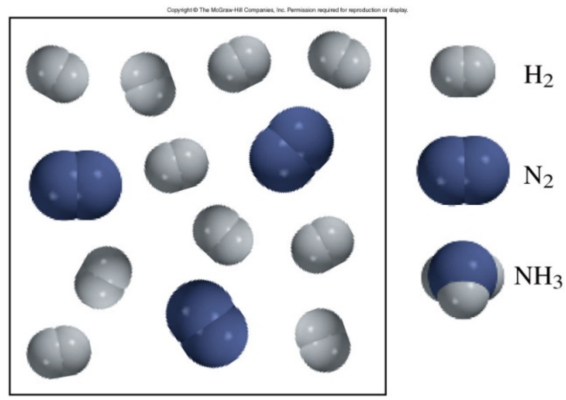
Chem 2 AP **Homework #3-5:** Problems pg. 109-110 #3.79, 3.82, 3.83, 3.84, 3.88, 3.92, 3.94, 3.108

79 **Define limiting reagent and excess reagent**.
Limiting reagent is the reactant that is depleted; excess reagent is a reagent that is not depleted.

What is the significance of the limiting reagent in predicting the amount of the product obtained in a reaction? In a reaction, since the reaction stops when one reactant is depleted, the amount of limiting reagent determines the amount of each product possible.

Can there be a limiting reagent if only one reactant is present? If there is only one reactant, then technically it must be the limiting reagent since it will be depleted.

82 **$N_2 + 3H_2 \rightarrow 2NH_3$ From diagram, one sees that you have 3 moles of N_2 and 10 moles of H_2**



Finding limiting reagent: The number of N_2 molecules shown in the diagram is 3. The balanced equation shows 3 moles $H_2 \bumpeq$ 1 mole N_2. Therefore, we need 9 molecules of H_2 to react completely with 3 molecules of N_2. There are 10 molecules of H_2 present in the diagram. H_2 is in excess.

N_2 is the limiting reagent.

Number of moles of product produced: The mole ratio between N_2 and NH_3 is 1:2. When 3 molecules of N_2 react, 6 molecules of NH_3 will be produced.

Number of moles of excess reagent left: 9 molecules of H_2 will react with 3 molecules of N_2, leaving 1 molecule of H_2 in excess.

83 **Nitric oxide reacts with oxygen gas to form nitrogen dioxide, a dark-brown gas:**

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

In one experiment 0.886 mole of NO is mixed with 0.503 mole of O_2. Calculate which of the two reactants is the limiting reagent. Calculate also the number of moles of NO_2 produced.

$$0.886 \text{ mol NO} \times \frac{2 \text{ mol NO}_2}{2 \text{ mol NO}} = 0.886 \text{ mol NO}_2$$

$$0.503 \text{ mol O}_2 \times \frac{2 \text{ mol NO}_2}{1 \text{ mol O}_2} = 1.01 \text{ mol NO}_2$$

NO is the **limiting reagent**; it limits the amount of product produced. The amount of product produced is **0.886 mole NO_2**.

84 The depletion of the ozone in the stratosphere has been a matter of great concern among scientists in recent years. $O_3 + NO \rightarrow O_2 + NO_2$ If 0.740 g of O_3 reacts with 0.679 g of NO…

- **How many grams of NO_2 produced?**

$$\text{equiv. } O_3 = \frac{0.740 \text{ g } O_3 \times \dfrac{1 \text{ mol } O_3}{48.00 \text{ g } O_3}}{1 \text{ mol } O_3} = 0.0154 \text{ equiv. } O_3$$

$$\text{equiv. NO} = \frac{0.670 \text{ g NO} \times \dfrac{1 \text{ mol NO}}{30.01 \text{ g NO}}}{1 \text{ mol NO}} = 0.0223 \text{ equiv. NO}$$

The equiv. of $\mathbf{O_3}$ < equiv. **NO**; therefore, it is the **limiting reagent**.

$$\mathbf{?\ g\ NO_2} = 0.0154 \text{ mol } O_3 \times \frac{1 \text{ mol } NO_2}{1 \text{ mol } O_3} \times \frac{46.01 \text{ g } NO_2}{1 \text{ mol } NO_2} = \mathbf{0.709\ g\ NO_2}$$

- **How many moles of excess reagent remain?**

$$\text{mol NO reacted} = 0.0154 \text{ mol } O_3 \times \frac{1 \text{ mol NO}}{1 \text{ mol } O_3} = 0.0154 \text{ mol NO reacted}$$

$$\text{mol NO initial} = 0.670 \text{ g NO} \times \frac{1 \text{ mol NO}}{30.01 \text{ g NO}} = 0.0223 \text{ mol NO}$$

$$\text{mol NO remaining} = \text{mol NO initial} - \text{mol NO reacted.}$$

$$\mathbf{mol\ NO\ remaining} = 0.0223 \text{ mol NO} - 0.0154 \text{ mol NO} = \mathbf{0.0069\ mol\ NO}$$

88 Why is the actual yield of a reaction almost always smaller than the theoretical yield?
The actual yield of a reaction is usually less than the theoretical yield because not all reactants form products (reversible reactions or side products), and not all product may be recovered.

92 Ethylene, can be prepared by heating hexane at 800°C: $C_6H_6 \rightarrow C_2H_4$ + other products If the yield of ethylene production is 42.5 percent, what mass of hexane must be reacted to produce 481 g of ethylene?

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

$$42.5\%\ yield = \frac{481 \text{ g } C_2H_4}{\text{theoretical yield}} \times 100\%$$

$$\text{theoretical yield } C_2H_4 = 1.13 \times 10^3 \text{ g } C_2H_4$$

The mass of hexane that must be reacted is:

$$(1.13 \times 10^3 \text{ g } C_2H_4) \times \frac{1 \text{ mol } C_2H_4}{28.05 \text{ g } C_2H_4} \times \frac{1 \text{ mol } C_6H_{14}}{1 \text{ mol } C_2H_4} \times \frac{86.15 \text{ g } C_6H_{14}}{1 \text{ mol } C_6H_{14}} = \mathbf{3.47 \times 10^3\ g\ C_6H_{14}}$$

94 Disulfide dichloride is used in the vulcanization of rubber. It is prepared by heating sulfur in an atmosphere of chlorine: $S_8(l) + 4Cl_2(g) \rightarrow 4S_2Cl_2(l)$

What is the theoretical yield of S_2Cl_2 when 4.06 g of S_8 are heated with 6.24g of Cl_2

$$4.06 \text{ g } S_8 \times \frac{1 \text{ mol } S_8}{256.6 \text{ g } S_8} \times \frac{4 \text{ mol } S_2Cl_2}{1 \text{ mol } S_8} = 0.0633 \text{ mol } S_2Cl_2$$

$$6.24 \text{ g } Cl_2 \times \frac{1 \text{ mol } Cl_2}{70.90 \text{ g } Cl_2} \times \frac{4 \text{ mol } S_2Cl_2}{4 \text{ mol } Cl_2} = 0.0880 \text{ mol } S_2Cl_2$$

$$? \text{ g } S_2Cl_2 = 0.0633 \text{ mol } S_2Cl_2 \times \frac{135.04 \text{ g } S_2Cl_2}{1 \text{ mol } S_2Cl_2} = \mathbf{8.55 \text{ g } S_2Cl_2}$$

If the actual yield of S_2Cl_2 is 6.55 g, what is the percent yield?

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{6.55\ g}{8.55\ g} \times 100\% = \mathbf{76.6\%}$$

108 $Fe_2O_3 + 3\ CO \rightarrow 2\ Fe + 3\ CO_2$

One obtains $1.64 \times 10^3\ kg\ Fe$ **from a 2.62 x 10^3 kg sample of Fe_2O_3. What is % purity of sample?**

ONE METHOD: Find mass of pure Fe_2O_3 needed to get amount of Fe obtained.

$$1.64 \times 10^3 \text{ kg Fe} \times \frac{1 \text{ kg-mol Fe}}{55.85 \text{ kg Fe}} \times \frac{1 \text{ kg-mol } Fe_2O_3}{2 \text{ kg-mol Fe}} \times \frac{159.7 \text{ kg } Fe_2O_3}{1 \text{ kg-mol } Fe_2O_3} = 2.34 \times 10^3 \text{ kg } Fe_2O_3(\text{pure})$$

$$\text{percent purity of sample} = \frac{\text{grams of pure } Fe_2O_3}{\text{grams of } Fe_2O_3 \text{ sample}} \times 100\%$$

$$\textbf{percent purity} = \frac{2.34 \times 10^3 \text{ kg pure } Fe_2O_3}{2.62 \times 10^3 \text{ kg sample of } Fe_2O_3} \times 100\% = \mathbf{89.6\% = purity\ of\ Fe_2O_3}$$

ANOTHER METHOD: Calculate kg of Fe that theoretically could be produced if the Fe_2O_3 sample was pure.

Theoretical yield of Fe =

$$(2.62 \times 10^3 \text{ kg } Fe_2O_3) \times \frac{1 \text{ kg-mol } Fe_2O_3}{159.7 \text{ kg } Fe_2O_3} \times \frac{2 \text{ kg-mol Fe}}{1 \text{ kg-mol } Fe_2O_3} \times \frac{55.85 \text{ kg Fe}}{1 \text{ kg-mol Fe}}$$
$$= 1.83 \times 10^3 \text{ kg Fe}$$

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

$$\textbf{percent yield} = \frac{1.64 \times 10^3\ kg\ Fe}{1.83 \times 10^3\ kg\ Fe} \times 100\% = \mathbf{89.6\% = purity\ of\ Fe_2O_3}$$