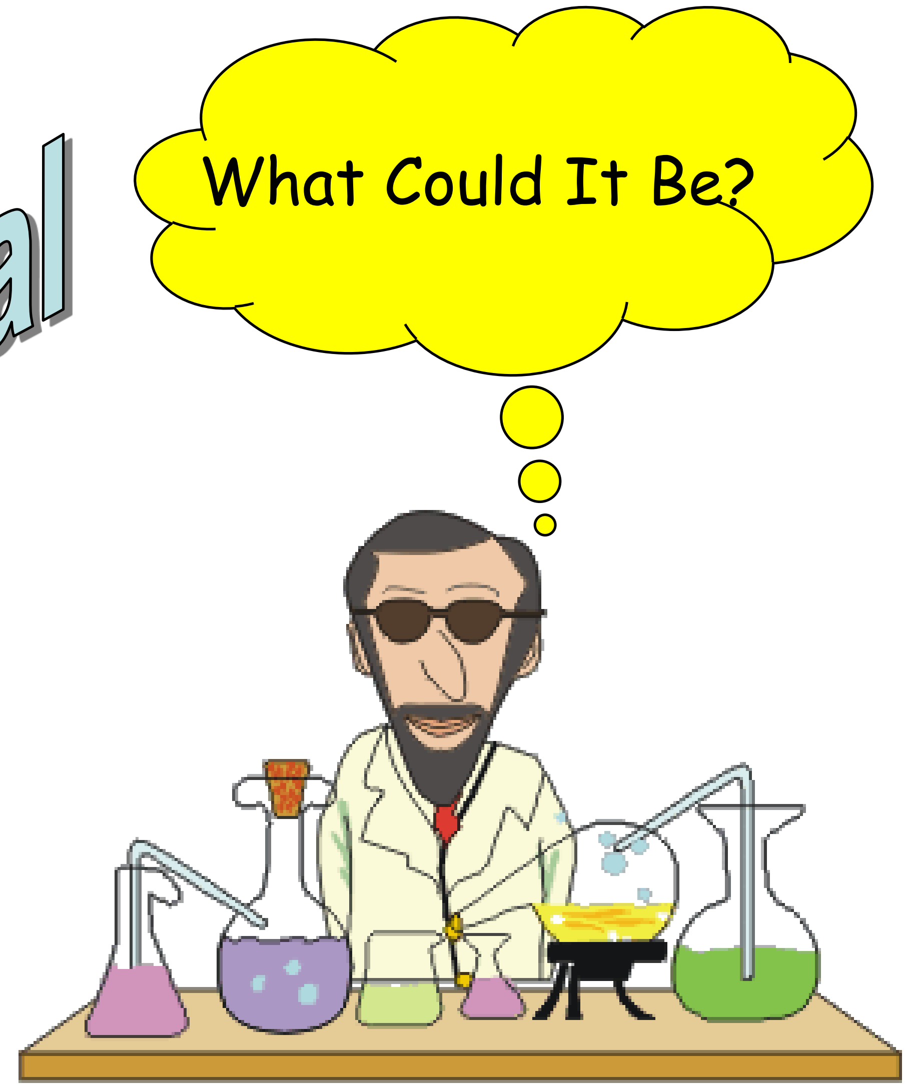
NOTES: 10.3 – Empirical and Molecular Formulas

Finding Empirical and Molecular Formulas



Empirical Formulas

- Indicate the lowest whole number ratio of the atoms in a compound:

1) Determine moles of each element present in the compound
2) Divide molar amounts by the smallest number of moles present
3) Obtain whole numbers by multiplying by integers if necessary

Calculating the Empirical Formula

Example: A compound is found to contain the following…

2.199 g Copper

0.277 g Oxygen

Calculate it's empirical formula.

Calculating the Empirical Formula

Step 1: Convert the masses to moles.

Copper: $$2.199 g\ Cu \frac{1 mole\, Cu}{63.5 g\, Cu} = 0.03460\ moles\, Cu$$

Oxygen: $$0.277 g\ O \frac{1 mole\ O}{16.0 g\ O} = 0.01731\ moles\ O$$

Calculating the Empirical Formula

Step 2: Divide all the moles by the smallest value. This gives the "mole ratio"

$$\frac{0.03460\ \text{mol Cu}}{0.0173\,\text{mol}} = 1.999\ \mathrm{C}u$$

$$\frac{0.0173\ m\text{ol}\,O}{0.0173\,\text{mol}} = 1.000O$$

Calculating the Empirical Formula

Step 3: Round off these numbers, they become the subscripts for the elements.

$$Cu_2O$$

Empirical Formula Example Problems:

1) A compound is analyzed and found to contain 79.8 g C and 20.2 g H. Determine the empirical formula.

2) What is the empirical formula of a compound that is 25.9 g nitrogen and 74.1 g oxygen?

1) A compound is analyzed and found to contain 79.8 g C and 20.2 g H. Determine the empirical formula.

$$79.8g\ C \bullet \frac{1 mole\, C}{12.0 g\, C} = 6.65\ moles\ C$$

$$20.2g\ H \bullet \frac{1 mole\, H}{1.0 g\, H} = 20.2\ moles\ H$$

1) A compound is analyzed and found to contain 79.8 g C and 20.2 g H. Determine the empirical formula.

$$\frac{6.65 \ \mathrm{mol\,C}}{6.65\,\mathrm{mol}} = 1.0 \ \mathrm{C}$$

$$\frac{20.2 \ mol\,H}{6.65\,\mathrm{mol}} = 3.0H$$

1) A compound is analyzed and found to contain 79.8 g C and 20.2 g H. Determine the empirical formula.

EMPIRICAL FORMULA = CH_3

2) What is the empirical formula of a compound that is 25.9 g nitrogen and 74.1 g oxygen?

$$25.9g\ \ N \bullet \frac{1 mole\, N}{14.0 g\ N} = 1.85\ moles\ N$$

$$74.1g\ \ O \bullet \frac{1 mole\, O}{16.0 g\, O} = 4.63\ moles\, O$$

2) What is the empirical formula of a compound that is 25.9 g nitrogen and 74.1 g oxygen?

$$\frac{1.85\ \mathrm{mol\,N}}{1.85\,\mathrm{mol}} = 1.0\ \mathrm{N}$$

$$\frac{4.63\ m\mathrm{ol}\,O}{1.85\,\mathrm{mol}} = 2.5O$$

2) What is the empirical formula of a compound that is 25.9 g nitrogen and 74.1 g oxygen?

EMPIRICAL FORMULA = $NO_{2.5}$

Is this formula acceptable?

2) What is the empirical formula of a compound that is 25.9 g nitrogen and 74.1 g oxygen?

EMPIRICAL FORMULA = $NO_{2.5}$

Is this formula acceptable? NO!

(must be WHOLE NUMBERS)

So, multiply by 2.

2) What is the empirical formula of a compound that is 25.9 g nitrogen and 74.1 g oxygen?

EMPIRICAL FORMULA:

$$2 \times (NO_{2.5}) \quad = \quad \boxed{N_2O_5}$$

Molecular Formulas

- Indicates the **actual** formula for a compound; the true formula
- Will be a multiple of the empirical formula

- **EXAMPLE**: glucose

-empirical: CH_2O

-molecular: $C_6H_{12}O_6$

Molecular Formula Example:

Calculate the molecular formula of the compound whose molar mass is 60.0 g and the empirical formula is CH_4N.

Molecular Formula Example:

Calculate the molecular formula of the compound whose molar mass is 60.0 g and the empirical formula is CH_4N.

CH_4N has a MW of: 12 + 1(4) + 14 = 30.0 g

Actual MW is: 60.0 g

Molecular Formula Example:

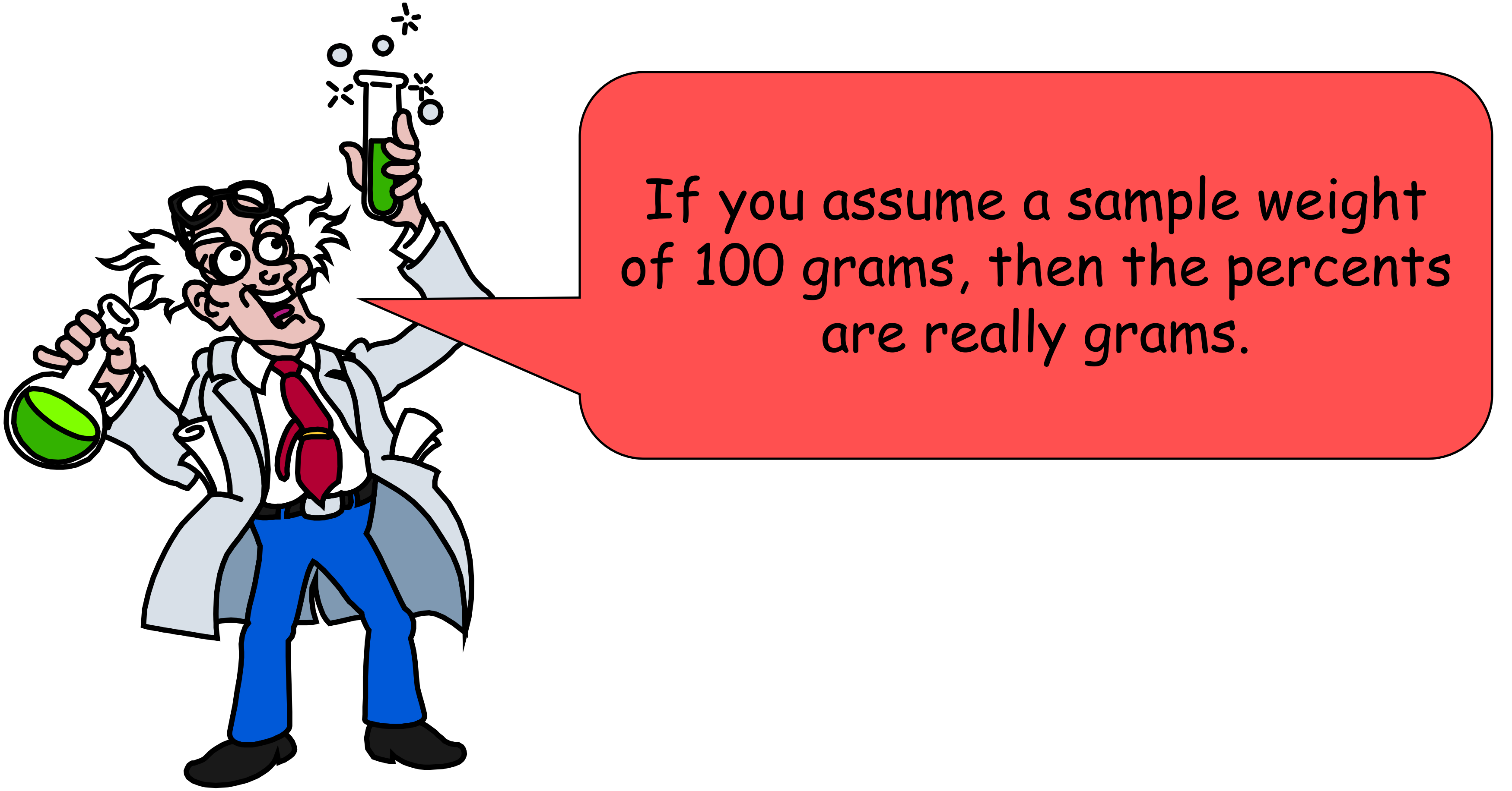
Calculate the molecular formula of the compound whose molar mass is 60.0 g and the empirical formula is CH_4N.

60.0 / 30.0 = 2

So, multiply the empirical form. by 2

2 x (CH_4N) = $\boxed{C_2H_8N_2}$

Example: A material is found to be composed of 38.7% Carbon, 51.6% Oxygen, and 9.7% Hydrogen. By other means, it is known that the molecular weight is 62.0 g/mol Calculate the empirical and molecular formula for the compound.



Example: A material is found to be composed of 38.7% Carbon, 51.6% Oxygen, and 9.7% Hydrogen. By other means, it is known that the molecular weight is 62.0 Calculate the empirical and molecular formula for the compound.

Carbon: $$38.7\,grams \cdot \frac{1\,mole}{12.0\,grams} = 3.23\,mol$$

Oxygen: $$51.6\,grams \cdot \frac{1\,mole}{16.0\,grams} = 3.23\,mol$$

Hydrogen: $$9.7\,grams \cdot \frac{1\,mole}{1.0\,grams} = 9.7\,mol$$

Carbon: $38.7\,grams \cdot \frac{1\,mole}{\ldots ms} = \frac{3.23\,mol}{3.23} = 1.00$

Now, divide all the moles by the smallest one, 3.23

Oxygen: $51.6\,grams \cdot \frac{\ldots}{16.0\,gram\ldots} = \frac{3.23\,mol}{3.23} = 1.00$

Hydrogen: $9.7\,grams \cdot \frac{1\,mole}{1.0\,grams} = \frac{9.7\,mol}{3.23} = 3.00$

Example: A material is found to be composed of 38.7% Carbon, 51.6% Oxygen, and 9.7% Hydrogen. By other means, it is known that the molecular weight is 62.0 Calculate the empirical and molecular formula for the compound.

So, the empirical formula must be: **$\underline{CH_3O}$**

The molecular weight of the empirical formula is….

C	12 x 1	=	12 g/mol
H	1 x 3	=	3 g/mol
O	16 x 1	=	16 g/mol
			<u>31 g/mol</u>

Remember, the empirical formula is not necessarily the molecular formula!

MW of the empirical formula = <u>**31**</u>

MW of the molecular formula = <u>**62**</u>

$$Multiplying\ Factor = \frac{Molecular}{Empirical}$$

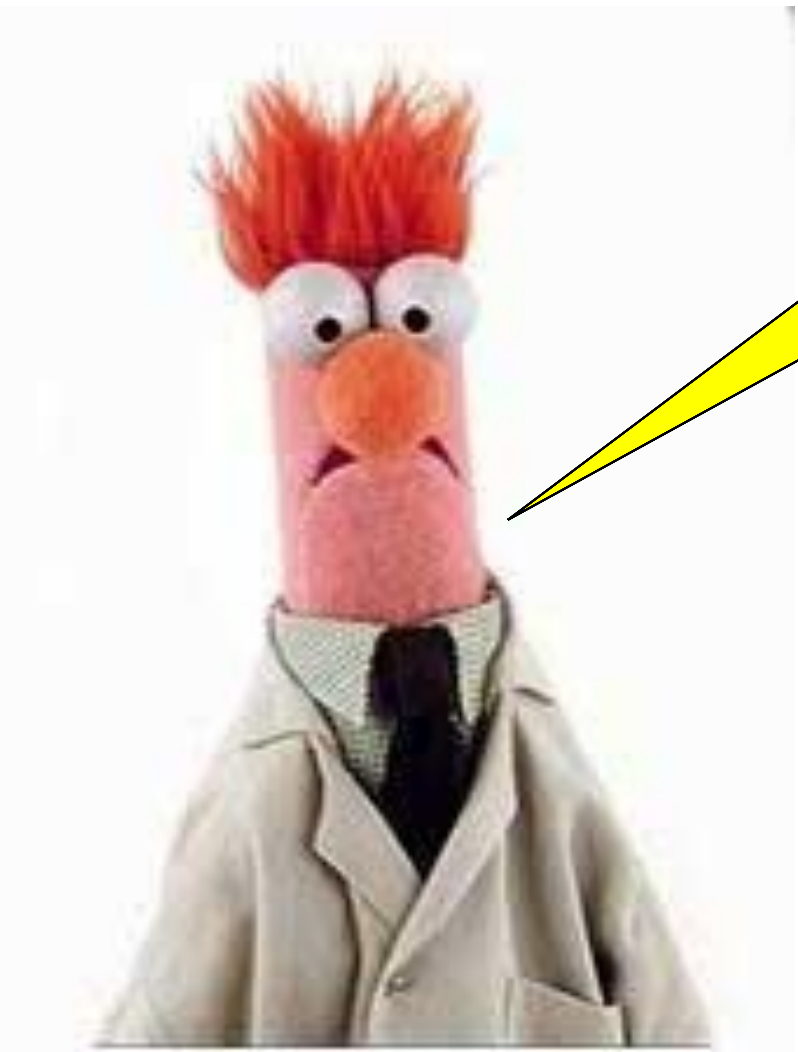
$$= \frac{62}{31} = 2$$

$$2 \times (CH_3O) = C_2H_6O_2$$

Empirical Formula

Molecular Formula

Remember, the molecular formula represents the actual formula.



What if the mole ratios don't come out even?

Example: A compound is analyzed and found to contain 2.42g aluminum and 2.15g oxygen. Calculate its empirical formula.

Aluminum: $2.42g \frac{1mol}{27g} = \frac{.0896 mol}{.0896} = 1.0$ X 2 = 2

Oxygen: $2.15g \frac{1mol}{16g} = \frac{.134 mol}{.0896} = 1.495$ X 2 = 3

Al_2O_3