

$$30. \text{ volume of CO}_2 \text{ (based on C}_5\text{H}_{12}) = 0.0250 \text{ L C}_5\text{H}_{12} \times \frac{626.0 \text{ g C}_5\text{H}_{12}}{1 \text{ L C}_5\text{H}_{12}} \times \frac{1 \text{ mol C}_5\text{H}_{12}}{72.0 \text{ g C}_5\text{H}_{12}} \times \frac{5 \text{ mol CO}_2}{1 \text{ mol C}_5\text{H}_{12}} \\ \times \frac{22.4 \text{ L CO}_2}{1 \text{ mol CO}_2} = 24.3 \text{ L}$$

$$\text{volume of CO}_2 \text{ (based on O}_2) = 40.0 \text{ L O}_2 \times \frac{1 \text{ mol O}_2}{22.4 \text{ L O}_2} \times \frac{5 \text{ mol CO}_2}{8 \text{ mol O}_2} \times \frac{22.4 \text{ L CO}_2}{1 \text{ mol CO}_2} = 25.0 \text{ L}$$

Hence, the C_5H_{12} is the limiting reactant and **24.3 L** of $\text{CO}_2(\text{g})$ will be produced.

$$31. \text{ moles of HCl} = 0.100 \frac{\text{mol}}{\text{L}} \times 0.0500 \text{ L} = 5.00 \times 10^{-3} \text{ mol}$$

$$\text{moles of NaCl (based on HCl)} = 5.00 \times 10^{-3} \text{ mol HCl} \times \frac{1 \text{ mol NaCl}}{1 \text{ mol HCl}} = 5.00 \times 10^{-3} \text{ mol}$$

$$\text{moles of NaOH} = 0.200 \frac{\text{mol}}{\text{L}} \times 0.0300 \text{ L} = 6.00 \times 10^{-3} \text{ mol}$$

$$\text{moles of NaCl (based on NaOH)} = 6.00 \times 10^{-3} \text{ mol NaOH} \times \frac{1 \text{ mol NaCl}}{1 \text{ mol NaOH}} = 6.00 \times 10^{-3} \text{ mol}$$

Since the NaOH can produce more NaCl, the **NaOH** is in excess.

$$32. \text{ mass of BaBr}_2 \text{ [based on Ba(OH)}_2] = 0.250 \text{ g Ba(OH)}_2 \times \frac{1 \text{ mol Ba(OH)}_2}{171.3 \text{ g Ba(OH)}_2} \times \frac{1 \text{ mol BaBr}_2}{1 \text{ mol Ba(OH)}_2} \\ \times \frac{297.1 \text{ g BaBr}_2}{1 \text{ mol BaBr}_2} = 0.434 \text{ g}$$

$$\text{moles of HBr} = 0.125 \frac{\text{mol}}{\text{L}} \times 0.0150 \text{ L} = 1.875 \times 10^{-3} \text{ mol}$$

$$\text{mass of BaBr}_2 \text{ (based on HBr)} = 1.875 \times 10^{-3} \text{ mol HBr} \times \frac{1 \text{ mol BaBr}_2}{2 \text{ mol HBr}} \times \frac{297.1 \text{ g BaBr}_2}{1 \text{ mol BaBr}_2} = 0.279 \text{ g}$$

Since HBr is the limiting reactant, **0.279 g** of BaBr_2 can be formed.

33. (a) First assume the FeCO_3 is 100 % pure.

$$\text{mass of Fe}_2\text{O}_3 = 15.0 \text{ g FeCO}_3 \times \frac{1 \text{ mol FeCO}_3}{115.8 \text{ g FeCO}_3} \times \frac{2 \text{ mol Fe}_2\text{O}_3}{4 \text{ mol FeCO}_3} \times \frac{159.6 \text{ g Fe}_2\text{O}_3}{1 \text{ mol Fe}_2\text{O}_3} = 10.3 \text{ g}$$

Since the FeCO_3 is only 42.0 % pure there will be less than 10.3 g.

$$\text{mass of pure FeCO}_3 = 0.420 \times 10.3 \text{ g} = \mathbf{4.34 \text{ g}}$$

(b) First calculate the mass of pure FeCO_3 required to produce 37.0 g of Fe_2O_3 .

$$\text{mass of FeCO}_3 = 37.0 \text{ g Fe}_2\text{O}_3 \times \frac{1 \text{ mol Fe}_2\text{O}_3}{159.6 \text{ g Fe}_2\text{O}_3} \times \frac{4 \text{ mol FeCO}_3}{2 \text{ mol Fe}_2\text{O}_3} \times \frac{115.8 \text{ g FeCO}_3}{1 \text{ mol FeCO}_3} = 53.69 \text{ g}$$

$$\text{Then: } \% \text{ purity} = \frac{\text{mass of pure FeCO}_3}{\text{mass of impure FeCO}_3} \times 100\% = \frac{53.69 \text{ g}}{55.0 \text{ g}} \times 100\% = \mathbf{97.6\%}$$

(c) First calculate the mass of Fe_2O_3 EXPECTED from the reaction.

$$\text{mass of Fe}_2\text{O}_3 = 35.0 \text{ g FeCO}_3 \times \frac{1 \text{ mol FeCO}_3}{115.8 \text{ g FeCO}_3} \times \frac{2 \text{ mol Fe}_2\text{O}_3}{4 \text{ mol FeCO}_3} \times \frac{159.6 \text{ g Fe}_2\text{O}_3}{1 \text{ mol Fe}_2\text{O}_3} = 24.12 \text{ g}$$

$$\text{Now: } \% \text{ yield} = \frac{\text{mass obtained}}{\text{mass expected}} \times 100\% = \frac{22.5 \text{ g}}{24.12 \text{ g}} \times 100\% = \mathbf{93.3\%}$$

(d) First find the mass of 100% pure FeCO_3 required to make $1.00 \times 10^3 \text{ g}$ of Fe_2O_3 .

$$\begin{aligned} \text{mass of FeCO}_3 &= 1.00 \times 10^3 \text{ g Fe}_2\text{O}_3 \times \frac{1 \text{ mol Fe}_2\text{O}_3}{159.6 \text{ g Fe}_2\text{O}_3} \times \frac{4 \text{ mol FeCO}_3}{2 \text{ mol Fe}_2\text{O}_3} \times \frac{115.8 \text{ g FeCO}_3}{1 \text{ mol FeCO}_3} \\ &= 1451 \text{ g} \end{aligned}$$

Since not all the ore was pure FeCO_3 , dividing by the percentage purity will increase the amount of ore to be used and allow for losses in forming products.

$$\text{mass of ore} = \frac{1451 \text{ g}}{0.628} = 2.31 \times 10^3 \text{ g}$$

34. First calculate the mass of pure FeS_2 required to produce 4.50 L of SO_2 .

$$\text{mass of FeS}_2 = 4.50 \text{ L SO}_2 \times \frac{1 \text{ mol SO}_2}{22.4 \text{ L SO}_2} \times \frac{4 \text{ mol FeS}_2}{8 \text{ mol SO}_2} \times \frac{120.0 \text{ g FeS}_2}{1 \text{ mol FeS}_2} = 12.05 \text{ g}$$

$$\% \text{ purity} = \frac{\text{mass of pure FeS}_2}{\text{mass of impure FeS}_2} \times 100\% = \frac{12.05 \text{ g}}{100.0 \text{ g}} \times 100\% = 12.1\%$$

$$\begin{aligned} 35. \text{ (a) volume of C}_6\text{H}_5\text{NO}_2 \text{ EXPECTED} &= 25.0 \text{ mL C}_6\text{H}_6 \times \frac{0.879 \text{ g C}_6\text{H}_6}{1 \text{ mL C}_6\text{H}_6} \times \frac{1 \text{ mol C}_6\text{H}_6}{78.0 \text{ g C}_6\text{H}_6} \times \\ &\quad \frac{1 \text{ mol C}_6\text{H}_5\text{NO}_2}{1 \text{ mol C}_6\text{H}_6} \times \frac{123.0 \text{ g C}_6\text{H}_5\text{NO}_2}{1 \text{ mol C}_6\text{H}_5\text{NO}_2} \times \frac{1 \text{ mL C}_6\text{H}_5\text{NO}_2}{1.204 \text{ g C}_6\text{H}_5\text{NO}_2} = 28.781 \text{ mL} \end{aligned}$$

$$\% \text{ yield} = \frac{\text{volume obtained}}{\text{volume expected}} \times 100\% = \frac{18.0 \text{ mL}}{28.781 \text{ mL}} \times 100\% = 62.54\% \text{ (which rounds to } \mathbf{62.5\%})$$

$$\begin{aligned} \text{(b) mass of C}_6\text{H}_6 \text{ reacted} &= 18.0 \text{ mL C}_6\text{H}_5\text{NO}_2 \times \frac{1.204 \text{ g C}_6\text{H}_5\text{NO}_2}{1 \text{ mL C}_6\text{H}_5\text{NO}_2} \times \frac{1 \text{ mol C}_6\text{H}_5\text{NO}_2}{123.0 \text{ g C}_6\text{H}_5\text{NO}_2} \\ &\quad \times \frac{1 \text{ mol C}_6\text{H}_6}{1 \text{ mol C}_6\text{H}_5\text{NO}_2} \times \frac{78.0 \text{ g C}_6\text{H}_6}{1 \text{ mol C}_6\text{H}_6} = 13.743 \text{ g} \end{aligned}$$

$$\text{mass of C}_6\text{H}_6 \text{ (originally)} = 25.0 \text{ mL} \times \frac{0.879 \text{ g}}{1 \text{ mL}} = 21.975 \text{ g}$$

$$\text{mass of C}_6\text{H}_6 \text{ (unreacted)} = 21.975 \text{ g} - 13.743 \text{ g} = \mathbf{8.23 \text{ g}}$$

Alternately: if the reaction has a 62.54% yield, then $100.0 - 62.54 = 37.46\%$ of the C_6H_6 is unreacted.

$$\text{mass of C}_6\text{H}_6 \text{ (originally)} = 21.975 \text{ g (as calculated above)}$$

$$\text{and: mass of C}_6\text{H}_6 \text{ (unreacted)} = 21.975 \text{ g} \times 0.3746 = \mathbf{8.23 \text{ g}}$$

$$36. \text{ (a) mass of SiF}_4 \text{ formed} = 2.50 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} \times \frac{1 \text{ mol SiF}_4}{2 \text{ mol H}_2\text{O}} \times \frac{104.1 \text{ g SiF}_4}{1 \text{ mol SiF}_4} = \mathbf{7.23 \text{ g}}$$

$$\text{(b) mass of SiO}_2 \text{ used} = 2.50 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} \times \frac{1 \text{ mol SiO}_2}{2 \text{ mol H}_2\text{O}} \times \frac{60.1 \text{ g SiO}_2}{1 \text{ mol SiO}_2} = 4.17 \text{ g}$$

$$\text{mass of SiO}_2 \text{ (unreacted)} = 12.20 - 4.17 = \mathbf{8.03 \text{ g}}$$

$$\text{(c) \% yield} = \frac{\text{mass of SiO}_2 \text{ used}}{\text{mass of SiO}_2 \text{ available}} \times 100\% = \frac{4.17 \text{ g}}{12.20 \text{ g}} \times 100\% = \mathbf{34.2\%}$$

$$\text{or: mass of SiF}_4 \text{ expected} = 12.20 \text{ g SiO}_2 \times \frac{1 \text{ mol SiO}_2}{60.1 \text{ g SiO}_2} \times \frac{1 \text{ mol SiF}_4}{1 \text{ mol SiO}_2} \times \frac{104.1 \text{ g SiF}_4}{1 \text{ mol SiF}_4} = 21.13 \text{ g}$$

$$\% \text{ yield} = \frac{\text{mass of SiF}_4 \text{ obtained}}{\text{mass of SiF}_4 \text{ expected}} \times 100\% = \frac{7.23 \text{ g}}{21.13 \text{ g}} \times 100\% = \mathbf{34.2\%}$$

37. (a) First find the mass of CuO produced if the purity and yield are both 100%.

$$\text{mass of CuO} = 5.00 \times 10^3 \text{ g malach} \times \frac{1 \text{ mol malach}}{221.0 \text{ g malach}} \times \frac{2 \text{ mol CuO}}{1 \text{ mol malach}} \times \frac{79.5 \text{ g CuO}}{1 \text{ mol CuO}} = 3597 \text{ g}$$

Next allow for a purity of 4.30 % and a yield of 84.0 % by decreasing the mass of products formed.

$$\text{mass of CuO formed} = 0.0430 \times 0.840 \times 3597 \text{ g} = \mathbf{1.30 \times 10^2 \text{ g}}$$

- (b) First find the mass of pure malachite required to produce 100.0 g of CuO, assuming 100 % yield.

$$\text{mass of malachite} = 100.0 \text{ g CuO} \times \frac{1 \text{ mol CuO}}{79.5 \text{ g CuO}} \times \frac{1 \text{ mol malachite}}{2 \text{ mol CuO}} \times \frac{221.0 \text{ g malachite}}{1 \text{ mol malachite}} = 139.0 \text{ g}$$

Both percent purity and percentage yield decrease the mass of products so that the mass of malachite ore used must be increased to compensate for these losses. Dividing the mass of pure malachite by the percent purity and percent yield gives the required increased mass.

$$\text{mass of ore used} = \frac{139.0 \text{ g}}{0.870 \times 0.0370} = \mathbf{4.32 \times 10^3 \text{ g}}$$

38. (a) First find the mass of Ag produced if the ore is 100% pure Ag₂S.

$$\text{mass of Ag} = 250.0 \times 10^3 \text{ g Ag}_2\text{S} \times \frac{1 \text{ mol Ag}_2\text{S}}{247.9 \text{ g Ag}_2\text{S}} \times \frac{2 \text{ mol Ag}}{1 \text{ mol Ag}_2\text{S}} \times \frac{107.9 \text{ g Ag}}{1 \text{ mol Ag}} = 2.18 \times 10^5 \text{ g}$$

Now find the mass of silver if the ore is only 0.135% pure.

$$\text{mass of Ag} = 0.00135 \times 2.18 \times 10^5 \text{ g} = \mathbf{294 \text{ g}}$$

- (b) Find the mass of Ag₂S which produces 0.261 g of silver.

$$\text{mass of Ag}_2\text{S} = 0.261 \text{ g Ag} \times \frac{1 \text{ mol Ag}}{107.9 \text{ g Ag}} \times \frac{1 \text{ mol Ag}_2\text{S}}{2 \text{ mol Ag}} \times \frac{247.9 \text{ g Ag}_2\text{S}}{1 \text{ mol Ag}_2\text{S}} = 0.300 \text{ g}$$

$$\text{Hence: \% purity} = \frac{\text{mass of pure Ag}_2\text{S}}{\text{mass of impure ore}} \times 100\% = \frac{0.300 \text{ g}}{76.4 \text{ g}} \times 100\% = \mathbf{0.392\%}$$

- (c) First find the mass of pure Ag which is produced by 152.6 g of pure Ag₂S.

$$\text{mass of Ag} = 152.6 \text{ g Ag}_2\text{S} \times \frac{1 \text{ mol Ag}_2\text{S}}{247.9 \text{ g Ag}_2\text{S}} \times \frac{2 \text{ mol Ag}}{1 \text{ mol Ag}_2\text{S}} \times \frac{107.9 \text{ g Ag}}{1 \text{ mol Ag}} = 132.8 \text{ g}$$

$$\text{Hence: \% yield} = \frac{\text{actual mass of pure Ag}}{\text{expected mass of pure Ag}} \times 100\% = \frac{117.4 \text{ g}}{132.8 \text{ g}} \times 100\% = \mathbf{88.38\%}$$

- (d) First find the mass of pure Ag₂S required to produce 50.0 kg of pure Ag.

$$\begin{aligned} \text{mass of pure Ag}_2\text{S} &= 50.0 \times 10^3 \text{ g Ag} \times \frac{1 \text{ mol Ag}}{107.9 \text{ g Ag}} \times \frac{1 \text{ mol Ag}_2\text{S}}{2 \text{ mol Ag}} \times \frac{247.9 \text{ g Ag}_2\text{S}}{1 \text{ mol Ag}_2\text{S}} \\ &= 5.744 \times 10^4 \text{ g} \end{aligned}$$

Now find the mass of ore needed if the ore only contains 0.795% Ag₂S.

$$\text{mass of ore} = \frac{5.744 \times 10^4 \text{ g}}{0.00795} = \mathbf{7.22 \times 10^6 \text{ g}}$$

- (e) First find the mass of Ag produced if the purity and yield are both 100%.

$$\begin{aligned} \text{mass of Ag expected} &= 3.50 \times 10^7 \text{ g Ag}_2\text{S} \times \frac{1 \text{ mol Ag}_2\text{S}}{247.9 \text{ g Ag}_2\text{S}} \times \frac{2 \text{ mol Ag}}{1 \text{ mol Ag}_2\text{S}} \times \frac{107.9 \text{ g Ag}}{1 \text{ mol Ag}} \\ &= 3.047 \times 10^7 \text{ g} \end{aligned}$$

Now, the ore is only 1.86% pure Ag₂S and only 89.2% of the Ag₂S is extracted.

$$\text{actual mass of Ag produced} = 3.047 \times 10^7 \text{ g} \times 0.0186 \times 0.892 = \mathbf{5.05 \times 10^5 \text{ g}}$$