Warm up

• How many liters of oxygen gas is contained in 458.3 grams of oxygen?

\[
\begin{array}{ccc}
458.3 \text{g O}_2 & 1 \text{mol O}_2 & 22.4 \text{ L O}_2 \\
32 \text{ g O}_2 & 1 \text{ mol O}_2 & \\
\end{array}
\]

\[= \boxed{320.8 \text{ L O}_2}\]
Percent composition:

- (remember: percent = part divided by whole x 100)
- the percent by mass of each element in a compound
- # of grams of the element per 100 grams of the compound
Ex. 1  Find the % composition for a compound that is formed from 222.6 g of Na and 77.4 g O.

\[
222.6 \text{g Na} + 77.4 \text{g O} = 300.0 \text{g total}
\]

\[
\frac{222.6 \text{g Na}}{300.0 \text{g total}} \times 100 = 74.20\% \text{ Na}
\]

\[
\frac{77.4 \text{g O}}{300.0 \text{g total}} \times 100 = 25.80\% \text{ O}
\]
OR solve for percent composition from the chemical formula:

\[
\% \text{ mass} = \frac{\text{molar mass of the element}}{\text{molar mass of compound}} \times 100
\]
#23. What is the percent composition of calcium acetate?

\[
\text{Ca(C}_2\text{H}_3\text{O}_2)_2
\]

\[
\begin{align*}
\text{Ca} & \quad 1 \times 40.08 = 40.08 \\
\text{C} & \quad 4 \times 12.01 = 48.04 \\
\text{H} & \quad 6 \times 1.01 = 6.06 \\
\text{O} & \quad 4 \times 16.0 = 64.0 \\
\end{align*}
\]

\[
\frac{158.18 \text{ g}}{}
\]
Ca = \( \frac{40.08}{158.18} \times 100 \) = 25.34% Ca

C = \( \frac{48.04}{158.18} \times 100 \) = 30.37% C

H = \( \frac{6.06}{158.18} \times 100 \) = 3.83% H

O = \( \frac{64.0}{158.18} \times 100 \) = 40.46% O
Application problem:

• How many grams of calcium can be recovered from 258 grams of calcium acetate?

• Given mass * % /100 (or change to decimal)

• 258 * .2534 = 65.4
CALCULATING EMPIRICAL FORMULAS

Empirical formula - lowest whole number ratio of the elements in a compound - may or may not be the same as the molecular formula.
Ex. H₂O is both the empirical & molecular formula.

H₂O₂ is a molecular formula.

HO is the empirical formula for H₂O₂.
Steps to calculate empirical formula:
1. Find moles of each element.
2. Set up mole ratio.
3. Simplify mole ratio (divide by smallest). If your answers are not in whole numbers, you must multiply by 2, 3, 4, or 5 to get whole numbers.

4. Use mole ratio as subscripts in the formula.

If given % composition, assume 100 g of compound.
#24. A compound is 79.8% C and 20.2% H. Find its empirical formula.

\[
\begin{align*}
79.8 \text{ g C} & \quad 1 \text{ mol C} = 6.65 \text{ mol C} \\
12.0 \text{ g C} & \\
20.2 \text{ g H} & \quad 1 \text{ mol H} = 20.0 \text{ mol H} \\
1.01 \text{ g H} & \\
\text{C:H} & \quad 6.65: 20.0 = 1 : 3.0 \\
\end{align*}
\]

\[
\frac{6.65}{6.65} \quad \frac{6.65}{6.65}
\]

\[\text{CH}_3\]
#25. An oxide of aluminum is formed by the complete reaction of 4.151g of aluminum with 3.692 g of oxygen. Calculate the empirical formula for this compound.

\[
\begin{align*}
4.151 \text{g Al} & \quad 1 \text{ mol Al} = 0.1537 \text{ mol Al} \\
27.0 \text{g Al} & \\
3.692 \text{g O} & \quad 1 \text{ mol O} = 0.23075 \text{ mol O} \\
16.0 \text{g O} & \\
\text{Al:O} & \quad \frac{0.1537}{0.1537} : \frac{0.23075}{0.1537} = 1 : 1.5 = 2 : 3
\end{align*}
\]

\[
\text{Al}_2\text{O}_3
\]
Molecular formulas are the actual formulas. They may be the same as the empirical formula or a multiple of it.
To find the multiple \((n)\), take the gram formulas mass \((gfm\) or molar mass) and divide by the empirical formula mass \((efm)\):

\[
n = \frac{gfm}{efm}
\]
Multiply each subscript in the empirical formula by n to get the molecular formula.
#26. A white powder is analyzed and found to have the empirical formula $\text{P}_2\text{O}_5$. The compound has a molar mass of 283.9 g. What is the compound’s molecular formula?

$\text{P}_2\text{O}_5 = 142.0 \text{ g/mol}$

gfm/efm

$\frac{283.9}{142.0} = 2$

$2(\text{P}_2\text{O}_5) = \text{P}_4\text{O}_{10}$
27. A compound used as an additive for gasoline to help prevent engine knock shows the following percentage composition:
71.65% Cl  24.27% C  4.07% H
The molar mass is known to be 98.96 g. Determine the empirical formula and the molecular formula for this compound.

\[
\begin{align*}
71.65 \text{ g Cl} & \quad 1 \text{ mol Cl} = 2.02 \text{ mol Cl} \\
35.5 \text{ g Cl} & \\
24.27 \text{ g C} & \quad 1 \text{ mol C} = 2.02 \text{ mol C} \\
12.0 \text{ g C} & \\
4.07 \text{ g H} & \quad 1 \text{ mol H} = 4.03 \text{ mol H} \\
1.01 \text{ g H} & \\
\text{Cl:C:H} &= 2.02 : 2.02 : 4.03 \\
&= \frac{2.02}{2.02} \quad \frac{2.02}{2.02} \quad \frac{2.02}{2.02}
\end{align*}
\]

1:1:2  ClCH₂ or CH₂Cl

efm = 12.0 + 2.0 + 35.5 = 49.5

98.96/49.5 = 2  C₂H₄Cl₂