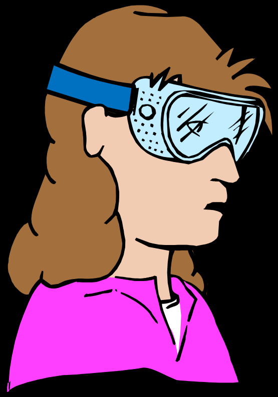
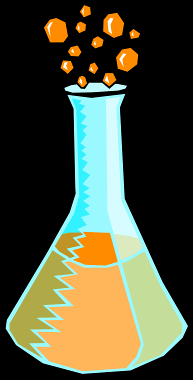
Calculating Yields





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Yields

Theoretical Yield:

Amount of product formed from *complete* conversion of a given amount of reactant to product (From stoichiometry calculations)

Prediction of EXPECTED Yield

Actual Yield:

Amount of product obtained in an experiment

Lower than theory:

- **Side Reactions**
- **Impure Reagents**
- **Mechanical Loss**
- **Errors (weighing, etc.)**

% Yield:

Actual yield expressed as a percentage of the theoretical yield

$$\% \text{ Yield} = \frac{\text{actual}}{\text{theoretical}} \times 100$$

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Theoretical vs. Actual

Theoretical Yield	Actual Yield
Calculated	Measured
Larger	(Should be) Smaller
Predicted	Obtained in the lab
may be supplied (given) OR calculated from $g(A) \rightarrow g(B)$	Measured in lab (given)

Determine Percent Yield From Data

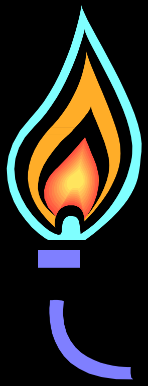
A chemist isolated 181 grams of CO_2 from the complete combustion of propane (C_3H_8). If the theoretical yield was 203 g of CO_2, what is the percent yield for this reaction?

$$C_3H_8 + 5\,O_2 \rightarrow 3\,CO_2 + 4\,H_2O$$

Given: actual yield = 181 g carbon dioxide

theoretical yield = 203 g carbon dioxide

Wanted: percent yield



% Yield = actual / theoretical x 100

$$\%\ \text{Yield} = \frac{181\ \text{g}}{203\ \text{g}} \times 100 = 89.2\ \%$$

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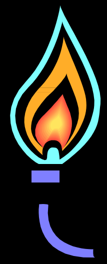
Determine Percent Yield From Stoichiometry

Calculate the percent yield if 70.6 grams of water were obtained when 52.5 grams of propane were burned in oxygen.



$$C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$$

Given:

actual yield = 70.6 g water

starting material = 52.5 g propane

Wanted: percent yield

Molar Mass: water = 18.02 g/mole; propane = 44.10 g/mole

Set-up conversion string for theoretical yield of water:

$$52.5\ g\ C_3H_8 \times \frac{1\ mole\ C_3H_8}{44.10\ g\ C_3H_8} \times \frac{4\ mole\ H_2O}{1\ mole\ C_3H_8} \times \frac{18.02\ g\ H_2O}{1\ mole\ H_2O} = 85.8\ g$$

% Yield = actual/theory x 100

$$\%\ Yield = \frac{70.6\ g}{85.8\ g} \times 100 = 82.3$$

Let the units drive the solution

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Determine Percent Yield From Stoichiometry

Calculate the theoretical yield of Copper (in grams) when 85.0 g of oxygen reacts with excess (xs) copper (I) sulfide.

$$Cu_2S_{(s)} + O_{2(g)} \rightarrow 2\ Cu_{(s)} + SO_{2(g)}$$

Molar Mass: O_2 = 32.00 g/mole; Cu = 63.55 g/mole

Set-up conversion string for theoretical yield of copper:

$$85.0\text{ g } O_2 \times \frac{1\text{ mol } O_2}{32.00\text{g } O_2} \times \frac{2\text{ mol Cu}}{1\text{ mol } O_2} \times \frac{63.55\text{ g Cu}}{1\text{ mol Cu}} = 338\text{ g}$$

If you isolated 231 g, what was your percent yield?

$$\frac{231\text{ g}}{338\text{ g}} \times 100 = 68.3\ \%$$

Let the units drive the solution

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Determine Percent Yield From Stoichiometry

Calculate the theoretical yield of H_2 gas in grams when 45.00 g of Al react with excess HBr.

$$2\,Al + 6\,HBr \rightarrow 2\,AlBr_3 + 3\,H_2$$

Molar Mass: Al = 26.98 g/mole; H2 = 2.016 g/mole

Theoretical Yield of Hydrogen gas:

$$45.00\text{ g Al} \times \frac{1\text{ mol Al}}{26.98\text{ g Al}} \times \frac{3\text{ mol } H_2}{2\text{ mol Al}} \times \frac{2.016\text{ g } H_2}{1\text{ mol } H_2} = 5.044\text{ g}$$

If you obtained 4.785 g H_2, what is the percent yield?

$$\frac{4.785\text{ g } H_2}{5.044\text{ g } H_2} \times 100 = 94.87\ \%$$

Let the units drive the solution

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Determine Percent Yield From Stoichiometry

Lead (II) nitrate reacts with sodium iodide to form lead(II) iodide and sodium nitrate. What is the theoretical yield of lead (II) iodide if 138.820 g of sodium iodide are reacted with an excess amount of Lead(II) nitrate?

Write the balanced reaction: $Pb(NO_3)_{2\,(aq)} + 2\ NaI_{\,(aq)} \rightarrow PbI_{2\,(s)} + 2\ NaNO_{3\,(aq)}$

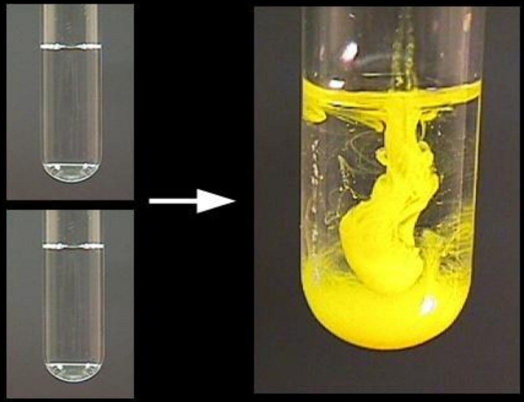
Molar Mass NaI: 149.89 g/mole; molar mass PbI_2 : 460.99 g/mole

Setting up the linear string starting with given mass of NaI:

$$138.820\text{ g NaI} \times \frac{1\text{ mole NaI}}{149.89\text{ g}} \times \frac{1\text{ mole } PbI_2}{2\text{ moles NaI}} \times \frac{460.99\text{ g}}{\text{mole } PbI_2} = 213.47\text{ g}$$

If 197.5 grams were isolated, what is the experimental yield?

$$\frac{197.5\text{ g}}{213.47\text{ g}} \times 100 = 92.5180 \rightarrow 92.52\ \%$$



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Typical Insert Problem

Calculate theoretical yield when 24.9 g of phosphorus pentachloride reacts with water to yield phosphoric and hydrochloric acids.

Write Balanced chemical reaction:

$$PCl_5 + 4\,H_2O \rightarrow H_3PO_4 + 5\,HCl$$

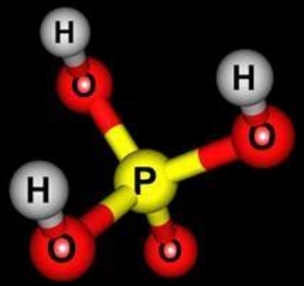
Use stoichiometry to determine theoretical yield:

$$24.9\text{ g } PCl_5 \times \frac{1\text{ mole } PCl_5}{208.22\text{ g}} \times \frac{5\text{ moles HCl}}{1\text{ mole } PCl_5} \times \frac{36.46\text{ g HCl}}{\text{mole HCl}} = 21.8\text{ g}$$

If a chemist isolated 19.234 g, what is the percent yield?

$$\frac{19.234\text{ g}}{21.8} \times 100 = 88.2\ \%$$

Let the units drive the solution



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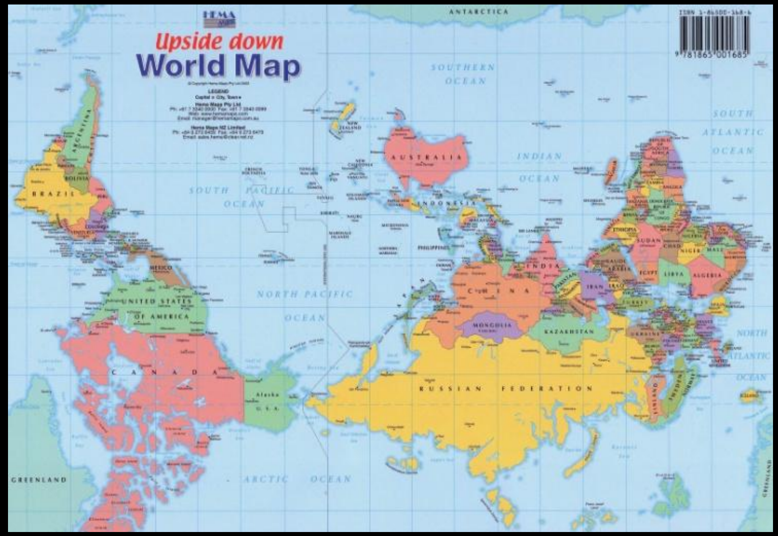
Most chemical processes yield less than 100%

This becomes a severe problem for multiple-step operations
Say 90% yield for each step, then after 5 steps:
$(0.90)^5$ = ~ 59% of wanted product: rest may be useless

Global manufacturing (and lab research)
centers around
Increasing % yields
Separating wanted products from reactants & unwanted products



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