Announcements for Monday, 28OCT2024

- Wednesday’s lecture will be a review
- Week 8 Homework Assignments available on eLearning
  - Graded and Timed Quiz 8 – “Bonding + review” due **tonight at 6:00 PM (EDT)**
- Mid-Semester Survey due **Monday, 04NOV2024, at 11:59 PM (EST)**
- Exam II is Wednesday, 30OCT2024, 7:45-9:05 **PM** (EDT)
  - Coverage: Chapters 3.6-6.5; exam consists of 19 multiple-choice questions and open-ended questions; see “Other Resources” on Canvas for periodic table and formula sheet to be used on the exam
  - Exam 2 Locations are posted on Canvas
- Exam II Calculator Policy
  - Scientific calculators and ***most*** graphing calculators are allowed
  - **TI-Nspire CX series & other calculators with *QWERTY* keyboards are NOT allowed**

ANY GENERAL QUESTIONS? Feel free to see me after class!

Try These On Your Own

Write **balanced** chemical equations for the following reactions:

Liquid dichlorine heptoxide is added to water to form liquid perchloric acid ($HClO_4$).

$$Cl_2O_7(\ell) + H_2O(\ell) \rightarrow 2\ HClO_4(\ell)$$

When aqueous solutions of aluminum sulfate and sodium hydroxide are mixed, aqueous sodium sulfate and solid aluminum hydroxide are formed.

$$Al_2(SO_4)_3(aq) + 6\ NaOH(aq) \rightarrow 3\ Na_2SO_4(aq) + 2\ Al(OH)_3(s)$$

Solid aluminum carbide (Al_4C_3), when mixed with water, produces solid aluminum hydroxide and methane (CH_4) gas.

$$Al_4C_3(s) + 12\ H_2O(\ell) \rightarrow 4\ Al(OH)_3(s) + 3\ CH_4(g)$$

Rust (i.e., iron(III) oxide) forms when iron is exposed to oxygen in the air.

$$4\ Fe(s) + 3\ O_2(g) \rightarrow 2\ Fe_2O_3(s)$$

Complete combustion of liquid propanol ($CH_3CH_2CH_2OH$) yields carbon dioxide and water vapor.

$$2\ CH_3CH_2CH_2OH(\ell) + 9\ O_2(g) \rightarrow 6\ CO_2(g) + 8\ H_2O(g)$$

Try This On Your Own

Ammonium nitrite (NH_4NO_2) is an unstable **solid** that readily decomposes into nitrogen **gas** and water **vapor**. How many molecules of **gas** will be produced by the decomposition of 100. g of ammonium nitrite? **2.82×10^{24}**

$$NH_4NO_2(s) \longrightarrow N_2(g) + 2\,H_2O(g)$$

from the balanced equation: for every 1 mol of **solid** decomposing, 3 mol of **gas** are produced (1 mol N_2 + 2 mol H_2O vapor)

starting units: **mass** $NH_4NO_2(s)$ → → → ending units: **number of molecules of gas**

$$100.\,\mathrm{g\;NH_4NO_2} \times \frac{1\,\mathrm{mol\;NH_4NO_2}}{64.05\,\mathrm{g}} \times \frac{3\,\mathrm{mol\;gas}}{1\,\mathrm{mol\;NH_4NO_2}} \times \frac{6.022 \times 10^{23}\,\mathrm{gas\;molecules}}{1\,\mathrm{mol\;gas}} = 2.82 \times 10^{24}\,\mathrm{molecules}$$

molar mass NH_4NO_2 **from balanced equation** **Avogadro’s number**

Reaction Stoichiometry – Try These On Your Own

Electrolysis of water leads to the formation of hydrogen and oxygen gases according to the reaction $H_2O(\ell) \rightarrow H_2(g) + O_2(g)$. How many *moles* of water must be electrolyzed to generate 0.231 *moles* of oxygen gas?

0.462 mol H_2O (ℓ)

What mass of oxygen is needed to completely combust 5.00 *moles* hexane (C_6H_{14})?

1.52 kg $O_2(g)$

Consider the reaction $2\ N_2(g) + 5\ O_2(g) \rightarrow 2\ N_2O_5(g)$. What masses of nitrogen gas and oxygen gas are needed to produce 100.0 g N_2O_5? MM N_2O_5 = 108.02 g/mol

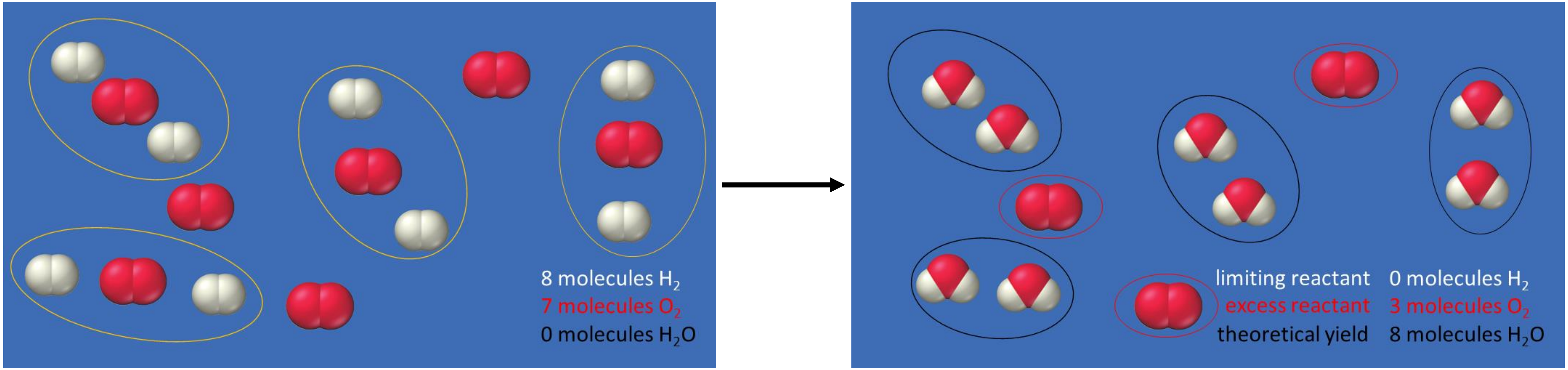
25.94 g $N_2(g)$ and 74.06 g $O_2(g)$

The complete decomposition of ClF_3 into Cl_2 and F_2 produced 142 g $Cl_2(g)$. <u>How many molecules</u> of fluorine gas was produced in the same reaction?

3.62×10^{24} molecules $F_2(g)$

Limiting Reactant Illustrated

$$2\ H_2(g) + 1\ O_2(g) \rightarrow 2\ H_2O(\ell)$$



$$8 \text{ molecules } H_2 \times \frac{2 \text{ molecules } H_2O}{2 \text{ molecules } H_2} = 8 \text{ molecules } H_2O \qquad \text{vs.} \qquad 7 \text{ molecules } O_2 \times \frac{2 \text{ molecules } H_2O}{1 \text{ molecule } O_2} = 14 \text{ molecules } H_2O$$

H_2 produces less H_2O than O_2, so **H_2 is limiting** and **8 molecules H_2O is theoretical yield**

$$7 \text{ molecules } O_2 \times \frac{2 \text{ molecules } H_2}{1 \text{ molecule } O_2} = 14 \text{ molecules } H_2 \text{ required to react with } O_2\text{; only 8 molecules } H_2 \text{ available so } \mathbf{H_2 \text{ is limiting}}$$

$$8 \text{ molecules } H_2 \times \frac{1 \text{ molecule } O_2}{2 \text{ molecules } H_2} = 4 \text{ molecules } O_2 \text{ required to react with } H_2\text{; 7 molecules } O_2 \text{ available so } \mathbf{H_2 \text{ is limiting}}$$

7 molecules O_2 available − 4 molecules O_2 required = **3 molecules O_2 in excess**

Limiting Reactant (aka Limiting Reagent)

limiting reactant = the ***reactant*** whose amount causes it to get completely consumed in a chemical reaction

- there can be more than one limiting reactant in a given reaction

★ the limiting reactant dictates how much product(s) can be produced

- the limiting reactant also dictates how much of the other reactant(s) gets used up
  - reactants that don't get completely consumed and have some left over at the end of the reaction are called "excess reactants/reagents"

- unless reactants are stated in a problem to be "in excess", you must do calculations to identify the limiting reactant(s) and use the amount of limiting reactant as the starting point for all subsequent stoichiometric calculations
  - this needs to be an automatic step in your procedure

Approaches for Determining Limiting Reactant

Consider the reaction 2 Mg(s) + O_2(g) → 2 MgO(s)

42.5 g Mg and 33.8 g O_2 are combined and allowed to react.

What mass of product is formed? What mass of excess reactant remains?

ALWAYS start by converting masses of reactants to moles

Approach 1:

a) separately determine the amount of product possible from each reactant (if more than one product possible, choose only one and use it for all calculations)

b) whichever reactant gives the least amount of product is limiting and the amount of product that was calculated is the theoretical yield

- works best for reactions that have more than two reactants and when you're asked to calculate amounts of products

Approach 2:

a) determine the amount of one reactant needed to completely react with the other reactant

b) compare the calculated amount to the amount actually stated in the problem

- works best for only two reactants and when you're asked to calculate the amount of excess reagent

Approaches for Determining Limiting Reactant

Consider the reaction 2 Mg(s) + O_2(g) $\rightarrow$ 2 MgO(s). 42.5 g Mg and 33.8 g O_2 are combined and allowed to react. **What mass of product is formed?** **What mass of excess reactant remains?**

first ALWAYS convert masses of reactants to moles

$$42.5\text{ g Mg} \times \frac{1\text{ mol Mg}}{24.31\text{ g}} = 1.748\text{ mol Mg} \quad \text{and} \quad 33.8\text{ g O}_2 \times \frac{1\text{ mol O}_2}{32.00\text{ g}} = 1.056\text{ mol O}_2$$

Approach 1: **separately determine the amount of product possible from each reactant**

$$1.748\text{ mol Mg} \times \frac{2\text{ mol MgO}}{2\text{ mol Mg}} = 1.748\text{ mol MgO} \quad \text{and} \quad 1.056\text{ mol O}_2 \times \frac{2\text{ mol MgO}}{1\text{ mol O}_2} = 2.112\text{ mol MgO}$$

since Mg makes less product, Mg is limiting and O_2 is in excess. 1.748 mol MgO is the theoretical yield of product.

$$1.748\text{ mol MgO} \times \frac{40.31\text{ g}}{1\text{ mol MgO}} = 70.5\text{ g MgO formed}$$

Approach 2: **determine the amount of one reactant needed to completely react with the other reactant**

$$1.056\text{ mol O}_2 \times \frac{2\text{ mol Mg}}{1\text{ mol O}_2} = 2.112\text{ mol Mg needed but we only have } 1.748\text{ mol Mg (Mg is limiting, O}_2\text{ in excess)}$$

OR

$$1.748\text{ mol Mg} \times \frac{1\text{ mol O}_2}{2\text{ mol Mg}} = 0.874\text{ mol O}_2\text{ needed and we have } 1.056\text{ mol O}_2\text{ (Mg is limiting, O}_2\text{ in excess)}$$

$$1.056\text{ mol O}_2\text{ available} - 0.874\text{ mol O}_2\text{ required} = 0.182\text{ mol O}_2\text{ left over} \times \frac{32.00\text{ g}}{1\text{ mol O}_2} = 5.82\text{ g O}_2\text{ in excess}$$

Percent Yield of a Reaction

- **theoretical yield =** the amount or mass of product(s) formed from the complete consumption of the limiting reactant
  - it is the best-case scenario based on stoichiometric calculations (i.e., in theory rather than in reality)
- **actual yield =** the amount or mass of products ***actually*** formed in a given reaction
  - cannot be greater than the theoretical yield
    - you can't get more than 100% of product
  - often times lower than the theoretical yield
    - can be due to many reasons (e.g., side reactions, experimental error, etc.)
- $percent\ yield\ of\ a\ reaction = \dfrac{actual\ yield}{theoretical\ yield} \times 100\%$

Some Reactions to Know

1. combustion reactions
   - reaction with O_2 to form one or more oxygen-containing compounds
   - $CH_4 + 2\ O_2(g) \longrightarrow CO_2 + 2\ H_2O$
   - very exothermic
2. alkali metals: Group 1A metals
   - react with nonmetals (X_2) to form ionic compounds ($2\ M(s) + X_2 \longrightarrow 2\ MX(s)$)
   - react with water to generate hydrogen gas ($2\ M(s) + 2\ H_2O(\ell) \longrightarrow 2\ MOH(aq) + H_2(g)$)
     - exothermic and explosive
3. halogens: Group 7A nonmetals
   - react with metals to form ionic compounds ($2\ M + n\ X_2(g) \longrightarrow 2\ MX_n(s)$)
   - react with H_2 to form hydrogen halides ($H_2(g) + X_2 \longrightarrow 2\ HX(g)$)
   - react with other halogens to form interhalogen compounds ($Br_2(\ell) + F_2(g) \longrightarrow 2\ BrF(g)$)

STUDENT NOTE: Cover on your own…know the different products that result

Try This

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

4.00 mol N_2 and 15.0 mol H_2 react with 65.0% yield in a sealed vessel. What mass of NH_3(molar mass = 17.03 g/mol) is produced?

limiting reactant

$$4.00\ \text{mol N}_2 \times \frac{2\ \text{mol NH}_3}{1\ \text{mol N}_2} \times \frac{17.03\ \text{g NH}_3}{1\ \text{mol NH}_3} = 136.2\ \text{g NH}_3$$

theoretical yield

$$15.0\ \text{mol H}_2 \times \frac{2\ \text{mol NH}_3}{3\ \text{mol H}_2} \times \frac{17.03\ \text{g NH}_3}{1\ \text{mol NH}_3} = 170.3\ \text{g NH}_3$$

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

$$65.0\% = \frac{\text{actual yield}}{136.2\ \text{g}} \times 100\%$$

$$\text{actual yield} = 0.650 \times 136.2\ \text{g} = \mathbf{88.5\ g}$$

Try This On Your Own

When heated, calcium carbonate can decompose into calcium oxide and carbon dioxide according to the balanced equation

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$$

What mass of $CaCO_3$ ($\mathcal{M}$ = 100.09 g/mol) must undergo decomposition to generate 10.0 g CO_2 ($\mathcal{M}$ = 44.01 g/mol) if the above reaction only has a yield of 75.0 %?

More Reaction Stoichiometry – Try These On Your Own

Consider the reaction $Cl_2(g) + 3\ F_2(g) \rightarrow 2\ ClF_3(g)$. What mass of $ClF_3(g)$ was produced from the reaction of 80.0 g $Cl_2(g)$ with 106 g $F_2(g)$? Assume 100% yield.

Consider the reaction $2\ N_2(g) + 5\ O_2(g) \rightarrow 2\ N_2O_5(g)$ in which 20.0 g N_2 is combined with 50.0 g O_2. What mass of excess reagent is left over once the reaction goes to completion?

10. g H_2 reacted with 40. g O_2 to produce water in 80.% yield. How many moles of water was actually produced?