Chapter 4: Chemical and Solution Stoichiometry

(Sections 4.1-4.4)

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Reaction Stoichiometry

- The coefficients in a balanced chemical equation specify the relative amounts in moles of each of the substances involved in the reaction

$$2\ C_4H_{10\,(g)} + 13\ O_{2\,(g)} \rightarrow 8\ CO_{2\,(g)} + 10\ H_2O_{\,(g)}$$

- 2 molecules of C_4H_{10} react with 13 molecules of O_2 to form 8 molecules of CO_2 and 10 molecules of H_2O
- **2 moles** of C_4H_{10} react with **13 moles** of O_2 to form **8 moles of CO_2** and **10 moles of H_2O**

Mole ratio

2 mol C_4H_{10} : **13** mol O_2 : **8** mol CO_2 : **10** mol H_2O

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Predicting Amounts from Stoichiometry

- The amount of any other substance in a chemical reaction can be determined from the amount of just one substance
- How much CO_2 can be made from 22.0 moles of C_4H_{10} in the combustion reaction of C_4H_{10}?

$$\mathbf{2\ C_4H_{10\,(g)}} + 13\ O_{2\,(g)} \rightarrow \mathbf{8\ CO_{2\,(g)}} + 10\ H_2O_{(g)}$$

$$22 \cancel{\text{ moles } C_4}H_{10} \times \frac{8 \text{ moles } CO_2}{2 \cancel{\text{ moles } C_4H_{10}}} = 88 \text{ moles } CO_2$$

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Practice

- According to the following equation, how many moles of water are made in the combustion of 0.10 moles of glucose?

$$C_6H_{12}O_6 + 6\ O_2 \rightarrow 6\ CO_2 + 6\ H_2O$$

Answer: **0.60 mol H_2O**

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Stoichiometry and Chemical Reactions

The most common stoichiometric problem will present you with a certain *mass of a reactant* and then ask the amount or *mass of product* that can be formed.

- This is called **mass-to-mass** stoichiometry problem

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Predicting Amounts from Stoichiometry – *Cont.*

- What if the mass of a substance is given, how do we know how much of another substance is needed (reactant) or is produced (product)?

For the balanced equation:
$$a\,\mathrm{A} + b\,\mathrm{B} \longrightarrow c\,\mathrm{C} + d\,\mathrm{D}$$

1. You cannot convert mass (g) of one substance directly to mass (g) of another substance in a given reaction.

$$\text{Grams of A} \not\longrightarrow \text{Grams of B}$$

2. However, you can convert mass to moles, then use their mole ratio to convert moles to grams of another substance.

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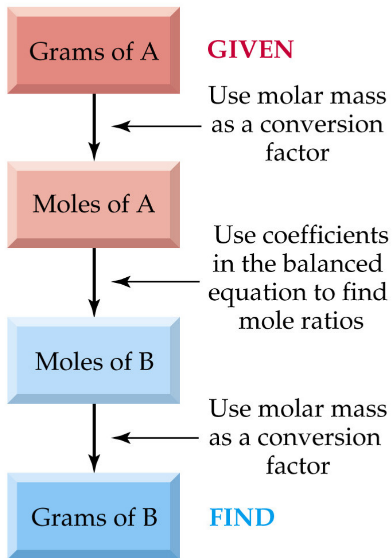
Predicting Amounts from Stoichiometry – *Cont.*

Mass-to-Mass Conversions



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http://wps.prenhall.com/

Solving Mass-Mass Stoichiometry

Goal: **To calculate the mass of product(s) given the mass of a reactant**

1. *Balance* the chemical equation.
2. Convert the known *mass* of reactant *to moles* using MM as a conversion factor.
3. Use the mole or *"stoichiometric "* ratio of reactant to product to convert moles of reactant to moles of product.
4. Finally, use the MM of the product to convert its *moles to mass.*

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Solving Mass-Mass Stoichiometry - *Cont.*

Example: Estimate the ***mass of CO_2*** produced in 2007 by the combustion of 3.5×10^{15} g gasoline.

- Assuming that gasoline is octane, C_8H_{18}, the equation for the reaction is

$$2\ C_8H_{18}\ (l) + 25\ O_2\ (g) \rightarrow 16\ CO_2\ (g) + 18\ H_2O\ (g)$$

- Follow the process:

g C_8H_{18} → **mol C_8H_{18}** → **mol CO_2** → **g CO_2**

Use MM | **Use mole ratio in balanced equation** | **Use MM**

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Example: Estimate the mass of CO_2 produced in 2007 by the combustion of 3.5×10^{15} g gasoline

Given:	3.4×10^{15} g C_8H_{18}
Find:	g CO_2
Conceptual Plan:	**g C_8H_{18}** → **mol C_8H_{18}** → **mol CO_2** → **g CO_2** $\frac{1\ mol}{114.22\ g}$ $\quad$ $\frac{16\ mol\ CO_2}{2\ mol\ C_8H_{18}}$ $\quad$ $\frac{44.01\ g}{1\ mol}$
Relationships:	$2\ C_8H_{18}\ (l) + 25\ O_2\ (g) \rightarrow 16\ CO_2\ (g) + 18\ H_2O\ (g)$ 1 mol C_8H_{18} = 114.22g; 1 mol CO_2 = 44.01g, 2 mol C_8H_{18}:16 mol CO_2

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$$\text{g } C_8H_{18} \xrightarrow{\frac{1\text{ mol}}{114.22\text{ g}}} \text{mol } C_8H_{18} \xrightarrow{\frac{16\text{ mol } CO_2}{2\text{ mol } C_8H_{18}}} \text{mol } CO_2 \xrightarrow{\frac{44.01\text{ g}}{1\text{ mol}}} \text{g } CO_2$$

Solution:

$$3.5\times10^{15}\ \cancel{\text{g } C_8H_{18}} \times \frac{1\text{ mol } C_8H_{18}}{114.22\ \cancel{\text{g } C_8H_{18}}} = 3.0643 \times 10^{13}\text{ mol } C_8H_{18}$$

$$3.0643 \times 10^{13}\ \cancel{\text{mol } C_8H_{18}} \times \frac{16\text{ mol } CO_2}{2\ \cancel{\text{mol } C_8H_{18}}} = 2.4514 \times 10^{14}\text{ mol } CO_2$$

$$2.4514 \times 10^{14}\ \cancel{\text{mol } CO_2} \times \frac{44.01\text{ g } CO_2}{1\ \cancel{\text{mol } CO_2}} = 1.0789 \times 10^{16}\text{ g } CO_2$$

$$= 1.1\times10^{16}\text{ g } CO_2$$

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Alternate Solution: One step with multiple conversion factors

Given: **Find:**	3.4×10^{15} g C_8H_{18} g CO_2
Conceptual Plan:	$\text{g } C_8H_{18} \xrightarrow{\frac{1\text{ mol}}{114.22\text{ g}}} \text{mol } C_8H_{18} \xrightarrow{\frac{16\text{ mol } CO_2}{2\text{ mol } C_8H_{18}}} \text{mol } CO_2 \xrightarrow{\frac{44.01\text{ g}}{1\text{ mol}}} \text{g } CO_2$
Relationships:	**MM C_8H_{18} = 114.22g/mol MM CO_2 = 44.01g/mol** **2 mol C_8H_{18}:16 mol CO_2** (fr. the balanced eq'n)
Solution:	

$$3.5\times10^{15}\ \cancel{\text{g } C_8H_{18}} \times \frac{1\ \cancel{\text{mol } C_8H_{18}}}{114.22\ \cancel{\text{g } C_8H_{18}}} \times \frac{16\ \cancel{\text{mol } CO_2}}{2\ \cancel{\text{mol } C_8H_{18}}} \times \frac{44.01\text{ g } CO_2}{1\ \cancel{\text{mol } CO_2}}$$

$$= 1.1\times10^{16}\text{ g } CO_2$$

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Mass-Mass Stoichiometry – *Cont.*

Your turn: (1) How many grams of chlorine gas can be liberated from the decomposition of 64.0 g of $AuCl_3$ by the reaction:

$$2\ AuCl_{3\ (s)} \rightarrow 2\ Au_{\ (s)} + 3\ Cl_{2\ (g)}$$

Answer: 22.4 g Cl_2

Relationships:

g $AuCl_3$ → mol $AuCl_3$ → mol Cl_2 → g Cl_2

$$\frac{1\ \text{mol}\ AuCl_3}{303.329\ \text{g}\ AuCl_3} \quad \frac{3\ \text{mol}\ Cl_2}{2\ \text{mol}\ AuCl_3} \quad \frac{70.906\ \text{g}\ Cl_2}{1\ \text{mol}\ Cl_2}$$

http://dbhs.wvusd.k12.ca.us/webdocs/Stoichiometry/Mass-Mass.html

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Limiting reactant, theoretical yield, and percent yield

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Tro: Chemistry: A Molecular Approach, 2/e

Limiting Reactant and Theoretical Yield

- In real life, we don't use the exact mole ratio of reactants as shown in the balanced equation
- In practice, an excess of one reactant is used for two reasons:
  - (1) To drive the reaction to completion
  - (2) To maximize the yield of products
- The reactant that is added in excess amount is called the ***excess reactant***, while the one in lower or limiting amount is called the *limiting reactant* or **limiting reagent**.

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Limiting Reactant and Theoretical Yield - *Cont.*

NOTE: The limiting reactant is:

(1) *completely consumed* in a chemical reaction. Thus, it

(2) *determines (or limits) the amount of product* formed.

- The max. amount of product that can form when all of the limiting reactant is used up is called the ***theoretical yield***.

Calculating theoretical yield:

1. Calculate yield assuming the first reactant is limiting.
2. Calculate yield assuming the 2nd reactant is limiting.
3. Choose the smaller of the two amounts.

- The *reactant that produces the smaller yield is the limiting reactant*

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Calc. Theoretical Yield - *Cont.*

Example: 31.84 g of aluminum and 73.15 g of sulfur are combined to form aluminum sulfide according to the equation:

$$\mathbf{Al_{(s)} + S_{(s)} \rightarrow Al_2S_{3\,(s)}}$$

Determine the limiting reactant and calculate the theoretical yield in grams of Al_2S_3.

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Theoretical vs. Actual Yield

The amount of product actually obtained is called the ***actual yield***.

Actual yield < Theoretical yield *WHY?*

(1) Not all the reactants may react
(2) Presence of significant side reactions
(3) Physical recovery of 100% of the sample may be impossible
- like getting all the peanut butter out of the jar

To calculate percent yield:

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

http://wine1.sb.fsu.edu/chm1045/notes/Stoich/Limiting/Stoich07.htm

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More Mass-Mass Stoichiometry Problems

Another example: ***Limiting reactant calculation***:
Consider the reaction: $2\,Al + 3\,I_2 \text{------>} 2\,AlI_3$

http://dbhs.wvusd.k12.ca.us/

aluminum iodine aluminum iodide

Determine the limiting reagent and the theoretical yield of the product if one starts with:
a) 1.20 mol Al and 2.40 mol iodine.
b) 1.20 g Al and 2.40 g iodine
c) How many grams of Al are left over in part b?

Solution:

a) Moles (instead of grams) are given so we use them directly. Use mole theoretical mole ratio (fr. the balanced eq.) and compare it with the actual ratio (given) to find the limiting reactant.

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$$2\,Al \quad + \quad 3\,I_2 \quad \text{------>} \quad 2\,AlI_3$$

a) Calculate yield of AlI_3 from 1.20 mol Al and 2.40 mol iodine.

MM Al = 26.982; MM I_2 = 253.80 and MM AlI_3 = 407.68 g/mol

Assume Al is limiting:

$$1.20 \text{ mol Al x } \frac{2 \text{ mol } AlI_3}{2 \text{ mol Al}} \text{ x } \frac{407.78 \text{ g } AlI_3}{1 \text{ mol } AlI_3} = 489 \text{ g } AlI_3$$

Now assume I_2 is limiting:

$$2.40 \text{ mol } I_2 \text{ x } \frac{2 \text{ mol } AlI_3}{3 \text{ mol } I_2} \text{ x } \frac{407.78 \text{ g } AlI_3}{1 \text{ mol } AlI_3} = 652 \text{ g } AlI_3$$

Obviously, 489 g is less than 652, so Al is the limiting reagent (lower yield from Al), making I_2 the excess reagent

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Solution – *Cont.*

(b) Determine the yield of the AlI_3 if one starts with 1.20 g Al and 2.40 g I_2. MM Al = 26.982; MM I_2 = 253.80 and MM AlI_3 = 407.68 g/mol

$$\mathbf{2\,Al + 3\,I_2 \text{ ------> } 2\,AlI_3}$$

WORK:

(i) Calc. yield of the AlI_3 assuming Al is limiting

$$1.20 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.982 \text{ g Al}} \times \frac{2 \text{ mol } AlI_3}{2 \text{ mol Al}} \times \frac{407.68 \text{ g } AlI_3}{1 \text{ mol } AlI_3} = 18.1 \text{ g } AlI_3$$

(ii) Calc. yield of the AlI_3 assuming I_2 is limiting

$$2.40 \text{ g } I_2 \times \frac{1 \text{ mol } I_2}{253.80 \text{ g } I_2} \times \frac{2 \text{ mol } AlI_3}{3 \text{ mol } I_2} \times \frac{407.68 \text{ g } AlI_3}{1 \text{ mol } AlI_3} = 2.57 \text{ g } AlI_3$$

(iii) Pick the lower of two yields because it came from the limiting reactant. This is the theoretical yield of product.

Thus:

- limiting reactant is I_2 and
- theoretical yield of AlI_3 is **2.57 g**

More Mass-Mass Stoichiometry Problems

Another HOMEWORK problem:
A 2.00 g sample of ammonia is mixed with 4.00 g of oxygen.

NH_3 (g) + O_2 (g) → NO (g) + H_2O (g)

Which is the limiting reactant? How much NO is produced?

First, balance the equation:

NH_3 (g) + O_2 (g) → NO (g) + H_2O (g)

Next we can use stoichiometry to calculate how much *NO* product is produced by each reactant. NOTE: It does not matter which product is chosen, but the same product must be used for both reactants so that the amounts can be compared.

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http://www.chem.tamu.edu/class/majors/tutorialnotefiles/limiting.htm

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4 NH_3 (g) + 5 O_2 (g) → 4 NO (g) + 6 H_2O (g)

Given: **2.00 g** **4.00 g**

$$2.00\ \text{g}\ \cancel{NH_3} \times \frac{1\ \cancel{\text{mol}\ NH_3}}{17.0\ \cancel{\text{g}\ NH_3}} \times \frac{4\ \cancel{\text{mol NO}}}{4\ \cancel{\text{mol}\ NH_3}} \times \frac{30.0\ \text{g NO}}{1\ \cancel{\text{mol NO}}} = 3.53\ \text{g NO} \quad \text{from } NH_3$$

$$4.00\ \text{g}\ \cancel{O_2} \times \frac{1\ \cancel{\text{mol}\ O_2}}{32.0\ \cancel{\text{g}\ O_2}} \times \frac{4\ \cancel{\text{mol NO}}}{5\ \cancel{\text{mol}\ O_2}} \times \frac{30.0\ \text{g NO}}{1\ \cancel{\text{mol NO}}} = \boxed{3.00\ \text{g NO}} \quad \text{from } O_2$$

The reactant that produces the lesser amount of product, in this case the **oxygen,** is the **limiting reactant**.

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