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 **6.022×10^{23} (*Avogadro's Number, N_A*)**

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

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

1 mol Pt = 6.022×10^{23} Pt atoms

$$3.50\ \cancel{\text{mol}} \times \frac{6.022\times10^{23}\ \text{atoms}}{1\ \cancel{\text{mol}}} = \mathbf{2.11 \times 10^{24}\ \text{atoms of Pt}}$$

Example 2: How many moles of sucrose are there if there are 3.54×10^{25} molecules of $C_{12}H_{22}O_{11}$?

1 mol $C_{12}H_{22}O_{11}$ = 6.022×10^{23} $C_{12}H_{22}O_{11}$ molecules

$$3.54 \times 10^{25}\ \cancel{\text{molecules}} \times \frac{1\ \text{mol}}{6.022\times10^{23}\ \cancel{\text{molecules}}} = \mathbf{58.8\ \text{mol of}\ C_{12}H_{22}O_{11}}$$

Example 3: Find the number of Br^- ion in 0.65 mol of $MgBr_2$.

1 mol $MgBr_2$ = 6.022×10^{23} $MgBr_2$ formula units

$$0.65\ \cancel{\text{mol}} \times \frac{6.022\times10^{23}\ \cancel{\text{units}}}{1\ \cancel{\text{mol}}} \times \frac{2\ Br^-\ \text{ion}}{1\ MgBr_2\ \cancel{\text{unit}}} = \mathbf{7.8 \times 10^{23}\ Br^-\ \text{ions}}$$

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

K = 39.10 g/mol

b. phosphorus (P_4)

$P_4 = 4 \times 30.97$ g/mol

P_4 = 123.88 g/mol

c. glucose ($C_6H_{12}O_6$)

$C_6H_{12}O_6 = (6 \times 12.01) + (12 \times 1.01) + (6 \times 16.00)$

$C_6H_{12}O_6$ = 180.18 g/mol

d. sodium chromate – Na_2CrO_4

$Na_2CrO_4 = (2 \times 22.99) + 52.00 + (4 \times 16.00)$

Na_2CrO_4 = 161.98 g/mol

e. iron (III) nitrite – $Fe(NO_2)_3$

$Fe(NO_2)_3 = 55.85 + (3 \times 14.01) + (6 \times 16.00)$

$Fe(NO_2)_3$ = 193.88 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 7.50×10^{23} molecules of sulphur (S_8).

$S_8 = 8 \times 32.07$ g/mol $= 256.56$ g/mol ➔ 256.56 g of S_8 = 1 mol S_8 = 6.022×10^{23} S_8 molecules

$$7.50 \times 10^{23}\ \cancel{\text{molecules}} \times \frac{256.56\ \text{g}}{6.022\times10^{23}\ \cancel{\text{molecules}}} = \textbf{320 g of } S_8$$

Example 6: Determine the number of phosphorus ***atoms*** in 60.0 g of solid phosphorus (P_4).

$P_4 = 4 \times 30.97$ g/mol $= 123.88$ g/mol ➔ 123.88 g of P_4 = 1 mol P_4 = 6.022×10^{23} P_4 molecules

There are four phosphorus atoms in one molecular unit of P_4.

$$60.0\ \cancel{\text{g}} \times \frac{6.022\times10^{23}\ \cancel{\text{molecules}}}{123.88\ \cancel{\text{g}}} \times \frac{4\ \text{P atoms}}{1\ \text{P}_4\ \cancel{\text{molecule}}} = \mathbf{1.17 \times 10^{24}}\ \textbf{phosphorus atoms}$$

Example 7: Determine the number of Na^+ ion in 92.3 g of sodium phosphate – Na_3PO_4.

$Na_3PO_4 = (3 \times 22.99) + 30.97 + (4 \times 16.00) = 163.94$ g/mol

➔ 163.94 g of Na_3PO_4 = 1 mol Na_3PO_4 = 6.022×10^{23} Na_3PO_4 formula units

There are three Na^+ ions in one formula unit of Na_3PO_4.

$$92.3\ \cancel{\text{g}} \times \frac{6.022\times10^{23}\ \cancel{\text{units}}}{163.94\ \cancel{\text{g}}} \times \frac{3\ \text{Na}^+\ \text{ions}}{1\ \text{Na}_3\text{PO}_4\ \cancel{\text{units}}} = \mathbf{1.02 \times 10^{24}\ Na^+}\ \textbf{ions}$$

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

$$\textbf{Moles (mol)} = \frac{\textbf{Mass (g)}}{\textbf{Molar Mass (g/mol)}} \qquad n = \frac{m}{M}$$

***n* = moles** ***m* = mass** ***M* = Molar mass**

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

Ethanol = C_2H_5OH = 2(12.01) + 6(1.01) + 16.00 $M = 46.08$ g/mol

$$n = \frac{m}{M} = \frac{20.0\ \cancel{\text{g}}}{46.08\ \cancel{\text{g}}/\text{mol}} \qquad \mathbf{n = 0.434\ mol}$$

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

Aluminum carbonate = $Al_2(CO_3)_3$ = 2(26.98) + 3(12.01) + 9(16.00) $M = 233.99$ g/mol

$$n = \frac{m}{M} \qquad m = nM = (0.852\ \cancel{\text{mol}})(233.99\ \text{g}/\cancel{\text{mol}}) \qquad \mathbf{m = 199\ g}$$

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 $\approx 1.67 \times 10^{-27}$ 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.

Average Atomic Mass = (Relative Abundance of Isotope A)(Mass Number of Isotope A)
+ (Relative Abundance of Isotope B)(Mass Number of Isotope B)
+ (Relative Abundance of Isotope C)(Mass Number of Isotope C) + …

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

% abundance of ^{54}Fe = 100% − 91.754% − 2.401% = **5.845%**

$$\text{Average amu of Iron} = \underbrace{(0.05845)(54)}_{5.845\% \text{ of Mass Number } 56} + \underbrace{(0.91754)(56)}_{91.754\% \text{ of Mass Number } 56} + \underbrace{(0.02401)(57)}_{2.401\% \text{ of Mass Number } 57}$$

Average amu of Iron = 55.91 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:

$$\%A = \frac{m_A}{m} \times 100\% \qquad \%B = \frac{m_B}{m} \times 100\% \qquad \%C = \frac{m_C}{m} \times 100\%$$

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

$$\%A = \frac{(x)(M_A)}{M} \times 100\% \qquad \%B = \frac{(y)(M_B)}{M} \times 100\% \qquad \%C = \frac{(z)(M_C)}{M} \times 100\%$$

Example 1: Calculate the mass percentage of sodium chromate.

Na_2CrO_4 $\quad M = 161.98$ g/mol

Assume we have 161.98 g (1 mole) of Na_2CrO_4, there are 2 moles of Na, 1 mole of Cr and 4 moles of O:

$$\%\ Na = \frac{(2\ \cancel{mol})(22.99\ \cancel{g}/\cancel{mol})}{161.98\ \cancel{g}} \times 100\% = 28.38622052\ \% \qquad \%\ Na = 28.39\ \%$$

$$\%\ Cr = \frac{(1\ \cancel{mol})(52.00\ \cancel{g}/\cancel{mol})}{161.98\ \cancel{g}} \times 100\% = 32.10272873\ \% \qquad \%\ Cr = 32.10\ \%$$

$$\%\ O = \frac{(4\ \cancel{mol})(16.00\ \cancel{g}/\cancel{mol})}{161.98\ \cancel{g}} \times 100\% = 39.51105075\ \% \qquad \%\ O = 39.51\ \%$$

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

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

$$\boxed{\textbf{Molecular Formula} = (\textbf{Empirical Formula})_n \quad \text{where } n = \text{natural number}}$$

Example: $C_6H_{12}O_6 \longleftarrow CH_2O$

Molecular Formula for Glucose — 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_3N has a molar mass of 116.12 g/mol. What is the molecular formula of the compound?

Empirical Formula = CH_3N (29.05 g/mol)

$$\frac{\text{Actual Molar Mass}}{\text{Emprical Molar Mass}} = \frac{116.12\ \cancel{g/mol}}{29.05\ \cancel{g/mol}} \approx 4$$

Molecular Formula = Empirical Formula × 4 = (CH_3N) × 4

Molecular Formula = $C_4H_{12}N_4$

Example 3: Cobalt (II) nitrate is a hydrate with a chemical formula of $Co(NO_3)_2 \bullet xH_2O$. 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

Mass of H_2O released = 2.45 g − 1.54 g = 0.91 g

$$n = \frac{1.54\ \cancel{g}}{182.95\ \cancel{g}/mol} = 0.0084176004\ mol\ Co(NO_3)_2$$

$$n = \frac{0.91\ \cancel{g}}{18.02\ \cancel{g}/mol} = 0.0504994451\ mol\ H_2O$$

$$\frac{n_{H_2O}}{n_{Co(NO_3)_2}} = \frac{0.0504994451\ mol\ H_2O}{0.0084176004\ mol\ Co(NO_3)_2} \approx \frac{6\ mol\ H_2O}{1\ mol\ Co(NO_3)_2}$$

Molecular Formula = $Co(NO_3)_2 \bullet 6\ H_2O$

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.

% C = 40.91% % O = 54.50% % H = 100% − 40.91% − 54.50% = 4.59%

Assume 100 g of Vitamin C. Then, there are

m_C = 100 g × 40.91% = 40.91 g m_O = 100 g × 54.50% = 54.50 g m_H = 100 g × 4.59% = 4.59 g

$$n_C = \frac{40.91\,\cancel{g}}{12.01\,\cancel{g}/mol} = 3.40632806\ mol_C \qquad n_H = \frac{4.59\,\cancel{g}}{1.008\,\cancel{g}/mol} = 4.553571429\ mol_H$$

$$n_O = \frac{54.50\,\cancel{g}}{16.00\,\cancel{g}/mol} = 3.40625\ mol_O$$

$$\frac{n_C}{n_O} = \frac{3.40632806\ mol\ C}{3.40625\ mol\ O} \approx \frac{1\ mol\ C}{1\ mol\ O} \qquad \frac{n_H}{n_O} = \frac{4.553571429\ mol\ H}{3.40625\ mol\ O} \approx 1.33 = \frac{4\ mol\ H}{3\ mol\ O}$$

$n_C : n_O = 1 : 1$ ⟵ **Combine Ratios** ⟶ $n_H : n_O = 4 : 3$

Empirical Formula = $C_3H_4O_3$ **(88.06 g/mol)**

$$\frac{\text{Actual Molar Mass}}{\text{Emprical Molar Mass}} = \frac{176.14\ \cancel{g/mol}}{88.06\ \cancel{g/mol}} = 2$$

Molecular Formula = Empirical Formula × 2

Molecular Formula = $C_6H_8O_6$

OR

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.

$$n_C = \frac{40.91\% \times 176.14\,\cancel{g}}{12.01\,\cancel{g}/mol} \approx 6.00\ mol_C \qquad n_H = \frac{4.59\% \times 176.14\,\cancel{g}}{1.01\,\cancel{g}/mol} \approx 8.00\ mol_H$$

$$n_O = \frac{54.50\% \times 176.14\,\cancel{g}}{16.00\,\cancel{g}/mol} \approx 6.00\ mol_O$$

Molecular Formula ($C_6H_8O_6$) will be found first, then the Empirical Formula ($C_3H_4O_3$) 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

