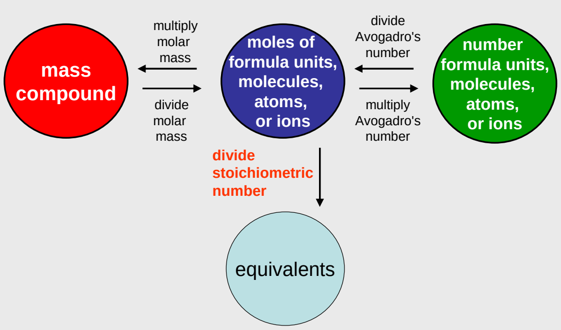
Lecture 17

Professor Hicks

General Chemistry (CHE131)

limiting reactant

- reactant that will run out first
- causes reaction to **STOP**
- LR is the reactant that is present in the smallest number of **equivalents**
- also called limiting reagent



equivalents are used to calculate the limiting reactant

equivalents are like sets

Ferrari factory

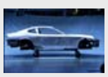
   → 

assembly process	1 car body +	4 wheels +	2 seats →	1 car
parts on hand in factory	1000 car bodies	4100 wheels	1900 seats	how many cars can be built?
***sets* on hand in factory**	$\frac{1000}{1}$ =1000	$\frac{4100}{4}$ = 1025	$\frac{1900}{2}$ = **950**	**950**

cars that can be built = smallest # **sets** of parts

when building the cars <u>sets of seats</u> will run out first

limiting reagent / equivalents

   → 

1 car body +	4 wheels +	2 seats →	1 car
1 equivalent of car bodies	**1 equivalent of wheels**	**1 equivalent of seats**	**1 equivalent of cars**

like

---

1 equivalent of N_2 (molecular nitrogen)		**1 equivalent of H_2 (molecular hydrogen)**		**1 equivalent NH_3 (ammonia)**
N_2	+	$3\ H_2$	→	$2\ NH_3$

1 equivalent of N_2 = 1 mole N_2
1 equivalent of NH_3 = 2 moles NH_3
1 equivalent of H_2 = 3 moles H_2
} for this reaction!

Q. why didn't we worry about limiting reactant here?

Example: How many moles of NH_3 will form **if 0.78 moles of H_2 reacts** in the reaction:

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

A. this means N_2 will not run out or limit the reaction→ H_2 is the LR

$$0.78\ \text{moles}\ H_2 \times \frac{2\ \text{moles}\ NH_3}{3\ \text{moles}\ H_2} = 0.52\ \text{moles}\ NH_3$$

limiting reagent — conversion factor — amount of product (theoretical yield)

$\frac{1\ \text{mole}\ N_2}{3\ \text{moles}\ H_2}$ $\frac{3\ \text{mole}\ H_2}{2\ \text{moles}\ NH_3}$ $\frac{1\ \text{mole}\ N_2}{2\ \text{moles}\ NH_3}$

$\frac{3\ \text{moles}\ H_2}{1\ \text{mole}\ N_2}$ $\frac{2\ \text{moles}\ NH_3}{3\ \text{mole}\ H_2}$ $\frac{2\ \text{moles}\ NH_3}{1\ \text{mole}\ N_2}$ **and their "one overs"**

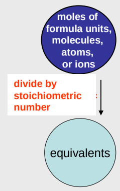
calculating equivalents

How many moles of $AlCl_3$ can be produced from 15.0 g HCl and 12.0 g $Al(OH)_3$?

calculate amounts based on limiting reactant

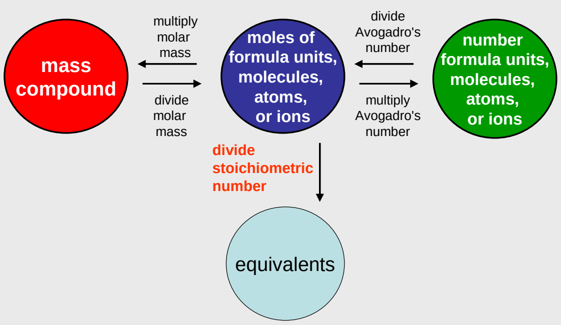
	36.46 g/mol		78.00 g/mol	
initial <u>rxnts</u>	$3HCl\ (aq) + 1\ Al(OH)_3\ (aq) \rightarrow 3H_2O\ (l) + AlCl_3\ (aq)$			
mass	15.0 g		12.0 g	
moles	$\frac{15.0}{36.46}$	= 0.411	$\frac{12.0}{78.00}$	= 0.154
equivalents	$\frac{0.420}{3}$	= 0.137	$\frac{0.154}{1}$	= 0.154

limiting reactant!
smaller # equivalents



$$0.411\ \text{moles HCl} \times \frac{1\ \text{mole } AlCl_3}{3\ \text{moles HCl}} = 0.137\ \text{moles } AlCl_3$$

limiting reactant — conversion factor from reaction — **theoretical yield**



theoretical yield of a reaction

- amount of product obtained based on limiting reactant
- often expressed in grams
- does not account for experimental errors

calculating theoretical yield

Example: Calculate the theoretical yield of CO_2 in grams If 10.0 grams lactic acid ($C_3H_6O_3$) undergoes combustion.

1) write combustion reaction $C_3H_6O_3$ (s) + O_2 (g) → CO_2 (g) + H_2O (l)

2) balance $C_3H_6O_3$ (s) + **3** O_2 (g) → **3** CO_2 (g) + **3** H_2O (l)

3) calculate all molar masses

4) rewrite as mass equation

90.0 g $C_3H_6O_3$ (s) + 3x32.0 g O_2 (g) → 3x44.0 g CO_2 (g) + 3x18.0 g H_2O (l)

$$10.0\ \text{g}\ C_3H_6O_3 \times \frac{3 \times 44.0\ \text{g}\ CO_2}{90.0\ \text{g}\ C_3H_6O_3} = \frac{10 \times 3 \times 44.0}{90.0} = 14.7\ \text{g}\ CO_2$$

molar mass CO_2

C 1 x 12.0
O 2 x 16.0
44.0 g/mol

conversion factor from mass reaction

etc.

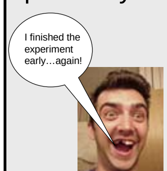
molar mass lactic acid $C_3H_6O_3$

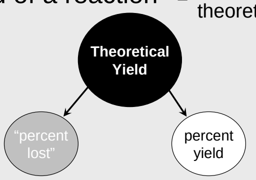
C 3 x 12.0
H 6 x 1.0
O 3 x 16.0
90.0 g/mol

---

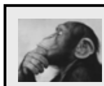
percent yield of a reaction

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

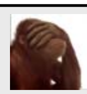




- the part of the theoretical yield obtained
- reflects all  experimental **error**
- usually 0 to 100%



calculating percent yield



Example: a) Calculate the theoretical yield of CO_2 in grams If 10.0 grams lactic acid ($C_3H_6O_3$) undergoes combustion.

$C_3H_6O_3$ (s) + **3** O_2 (g) → **3** CO_2 (g) + **3** H_2O (l)

90.0 g $C_3H_6O_3$ (s) + 3x32.0 g O_2 (g) → 3x44.0 g CO_2 (g) + 3x18.0 g H_2O (l)

$$10.0\ \text{g}\ C_3H_6O_3 \times \frac{3 \times 44.0\ \text{g}\ CO_2}{90.0\ \text{g}\ C_3H_6O_3} = \frac{10 \times 3 \times 44.0}{90.0} = 14.7\ \text{g}\ CO_2$$

b) If 0.88 grams of CO_2 were obtained what was the percent yield of the reaction?

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \qquad \frac{0.88\ \text{g}}{14.7\ \text{g}} \times 100\% = 6.0\ \%$$

Heroin has the molecular formula $C_{21}H_{23}NO_5$. It is synthesized from morphine a natural product isolated from the poppy plant. Morphine has the molecular formula $C_{17}H_{19}NO_3$. Heroin is more potent than morphine because it enters the brain more efficiently, though once in the brain it is converted into morphine. Heroin has numerous street names including: *horse*, *honk*, *junk*, *dope*, *H*, and *smack*.[1]

The reaction for synthesis of *smack* from morphine can be written,

$$C_{17}H_{19}NO_3\ (s) \quad + \quad 2\ C_2H_4O_2\ (l) \quad \rightarrow C_{21}H_{23}NO_5(s) \quad + \quad 2\ H_2O\ (l)$$

Mass (g)				
Moles				
Equivalents				

a) What is the limiting reagent if 50.0 grams of morphine is reacted with 45.0 grams of $C_2H_4O_2$?

b) Write the chemical reaction for the synthesis of smack from morphine using the mass interpretation.

c) Calculate the theoretical yield of smack based on the limiting reagent from part a.

d) Calculate the percent yield if 48.55 g of smack was obtained.

1) http://en.wikipedia.org/wiki/Heroin#Regulation

heroin

handout

48. For each of the following *unbalanced* chemical equations, suppose that exactly 1.00 g of *each* reactant is taken. Determine which reactant is limiting, and calculate what mass of the product in boldface is expected (assuming that the limiting reactant is completely consumed).

a. $CS_2(l) + O_2(g) \rightarrow \mathbf{CO_2}(g) + SO_2(g)$

48. For each of the following *unbalanced* chemical equations, suppose that exactly 1.00 g of *each* reactant is taken. Determine which reactant is limiting, and calculate what mass of the product in boldface is expected (assuming that the limiting reactant is completely consumed).

b. $NH_3(g) + CO_2(g) \rightarrow CN_2H_4O(s) + \mathbf{H_2O}(g)$

45. For each of the following *unbalanced* reactions, suppose exactly 5.00 g of *each reactant* is taken. Determine which reactant is limiting, and also determine what mass of the excess reagent will remain after the limiting reactant is consumed.

a. $Na_2B_4O_7(s) + H_2SO_4(aq) + H_2O(l) \rightarrow H_3BO_3(s) + Na_2SO_4(aq)$

61. According to his prelaboratory theoretical yield calculations, a student's experiment should have produced 1.44 g of magnesium oxide. When he weighed his product after reaction, only 1.23 g of magnesium oxide was present. What is the student's percent yield?

66. A common undergraduate laboratory analysis for the amount of sulfate ion in an unknown sample is to precipitate and weigh the sulfate ion as barium sulfate.

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

The precipitate produced, however, is very finely divided, and frequently some is lost during filtration before weighing. If a sample containing 1.12 g of sulfate ion is treated with 5.02 g of barium chloride, what is the theoretical yield of barium sulfate to be expected? If only 2.02 g of barium sulfate is actually collected, what is the percent yield?