Limiting Reagent Examples from Lecture of 9/30/05

An efficient method for identifying the limiting reagent uses the idea that reactants combine in what might be called "sets" of moles indicated by the stoichiometric coefficients in the balanced equation. If we divide the moles of each reagent by its stoichiometric coefficient, the result with the lowest number will indicate the reagent with the fewest "sets", which is the limiting reagent. We then use the number of moles of this species (not its number of "sets") in all subsequent stoichiometric calculations. The following example illustrates this method.

Example: What is the theoretical yield of the reaction

$$3Ca(OH)_2(s) + 2H_3PO_4(l)\ 6\ Ca_3(PO_4)_2(s) + 3H_2O(l)$$

when 10.00 g $Ca(OH)_2$ and 10.00 g H_3PO_4 are mixed? [f.w. $Ca(OH)_2$ = 74.10 u; m.w. H_3PO_4 = 97.99 u; f.w. $Ca_3(PO_4)_2$ = 310.18 u]

mol $Ca(OH)_2$ = (10.00 g $Ca(OH)_2$)(1 mol $Ca(OH)_2$/74.10 g $Ca(OH)_2$) = 0.1349_{53} mol $Ca(OH)_2$

mol H_3PO_4 = (10.00 g H_3PO_4)(1 mol H_3PO_4/97.99 g H_3PO_4) = 0.1020_{51} mol H_3PO_4

To identify the limiting reagent divide each number of moles by its stoichiometric coefficient in the balanced equation, thereby determining the number of "sets" of each.

"sets" $Ca(OH)_2$ = (0.1349_{53} mol $Ca(OH)_2$)(1 "set" $Ca(OH)_2$/3 mol $Ca(OH)_2$)
= 0.04498 "set" $Ca(OH)_2$

"sets" H_3PO_4 = (0.1020_{51} mol H_3PO_4)(1 "set" H_3PO_4/2 mol H_3PO_4)
= 0.05103 "set" H_3PO_4

$\Rightarrow$ $Ca(OH)_2$ limits, because it has the fewer "sets".

Now, use the moles of $Ca(OH)_2$ (*not the number of "sets"!*) in all subsequent calculations.

g $Ca_3(PO_4)_2$ = (0.1349_{53} mol $Ca(OH)_2$)(1 mol $Ca_3(PO_4)_2$/3 mol $Ca(OH)_2$)
x (310.18 g $Ca_3(PO_4)_2$/mol $Ca_3(PO_4)_2$)
= 13.95 g $Ca_3(PO_4)_2$

How many grams of H_3PO_4 are left over? Calculate the moles H_3PO_4 used, subtract this from the moles H_3PO_4 present, and convert the remaining number of moles to grams.

mol H_3PO_4 used = (0.1349_{53} mol $Ca(OH)_2$)(2 mol H_3PO_4/3 mol $Ca(OH)_2$) = 0.08996_{87} mol H_3PO_4

mol H_3PO_4 left = mol initial - mol used = 0.1020_{51} mol – 0.08996_{87} mol = 0.0121 mol H_3PO_4

g H_3PO_4 left = (0.0121 mol H_3PO_4)(97.99 g H_3PO_4/ mol H_3PO_4) = 1.18 g H_3PO_4

Example: Iodic acid, HIO_3, can be prepared by the following reaction

$$I_2(s) + 5H_2O_2(l) \rightarrow 2HIO_3(aq) + 4H_2O(l)$$

What is the theoretical yield in grams of iodic acid in the reaction of 16.00 g I_2 and 10.00 g H_2O_2? How many grams of the non-limiting reagent will be left over? [Molecular weights: I_2 = 253.8 u, H_2O_2 = 34.01u, HIO_3 = 175.9 u]

First calculate the moles of each reactant, then divide each by its stoichiometric coefficient..

mol I_2 = (16.00 g I_2)(1 mol I_2/253.8 g I_2) = 0.06304 mol I_2

“sets” I_2 = (0.06304 mol I_2)(“set” I_2/1 mol I_2) = 0.06304 “set” I_2

mol H_2O_2 = (10.00 g H_2O_2)(1 mol H_2O_2/34.01 g H_2O_2) = 0.2940 mol H_2O_2

“sets” H_2O_2 = (0.2940 mol H_2O_2)(“set H_2O_2/5 mol H_2O_2) = 0.05880 “set” H_2O_2

Therefore, moles of H_2O_2 limits and we will base our calculation of grams of HIO_3 on it, not the moles of I_2.

$$\text{g HIO}_3 = (0.2940\ \text{mol H}_2\text{O}_2)\left(\frac{2\ \text{mol HIO}_3}{5\ \text{mol H}_2\text{O}_2}\right)\left(\frac{175.9\ \text{g HIO}_3}{\text{mol HIO}_3}\right) = 20.69\ \text{g HIO}_3$$

The I_2 is present in excess, so some of it will be left over after the reaction is complete. As in the preceding example, calculate how many moles of I_2 is used, subtract from the moles initially present, and convert the remaining number of moles to grams.

mol I_2 used = (0.2940 mol H_2O_2)(1 mol I_2/5 mol H_2O_2) = 0.05880 mol I_2

mol I_2 left = mol initial - mol used = 0.06304 mol – 0.05880 mol = 0.00424 mol I_2

g I_2 left = (0.00424 mol I_2)(253.8 g I_2/mol I_2) = 1.07_{61} g I_2 = 1.08 I_2