CHEMISTRY 109 – Help Sheet #5
REVIEW (Part V): Stoichiometry (Part II)

***** Review the appropriate topics for your lecture*****

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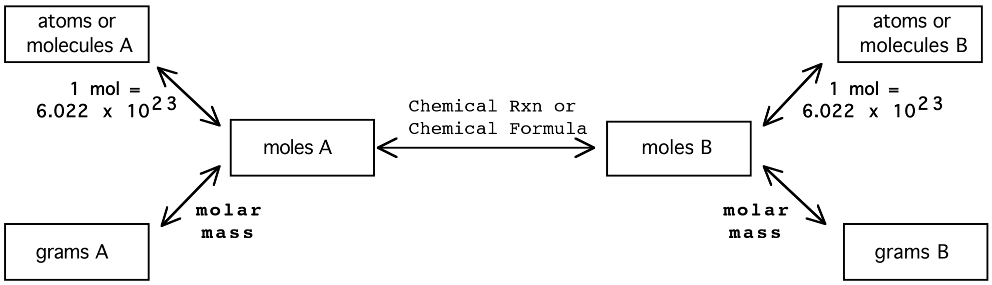
https://clc.chem.wisc.edu (Resources page)

Nuggets: *Stoichiometry, Limiting reagents, Percent yield, Empirical formula from combustion*



STOICHIOMETRY – calculations using stoichiometric coefficients from balanced reactions

Example 1: How many grams $NH_3(g)$ can be prepared from 84.0g $N_2(g)$ in this reaction: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

Answer 1: This is a "grams A → grams B" calculation (see diagram above); it requires 3 conversions (the steps are shown in the flow chart)

Step 1: $N_2 + 3H_2 \longrightarrow 2NH_3$

Step 2: 3-step conversion: $gA\left(\frac{1molA}{gA}\right)\left(\frac{molB}{molA}\right)\left(\frac{gB}{1molB}\right) = gB$

Step 3: 3-step conversion: $84.0gN_2\left(\frac{1molN_2}{28.0gN_2}\right)\left(\frac{2molNH_3}{1molN_2}\right)\left(\frac{17.0gNH_3}{1molNH_3}\right) = 102gNH_3$

LIMITING REAGENTS: one reagent runs out first – this is the limiting reagent; *a limiting reagent problem can be identified when 2 reactant quantities are given in the problem;* many ways to solve these types of problems - one way: calculate the amount of products possible from each reactant quantity; the smaller amount produced is the theoretical amount that can be made; the reactant that gives this smaller amount is the limiting reagent

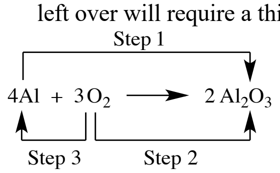
Example 2: a. How many grams $Al_2O_3(s)$ can be made from 25.0g Al(s) and 20.0g $O_2(g)$ using: $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$?

b. Which reactant is the limiting reagent?

c. Which reagent is the excess reagent?

d. How much of the excess reagent remains after the reaction is complete?

Answer 2: This is a limiting reagent problem since ***2 reactant quantities were given in the problem.*** The amount produced can be solved by doing two "grams A → grams B" calculations, and then comparing the two possible amounts of $Al_2O_3(s)$ produced. The amount of excess reagent left over will require a third "grams A → grams B" calculation requiring 3 steps.



Step 1: gA → gB calculation: g Al → g Al_2O_3

$25.0gAl\left(\frac{1molAl}{27.0gAl}\right)\left(\frac{2molAl_2O_3}{4molAl}\right)\left(\frac{102gAl_2O_3}{1molAl_2O_3}\right) = 47.2gAl_2O_3$

Step 2: gA → gB calculation: g O_2 → g Al_2O_3

$20.0gO_2\left(\frac{1molO_2}{32.0gO_2}\right)\left(\frac{2molAl_2O_3}{3molO_2}\right)\left(\frac{102gAl_2O_3}{1molAl_2O_3}\right) = 42.5gAl_2O_3$

a. 42.5g Al_2O_3 {the smaller quantity of $Al_2O_3(s)$ is how much can theoretically be produced}

b. $O_2(g)$ is the LR {since it produced the smaller amount $Al_2O_3(s)$}

c. Al(s) is the excess reagent {since it produced the larger amount $Al_2O_3(s)$}

d. 2.5g Al left over {Amount Excess Reagent Left Over = Starting Amount Excess Reagent – Used Amount Excess Reagent; Used amount can be calculated with "grams A → grams B" with the calculation: LR → Excess reagent:

Step 3: gA → gB calculation: g O_2 → g Al ***used***

$20.0gO_2\left(\frac{1molO_2}{32.0gO_2}\right)\left(\frac{4molAl}{3molO_2}\right)\left(\frac{27.0gAl}{1molAl}\right) = 22.5gAl$ used

22.5g Al is the amount of the Excess Reagent ***used***

Amount Left Over = 25.0g Al – 22.5g Al = 2.5g Al left over}

$$\textbf{PERCENT YIELD} = \left(\frac{\textbf{actual yield}}{\textbf{theoretical yield}}\right) \textbf{x 100\%}$$

Actual yield is the actual amount obtained and is always given in the problem; theoretical yield is usually calculated. In a lab setting, if the %yield > 100% then there is an error (e.g., the sample may be wet, etc.)

EMPIRICAL FORMULA from mass CO_2 and H_2O (combustion reaction)

A. Compound contains **C and H only**

1. Convert $gCO_2 \rightarrow$ mol $CO_2 \rightarrow$ **mol C**
2. Convert $gH_2O \rightarrow$ mol $H_2O \rightarrow$ **mol H**
3. Write formula and divide by smallest moles
4. If needed, fractions: $^1/_2$ (0.5) → x 2; $^1/_3$ or $^2/_3$ (0.33, 0.66) → x 3; $^1/_4$ or $^3/_4$ (0.25, 0.75) → x 4

B. Compound contains **C, H, and X** (used for a C, H, and X compound; X is a third element often O or N; example below has X = O)

1. Convert $gCO_2 \rightarrow$ mol $CO_2 \rightarrow$ **mol C**; mol C → **gC** (need both mol C *and* gC)
2. Convert $gH_2O \rightarrow$ mol $H_2O \rightarrow$ **mol H**; mol H → **gH** (need both mol H *and* gH)
3. Calculate gO from: total g sample = gC + gH + gO (gO = total g sample - gC - gH)
4. Convert gO → **mol O**
5. Write formula and divide by smallest moles
6. If needed, fractions: $^1/_2$ (0.5) → x 2; $^1/_3$ or $^2/_3$ (0.33, 0.66) → x 3; $^1/_4$ or $^3/_4$ (0.25, 0.75) → x 4

Example 3: Butane, a ***hydrocarbon,*** was burned and 15.14g CO_2 and 7.751g H_2O are recovered. What is the empirical formula of butane?
Answer 3: C_2H_5

1. **mol C:** $15.14gCO_2\left(\frac{1molCO_2}{44.01gCO_2}\right)\left(\frac{1molC}{1molCO_2}\right) = 0.3441molC$

2. **mol H:** $7.751gH_2O\left(\frac{1molH_2O}{18.02gH_2O}\right)\left(\frac{2molH}{1molH_2O}\right) = 0.8603molH$

3. **Write formula:** $C_{0.3441}H_{0.8603}$ and divide by smallest number of mol = 0.3441: $C_{\frac{0.3441}{0.3441}}H_{\frac{0.8603}{0.3441}} \rightarrow C_1H_{2.500}$

4. **Fractions:** multiply by 2: $C_{(1 \text{ x } 2)}H_{(2.500 \text{ x } 2)} \rightarrow$ empirical formula = $\mathbf{C_2H_5}$

Example 4: When 2.000g of a compound containing ***carbon, hydrogen, and oxygen*** is combusted, 1.912g CO_2 and 0.7830g H_2O are recovered. What is the empirical formula?
Answer 4: CH_2O_2

1. **mol C:** $1.912gCO_2\left(\frac{1molCO_2}{44.01gCO_2}\right)\left(\frac{1molC}{1molCO_2}\right) = 0.04344molC$ and **gC:** $0.04344molC\left(\frac{12.01gC}{1molC}\right) = 0.5217gC$

2. **mol H:** $0.7830gH_2O\left(\frac{1molH_2O}{18.02gH_2O}\right)\left(\frac{2molH}{1molH_2O}\right) = 0.08690molH$ and **gH:** $0.08690molH\left(\frac{1.008gH}{1molH}\right) = 0.08760gH$

3. **gO:** g_{sample} = gO + gC + gH → solve for gO: gO = g_{sample} – gC – gH = 2.000 – 0.5217 – 0.08760 = 1.3907g O

4. **mol O:** $1.3907gO\left(\frac{1molO}{16.00gO}\right) = 0.08692molO$

5. **Write formula:** $C_{0.04344}H_{0.08690}O_{0.08692}$; divide by smallest number of mol = 0.04344: $C_{\frac{0.04344}{0.04344}}H_{\frac{0.08690}{0.04344}}O_{\frac{0.08692}{0.04344}} = C_1H_{2.000}O_{2.000}$;

Empirical formula = $\mathbf{CH_2O_2}$

6. **Fractions:** no fractions; this step not needed

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1. Nitric acid, HNO_3, is manufactured by the Oswald process, in which nitrogen dioxide, NO_2, reacts with H_2O. How many grams of NO_2 are required to produce 5.00g HNO_3.

$$3NO_2(g) + H_2O(l) \rightarrow 2HNO_3(aq) + NO(g)$$

a. 7.50 b. 15.0 c. 3.65 d. 0.120 e. 5.48

2. Calculate the masses (in grams) of Cr_2O_3 (chromium(III) oxide), N_2, and H_2O produced from 10.8g of $(NH_4)_2Cr_2O_7$ (ammonium dichromate) in the following balanced reaction:

$$(NH_4)_2Cr_2O_7(s) \rightarrow Cr_2O_3(s) + N_2(g) + 4H_2O(l)$$

3. a. When 9.20g of C_2H_6O are reacted with 40.6g of PBr_3, what mass of C_2H_5Br can be produced?

$$3C_2H_6O + PBr_3 \rightarrow 3C_2H_5Br + H_3PO_3$$

b. Which reactant is the limiting reagent?
c. Which reactant is the excess reagent?
d. How much of the excess reagent is left over?
e. If 10.9g PBr_3 was recovered after the experiment, what was the %yield of this reaction?

4. If 75.0g of SiO_2 and 30.0g C react according to the equation below, what is the maximum number of moles of CO that can be produced? $SiO_2 + C \rightarrow CO + SiO$

a. 1.25 b. 1.67 c. 2.25 d. 2.50 e. none of these

5. I. Given the following balanced combustion reaction below, if there were 3.0mol of C_2H_4 and 6.0mol of O_2, how many moles of CO_2 could theoretically be produced?

$$C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$$

a. 6mol b. 4mol c. 9mol d. 3mol e. 7mol

II. After the above reaction was completed, 3.5mol of CO_2 were actually obtained. What is the percent yield of this reaction? (Note: You can only do this part with the correct answer to part I.)

6. Consider the reaction:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

Identify the limiting reagent in each of the reaction mixtures below. Note: This is not a multiple choice question.

a. 0.50mol of H_2 and 0.75mol of O_2
b. 0.80mol of H_2 and 0.75mol of O_2
c. 1.00g of H_2 and 0.35mol of O_2
d. 5.00g of H_2 and 64.00g of O_2

7. When a compound containing carbon and hydrogen is combusted, 3.38g CO_2 and 0.692g H_2O are recovered. a. What is the empirical formula? b. The molar mass of the compound is 78.1g/mol. What is the molecular formula?

8. When 5.000g of a compound containing carbon, hydrogen, and oxygen is combusted, 8.910g CO_2 and 3.648g H_2O are recovered. a. What is the empirical formula? b. The molar mass of the compound is 74.1g/mol. What is the molecular formula?

9. A compound contains carbon, hydrogen, and nitrogen. When 0.9731 gram of this substance is combusted, 2.3744grams CO_2 and 1.2153grams of H_2O are collected. What is the empirical formula?

ANSWERS

1. e $\{5.00\text{g HNO}_3\left(\frac{1\text{mol HNO}_3}{63.02\text{g HNO}_3}\right)\left(\frac{3\text{mol NO}_2}{2\text{mol HNO}_3}\right)\left(\frac{46.01\text{g NO}_2}{1\text{mol NO}_2}\right)=5.476\text{g NO}_2\}$

2. 6.52g Cr_2O_3 $\{10.8g\,(NH_4)_2Cr_2O_7\left(\frac{1mol\,(NH_4)_2Cr_2O_7}{252.08g\,(NH_4)_2Cr_2O_7}\right)\left(\frac{1mol\,Cr_2O_3}{1mol\,(NH_4)_2Cr_2O_7}\right)\left(\frac{152.00g\,Cr_2O_3}{1mol\,Cr_2O_3}\right)=65.12g\,Cr_2O_3\}$

1.20g N_2 $\{10.8g\,(NH_4)_2Cr_2O_7\left(\frac{1mol\,(NH_4)_2Cr_2O_7}{252.08g\,(NH_4)_2Cr_2O_7}\right)\left(\frac{1mol\,N_2}{1mol\,(NH_4)_2Cr_2O_7}\right)\left(\frac{28.02g\,N_2}{1mol\,N_2}\right)=1.200g\,N_2\}$

3.09g H_2O $\{10.8g\,(NH_4)_2Cr_2O_7\left(\frac{1mol\,(NH_4)_2Cr_2O_7}{252.08g\,(NH_4)_2Cr_2O_7}\right)\left(\frac{4mol\,H_2O}{1mol\,(NH_4)_2Cr_2O_7}\right)\left(\frac{18.02g\,H_2O}{1mol\,H_2O}\right)=3.088g\,H_2O\}$

3. a. 21.8g C_2H_5Br {limiting reagent problem since 2 reactant quantities are given;

$9.20g\,C_2H_6O\left(\frac{1mol\,C_2H_6O}{46.07g\,C_2H_6O}\right)\left(\frac{3mol\,C_2H_5Br}{3mol\,C_2H_6O}\right)\left(\frac{108.96g\,C_2H_5Br}{1mol\,C_2H_5Br}\right)=21.76g\,C_2H_5Br$;

$40.6g\,PBr_3\left(\frac{1mol\,PBr_3}{270.67g\,PBr_3}\right)\left(\frac{3mol\,C_2H_5Br}{1mol\,PBr_3}\right)\left(\frac{108.96g\,C_2H_5Br}{1mol\,C_2H_5Br}\right)=49.03g\,C_2H_5Br$;

the smaller amount of C_2H_5Br is the theoretical yield that can be produced}

b. C_2H_6O is LR {the chemical that produced the smaller amount of C_2H_5Br}

c. PBr_3 is the excess reagent {the chemical that produced the larger amount of C_2H_5Br}

d. 22.6g PBr_3 {Left over = starting amount – amount used; amount used = g A → g B going from LR → Excess Reagent;

$9.20g\,C_2H_6O\left(\frac{1mol\,C_2H_6O}{46.07g\,C_2H_6O}\right)\left(\frac{1mol\,PBr_3}{3mol\,C_2H_6O}\right)\left(\frac{270.67g\,PBr_3}{1mol\,PBr_3}\right)=18.02g\,C_2H_5Br$ ***used***;

Left over = 40.6g PBr_3 – 18.02g PBr_3 = 22.58g PBr_3 left over}

e. 50.0% {%yield = (actual yield/theoretical yield) x 100%; %yield = (10.9g C_2H_5Br/21.8g C_2H_5Br) x 100% = 50.0%}

4. a {limiting reagent problem since 2 reactant quantities are given;

$75.0g\,SiO_2\left(\frac{1mol\,SiO_2}{60.09g\,SiO_2}\right)\left(\frac{1mol\,CO}{1mol\,SiO_2}\right)=1.248mol\,CO$; $30.0g\,C\left(\frac{1mol\,C}{12.01g\,C}\right)\left(\frac{1mol\,CO}{1mol\,C}\right)=2.498mol\,CO$;

the smaller amt of CO is the theoretical yield}

5. I. b {limiting reagent problem; $(3mol\,C_2H_4)\left(\frac{2mol\,CO_2}{1mol\,C_2H_4}\right)=6mol\,CO_2$; $(6mol\,O_2)\left(\frac{2mol\,CO_2}{3mol\,O_2}\right)=4mol\,CO_2$;

choose smaller amount → 4mol CO_2}

II. 87.5% $\{\%yield=\frac{\text{actual yield}}{\text{theoretical yield}}\ x\ 100\%$; $\%yield=\left(\frac{3.5mol\,CO_2}{4.0mol\,CO_2}\right)x\ 100\%=87.5\%\}$

6. a. H_2 b. H_2 c. O_2 d. H_2

7. a. CH $\{3.38g\,CO_2\left(\frac{1mol\,CO_2}{44.01g\,CO_2}\right)\left(\frac{1mol\,C}{1mol\,CO_2}\right)=0.07680mol\,C$; $0.692g\,H_2O\left(\frac{1mol\,H_2O}{18.02g\,H_2O}\right)\left(\frac{2mol\,H}{1mol\,H_2O}\right)=0.07680mol\,H$;

write formula: $C_{0.07680}H_{0.07680}$; divide by 0.07680: $C_{\frac{0.07680}{0.07680}}H_{\frac{0.07680}{0.07680}}\rightarrow C_1H_1$ = EF}

b. C_6H_6 $\{ratio=\frac{\text{molar mass}_{\text{molecular formula}}}{\text{molar mass}_{\text{empirical formula}}}=\frac{78.1}{13.02}=5.998=6$; multiply EF by 6; $C_{(1\ x\ 6)}H_{(1\ x\ 6)}=C_6H_6$ = MF}

8. a. $C_3H_6O_2$ $\{8.910g\ CO_2\left(\frac{1mol\ CO_2}{44.01g\ CO_2}\right)\left(\frac{1mol\ C}{1mol\ CO_2}\right) = 0.2025mol\ C$; gC = 0.2025mol C x (12.01g C/1mol C) = 2.4320g C;

$3.648g\ H_2O\left(\frac{1mol\ H_2O}{18.02g\ H_2O}\right)\left(\frac{2mol\ H}{1mol\ H_2O}\right) = 0.4049mol\ H$; gH = 0.4049mol H x (1.008g H/1mol H) = 0.4081g H;

$g_{sample} = gO + gC + gH \rightarrow$ solve for gO: $gO = g_{sample} - gC - gH = 5.00 - 2.4320 - 0.4081 = 2.1599g\ O$

$2.1599g\ O\left(\frac{1mol\ O}{16.00g\ O}\right) = 0.1350mol\ O$;

$C_{0.2025}H_{0.4049}O_{0.1350}$; divide by 0.1350: $C_{\frac{0.2025}{0.1350}}H_{\frac{0.4049}{0.1350}}O_{\frac{0.1350}{0.1350}} = C_{1.500}H_{2.999}O_1$; fraction of 0.5 $\rightarrow$ $^1/_2$ $\rightarrow$ x2: $C_{(1.500 \times 2)}H_{(2.999 \times 2)}O_{(1 \times 2)} = C_3H_{5.998}O_2 = C_3H_6O_2$ = EF}

b. $C_3H_6O_2$ $\{ratio = \frac{molar\ mass_{molecular\ formula}}{molar\ mass_{empirical\ formula}} = \frac{74.1}{74.08} = 1.000$; multiply EF by 1; $C_{(3 \times 1)}H_{(6 \times 1)}O_{(2 \times 1)} = C_3H_6O_2$ = MF}

9. $C_4H_{10}N$ $\{2.3744g\ CO_2\left(\frac{1mol\ CO_2}{44.01g\ CO_2}\right)\left(\frac{1mol\ C}{1mol\ CO_2}\right) = 0.053951mol\ C$; $gC = 0.053951mol\ C\left(\frac{12.01g\ C}{1mol\ C}\right) = 0.64795g\ C$;

$1.2153g\ H_2O\left(\frac{1mol\ H_2O}{18.02g\ H_2O}\right)\left(\frac{2mol\ H}{1mol\ H_2O}\right) = 0.13488mol\ H$; $gH = 0.13488mol\ H\left(\frac{1.008g\ H}{1mol\ H}\right) = 0.13596g\ H$;

$g_{sample} = gN + gC + gH \rightarrow$ solve for gN: $gN = g_{sample} - gC - gH = 0.9731 - 0.64795 - 0.13596 = 0.18919g\ N$

$mol\ N = 0.18919g\ N\left(\frac{1mol\ N}{14.01g\ N}\right) = 0.013504mol\ N$; $C_{0.053951}H_{0.13488}N_{0.013504}$;

divide by 0.013504: $C_{\frac{0.053951}{0.013504}}H_{\frac{0.13488}{0.013504}}N_{\frac{0.013504}{0.013504}} = C_{3.9952}H_{9.9882}N_1 = C_4H_{10}N_1$ = EF}