

CONCEPT: PERCENT YIELD

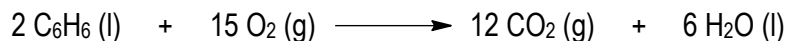
- **Percent Yield** determines how successful the scientist was in creating their desired product.
 - The higher the percent yield then the _____ the efficiency of a chemical reaction.
 - In terms of percent yield values: _____ = Excellent, _____ = Very Good, _____ = Good, & _____ = Poor

Percent Yield Formula

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

- **Actual Yield:** The amount of pure product _____ created when the experiment is done in a laboratory.
 - The units used in the formula are based on the units of the actual yield.
 - No chemical reaction is 100% efficient so the actual yield is always _____ than the theoretical yield.

EXAMPLE: Consider the following balanced chemical reaction:



If a 2.6 g sample of C_6H_6 reacted with excess O_2 to produce 1.25 g of water, what is the percent yield of water?

STEP 1: Map out the portion of the stoichiometric chart you will use.

STEP 2: Convert the **Given** quantity into moles of **Given**.

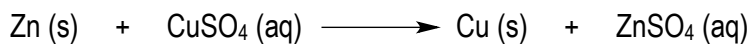
STEP 3: Do a **Mole to Mole comparison** to convert moles of **Given** into moles of **Unknown**.

STEP 4: If necessary, convert the moles of **Unknown** into the final desired units.

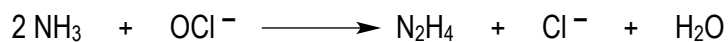
STEP 5: Plug in the actual yield and theoretical yield into the formula to determine the percent yield.

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PRACTICE: What is the percent yield for a reaction in which 22.1 g Cu is isolated by reacting 45.5 g Zn with 70.1 g CuSO₄?



PRACTICE: Ammonia, NH₃, reacts with hypochlorite ion, OCl⁻, to produce hydrazine, N₂H₄. How many grams of hydrazine are produced from 115.0 g NH₃ if the reaction has a 81.5% yield?



PRACTICE: The reduction of iron (III) oxide creates the following reaction:



If the above reaction only went to 75% completion, how many moles of Fe₂O₃ were require to produce 0.850 moles of Fe?