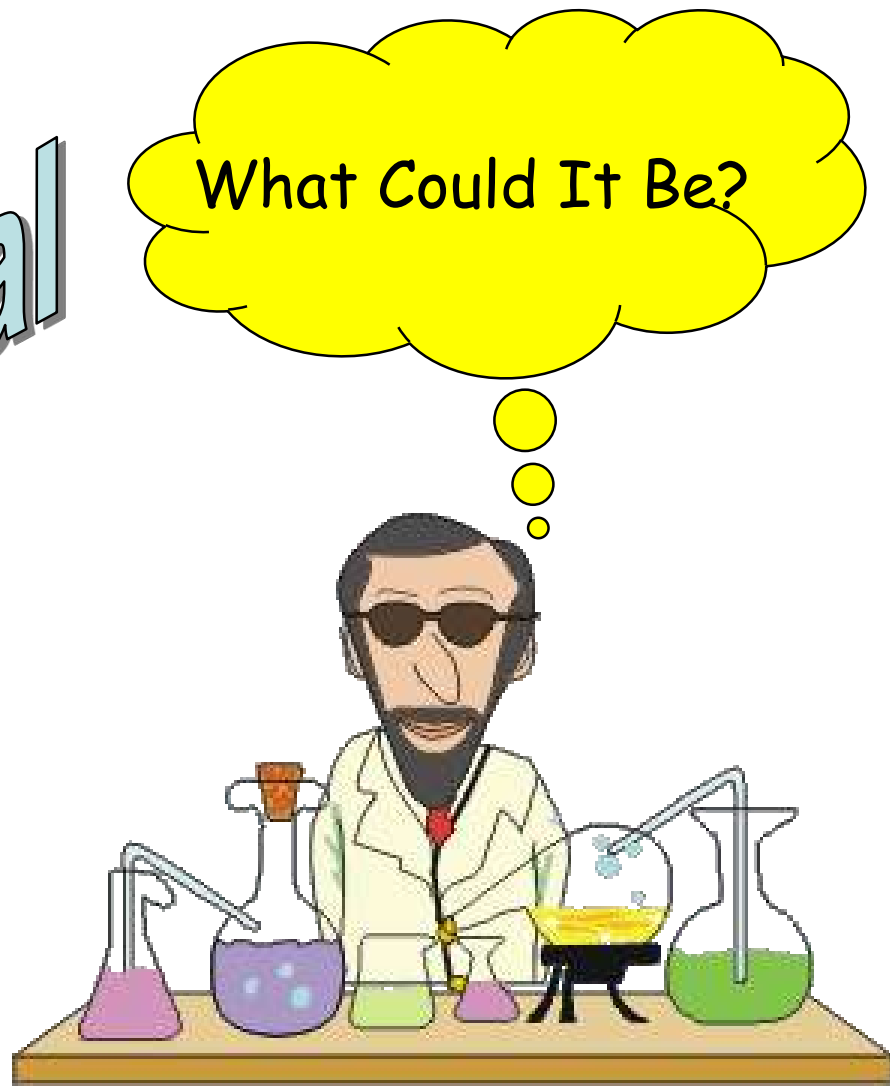


NOTES: 10.3 – Empirical and Molecular Formulas



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.

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

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

Calculating the Empirical Formula

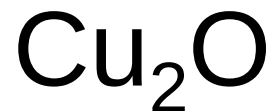
Step 2: Divide all the moles by the smallest value.
This gives the “mole ratio”

$$\frac{0.03460 \text{ mol Cu}}{0.0173 \text{ mol}} = 1.999 \text{ Cu}$$

$$\frac{0.0173 \text{ mol O}}{0.0173 \text{ mol}} = 1.000 \text{ O}$$

Calculating the Empirical Formula

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



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.

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

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

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

$$\frac{6.65 \text{ mol C}}{6.65 \text{ mol}} = 1.0 \text{ C}$$

$$\frac{20.2 \text{ mol H}}{6.65 \text{ mol}} = 3.0 \text{ H}$$

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

EMPIRICAL FORMULA = CH₃

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

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

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

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

$$\frac{1.85 \text{ mol N}}{1.85 \text{ mol}} = 1.0 \text{ N}$$

$$\frac{4.63 \text{ mol O}}{1.85 \text{ mol}} = 2.5\text{O}$$

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 = $\text{NO}_{2.5}$

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:



=



Molecular Formulas

- Indicates the actual formula for a compound; the true formula
- Will be a multiple of the empirical formula
- **EXAMPLE:** glucose
 - empirical: $\underline{\text{CH}_2\text{O}}$
 - molecular: $\underline{\text{C}_6}\underline{\text{H}_{12}}\underline{\text{O}_6}$

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: $12 + 1(4) + 14 = 30.0 \text{ g}$

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



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.



If you assume a sample weight of 100 grams, then the percents are really grams.

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.

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

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

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

$$\text{Carbon: } 38.7 \text{ grams} \cdot \frac{1 \text{ mole}}{12.01 \text{ grams}} = \frac{3.23 \text{ mol}}{3.23} = 1.00$$

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

$$\text{Oxygen: } 51.6 \text{ grams} \cdot \frac{1 \text{ mole}}{16.00 \text{ grams}} = \frac{3.23 \text{ mol}}{3.23} = 1.00$$

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



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: **CH₃O**

The molecular weight of the empirical formula is....

$$\text{C} \quad 12 \times 1 \quad = \quad 12 \text{ g/mol}$$

$$\text{H} \quad 1 \times 3 \quad = \quad 3 \text{ g/mol}$$

$$\text{O} \quad 16 \times 1 \quad = \quad 16 \text{ g/mol}$$

$$\mathbf{\underline{31 \text{ g/mol}}}$$

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



MW of the empirical formula = 31

MW of the molecular formula = 62

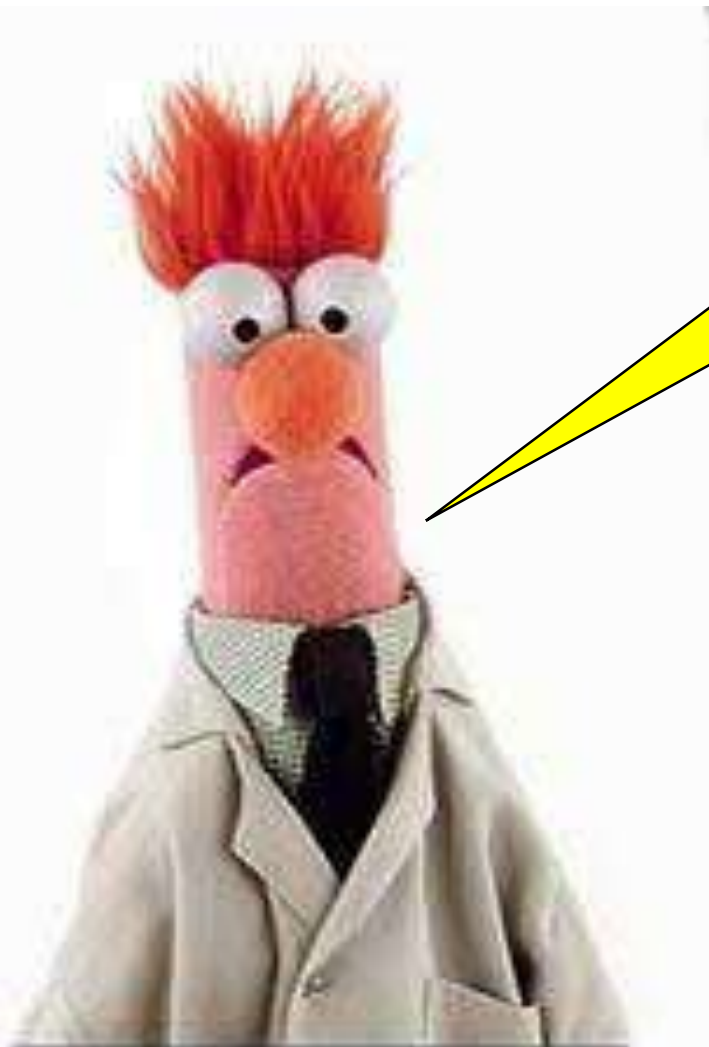
$$\text{Multipliyin g Factor} = \frac{\text{Molecular}}{\text{Empirical}}$$

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



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.

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

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

