Chemical Reactions

Chapter 4 Part 1

Reactants: $Zn + I_2$ ⟶ Product: ZnI_2

Chemistry 221

Professor Michael Russell

MAR

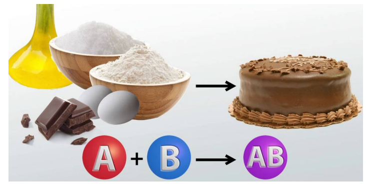
Last update: 4/10/23

Chemistry as Cooking! - the Chemical Reaction

"Recipe" and technique leads to successful creations

Must know amounts to add, how much will be produced

Haphazard additions can be disastrous!



Chemical Equations

Depict the kind of reactants and products and their relative amounts in a reaction.

$4\ Al(s) + 3\ O_2(g) \text{ ---> } 2\ Al_2O_3(s)$

The numbers in the front are called stoichiometric coefficients

The letters (s), (g), (aq) and (l) are the physical states of compounds.

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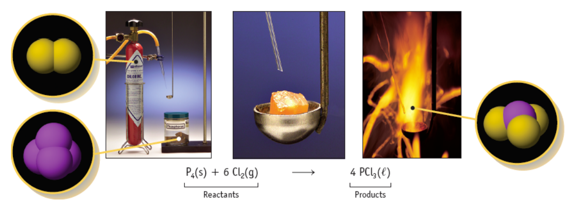
Reaction of Phosphorus with Cl_2



Notice the stoichiometric coefficients and the physical states of the reactants and products.

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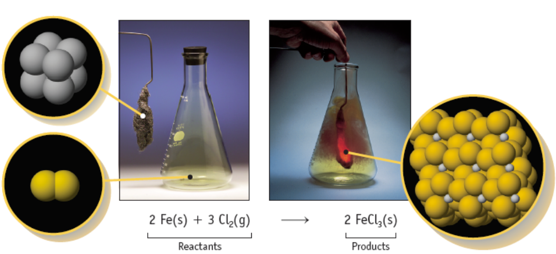
Reaction of Iron with Cl_2



Evidence of a chemical reaction:

heat change, precipitate formation, gas evolution, color change

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Chemical Equations

$4\ Al(s) + 3\ O_2(g) \rightarrow 2\ Al_2O_3(s)$

This equation means:

4 Al atoms + 3 O_2 molecules
---give--->
2 molecules of Al_2O_3

or

4 moles of Al + 3 moles of O_2
---give--->
2 moles of Al_2O_3

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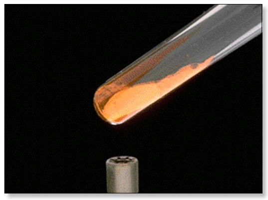
Chemical Equations

Because the same atoms are present in a reaction at the beginning and at the end, the amount of matter in a system does not change.

The Law of the Conservation of Matter

Also known as the Law of Mass Action



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Chemical Equations / Lavoisier

Because of the principle of the conservation of matter,

an equation must be balanced.

It must have the same number of atoms of the same kind on both sides.

Lavoisier, 1788

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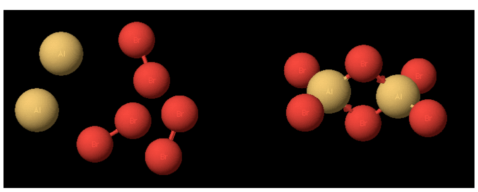
Balancing Equations

2 Al(s) + 3 Br_2(liq) ---> Al_2Br_6(s)



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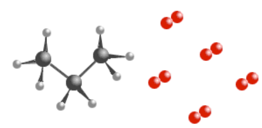

Balancing Equations

C_3H_8(g) + 5 O_2(g) ---->
3 CO_2(g) + 4 H_2O(g)

2 B_4H_{10}(g) + 11 O_2(g) ---->
4 B_2O_3(g) + 10 H_2O(g)

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Balancing Equations - *Hints*

Balance those atoms which occur in only one compound on each side last *(i.e. O_2 in previous examples)*

Balance the remaining atoms first

Reduce coefficients to smallest whole integers

Check your answer *if uncertain*

Helpful but optional: Check that charges are balanced

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STOICHIOMETRY

Stoichiometry is the study of the quantitative aspects of chemical reactions.

Stoichiometry rests on the principle of the conservation of matter.

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Stoichiometry

The balanced chemical equation

$4\ Al(s) + 3\ O_2(g) \text{ ---> } 2\ Al_2O_3(s)$

implies *all* of the following ratios:

$$\frac{4 \text{ mol Al}}{3 \text{ mol } O_2} \quad \frac{4 \text{ mol Al}}{2 \text{ mol } Al_2O_3} \quad \frac{3 \text{ mol } O_2}{2 \text{ mol } Al_2O_3}$$

$$\frac{3 \text{ mol } O_2}{4 \text{ mol Al}} \quad \frac{2 \text{ mol } Al_2O_3}{4 \text{ mol Al}} \quad \frac{2 \text{ mol } Al_2O_3}{3 \text{ mol } O_2}$$

These are nothing more than "conversion units" in dimensional analysis!

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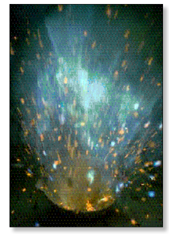
PROBLEM: If 454 g of NH_4NO_3 decomposes, how much N_2O and H_2O are formed? What is the theoretical yield of products?



STEP 1

Write the balanced chemical equation

NH_4NO_3 ---> N_2O + 2 H_2O

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454 g *of* NH_4NO_3 --> N_2O + 2 H_2O

STEP 2 Convert mass reactant (454 g) --> moles

$$454 \text{ g} \bullet \frac{1 \text{ mol}}{80.04 \text{ g}} = 5.68 \text{ mol } NH_4NO_3$$

80.04 g/mol = molar mass of NH_4NO_3

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454 g *of* NH_4NO_3 --> N_2O + 2 H_2O

STEP 3 Convert moles reactant --> moles product

Relate moles NH_4NO_3 to moles product expected.

1 mol NH_4NO_3 --> 2 mol H_2O

Express as a STOICHIOMETRIC FACTOR:

$$\frac{2 \text{ mol } H_2O \text{ produced}}{1 \text{ mol } NH_4NO_3 \text{ used}}$$

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454 g *of* NH_4NO_3 --> N_2O + 2 H_2O

STEP 3 Convert moles reactant (5.68 mol) --> moles product

$$5.68 \text{ mol } NH_4NO_3 \bullet \frac{2 \text{ mol } H_2O \text{ produced}}{1 \text{ mol } NH_4NO_3 \text{ used}}$$

= 11.4 mol H_2O produced

How many moles of N_2O produced?

Answer = **5.68 mol N_2O**

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454 g *of* NH_4NO_3 --> N_2O + 2 H_2O

STEP 4 Convert moles product (11.4 mol) --> mass product

This is called the THEORETICAL YIELD

$$11.4 \text{ mol } H_2O \bullet \frac{18.02 \text{ g}}{1 \text{ mol}} = 204 \text{ g } H_2O$$

ALWAYS FOLLOW THESE STEPS IN SOLVING STOICHIOMETRY PROBLEMS!

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454 g *of* NH_4NO_3 --> N_2O + 2 H_2O

STEP 5 How much N_2O is formed?

Total mass of reactants =

total mass of products

454 g NH_4NO_3 = ___ g N_2O + 204 g H_2O

mass of N_2O = 250. g ***law of mass action!***

could also turn mol NH_4NO_3 into mol N_2O, then grams of N_2O:

5.68 mol N_2O * 44.01 g/mol = 250. g

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454 g *of* NH_4NO_3 --> N_2O + 2 H_2O

Compound	NH_4NO_3	N_2O	H_2O
Initial (g)	454 g	0	0
Initial (mol)	5.68mol	0	0
Change (mol)	-5.68	+5.68	+2(5.68)
Final (mol)	0	5.68	11.4
Final (g)	0	250.	204

Mass is conserved!

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454 g *of* NH_4NO_3 --> N_2O + 2 H_2O

STEP 6 Calculate the percent yield

We predicted a yield of 250. g of N_2O. If you isolated only 131 g of N_2O, what is the percent yield of N_2O?

This compares the **theoretical yield** *(250. g) and* **actual yield** *(131 g) of N_2O.*

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454 g *of* NH_4NO_3 --> N_2O + 2 H_2O

STEP 6 Calculate the percent yield

$$\%\text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \bullet 100\%$$

$$\%\text{ yield} = \frac{131\text{ g}}{250.\text{ g}} \bullet 100\% = 52.4\%$$

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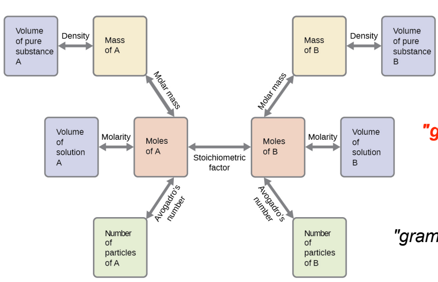
GENERAL PLAN FOR STOICHIOMETRY CALCULATIONS



We use the ***"grams - moles - moles - grams"*** *conversion often in chemistry (mass is grams)*

Sometimes I call this the "grams - moles - moles - grams" dance lol

Molarity in next chapter - See Stoichiometry Guide

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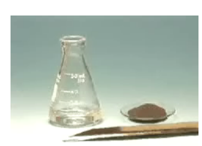
PROBLEM: Using 5.00 g of H_2O_2, what mass of O_2 and of H_2O can be obtained?

2 H_2O_2(liq) ---> 2 H_2O(g) + O_2(g)

Reaction is catalyzed by MnO_2



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PROBLEM: Using 5.00 g of H_2O_2, what mass of O_2 and of H_2O can be obtained?

$2\ H_2O_2$(liq) ---> $2\ H_2O$(g) + O_2(g)

Reaction is catalyzed by MnO_2

Step 1: moles of H_2O_2

Step 2: use STOICHIOMETRIC FACTOR to calculate moles of O_2

Step 3: mass of O_2 (2.35 g)

Step 4: mass of H_2O (2.65 g)

Try this problem yourself!

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Reactions Involving a LIMITING REACTANT

In a given reaction, there is not enough of one reagent to use up the other reagent completely.

The reagent in short supply LIMITS the quantity of product that can be formed.

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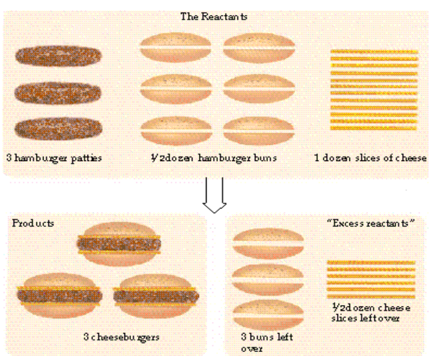
LIMITING REACTANTS



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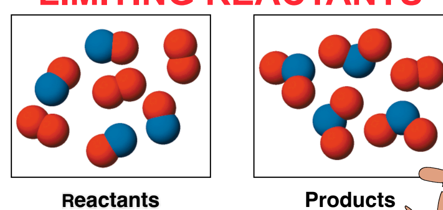
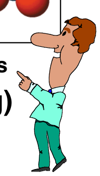
LIMITING REACTANTS





$2\ NO(g) + O_2\ (g) \longrightarrow 2\ NO_2(g)$

Limiting reactant = ___________

Excess reactant = ____________

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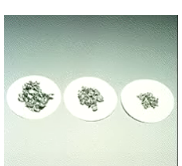
LIMITING REACTANTS



React solid Zn with 0.100 mol HCl (aq)

$Zn_{(s)} + 2\ HCl_{(aq)}$ ---> $ZnCl_{2(aq)} + H_{2(g)}$

Left: Balloon inflates fully, some Zn left

* More than enough Zn to use up the 0.100 mol HCl

Center: Balloon inflates fully, no Zn left

* Right amount of each (HCl and Zn)

Right: Balloon does not inflate fully, no Zn left.

* Not enough Zn to use up 0.100 mol HCl

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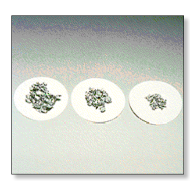
LIMITING REACTANTS



React solid Zn with 0.100 mol HCl (aq)

$Zn_{(s)} + 2\ HCl_{(aq)}$ ---> $ZnCl_{2(aq)} + H_{2(g)}$

0.100 mol HCl [1 mol Zn/2 mol HCl] = 0.0500 mol Zn

	Left	Center	Right
mass Zn (g)	7.00	3.27	1.31
mol Zn	0.107	0.050	0.020
mol HCl	0.100	0.100	0.100
mol HCl/mol Zn	0.93	2.00	5.00
Lim Reactant	LR = HCl	no LR	LR = Zn

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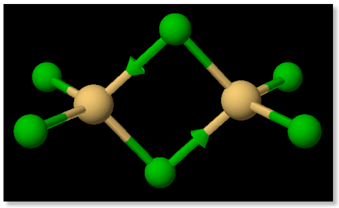
Reaction to be Studied:

$2\ Al\ +\ 3\ Cl_2\ \text{--->}\ Al_2Cl_6$



PROBLEM: Mix 5.40 g of Al with 8.10 g of Cl_2. How many grams of Al_2Cl_6 can form?

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Step 1 of the Limiting Reactant problem: **Compare actual mole ratio of reactants to theoretical mole ratio.**

$2\ Al\ +\ 3\ Cl_2\ \text{--->}\ Al_2Cl_6$

Reactants must be in the mole ratio

$$\frac{\text{mol } Cl_2}{\text{mol Al}} = \frac{3}{2}$$

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Deciding on the Limiting Reactant

$2\ Al\ +\ 3\ Cl_2\ \text{--->}\ Al_2Cl_6$

If $\frac{\text{mol } Cl_2}{\text{mol Al}} > \frac{3}{2}$

then there is not enough Al to use up all the Cl_2, and the limiting reagent is Al

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Deciding on the Limiting Reactant

$2\ Al\ +\ 3\ Cl_2\ \text{--->}\ Al_2Cl_6$

If $\frac{\text{mol } Cl_2}{\text{mol Al}} < \frac{3}{2}$

then there is not enough Cl_2 to use up all the Al, and the limiting reagent is Cl_2

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Step 2 of the Limiting Reactant problem: **Calculate moles of each reactant**

We have 5.40 g of Al and 8.10 g of Cl_2. How much Al_2Cl_6 can form?

$$5.40 \text{ g Al} \bullet \frac{1 \text{ mol}}{27.0 \text{ g}} = 0.200 \text{ mol Al}$$

$$8.10 \text{ g } Cl_2 \bullet \frac{1 \text{ mol}}{70.9 \text{ g}} = 0.114 \text{ mol } Cl_2$$

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$2\ Al\ +\ 3\ Cl_2\ \text{--->}\ Al_2Cl_6$

Step 3 of the Limiting Reactant problem: **Compare moles to find limiting reactant**

$$\frac{\text{mol } Cl_2}{\text{mol Al}} = \frac{0.114 \text{ mol}}{0.200 \text{ mol}} = 0.570$$

This should be 3/2 or 1.5/1 if reactants are present in the exact stoichiometric ratio.

Limiting reagent is Cl_2

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$2\ Al\ +\ 3\ Cl_2\ \text{--->}\ Al_2Cl_6$

Mix 5.40 g of Al with 8.10 g of Cl_2. What mass of Al_2Cl_6 can form?

Limiting reactant = Cl_2
Base all calculations on Cl_2

grams Cl_2 → moles Cl_2 → (1 mol Al_2Cl_6 / 3 mol Cl_2) → moles Al_2Cl_6 → grams Al_2Cl_6

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$2 Al + 3 Cl_2 \text{ ---> } Al_2Cl_6$

CALCULATIONS: calculate mass of Al_2Cl_6 expected using limiting reactant.

Step 1: Calculate moles of Al_2Cl_6 expected using chlorine:

$$0.114 \text{ mol } Cl_2 \bullet \frac{1 \text{ mol } Al_2Cl_6}{3 \text{ mol } Cl_2} = 0.0380 \text{ mol } Al_2Cl_6$$

Step 2: Calculate mass of Al_2Cl_6 expected based on chlorine:

$$0.0380 \text{ mol } Al_2Cl_6 \bullet \frac{266.4 \text{ g } Al_2Cl_6}{\text{mol}} = 10.1 \text{ g } Al_2Cl_6$$

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$2 Al + 3 Cl_2 \text{ ---> } Al_2Cl_6$

Alternate Limiting Reactant Method

Calculate theoretical yield of product based on *both* reactants.

Smaller theoretical yield comes from limiting reactant, greater yield from excess reactant.

$$8.10 \text{ g } Cl_2 \bullet \frac{1 \text{ mol}}{70.9 \text{ g}} \bullet \frac{1 \text{ mol } Al_2Cl_6}{3 \text{ mol } Cl_2} \bullet \frac{266.4 \text{ g}}{1 \text{ mol}} = 10.1 \text{ g } Al_2Cl_6$$

$$5.40 \text{ g Al} \bullet \frac{1 \text{ mol}}{27.0 \text{ g}} \bullet \frac{1 \text{ mol } Al_2Cl_6}{2 \text{ mol Al}} \bullet \frac{266.4 \text{ g}}{1 \text{ mol}} = 26.6 \text{ g } Al_2Cl_6$$

10.1 g < 26.6 g, *so:* limiting reactant = Cl_2, theoretical yield = 10.1 g, excess reactant = Al

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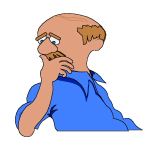
$2 Al + 3 Cl_2 \text{ ---> } Al_2Cl_6$

How much of which reactant will remain when reaction is complete?



- Cl_2 was the limiting reactant. Therefore, Al was present in excess. But by how much?
- *First* find how much Al was required based on limiting reactant (Cl_2).
- *Then* find how much Al is in excess.

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$2 Al + 3 Cl_2 \text{ ---> } Al_2Cl_6$

Calculating Excess Al

$2 Al + 3 Cl_2 \longrightarrow$ products

0.200 mol (Al) 0.114 mol = LR (Cl_2)

$$8.10 \text{ g } Cl_2 \bullet \frac{1 \text{ mol}}{70.9 \text{ g}} \bullet \frac{2 \text{ mol Al}}{3 \text{ mol } Cl_2} \bullet \frac{26.98 \text{ g}}{1 \text{ mol}} = 2.05 \text{ g Al}$$

Excess Al = Al available - Al required
= 5.40 g - 2.05 g
= 3.35 g Al unused in reaction

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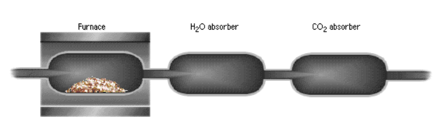
$2 Al + 3 Cl_2 \text{ ---> } Al_2Cl_6$

Using Stoichiometry to Determine a Formula

- Hydrocarbons, C_xH_y, can be burned in oxygen to give CO_2 and H_2O (combustion reaction).
- The CO_2 and H_2O can be collected to determine the empirical formula of the hydrocarbon.

$C_xH_y + O_2 \text{ ---> } CO_2 + H_2O$



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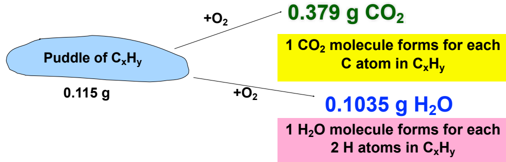
Using Stoichiometry to Determine a Formula

What is the empirical formula of a hydrocarbon, C_xH_y, if burning 0.115 g produces 0.379 g CO_2 and 0.1035 g H_2O?

C_xH_y + some O_2 ---> 0.379 g CO_2 + 0.1035 g H_2O



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Using Stoichiometry to Determine a Formula

C_xH_y + some oxygen --->
0.379 g CO_2 + 0.1035 g H_2O

***First*, recognize that all C in CO_2 and all H in H_2O comes from C_xH_y.**

1. Calculate amount of C in CO_2

8.61×10^{-3} mol CO_2 --> 8.61×10^{-3} mol C

1 mol C per 1 mol CO_2

2. Calculate amount of H in H_2O

5.744×10^{-3} mol H_2O -- >1.149×10^{-2} mol H

2 mol H per 1 mol water!

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Using Stoichiometry to Determine a Formula

C_xH_y + some oxygen --->
0.379 g CO_2 + 0.1035 g H_2O

Now find ratio of mol H/mol C to find values of x and y in C_xH_y.

1.149×10^{-2} mol H/ 8.61×10^{-3} mol C

= 1.33 mol H / 1.00 mol C

= 4 mol H / 3 mol C

Empirical formula = C_3H_4

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Formulas with C, H and O

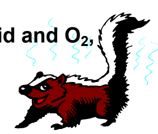
Caproic acid, the substance responsible for "dirty gym socks" smell, contains C, H *and* O.

Combustion analysis of 0.450 g caproic acid gives 0.418 g H_2O and 1.023 g CO_2, and the molar mass was found to be 116.2 g mol^{-1}.

What is the molecular formula of caproic acid?

$C_xH_yO_z$ + some oxygen ---> 1.023 g CO_2 + 0.418 g H_2O

***Careful:* oxygen comes from caproic acid and O_2, need special technique**



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Formulas with C, H and O

Combustion analysis of 0.450 g caproic acid gives 0.418 g H_2O and 1.023 g CO_2, and the molar mass is 116.2 g mol^{-1}. What is the molecular formula?

Start with "regular" approach for mol H & mol C:

0.418 g H_2O * (mol/18.02 g) * (2 mol H/mol H_2O) = 0.0464 mol H

0.0464 mol H * (1.01 g/mol H) = 0.0469 g H

1.023 g CO_2 * (mol/44.01 g) * (1 mol C/mol CO_2) = 0.02324 mol C

0.02324 mol C * (12.01 g/mol C) = 0.2791 g C

***Why did we convert to grams?* Law of Mass Action!**

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Formulas with C, H and O

0.450 g caproic acid: 0.418 g H_2O (0.0464 mol H, 0.0469 g H) and 1.023 g CO_2 (0.02324 mol C, 0.2791 g C), molar mass = 116.2 g/mol. What is the molecular formula?

Realize that 0.450 g of caproic acid equals all the g C, g H *and g O* in the complex.

Converting mol H and mol C to grams, then subtracting from 0.450 g, gives g O in caproic acid:

0.450 g - 0.0469 g - 0.2791 g = 0.124 g O

caproic acid   g of H in acid   g of C in acid   g of O in acid

0.124 g O * (mol O / 16.00 g) = 0.00775 mol O

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Formulas with C, H and O

0.450 g caproic acid: 0.418 g H_2O (0.0464 mol H) and 1.023 g CO_2 (0.02324 mol C), molar mass = 116.2 g/mol, 0.00775 mol O. What is the molecular formula?

Now compare moles:

$C_{0.02324}H_{0.0464}O_{0.00775}$ *gives* C_3H_6O = *empirical formula*

C_3H_6O has a molar mass of 58.1 g/mol, which is half of the 116.2 g/mol value

Molecular Formula = $(C_3H_6O)_2$, or

$C_6H_{12}O_2$

You can now find empirical formulas based on combustion analysis (this chapter) and elemental percentages (previous chapter)!

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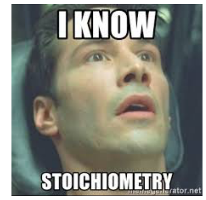
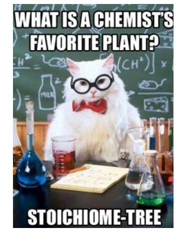
End of Chapter 4 Part 1

See also:

- *Chapter Four Part 1 Study Guide*
- *Chapter Four Part 1 Concept Guide*
- *Important Equations (following this slide)*
- *End of Chapter Problems (following this slide)*





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Important Equations, Constants, and Handouts from this Chapter:

- **be able to find the theoretical yield, actual yield, percent yield**
- **be able to determine the limiting reactant, excess reactant, excess reactant remaining at end of reaction**
- **understand how to calculate empirical formula (EF) and molecular formula (MF) using organic compounds containing oxygen**

Balancing Equations:
Reactants, Products, states of matter (s, l, g, aq), stoichiometric coefficients, Law of Conservation of Matter ("mass action")

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End of Chapter Problems: *Test Yourself*

See practice problem set #4 and self quizzes for ***balancing chemical equations*** *examples and practice*

1. What mass of Br_2, in grams, is required for complete reaction with 2.56 g of Al? What mass of white, solid Al_2Br_6 is expected? The equation: 2 Al(s) + 3 Br_2(l) → Al_2Br_6(s)
2. Aluminum chloride is made by treating aluminum with chlorine: **2 Al(s) + 3 Cl_2(g) → 2 $AlCl_3$(s)** If you begin with 2.70 g of Al and 4.05 g of Cl_2, which reactant is limiting? What mass of $AlCl_3$ can be produced? What mass of the excess reactant remains when the reaction is completed?
3. $Cu(NH_3)_4SO_4$ is made via: **$CuSO_4$(aq) + 4 NH_3(aq) → $Cu(NH_3)_4SO_4$(aq)** If you use 10.0 g of $CuSO_4$ and excess NH_3, what is the theoretical yield of $Cu(NH_3)_4SO_4$? If you isolate 12.6 g of $Cu(NH_3)_4SO_4$, what is the percent yield of $Cu(NH_3)_4SO_4$?
4. An unknown compound has the formula $C_xH_yO_z$. You burn 0.0956 g of the compound and isolate 0.1356 g of CO_2 and 0.0833 g of H_2O. What is the empirical formula of the compound? If the molar mass is 62.1 g/mol, what is the molecular formula?

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End of Chapter Problems: *Answers*

1. 22.7 g Br_2 , 25.3 g Al_2Br_6
2. Chlorine is limiting; 5.09 g $AlCl_3$; 1.67 g Al remains
3. 14.3 g $Cu(NH_3)_4SO_4$, 88.3%
4. EF = CH_3O, MF = $C_2H_6O_2$

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