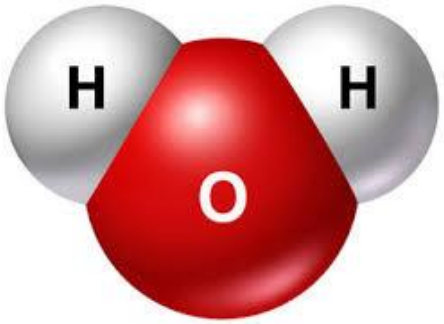
[6.3]
Percent Composition, Empirical Formula, Molecular Formula

Percent Composition of Compounds

- We can find the percent composition, **how much** of each element, is in a compound
- Steps:

1. Calculate total molar mass
2. Divide the mass of each element over the total molar mass
3. Multiply by 100%



<u>In a water molecule:</u>

Total molar mass = 16.0 + 1.0 + 1.0 g/mol
= 18.0 g/mol

Mass of O = 16.0 g/mol
Mass of H = 1.0 + 1.0 g/mol
= **2.0 g/mol**

Percent Composition of Water

Mass of O = 16.0 g/mol

Mass of H = 1.0 + 1.0 g/mol

= **2.0 g/mol**

1. **Total molar mass** = 16.0g/mol + 2.0 g/mol

   = 18.0 g/mol

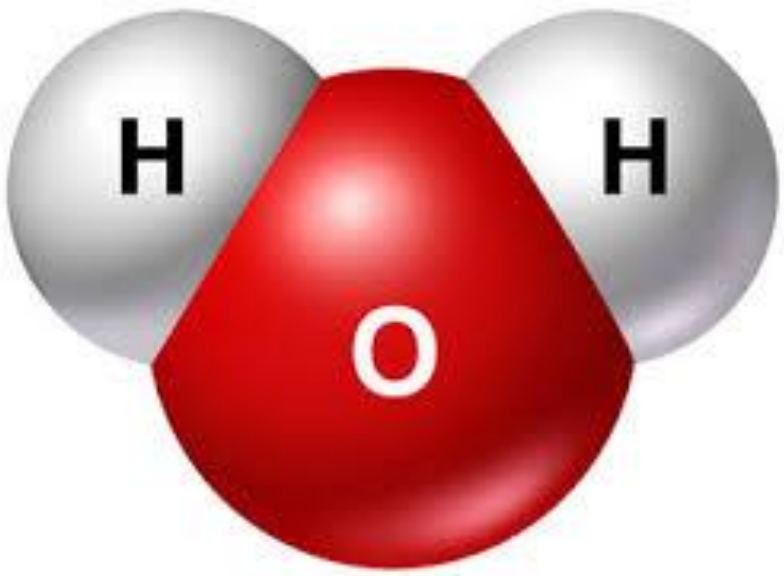
2. **O** = (16.0 ÷ 18.0) g/mol = 0.889

   H = (2.0 ÷ 18.0) g/mol =0.111

3. **%O =** 0.889 x 100% = 88.9%

   %H = 0.111 x 100% = 11.1%

   =**100.0%**



Sometimes, the total percentage may be 99.9% or 99.9%, that is okay, just round to 100.0%

Percent Composition of Methane, CH_4

- **Find the percent composition of each element in CH_4**

Percent Composition of Methane, CH_4

- **Find the percent composition of each element in CH_4**

C= 12.0 g /mol

H = (4 x 1.0 g/mol) = 4.0 g/mol

Total Molar Mass: 12.0 g /mol + 4.0 g/mol = **16.0 g/mol**

% of C = (12.0 g/mol ÷ 16.0 g/mol) x 100% = **75.0% is C**

% of H = (4.0 g/mol ÷ 16.0 g/mol) x 100% = **25.0% is H**

Practice Problems

- Find the percent composition of each element in KCN

KCN= (39.1+12.0+14.0) g/mol=65.1 g/mol

K: (39.1g/mol÷65.1g/mol)x100%=60.1%

C: (12.0 g/mol ÷65.1 g/mol) x 100% = 18.4%

N: (14.0 g/mol ÷65.1 g/mol) x 100% =21.5%

- Find the percent composition of each element in Na_3PO_4

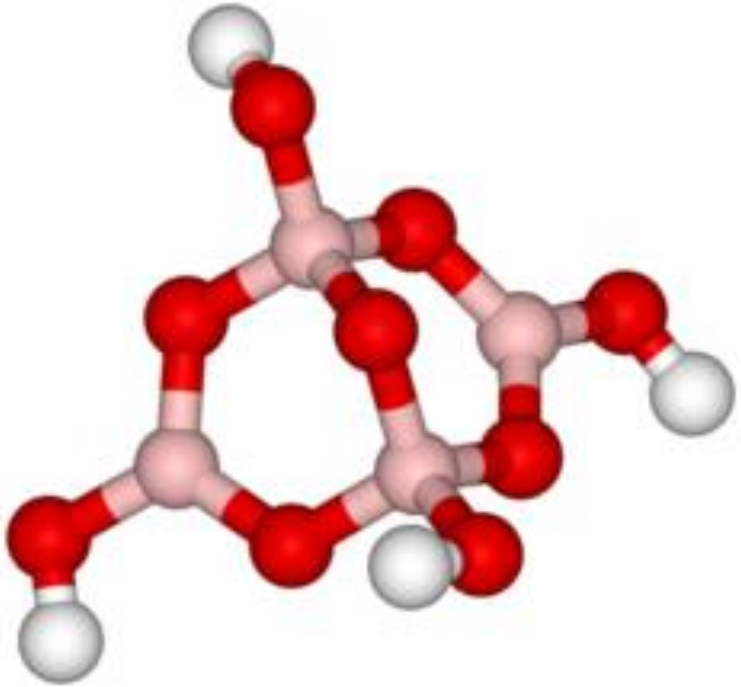
Na_3PO_4= [(3x 23.0)+31.0+(4x16.0)] g/mol = 164 g/mol

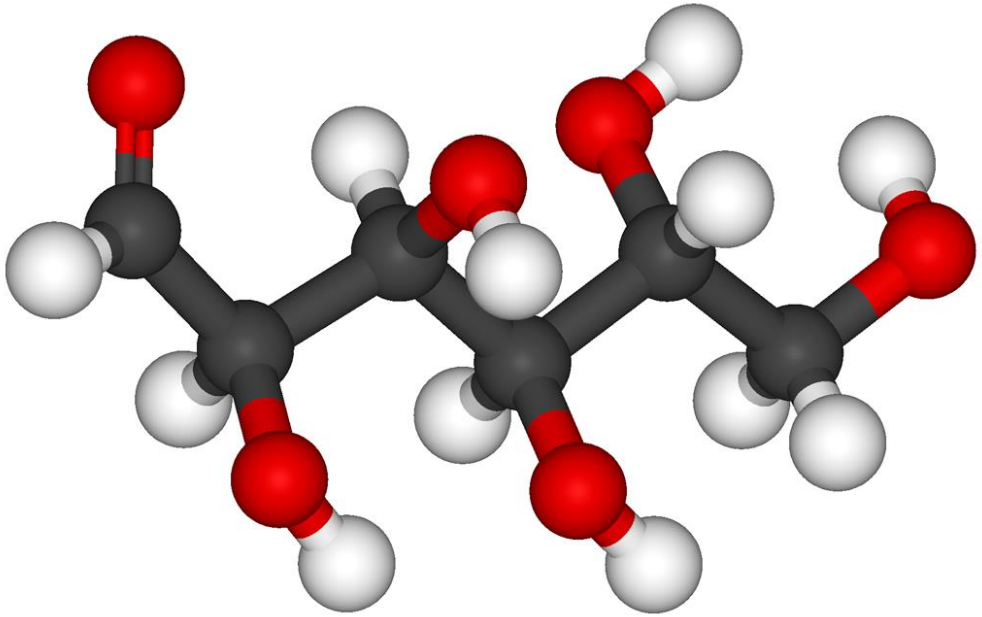
Na: [(3x 23.0g/mol) ÷ 164 g/mol]x 100%= 42.1%

P: (31.0 g/mol ÷ 164 g/mol) x 100%=18.9%

O: [(4 x 16.0 g/mol) ÷ 164 g/mol] x 100% = 39.0%

Empirical Formulas & Molecular Formulas





Empirical Formula and Molecular Formula

Empirical Formula **vs.** **Molecular Formula**

The <u>simplest ratio</u> of a compound reduced to their lowest form

Example: HO

The <u>actual chemical composition</u> if the compound, not necessarily reduced to the simplest ratio

Example: H_2O_2

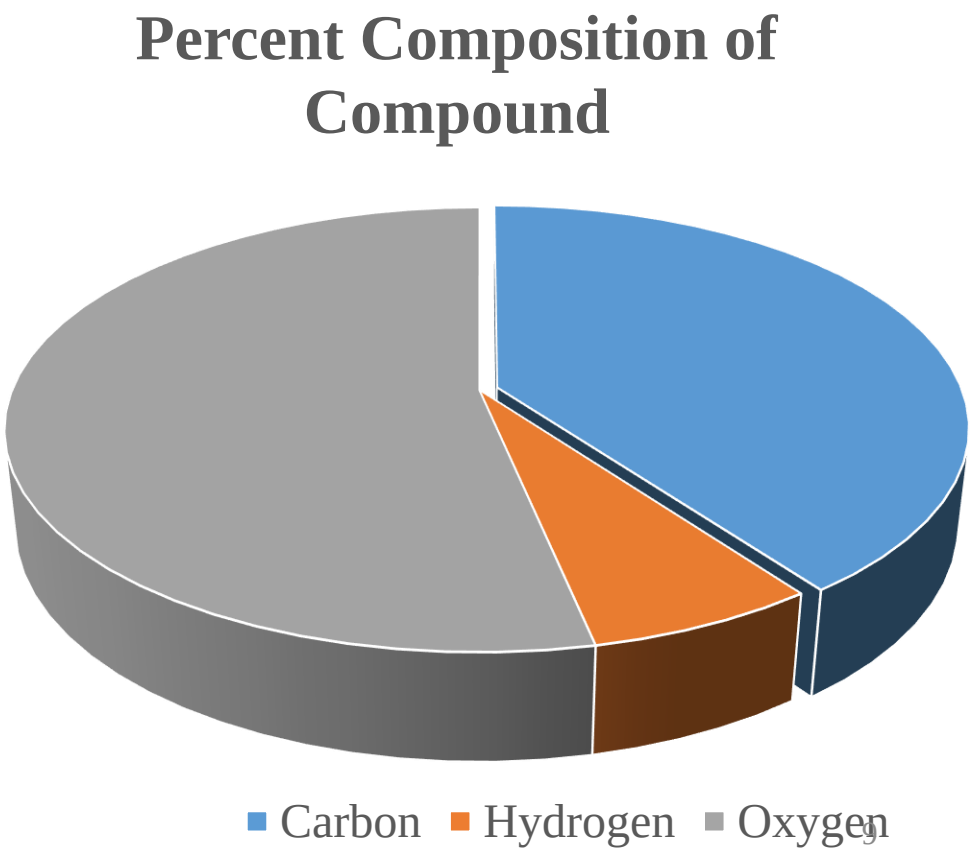
Empirical Formula

Let's say, you are given the following percent compositions of a compound:

40.00 % Carbon

6.72 % Hydrogen

53.29 % Oxygen

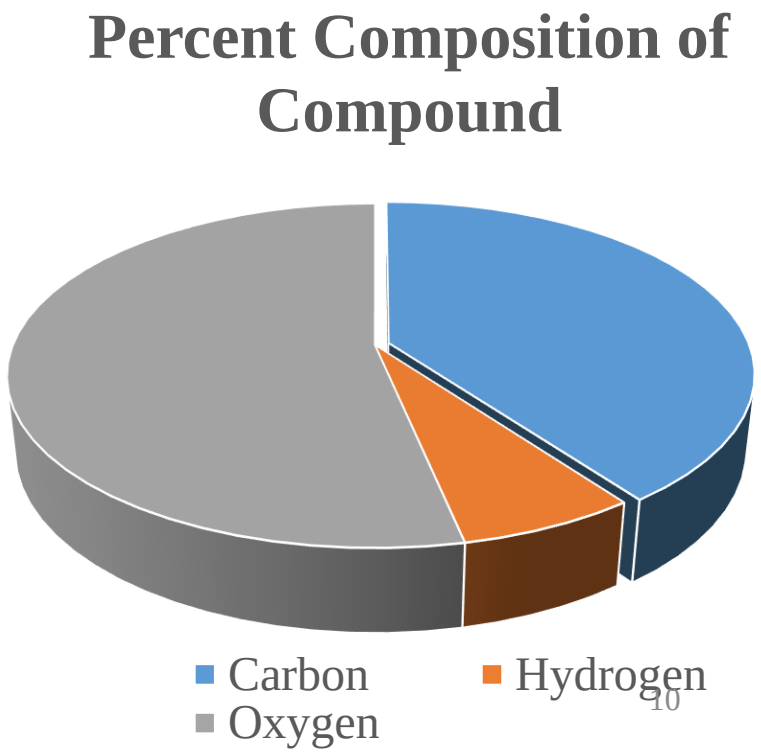


Empirical Formula

Let's pretend instead of 100% of the compound, we have 100 g of the compound.

Out of this 100 g, how many grams would be carbon, hydrogen and oxygen?

40.00 % Carbon →40.00 g
6.72 % Hydrogen → 6.72 g
53.29 % Oxygen → 53.29 g



Empirical Formulas

Now, let's use the new **mass** values that we found and try to find the **number of moles**!

What is the relationship between **mass and moles**?
How can we figure out the moles value?

Use MOLAR MASS

$$\boxed{n = \frac{m}{MM}}$$

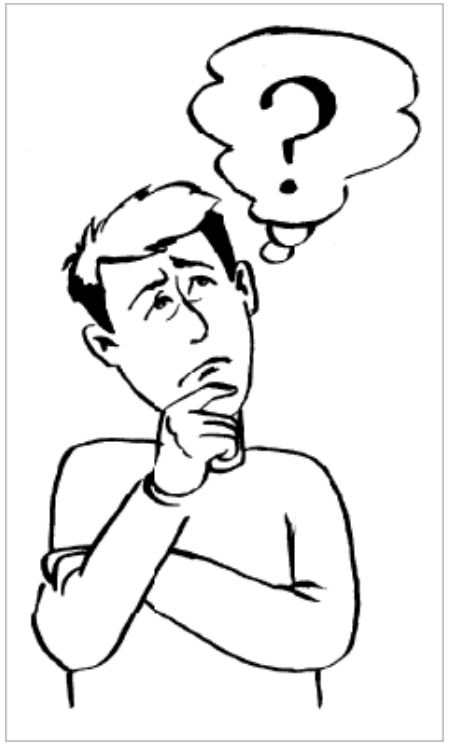
Empirical Formulas

Find the mole values based on the masses and molar masses of each element:

moles of C = 40.00 g ÷12.0 g/mol = **3.33 mol**

moles of H = 6.72 g ÷ 1.0 g/mol = **6.72 mol**

moles of O = 53.29 g ÷ 16.0 g/mol = **3.33 mol**



How can we get a ratio of the elements using these mole values?

Empirical Formulas

Divide **all** of the mole values by the lowest mole value to find the ratios! In this case, the lowest value is 3.33 mol.

moles of C = 40.00 g ÷ 12.0 g/mol = **3.33 mol**

moles of H = 6.72 g ÷ 1.0 g/mol = **6.72 mol**

moles of O = 53.29 g ÷ 16.0 g/mol = **3.33 mol**

Ratio of C = 3.33 mol/3.33 mol = 1

Ratio of H = 6.72 mol/3.33 mol = 2

Ratio of O = 3.33 mol/3.33 mol = 1

Therefore, the empirical formula for the compound is:

$$\boxed{CH_2O}$$

Empirical Formula Example

A compound is **75.46% carbon**, **4.43% hydrogen**, and **20.10% oxygen** by mass. What is its **empirical formula**?

<u>Step 1:</u> Change the percentages to masses per 100.00 g

mass of C = 75.46 g
mass of H = 4.43 g
mass of O = 20.10 g

<u>Step 2:</u> Find the moles of each element using mass & molar mass

moles of C = 75.46 g / 12.0 g = 6.288 mol
moles of H = 4.43 g /1.0 g = 4.43 mol
moles of O = 20.10 g / 16.0 g= 1.256 mol

<u>Step 3</u>: Divide all the mole values by the lowest mole value (In this case it is 1.256 mol)

Ratio of C = 6.288 mol /1.256 mol
Ratio of H = 4.43 mol / 1.256 mol
Ratio of O = 1.256 mol / 1.256 mol

<u>Step 4</u>: We don't have all whole numbers! So we need to multiply to get whole numbers.
Ratio of C = 5.006 mol x 2 = 10
Ratio of H = 3.52 x 2= 7
Ratio of O = 1.00 x 2 = 2

<u>Step 5</u>: Write your empirical formula with the correct subscripts:

$$\boxed{C_{10}H_7O_2}$$

Practice Problem #1

A compound is found to contain **23.3% magnesium**, **30.7% sulfur** and **46.0% oxygen**. What is the **empirical formula** of this compound?

Mg: 23.3% → 23.3 g
S:30.7% → 30.7 g
O: 46.0% → 46.0 g

Mg: 23.3 g÷24.3g/mol=0.959 mol
S:30.7 g ÷32.1g/mol = 0.956 mol
O: 46.0 g ÷16.0g/mol = 2.88 mol

Mg: 0.959 mol ÷ 0.956 mol = 1
S: 0.956 mol ÷ 0.956 mol = 1
O: 2.88 mol ÷ 0.956 mol = 3

$MgSO_3$

Practice Problem #2:

What is the empirical formula of a compound containing 58.5% C, 7.3% H, and 34.1% N?

C: 58.5% → 58.5 g

H: 7.30% → 7.3 g

N: 34.1% → 34.1 g

C: 58.5 g ÷ 12.0 g/mol = 4.875 mol

H: 7.3 g ÷ 1.0 g/mol = 7.3 mol

N: 34.1 g ÷ 14.0 g/mol = 2.44 mol

C: 4.875 mol ÷ 2.44 mol = 1.99 ≈ 2.00

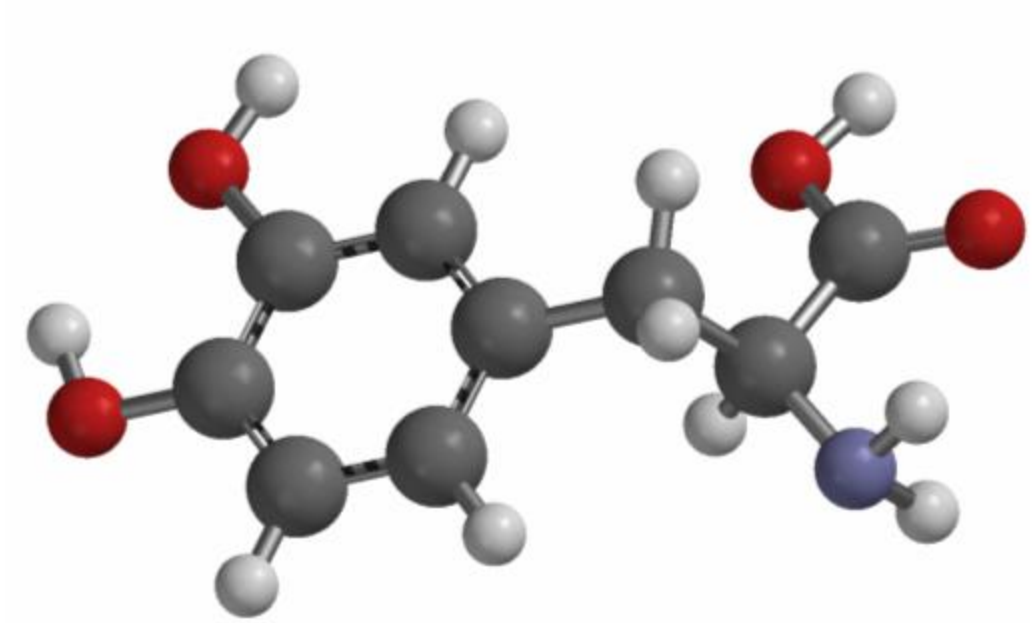
H: 7.3 mol ÷ 2.44 mol = 2.99 ≈ 3.00

N: 2.44 mol ÷ 2.44 mol = 1.00

C_2H_3N

Molecular Formula

- **Molecular Formula**: the actual composition of the compound
- To solve for molecular formula we need to know:
  - Empirical formula
  - Molar mass of the actual compound



Calculating for Molecular Formula

$$Z = \frac{\textbf{Molar mass of actual compound}}{\textbf{Mass of Empirical Formula}}$$

Molecular formula= **Z** x (Subscripts of Empirical Formula)

Example: A molecule has an empirical formula of HO and a molar mass of 34.0 g/mol. What is the molecular formula?

1. $Z = \frac{\textbf{34.0 g/mol}}{\textbf{17.0 g/mol}} = 2$

2. Molecular Formula = 2 x ($H_{\mathbf{1}}O_{\mathbf{1}}$) = $\boxed{H_2O_2}$

Practice Problem #1

A molecule has the empirical formula of CH_2O, and a molar mass of 180.0 g/mol. What is the molecular formula of the compound?

1. $$\mathbf{Z} = \frac{\textbf{Molar mass of actual compound}}{\textbf{Mass of Empirical Formula}} = \frac{\textbf{180.0g}}{\textbf{(12.0 g + 2.0 g + 16.0g)}} = 6$$

2. Molecular formula= **6** x (CH_2O) = $\boxed{C_6H_{12}O_6}$

Practice Problem #2:

Caffeine has the following percent composition: **carbon 49.48%**, **hydrogen 5.19%**, **oxygen 16.48%**, and **nitrogen 28.85%.** Its **molecular weight is 194.19 g/mol**. What is its **molecular formula**?

Practice Problem #2:

Caffeine has the following percent composition: **carbon 49.48%**, **hydrogen 5.19%**, **oxygen 16.48%**, and **nitrogen 28.85%.** Its **molecular weight is 194.19 g/mol**. What is its **molecular formula**?

C: 49.48% → 49.48 g

H:5.19% → 5.19 g

O: 16.48% → 16.48 g

N:28.85% → 28.85 g

C: 49.48 g ÷12.0g/mol= 4.12 mol

H: 5.19 g ÷ 1.0 g/mol = 5.19 mol

O: 16.48 g ÷ 16.0 g/mol= 1.03 mol

N: 28.85 g ÷14.0g/mol = 2.06 mol

C: 4.12 mol ÷1.03 mol = 4

H: 5.19 mol ÷1.03mol = 5.03 ≈5

O: 1.03 mol ÷1.03 mol=1

N: 2.06 mol ÷1.03 mol= 2

$C_4H_5ON_2$

Practice Problem #2… continued

- Caffeine has the following percent composition: **carbon 49.48%**, **hydrogen 5.19%**, **oxygen 16.48%**, and **nitrogen 28.85%.** Its **molecular weight is 194.19 g/mol**. What is its **molecular formula**?

$\boxed{\mathbf{C_4H_5ON_2}}$ =97.0 g/mol

1. $Z = \dfrac{\mathbf{194.19\ g/mol}}{\mathbf{97.0\ g/mol}} = 2$

2. Molecular Formula= 2 x ($C_4H_5ON_2$)= $\boxed{\mathbf{C_8H_{10}O_2N_4}}$

HOMEWORK

- Pg. 91 #41 A&B
- Pg. 93 #46 F,K,L
- Pg. 95 #47-50