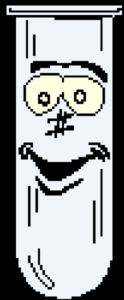


Unit 07 Outcomes

For a chemical equation, develop "per expressions" from mole ratios

For $2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$

"per expressions" (Conversion factors) based on coefficients of balanced equation

$\frac{2 \text{ moles } C_2H_6}{7 \text{ moles } O_2}$ $\frac{2 \text{ moles } C_2H_6}{6 \text{ moles } H_2O}$ $\frac{2 \text{ moles } C_2H_6}{4 \text{ moles } CO_2}$

$\frac{7 \text{ moles } O_2}{2 \text{ moles } C_2H_6}$ $\frac{7 \text{ moles } O_2}{4 \text{ moles } CO_2}$ $\frac{7 \text{ moles } O_2}{6 \text{ moles } H_2O}$

$\frac{4 \text{ moles } CO_2}{2 \text{ moles } C_2H_6}$ $\frac{4 \text{ moles } CO_2}{7 \text{ moles } O_2}$ $\frac{4 \text{ moles } CO_2}{6 \text{ moles } H_2O}$

$\frac{6 \text{ moles } H_2O}{2 \text{ moles } C_2H_6}$ $\frac{6 \text{ moles } H_2O}{7 \text{ moles } O_2}$ $\frac{6 \text{ moles } H_2O}{4 \text{ moles } CO_2}$

Always:

$\frac{\text{Moles Wanted}}{\text{Moles Given}}$

All products/reactants related by their coefficients (molar ratios)

LPT

Given a balanced chemical equation, calculate the # of moles of any component given the moles of any other species.

$$\text{Moles given} \times \frac{\text{Coefficients wanted}}{\text{Coefficients given}} = \text{moles wanted}$$

Where coefficients wanted and given is the ratio derived from balanced equation coefficients

Given a balanced chemical equation, calculate the # of grams of any component given the moles of any other species.

$$\text{Moles given} \times \frac{\text{moles wanted}}{\text{moles given}} \times \frac{\text{molar mass wanted}}{\text{1 mole wanted}} = \text{grams wanted}$$

Where moles wanted and given is the ratio derived from balanced equation coefficients

How many moles of aluminum chloride are produced when 1.35 moles of aluminum oxide are reacted with hydrochloric acid

$$Al_2O_{3(s)} + 6\ HCl_{(aq)} \rightarrow 2\ AlCl_{3(aq)} + 3\ H_2O_{(l)}$$

$$1.35 \text{ moles } Al_2O_3 \times \frac{2 \text{ moles } AlCl_3}{1 \text{ mole } Al_2O_3} = 2.70 \text{ moles } AlCl_3$$

The path:

Moles Al_2O_3 ➔ moles $AlCl_3$

LPT

The # of moles of Fe that react with 15.7 g of S to form Iron (III) Sulfide

$$2\ Fe + 3\ S \rightarrow Fe_2S_3$$

Given: 15.7 g S
Wanted: moles Fe

$$15.7\ g\ S \times \frac{1\ mole\ S}{32.06\ g} \times \frac{2\ moles\ Fe}{3\ moles\ S} = 0.326\ moles$$

The path:

Grams S → Moles S → Moles Fe

LPT

Given a balanced chemical equation, calculate the # moles of any component given the # grams of any other species.

$$\text{Grams given} \times \frac{1 \text{ mole}}{\text{molar mass given}} \times \frac{\text{moles wanted}}{\text{moles given}} = \text{moles wanted}$$

Where moles wanted and given is the ratio derived from balanced equation coefficients

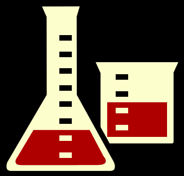
Given a balanced chemical equation, calculate the # gram of any component given the # grams of any other species.

$$\#\text{g given} \times \frac{1 \text{ mole}}{\text{molar mass given}} \times \frac{\text{moles wanted}}{\text{moles given}} \times \frac{\text{molar mass wanted}}{1 \text{ mole}} = \#\text{g wanted}$$

Where moles wanted and given is the ratio derived from balanced equation coefficients

LPT

How many grams of aluminum chloride are produced if 17.75 grams of aluminum are reacted with chlorine?



$$2\,Al + 3\,Cl_2 \rightarrow 2\,AlCl_3$$

$$17.75\text{ g Al} \times \frac{1\text{ mole Al}}{26.98\text{ g}} \times \frac{2\text{ mole }AlCl_3}{2\text{ moles Al}} \times \frac{133.3\text{ g}}{1\text{ mole }AlCl_3} = 87.70\text{ g }AlCl_3$$

The path:

Grams Al → Moles Al → Moles $AlCl_3$ → Grams $AlCl_3$

LPT

Define the following terms as they apply to chemical reactions: theoretical yield; actual yield; % yield

Theoretical Yield:

Amount of product formed from *complete* conversion of a given amount of reactant to product

Actual Yield:

Amount of product obtained in an experiment

Typically less than theoretical

% Yield:

Actual yield expressed as a percentage of the theoretical yield

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

For the reaction $BaCl_{2\,(aq)}$ + $Na_2SO_{4(aq)}$ → $BaSO_{4(s)}$ + 2 $NaCl_{(aq)}$, the theoretical yield of barium sulfate was calculated to be 27.85 grams. If 23.45 grams of $BaSO_4$ were actually obtained, what was the percent yield?

$$\frac{23.45\text{ g}}{27.85\text{ g}} \times 100 = 84.20\ \%$$

For the reaction of sodium chlorate decomposing to oxygen and sodium chloride, the theoretical yield of sodium chloride was calculated to be 116.8 grams. If 85.29 g of NaCl were actually obtained, what was the percent yield?

$2\ NaClO_3 \rightarrow 2\ NaCl + 3\ O_2$ **(not needed to solve the problem)**

$$\frac{85.29\text{ g}}{116.8\text{ g}} \times 100 = 73.02\ \%$$

LPT

Given the actual yield & information from which the theoretical yield can be calculated, determine the % yield.

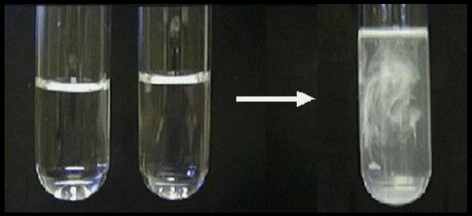
- **Use Correct Formulas for Reactants & Products**
- **Balance Equation**
- **Set-up conversion string for theoretical yield of wanted compound**
  - **Usually involves determining molar mass**
- **% Yield:**

Actual yield expressed as a percentage of the theoretical yield

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

LPT

For the precipitation reaction: lead (II) nitrate plus sodium iodide yields lead (II) iodide precipitate and sodium nitrate.



$$Pb(NO_3)_{2(aq)} + 2\,NaI_{(aq)} \rightarrow PbI_{2(s)} + 2\,NaNO_{3(aq)}$$

If 0.925 moles of sodium iodide are reacted with excess lead(II) nitrate, what is the theoretical yield of lead (II) iodide?

$$0.925 \text{ moles NaI} \times \frac{1 \text{ mole } PbI_2}{2 \text{ moles NaI}} \times \frac{460.99 \text{ g}}{1 \text{ mole } PbI_2} = 213.20 \rightarrow 213 \text{ g } PbI_2$$

If 197.50 grams of lead (II) iodide were actually produced, what was the percent yield?

$$\frac{197.50 \text{ g}}{213} \times 100 = 92.7\ \%$$

LPT