Welcome to our CHEM 4 lecture

Review question: Mole-to-mole ratios from balanced reactions
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1) Fermentation is a complex process of wine making in which glucose, $C_6H_{12}O_6$, is converted into ethanol, C_2H_5OH, and carbon dioxide. Starting with 500.4 g of glucose (molar mass = 180.18 g/mol), what is the maximum volume, in L, of ethanol (molar mass = 46.08 g/mol) that can be obtained by this process? The density of ethanol is 0.789 g/mL.

A) 12.0 L
B) 0.324 L
C) 0.0698 L
D) 20.8 L
E) 0.781 L
F) 2.95 L

See work shown on next slide…

Question continued from previous slide: Mole-to-mole ratios from balanced reactions

1) During fermentation glucose, $C_6H_{12}O_6$, is converted into ethanol, C_2H_5OH, and carbon dioxide. Starting with 500.4 g of glucose (molar mass = 180.18 g/mol), what is the maximum volume, in L, of ethanol (molar mass = 46.08 g/mol) that can be obtained by this process? The density of ethanol is 0.789 g/mL.

Answer:

Balanced equation: $C_6H_{12}O_6(s) \rightarrow 2\ C_2H_5OH(l) + 2\ CO_2(g)$

Flowchart: g $C_6H_{12}O_6$ $\rightarrow$ $\boxed{\text{mol } C_6H_{12}O_6 \rightarrow \text{mol } C_2H_5OH}$ $\rightarrow$ g C_2H_5OH $\rightarrow$ mL C_2H_5OH $\rightarrow$ L C_2H_5OH

from balanced reaction

Work:

$$(500.4\ \text{g}\ C_6H_{12}O_6)\left(\frac{1\ \text{mol}\ C_6H_{12}O_6}{180.18\ \text{g}\ C_6H_{12}O_6}\right)\boxed{\left(\frac{2\ \text{mol}\ C_2H_5OH}{1\ \text{mol}\ C_6H_{12}O_6}\right)}\left(\frac{46.08\ \text{g}\ C_2H_5OH}{1\ \text{mol}\ C_2H_5OH}\right)\left(\frac{1\ \text{mL}}{0.789\ \text{g}}\right)\left(\frac{10^{-3}\ \text{L}}{1\ \text{mL}}\right) = 0.324\ \text{L}$$

4sf *5sf* *∞sf* *4sf* *3sf* *∞sf* *keep 3sf*

CHEM 4 website: tinyurl.com/SacStateChem4

<table>
<tr><th>Week 15: December 7 (Monday)</th><th>December 9 (Wednesday)</th><th>December 11 (Friday)</th></tr>
<tr><td>Before class:
Read 8.1-8.4 [reaction calculations] (90 min)
PAL worksheets for week 15: A and B</td><td>Before class:
Read 8.5-8.6 [limiting reactants] (2 hours)</td><td>Before class:
I'll spend the review session answering your questions from Practice Final exams (A and B).
Today, before class is the last day to submit late homework for credit.</td></tr>
<tr><td>After class:
Today's PowerPoint slides and recording (45 min)
MasteringChemistry #30 (40 min) [Due: W, 12/9]
Prepare for our review session [F, 12/11] and final exam [see dates next week]. Practice finals: A and B (2 hours each).
You have until Dec 11 to complete your online CHEM 4 student evaluation in Canvas. Here is a video explaining the process.</td><td>After class:
Today's PowerPoint slides and recording (45 min)
MasteringChemistry #32 (40 min) [Due: F, 12/11]
Prepare for our review session [F, 12/11] and final exam [see dates next week]. Practice finals: A and B (2 hours each).
Before class on F, 12/11 is the last day to submit late homework for credit.</td><td>After class:
Finish preparing for our final exam [see dates next week]. Practice: A and B (2 hours each)
Verify your updated homework and clicker grades on Canvas (posted by 12 midnight).
Verify that you have credit for completing the Commit to Study program on Canvas (posted by 12 midnight).</td></tr>
<tr><th>Week 16: December 14 (Monday)</th><th></th><th>December 18 (Friday)</th></tr>
<tr><td>CHEM 4, Sec 01 (meets MFW @ 8 am)
Final exam time = 8:00 - 10:00 am
Covers: Cumulative, with a slight stress on material since last exam (sections 8.1 - 8.6).
Practice: A, B (2 hours each)
Log onto our Final using Canvas</td><td></td><td>CHEM 4, Sec 03 (meets MWF @ 10 am)
Final exam time = 8:00 - 10:00 am
Covers: Cumulative, with a slight stress on material since last exam (sections 8.1 - 8.6).
Practice: A, B (2 hours each)
Log onto our Final using Canvas</td></tr>
</table>

Prerequisites for CHEM 1A/1E

Students can meet the *chemistry prerequisite* in any of the following ways:

- ~~Having a Chemistry Diagnostic Score of 35 or higher.~~ (not available during COVID)
- Completed CHEM ALEKS (CARA) with 85% of the topics completed.
- Passing CHEM 4 or CHEM 6A with a ***grade of C or better***.

Students can meet the *math prerequisite* in any of the following ways:

Math Prerequisite for CHEM 1A:	Math Prerequisite for CHEM 1E:
• A Math ALEKS PPL Score of 61 or higher • Successful completion of Math 12 or the equivalent • Current enrollment in Math 26A, Math 29 or a higher • Score of a 3 or higher on AB or BC Calculus AP Test • Ability to enroll in Math 26A or Math 29	• A Math ALEKS PPL score of 76 or higher • Successful completion of Math 29 or equivalent • Enrollment in a math course of Math 30 or higher • Score of a 3 or higher on AB or BC Calculus AP Test

- Questions can be directed to **Dr. Susan Crawford (crawford@csus.edu)** or **Dr. Roy Dixon (rdixon@csus.edu)**
- Chem department: https://www.csus.edu/college/natural-sciences-mathematics/chemistry/
- Math dept ALEKS PPL: https://www.csus.edu/college/natural-sciences-mathematics/math-placement-exam/


grazas
dankie
cảm ơn bạn
תודה
mochchakkeram
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CHEM 4 lecture

Wednesday – December 9, 2020

Sec 8.5 – 8.6

Limiting reactants, theoretical yield and % yield

Reading question: Based on your assigned reading (Sec 8.5-8.6)
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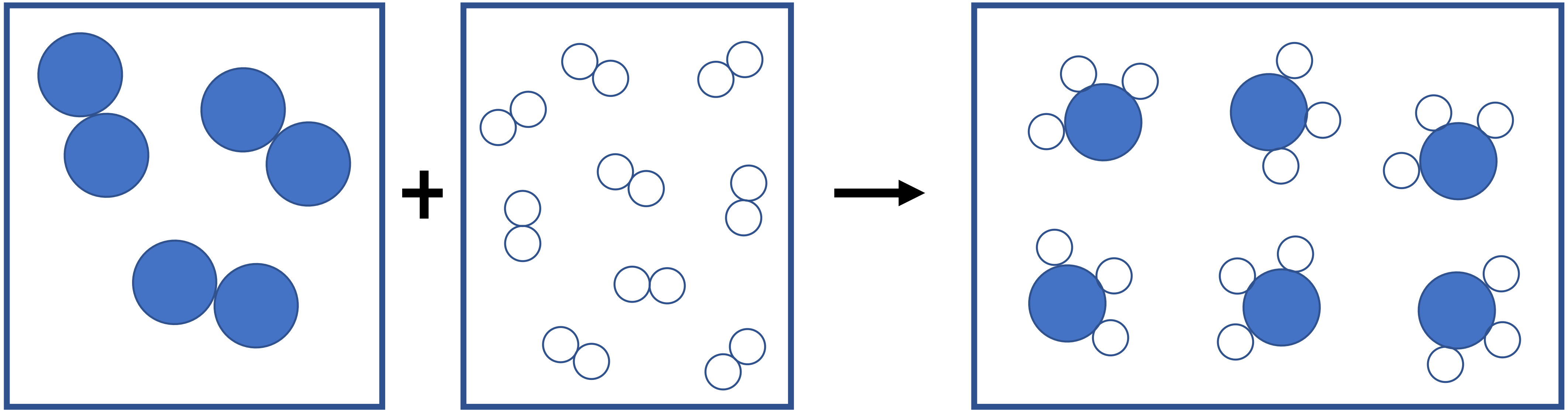
2) Which of the following statements is false?

A) The *limiting reactant* is the reactant that is completely used up during the reaction.
B) The *limiting reactant* is the reactant with the smallest starting mass.
C) The *limiting reactant* is the reactant that makes the least amount of product.
D) The *theoretical yield* is the amount of product that can be made based on the amount of limiting reactant available.
E) The *actual yield* is the amount of product that actually ends up being produced.
F) The percent yield = $\frac{\text{actual yield}}{\text{theoretical yield}}$ x 100%

Background information: Limiting reactants

Stoichiometric ratio: When you have exactly the right ratio of reactants described in the balanced chemical reaction.

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



1 mole N_2 + *3 moles H_2* → *2 moles NH_3*

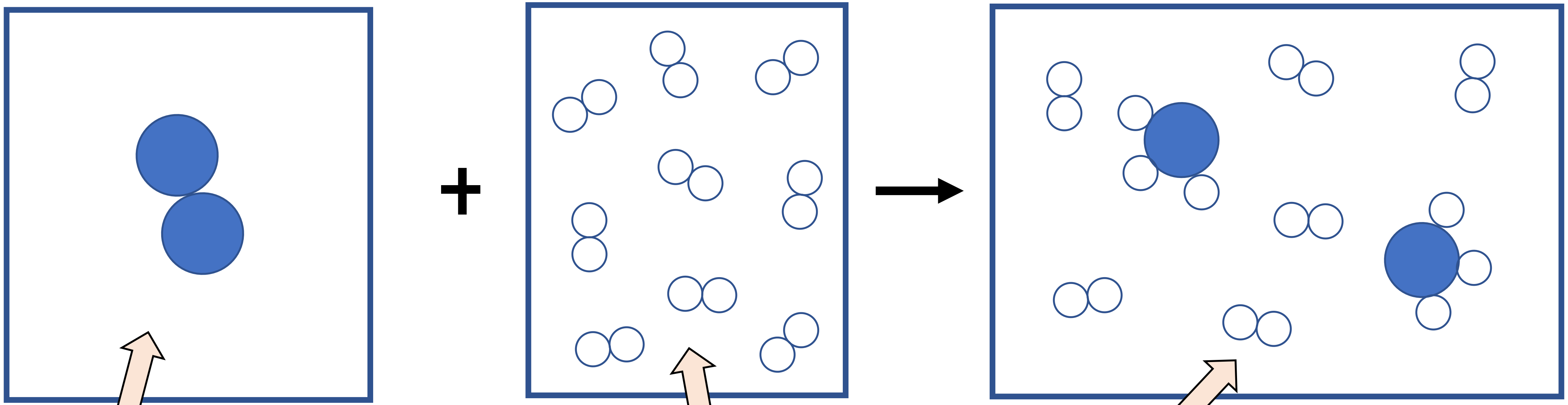
In lab: 28.02 g + 6.048 g → 34.07 g

Note: Even though we have more H_2 molecules, they don't necessarily weigh more than the N_2.

Background information: Limiting reactants

Non-stoichiometric ratio: When you don't have exactly the right ratio of reactants described in the balanced chemical reaction. You run out of a reactant (it is limiting).

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



N_2 is the *limiting reactant*.
It controls how much product you can make and you have 0 g left after the reaction is done.

The H_2 is the *excess reactant*.

After the reaction is done, all of the *limiting reactant* (N_2) is used up to make product (NH_3) but there is *excess reactant* (H_2) left over.

Sample calculation: Limiting reactant

Example: What is the maximum mass of solid barium phosphate (601.84 g/mol) you can make if you start with 6.40 g of barium nitrate (261.32 g/mol) and 3.50 g of sodium phosphate (163.94 g/mol)?

Answer: $2Na_3PO_4(aq) + 3Ba(NO_3)_2(aq) \rightarrow Ba_3(PO_4)_2(s) + 6NaNO_3(aq)$

Must do 2 calculations… one for each reactant:

Limiting reactant = makes the smaller amount of product

#1) Calculation based on starting mass of $\boxed{Ba(NO_3)_2:}$

$$(6.40 \text{ g } Ba(NO_3)_2)\left(\frac{1 \text{ mol } Ba(NO_3)_2}{261.32 \text{ g } Ba(NO_3)_2}\right)\left(\frac{1 \text{ mol } Ba_3(PO_4)_2}{3 \text{ mol } Ba(NO_3)_2}\right)\left(\frac{601.84 \text{ g } Ba_3(PO_4)_2}{1 \text{ mol } Ba_3(PO_4)_2}\right) = \boxed{4.91 \text{ g } Ba_3(PO_4)_2}$$

Theoretical yield = the smaller calculated mass. It is the max amount that can be made.

#2) Calculation based on starting mass of $\boxed{Na_3PO_4:}$

$$(3.50 \text{ g } Na_3PO_4)\left(\frac{1 \text{ mol } Na_3PO_4}{163.94 \text{ g } Na_3PO_4}\right)\left(\frac{1 \text{ mol } Ba_3(PO_4)_2}{2 \text{ mol } Na_3PO_4}\right)\left(\frac{601.84 \text{ g } Ba_3(PO_4)_2}{1 \text{ mol } Ba_3(PO_4)_2}\right) = 6.42 \text{ g } Ba_3(PO_4)_2$$

The Na_3PO_4 is in excess. We have enough Na_3PO_4 to make 6.42 g of product, but we run out of $Ba(NO_3)_2$ first.

Sample calculation continued: % yield

- When you are actually in CHEM 1a/1e lab, there are reasons why you don't make as much product as the theoretical yield would predict (for example, if you leave behind a few drops when you are combining two solutions or if you spill some).
- So we consider the *% yield* which compares how much you *actually* made to the maximum, *theoretical* amount that you calculate you could have made:

$$\text{percent yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \text{ x } 100\%$$

Example: What is the % yield of $Ba_3(PO_4)_2$ from the previous problem, if we actually made 4.55 g in lab? Remember, our theoretical yield was 4.91 g.

Answer:

$$\text{percent yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \text{ x } 100\% = \left(\frac{4.55\text{ g}}{4.91\text{ g}}\right) \text{ x } 100\% = 92.7\ \%$$

Progress question: Limiting reactant and % yield
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3) When ignited, hydrogen gas and oxygen gas react explosively to produce water. Find the limiting reactant, the theoretical yield of H_2O, and the percent yield if you make 9.15 g of H_2O when starting with 5.00 g hydrogen gas and 10.0 g oxygen gas.

A) O_2 is limiting; 11.3 g H_2O; 123% yield

B) O_2 is limiting; 11.3 g H_2O; 81.0% yield

C) O_2 is limiting; 44.7 g H_2O; 20.5% yield

D) O_2 is limiting; 44.7 g H_2O; 48.9% yield

E) H_2 is limiting; 11.3 g H_2O; 81.0 % yield

F) H_2 is limiting; 11.3 g H_2O; 123% yield

G) H_2 is limiting; 44.7 g H_2O; 20.5% yield

H) H_2 is limiting; 44.7 g H_2O; 48.9% yield

See work shown on next slide…

Question continued from previous slide: Limiting reactant and % yield

3) When ignited, hydrogen gas and oxygen gas react explosively to produce water. Find the limiting reactant, the theoretical yield of H_2O, and the percent yield if you make 9.15 g of H_2O when starting with 5.00 g hydrogen gas and 10.0 g oxygen gas.

Answer: $2\ H_2(g) + O_2\ (g) \rightarrow 2\ H_2O(l)$

$$(5.00\text{ g H}_2)\left(\frac{1\text{ mol H}_2}{2.016\text{ g H}_2}\right)\left(\frac{2\text{ mol H}_2\text{O}}{2\text{ mol H}_2}\right)\left(\frac{18.02\text{ g H}_2\text{O}}{1\text{ mol H}_2\text{O}}\right) = 44.7\text{ g H}_2\text{O}$$

$$(10.0\text{ g }\boxed{\text{O}_2})\left(\frac{1\text{ mol O}_2}{32.00\text{ g O}_2}\right)\left(\frac{2\text{ mol H}_2\text{O}}{1\text{ mol O}_2}\right)\left(\frac{18.02\text{ g H}_2\text{O}}{1\text{ mol H}_2\text{O}}\right) = \boxed{11.3\text{ g H}_2\text{O}}$$

Theoretical yield

Limiting reactant

$$\%\text{ yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \times 100\% = \left(\frac{9.15\text{ g}}{11.3\text{ g}}\right) \times 100\% = 81.0\ \%$$

Clicker question: Limiting reactant and % yield
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4) Solid lithium metal and nitrogen gas combine in a synthesis-type reaction to produce solid lithium nitride. If 5.00 g of each reactant undergo a reaction with 80.5% yield, how many grams of product are obtained?

A) 6.73 g lithium nitride
B) 10.4 g lithium nitride
C) 9.98 g lithium nitride
D) 15.4 g lithium nitride
E) 6.22 g lithium nitride
F) 5.00 g lithium nitride

See work shown on next slide…

Question continued from previous slide: Limiting reactant and % yield

4) Solid lithium metal and nitrogen gas combine in a synthesis-type reaction to produce solid lithium nitride. If 5.00 g of each reactant undergo a reaction with 80.5% yield, how many grams of product are obtained?

Answer: $6\ Li(s) + N_2(g) \rightarrow 2\ Li_3N(s)$

Limiting reactant

Theoretical yield

$$(5.00\ \text{g}\ \boxed{\text{Li}})\left(\frac{1\ \text{mol Li}}{6.941\ \text{g Li}}\right)\left(\frac{2\ \text{mol Li}_3\text{N}}{6\ \text{mol Li}}\right)\left(\frac{34.83\ \text{g Li}_3\text{N}}{1\ \text{mol Li}_3\text{N}}\right) = \boxed{8.36\ \text{g Li}_3\text{N}}$$

$$(5.00\ \text{g N}_2)\left(\frac{1\ \text{mol N}_2}{28.02\ \text{g N}_2}\right)\left(\frac{2\ \text{mol Li}_3\text{N}}{1\ \text{mol N}_2}\right)\left(\frac{34.83\ \text{g Li}_3\text{N}}{1\ \text{mol Li}_3\text{N}}\right) = 12.4\ \text{g Li}_3\text{N}$$

This time we know the % yield and theoretical yield and want to find the actual yield:

$$80.5\ \% = \left(\frac{\text{actual yield}}{8.36\ \text{g}}\right) \times 100\%$$

so, actual yield = $6.73\ g\ Li_3N$