## Chemical Reactions

*Chapter 4 Part 1*

**Reactants: Zn + $I_2$ ⟶ Product: $ZnI_2$**

**Chemistry 221**
**Professor Michael Russell**

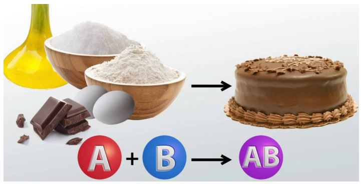
*MAR*

## *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!**



*Last update: 4/29/24*

## Chemical Equations

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

**$4 Al(s) + 3 O_2(g) ---> 2 Al_2O_3(s)$**

**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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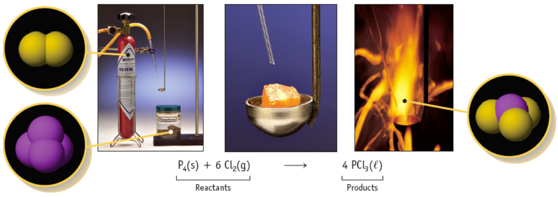
## Reaction of Phosphorus with $Cl_2$



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

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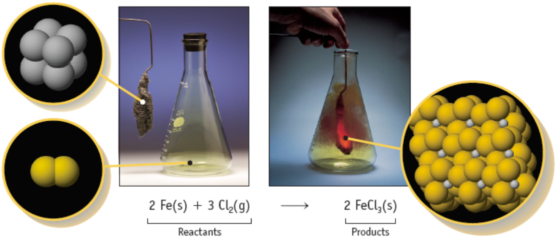
## Reaction of Iron with $Cl_2$



***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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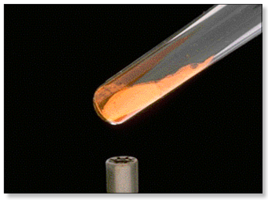
## Chemical Equations

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



## Chemical Equations / Lavoisier

Because of the principle of the conservation of matter,

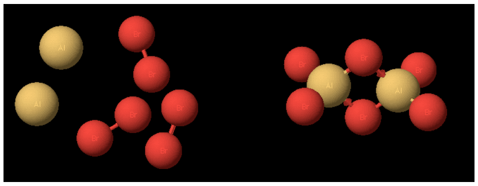
an equation must be balanced.

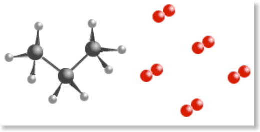
It must have the same number of atoms of the same kind on both sides.

Lavoisier, 1788

## Balancing Equations

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





## 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)

## 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

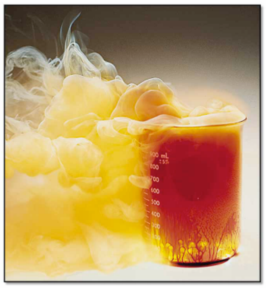
Check your answer *if uncertain*

***Helpful but optional:*** Check that charges are balanced

## STOICHIOMETRY

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

Stoichiometry rests on the principle of the conservation of matter.



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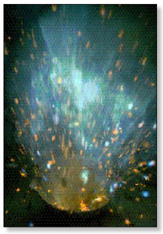
**PROBLEM: If 454 g of $NH_4NO_3$ decomposes, how much $N_2O$ and $H_2O$ are formed? What is the theoretical yield of products?**



**STEP 1**

**Write the balanced chemical equation**

**$NH_4NO_3$ --->**
**$N_2O\ +\ 2\ H_2O$**

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**454 g *of* $NH_4NO_3 \text{ --> } N_2O + 2\ H_2O$**

**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* $NH_4NO_3 \text{ --> } N_2O + 2\ H_2O$**

**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:**

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

*MAR*

**454 g *of* $NH_4NO_3 \text{ --> } N_2O + 2\ H_2O$**

**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* $NH_4NO_3 \text{ --> } N_2O + 2\ H_2O$**

**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!**

*MAR*

## 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**

## 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!***

## 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 $N_2O$.*

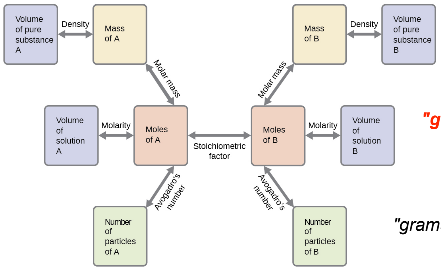
## 454 g *of* $NH_4NO_3$ --> $N_2O$ + 2 $H_2O$

**STEP 6 Calculate the percent yield**

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

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

## GENERAL PLAN FOR STOICHIOMETRY CALCULATIONS



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

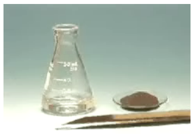
*Sometimes I call this the "grams - moles - moles - grams" dance lol*

***Molarity in next chapter - See Stoichiometry Guide***

## 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 $MnO_2$**



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

$2\,H_2O_2(liq) \text{ ---> } 2\,H_2O(g) + O_2(g)$

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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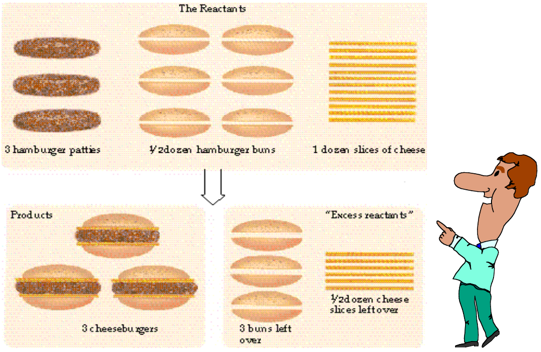
## LIMITING REACTANTS



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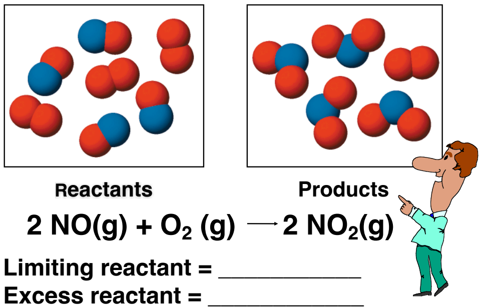
## LIMITING REACTANTS



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

Limiting reactant = ___________

Excess reactant = ___________

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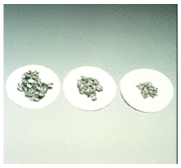
## LIMITING REACTANTS



React solid Zn with 0.100 mol HCl (aq)

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

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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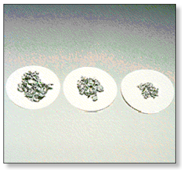
## LIMITING REACTANTS



React solid Zn with 0.100 mol HCl (aq)

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

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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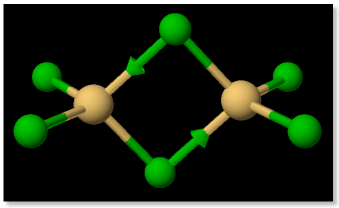
## *Reaction to be Studied:*

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



**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**

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

*MAR*

## 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**

*MAR*

## 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 $Cl_2$**

*MAR*

*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 $Cl_2$**

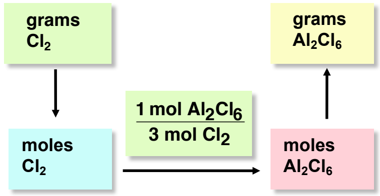
*MAR*

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

## 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 $Cl_2$***



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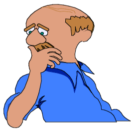
***$2\,Al + 3\,Cl_2 \text{ ---> } Al_2Cl_6$***

## 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 $Cl_2$ ⟶ products**

**0.200 mol**

**0.114 mol = LR**

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

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

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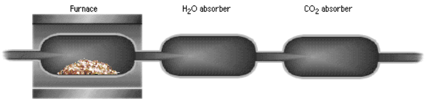
***$2\,Al + 3\,Cl_2 \text{ ---> } Al_2Cl_6$***

## 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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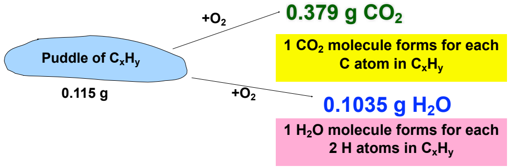
## Using Stoichiometry to Determine a Formula

**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 $H_2O$**



+$O_2$

**0.379 g $CO_2$**

Puddle of $C_xH_y$

0.115 g

**1 $CO_2$ molecule forms for each C atom in $C_xH_y$**

+$O_2$

**0.1035 g $H_2O$**

**1 $H_2O$ molecule forms for each 2 H atoms in $C_xH_y$**

*MAR*

## 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 x $10^{-3}$ mol $CO_2$ --> 8.61 x $10^{-3}$ mol C**

***1 mol C per 1 mol $CO_2$***

**2. Calculate amount of H in $H_2O$**

**5.744 x $10^{-3}$ mol $H_2O$ -- >1.149 x $10^{-2}$ mol H**

***2 mol H per 1 mol water!***

*MAR*

## Using Stoichiometry to Determine a Formula

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

**Now find <u>ratio</u> of mol H/mol C to find values of x and y in $C_xH_y$.**

**1.149 x 10 $^{-2}$ mol H/ 8.61 x $10^{-3}$ mol C**

**= 1.33 mol H / 1.00 mol C**

**= 4 mol H / 3 mol C**

**Empirical formula = $C_3H_4$**

*MAR*

## 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**

*MAR*

## 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!**

*MAR*

## 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:**

$$\underbrace{0.450\text{ g}}_{\text{caproic acid}} - \underbrace{0.0469\text{ g}}_{\text{g of H in acid}} - \underbrace{0.2791\text{ g}}_{\text{g of C in acid}} = \underbrace{0.124\text{ g O}}_{\text{g of O in acid}}$$

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

*MAR*

## 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**

**$C_6H_{12}O_2$**

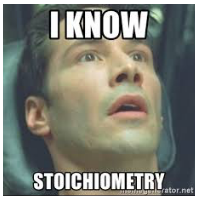
***You can now find empirical formulas based on combustion analysis (this chapter) and elemental percentages (previous chapter)!***

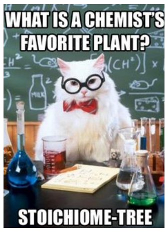
MAR

## 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)*





MAR

### *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")**

MAR

### 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 = $C_2H_6O_2$

*Be sure to view practice problem set #4 and self quizzes for* ***balancing chemical equations*** *examples and practice*

MAR