CHEM 205 section 03

LECTURE #9 Thurs. Jan.31, 2008

ASSIGNED READINGS:

TODAY'S CLASS: continue Ch.4

NEXT CLASS: **start Ch.5**

Last chance for Chem 101 seminar: Mon. Feb.11th
5 pm, CC-320
sign up in SP-201.01

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Chapter 4: Chemical Equations & Stoichiometry

4.1 Chemical equations
4.2 Balancing chemical equations
4.3 Mass relationships in chemical reactions: stoichiometry
4.4 Reactions in which one reactant is present in limited supply
4.5 Percent yield
4.6 Chemical equations and chemical analysis

Chapter Goals:

- Balance equations
- Perform stoichiometric calculations
- Understand limiting reactant
- Calculate theoretical & percent yield
- Use stoichiometry to analyze a mixture of compounds or to determine formula of a compound

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Consider a rusting nail…

Imagine you place a nail (composed of iron) into an open beaker containing acidic water, which makes the nail gradually rust. Rust is composed of iron (III) oxide and iron (III) hydroxide.

You then remove the nail from the water, and all of the rust remains attached to the nail's surface.
Is the nail lighter, heavier, or the same as before it rusted?

a) lighter
b) same mass
c) heavier

- Explain briefly in words.

- Write an equation to explain the formation of iron (III) oxide.
  (The rxn that forms iron (III) hydroxide is harder to predict.)

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4.1 Reactions follow conservation of mass

John Dalton: "Chemical change involves a reorganization of the atoms in one or more substances."

← Antoine Lavoisier: "Matter can neither be created nor destroyed."
(18th century)

LAW OF CONSERVATION OF MATTER
applies during chemical reactions:

Total # atoms of each element
is the SAME in reactants and products

→ Two sides of chemical equations are "balanced" atom-by-atom

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4.2 Balancing chemical equations

Symbolic representation of a chemical reaction:

$$1\ CH_4(g) + 2\ O_2(g) \rightarrow 1\ CO_2(g) + 2\ H_2O(g)$$

"stoichiometric coefficients" indicate mole ratio in which substances react & form

Physical states of substances also usually indicated

Summarizes many ways to think about what's happening:

Zumdahl's **TABLE 3.2**

Reactants		Products
$CH_4(g) + 2O_2(g)$	⟶	$CO_2(g) + 2H_2O(g)$
1 molecule + 2 molecules	⟶	1 molecule + 2 molecules
1 mole + 2 moles	⟶	1 mole + 2 moles
6.022×10^{23} molecules + 2 (6.022×10^{23} molecules)	⟶	6.022×10^{23} molecules + 2 (6.022×10^{23} molecules)
16 g + 2 (32 g)		44 g + 2 (18 g)
80 g reactants	⟶	80 g products

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Equations can often be balanced "by inspection"

- START: find easiest element (*in <u>one</u> species on each side…*)
- Balance its "partner" next
- Hop LOGICALLY back & forth, from reactants to products, from one element to the next

Method of making sodium sulfate on industrial scale:
(it is used in paper manufacturing)

$$\underline{2}\ NaCl(aq) + \underline{1}\ SO_2(g) + \underline{1}\ H_2O(aq) + \underline{{}^1/_2}\ O_2(g) \rightarrow \underline{1}\ Na_2SO_4(aq) + \underline{2}\ HCl(aq)$$

Oxygen: **products = 4 O's total** (atoms…or moles…)
reactants = 3 → need 1 more only…from O_2
→ SO: need "half" an O_2
→ remember, really means ½ a MOLE!

$$2NaCl(aq) + SO_2(g) + H_2O(aq) + {}^1/_2O_2(g) \rightarrow Na_2SO_4(aq) + 2HCl(aq)$$

Or, avoid fractions by multiplying all coefficients by denominator:

$$4NaCl(aq) + 2SO_2(g) + 2H_2O(aq) + O_2(g) \rightarrow 2Na_2SO_4(aq) + 4HCl(aq)$$

Always verify that your final equation is actually balanced

- Fool-proof check:
  tabulate # atoms / element in reactants *vs.* products

$$2NaCl(aq) + SO_2(g) + H_2O(aq) + {}^1/_2O_2(g) \rightarrow Na_2SO_4(aq) + 2HCl(aq)$$

Or, to avoid fractions:

$$4NaCl(aq) + 2SO_2(g) + 2H_2O(aq) + O_2(g) \rightarrow 2Na_2SO_4(aq) + 4HCl(aq)$$

Element	# atoms in reactants	# atoms in products
✓ Na	4x1 = 4	2x2 = 4
✓ Cl	4x1 = 4	4x1 = 4
✓ S	2x1 = 2	2x1 = 2
✓ O	(2x2)+(2x1)+(1x2) = 8	2x4 = 8
✓ H	2x2 = 4	4x1 = 4

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Write balanced equations for these rxns (on your own)

Sulfuric acid ($H_2SO_{4\,(\ell)}$) can be formed via the reaction of sulfur dioxide gas, oxygen gas and water (liquid).

Sodium hypochlorite is used as a bleaching agent. It is produced by treating an aqueous solution of sodium hydroxide with gaseous chlorine. The products of the reaction are sodium hypochlorite, sodium chloride and water.

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Balance these reactions… **(on your own)**

A protective film of an extremely hard compound is deposited onto cutting tools using this reaction: *at 1000°C*

$$\_\_TiCl_4(g) + \_\_BCl_3(g) + \_\_H_2(g) \rightarrow \_\_TiB_2(s) + \_\_HCl(g)$$

Combustion of butane:

$$\_\_C_4H_{10}(g) + \_\_O_2(g) \rightarrow \_\_CO_2(g) + \_\_H_2O(l)$$

Industrial formation of elemental phosphorus:

$$\_\_Ca_3(PO_4)_2(s) + \_\_SiO_2(s) + \_\_C(s)$$

$$\rightarrow \_\_CaSiO_3(s) + \_\_P_4(s) + \_\_CO(g)$$

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4.3 Mass relationships in chemical reactions: *"Stoichiometry"*

USING BALANCED CHEMICAL EQUATIONS…
(and some math…)

We can figure out USEFUL QUANTITIES:

- How much reactant we need…
  …to make a certain amount of products
- How much product we should get…
  …starting from a certain amount of reactants

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Reactions occur in fixed mole ratios:
i.e., according to the balanced chemical equation

$$a\,\mathbf{A} + b\,\mathbf{B} \longrightarrow c\,\mathbf{C}$$

Stoichiometry of a reaction

= item-by-item ratio in which reactants react to form product

→ Molecule-by-molecule... *OR...* mole-by-mole...

→ Stoichiometric coefficients (a, b, c) from balanced reaction equation tell us mole ratios involved

Mix exact "stoichiometric quantities"

all reactants used up at same time

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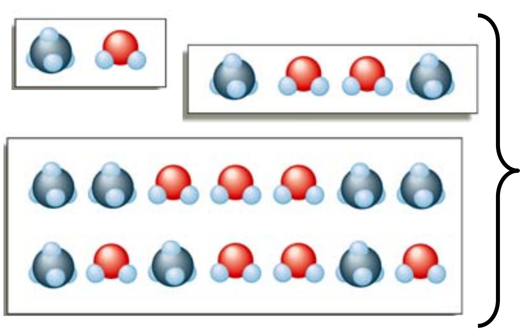
EXAMPLE: At high temperatures, methane and water react "one-to-one" according to the equation:

$$CH_4(g) + H_2O(g) \rightarrow 3H_2(g) + CO(g)$$



Three different "stoichiometric mixtures" for a 1:1 reaction

Zumdahl's Figure 3.9

If reactants present in exact ratios required by reaction, then when reaction is complete: no reactants will remain

only products will be present

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What quantity of CH_4 would generate 2.1 kg of H_2?

$$CH_{4}(g) \quad + \quad H_{2}O(g) \quad \rightarrow \quad 3\,H_{2}(g) \quad + \quad CO(g)$$

- *MM = 16.043 g/mol*
- *n = ?*
- *m = ?*

❸ From rxn stoichiometry:

$n_{CH4} = {}^{1}/_{3}\; n_{H2} = 347$ mol

$m = (347 \text{ mol})(16.043 \text{ g·mol}^{-1})$
$= 5.57\times10^{3}$ g *(1 extra SF…)*
$\approx$ **56 kg CH_4** *(2 SF)*

- *MM = 2.016 g/mol*
- Want: m = 2.1 kg H_2

❶ $n_{H2} = \dfrac{2100 \text{ g}}{2.016 \text{ g·mol}^{-1}}$

$= 1042$ mol *(2 extra SF…)*

❷ Reaction stoichiometry:

$$\frac{\text{need 1 mol } CH_4}{\text{to make 3 mol } H_2} = \frac{x \text{ mol } CH_4}{1042 \text{ mol } H_2}$$

⇨ x = 347 mol CH_4 *(1 extra SF)*
this much is required…

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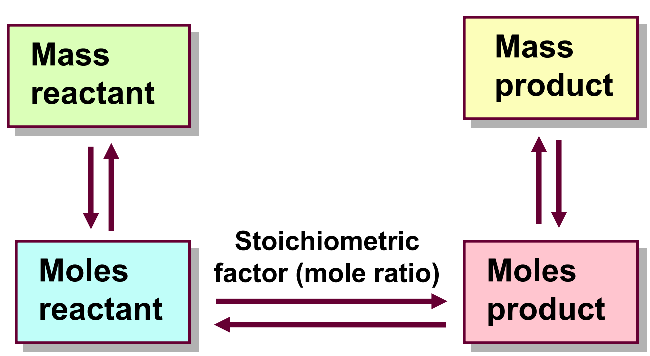
General strategy for stoichiometry problems

1. Write a balanced reaction equation.
2. Convert mass to moles *(practical link to # of molecules…).*
3. Set up mole ratios *i.e.,* "stoichiometric factor" for rxn.
4. Use mole ratios to calculate moles of desired substance.
5. Convert moles to grams.



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4.4 Reactions in which one reactant is present in limited supply…

What if we're sloppy?

$$a\,A + b\,B \longrightarrow c\,C$$

Reaction still obeys its stoichiometry!

- Approximate or unmeasured amounts used
  - One reactant: completely used up
  - Other reactant: some left over

Note: it's not just that we're sometimes sloppy…
It is common to use one reactant in large excess, because it can help "drive" a reaction towards products…
…more about that in Chem206

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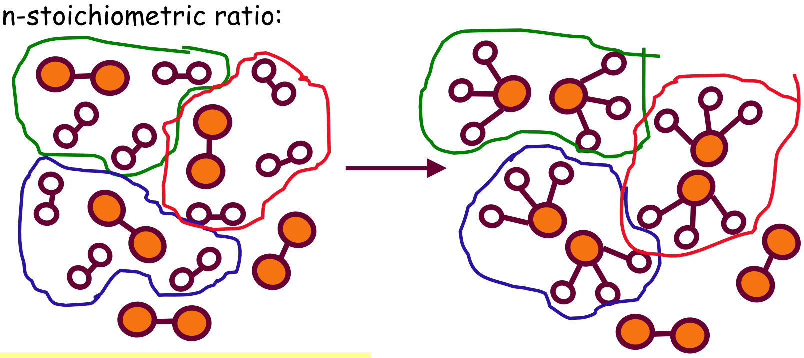
A real example: the Haber Process
reaction for industrial-scale production of ammonia

$$N_{2(g)} + 3\,H_{2(g)} \xrightarrow[\text{High pressure}]{\text{High temperature}} 2\,NH_{3(g)}$$

Reactants mixed in non-stoichiometric ratio:



Which reactant runs out first?
= **"limiting reactant"**
limits the yield of product

Some of one reactant is left over, **unreacted** = reactant **"in excess"**

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Learning to identify limiting reactant

If 2.00 moles of nitrogen and 2.00 moles of hydrogen react according to the equation below, what is the maximum amount of ammonia that could be produced?

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$$

a) 0.75 mol

b) 1.33 mol

c) 2.00 mol

d) 4.00 mol

e) 6.00 mol

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	Different ways to identify a rxn's limiting reactant	Advantages	Disadvantages
①	Determine how much "A" is required to use up all the "B" (in moles). **Do you have enough?**	***Intuitive, fool-proof.*** Linked to finding out how much unreacted A/B will remain.	Requires a stoichiometric calculation.
②	Determine actual A:B mole ratio & compare to rxn's requirement. **Do you have more of one reactant than you need?**	***Relatively fast.*** *Simplest if normalize in way that yields ratio with all #s > 1.*	***Be careful*** – easy to mix up ratios when in a hurry. ***No direct link*** to amounts used vs. left unreacted.
③	Calculate quantity of product formed if all of "A" used *vs* all of "B" used. **Which yields <u>less</u> product?**	No need to recalculate product yield.	***Does not require you to understand*** *reacted vs excess* *…try other ways instead…*

Use method ① or ② – whichever helps you visualize what really happens.

Q: What will be the yield of ammonia if we react 25.0 kg of nitrogen and 5.00 kg of hydrogen?

$N_{2(g)} + 3\ H_{2(g)} \longrightarrow 2\ NH_{3(g)}$

N_2:
- MM= 28.01 g/mol
- m = 25.0 kg = 25000 g

❶ $n_{N2} = \dfrac{25000\ g}{28.01\ g/mol}$ = 8.925×10² mol *(1 extra SF)*

H_2:
- MM = 2.016 g/mol
- m= 5.00 kg = 5000 g

❷ $n_{H2} = \dfrac{5000\ g}{2.016\ g/mol}$ = 2.480×10³ mol *(1 extra SF)*

NH_3:
- MM = 17.03 g/mol

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$N_{2(g)} + 3\ H_{2(g)} \longrightarrow 2\ NH_{3(g)}$

N_2:
- MM= 28.01 g/mol
- m = 25.0 kg = 25000 g

❶ $n_{N2} = \dfrac{25000\ g}{28.01\ g/mol}$ = 8.925×10² mol *(1 extra SF)*

❷ **To identify the limiting reactant:**

How much H_2 is required to use up all the N_2?
& Do we have that much?

H_2:
- MM = 2.016 g/mol
- m= 5.00 kg = 5000 g

❹ $n_{H2} = \dfrac{5000\ g}{2.016\ g/mol}$ = 2.480×10³ mol

❸ **To use up N_2, need:**

$\dfrac{3\ mol\ H_2}{1\ mol\ N_2} = \dfrac{x\ mol\ H_2}{8.925\times10^2\ mol\ N_2}$

x = (3/1)(8.925×10² mol) = 2.678×10³ mol H_2

❺ **Don't have enough!**
Thus: H_2 is limiting, N_2 is in excess

NH_3:
- MM = 17.03 g/mol

❻ Maximum NH_3 yield:

$\dfrac{2\ mol\ NH_3}{3\ mol\ H_2} = \dfrac{x\ mol\ NH_3}{2480.\ mol\ H_2}$

x = (2/3)(2480. mol) = 1.653×10³ mol NH_3

❼ Theoretical yield is:

= (1653mol)×(17.03g/mol)
= 28151 g
= 28.2 kg NH_3 *(3 SF)*

❽ *Remember to round off final answer to correct # SF*

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4.5 Percent yield ...less product than expected?

What if we **obtained 21.5 kg of NH_3** from our rxn mixture, but we had **expected 28.2 kg (theoretical yield)**?

$$\% \text{ yield} = \frac{\text{ACTUAL YIELD}}{\text{THEORETICAL YIELD}} \times 100\%$$

$$= (21.5 \text{ kg} / 28.2 \text{ kg}) \times 100\%$$
$$= 86.9\ \%$$

- **Reactions don't always proceed to completion! (<100% yield)**
  - Some prefer specific ratios of reactants/products; and/or
  - Side reactions can consume reactants or products.

(21) **NOTE: >100% yield can result from experimental error... Always make sure your sample is dry before weighing!**

On your own: applying % yield in problems

Bornite (Cu_3FeS_3) is a copper ore used in the production of copper. When heated in air, the following reaction occurs: **(unbalanced)**

$$Cu_3FeS_3(s) + O_2(g) \rightarrow Cu(s) + FeO(s) + SO_2(g)$$

If 2.50 tonnes of bornite is reacted with excess oxygen and the process has an 86.3% yield of copper, how much Cu(s) is produced?

Interpretation: a) bornite is limiting reactant (O_2 in excess);
b) reaction does not proceed to completion.

Approach:
1) Balance the equation.
2) Identify mole relationship between bornite & Cu
3) Calculate #mol of bornite used
4) Convert to theoretical yield of Cu
5) Correct this yield to 86.3% of theoretical yield.

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$Cu_3FeS_3(s)$ + $^7/_2$ $O_2(g)$ → 3 $Cu(s)$ + $FeO(s)$ + 3 $SO_2(g)$

$\boxed{2\ Cu_3FeS_3(s) + 7\ O_2(g) \rightarrow 6\ Cu(s) + 2\ FeO(s) + 6\ SO_2(g)}$

Bornite:
MM= 342.68 g/mol

m = 2.50 tonnes
= $2.50\text{x}10^3$ kg
= $2.50\text{x}10^6$ g

Thus: start with

$n = \frac{2.50\text{x}10^6 \text{ g}}{342.68 \text{ g/mol}}$

= 7295.5 mol

Oxygen:
excess

Note: can do the % yield correction at the mole quantity stage OR the mass stage …gives same result.

Copper:
MM= 63.546 g/mol
Note: 86.3% yield…

$\frac{2 \text{ mol } Cu_3FeS_3}{7295.5 \text{ mol } Cu_3FeS_3} = \frac{6 \text{ mol Cu}}{x \text{ mol Cu}}$

Thus, x = (6/2) x 7295.5
= 21886 mol Cu

m = 21886 mol x 63.546 g/mol
= 1390800 g theoretical yield

Actual yield = 86.3%
So, mass copper produced
= 0.863 x 1390800 g
= 1200260 g
= $1.20\text{x}10^6$ g
= 1.20 tonnes *3 SF*

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ASSIGNED READINGS

- BEFORE NEXT CLASS:

  Read rest of Ch. 3 & Ch. 4 (all)

  Master Ch.1-4 material & exercises

- Practice: naming/formulae of compounds
  determining formulae from mass %
  doing stoichiometry problems

Don't wait for problems to be assigned for tutorials!

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