b) How many grams of Fe$_2$O$_3$ are produced when 475 kJ of heat are released?

$$\text{475 kJ} \times \frac{2 \text{ mol Fe}_2\text{O}_3}{1.7 \times 10^3 \text{ kJ}} \times \frac{159.7 \text{ g}}{1 \text{ mol Fe}_2\text{O}_3} = 89.2 \text{ g Fe}_2\text{O}_3$$