**Additional Chemical Math Problems:**

Percent composition – the percentage by mass of each element present in a compound

**Example 1:** Calculate the % Comp (to two decimal places) of DDT (a banned insecticide) from its formula, $C_{14}H_9Cl_5$

C: $14 \times 12.01\text{amu} = 168.14\text{amu}$
H: $9 \times 1.01\text{amu} = 9.09\text{amu}$
Cl: $5 \times 35.45\text{amu} = 177.25\text{amu}$

$\text{FW} = 354.48\text{amu}$

$\% \text{ Element} = \frac{\text{mass of element in one formula unit}}{\text{FW}} \times 100$

$\% C = \frac{168.14\text{amu}}{354.48\text{amu}} \times 100 = 47.43\%$

$\% H = \frac{9.09\text{amu}}{354.48\text{amu}} \times 100 = 2.56\%$

$\% Cl = \frac{177.25\text{amu}}{354.48\text{amu}} \times 100 = 50.00\%$

Note: % Composition may also be calculated from mass data obtained from
1) Compound Decomposition
2) Compound Synthesis experiments

**Example 2:** Decomposition of a sample of a compound yields the following:
19.43g Na
13.55g S
27.04g O

What is the % composition for this compound?
Total mass of sample = 60.02g

\[ \% \text{Na} = \frac{19.43g}{60.02g} \times 100 = 32.37\% \]

\[ \% \text{S} = \frac{13.55g}{60.02g} \times 100 = 22.58\% \]

\[ \% \text{O} = \frac{27.04g}{60.02g} \times 100 = 45.04\% \]

**Example 3:** You learn that 100.0g of a compound, C\textsubscript{6}H\textsubscript{16}N\textsubscript{2}, is decomposed into 62.01g C, 13.87g H, 24.11g N. Calculate the \% Composition.

**The MOLE**

1 mole = 6.02 X \(10^{23}\) particles, atoms, objects, molecules, formula units

**Example 4:** Calculate the number of molecules in 2.67 mole of CO\textsubscript{2}.

\[
\begin{array}{c|c}
2.67 \text{ mole CO}_2 & 6.02 \times 10^{23} \text{ molecules CO}_2 \\
1 \text{ mole CO}_2 & 1 \text{ mole CO}_2
\end{array}
\]

\[= 1.61 \text{ molecules CO}_2 \]

**Molar Mass**

1 mole of any substance = FW in grams of that substance

**Example 5:** What is the mass in grams of 4.33 mole of O atoms?

We know that the atomic weight of O = 16 amu

Apply definition of Molar Mass: 1 mole O = 16 g O

Thus, \[
\begin{array}{c|c}
4.33 \text{ mole O atoms} & 16.0 \text{g O atoms} \\
1 \text{ mole O atoms} & 1 \text{ mole O atoms}
\end{array}
\]

\[= 69.3 \text{g O atoms} \]
**Example 6:** What is the mass in grams of 3.25 mole of CO molecules?

Must first calculate FW of CO

\[
\begin{align*}
1 \times 12.0 \text{ amu} &= 12.0 \text{ amu} \\
1 \times 16.0 \text{ amu} &= 16.0 \text{ amu} \\
\text{FW} &= 28.0 \text{ amu}
\end{align*}
\]

Now, apply definition of Molar Mass

1 mole of CO molecules = 28.0g CO molecules

Solve problem:

\[
\begin{align*}
3.25 \text{ mole CO molecules} & \quad 28.0 \text{g CO molecules} \\
1 \text{ mole CO molecules} & \quad 91.0 \text{g CO molecules}
\end{align*}
\]

**Example 7:** Calculate the number of Cu atoms present in a 37.83 sample of Cu.

<table>
<thead>
<tr>
<th>Grams Cu</th>
<th>Molar Mass</th>
<th>mole Cu</th>
<th>Avogadro’s #</th>
</tr>
</thead>
<tbody>
<tr>
<td>37.83 g Cu</td>
<td>63.55g Cu</td>
<td>1 mole Cu</td>
<td>6.02 X 10^23 atoms Cu</td>
</tr>
</tbody>
</table>

6.02 X 10^23 atoms Cu = 3.585 X 10^23 atoms Cu

Other info:

**Empirical formula (simplest formula)** – smallest whole number ratio of atoms present in a formula unit of a compound

**Molecular Formula (true formula)** – gives the actual number of atoms present in a formula unit of a compound

Molecular Formula = whole # multiplier X empirical formula

\[
\begin{align*}
\text{N}_2\text{F}_4 &= 2 \times \text{NF}_2
\end{align*}
\]