UNIT 5: MOLES & STOICHIOMETRY

VOCABULARY:

1. Mole
2. Formula mass (FM)
3. Gram formula mass (GFM)
4. Coefficient
5. Subscript
6. Species
7. Law of conservation of mass
8. Law of conservation of energy
9. Balanced equation
10. Synthesis reaction
11. Decomposition reaction
12. Single-replacement reaction
13. Double-replacement reaction
14. Molecular formula
15. Empirical formula
16. Percent mass

INTRODUCTION:

Before we can even begin to understand what this unit is about, we need to be able to find the mass of different compounds. Open your Periodic Table and we'll get started…

- First, what are the units we use for the mass atoms?
  ATOMIC MASS UNITS (amu) → $\frac{1}{12}$ of the mass of C-12 atom
- What is the mass of one atom of oxygen?
  15.9994 amu
- Why don't we use grams as the units for massing atoms? Atoms are too SMALL—the number would be VERY BULKY
  Ex: If we used grams to mass atoms, the mass of oxygen would be 0.000000000000000000000027 g or 2.7×10^{-23}g

Find the mass of the following *atoms*:

1) Mg = 24.305 amu 3) Cl = 35.453 amu 5) Ca = 40.08 amu

2) Li = 6.941 amu 4) Al = 26.9815 amu 6) H = 1.00794 amu

- **MONOATOMIC** ELEMENTS = one atom of an element that's stable enough to stand on its own (VERY RARE)—not bonded to anything
- **DIATOMIC** ELEMENTS or DIATOMS = elements whose atoms always travel in pairs (N_2, O_2, F_2, Cl_2, Br_2, I_2, At_2, H_2)—bonded to another atom of the same element

So, what would the mass be of one ***molecule*** of oxygen (O_2)?

O_2 ← subscript = tells you the total number of atoms in the compound/molecule

This means that the mass of O_2 = 2 x 15.994 amu = 31.9988 amu

Calculating Formula Mass & Gram Formula Mass of Compounds:

- **FORMULA MASS**: the mass of an atom, molecule or compound in ATOMIC MASS UNITS (amu)
  Ex: formula mass of a hydrogen atom is 1.00794 amu
- **GRAM FORMULA MASS**: the mass of one **MOLE** of an atom, molecule or compound in GRAMS (g)
  Ex: GFM of hydrogen is 1.00794 g
- **MOLE**: 6.02×10^{23} units of a substance (like a really big dozen)
  Avogadro's
  Ex: 1 mol of C = 6.02×10^{23} atoms of C = 12.011 g of C
  Quantity Mass

Practice – SHOW ALL WORK!

→ mass of 1 molecule (amu)

1) What is the formula mass of K_2CO_3?

K = 2 × 39.0983 = 78.1966
C = 1 × 12.011 = 12.011
O = 3 × 15.9994 = 47.9982

138.2059 amu

→ mass of 1 mole (g)

2) What is the gram formula mass of $CuSO_4 \cdot 5H_2O$?

Cu = 1 × 63.546 = 63.546
S = 1 × 32.06 = 32.06
O = 4 × 15.9994 = 63.9976
H = 10 × 1.00794 = 10.0794
O = 5 × 15.994 = 79.997 +

249.68 g

Calculating Percent Composition

Step 1: Calculate the GFM for the compound (or the FM).
Ex: $CaCl_2$

Ca = 1 x 40.08 = 40.08 ← mass of the part (this is the "part" Ca)

Cl = 2 x 35.453 = 70.906 + (this is the "part" Cl)

110.986 g ← mass of the whole

Step 2: Check the last page of your periodic table for the formula for percent composition. Write the formula below:

Ca

$$\% \text{ composition by mass} = \frac{\text{mass of part}}{\text{mass of whole}} \times 100$$

Now, use the formula to find the percent composition of each element or "part" in our compound (to the nearest 0.1 %).

Ca: $\left(\frac{40.08}{110.986}\right) \times 100 = \boxed{36.1\% \text{ Ca}}$

Cl: $\left(\frac{70.906}{110.986}\right) \times 100 = \boxed{63.9\% \text{ Cl}}$

Practice:

1) What is the percentage by mass of carbon in CO_2?

$C = 1 \times 12.011 = 12.011$

$O = 2 \times 15.9994 = 31.9988 +$

44.0098

$\% \text{comp.} = \left(\frac{12.011}{44.0098}\right) \times 100$

$= \boxed{27.3\%\ C}$

2) What is the percent by mass of nitrogen in NH_4NO_3?

$N = 2 \times 14.0067 = 28.0134 \leftarrow$ part

$H = 4 \times 1.00794 = 4.03176$

$O = 3 \times 15.9994 = 47.9982$

$80.04336\,g \leftarrow$ whole

$\% \text{Comp} = \frac{28.0134}{80.04336} \times 100$

$= \boxed{34.9\%\ N}$

3) What is the percent by mass of oxygen in magnesium oxide?

$\overset{+2}{Mg}\overset{-2}{O}$

$Mg = 24.305$

$O = 15.9994 \leftarrow$ part

$40.3044 \leftarrow$ whole

$\% \text{Comp} = \frac{15.9994}{40.3044} \times 100 = \boxed{39.7\%\ O}$

A Special Type of Percent Composition: CRYSTAL HYDRATES

A **HYDRATE** is a **CRYSTALLINE** compound in which ions are attached to one or more **WATER** molecules

Example: $Na_2CO_3 \cdot 10H_2O$

- Notice how **WATER** molecules are BUILT INTO the chemical formula
- Substances ***without*** water built into the formula are called **ANHYDRATES**

Problem: What is the percentage by mass of water in sodium carbonate crystals ($Na_2CO_3 \cdot 10H_2O$)?

Step 1: Calculate the formula mass for the hydrate.

$$\begin{array}{r} 2\times(1.00794) \\ + \; 1\times(15.9994) \\ \hline 18.01528 \end{array}$$

whole

Na =	2 x 22.98977	=	45.97954
C =	1 x 12.0111	=	12.0111
O =	3 x 15.9994	=	47.9982
H_2O =	10 x 18.01528	=	180.1528
	Formula mass of hydrate	=	286.14164

part

Step 2: Check the last page of your periodic table for the formula for percent composition. Write the formula below:

$$\text{\% composition by mass} = \frac{\text{mass of part}}{\text{mass of whole}} \times 100$$

$$\text{\% } H_2O \text{ by mass} = \frac{180.1528}{286.1416} \times 100 = \boxed{63.0\%}$$

Practice:

1) What is the percent by mass of water in $BaCl_2 \cdot 2H_2O$?

$Ba = 137.33$

$Cl = 2 \times (35.453)$

208.236

$H_2O = 2 \times (18.01528)$

36.03056

$\frac{36.03056}{244.26656} \times 100$

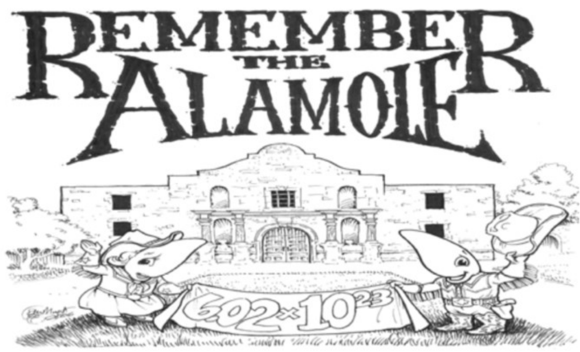
$\boxed{= 14.8\%}$

2) Which species contains the greatest percent by mass of oxygen?

a) CO_2 b) H_2O c) NO_2 d) MgO

CO_2 $C = 12$ $2O = 32 +$ 44 $O = \frac{32}{44} \times 100 = \boxed{72.7\%}$	H_2O $2H = 2$ $O = 16 +$ 18 $O = \frac{16}{18} \times 100 = \boxed{88.9\%}$
NO_2 $N = 14$ $2O = 32 +$ 46 $O = \frac{32}{46} \times 100 = \boxed{70.0\%}$	MgO $Mg = 24$ $O = 16 +$ 40 $O = \frac{16}{40} \times 100 = \boxed{40.0\%}$

* round atomic masses to nearest whole number



Application of the Mole: The Math of Chemistry

We will need to convert from grams to moles and vice versa for this class. The diagram below summarizes these processes:

Mass in grams (g) $\xrightarrow{\text{Divide by GFM}}$ **Number of moles (mol)**

Number of moles (mol) $\xrightarrow{\text{Multiply by GFM}}$ **Mass in grams (g)**

Converting from Grams to Moles:

From Table T, you would use the Mole Calculations Formula:

$$GFM \times \text{\# of moles} = \frac{\text{given mass (g)}}{\cancel{\text{GFM}}\text{ (g/mol)}} \cancel{\times GFM}$$

given mass = GFM x # of moles

X Given

Problem: How many moles are in 4.75 g of sodium hydroxide? (NaOH)

Step 1: Calculate the GFM for the compound.

Na = 1 x 22.98977
O = 1 x 15.9994
H = 1 x 1.00794
39.99711 g/mol

Step 2: Plug the given value and the GFM into the "mole calculations" formula and solve for the number of moles.

$$\# \text{ of moles} = \frac{\text{given mass (g)}}{\text{GFM (g/mol)}} = \frac{4.75}{39.99711} = \boxed{0.119 \text{ mol}}$$

Practice: X Given

1) How many moles are in 39.0 grams of LiF?

GFM

Li 1 x 6.941
F 1 x 18.9984 +
25.9394 g/mol

$$\# \text{ of moles} = \frac{39.0}{25.9394} = \boxed{1.504 \text{ mol of LiF}}$$

2) What is the number of moles of potassium chloride present in 148 g?

+1 -1

KCl

GFM

K 1 x 39.0983
Cl 1 x 35.453 +
74.5513 g/mol

$$\# \text{ of mol KCl} = \frac{148}{74.5513} = \boxed{1.98 \text{ mol of KCl}}$$

3) How many moles are in 168 g of KOH?

GFM

56.10564

$$\# \text{ of moles} = \frac{168}{56.10564} = \boxed{2.994 \text{ mol of KOH}}$$

Converting from Moles to Grams:

From Table T, you would still use the Mole Calculations Formula, but you must rearrange it since you are solving for **GRAMS** now:

mass of sample (g) = # of moles (mol) x GFM (g/mol)

Problem: You have a 2.50 mole sample of sulfuric acid. What is the mass of your sample in grams? (H_2SO_4)

Step 1: Calculate the GFM for the compound.

GFM

H 2 × 1.00794 = 2.01588
S 1 × 32.06 = 32.06
O 4 × 15.9994 = 63.9976 +
98.07348 g/mol

Mass = mol × GFM
= 2.5 × 98.07348
= 98.07348 g

Step 2: Plug the given value and the GFM into the "mole calculations" formula and solve for the mass of the sample.

mass of sample (g) = # of moles (mol) x GFM (g/mol)

Practice: X Given

1) What is the mass of 4.5 moles of KOH?

K 1 × 39.0983
O 1 × 15.9994
H 1 × 1.00794
56.1056

mass = 4.5 mol × 56.1056 g/mol
= $\boxed{252.4752 \text{ g KOH}}$

2) What is the mass of 0.50 mol of $CuSO_4$?

GFM
159.6086 g/mol

$mass = 0.50\ \cancel{mol} \times 159.6086\ g/\cancel{mol}$
$= \boxed{79.8043\ g\ CuSO_4}$

3) What is the mass of 1.50 mole of nitrogen N_2 gas?

GFM
28.0134

$mass = 1.50\ \cancel{mol} \times 28.0134\ g/\cancel{mol}$
$= \boxed{42.0\ g\ N_2}$

CHALLENGE: Convert from grams to atoms/molecules or vice versa.

4) How many molecules of SO_2 are there in a 1.75 g sample?

steps
① First convert mass to moles
② Then convert moles to molecules

Avogadro's number

① GFM
1 S = 32.065
2 O = 2(15.9994)
(Table T) 64.0638 g/mol

$moles = \frac{given\ mass}{gfm} = \frac{1.75}{64.06385} = 0.2731\ mol$

② $molecules = moles \times 6.02 \times 10^{23}$
$= 0.2731 \times 6.02 \times 10^{23}$
$= \boxed{1.64 \times 10^{22}\ molecules}$

5) What is the mass of 3.01×10^{23} atoms of carbon?

(Factor-label method)

$$3.01 \times 10^{23}\ \cancel{atoms\ C} \times \frac{1\ \cancel{mol\ C}}{6.02 \times 10^{23}\ \cancel{atoms\ C}} \times \frac{12.011\ g\ C}{1\ \cancel{mole\ C}} = \boxed{6.01\ g\ C}$$

Chemical Equations:

- A CHEMICAL EQUATION is a set symbols that state the **PRODUCTS** and **REACTANTS** in a chemical reaction.

REACTANTS = the starting substances in a chemical reaction (found to the **LEFT** of the arrow)

PRODUCTS = a substance produced by a chemical reaction (found to the **RIGHT** of the arrow)

Example:

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

- Chemical equations must be **BALANCED**. Think of the arrow (→) as an equal sign.
- LAW of CONSERVATION of MASS: mass can neither be **CREATED** nor **DESTROYED** in a chemical reaction

Balancing Equations:

The number of **MOLES** of each **ELEMENT** on the **REACTANTS** (left) side of the equation must be the same as the number of **MOLES** of each **ELEMENT** on the **PRODUCTS** (right) side of the equation.

***COEFFICIENTS** and **SUBSCRIPTS** tell us how many moles we have for each element

Let's look at the BALANCED equation below:

$$1C + 1O_2 \rightarrow 1CO_2$$

*Note that there is 1 mol of carbon and 2 mol of oxygen on each side of the arrow. That's what it means to be BALANCED.

1 mol C = 1 mol C
2 mol O = 2 mol O

Click here to watch vodcast: https://www.youtube.com/watch?v=YXTMfwEZAyU

Now, let's examine the following UNBALANCED equation:

$$H_2 + O_2 \rightarrow H_2O$$

Q: How does this unbalanced equation violate the Law of Conservation of Mass?
A: In this equation, oxygen would have to be **DESTROYED** (there's one less on the products side)

- **COEFFICIENT** = the integer in front of an element or compound which indicates the number of moles present
- **SUBSCRIPT** = the integer to the lower right of an element which indicates the number of atoms present
- **SPECIES** = the individual reactants and products in a chemical reaction.

Q: What do we use to balance equations?
A: **COEFFICIENTS**

NOTE: WE <u>NEVER</u> CHANGE THE SUBSCRIPTS IN A FORMULA!

Example:

$$2Ag + S \rightarrow Ag_2S$$

COEFFICIENTS:
Ag = 2
S = 1
Ag_2S = 1

SUBSCRIPTS:
Ag = 1
S = 1
Ag_2S:
Ag = 2
S = 1

Method for Balancing Equations:

Step 1: Draw a line to separate products from reactants

Step 2: List each of the different elements on each side of the line

Step 3: Count up the number of atoms on each side & record next to the element symbol

Step 4: Find the most complex compound in the equation. Balance the elements found in that compound on the opposite side of the arrow by changing the coefficients for the different species. Every time you change a coefficient, you must update the number of each element.

Step 5: Now, continue balancing the elements by changing coefficients until you have the same number of each element on both sides of the equation.

Example: $\underline{2}\ H_2 + \underline{1}\ O_2 \rightarrow \underline{2}\ H_2O$

H ~~2~~ 4	H ~~2~~ 4
O 2	O ~~1~~ 2

Example: $\underline{2}\ Na + \underline{2}\ H_2O \rightarrow \underline{2}\ NaOH + \underline{1}\ H_2$

Na ~~1~~ 2	Na ~~1~~ 2
H ~~2~~ 4	H ~~3~~ 4
O ~~1~~ 2	O ~~1~~ 2

Example: $\underline{6}\ CO_2 + \underline{6}\ H_2O \rightarrow \underline{1}\ C_6H_{12}O_6 + \underline{6}\ O_2$

C ~~1~~ 6	C 6
O ~~3~~ ~~13~~ 18	O ~~8~~ 18
H ~~2~~ 12	H 12

ONE LAST NOTE:

When balancing chemical equations, **POLYATOMIC IONS** may be balanced as a **SINGLE ELEMENT** rather than as separate elements as long as they stay intact during the reaction.

Example: $Al_2(SO_4)_3 + Ca(OH)_2 \rightarrow Al(OH)_3 + CaSO_4$

In this equation, we have the polyatomic ions SULFATE & HYDROXIDE, and both remain intact during the reaction. Since SO_4 has the subscript of 3, we *could* think of it as 3 x 1 = 3 sulfur atoms and 3 x 4 = 12 oxygen atoms. OR, we can just look at the UNIT and say there are 3 (SO_4)'s on the reactant side and 1 (SO_4) on the product side.

Now let's balance the equation:

1 $Al_2(SO_4)_3$ + 3 $Ca(OH)_2$ $\rightarrow$ 2 $Al(OH)_3$ + 3 $CaSO_4$

Al	2	Al	~~1~~ 2
SO_4	3	SO_4	~~1~~ 3
Ca	~~1~~ 3	Ca	~~1~~ 3
OH	~~2~~ 6	OH	~~3~~ 6

TYPES OF CHEMICAL REACTIONS:

Type 1: SINGLE REPLACEMENT	Type 2: DOUBLE REPLACEMENT
Definition: Reaction where one species replaces another (one species alone on one side and combined on the other). Ex: $3Ag + AuCl_3 \rightarrow 3AgCl + Au$ $2Cr + 3H_2SO_4 \rightarrow Cr_2(SO_4)_3 + 3H_2$ $2Cr + 3FeCO_3 \rightarrow Cr_2(CO_3)_3 + 3Fe$ Will look like: $A + BC \rightarrow AC + B$	Definition: Reaction where compounds react, switch partners and produce 2 new compounds. Ex: $Pb(NO_3)_2 + 2NaCl \rightarrow PbCl_2 + 2NaNO_3$ $Na_3PO_4 + 3AgNO_3 \rightarrow Ag_3PO_4 + 3NaNO_3$ $K_2CO_3 + 2AgNO_3 \rightarrow Ag_2CO_3 + 2KNO_3$ Will look like: $AB + CD \rightarrow AC + BD$
Type 3: SYNTHESIS Definition: Reaction where we take more than one reactant and create one product Ex: $4Al + 3O_2 \rightarrow 2Al_2O_3$ $2H_2 + O_2 \rightarrow 2H_2O$ Will look like: $A + B \rightarrow AB$	**Type 4: DECOMPOSITION** Definition: Reaction where we take one reactant and create 2 products. Ex: $BaCO_3 \rightarrow BaO + CO_2$ $2H_2O_2 \rightarrow 2H_2O + O_2$ $2Bi(OH)_3 \rightarrow Bi_2O_3 + 3H_2O$ Will look like: $AB \rightarrow A + B$

#5. Combustion

Definition: a chemical reaction between oxygen and an organic species

Ex: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + \text{Heat}$

Mole-Mole Problems: An Introduction

A chemical equation is basically the "recipe" for a reaction. The **COEFFICIENTS** in an equation tell us the amounts of **REACTANTS** and **PRODUCTS** we need to make the recipe work. Reactants in an equation react in specific **RATIOS** to produce a specific amount of products. Below is a recipe for sugar cookies:

3 eggs + 1 cup of flour + 2 cups sugar → 24 cookies

Let's simplify this to: 3E + 1F + 2S → 24C

If we massed the eggs, flour and sugar, they should (in a perfect world) equal the mass of the cookies. (This illustrates the LAW OF CONSERVATION OF MASS)

3E + 1F + 2S = 24C
(200 g + 160g + 240g = 600g)

Q: If you had to bake 48 cookies, how many eggs would you need?
A: 6 (double it)

- **Method for solving mole-mole problems**: set up a proportion using your known and unknown values, then cross-multiply and solve for your unknown.

Ex 1: Set up the proportion from the Q & A above and solve.

$$\overset{x}{3E} + 1F + 2S \rightarrow \overset{48}{24C}$$

$$\frac{x}{3} = \frac{48}{24}$$

$$\frac{24x}{24} = \frac{142}{24} = 6 \text{ eggs}$$

Ex 2: If you have 10 eggs and an infinite amount of sugar and flour, what is the greatest number of cookies you can make?

$$\frac{10}{3E} + 1F + 2S \rightarrow \frac{x}{24}C$$

$$\frac{10}{3} = \frac{x}{24} \qquad \frac{\cancel{3}x}{\cancel{3}} = \frac{240}{3}$$

$$\boxed{x = 80 \text{ cookies}}$$

We can use the process we used with the cookie recipe and apply it to chemical equations. The only difference is we* ***ALWAYS* *check to make sure we are starting with a BALANCED CHEMICAL EQUATION*

Ex 3: Consider the following formula:

$$\frac{x}{1\,N_2} + 3H_2 \rightarrow \frac{10}{2NH_3}$$

How many moles of nitrogen gas (N_2) would be needed to produce 10 moles of ammonia (NH_3)?

$$\frac{x}{1} = \frac{10}{2} \qquad 2x = 10$$

$$\boxed{x = 5 \text{ mol } N_2}$$

Mole-Mole Practice

Use the following equation to answer questions 1-3:

$$C_3H_8 + \overset{x}{5O_2} \rightarrow \overset{20}{3CO_2} + 4H_2O$$

1) If 12 moles of C_3H_8 react completely, how many moles of H_2O are formed?

$\frac{12}{1} = \frac{x}{4}$ $x = 12 \times 4$
$= 48$ mol of H_2O

2) If 20 moles of CO_2 are formed, how many moles of O_2 reacted?

$\frac{20}{3} = \frac{x}{5}$ $\frac{\cancel{3}x}{\cancel{3}} = \frac{100}{3} = \boxed{33.3 \text{ mol } O_2}$

3) If 8 moles of O_2 react completely, how many moles of H_2O are formed?

$\frac{8}{5} = \frac{x}{4}$ $\frac{\cancel{5}x}{\cancel{5}} = \frac{32}{5} = \boxed{6.4 \text{ mol } H_2O}$

Use the following equation to answer questions 4-7:

$$N_2 + 3H_2 \rightarrow 2NH_3$$

4) If 2.5 moles of N_2 react completely, how many moles of NH_3 are formed?

$\frac{2.5}{1} = \frac{x}{2}$ $x = 2 \times 2.5 = \boxed{5.0 \text{ mol } NH_3}$

5) If 9 moles of NH_3 are formed, how many moles of H_2 reacted?

$\frac{9}{2} = \frac{x}{3}$ $\frac{\cancel{2}x}{\cancel{2}} = \frac{27}{2} = \boxed{13.5 \text{ mol } H_2}$

6) If 3.5 moles of NH_3 are formed, how many moles of N_2 reacted?

$\frac{3.5}{2} = \frac{x}{1}$ $\frac{\cancel{2}x}{\cancel{2}} = \frac{3.5}{2} = \boxed{1.75 \text{ mol } N_2}$

7) How many grams of N_2 are reacted when 3.5 moles of NH_3 are formed?

step ① $\frac{x}{1} = \frac{3.5}{2}$ $\frac{\cancel{2}x}{\cancel{2}} = \frac{3.5}{2} = 1.75$ mol N_2

step ② gfm
$2N = 2(14.0067)$
$= 28.0134$ g/mol

Tablet

~~gfm~~ moles = $\frac{\text{given mass}}{\cancel{\text{gfm}}}$ ~~gfm~~

given mass = moles × gfm
$= 1.75 \times 28.0134$
$= \boxed{49.02 \text{ g of } N_2}$

Determining EMPIRICAL Formulas: simplest reduced

Empirical Formula = the reduced formula; a formula whose subscripts cannot be reduced any further

Molecular Formula = the actual formula for a compound; subscripts represent actual quantity of atoms present

*divide subscripts by common factor

Molecular Formula	Empirical Formula
N_2O_4 →	NO_2
C_3H_9	CH_3
$C_6H_{12}O_6$	CH_2O
B_4H_{10}	B_2H_5
C_5H_{12}	C_5H_{12}

Practice: Determining empirical formula from molecular formula.

a) H_2O same	h) C_2H_6 CH_3
b) N_2O_2 NO	i) Na_2SO_4 same
c) C_3H_8 same	j) C_6H_5N same
d) $Fe(CO)_3$ same	k) P_2O_5 same
e) C_5H_{10} CH_2	l) H_2O_2 HO
f) NH_3 same	m) SeO_3 same
g) $CaBr_2$ same	n) LiCl same

Calculating Empirical Formula from % Mass (100%)

Step 1: Always assume you have a 100 g sample (The total % for the compound must = 100, so we can just change the units from % to *g*)
Step 2: Convert grams to moles. (formulas = mole ratios)
Step 3: Divide all mole numbers by the smallest mole number.

Get whole #'s → Keep the same ratio

Ex. A compound is 46.2 % mass carbon and 53.8 % mass nitrogen. What is its empirical formula?

Step 1: Assume a 100 g sample.

$$\boxed{\begin{array}{l} 46.2\ \%\ C = 46.2\ g\ C \\ 53.8\ \%\ N = 53.8\ g\ N \end{array}}$$

$$mol = \frac{mass}{gfm}$$

Step 2: Convert grams to moles (have grams, need moles)

$$mol\ of\ C = \frac{46.2}{12.011} = 3.85\ mol\ C$$

$$mol\ of\ N = \frac{53.8}{14.0067} = 3.84\ mol\ N$$

But we must have **WHOLE NUMBERS** for **SUBSCRIPTS**.

Step 3: Divide each mole number by the smallest mole number (We will round in this step to the nearest integer if it's super close).

For C: = $\frac{3.85}{3.84}$ mol C = 1.0026 ≈ 1:1 ratio

For N: = $\frac{3.84}{3.84}$ mol N = 1.00

So, the empirical formula for our compound is CN

Practice: Determine the empirical formula for each compound below.

1) A compound contains 24.0 g C and 32.0 g O. Calculate its empirical formula. (Hint: start with step 2)

mol of $C = \frac{24}{12.011} = 1.998$ $\quad$ C $\frac{1.998}{1.998} = 1$

mol of $O = \frac{32}{15.9994} = 2.000$ $\quad$ O $\frac{2.000}{1.998} = 1.001$

$\approx 1:1$ ratio $\quad \boxed{CO}$

2) A compound contains 0.50 moles of carbon for each 1.0 mole of hydrogen. Calculate the empirical formula of this compound. (Hint: start with step 3)

$C = \frac{0.5}{0.5} = 1$ mol C

$H = \frac{1.0}{0.5} = 2$ mol H

$\boxed{CH_2}$

3) A compound contains 14.6% C and 85.4% Cl by mass. Calculate the empirical formula of this compound.

mol of $C = \frac{14.6}{12.0111} = 1.216$ $\quad$ C $\frac{1.216}{1.216} = 1$

mol of $Cl = \frac{85.4}{35.453} = 2.409$ $\quad$ Cl $\frac{2.409}{1.216} = 1.98$

$\approx 1:2$ ratio $\quad \boxed{CCl_2}$

4) 32.8% chromium and 67.2% chlorine.

mol of $Cr = \frac{32.8}{51.996} = 0.6308$ $\quad$ Cr $\frac{0.6308}{0.6308} = 1$

mol of $Cl = \frac{67.2}{35.453} = 1.895$ $\quad$ Cl $\frac{1.895}{0.6308} = 3.004$

$\approx 1:3$ ratio $\quad \boxed{CrCl_3}$

5) 67.1% zinc and the rest is oxygen.

mol of $Zn = \frac{67.1}{65.39} = 1.026$ $\quad$ $Zn = \frac{1.026}{1.026} = 1$

$100 - 67.1 = 32.9$

mol of $O = \frac{32.9}{15.9994} = 2.056$ $\quad$ $O = \frac{2.056}{1.026} = 2.003$

$\approx 1:2$ ratio $\quad \boxed{ZnO_2}$

Determining MOLECULAR Formulas:

So far, we know how to:
1. Find an empirical formula from percent mass
2. Find an empirical formula from a molecular formula

But how do we find out the molecular formula from an empirical formula?

Ex: A compound is 80.0 % C and 20.0 % H by mass. If its molecular mass is 75.0 g, what is its empirical formula? What is its molecular formula?

First, we must determine the empirical formula using the 3-step process.

Step 1: Assume a 100 g sample.
80.0 % C = 80.0 g C
20.0 % H = 20.0 g H

Step 2: Convert grams to moles (have grams, need moles)

mol of C $= \frac{80}{12.011} = 6.66$ C $\frac{6.66}{6.66} = 1$

mol of H $= \frac{20}{1.00794} = 19.8$ H $\frac{19.8}{6.66} = 2.97$

≈ 1:3 ratio

Step 3: Divide each mole number by the smallest mole number and round to the nearest integer

For C: = 1

For H: = 3

So, the empirical formula for our compound is CH_3.

Continued on next page ↓

$(12.0)+3(1.0)=15.0g$

Now we can determine the molecular formula:

- Empirical mass (the mass of 1 mole of CH_3) = 15.0 g
- Molecular mass = 75 g * from original question
- Molecular mass is 5 times larger than empirical mass $\frac{75}{15}=5$
- Molecular formula must be 5 times larger than empirical formula
- Multiply **ALL** the subscripts in our empirical formula by 5 CH_3

CH_3

Thus, our molecular formula is C_5H_{15}

Practice: Answer the questions below in the space provided. SHOW ALL WORK.

1) What is the molecular formula of a compound that has an empirical formula of NO_2 and molecular mass of 92.0 g?

Emp form mass

$NO_2 = 46.0g$

$(14)+2(16)$

$\frac{92.0}{46.0}=2$

NO_2 2 $\rightarrow \boxed{N_2O_4}$

2) A compound is 50% sulfur and 50% oxygen by mass. Calculate the empirical formula. If its molecular mass is 128 g, determine its molecular formula.

E.f. mass $SO_2 = 64g$

mol of S = $\frac{50g}{32.065} = 1.56$ mol

mol of O = $\frac{50g}{15.9994} = 3.13$ mol

$S=\frac{1.56}{1.56}=1$

$O=\frac{3.13}{1.56}=2.006$

≈ 1:2 ratio SO_2

$\frac{128}{64}=2$ SO_2 $\rightarrow \boxed{S_2O_4}$

3) A compound is 63.6% N and 36.4% O by mass. Calculate its empirical formula. List three possible molecular formulas for this compound.

$\boxed{N_2O, N_4O_2, N_6O_3}$

$\frac{63.6g\ N}{} \left| \frac{1\ mol\ N}{14.0067g} \right. = \frac{4.54\ mol\ N}{2.28}$

≈ 2:1 ratio

$\frac{36.4g\ O}{} \left| \frac{mol\ O}{15.9994g} \right. = \frac{2.28}{2.28}$ mol O

$\boxed{N_2O}$ EF

4) A compound is 92.3% carbon and 7.7% hydrogen by mass. Calculate its empirical formula. If the molecular mass is 78.0 g, determine its molecular formula.

5) A compound is 74.0% C, 8.7% H, and 17.3% N. Calculate its empirical formula. Its molecular mass is 162 g. Determine its molecular formula.