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

**STOICHIOMETRY** – calculations with balanced reactions and the stoichiometric coefficients in those reactions

*Example 1:* How many grams NH$_3$(g) can be prepared from 84.0g N$_2$(g) using: N$_2$(g) + 3H$_2$(g) → 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$_2$ → 2NH$_3$

**Step 2 (3 conversions).** $g_A \left( \frac{1 \text{mol} A}{6.022 \times 10^{23} \text{atoms or molecules A}} \right) \left( \frac{\text{mol} B}{1 \text{mol} A} \right) = g_B$

**Step 3 (3 conversions).** 84.0gN$_2$ (28.0gN$_2$) $\left( \frac{2 \text{mol} NH_3}{1 \text{mol} N_2} \right) \left( \frac{17.0gNH_3}{1 \text{mol} NH_3} \right) = 102gNH_3$

**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$_2$O$_3$(s) can be made from 25.0g Al(s) and 20.0g O$_2$(g) using: 4Al(s) + 3O$_2$(g) → 2Al$_2$O$_3$(s)?

  b. Which reactant is the limiting reagent?
  
  c. Which reactant 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 each requiring 3 steps/conversions, and then comparing the two possible amounts of Al$_2$O$_3$(s) produced. The amount of excess reagent left over will require a third “grams A → grams B” calculation requiring 3 steps/conversions.

**Step 1 calculation**

\[
\begin{align*}
4 \text{Al} + 3 \text{O}_2 & \rightarrow 2 \text{Al}_2\text{O}_3 \\
25.0 \text{g Al} & \text{ (1 mol Al/27.0 g Al) } (2 \text{ mol Al}_2\text{O}_3/4 \text{ mol Al} ) (102 \text{ g Al}_2\text{O}_3/1 \text{ mol Al}_2\text{O}_3) = 47.2 \text{ g Al}_2\text{O}_3
\end{align*}
\]

**Step 2 calculation**

\[
\begin{align*}
20.0 \text{ g O}_2 & \text{ (1 mol O}_2/32.0 \text{ g O}_2) \text{ (2 mol Al}_2\text{O}_3/3 \text{ mol O}_2) \text{ (102 g Al}_2\text{O}_3/1 \text{ mol Al}_2\text{O}_3) = 42.5 \text{ g Al}_2\text{O}_3
\end{align*}
\]

a. 42.5g Al$_2$O$_3$ (the smaller quantity of Al$_2$O$_3$(s) is how much can theoretically be produced)

b. O$_2$(g) is the LR (since it produced the smaller amount Al$_2$O$_3$(s))

c. Al(s) is the excess reagent (since it produced the larger amount Al$_2$O$_3$(s))

d. **Amount Excess Reagent Left Over = Starting Amount Excess Reagent – Amount Excess Reagent Used**; amount used is calculated with “grams A → grams B” with the calculation going from the LR to the Excess reagent:

**Step 3 calculation**

\[
\begin{align*}
20.0 \text{ g O}_2 & \text{ (1 mol O}_2/32.0 \text{ g O}_2) \text{ (4 mol Al/3 mol O}_2) \text{ (27.0 g Al/1 mol Al) = 22.5 g Al}
\end{align*}
\]

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
PERCENT YIELD = \( \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \)

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 CO₂ and H₂O (combustion)**

A. Compound contains C and H only
   1. Convert \( g\text{CO}_2 \rightarrow \text{mol CO}_2 \rightarrow \text{mol C} \)
   2. Convert \( g\text{H}_2\text{O} \rightarrow \text{mol H}_2\text{O} \rightarrow \text{mol H} \)
   3. Write formula and divide by smallest moles
   4. If needed, fractions: \( \frac{1}{2} (0.5) \rightarrow x 2; \frac{1}{3} or \frac{2}{3} (0.33, 0.66) \rightarrow x 3; \frac{1}{4} or \frac{3}{4} (0.25, 0.75) \rightarrow x 4 \)

B. Compound contains C, H, and O
   1. Convert \( g\text{CO}_2 \rightarrow \text{mol CO}_2 \rightarrow \text{mol C} \rightarrow g\text{C} \) (need both mol C and gC)
   2. Convert \( g\text{H}_2\text{O} \rightarrow \text{mol H}_2\text{O} \rightarrow \text{mol H} \rightarrow g\text{H} \) (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 \rightarrow \text{mol O} \)
   5. Write formula and divide by smallest moles
   6. If needed, fractions: \( \frac{1}{2} (0.5) \rightarrow x 2; \frac{1}{3} or \frac{2}{3} (0.33, 0.66) \rightarrow x 3; \frac{1}{4} or \frac{3}{4} (0.25, 0.75) \rightarrow x 4 \)

**Example 3:** Butane, a hydrocarbon, was burned and 15.14g CO₂ and 7.751g H₂O are recovered. What is the empirical formula of butane?

**Answer 3:** 1. mol C: \( 15.14\text{gCO}_2 \times \frac{1\text{molCO}_2}{44.01\text{gCO}_2} = 0.3441\text{molC} \)
   2. mol H: \( 7.751\text{gH}_2\text{O} \times \frac{2\text{molH}}{18.02\text{gH}_2\text{O}} = 0.8603\text{molH} \)
   3. Write formula: \( \text{C}_0.3441\text{H}_{0.8603} \) and divide by smallest #mol: 0.3441: \( \text{C}_{0.3441} \text{H}_{0.8603} \rightarrow \text{C}_1 \text{H}_2 \text{.500} \)
   4. Fractions: \( \text{C}_1 \text{H}_2 \times 2 \rightarrow \text{C}_3 \text{H}_5 \)

**Example 4:** When 2.000g of a compound containing carbon, hydrogen, and oxygen is combusted, 1.912g CO₂ and 0.7830g H₂O are recovered. What is the empirical formula?

**Answer 4:** 1. mol C: \( 1.912\text{gCO}_2 \times \frac{1\text{molCO}_2}{44.01\text{gCO}_2} = 0.04344\text{molC} \), \( g\text{C} : 0.04344\text{molC} \times \frac{12.01\text{gC}}{1\text{molC}} = 0.5217\text{gC} \)
   2. mol H: \( 0.7830\text{gH}_2\text{O} \times \frac{2\text{molH}}{18.02\text{gH}_2\text{O}} = 0.08690\text{molH} \), \( g\text{H} : 0.08690\text{molH} \times \frac{1.008\text{gH}}{1\text{molH}} = 0.08760\text{gH} \)
   3. \( g\text{O} : g\text{sample} = g\text{C} + g\text{H} + g\text{O} \rightarrow \text{solve for } g\text{O} : g\text{O} = g\text{sample} - g\text{C} - g\text{H} = 2.000 - 0.5217 - 0.08760 = 1.3907\text{gO} \)
   4. mol O: \( 1.3907\text{gO} \times \frac{1\text{molO}}{16.00\text{gO}} = 0.08692\text{molO} \)
   5. Write formula: \( \text{C}_{0.04344} \text{H}_{0.8609} \text{O}_{0.08692} \); divide by smallest #mol: 0.04344: \( \text{C}_{0.04344} \text{H}_{0.8690} \text{O}_{0.08692} \rightarrow \text{C}_1 \text{H}_2 \text{.0000} \text{O}_{2.000} \rightarrow \text{CH}_2\text{O}_2 \)

1. If 25.0g N₂ reacts with excess H₂, how many grams of NH₃ would be produced?
   \( \text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g) \)

2. Calculate the masses (in grams) of Cr₂O₃ (chromium(III) oxide), N₂, and H₂O produced from 10.8g of \( (\text{NH}_4)_2\text{Cr}_2\text{O}_7 \) (ammonium dichromate) in the following balanced reaction:
   \( (\text{NH}_4)_2\text{Cr}_2\text{O}_7(s) \rightarrow \text{Cr}_2\text{O}_3(s) + \text{N}_2(g) + 4\text{H}_2\text{O(l)} \)

3. Nitric acid, HNO₃, is manufactured by the Oswald process, in which nitrogen dioxide, NO₂, reacts with H₂O. How many grams of NO₂ are required to produce 5.00g HNO₃.
   \( 3\text{NO}_2(g) + \text{H}_2\text{O(l)} \rightarrow 2\text{HNO}_3(aq) + \text{NO(g)} \)

   a. 7.50   b. 15.0   c. 3.65   d. 0.120   e. 5.48
4. a. When 9.20g of C₂H₆O are reacted with 40.6g of PBr₃, what mass of C₂H₂Br can be produced?
   \[3C₂H₆O + PBr₃ \rightarrow 3C₂H₂Br + H₃PO₃\]
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 PBr₃ was recovered after the experiment, what was the %yield of this reaction?

5. Tungsten metal can be produced by reacting yellow tungsten oxide, WO₃, with hydrogen.
   \[WO₃(s) + 3H₂(g) \rightarrow W(s) + 3H₂O(l)\]
How many grams of tungsten can be obtained from 1.25 x 10⁵g H₂ and 6.55 x 10⁶g WO₃?

6. If 75.0g of SiO₂ and 30.0g C react according to the equation below, what is the maximum number of moles of CO that can be produced? \(SiO₂ + C \rightarrow CO + SiO\)
a. 1.25  
b. 1.67  
c. 2.25  
d. 2.50  
e. none of these

7. I. Given the following balanced combustion reaction below, if there were 3.0mol of C₄H₄ and 6.0mol of O₂, how many moles of CO₂ could theoretically be produced?
   \[C₂H₄ + 3O₂ \rightarrow 2CO₂ + 2H₂O\]
a. 6mol  
b. 4mol  
c. 9mol  
d. 3mol  
e. 7mol
II. After the above reaction was completed, 3.5mol of CO₂ 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.)

8. Consider the reaction:
   \[2H₂(g) + O₂(g) \rightarrow 2H₂O(g)\]
Identify the limiting reagent in each of the reaction mixtures below. Note: This is not a multiple choice question.  
a. 0.50mol of H₂ and 0.75mol of O₂  
b. 0.80mol of H₂ and 0.75mol of O₂  
c. 1.00g of H₂ and 0.35mol of O₂  
d. 5.00g of H₂ and 64.00g of O₂

9. When a compound containing carbon and hydrogen is combusted, 3.38g CO₂ and 0.692g H₂O are recovered.  
a. What is the empirical formula?  
b. The molar mass of the compound is 78.1g/mol. What is the molecular formula?

10. When 5.000g of a compound containing carbon, hydrogen, and oxygen is combusted, 8.910g CO₂ and 3.648g H₂O are recovered.  
a. What is the empirical formula?  
b. The molar mass of the compound is 74.1g/mol. What is the molecular formula?

11. A compound contains carbon, hydrogen, and oxygen. When 1.0000 gram of this substance is combusted, 2.3744grams CO₂ and 1.2153grams of H₂O are collected. What is the empirical formula?

12. A 1.257g sample of a compound, BₓHᵧ, is reacted in pure oxygen to form 3.163g of B₂O₃. What is the empirical formula for BₓHᵧ?
ANSWERS

1. 30.4g NH₃  \( \{25.0gN_2 \times (1molN_2/28.0gN_2) \times (2molNH_3/1molN_2) \times (17.0gNH_3/1molNH_3) = 30.36gNH_3 \} \)

2. 6.52g Cr₂O₃  \( \{10.8g(NH_4)_2Cr_2O_7 \times (1mol(NH_4)_2Cr_2O_7/252.08g(NH_4)_2Cr_2O_7) \times (1molCr_2O_3/1mol(NH_4)_2Cr_2O_7) \times (152.00gCr_2O_3/1molCr_2O_3) = 65.12gCr_2O_3 \} \)

3. 2.65g Cr  \( \{6.52gCr_2O_3 \times (1molCr_2O_3/211.58gCr_2O_3) \times (2molCr/1molCr_2O_3) \times (52.00gCr/1molCr) = 21.76gCr \} \)

5. 3.80 x 10^6g W  \( \{1.25 x 10^5gH_2 \times (1molH_2/2.0166gH_2) \times (1molW/3molH_2) \times (183.9gW/1molW) = 3.801 x 10^6g W; \}

6. 3.09g H₂O  \( \{10.8g(NH_4)_2Cr_2O_7 \times (1mol(NH_4)_2Cr_2O_7/252.08g(NH_4)_2Cr_2O_7) \times (4molH_2O/1mol(NH_4)_2Cr_2O_7) \times (18.02gH_2O/1molH_2O) = 3.088gH_2O \} \)

7. II. c  \( \{\%yield = (actualyield/theoreticalyield) \times 100\%; \%yield = (10.9gC_2H_5Br/21.8gC_2H_5Br) \times 100\% = 50.0\% \} \)

8. a. H₂  \( \{3.38gCO_2 \times (1molCO_2/44.01gCO_2) \times (1molC/1molCO_2) = 0.07680molC; \}

9. a. CH  \( \{3.89gCO_2 \times (1molCO_2/44.01gCO_2) \times (1molC/1molCO_2) = 0.07680molC; \}

10. a. C₃H₆O₂  \( \{8.910gCO_2 \times (1molCO_2/44.01gCO_2) \times (1molC/1molCO_2) = 0.2025molC; \}

b. C₂H₆O₂  \( \{molarmass_{MF}/molarmass_{EF} = 74.1/74.08 = 1.000; C_3H_6O_2 \times 1 = C_3H_6O_2 \} \)

b. C₂H₆O₂  \( \{molarmass_{MF}/molarmass_{EF} = 74.1/74.08 = 1.000; C_3H_6O_2 \times 1 = C_3H_6O_2 \} \)
11. **C₄H₁₀O**  
\[
2.3744 \text{g CO}_2 \times \left(\frac{1 \text{mol CO}_2}{44.01 \text{g CO}_2}\right) \times \left(\frac{1 \text{mol C}}{1 \text{mol CO}_2}\right) = 0.053951 \text{mol C};
\]
\[
g_\text{C} = 0.053951 \text{mol C} \times (12.01 \text{g C}/1 \text{mol C}) = 0.64795 \text{g C};
\]
\[
1.2153 \text{g H}_2\text{O} \times \left(\frac{1 \text{mol H}_2\text{O}}{18.02 \text{g H}_2\text{O}}\right) \times \left(\frac{2 \text{mol H}}{1 \text{mol H}_2\text{O}}\right) = 0.13488 \text{mol H};
\]
\[
g_\text{H} = 0.13488 \text{mol H} \times (1.008 \text{g H}/1 \text{mol H}) = 0.13596 \text{g H};
\]
\[
g_\text{sample} = g_\text{O} + g_\text{C} + g_\text{H} \rightarrow \text{solve for } g_\text{O}:
\]
\[
g_\text{O} = g_\text{sample} - g_\text{C} - g_\text{H} = 1.0000 - 0.64795 - 0.13596 = 0.21609 \text{g O};
\]
\[
\text{Mol O} = 0.21609 \text{g O} \times \left(\frac{1 \text{mol O}}{16.00 \text{g O}}\right) = 0.013506 \text{mol O};
\]
\[
\text{C}_0.053951\text{H}_{0.13488}\text{O}_{0.013506}; \text{ divide by 0.013506; } = \text{C}_3.9946\text{H}_{9.9867}\text{O}_1 = \text{C}_4\text{H}_{10}\text{O}_1 = \text{EF}
\]

12. **BH₃**  
\[
\text{find mol B: } 3.163 \text{g B}_2\text{O}_3 \times \left(\frac{1 \text{mol B}_2\text{O}_3}{69.62 \text{g B}_2\text{O}_3}\right) \times \left(\frac{2 \text{mol B}}{1 \text{mol B}_2\text{O}_3}\right) = 0.09086 \text{mol B};
\]
\[
g_\text{B} = 0.09086 \text{mol B} \times (10.81 \text{g B}/1 \text{mol B}) = 0.9822 \text{g B};
\]
\[
g_\text{sample} = g_\text{B} + g_\text{H};
\]
\[
g_\text{H} = 1.257 - 0.9822 = 0.2748 \text{g H};
\]
\[
\text{find mol H: } 0.2748 \text{g H} \times \left(\frac{1 \text{mol H}}{1.008 \text{g H}}\right) = 0.2726 \text{mol H};
\]
\[
\text{write formula: } \text{B}_{0.09086}\text{H}_{0.2726} \text{ and divide by smallest number of moles: } \text{B}_1\text{H}_{3.024} = \text{BH}_3\}