Ch. 9 Notes – THE MOLE

NOTE: Vocabulary terms are in **boldface and underlined**. Supporting details are in *italics.*

I. Measuring Matter

A. SI unit of chemical quantity = the **<u>mole</u>** (abbreviated *mol*)

1) 6.0221367×10^{23} *** to make it easier, we use 6.02×10^{23} ***
2) **<u>6.02×10^{23} = Avogadro's number</u>** (Amedeo Avogadro, 1776-1856)
3) the mass of 12 g of pure C-12
4) the **<u>mole</u>** is a *counting unit used in science to count particles*
   - **<u>representative particle (r.p.)</u>** = *atom, ion, molecule, formula unit*

- **<u>atom</u>** = *one symbol, no charge: Ne, U, Cs*
- **<u>ion</u>** = *one symbol with charge (monatomic) or more than one symbol with charge (polyatomic: Na^+, N^{3-}, $(C_2H_3O_2)^-$*
- **<u>molecule</u>** = *compound with all nonmetals (BM): CO, BF_3, Cl_2*
- **<u>formula unit</u>** **("fun")** = *compound with metal and nonmetal (BI, TI, OTHER): KI, Na_2SO_4*

1 MOLE = 6.02×10^{23} representative particles

6.02×10^{23} atoms = 1 mol atoms	6.02×10^{23} molecules = 1 mol molecules
$6.02\ 10^{23}$ ions = 1 mol ions	6.02×10^{23} formula units = 1 mol fun

B. **<u>stoichiometry</u>**—using balanced chemical equations to obtain info.

C. mole-to-r.p. and r.p.-to-mole example problems:

EXAMPLE 1) How many moles of Ca are in 9.00×10^{16} atoms of calcium?

$$9.00 \times 10^{16}\ \cancel{\text{atoms Ca}} \times \frac{1\ \text{mol Ca}}{6.02 \times 10^{23}\ \cancel{\text{atoms Ca}}} = \boxed{1.50 \times 10^{-7}\ \text{mol Ca}}$$

EXAMPLE 2) How many moles of sulfide ions are in 4.14×10^{30} sulfide ions?

$$4.14 \times 10^{30}\ \cancel{\text{ions } S^{2-}} \times \frac{1\ \text{mol S}}{6.02 \times 10^{23}\ \cancel{\text{ions } S^{2-}}} = \boxed{6.88 \times 10^{6}\ \text{mol } S^{2-}}$$

EXAMPLE 3) How many molecules are in 0.0221 mol oxygen gas?

$$0.0221\ \cancel{\text{mol } O_2} \times \frac{6.02 \times 10^{23}\ \text{molecules } O_2}{1\ \cancel{\text{mol } O_2}} = \boxed{1.33 \times 10^{22}\ \text{molecules } O_2}$$

D. Finding the number of atoms in a compound—look at the subscripts

EXAMPLE 4) How many hydrogen atoms are in 0.89 mol water?

$$0.89\ \cancel{\text{mol } H_2O} \times \frac{6.02 \times 10^{23}\ \cancel{\text{molecules } H_2O}}{1\ \cancel{\text{mol } H_2O}} \times \frac{2\ \text{H atoms}}{1\ \cancel{\text{molecule } H_2O}} = \boxed{1.1 \times 10^{24}\ \text{H atoms}}$$

EXAMPLE 5) How many sodium ions are found in 0.129 mol of sodium phosphate?

$$0.129\ \cancel{\text{mol } Na_3PO_4} \times \frac{6.02 \times 10^{23}\ \cancel{\text{fun } Na_3PO_4}}{1\ \cancel{\text{mol } Na_3PO_4}} \times \frac{3\ Na^+\ \text{ions}}{1\ \cancel{\text{fun } Na_3PO_4}} = \boxed{2.33 \times 10^{23}\ Na^+\ \text{ions}}$$

II. Mass and the Mole

[The atomic masses on the periodic table have a unit of **atomic mass unit** (amu, or u).]

A. **GAM** = **gram atomic mass**

1) *the atomic mass (listed on the periodic table) written in grams*

1 atom Xe = 131.30 u & 1 mol Xe = 131.30 g & GAM of Xe = 131.30 g

2) *these numbers are usually rounded to 0.01 (hundredths)*

B. **GMM** = **gram molecular mass**

1) *the sum of all masses of atoms in a molecular compound*

1 molecule Cl_2 = 70.90 u & 1 mol Cl_2 = 70.90 g & GMM Cl_2 = 70.90 g

2) example

EXAMPLE 6) Find the GMM of methane, CH_4.

$$CH_4 = 1(12.01) + 4(1.01) = 12.01 + 4.04 = \boxed{16.05\text{ g}}$$

C. **GFM** = **gram formula mass**

1) *the mass of one mole of ionic compound*

1 fun NaCl = 58.44 amu & 1 mol NaCl = 58.44 g & GFM NaCl = 58.44 g

2) example:

EXAMPLE 7) Find the GFM of calcium hydroxide.

$$Ca(OH)_2 = 1(40.08) + 2(16.00) + 2(1.01) = \boxed{74.10\text{ g}}$$

D. **molar mass**—*the mass, in g, of 1 mole of a substance*

- molar mass is a general term for doing GAM, GMM, or GFM

*** WE ROUND OUR MOLAR MASSES TO HUNDREDTHS (0.01 g), TWO DECIMAL PLACES ***

III. Mole-to-Mass and Mass-to-Mole Conversions (DA; dimensional analysis)

$\cancel{\text{grams A}} \times \dfrac{\text{1 mol A}}{\cancel{\text{MOLAR MASS (g) A}}} = \boxed{\text{mol A}}$		$\cancel{\text{mol B}} \times \dfrac{\text{MOLAR MASS (g) B}}{\cancel{\text{1 mol B}}} = \boxed{\text{grams B}}$

examples:

EXAMPLE 8) How many grams are in 0.70 mol of carbon dioxide?

$$0.70\ \cancel{\text{mol } CO_2} \times \frac{\overset{44.01}{[1(12.01) + 2(16.00)]}\text{ g } CO_2}{1\ \cancel{\text{mol } CO_2}} = \boxed{31\text{ g } CO_2}$$

EXAMPLE 9) How many moles are in 362 g of sodium bromide?

$$362\ \cancel{\text{g NaBr}} \times \frac{\text{1 mol NaBr}}{\underset{102.89}{(22.99 + 79.90)}\ \cancel{\text{g NaBr}}} = \boxed{\text{3.52 mol NaBr}}$$

Preview of gas laws…

IV. Molar Volume: volume-to-mole and mole-to-volume conversions

A. **STP** = **standard temperature and pressure**

1) *standard temperature = 0.00 °C, 273.15 K*

2) *standard pressure = 101.3 kPa, 1 atm, 760 mm Hg, 760 torr, 14.7 psi*

(the standard values for atm, mm Hg, and torr are exact numbers)

B. At STP, all gases occupy the same amount of space:

MOLAR VOLUME of any gas at STP: 22.4 L = 1 mol

C. examples

EXAMPLE 10) What is the volume, in L, of 0.495 mol of nitrogen dioxide gas at STP?

$$0.495\ \cancel{\text{mol } NO_2} \times \frac{22.4\text{ L } NO_2}{1\ \cancel{\text{mol } NO_2}} = \boxed{11.1\text{ L } NO_2}$$

EXAMPLE 11) How many moles are found in 84 L of neon gas at STP?

$$84\ \cancel{\text{L Ne}} \times \frac{1\text{ mol Ne}}{22.4\ \cancel{\text{L Ne}}} = \boxed{3.8\text{ mol Ne}}$$

MOLE CONVERSION FACTOR SUMMARY:

$$\frac{\textbf{6.02 x 10}^{23}\textbf{ representative particles}}{\textbf{1 MOLE}} \qquad \frac{\textbf{1 MOLE}}{\textbf{6.02 x 10}^{23}\textbf{ representative particles}}$$

$$\frac{\textbf{MOLAR MASS (g)}}{\textbf{1 MOLE}} \qquad \frac{\textbf{1 MOLE}}{\textbf{MOLAR MASS (g)}}$$

$$\frac{\textbf{22.4 L (for a gas at STP)}}{\textbf{1 MOLE}} \qquad \frac{\textbf{1 MOLE}}{\textbf{22.4 L (for a gas at STP)}}$$

V. Mass-to-r.p. and r.p.-to-mass conversions

EXAMPLE 12) How many grams of barium sulfide are in 3.39 x 10^{25} r.p. of barium sulfide?

$$3.39 \times 10^{25}\ \cancel{\text{fun BaS}} \times \frac{1\ \cancel{\text{mol BaS}}}{6.02 \times 10^{23}\ \cancel{\text{fun BaS}}} \times \frac{\overset{169.40}{(137.33 + 32.07)}\text{ g BaS}}{1\ \cancel{\text{mol BaS}}} = \boxed{9540\text{ g BaS}}$$

EXAMPLE 13) How many particles of rubidium nitrate are in 45.00 g of rubidium nitrate?

$$45.00\ \cancel{\text{g } RbNO_3} \times \frac{1\ \cancel{\text{mol } RbNO_3}}{\underset{146.48}{[84.47 + 14.01 + 3(16.00)]}\ \cancel{\text{g } RbNO_3}} \times \frac{6.02 \times 10^{23}\ \cancel{\text{fun } RbNO_3}}{1\ \cancel{\text{mol } RbNO_3}} = \boxed{1.85 \times 10^{23}\text{ fun } RbNO_3}$$

There are many types of mole problems:

1 step:	*r.p. → mol & mol → r.p.*	
	mass → mol & mol → mass	
	volume → mol & mol → volume	*(of gas at STP)*
2 step:	*mass → r.p. & r.p. → mass*	
	mass → volume & volume → mass	*(of gas at STP)*
	r.p. → volume & volume → r.p.	*(of gas at STP)*

Preview of gas laws…

VI. Gas Density and Molar Mass

A. Density *D = M / V*

B. *gas density usually measured in g/L*

C. *use 22.4 L = 1 mol to calculate molar masses (g/mol, the mass of 1 mole)*

D. examples

EXAMPLE 14) The density of a gas is 3.64 g/L in STP conditions. What is its molar mass?

$$\frac{3.64\ \text{g}}{\cancel{\text{L}}} \times \frac{22.4\ \cancel{\text{L}}}{1\ \text{mol}} = \boxed{81.5\ \text{g/mol}}$$

EXAMPLE 15) At STP, 6.00 L of a gas has a mass of 25.10 g. Calculate the density of the gas and its molar mass.

$$D = \frac{M}{V} = \frac{25.10\ \text{g}}{6.00\ \text{L}} = \boxed{4.18\ \text{g/L}}$$

$$\text{MOLAR MASS} = \frac{25.10\ \text{g}}{6.00\ \cancel{\text{L}}} \times \frac{22.4\ \cancel{\text{L}}}{1\ \text{mol}} = \boxed{93.7\ \text{g/mol}}$$

VII. Percent Composition

A. **<u>Percent composition</u>**—*% by mass of each individual element in a compound*

B. FORMULA → PERCENTS

C. remember to list all percentages

D. double-check that the % total is 100% (or very close if rounding)

E. formulas

$$\% = \frac{\#\ \text{g element}}{\#\ \text{g cmpd.}} \times 100 \qquad\qquad \% = \frac{\text{MOLAR MASS of element}}{\text{MOLAR MASS of cmpd.}} \times 100$$

F. examples

EXAMPLE 16) Find the % by mass of hydrogen and oxygen in water.

MOLAR MASS H_2O = 2(1.01) + 1(16.00) = 18.02 g

2 H + 1 O = total mass H_2O

$$\%\ \text{H} = \frac{2.02\ \text{g}}{18.02\ \text{g}} \times 100 = \boxed{11.2\%\ \text{H}} \qquad \%\ \text{O} = \frac{16.00\ \text{g}}{18.02\ \text{g}} \times 100 = \boxed{88.79\ \%\ \text{O}}$$

EXAMPLE 17) Calculate the % composition of sulfuric acid.

MOLAR MASS H_2SO_4 = 2(1.01) + 32.07 + 4(16.00) = 98.09 g
2 H + 1 S + 4 O = total mass H_2SO_4

$\% H = \frac{2.02 \text{ g}}{98.09 \text{ g}} \times 100 = \boxed{2.06\ \% H}$

$\% S = \frac{32.07 \text{ g}}{98.09 \text{ g}} \times 100 = \boxed{32.69\ \% S}$

$\% O = \frac{4(16.00) \text{ g}}{98.09 \text{ g}} \times 100 = \boxed{65.25\ \% O}$

VIII. Empirical Formula

A. **Empirical formula**—*the simplest whole-number ratio of elements in a cmpd.*
B. it is a non-reducible ratio of moles
C. PERCENTS → FORMULA
D. it is the reverse process of percent composition
E. sometimes empirical formulas can be found in nature
F. problem procedure

> *1. convert % to grams directly*
> *2. find numbers of moles*
> *3. make mole ratios using the smallest mole number as the denominator*
> *4. use these whole number ratios as the subscripts of the formula*
> *(If a ratio with 0.5 is observed, multiply everything by 2. Don't round up!)*
> *(If a ratio with 0.33 is observed, multiply everything by 3. Don't round up!)*

G. examples

EXAMPLE 18) Calculate the empirical formula of a compound composed of 67.6% mercury, 10.8% sulfur, and 21.6% oxygen.

STEP 1… mercury: 67.6% Hg = 67.6 g Hg out of 100 g cmpd.
sulfur: 10.8% S = 10.8 g S out of 100 g cmpd.
oxygen: 21.6% O = 21.6 g O out of 100 g cmpd.

STEP 2... Hg: $67.6 \text{ g Hg} \times \frac{1 \text{ mol Hg}}{200.59 \text{ g Hg}} = \mathit{0.337\ mol\ Hg}$

S: $10.8 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = \mathit{0.337\ mol\ S}$

O: $21.6 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g}} = \mathit{1.35\ mol\ O}$

STEP 3… $Hg = \frac{0.337}{0.337} = 1$ $S = \frac{0.337}{0.337} = 1$ $O = \frac{1.35 \text{ mol}}{0.337 \text{ mol}} = 4$

STEP 4… $Hg_1S_1O_4 = \boxed{HgSO_4}$

EXAMPLE 19) What is the empirical formula of a 25.9% nitrogen and 74.1% oxygen compound?

25.9% N = 25.9 g N out of 100 g cmpd. 74.1% O = 74.1 g O out of 100 g cmpd.

N: $25.9 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = \mathit{1.85\ mol\ N}$ O: $74.1 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = \mathit{4.63\ mol\ O}$

N: $\frac{1.85}{1.85} = 1$ O: $\frac{4.63}{1.85} = 2.5$ $N_1O_{2.5}$ – can't have .5 subscripts x 2 = $\boxed{N_2O_5}$

IX. Molecular Formula

A. **<u>Molecular formula</u>**—*a multiple of the empirical formula*

B. still whole number ratios

C. examples

EXAMPLE 20) A compound with an empirical formula of CH has a molecular weight of 78.0 g/mol. What is the molecular formula?

molar mass CH = 12.01 + 1.01 = 13.02 g   78.0 / 13.02 = 6   molecular formula = $\boxed{C_6H_6}$

EXAMPLE 21) A compound is 75.46% carbon, 4.43% hydrogen, and 20.10% oxygen by mass. It has a molecular mass of 318.31 g/mol. What is the molecular formula for this compound?

75.46% C = 75.46 g C out of 100 g cmpd.
4.43% H = 4.43 g H out of 100 g cmpd.
20.10% O = 20.10 g O out of 100 g cmpd.

(75.46 g C) (1 mol/ 12.01 g C) = *6.283 mol C*
(4.43 g H) (1 mol/ 1.01 g H) = *4.39 mol H*
(20.10 g O) (1 mol/ 16.00 g O) = *1.256 mol O*

(6.28 mol C)/ (1.26) = 4.99 = 5 mol C
(4.39 mol H)/ (1.26) = 3.49 = 3.5 mol H
(1.26 mol O)/ (1.26) = 1 mol O
.5 value means multiply subscripts by 2:   empirical formula = $C_{10}H_7O_2$

Now that you have the emp.fmla., you can find the molecular fmla like in problem E20).
emp. fmla. mass = 10(12.01) + 7(1.01) + 2(16.00) = 159.17 g/mol
The problem says the molecular mass is 318.31 g per mole.

$$\frac{318.31 \text{ g/mol}}{159.17 \text{ g/mol}} = 2 \text{ ratio}$$

Since there are two empirical units in a molecular unit, the molecular formula = $\boxed{C_{20}H_{14}O_4}$