Theoretical Yield: Example 1

What is the theoretical yield of ethylene in the acid-catalyzed production of ethylene from ethanol, if you start with 100 g of ethanol?

\[
\text{CH}_3\text{CH}_2\text{OH} \xrightarrow{H^+} \text{H}_2\text{C}=\text{CH}_2 + \text{H}_2\text{O}
\]

ethanol \hspace{1cm} \text{ethylene}

The reaction as written above is balanced, with one mole of ethanol producing one mole of ethylene, therefore the stoichiometry is 1:1. **The acid, written over the arrow, is a catalyst and does not enter into the theoretical yield calculations.** To calculate the theoretical yield, determine the number of moles of each reactant, in this case the sole reactant ethanol. Convert the 100 g to moles; the molecular weight of ethanol is 46 g/mole, therefore:

\[
100 \text{ g} \times \frac{1 \text{ mole}}{46 \text{ g}} = 2.17 \text{ mole}
\]

Since there is only one reactant, it is also the limiting reagent. The theoretical number of moles of ethylene is 2.17. Since the molecular weight of ethylene is 28 g/mole, this corresponds to 61 g from the following calculation:

\[
2.17 \text{ mole} \times \frac{28 \text{ g}}{1 \text{ mole}} = 61 \text{ g}
\]

The theoretical yield is therefore “61 g”. Note that theoretical yield is expressed as mass or “grams”.

Actual yields are expressed as percentages. For instance, if you start with 100 g of ethanol and isolate 50 g of ethylene, the actual yield is 82%:

\[
\frac{50 \text{ g}}{61 \text{ g}} = 82\%
\]

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