Chapter 3: STOICHIOMETRY: MASS, FORMULAS, AND REACTIONS


3.2 THE MOLE

Stoichiometry (STOY-key-OM-e-tree): quantitative study of reactants and products in a chemical reaction

Interpreting a Chemical Equation

\[ \text{H}_2 (g) + \text{Cl}_2 (g) \rightarrow 2 \text{HCl} (g) \]

1 molecule 1 molecule 2 molecules

It follows that any multiples of these coefficients will be in same ratio!

\[ 2 \text{H}_2 (g) + \text{O}_2 (g) \rightarrow 2 \text{H}_2\text{O}(g) \]

\[
\times 1000 \quad \text{molecule(s)} \quad \text{molecule(s)} \quad \text{molecule(s)}
\]

\[
\times N \quad \text{molecule(s)} \quad \text{molecule(s)} \quad \text{molecule(s)}
\]

Since \( N = \text{Avogadro's \#} = 6.022 \times 10^{23} \text{ molecules} = 1 \text{ mole} \)

\[ 2 \text{H}_2 (g) + \text{O}_2 (g) \rightarrow 2 \text{H}_2\text{O}(g) \]

\[
\text{mole(s)} \quad \text{mole(s)} \quad \text{mole(s)}
\]

Thus, the coefficients in a chemical equation give the \textit{mole ratios} of reactants and products.

Consider the combustion of butane: \[ 2 \text{C}_4\text{H}_{10}(g) + 13 \text{O}_2(g) \rightarrow 8 \text{CO}_2(g) + 10 \text{H}_2\text{O}(g) \]

1. How many moles of \text{O}_2 will react with 2.50 moles of butane?

2. How many moles of \text{CO}_2 form when 3.50 moles of \text{O}_2 completely react?
3.5 STOICHIOMETRIC CALCULATIONS AND THE CARBON CYCLE

Mass-Mass Stoichiometry Problems

3.4 Combustion Reactions:

\[ \text{C}_x\text{H}_y + \text{O}_2(g) \xrightarrow{\Delta} \text{CO}_2(g) + \text{H}_2\text{O}(g) \]
\[ \text{C}_x\text{H}_y\text{O}_z + \text{O}_2(g) \xrightarrow{\Delta} \text{CO}_2(g) + \text{H}_2\text{O}(g) \]

Hydrocarbons (compounds with only C and H) and hydrocarbon derivatives (compounds with only C, H and O) burn in O\(_2\) to produce CO\(_2\) gas and steam, H\(_2\)O(g).

Ex. 1: Many home barbecues are fueled with propane gas (C\(_3\)H\(_8\)).
   a. Write the balanced equation for the combustion of propane.
   b. Calculate the mass (in kg) of carbon dioxide produced upon complete combustion of liquid propane from a 5.0 gal tank. (Note: Liquid propane’s density at 60°F is about 4.2 lbs. per gallon, and 1 lb. = 453.6 g.)

Ex. 2 Gasoline used in cars is a mixture containing octane (C\(_8\)H\(_{18}\)), other hydrocarbons, and often some ethanol. To simplify matters, assume a sample of gasoline is pure octane, and calculate the mass (in kg) of carbon dioxide produced upon complete combustion of 1.00 gallon of gasoline. (The density of gasoline at 60°F is 0.737 g/mL, 1 gal. = 4 qt., and 1 qt. = 946 mL) Write the balanced chemical equation for the combustion of octane.
3.9 LIMITING REACTANTS (or LIMITING REAGENTS) AND PERCENT YIELD

In practice, reactants will not always be present in the exact amounts necessary to be converted completely into products.

Some reactants (usually the more expensive) are only present in a limited supply, so these are almost always completely used up → “limiting reactant” (or limiting reagent) since it limits the amount of product made

Some reactants (usually the less expensive) are present in larger amounts and are never completely used up → “reactant(s) in excess”

GUIDELINES for Solving Limiting Reactant Problems

1. Calculate the mass or the # of moles of the 2nd reactant needed to completely react with the 1st reactant.
   - If the moles needed is greater than the number of moles present for the 2nd reactant → That 2nd reactant will run out before the 1st reactant.
     → The 2nd reactant = the limiting reactant, and the 1st reactant is in excess.
   - If the moles needed is less than the number of moles present for the reactant, → The 1st reactant = the limiting reactant, and the 2nd reactant is in excess.

2. Use the amount of the limiting reactant present to solve for the mass or # of moles of product that can be made.

Note: You may use any method to solve limiting reactant problems, but you must show sufficient work to determine the limiting reactant—e.g. comparing the moles or mass of reactants present with the amount needed, solving for the moles or mass of product produced by each reactant, etc.

CALCULATING PERCENT YIELD: Percent yield = \( \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \)

theoretical yield: Amount of product one should get based on the chemical equation and the amount of reactants present
   - One generally calculates this in grams from info given

actual yield: Amount of product one actually obtains
   - Generally smaller than the theoretical yield because of impurities and other adverse conditions in the lab
   - This is generally determined experimentally in the lab or given for a problem in lecture.
Consider the reaction to produce ammonia: \( \text{N}_2(g) + 3 \text{H}_2(g) \rightarrow 2 \text{NH}_3(g) \)

Ex. 1. a. If 40.0 g of N\(_2\) react with 10.0 g of H\(_2\), what mass of ammonia is produced?

b. The limiting reactant is ____________, and the reactant in excess is__________.

c. What mass of the reactant in excess remains after the reaction?

b. If 45.7 g of ammonia were produced, calculate the percent yield for the reaction.

Ex. 2: **Nitrogen triiodide reacts explosively on contact** to produce a bright pinkish-purple cloud of iodine gas:

\[
2 \text{NI}_3(s) \rightarrow \text{N}_2(g) + 3 \text{I}_2(g)
\]

What is the percent yield if 50.0 g of nitrogen triiodide decomposes to produce 46.1 g of iodine gas?
Ex. 3  The presence of iron(III) ions in water is evident by their characteristic orange-brown color. The ions can be precipitated out using hydroxide ions, as shown below:

\[
\text{Fe(NO}_3\text{)}_3(\text{aq}) + 3 \text{NaOH}(\text{aq}) \rightarrow \text{Fe(OH)}_3(\text{s}) + 3 \text{NaNO}_3(\text{aq})
\]

a. When 100.0 g of iron(III) nitrate react with 100.0 g of sodium hydroxide, what mass of precipitate can be produced? Indicate the limiting reactant and the reactant in excess.

Animas River before and after iron contamination

https://www.reddit.com/r/pics/comments/3gb02n/before_and_after_pictures_of_the_animas_river/

mass of precipitate produced = ______________________________

limiting reagent = _____________ reactant in excess = _____________

b. What mass of the reactant in excess remains after the reaction?

c. What is the percent yield if 42.55 g of precipitate are actually produced?
Ex. 4 Methane is the primary component (about 70-95%) of natural gas, which burns to produce carbon dioxide. In Washington state burning natural gas produces 10.7% of the state's electricity, and Washington State Ferries (WSF) are considering conversion of their diesel fuel to liquid natural gas (LNG). If the typical WSF ferry burns 28,000 gallons of LNG per week and emits 123 tons of CO$_2$ from the combustion of methane, calculate percentage of CO$_2$ emitted relative to the total amount produced by CH$_4$'s combustion each week. (Use $d_{\text{LNG}}=0.45$ g/cm$^3$, the LNG is 95% CH$_4$, 1 gal. ≡ 4 qt., 1 qt. = 946 mL, 1 lb. = 453.6 g, and 1 ton = 2000 lb.) Write the balanced equation for CH$_4$'s combustion.

"Gases that trap heat in the atmosphere are called greenhouse gases." Consider the image below indicating the primary greenhouse gases and the main sources of these gases. Note that methane not only burns to produce carbon dioxide, but it is itself a greenhouse gas that traps heat more effectively than carbon dioxide.

https://www3.epa.gov/climatechange/ghgemissions/inventoryexplorer/

3 https://www.epa.gov/ghgemissions/overview-greenhouse-gases