3.36 Describe how the knowledge of the percent composition by mass of an unknown compound can help us identify the compound.

For a pure material, the percent composition is constant. So, to determine what a sample is, the percent composition will provide evidence of its formula.

3.39 Tin (Sn) exists in Earth’s crust as SnO₂. Calculate the percent composition by mass of Sn and O in SnO₂.

\[
\%Sn = \frac{118.7 \text{ g/mol}}{150.7 \text{ g/mol}} \times 100\% = 78.77\%
\]

\[
\%O = \frac{2 \times (16.00 \text{ g/mol})}{150.7 \text{ g/mol}} \times 100\% = 21.23\%
\]

3.40 Calculate the percent composition of chloroform (CHCl₃).

\[
\%C = \frac{12.01 \text{ g/mol}}{119.4 \text{ g/mol}} \times 100\% = 10.06\%
\]

\[
\%H = \frac{1.008 \text{ g/mol}}{119.4 \text{ g/mol}} \times 100\% = 0.8442\%
\]

\[
\%Cl = \frac{3 \times (35.45) \text{ g/mol}}{119.4 \text{ g/mol}} \times 100\% = 89.07\%
\]

**Check:** The sum of the percentages is (10.06% + 0.8442% + 89.07%) = 99.97%. The small discrepancy from 100% is due to the way we rounded off.

3.48 What is the mass of F, in grams, in 24.6 g of tin(II) fluoride (SnF₂)?

\[
\%F = \frac{\text{mass of F in 1 mol SnF}_2}{\text{molar mass of SnF}_2} \times 100\% = \frac{2 \times (19.00 \text{ g})}{156.7 \text{ g}} \times 100\% = 24.25\% 
\]

\[
? \text{ g F} = \left( \frac{24.25 \text{ g F}}{100 \text{ g SnF}_2} \right) (24.6 \text{ g SnF}_2) = 5.97 \text{ g F}
\]

**Note:** This problem could have been worked by the following conversions:

\[
g \text{ of SnF}_2 \rightarrow \text{ mol of SnF}_2 \rightarrow \text{ mol of F} \rightarrow \text{ g of F}
\]

3.50 What are the empirical formulas of the compounds with the following compositions?

(a) 40.1 % C, 6.6 % H, 53.3 % O

\[
\begin{align*}
n_c &= 40.1 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 3.34 \text{ mol C} \times 3.33 = 1 \\
n_h &= 6.6 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 6.5 \text{ mol H} \times 3.33 = 2 \\
n_o &= 53.3 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.33 \text{ mol O} \times 3.33 = 1
\end{align*}
\]

This gives the empirical formula, **CH₂O**.

(b) 60.1 % K, 18.4 % C, 21.5 % N

\[
\begin{align*}
n_k &= 60.1 \text{ g K} \times \frac{1 \text{ mol K}}{39.10 \text{ g K}} = 1.54 \text{ mol K} \\
n_c &= 18.4 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 1.53 \text{ mol C} \\
n_n &= 21.5 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 1.53 \text{ mol N}
\end{align*}
\]

Dividing by the smallest number of moles (1.53 mol) gives the empirical formula, **KCN**.
A. An unknown hydrocarbon is found to contain 84.21% C by mass. What is its empirical formula?

Since it is a hydrocarbon, it contains only C & H. Thus it contains 15.79% H by mass.

\[ n_C = 84.21 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 7.012 \text{ mol C} \approx 7 \times 4 = 4 \]

\[ n_H = 15.79 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 15.66 \text{ mol H} \approx 2.233 \times 4 = 8.933 \approx 9 \]

2.233 ≈ 2.25, so we multiply both moles by 4 to get \( C_4H_9 \).

B. The general formula of Epson salts can be written as \( \text{MgSO}_4 \cdot x \text{ H}_2\text{O} \). When 5.061 g of this hydrate is heated to 250°C, all the water of hydration is lost, leaving 2.472 g of \( \text{MgSO}_4 \). What is the value of \( x \)?

Moles \( \text{MgSO}_4 \) = 2.472 g \( \text{MgSO}_4 \) \times \frac{1 \text{ mol } \text{MgSO}_4}{120.36 \text{ g}} = 2.05 \times 10^{-2} \text{ mol } \text{MgSO}_4

Mass \( \text{H}_2\text{O} \) lost = 5.061 g – 2.472 g = 2.589 g \( \text{H}_2\text{O} \)

Moles \( \text{H}_2\text{O} \) = 2.589 g \( \text{H}_2\text{O} \) \times \frac{1 \text{ mol } \text{H}_2\text{O}}{18.02 \text{ g } \text{H}_2\text{O}} = 1.44 \times 10^{-1} \text{ mol } \text{H}_2\text{O}

\[ x = \frac{1.44 \times 10^{-1} \text{ mol } \text{H}_2\text{O}}{2.05 \times 10^{-2} \text{ mol } \text{MgSO}_4} = 7.01 \approx 7 \] (\( \text{MgSO}_4 \cdot 7 \text{ H}_2\text{O} \), magnesium sulfate heptahydrate)