

## Chem. I Notes – Ch. 12, part 2 – Using Moles

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

**1 MOLE =  $6.02 \times 10^{23}$  representative particles**  
**(representative particles = ATOMS, IONS, MOLECULES, & FORMULA UNITS)**

### CONVERSION FACTOR SUMMARY:

<u><math>6.02 \times 10^{23}</math> representative particles</u>	<u>1 MOLE</u>
1 MOLE	<u><math>6.02 \times 10^{23}</math> representative particles</u>
<u>MOLAR MASS (g)</u>	<u>1 MOLE</u>
1 MOLE	<u>MOLAR MASS (g)</u>
<u>22.4 L (for a gas at STP)</u>	<u>1 MOLE</u>
1 MOLE	<u>22.4 L (for a gas at STP)</u>

**stoichiometry**—using balanced chemical equations to obtain info.

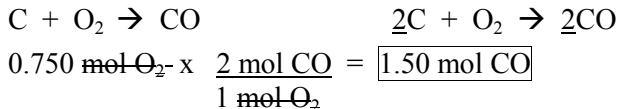
### I. **Mole - Mole (MOL – MOL) Conversions**

- A. the most important, most basic stoich calculation
- B. *uses the coefficients of a balanced equation* to compare the amounts of reactants and products
- C. *coefficients are mole ratios*
- D. *the way to go from substance A to substance B*
- E. **mol – mol is the only time the mole number in the conversion is not automatically 1.** (Avogadro's #, molar mass, and 22.4 L (STP) are all equal to 1 mole.)

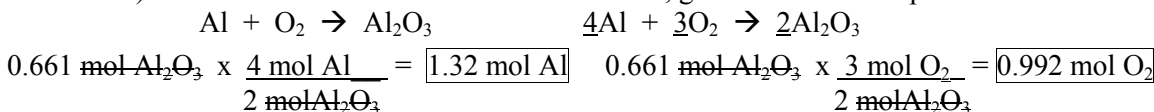
<b>MOL – MOL :</b>	<u><math>\frac{\# \text{ mol A}}{\# \text{ mol B}}</math></u>	<u><math>\frac{\# \text{ mol B}}{\# \text{ mol A}}</math></u>	<b># = coefficients</b>
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F. examples

- E1) How many moles of carbon monoxide are produced when 0.750 mol of oxygen reacts with carbon?



- E2) Find the number of moles of the reactants, given 0.661 mol of product is formed.



### II. **MASS – MASS Conversions** – Using molar mass in stoich problems to predict masses of reactants and/or products

- A. a balanced chemical equation can be used to compare masses of reactants and products
- B. *mass – mass* cannot change which substance you are dealing with; only *mol – mol* can do that

$$\text{"MASS - MASS": } \frac{\text{GIVEN g A}}{\text{PT g A}} \times \frac{1 \text{ mol A}}{\text{CE mol A}} \times \frac{\text{CE mol B}}{\text{CE mol A}} \times \frac{\text{PT g B}}{1 \text{ mol B}}$$

*PT = periodic table, molar mass*

*CE = coefficients*

C. examples

- E3) How many grams of hydrochloric acid are made from the reaction of 0.500 g of hydrogen gas with excess chlorine gas?



GAME PLAN: KNOWN, PER.TABLE, COEFF., PER.TABLE

$$0.500 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.0 \text{ g H}_2} \times \frac{2 \text{ mol HCl}}{1 \text{ mol H}_2} \times \frac{36.5 \text{ g HCl}}{1 \text{ mol HCl}} = \boxed{18 \text{ g HCl}}$$

- E4) Calculate the numbers of grams of products formed when 25.0 g of sodium nitrate decomposes into sodium nitrite and oxygen.



$$25.0 \text{ g NaNO}_3 \times \frac{1 \text{ mol NaNO}_3}{85.0 \text{ g NaNO}_3} \times \frac{2 \text{ mol NaNO}_2}{2 \text{ mol NaNO}_3} \times \frac{69.0 \text{ g NaNO}_2}{1 \text{ mol NaNO}_2} = \boxed{20.3 \text{ g NaNO}_2}$$

$$25.0 \text{ g NaNO}_3 \times \frac{1 \text{ mol NaNO}_3}{85.0 \text{ g NaNO}_3} \times \frac{1 \text{ mol O}_2}{2 \text{ mol NaNO}_3} \times \frac{32.0 \text{ g O}_2}{1 \text{ mol O}_2} = \boxed{4.71 \text{ g O}_2}$$

III. Other Stoich Calculations

A. "mole - mass" (mass - mole) calculations

"MASS - MOLE":

$$\frac{\text{GIVEN g A}}{\text{PT g A}} \times \frac{1 \text{ mol A}}{\text{CE mol A}} \times \frac{\text{CE mol B}}{\text{CE mol A}}$$

"MOLE - MASS":

$$\frac{\text{GIVEN mol A}}{\text{CE mol A}} \times \frac{\text{CE mol B}}{\text{CE mol A}} \times \frac{\text{PT g B}}{1 \text{ mol B}}$$

*PT = periodic table, molar mass*      *CE = coefficients*

- E5) How many g of water are produced from the complete combustion of 0.6829 mol of C<sub>2</sub>H<sub>2</sub>?



$$0.6829 \text{ mol C}_2\text{H}_2 \times \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol C}_2\text{H}_2} \times \frac{18.0 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = \boxed{12.3 \text{ g H}_2\text{O}}$$

- E6) Using the equation  $\underline{2}\text{C}_2\text{H}_2 + \underline{5}\text{O}_2 \rightarrow \underline{4}\text{CO}_2 + \underline{2}\text{H}_2\text{O}$  how many moles of O<sub>2</sub> would be needed to produce 56.09 g of CO<sub>2</sub>?

$$56.09 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.0 \text{ g CO}_2} \times \frac{5 \text{ mol O}_2}{4 \text{ mol CO}_2} = \boxed{1.59 \text{ mol O}_2}$$

B. “mass – volume” (volume – mass) calculations – Using molar volume in stoich problems

**“MASS – VOLUME”:** (gases @ STP)

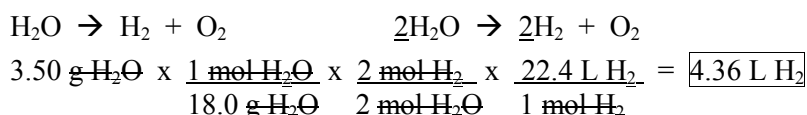
$$\frac{\text{GIVEN g A}}{PT \text{ g A}} \times \frac{1 \text{ mol A}}{CE \text{ mol A}} \times \frac{CE \text{ mol B}}{1 \text{ mol B}} \times \frac{22.4 \text{ L B}}{1 \text{ mol B}}$$

**“VOLUME – MASS”:** (gases @ STP)

$$\frac{\text{GIVEN L A}}{22.4 \text{ L A}} \times \frac{1 \text{ mol A}}{CE \text{ mol A}} \times \frac{CE \text{ mol B}}{1 \text{ mol B}} \times \frac{PT \text{ g B}}{1 \text{ mol B}}$$

*PT = periodic table, molar mass      CE = coefficients*

- E7) How many L of hydrogen are produced from the decomposition of 3.50 g of water at STP?



- E8) Using the equation  $\underline{2}\text{C}_2\text{H}_2 + \underline{5}\text{O}_2 \rightarrow \underline{4}\text{CO}_2 + \underline{2}\text{H}_2\text{O}$  how many liters of water vapor are produced when 5.02 g of  $\text{C}_2\text{H}_2$  undergoes complete combustion under STP conditions?

$$5.02 \text{ g C}_2\text{H}_2 \times \frac{1 \text{ mol C}_2\text{H}_2}{26.0 \text{ g C}_2\text{H}_2} \times \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol C}_2\text{H}_2} \times \frac{22.4 \text{ L H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = \boxed{4.32 \text{ L H}_2\text{O}}$$

C. “volume – volume” calculations

**“VOLUME – VOLUME”:** (gases @ STP)

$$\frac{\text{GIVEN L A}}{22.4 \text{ L A}} \times \frac{1 \text{ mol A}}{CE \text{ mol A}} \times \frac{CE \text{ mol B}}{1 \text{ mol B}} \times \frac{22.4 \text{ L B}}{1 \text{ mol B}}$$

*CE = coefficients      (SHORT CUT: compare coefficients!)*

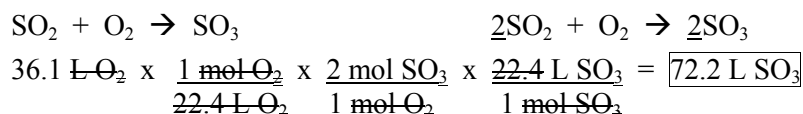
- E9) How many liters of  $\text{CO}_2$  are produced from 0.252 L of  $\text{HCl}$  reacting with  $\text{NaHCO}_3$ ?

Use the equation  $\text{NaHCO}_3 + \text{HCl} \rightarrow \text{NaCl} + \text{CO}_2 + \text{H}_2\text{O}$

$$0.252 \text{ L HCl} \times \frac{1 \text{ mol HCl}}{22.4 \text{ L HCl}} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol HCl}} \times \frac{22.4 \text{ L CO}_2}{1 \text{ mol CO}_2} = \boxed{0.252 \text{ L CO}_2}$$

*SHORTCUT: coefficients = 1 mol HCl to 1 mol CO<sub>2</sub>, so 0.252 L HCl = 0.252 L CO<sub>2</sub>*

- E10) How many L of sulfur trioxide are produced from the reaction of 36.1 L of oxygen with sulfur dioxide at STP?



SHORTCUT: coefficient of  $\text{O}_2 = 1$ ; coefficient of  $\text{SO}_3 = 2$        $36.1 \text{ L} \times 2 = \boxed{72.2 \text{ L SO}_3}$

D. **“mass – particle” (particle – mass) calculations**

“MASS – PARTICLE”:

$$\frac{\text{GIVEN g A}}{PT \text{ g A}} \times \frac{1 \text{ mol A}}{CE \text{ mol A}} \times \frac{CE \text{ mol B}}{1 \text{ mol B}} \times \frac{(6.02 \times 10^{23}) \text{ r.p. B}}{1 \text{ mol B}}$$

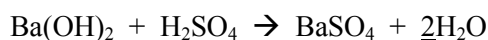
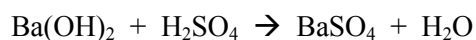
“PARTICLE – MASS”:

$$\frac{\text{GIVEN r.p. A}}{(6.02 \times 10^{23}) \text{ r.p. A}} \times \frac{1 \text{ mol A}}{CE \text{ mol A}} \times \frac{CE \text{ mol B}}{1 \text{ mol B}} \times \frac{PT \text{ g B}}{1 \text{ mol B}}$$

*PT = periodic table, molar mass*

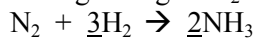
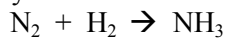
*CE = coefficients*

E11) How many formula units of barium sulfate are made from reacting 5.33 g of barium hydroxide with sulfuric acid?



$$5.33 \text{ g Ba(OH)}_2 \times \frac{1 \text{ mol Ba(OH)}_2}{171.3 \text{ g Ba(OH)}_2} \times \frac{1 \text{ mol BaSO}_4}{1 \text{ mol Ba(OH)}_2} \times \frac{6.02 \times 10^{23} \text{ f.units BaSO}_4}{1 \text{ mol BaSO}_4} = \boxed{1.87 \times 10^{22} \text{ f.units BaSO}_4}$$

E12) How many molecules of NH<sub>3</sub> are produced from reacting 2.07 g of H<sub>2</sub> with excess N<sub>2</sub>?



$$2.07 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.0 \text{ g H}_2} \times \frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2} \times \frac{6.02 \times 10^{23} \text{ molecules NH}_3}{1 \text{ mol NH}_3} = \boxed{4.2 \times 10^{23} \text{ molecules NH}_3}$$

IV. Percent Yield

- A. **percent yield**—percentage of product recovered; comparison of actual and theoretical yields
- B. **actual yield**—amount of product obtained in lab
- C. **theoretical yield**—amount of product predicted by the math (theory)

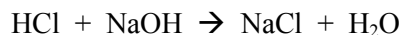
$$\% \text{ YIELD} = \frac{\text{ACTUAL YIELD}}{\text{THEORETICAL YIELD}} \times 100$$

D. examples

E14) 35.0 g of product should be recovered from an experiment. A student collects 22.9 g at the end of the lab. What is the percent yield?

$$\frac{22.9 \text{ g}}{35.0 \text{ g}} \times 100 = \boxed{65.4\%}$$

E15) What is the percent yield if 2.89 g of NaCl is produced when 1.99 g of HCl reacts with excess NaOH? Water is the other product.



*Percent yield implies product yield.*

Actual yield = 2.89 g NaCl

Theoretical yield = ?

$$1.99 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.5 \text{ g HCl}} \times \frac{1 \text{ mol NaCl}}{1 \text{ mol HCl}} \times \frac{58.5 \text{ g NaCl}}{1 \text{ mol NaCl}} = 3.19 \text{ g NaCl theoretical yield}$$

$$\% \text{ YIELD} = \frac{\text{ACTUAL YIELD}}{\text{THEORETICAL YIELD}} \times 100 = \frac{2.89 \text{ g NaCl}}{3.19 \text{ g NaCl}} = \boxed{90.6\%}$$

Ideal Gas Law will be covered later

Percent Composition already covered in part 1

V. Empirical Formula – Determining Chemical Formulas

- A. **Empirical Formula**—the simplest whole-number ratio of elements in a compd.
- B. it is a non-reducible ration of moles
- C. 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 number with 0.5 is observed, multiply everything by 2. Don't round up!)

D. examples

E16) Calculate the empirical formula of a compound composed of 67.6% mercury, 10.8% sulfur, and 21.6% oxygen. (check the PROBLEM PROCEDURE steps)

STEP 1... mercury: 67.6% Hg = 67.6 g Hg out of 100 g compd.

sulfur: 10.8% S = 10.8 g S out of 100 g compd.

oxygen: 21.6% O = 21.6 g C out of 100 g compd.

STEP 2... Hg:  $67.6 \text{ g C} \times \frac{1 \text{ mol Hg}}{200.6 \text{ g Hg}} = 0.337 \text{ mol Hg}$     S:  $10.8 \text{ g S} \times \frac{1 \text{ mol S}}{32.1 \text{ g S}} = 0.336 \text{ mol S}$

O:  $21.6 \text{ g O} \times \frac{1 \text{ mol O}}{16.0 \text{ g}} = 1.35 \text{ mol O}$

STEP 3... Hg =  $\frac{0.337}{0.336} = 1$     S =  $\frac{0.336}{0.336} = 1$     O =  $\frac{1.35 \text{ mol}}{0.336 \text{ mol}} = 4$

STEP 4...  $\text{Hg}_1\text{S}_1\text{O}_4 = \boxed{\text{HgSO}_4}$

E17) What is the empirical formula of a compound with 25.9% nitrogen and 74.1% oxygen?

25.9% N = 25.9 g N out of 100 g compd.

74.1% O = 74.1 g O out of 100 g compd.

N:  $25.9 \text{ g N} \times \frac{1 \text{ mol N}}{14.0 \text{ g N}} = 1.85 \text{ mol N}$

O:  $74.1 \text{ g O} \times \frac{1 \text{ mol O}}{16.0 \text{ g O}} = 4.63 \text{ mol O}$

N:  $\frac{1.85}{1.85} = 1$     O:  $\frac{4.63}{1.85} = 2.5$      $\text{N}_1\text{O}_{2.5}$  – can't have .5 subscripts x 2 =  $\boxed{\text{N}_2\text{O}_5}$

VI. **Molecular Formula**

A. *a multiple of the empirical formula*

B. still whole number ratios

C. examples

E18) A compound with an empirical formula of CH has a molecular weight of 78.0 g/mol. What is the molecular formula?

$$\text{molar mass of CH} = 12.0 + 1.0 = 13.0 \text{ g} \quad 78 / 13 = 6 \quad \text{molecular formula} = \boxed{\text{C}_6\text{H}_6}$$

E19) 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?

$$67.6\% \text{ C} = 75.46 \text{ g C out of 100 g compd.} \quad 4.43\% \text{ H} = 4.43 \text{ g H out of 100 g compd.}$$

$$20.10\% \text{ O} = 20.10 \text{ g O out of 100 g compd.}$$

$$(75.46 \text{ g C}) (1 \text{ mol} / 12.0 \text{ g C}) = 6.29 \text{ mol C} \quad (4.43 \text{ g H}) (1 \text{ mol} / 1.0 \text{ g H}) = 4.4 \text{ mol H}$$

$$(20.10 \text{ g O}) (1 \text{ mol} / 16.0 \text{ g O}) = 1.26 \text{ mol O}$$

$$(6.29 \text{ mol C}) / (1.26) = 4.99 = 5 \text{ mol C}$$

$$(4.4 \text{ mol H}) / (1.26) = 3.49 = 3.5 \text{ mol H}$$

$$(1.26 \text{ mol O}) / (1.26) = 1 \text{ mol O}$$

.5 value means multiply subscripts by 2: empirical formula =  $\text{C}_{10}\text{H}_7\text{O}_2$

Now that you have the empirical formula, you can find the molecular formula like in problem E18.

$$\text{emp. formula mass} = 10(12.0) + 7(1.0) + 2(16.0) = 159.1 \text{ g/mol}$$

The problem says the molecular mass is 318.31 g per mole.

$$\frac{318.31 \text{ g/mol}}{159.1 \text{ g/mol}} = 2.001 = 2 \text{ ratio}$$

$$159.1 \text{ g/mol}$$

Since there are two empirical units in a molecular unit, the molecular formula =  $\boxed{\text{C}_{20}\text{H}_{14}\text{O}_4}$