1. One of the steps in the commercial process for converting ammonia to nitric acid is the conversion of \( \text{NH}_3 \) to \( \text{NO} \). In a certain experiment, 1.50g of \( \text{NH}_3 \) reacts with 2.75g of \( \text{O}_2 \) according to the following equation:

\[
4\text{NH}_3(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 4\text{NO}(\text{g}) + 6\text{H}_2\text{O}(\text{g})
\]

1.1 Determine the limiting reactant.

\[
\frac{n_{\text{NO}}}{4} = \frac{n_{\text{O}_2}}{5} \quad \text{nNO} = n_{\text{O}_2} \times 4 = \frac{2.75}{2(15.999\text{g mol}^{-1})} \times 4 = 0.0688\text{mol}
\]

\[
\frac{n_{\text{NO}}}{4} = \frac{n_{\text{NH}_3}}{4} \quad \text{nNO} = n_{\text{NH}_3} \times 4 = \frac{1.50}{17.031\text{g mol}^{-1}} = 0.0881\text{mol}
\]

The limiting reagent is therefore \( \text{O}_2 \)

Alternative Method:

\[
\text{mol} \text{NH}_3 = 4
\]
\[
\text{mol} \text{O}_2 = 5
\]

\[
n_{\text{NH}_3} = 4/5 \times n_{\text{O}_2} = 4/5 \times 2.75/31.998\text{g mol}^{-1} = 0.8 \times 0.0859 = 0.0687 \text{mol}
\]

Available at the start \( n_{\text{NH}_3} = 1.5\text{mol}/17.031\text{g mol}^{-1} = 0.0881\text{mol} \)

Therefore the limiting reactant is therefore \( \text{O}_2 \) since \( \text{NH}_3 \) is in excess.

1.2 Calculate how many grams of NO and of \( \text{H}_2\text{O} \) are formed?

The limiting reagent determines the amount of product formed therefore

\[
\frac{n_{\text{NO}}}{4} = \frac{n_{\text{O}_2}}{5} \quad m_{\text{NO}} = \frac{2.75g/2(15.999\text{g mol}^{-1})}{5} \times 4 \times 30.006\text{g mol}^{-1} = 2.06\text{g}
\]
\[
\frac{n_{\text{H}_2\text{O}}}{6} = \frac{n_{\text{O}_2}}{5}
\]
\[
m_{\text{H}_2\text{O}} = \frac{2.75g}{2(15.999\text{ g mol}^{-1})} \times 6 \times 18.015\text{ g mol}^{-1} = 1.86\text{ g}
\]

1.3 Calculate the mass, (in grams) of the excess reactant remaining after the limiting reactant is completely consumed?

\[
\frac{n_{\text{NH}_3}}{4} = \frac{n_{\text{NH}_3}}{4} - \frac{n_{\text{O}_2}}{5}
\]

Mass of NH\textsubscript{3} = \left(\frac{1.50g}{17.031\text{ g mol}^{-1}} - \frac{2.75g}{31.998\text{ g mol}^{-1}}\right)x 4 \times 17.031\text{ g mol}^{-1} = \boxed{0.327\text{ g}}

2. When benzene, C\textsubscript{6}H\textsubscript{6}, reacts with bromine, Br\textsubscript{2}, bromobenzene (C\textsubscript{6}H\textsubscript{5}Br) is formed according to the equation below:

\[
\text{C}_6\text{H}_6 + \text{Br}_2 \rightarrow \text{C}_6\text{H}_5\text{Br} + \text{HBr}
\]

2.1 Calculate the theoretical yield of bromobenzene in the reaction when 3.0g of benzene reacts with 65.0g of bromine?

\[
\frac{n_{\text{C}_6\text{H}_5\text{Br}}}{1} = \frac{n_{\text{C}_6\text{H}_6}}{1} \quad n_{\text{C}_6\text{H}_5\text{Br}} = \frac{3.0g}{78.05\text{ g mol}^{-1}} = 0.04 \text{ mol}
\]

\[
\frac{n_{\text{C}_6\text{H}_5\text{Br}}}{1} = \frac{n_{\text{Br}_2}}{1} \quad n_{\text{C}_6\text{H}_5\text{Br}} = \frac{65.0g}{159.8\text{ g mol}^{-1}} = 0.41 \text{ mol}
\]

Therefore C\textsubscript{6}H\textsubscript{6} is the limiting reagent.

Theoretical Yield of C\textsubscript{6}H\textsubscript{5}Br = \frac{n_{\text{C}_6\text{H}_5\text{Br}}}{1} \times M_{\text{C}_6\text{H}_5\text{Br}} = 0.04 \times 156.94\text{ g mol}^{-1} = 6.28\text{ g}
2.2 If the actual yield of bromobenzene was 5.67g what was the percentage yield?

\[
\% \text{ Yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100
\]

\[
= \frac{5.67g}{6.28g} \times 100
\]

\[
= 90.29 \%
\]

3. Write the balanced molecular, total ionic and net ionic equations for the reaction below, showing the states for all products (s, l, g, aq) formed:

*An aqueous solution of silver nitrate was mixed with an aqueous solution of copper(II) chloride (CuCl}_2).

**Molecular equation:**

\[
2\text{AgNO}_3(\text{aq}) + \text{CuCl}_2(\text{aq}) \rightarrow 2\text{AgCl(s)} + \text{Cu(NO}_3)_2(\text{aq})
\]

**Total ionic eqn:**

\[
2\text{Ag}^+ (\text{aq}) + 2\text{NO}_3^- (\text{aq}) + \text{Cu}^{2+} (\text{aq}) + 2\text{Cl}^- (\text{aq}) \rightarrow 2\text{AgCl(s)} + \text{Cu}^{2+} (\text{aq}) + 2\text{NO}_3^- (\text{aq})
\]

**Net ionic eqn:**

\[
2\text{Ag}^+ (\text{aq}) + 2\text{Cl}^- (\text{aq}) \rightarrow 2\text{AgCl(s)}
\]

4. Predict (using solubility rules) whether a reaction will occur in each of the following cases. If so, write the balanced molecular, total ionic and net ionic equations for the following reaction: \( \text{KNO}_3(\text{aq}) + \text{NaCl(aq)} \rightarrow \)

**Molecular equation:**

\[
\text{KNO}_3(\text{aq}) + \text{NaCl(aq)} \rightarrow \text{NaNO}_3(\text{aq}) + \text{KCl(aq)}
\]

**Total ionic equation:**

\[
\text{K}^+(\text{aq}) + \text{NO}_3^-(\text{aq}) + \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{Na}^+(\text{aq}) + \text{NO}_3^-(\text{aq}) + \text{K}^+(\text{aq}) + \text{Cl}^-(\text{aq})
\]

No precipitate forms. All ions remain in solution.
5. A 0.1216 g sample of an acid is dissolved in 25.00 mL water, and the resulting solution is titrated with 0.1104 mol dm$^{-3}$ NaOH solution. A 12.50 mL volume of the base is required to neutralize the acid. The reacting ratio of the acid to NaOH is 1:1. Calculate the molar mass of the acid.

\[
\text{moles of base, } n_{\text{NaOH}} = (0.01250 \text{ dm}^3) \times 0.1104 \text{ mol dm}^{-3} \\
= 0.001380 \\
= \text{moles of acid}
\]

\[
\text{molar mass of acid} = \frac{0.1216 \text{ g}}{0.001380 \text{ mol}} = 88.11 \text{ g mol}^{-1}
\]