Chapter 3: CHEMICAL REACTIONS AND EARTH’S COMPOSITION
(Topics to Review)

3.2 The Mole

Avogadro’s Number \((N_A) = 6.022 \times 10^{23}\) (to 4 sig figs)

How big is this?
- If \(6.022 \times 10^{23}\) hydrogen atoms were laid side by side, the total length would encircle the earth about a million times.
- The mass of \(6.022 \times 10^{23}\) Olympic shotput balls is about equal to the mass of the Earth.
- The volume of \(6.022 \times 10^{23}\) softballs is about equal to the volume of the Earth.

1 mole (abbreviated mol) = \(6.022 \times 10^{23}\) entities

Similar to: 1 dozen = 12 entities:

- 1 dozen doughnuts = 12 doughnuts
- 1 mole of doughnuts = \(6.022 \times 10^{23}\) doughnuts

Ex. 1 How many eggs are in 3 dozen eggs?  
\[3 \times 12 \text{ eggs} = 36 \text{ eggs}\]

Ex. 2 How many eggs are in 3 moles of eggs?  
\[3 \times 6.022 \times 10^{23} \text{ eggs} = 1.807 \times 10^{24} \text{ eggs}\]

Ex. 3 How many moles of carbon contain \(7.25 \times 10^{24}\) carbon atoms?  
\[
\frac{7.25 \times 10^{24} \text{ C atoms}}{6.022 \times 10^{23} \text{ C atoms}} \times \frac{1 \text{ mol C}}{6.022 \times 10^{23} \text{ C atoms}} = 12.0 \text{ mol C}
\]

Atomic weights and molar masses:
- The mass of 1 C atom (on average) is 12.01 amu
- The mass of 1 mole of C atoms is 12.01 g (or 12.01 g/mol)

→ 1 mole (\(6.022 \times 10^{23}\)) is the amount of atoms of any element that has a mass in grams equal to the mass of ONE atom in amu.

→ The atomic masses reported in the Periodic Table give the atomic weight (or molecular/formula weight for compounds) in amu and the molar mass in g/mol.
MOLAR MASS (MM): Mass in grams of 1 mole of any element/compound
- To obtain, multiply the molar mass of each element by the number of each present, then add up all the constituent parts.

Example: Determine the molar mass of each of the following compounds:

a. \( \text{O}_2: \) 2 (molar mass of O) = 2 (16.00 g/mol) = \( 32.00 \text{ g/mol} \)

b. \( \text{H}_3\text{PO}_4: \)

\[
\begin{align*}
\text{H:} & \quad 3(1.008) \\
\text{P:} & \quad 30.97 \\
\text{O:} & \quad 6(16.00)
\end{align*}
\]

\( 97.994 \text{ g/mol} \)

c. \( \text{Al}_2(\text{SO}_4)_3: \)

\[
\begin{align*}
\text{Al:} & \quad 2(26.98) \\
\text{S:} & \quad 3(32.07) \\
\text{O:} & \quad 12(16.00)
\end{align*}
\]

\( 310.18 \text{ g/mol} \)

Note: Don’t worry about rounding molar masses since they rarely limit the number of sig figs for your final calculated answer.

Mole Calculations

Ex. 1 How many moles of Ne are in 0.500 g Ne?

\[
0.500 \text{ g Ne} \times \frac{1 \text{ mol Ne}}{20.18 \text{ g Ne}} = 0.0248 \text{ mol Ne}
\]

Ex. 2 How many Ne atoms are in 0.500 g of Ne?

\[
0.500 \text{ g Ne} \times \frac{1 \text{ mol Ne}}{20.18 \text{ g Ne}} \times \frac{6.022 \times 10^{23} \text{ Ne atoms}}{1 \text{ mol Ne}} = 1.49 \times 10^{22} \text{ Ne atoms}
\]

Ex. 3 How many moles of \( \text{CO}_2 \) are in 0.500 g of \( \text{CO}_2 \)?

\[
0.500 \text{ g } \text{CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} = 0.0114 \text{ mol CO}_2
\]

Ex. 4 How many \( \text{CO}_2 \) molecules are in 0.500 g of \( \text{CO}_2 \)?

\[
0.500 \text{ g } \text{CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{6.022 \times 10^{23} \text{ CO}_2 \text{ molecules}}{1 \text{ mol CO}_2} = 6.84 \times 10^{21} \text{ CO}_2 \text{ molecules}
\]

Ex. 5 How many oxygen atoms are in 0.500 g of \( \text{CO}_2 \)?

\[
0.500 \text{ g } \text{CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{6.022 \times 10^{23} \text{ CO}_2 \text{ molecules}}{1 \text{ mol CO}_2} \times \frac{2 \text{ O atoms}}{\text{CO}_2 \text{ molecule}} = 1.37 \times 10^{22} \text{ O atoms}
\]
3.5 Percent Composition of Compounds

**Percent composition by mass:** The mass of one element in a compound divided by the mass of the entire compound.

Steps to determine percentage composition:
1. Calculate the mass of each individual element in the compound.
2. Add up all the masses of each element to get the total mass of compound.
3. Divide the mass of each individual element by the total mass of compound.

Ex. 1 What is the percent composition of S and O in SO₃?

\[
S: \frac{32.07}{80.07} \times 100\% = 40.05\% \text{ S} \\
O: \frac{48.00}{80.07} \times 100\% = 60.00\% \text{ O}
\]

Ex. 2 What is the percent composition of Ca, P, and O in Ca₃(PO₄)₂?

\[
\begin{align*}
\text{Ca: } & \frac{3(40.08)}{3(100.18)} \times 100\% = 120.24\% \\
\text{P: } & \frac{2(30.97)}{3(100.18)} \times 100\% = 61.97\% \\
\text{O: } & \frac{8(16.00)}{3(100.18)} \times 100\% = 19.79\%
\end{align*}
\]

3.6 Determining the Formula of a Compound

**Empirical Formula:** Simplest whole-number ratio of atoms in a compound — where the word “empirical” means derived from experiment.

**Molecular Formula:** Chemical formula of a compound that expresses the actual number of atoms present in one molecule. — The molecular formula will either be exactly the same or some multiple of the empirical formula!

Fill in the table below:

<table>
<thead>
<tr>
<th></th>
<th>glucose, C₆H₁₂O₆</th>
<th>caffeine, C₆H₁₀N₄O₂</th>
<th>acrylonitrile, C₃H₅N</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Empirical Formula</strong></td>
<td>CH₂O</td>
<td>C₃H₅N₂O</td>
<td>C₃H₅N</td>
</tr>
<tr>
<td><strong>Molecular Formula</strong></td>
<td>C₆H₁₂O₆</td>
<td>C₆H₁₀N₄O₂</td>
<td>C₃H₃N</td>
</tr>
</tbody>
</table>
Guidelines for Determining the Empirical Formula of a Compound

1. Find the # of moles of each element in the compound.

2. Divide each # of moles of each element by smallest # of moles to get ratio of atoms.

3. Get a whole number ratio for all atoms in the compound:
   - If within 0.1 of a whole number, round to that whole number
   - If any ratio ends close to 0.5 → multiply ALL subscripts by 2
   - If any ratio ends close to 0.33 or 0.66 → multiply ALL subscripts by 3

Ex. 1: Determine the empirical formula for nickel oxide if a sample of nickel oxide consists of 17.74 g of nickel and 7.26 g of oxygen?

\[ 17.74 \text{ g Ni} \times \frac{1 \text{ mol Ni}}{58.69 \text{ g Ni}} = 0.3023 \text{ mol Ni} \]
\[ 7.26 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.454 \text{ mol O} \]

Comparison:
\[ \text{Ni}_0.3023 \text{ O}_{0.454} \]

Multiplying by 2:
\[ \text{Ni}_0.3023 \times 2 \text{ O}_{0.454} \times 2 = \text{Ni}_1 \text{ O}_{1.50} \]

Empirical formula: \( \text{Ni}_2\text{O}_3 \)

Name: nickel (III) oxide

Ex. 2: Ascorbic acid is known more commonly as Vitamin C. Ascorbic acid is a hydrocarbon derivative, a compound that consists of carbon, hydrogen, and oxygen. If a 25.00 g sample of ascorbic acid contains 10.23 g of carbon and 1.14 g of hydrogen, determine the empirical formula for ascorbic acid.

\[ 10.23 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.8518 \text{ mol C} \]
\[ 1.14 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 1.13 \text{ mol H} \]
\[ 13.63 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.8519 \text{ mol O} \]

Comparison:
\[ \text{C}_{0.8518} \text{ H}_{1.13} \text{ O}_{0.8519} \]

Multiplying by 3:
\[ \text{C}_{0.8518} \times 3 \text{ H}_{1.13} \times 3 \text{ O}_{0.8519} \times 3 = \text{C}_3 \text{ H}_{3.39} \text{ O}_{3} \]

\[ \text{C}_3\text{H}_3\text{O}_3 \Rightarrow \text{C}_3\text{H}_4\text{O}_3 \]
Determining the Molecular Formula of a Compound

1. For the Molecular Formula, you will be given the molar mass of the compound, and you need to calculate the molar mass of the empirical formula.

2. Divide the molar mass of the compound by the molar mass of the empirical formula to get the factor by which to multiply each subscript in the empirical formula.

Ex. 3: If the ascorbic acid in Ex. 2 on the previous page has a molar mass of 176.1 g/mol, determine its molecular formula.

\[
\begin{align*}
\text{C: } & 3(12.01) \\
\text{H: } & 4(1.008) \\
\text{O: } & 3(16.00) \\
\hline
\text{Total Molar Mass: } & 88.062 \text{ g} \\
\frac{176.1 \text{ g}}{88.062 \text{ g}} & \approx 2 \\
\Rightarrow & \text{ } C_3H_4O_3 \\
\times 2 & \Rightarrow C_6H_8O_6
\end{align*}
\]

Empirical and Molecular Formulas from Percent Composition:

1. Assume 100.0 g of the compound is present, so change percent units to grams.

2. Follow the same steps for determining empirical formula and molecular formula.

Ex. 4: Quinine is used as an antimalarial drug. An analysis of quinine indicates the compound consists of 74.03% carbon, 7.47% hydrogen, 8.64% nitrogen, and the remainder is oxygen. If quinine’s molar mass is 325 g/mol, determine the empirical and molecular formulas for quinine.

\[
\begin{align*}
74.03 \text{ g C} & \times \frac{\text{mol C}}{12.01 \text{ g C}} = 6.164 \text{ mol C} = \frac{6.164 \text{ mol C}}{0.016 \text{ mol C}} = 10.0 \\
7.47 \text{ g H} & \times \frac{\text{mol H}}{1.008 \text{ g H}} = 7.41 \text{ mol H} = \frac{7.41 \text{ mol H}}{0.016 \text{ mol H}} = 12.0 \\
8.64 \text{ g N} & \times \frac{\text{mol N}}{14.01 \text{ g N}} = 0.617 \text{ mol N} = \frac{0.617 \text{ mol N}}{0.016 \text{ mol N}} = 1.00 \\
9.86 \text{ g O} & \times \frac{\text{mol O}}{16.00 \text{ g O}} = 0.616 \text{ mol O} = \frac{0.616 \text{ mol O}}{0.016 \text{ mol O}} = 1
\end{align*}
\]

\[
\text{empirical formula } = \text{C}_{10}\text{H}_{12}\text{NO} \times 2 \rightarrow \text{C}_{20}\text{H}_{24}\text{N}_{2}\text{O}_{2}
\]
3.7 CHEMICAL EQUATIONS: formulas and symbols describing a chemical reaction

\[ \text{A} + \text{B} \rightarrow \text{C} + \text{D} \]

 reactants                           products
 starting materials       substance(s) resulting
 from chemical reaction

The Physical state of all the reactants and products are also indicated:
(s) = solid     (l) = liquid     (g) = gas
(aq) = aqueous (ions or compounds dissolved in water)

Example: \[ 2 \text{Al (s)} + 6 \text{HCl (aq)} \rightarrow 2 \text{AlCl}_3 \text{(aq)} + 3 \text{H}_2 \text{(g)} \]

3.8 Balancing Chemical Equations

**Coefficient:** Whole #s in front of each reactant/product, indicating the number of each present.

**Subscript:** Whole # after each element in a compound, indicating # of each element present.

**GUIDELINES for Balancing by Inspection**

1. Count the # of each element on both sides of the equation.

2. Change the coefficients (NEVER the subscripts) to get the same # of elements on both sides of the equation
   - Balance the equation using the following order:
     - Metals
     - Polyatomic ions – Balance as a whole!
     - Hydrogen
     - Carbon
     - Oxygen
     - All other atoms

Examples:

\[ \frac{4}{2} \text{Al (s)} + 3 \text{O}_2 \text{(g)} \rightarrow 2 \text{Al}_2\text{O}_3 \text{(s)} \]

\[ 2 \text{KClO}_3 \text{(s)} \xrightarrow{\Delta} 2 \text{KCl (s)} + 3 \text{O}_2 \text{(g)} \]

\[ \frac{2}{2} \text{C}_5\text{H}_12 \text{(l)} + 8 \text{O}_2 \text{(g)} \xrightarrow{\Delta} \frac{16}{6} \text{H}_2\text{O (g)} + \frac{5}{4} \text{CO}_2 \text{(g)} \]

\[ 2 \text{HCl(aq)} + 8 \text{CaCO}_3\text{(s)} \rightarrow \frac{8}{4} \text{CaCl}_2\text{(aq)} + \frac{16}{6} \text{H}_2\text{O(l)} + \frac{5}{4} \text{CO}_2\text{(g)} \]

Treat polyatomic ions as ONE UNIT—Do not break them up into atoms!

\[ \frac{3}{3} \text{Al}_2\text{(SO}_4\text{)}_3 \text{(aq)} + \frac{3}{2} \text{Ba(NO}_3\text{)}_2 \text{(aq)} \rightarrow \frac{3}{2} \text{BaSO}_4 \text{(s)} + \frac{3}{2} \text{Al(NO}_3\text{)}_3 \text{(aq)} \]

\[ \frac{3}{2} \text{Sr(OH)}_2 \text{(aq)} + \frac{2}{1} \text{H}_3\text{PO}_4 \text{(aq)} \rightarrow \frac{6}{1} \text{H}_2\text{O (l)} + \frac{3}{4} \text{Sr}_3\text{(PO}_4\text{)}_2 \text{(s)} \]