# CHEM 103
## Limiting Reagents

Lecture Notes
February 21, 2006
Prof. Sevian

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## Chem 103

Please sit with your groups today.

We will be doing a group problem at the end of class.

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## Agenda

- Limiting reagent problems
- Combustion analysis

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## Announcements

- First exam is this Thursday, Feb 23
  - Most important rules to know…
- One-third of the class will be in an overflow room.
  - Last names beginning A-H in room …
  - Last names beginning I-Z here in small science auditorium
- Posted on the course website:
  - Practice exam #1
  - Answer key to practice exam #1
  - Study guide
- ACS exam study guide is/is not at the bookstore?

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## Quantitative Analysis in the Lab

- Determine the composition of a mixture
  - React one substance in the mixture according to a known reaction.
  - Measure results of reaction.
  - Use stoichiometry to determine how much of original substance was present.
- Determine the empirical formula of an unknown chemical
  - React the substance and measure amounts of different chemical products.
  - Use stoichiometry to determine moles of elements present in original sample.
  - Find empirical formula from ratio of moles.

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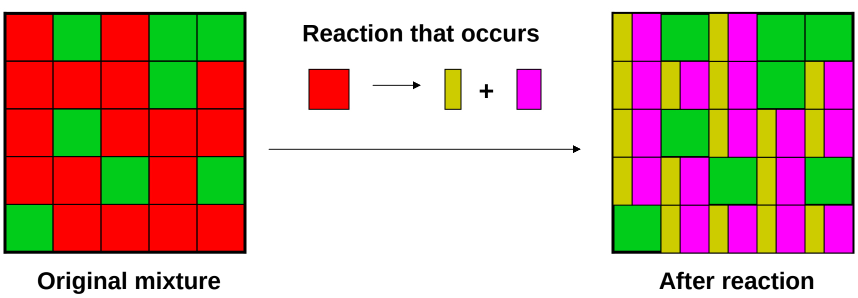
## Composition of a Mixture




- Products of the reaction must be known
- If you can measure the mass of one of the products, then you can use stoichiometry to determine how much of the original reactant was present
- Examples: product is a gas or precipitate

## Example: Gravimetric Analysis

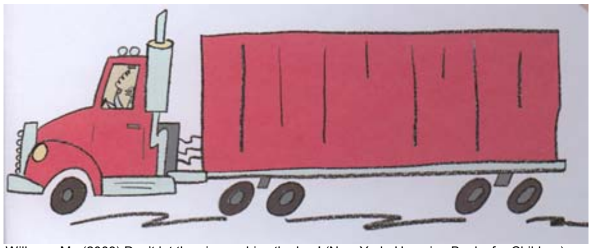
A mixture of unknown percentage composition of white $BaCl_2 \bullet 2H_2O$ crystals and white $Na_2SO_4 \bullet 10H_2O$ crystals is provided. How could you determine the percent composition of the mixture?

1. How could you get a reaction to occur?
2. What reaction would occur?
3. What would you need to measure?
4. How would you measure it?
5. What information could you calculate by knowing that?
6. What additional information might you need?



## Limiting Reactants

- If you use an excess of one reactant, it is more likely that all of the other reactant will be used up in the chemical reaction.
- The reactant that is in excess will be present in the final products, since it didn’t all get used up.



Willems, Mo (2003) Don't let the pigeon drive the bus! (New York: Hyperion Books for Children).

## An analogy to illustrate limiting reactants

To build a semi tractor-trailer requires:

18 wheels + 5 axles + 4 cylinders + 1 pigeon → 1 semi

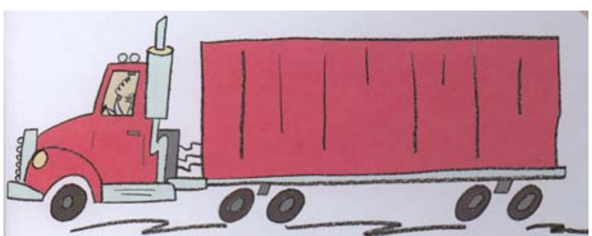
Let's say you have in stock:

- 198 wheels × (1 semi)/(18 wheels) = 11 semis possible
- 100 axles × (1 semi)/(5 axles) = 20 semis possible
- 100 cylinders × (1 semi)/(4 cylinders) = 25 semis possible
- 18 pigeons × (1 semi)/(1 pigeon) = 18 semis possible

Maximum quantity of semis you could produce is 11

How many semis could you produce?

LIMITING REACTANT



Willems, Mo (2003) Don't let the pigeon drive the bus! (New York: Hyperion Books for Children).

## An analogy to illustrate limiting reactants

18 wheels + 5 axles + 4 cylinders + 1 pigeon → 1 semi

After you build 11 semis, what remains of the 198 wheels, 100 axles, 100 cylinders, and 18 pigeons?

| | 18 wheels | 5 axles | 4 cylinders | 1 pigeon | → | 1 semi |
|---|---|---|---|---|---|---|
| Initial | 198 | 100 | 100 | 18 | | 0 |
| Change | -198 | -(11×5) | -(11×4) | -(11×1) | | +11 |
| Final | 0 | 45 | 56 | 7 | | 11 |

## What was actually in the barium chloride + sodium sulfate mixture

$BaCl_2\ (aq)\ +\ Na_2SO_4\ (aq) \rightarrow 2\ NaCl\ (aq)\ +\ BaSO_4\ (s)$

$$5.00g\ BaCl_2 \bullet 2H_2O \times \frac{1\ mol\ BaCl_2 \bullet 2H_2O}{244.2g\ BaCl_2 \bullet 2H_2O} \times \frac{1\ mol\ BaCl_2}{1\ mol\ BaCl_2 \bullet 2H_2O} = 0.0205\ mol\ BaCl_2$$

$$5.00g\ Na_2SO_4 \bullet 10H_2O \times \frac{1\ mol\ Na_2SO_4 \bullet 10H_2O}{322.3g\ Na_2SO_4 \bullet 10H_2O} \times \frac{1\ mol\ Na_2SO_4}{1\ mol\ Na_2SO_4 \bullet 10H_2O} = 0.0155\ mol\ Na_2SO_4$$

| moles | $BaCl_2$ | $Na_2SO_4$ | $\rightarrow$ | 2 NaCl | $BaSO_4$ |
|---|---|---|---|---|---|
| Initial | 0.0205 | 0.0155 | | 0 | 0 |
| Change | | | | | |
| Final | | | | | |

## Excess Reactant Logic

Amounts of reactants

Molar masses

How many moles present of each reactant?

Stoichiometric coefficients

How many moles of product could be produced?

Which reactant produces least product? Limiting reactant

All other reactants are present in excess

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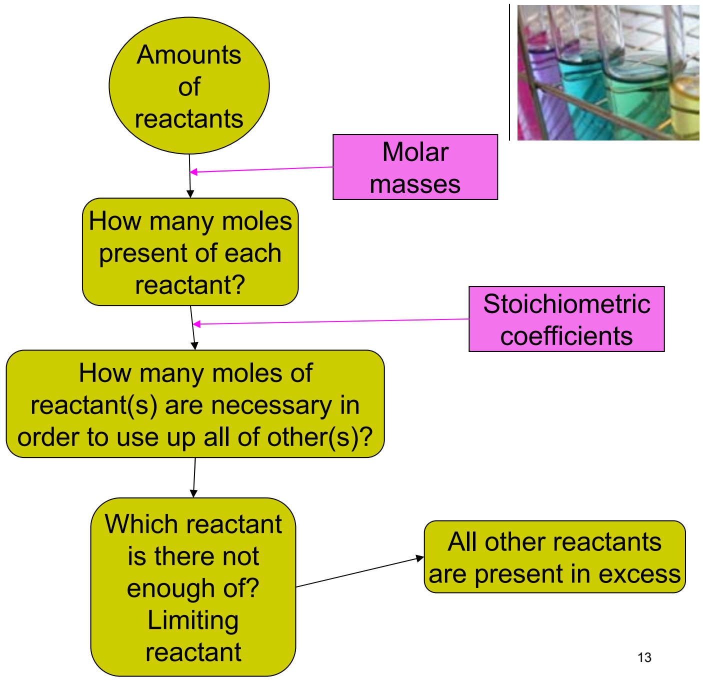
## Excess Reactant Logic
(alternative)

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## Chemical Analysis

1. Unknown quantity of a known substance
   - Two known chemicals react to exactly the point of completion
   - Quantity of one of them is known
   - Determine quantity that must have been present of the other one
   - Example: titration
2. Known quantity of an unknown substance
   - A specific amount of an unknown substance reacts
   - Known products are formed, and their quantities are measured
   - Determine empirical formula of the unknown reactant from this information
   - Examples: gravimetric analysis, combustion analysis

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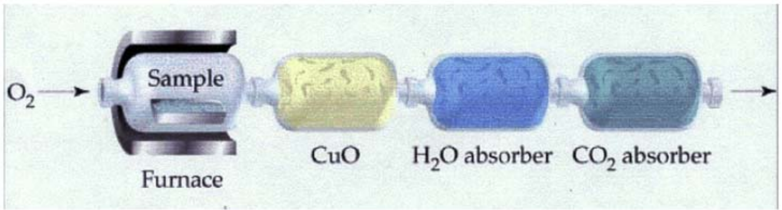
## Combustion Analysis



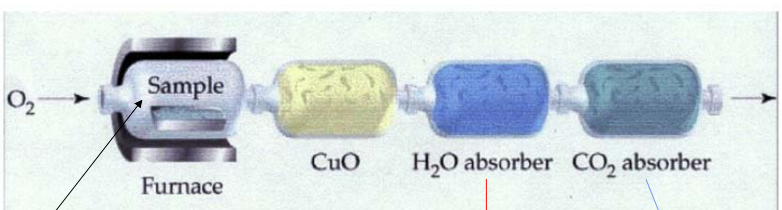

Unbalanced reaction is

Sample (containing C, H, and O) + $O_2 \rightarrow H_2O + CO_2$

- Excess $O_2$ used, so that all of the sample reacts
- CuO ensures that any carbon monoxide converts to carbon dioxide
- Water absorber is usually $Mg(ClO_4)_2$
- Carbon dioxide is usually collected by sodium hydroxide on asbestos

Picture from http://itl.chem.ufl.edu/2045_s00/lectures/lec_4.html

## Combustion Analysis




**Known mass of sample reacts**

**Mass of $H_2O$ is measured**

**Mass of $CO_2$ is measured**

**Mass of $H_2O$ $\rightarrow$ Moles of $H_2O$ $\rightarrow$ Moles of H atoms present in original sample**

**Mass of $CO_2$ $\rightarrow$ Moles of $CO_2$ $\rightarrow$ Moles of C atoms present in original sample**

**(Mass of original sample) – (Mass of H atoms + Mass of C atoms)**

**= Mass of O atoms, if present in original sample**

**$\rightarrow$ Moles of O atoms present in original sample**

## Combustion Analysis Example

The combustion of 0.5320 g butane (in a lighter) produces 1.6114 g of carbon dioxide and 0.8247 g of water.

(a) What is the empirical formula for butane?

Chemical reaction is

butane + $O_2 \rightarrow CO_2 + H_2O$

Strategy:

- Empirical formula means ratio of moles
- Need to know moles of C, moles of H (and possibly moles of O)
- Get moles of C from mass of $CO_2$
- Get moles of H from mass of $H_2O$
- Get moles of O (if any) by determining if C and H account for entire mass of butane
- Determine empirical formula of butane from ratio of moles C:H(:O)

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## Combustion Analysis Example

The combustion of 0.5320 g butane (in a lighter) produces 1.6114 g of carbon dioxide and 0.8247 g of water.

(a) What is the empirical formula for butane?

| Start | End |
|---|---|
| **1.6114 g $CO_2$** | **? mol C** |

| Start | End |
|---|---|
| **0.8247 g $H_2O$** | **? mol H** |

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## Combustion Analysis Example

The combustion of 0.5320 g butane (in a lighter) produces 0.4028 g of carbon dioxide and 0.8247 g of water.

(a) What is the empirical formula for butane?

Is there any O in butane?

$$\frac{0.036614 \text{ mol C} \mid 12.011 \text{ g C}}{\mid 1 \text{ mol C}} = 0.43973 \text{ g C}$$

$$\frac{0.09153 \text{ mol H} \mid 1.008 \text{ g H}}{\mid 1 \text{ mol H}} = 0.09226 \text{ g H}$$

**Sum is**

**0.5320 g**

**So what?**

## Combustion Analysis Example

The combustion of 0.5320 g butane (in a lighter) produces 0.4028 g of carbon dioxide and 0.8247 g of water.

(a) What is the empirical formula for butane?

Ratio of Moles C : Moles H

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## Combustion Analysis Example

**The combustion of 0.5320 g butane (in a lighter) produces 1.6114 g of carbon dioxide and 0.8247 g of water.**

**(b) A different analytical procedure indicated that 9.153 $\times$ $10^{-3}$ moles of butane were present in the original sample of butane. What is the molecular formula for butane?**

**Strategy:**

- **Start with empirical formula from part (a)**
- **Use mass and moles to calculate molar mass of butane**
- **Find out how many empirical units are in butane**
- **Molecular formula of butane is same factor $\times$ empirical formula**

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## Combustion Analysis Example

**The combustion of 0.5320 g butane (in a lighter) produces 1.6114 g of carbon dioxide and 0.8247 g of water.**

**(b) A different analytical procedure indicated that 9.153 $\times$ $10^{-3}$ moles of butane were present in the original sample of butane. What is the molecular formula for butane?**

Molar mass is measured in g/mol

Molar mass of butane =

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## Combustion Analysis Example

**The combustion of 0.5320 g butane (in a lighter) produces 1.6114 g of carbon dioxide and 0.8247 g of water.**

**(b) A different analytical procedure indicated that 9.153 × 10$^{-3}$ moles of butane were present in the original sample of butane. What is the molecular formula for butane?**

| | Empirical | Molecular |
|---|---|---|
| Formula | $C_2H_5$ | |
| Molar mass | 29.06 g/mol | 58.12 g/mol |

Where did this come from?

Where did this come from?

## Percent Yield

- "Yield" = what is produced
- "Percent" = (part/whole) × 100%
- Percent yield = comparison between

what was actually produced

and

the stoichiometric amount that could have been produced if the world was perfect

$$Percent\ yield = \frac{Actual\ amount\ produced}{Amount\ that\ could\ have\ been\ produced} \times 100\%$$

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## Percent Yield Example

Reaction of 5.00 g of nitrogen gas with hydrogen produced 5.50 g of ammonia. What was the percent yield of the reaction?

$$N_2 + 3\,H_2 \rightarrow 2\,NH_3$$

Mass of ammonia possible in a perfect world (theoretical yield):

$$5.00\ g\ N_2 \times \frac{1\ mol\ N_2}{28.02\ g\ N_2} \times \frac{2\ mol\ NH_3}{1\ mol\ N_2} \times \frac{17.03\ g\ NH_3}{1\ mol\ NH_3} = 6.08\ g\ NH_3$$

Mass of ammonia actually produced (actual yield): 5.50 g

$$\%\ yield = \frac{actual}{theoretical} \times 100\%$$

$$= \frac{5.50g}{6.08g} \times 100\% = 90.5\%$$

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## Exam 1: Most Common Difficulties Last Year

- Naming compounds
- Writing formulas of compounds
- Distinguishing extensive vs. intensive properties
- Explaining the differences between element, compound, mixture
- Doing metric conversions
- Calculating number of molecules
- Locating elements in Periodic Table
- Answering everything that is asked for in the problems

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## Naming Compounds

Keys to naming:

- Recognize what kind of compound (ionic, molecular, acid)
- Figure out the parts

1. Ionic compounds $NaCH_3COO$ $Ca(OH)_2$ $Mg(OCl)_2$
   - Identify the ions
   - Name the ions
2. Molecular compounds $P_2O_5$
   - Identify the elements
   - Use prefixes
3. Acids $HNO_3$ $HClO_4$
   - Identify the negatively charged ion (anion)
   - Morph the anion name into the acid name

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## Writing Formulas

Keys to writing formulas:

- Recognize what kind of name (ionic, molecular, acid)
- Figure out the parts

1. Ionic compounds
   - Determine the ions
   - Use charges to find the ratio of ions
2. Molecular compounds
   - Determine the elements
   - Change prefixes to subscripts
3. Acids
   - Morph the acid name into the anion name
   - Use charges to find the ratio of ions

Copper (II) phosphate
Hypochlorous acid
Dinitrogen monoxide
Chromium (III) carbonate
Nitrous acid
Diarsenic pentoxide

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## Metric Conversions and Sig Figs

- Metric conversions
  - Figure out what units converting from and to
  - Find equivalent measures (dominoes) and orient conversion factors properly so units cancel
- Significant figures
  - Addition rule
  - Multiplication rule
  - Order of operations (mathematically)

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## How Many Particles?

How many molecules of $N_2$ are in 10.0 grams of $N_2$?

Strategy:

1. What are you starting with? Grams
2. Where do you need to end? Molecules
3. What kind(s) of conversions are necessary to get there?

   Grams → Moles (molar mass)

   Moles → Molecules (Avogadro's number)

$$10.0g\ \mathrm{N_2} \times \frac{1\,mol}{28.02g} \times \frac{6.022\times10^{23}\,particles}{1\,mol} = 2.15\times10^{23}\,molecules\ \mathrm{of\ N_2}$$

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