UNIT 4: THE MOLE AND CHEMICAL COMPOSITION

Chapter 7: The Mole and Chemical Composition

7.1 & 7.2: Avogadro’s Number and Molar Conversion & Relative Atomic Mass and Chemical Formulas

**Mole (mol):** - a group of atoms or molecules numbered \(6.022 \times 10^{23}\) (Avogadro’s Number, \(N_A\))

**Examples:**
1. mol of carbon (C) = \(6.022 \times 10^{23}\) carbon atoms = 12.01 g (same as the amu)
2. mol of fluorine (\(F_2\)) = \(6.022 \times 10^{23}\) fluorine molecules = 38.00 g (include subscripts with amu)

**Example 1:** Find the number of molecules in 3.50 moles of platinum.

\[
\text{1 mol Pt} = 6.022 \times 10^{23} \text{ Pt atoms} \\
3.50 \text{ mol} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}} = 2.11 \times 10^{24} \text{ atoms of Pt}
\]

**Example 2:** How many moles of sucrose are there if there are \(3.54 \times 10^{25}\) molecules of \(\text{C}_{12}\text{H}_{22}\text{O}_{11}\)?

\[
\text{1 mol } \text{C}_{12}\text{H}_{22}\text{O}_{11} = 6.022 \times 10^{23} \text{ C}_{12}\text{H}_{22}\text{O}_{11} \text{ molecules} \\
3.54 \times 10^{25} \text{ molecules} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} = 58.8 \text{ mol of } \text{C}_{12}\text{H}_{22}\text{O}_{11}
\]

**Example 3:** Find the number of \(\text{Br}^-\) ion in 0.65 mol of \(\text{MgBr}_2\).

\[
\text{1 mol MgBr}_2 = 6.022 \times 10^{23} \text{ MgBr}_2 \text{ formula units} \\
0.65 \text{ mol} \times \frac{6.022 \times 10^{23} \text{ units}}{1 \text{ mol}} \times \frac{2 \text{ Br}^- \text{ ion}}{1 \text{ MgBr}_2 \text{ unit}} = 7.8 \times 10^{23} \text{ Br}^- \text{ ions}
\]

**Molar Mass (g/mol):** - sometimes refer to as **formula mass, molecular weight**, is the mass per one mole of atoms or molecules.
- molar mass of a mono-atomic element is the same as the atomic mass.
- molar mass of a compound, diatomic element, or polyatomic element is the same as the combine atomic masses of all atoms in the molecule. (Be careful counting number of atoms of a polyatomic ion in parenthesis.)

**Example 4:** Find the molar mass of the following.

a. potassium

\[\text{K} = 39.10 \text{ g/mol}\]

b. phosphorus (\(P_4\))

\[P_4 = 4 \times 30.97 \text{ g/mol} \Rightarrow P_4 = 123.88 \text{ g/mol}\]

c. glucose (\(\text{C}_6\text{H}_{12}\text{O}_6\))

\[\text{C}_6\text{H}_{12}\text{O}_6 = (6 \times 12.01) + (12 \times 1.01) + (6 \times 16.00) \Rightarrow \text{C}_6\text{H}_{12}\text{O}_6 = 180.18 \text{ g/mol}\]

d. sodium chromate – \(\text{Na}_2\text{CrO}_4\)

\[\text{Na}_2\text{CrO}_4 = (2 \times 22.99) + 52.00 + (4 \times 16.00) \Rightarrow \text{Na}_2\text{CrO}_4 = 161.98 \text{ g/mol}\]

e. iron (III) nitrite – \(\text{Fe(NO}_2)_3\)

\[\text{Fe(NO}_2)_3 = 55.85 + (3 \times 14.01) + (6 \times 16.00) \Rightarrow \text{Fe(NO}_2)_3 = 193.88 \text{ g/mol}\]
Converting Between Mass and Number of Particles:
1. Find the Molar Mass (in g/mol).
2. Set up a Conversion Factor between Molar Mass and the Avogadro’s Number and Solve.

**Example 5**: Calculate the mass of \(7.50 \times 10^{23}\) molecules of sulphur (S\(_8\)).

\[
S_8 = 8 \times 32.07 \text{ g/mol} = 256.56 \text{ g/mol} \Rightarrow 256.56 \text{ g of } S_8 = 1 \text{ mol } S_8 = 6.022 \times 10^{23} S_8 \text{ molecules}
\]

\[
7.50 \times 10^{23} \text{ molecules} \times \frac{256.56 \text{ g}}{6.022 \times 10^{23} \text{ molecules}} = 320 \text{ g of } S_8
\]

**Example 6**: Determine the number of phosphorus atoms in 60.0 g of solid phosphorus (P\(_4\)).

\[
P_4 = 4 \times 30.97 \text{ g/mol} = 123.88 \text{ g/mol} \Rightarrow 123.88 \text{ g of } P_4 = 1 \text{ mol } P_4 = 6.022 \times 10^{23} P_4 \text{ molecules}
\]

There are four phosphorus atoms in one molecular unit of P\(_4\).

\[
60.0 \text{ g} \times \frac{6.022 \times 10^{23} \text{ molecules}}{123.88 \text{ g}} \times \frac{4 \text{ P atoms}}{1 \text{ P}_4 \text{ molecule}} = 1.17 \times 10^{24} \text{ phosphorus atoms}
\]

**Example 7**: Determine the number of Na\(^+\) ion in 92.3 g of sodium phosphate – Na\(_3\)PO\(_4\).

\[
\text{Na}_3\text{PO}_4 = (3 \times 22.99) + 30.97 + (4 \times 16.00) = 163.94 \text{ g/mol} \Rightarrow 163.94 \text{ g of } \text{Na}_3\text{PO}_4 = 1 \text{ mol } \text{Na}_3\text{PO}_4 = 6.022 \times 10^{23} \text{ Na}_3\text{PO}_4 \text{ formula units}
\]

There are three Na\(^+\) ions in one formula unit of Na\(_3\)PO\(_4\).

\[
92.3 \text{ g} \times \frac{6.022 \times 10^{23} \text{ units}}{163.94 \text{ g}} \times \frac{3 \text{ Na}^+ \text{ ions}}{1 \text{ Na}_3\text{PO}_4 \text{ units}} = 1.02 \times 10^{24} \text{ Na}^+ \text{ ions}
\]

**Converting between Mass and Moles**: (need to find Molar Mass first!)

\[
\text{Moles (mol)} = \frac{\text{Mass (g)}}{\text{Molar Mass (g/mol)}} \quad n = \frac{m}{M}
\]

\[
n = \text{moles} \quad m = \text{mass} \quad M = \text{Molar mass}
\]

**Example 8**: Calculate the number of moles for 20.0 g of ethanol.

\[
\text{Ethanol} = \text{C}_2\text{H}_5\text{OH} = 2(12.01) + 6(1.01) + 16.00 \quad M = 46.08 \text{ g/mol}
\]

\[
n = \frac{m}{M} = \frac{20.0 \text{ g}}{46.08 \text{ g/mol}} = 0.434 \text{ mol}
\]

**Example 9**: Determine the mass of 0.852 mol of aluminum carbonate

\[
\text{Aluminum carbonate} = \text{Al}_2(\text{CO}_3)_3 = 2(26.98) + 3(12.01) + 9(16.00) \quad M = 233.99 \text{ g/mol}
\]

\[
n = \frac{m}{M} \quad m = nM = (0.852 \text{ mol})(233.99 \text{ g/mol}) = 199 \text{ g}
\]
Atomic Mass: - sometimes called atomic weight.
- the mass of the atom in atomic mass unit (amu).
- 1 amu = exactly one-twelfth the mass of one carbon-12 atom ≈ 1.67 × 10^{-27} kg.

Average Atomic Mass: - Average Mass of an atom and its isotopes after accounting their proportions of abundance (as stated on the Periodic Table of Elements).

Relative Abundance: - the relative proportion of various isotopes of an element.

Relative Percentage Abundance: - the relative proportion of various isotopes of an element in percentage.

\[
\text{Average Atomic Mass} = (\text{Relative Abundance of Isotope A})(\text{Mass Number of Isotope A}) + (\text{Relative Abundance of Isotope B})(\text{Mass Number of Isotope B}) + (\text{Relative Abundance of Isotope C})(\text{Mass Number of Isotope C}) + \ldots
\]

Example 10: Iron has three natural isotopes. If $^{56}\text{Fe}$ and $^{57}\text{Fe}$ have relative percentage of abundance at 91.754% and 2.401% respectively, and the rest is $^{54}\text{Fe}$, calculate the average atomic mass of iron?

\[
\% \text{ abundance of } ^{54}\text{Fe} = 100\% - 91.754\% - 2.401\% = 5.845\%
\]
\[
\text{Average amu of Iron} = (0.05845)(54) + (0.91754)(56) + (0.02401)(57)
\]
\[
\text{Average amu of Iron} = 55.91 \text{ amu}
\]

Assignment

7.1 pg. 229 #1 to 5 (Practice); pg. 231 #1 to 4 (Practice); pg. 232 #1 to 3 (Practice); pg. 233 #1 to 13
7.2 pg. 239–240 #1 to 4 (Practice); pg. 236 #1 and 2 (Practice); pg. 240 #1, 3 to 12, 14 to 16

7.3: Formulas and Percentage Composition

Percentage Composition: - also called mass percent or percentage mass.
- it is the mass percentage of each element in a compound.

For Compound $A_xB_yC_z$ with its Total Mass ($m$), the Mass Percentages are:

\[
\%A = \frac{m_A}{m} \times 100\% \quad \%B = \frac{m_B}{m} \times 100\% \quad \%C = \frac{m_C}{m} \times 100\%
\]

For Compound $A_xB_yC_z$ with its Molar Mass ($M$), the Mass Percentages are:

\[
\%A = \frac{x(M_A)}{M} \times 100\% \quad \%B = \frac{y(M_B)}{M} \times 100\% \quad \%C = \frac{z(M_C)}{M} \times 100\%
\]
Example 1: Calculate the mass percentage of sodium chromate.

\[ \text{Na}_2\text{CrO}_4 \quad M = 161.98 \text{ g/mol} \]

Assume we have 161.98 g (1 mole) of Na\(_2\)CrO\(_4\), there are 2 moles of Na, 1 mole of Cr and 4 moles of O:

\[
\% \text{Na} = \frac{2 \text{ mol} \times (22.99 \text{ g/mol})}{161.98 \text{ g}} \times 100\% = 28.39 \%
\]

\[
\% \text{Cr} = \frac{1 \text{ mol} \times (52.00 \text{ g/mol})}{161.98 \text{ g}} \times 100\% = 32.10 \%
\]

\[
\% \text{O} = \frac{4 \text{ mol} \times (16.00 \text{ g/mol})}{161.98 \text{ g}} \times 100\% = 39.51 \%
\]

**Empirical Formula:** - the simplest ratio between the elements in a chemical formula.

**Molecular Formula:** - the actual chemical formula of a compound.

\[
\text{Molecular Formula} = (\text{Empirical Formula})_n
\]

where \( n = \text{natural number} \)

Example:

\[
\text{C}_6\text{H}_12\text{O}_6 \quad \text{CH}_2\text{O}
\]

Molecular Formula for Glucose \( \rightarrow \) Empirical Formula

Note: Knowing the mass percentages of a compound allow us to find the empirical formula. To know the molecular formula, we must also know the molar mass.

Example 2: A compound has an empirical formula of CH\(_3\)N has a molar mass of 116.12 g/mol. What is the molecular formula of the compound?

**Empirical Formula** = CH\(_3\)N \( (29.05 \text{ g/mol}) \)

\[
\frac{\text{Actual Molar Mass}}{\text{Empirical Molar Mass}} = \frac{116.12 \text{ g/mol}}{29.05 \text{ g/mol}} \approx 4
\]

Molecular Formula = Empirical Formula \( \times 4 = (\text{CH}_3\text{N}) \times 4 \)

\[
\text{Molecular Formula} = \text{C}_4\text{H}_{12}\text{N}_4
\]

Example 3: Cobalt (II) nitrate is a hydrate with a chemical formula of Co(NO\(_3\))\(_2\) \( \cdot \) \(x\)H\(_2\)O. When the 2.45 g of hydrate is heated, 1.54 g of residual is left behind. Determine the number of hydrate unit for cobalt (II) nitrate.

Mass of Co(NO\(_3\))\(_2\) = 1.54 g

n = \[
\frac{1.54 \text{ g}}{182.95 \text{ g/mol}} = 0.0084176004 \text{ mol Co(NO}_3\text{)}_2
\]

Mass of H\(_2\)O released = 2.45 g – 1.54 g = 0.91 g

n = \[
\frac{0.91 \text{ g}}{18.02 \text{ g/mol}} = 0.0504994451 \text{ mol H}_2\text{O}
\]

\[
\frac{n_{\text{H}_2\text{O}}}{n_{\text{Co(NO}_3\text{)}_2}} = \frac{0.0504994451 \text{ mol H}_2\text{O}}{0.0084176004 \text{ mol Co(NO}_3\text{)}_2} \approx \frac{6 \text{ mol H}_2\text{O}}{1 \text{ mol Co(NO}_3\text{)}_2}
\]

\[
\text{Molecular Formula} = \text{Co(NO}_3\text{)}_2 \cdot 6 \text{ H}_2\text{O}
\]
Example 4: Vitamin C has a molar mass of 176.14 g/mol and contains carbon, hydrogen, and oxygen atoms. If the % mass of carbon and oxygen are 40.91% and 54.50% respectively, determine the empirical and molecular formula of vitamin C.

\[ \% \text{C} = 40.91\% \quad \% \text{O} = 54.50\% \quad \% \text{H} = 100\% - 40.91\% - 54.50\% = 4.59\% \]

Assume 100 g of Vitamin C. Then, there are

\[ m_{\text{C}} = 100 \text{ g} \times 40.91\% = 40.91 \text{ g} \quad m_{\text{O}} = 100 \text{ g} \times 54.50\% = 54.50 \text{ g} \quad m_{\text{H}} = 100 \text{ g} \times 4.59\% = 4.59 \text{ g} \]

\[ n_{\text{C}} = \frac{40.91 \text{ g}}{12.01 \text{ g/mol}} = 3.40632806 \text{ mol}_\text{C} \]

\[ n_{\text{O}} = \frac{54.50 \text{ g}}{16.00 \text{ g/mol}} = 3.40625 \text{ mol}_\text{O} \]

\[ \frac{n_{\text{C}}}{n_{\text{O}}} = \frac{3.40632806 \text{ mol C}}{3.40625 \text{ mol O}} \approx 1 \text{ mol C} \quad \frac{n_{\text{H}}}{n_{\text{O}}} = \frac{4.553571429 \text{ mol H}}{3.40625 \text{ mol O}} \approx 1.33 = 4 \text{ mol H} \]

Combine Ratios

\[ n_{\text{C}} : n_{\text{O}} = 1 : 1 \]

Empirical Formula = \( \text{C}_3\text{H}_4\text{O}_3 \) (88.06 g/mol)

Molecular Formula = Empirical Formula \( \times 2 \)

Molecular Formula = \( \text{C}_6\text{H}_8\text{O}_6 \)

Another Method may be used where the Actual Molar Mass becomes the Mass of Vitamin used. Then, the Mole of each Atom is calculated to determine the Molecular Formula first.

\[ n_{\text{C}} = \frac{40.91\% \times 176.14 \text{ g}}{12.01 \text{ g/mol}} \approx 6.00 \text{ mol}_\text{C} \]

\[ n_{\text{H}} = \frac{4.59\% \times 176.14 \text{ g}}{1.01 \text{ g/mol}} \approx 8.00 \text{ mol}_\text{H} \]

\[ n_{\text{O}} = \frac{54.50\% \times 176.14 \text{ g}}{16.00 \text{ g/mol}} \approx 6.00 \text{ mol}_\text{O} \]

Molecular Formula \( \text{(C}_6\text{H}_8\text{O}_6 \) will be found first, then the Empirical Formula \( \text{(C}_3\text{H}_4\text{O}_3 \) will be stated.

Assignment

7.3 pg. 243 #1 to 4 (Practice); pg. 25 #1 to 3 (Practice); pg. 248 #1 to 5 (Practice); pg. 248 #1 to 10
Chapter 7 Review pg. 251–254 #19 to 66