

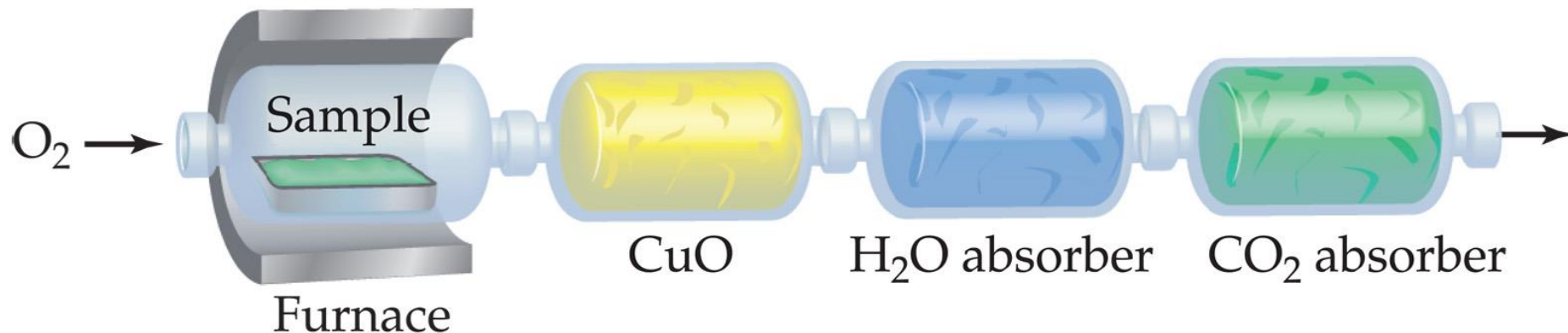
N₃₀ – COMBUSTION ANALYSIS

It's just a more involved form
of empirical formulas!

Target:

I can apply my knowledge of empirical formulas to data obtained from combustion analysis

Combustion Analysis



Compounds containing C, H and O are routinely analyzed through combustion in a chamber like this.

- C is determined from the mass of CO₂ produced.
- H is determined from the mass of H₂O produced.
- O is determined by difference after the C and H have been determined.

We have been working problems where we start with a % composition and doing this:

- **% to mass**
- **Mass to mole**
- **Divide by small**
- **Multiply until whole**

We don't HAVE to start with a % composition though...

- **As long as we can find the number of grams of each element, then we can find the empirical formula!**


So...in combustion analysis problems...

- **You will be figuring out the grams of each element in the sample using data and dimensional analysis, and do the normal empirical formula calculation!**

So now it will be like this!

- ~~% to mass~~
- Mass to mole
- Divide by small
- Multiply til whole

Use Combustion
Analysis Data and
Dimensional Analysis
to find grams



The amount of CO_2 gives the amount of C originally present in the sample compound

- All the carbon atoms in the unknown starting sample are rearranged into CO_2 product
- Why you ask? Because the law of conservation of mass is ALWAYS TRUE!

The amount of H_2O gives the amount of H originally present in the sample

- Why you ask? Why yes, that is correct. Because the law of conservation of mass is ALWAYS TRUE!
- Watch the subscript stoichiometry:
1 mol H_2O contains 2 mol H.

The amount of O originally present in the sample can be found by simple subtraction

- **Mass of sample
Mass of C
– Mass of H
= Mass of Oxygen!**
- **Why you ask? You know the answer!**

Important Points to Know

- Must know the mass of the unknown substance before burning it
- The unknown will be burnt in pure oxygen, present in large excess
- The amount of oxygen will be determined by subtraction.
- The combustion products always have CO_2 and H_2O . Might have extra products if other elements are present!
- Nitrogen product can come in different forms. N_2 , NH_3 , etc. Will be given more info if needed. Often given as a separate experiment – will need to convert all to %'s if this is the case! **Nitrogen is the problem child in combustion analysis.**
- All the carbon winds up as CO_2 and all the hydrogen winds up as H_2O .

Steps to Solve

- 1) Determine the mass of each element present in the original compound using dimensional analysis
 - Carbon is always in CO_2 in the ratio of 1 mole $\text{CO}_2 = 1$ mole C
 - Hydrogen is always in H_2O in the ratio of 1 mole $\text{H}_2\text{O} = 2$ mole H
 - Nitrogen can be (NH_3 , N_2 , N, NO_2 , etc...). If data from a separate experiment, make sure to convert masses to % values!
- 2) Subtract to solve for oxygen
Sample mass – ($\text{C}_{\text{mass}} + \text{H}_{\text{mass}} + \text{N}_{\text{mass}}$ if necessary, or any other random element)
- 3) Now continue with the Rhyme from before!
 - Mass to moles
 - Divide by small
 - Multiply until whole

Example #1

A sample of a compound that is known to contain only carbon, hydrogen, and oxygen is combusted, and the CO_2 and H_2O produced are trapped and weighed. The original sample weighed 8.38 g and yielded 16.0 g CO_2 and 9.8 g H_2O . What is the empirical formula?

Example #1

Original sample = 8.38 g and yielded 16.0 g CO₂ and 9.80 g H₂O

Moles of Carbon

16.0 g CO ₂	1 mole CO ₂	1 mole C	= 0.364 mole C
	44.0 g CO ₂	1 mole CO ₂	

Moles of Hydrogen

9.80 g H ₂ O	1 mole H ₂ O	2 mole H	= 1.09 mole H
	18.0 g H ₂ O	1 mole H ₂ O	

Example #1

Original sample = 8.38 g and yielded 16.0 g CO₂ and 9.80 g H₂O

Moles to Mass to Calculate Oxygen

0.364 mole C	12.0 g C	= 4.37 g C
	1 mole C	
1.09 mole H	1.01 g H	= 1.10 g H
	1 mole H	

Grams of Oxygen

$$8.38 \text{ g Sample} - 4.37 \text{ g C} - 1.10 \text{ g H} = 2.91 \text{ g Oxygen}$$

Example #1 Original sample = 8.38 g and yielded 16.0 g CO₂ and 9.80 g H₂O

Back to the Rhyme! Mass to moles, divide by small, multiply till whole!

$$\frac{2.91 \text{ g O}}{16.00 \text{ g O}} \times 1 \text{ mole O} = 0.182 \text{ mole O}$$

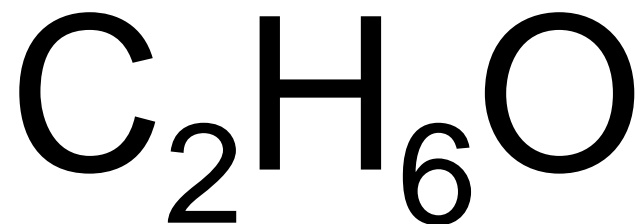
Therefore
0.364 mole C
1.09 mole H
0.182 mole O

Divide by small, multiply till whole (if needed)

$$\frac{0.364 \text{ C}}{0.182} = 2$$

$$\frac{1.09 \text{ H}}{0.182} = 5.989 \rightarrow 6$$

$$\frac{0.182 \text{ O}}{0.182} = 1$$



Example #2

Lysine is an amino acid which has the following elemental composition: C, H, O, N. In one experiment, 2.175 g of lysine was combusted to produce 3.94 g of CO_2 and 1.89 g H_2O . In a separate experiment, 1.873 g of lysine was burned to produce 0.436 g of NH_3 . The molar mass of lysine is approximately 150 g/mol. Determine the empirical and molecular formula of lysine.

Example #2

Original sample = 2.175 g and yielded 3.94 g CO₂ and 1.89 g H₂O
Nitrogen Sample = 1.873g → 0.436 g NH₂

Moles of Carbon

3.94 g CO ₂	1 mole CO ₂	1 mole C	= 0.0895 mole C
	44.01 g CO ₂	1 mole CO ₂	

Moles of Hydrogen

1.89 g H ₂ O	1 mole H ₂ O	2 mole H	= 0.2098 mole H
	18.015 g H ₂ O	1 mole H ₂ O	

Moles of Nitrogen – data from other experiment!

0.436 g NH ₂	1 mole NH ₂	1 mole N	= 0.02721 mole N
	16.023 g NH ₂	1 mole NH ₂	

Example #2

Original sample = 2.175 g and yielded 3.94 g CO_2 and 1.89 g H_2O
Nitrogen Sample = 1.873g \rightarrow 0.436 g NH_2

Moles to Mass

$$\frac{0.0895 \text{ mole C}}{1 \text{ mole C}} \times 12.011 \text{ g C} = 1.074 \text{ g C}$$

$$\frac{0.2098 \text{ mole H}}{1 \text{ mole H}} \times 1.008 \text{ g H} = 0.2115 \text{ g H}$$

$$\frac{0.0272 \text{ mole N}}{1 \text{ mole N}} \times 14.007 \text{ g N} = 0.38114 \text{ g N}$$

Example #2

Original sample = 2.175 g and yielded 3.94 g CO₂ and 1.89 g H₂O
Nitrogen Sample = 1.873g → 0.436 g NH₂

Convert to % values because N is from another experiment!

$$\frac{1.0753 \text{ g C}}{2.175 \text{ g Sample}} \times 100 = 49.44\% \text{ C}$$

$$\frac{0.2115 \text{ g H}}{2.175 \text{ g Sample}} \times 100 = 9.72\% \text{ H}$$

$$\frac{0.38114 \text{ g N}}{1.873 \text{ g Sample}} \times 100 = 19.17\% \text{ N}$$

Example #2

Original sample = 2.175 g and yielded 3.94 g CO₂ and 1.89 g H₂O
Nitrogen Sample = 1.873g → 0.436 g NH₂

Subtract the % values from 100 to find how much Oxygen!

$$100 - 49.44\text{C} - 9.72\text{H} - 19.17\text{N} = 21.67\% \text{ Oxygen}$$

Back to the Rhyme!

% to mass, mass to moles, divide by small, multiply till whole!

$$49.44\% \text{ C} \rightarrow 49.44 \text{ g C}$$

$$9.72\% \text{ H} \rightarrow 9.72 \text{ g H}$$

$$19.17\% \text{ N} \rightarrow 19.17 \text{ g N}$$

$$21.67\% \text{ O} \rightarrow 21.67 \text{ g O}$$

Example #2

Original sample = 2.175 g and yielded 3.94 g CO₂ and 1.89 g H₂O
Nitrogen Sample = 1.873g → 0.436 g NH₂

Back to the Rhyme!

% to mass, **mass to moles**, divide by small, multiply till whole!

49.44 % C → 49.44 g C
9.72 % H → 9.72 g H
19.17 % N → 19.17 g N
21.67 % O → 21.67 g O

$$\begin{array}{r|l} 21.67 \text{ g O} & 1 \text{ mole O} \\ \hline & 16.0 \text{ g O} \end{array}$$

$$\begin{array}{r|l} 49.44 \text{ g C} & 1 \text{ mole C} \\ \hline & 12.01 \text{ g C} \end{array} = 4.116 \text{ mole C}$$

$$\begin{array}{r|l} 9.72 \text{ g H} & 1 \text{ mole H} \\ \hline & 1.008 \text{ g H} \end{array} = 9.643 \text{ mole H}$$

$$\begin{array}{r|l} 19.17 \text{ g N} & 1 \text{ mole N} \\ \hline & 14.01 \text{ g N} \end{array} = 1.3686 \text{ mole N}$$

Example #2

Original sample = 2.175 g and yielded 3.94 g CO₂ and 1.89 g H₂O
Nitrogen Sample = 1.873g → 0.436 g NH₂

Back to the Rhyme!

% to mass, mass to moles, **divide by small**, multiply till whole!

4.115 mole C
9.643 mole H
1.3686 mole N
1.3544 mole O

$$\frac{4.115 \text{ mole C}}{1.3544} = 3.04 \text{ C}$$

$$\frac{9.643 \text{ mole H}}{1.3544} = 7.12 \text{ H}$$

$$\frac{1.3686 \text{ mole N}}{1.3544} = 1.01 \text{ N}$$

$$\frac{1.3544 \text{ mole O}}{1.3544} = 1 \text{ O}$$

Example #2

Original sample = 2.175 g and yielded 3.94 g CO_2 and 1.89 g H_2O
Nitrogen Sample = 1.873g \rightarrow 0.436 g NH_2

Back to the Rhyme!

% to mass, mass to moles, divide by small, **multiply till whole!**

$$3.04 \text{ C} \rightarrow 3$$

$$7.12 \text{ H} \rightarrow 7$$

$$1.01 \text{ N} \rightarrow 1$$

$$1 \text{ O} \rightarrow 1$$



Example #2

Original sample = 2.175 g and yielded 3.94 g CO_2 and 1.89 g H_2O
Nitrogen Sample = 1.873g \rightarrow 0.436 g NH_2

Find Molecular Formula Now!

Molecular Mass/Empirical Mass = multiplier

Molecular Mass = 150 g/mol

Empirical Mass $\text{C}_3\text{H}_7\text{NO}$ = 73.11 g/mol

$$150/73.11 = 2.05 \rightarrow 2$$

$$\text{C}_3\text{H}_7\text{NO} \times 2 =$$

