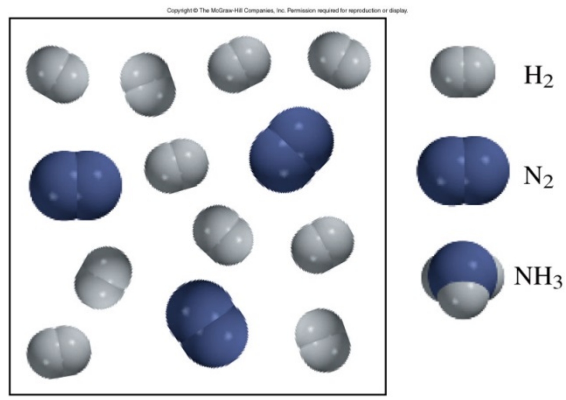
WKS 3-5 – Limiting & Excess Reagents; Reaction Yields **Name:__________________________**

3.79 **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.

3.82 **$N_2 + 3H_2 \rightarrow 2NH_3$ From diagram, one sees that there are 3 moles of N_2 and 10 moles of H_2**



Fill in the "ICF" table to help answer all questions:

	N_2	+	$3\ H_2$	$\rightarrow$	$2\ NH_3$
Initial	**3 mc**		**10 mc**		**0 mc**
Change	−3 mc		−9 mc		+6 mc
Final	0 mc		1 mc		6 mc

What is the limiting reagent? Explain. There are 3 molecules of N_2 in the diagram. We need 9 molecules of H_2 to react completely with 3 molecules of N_2. Since there are 10 molecules of H_2, H_2 is in excess.

Number of moles of product produced: 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.

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

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

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.

$$\frac{0.886 \text{ mol NO}}{2} = 0.443 \text{ mol Rxn}; \quad \frac{0.503 \text{ mol } O_2}{1} = 0.503 \text{ mol Rxn}; \quad \textbf{NO} \text{ is the } \textbf{limiting reagent}$$

$$\text{mol } NO_2 = 0.443 \text{ mol Rxn} \times \frac{2 \text{ mol } NO_2}{1 \text{ mol Rxn}} = \boxed{\mathbf{0.886\ mol\ NO_2}}$$

Here is an ICF table, to see the information in an organized fashion:

	2 NO	+	**O_2**	$\rightarrow$	**2 NO_2**
Initial	**0.886 mol**		**0.503 mol**		**0 mol**
Change	−0.886 mol		−0.443 mol		+0.886 mol
Final	0 mol		0.060 mol		0.886 mol

3.88 **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.

3.84 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?

$$\frac{0.740\text{ g }O_3 \times \frac{1\text{ mol }O_3}{48.00\text{ g }O_3}}{1} = 0.0154\text{ mol Rxn}; \quad \frac{0.679\text{ g NO} \times \frac{1\text{ mol NO}}{30.01\text{ g NO}}}{1} = 0.0226\text{ mol Rxn}$$

Therefore **O_3** is the **limiting reagent**.

$$\textbf{? g NO}_2 = 0.0154\text{ mol Rxn} \times \frac{1\text{mol NO}_2}{1\text{ mol Rxn}} \times \frac{46.01\text{ g NO}_2}{1\text{ mol NO}_2} = \boxed{\textbf{0.709 g NO}_2}$$

How many moles of excess reagent remain?

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

$$0.679\text{ g NO} \times \frac{1\text{ mol NO}}{30.01\text{ g NO}} = \boxed{0.0226\text{ mol NO}}\text{ initially present}$$

$$\textbf{mol NO remaining} = 0.0226\text{ mol NO} - 0.0154\text{ mol NO} = \textbf{0.0072 mol NO}$$

3.92 Ethylene (C_2H_4), can be prepared by heating hexane (C_6H_{14}) at 800°C: $C_6H_{14} \rightarrow C_2H_4$ + 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?

$$42.5\%\ \textit{yield} = \frac{481\text{ g }C_2H_4}{\text{theoretical yield}} \times 100\%$$; theoretical yield C_2H_4 = 1.13×10^3 g C_2H_4

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

3.94 Disulfur dichloride is used in the vulcanization of rubber. It is prepared by heating sulfur in an atmosphere of chlorine. What is the theoretical yield of S_2Cl_2 when 4.06 g of S_8 is heated with 6.24g of Cl_2 $S_8(l) + 4Cl_2(g) \rightarrow 4S_2Cl_2(l)$

$$\frac{4.06\text{ g }S_8 \times \frac{1\text{ mol }S_8}{256.6\text{ g }S_8}}{1} = \textbf{0.0158 mol Rxn}; \quad \frac{6.24\text{ g }Cl_2 \times \frac{1\text{ mol }Cl_2}{70.90\text{ g }Cl_2}}{4} = 0.0220\text{ mol Rxn}$$

$$?\text{ g }S_2Cl_2 = 0.0158\text{ mol Rxn} \times \frac{4\text{ mol }S_2Cl_2}{1\text{ mol Rxn}} \times \frac{135.04\text{ g }S_2Cl_2}{1\text{ mol }S_2Cl_2} = \boxed{\textbf{8.55 g S}_2\textbf{Cl}_2}$$

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

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

3.108 $Fe_2O_3 + 3\,CO \rightarrow 2\,Fe + 3\,CO_2$ One obtains 1.64×10^3 kg Fe from 2.62×10^3 kg of an impure sample of Fe_2O_3. What is % purity of sample? [Hint: Find the mass of pure Fe_2O_3 needed to produce the amount of Fe obtained]

$$1.64 \times 10^3\text{ kg Fe} \times \frac{1\text{kg-mol Fe}}{55.85\text{kg Fe}} \times \frac{1\text{kg-mol Fe}_2O_3}{2\text{kg-mol Fe}} \times \frac{159.7\text{kg Fe}_2O_3}{1\text{kg-mol Fe}_2O_3} = 2.34 \times 10^3\text{ kg Fe}_2O_3\text{ (pure)}$$

$$\textbf{percent purity Fe}_2\textbf{O}_3 = \frac{2.34 \times 10^3\text{ kg pure Fe}_2O_3}{2.62 \times 10^3\text{ kg impure Fe}_2O_3} \times 100\% = \boxed{\textbf{89.6\%}}$$