Chapter 8: Quantities in Chemical Reactions

Read Chapter 8 and Check MasteringChemistry due dates.

**Stoichiometry:** Chemical arithmetic, numerical relationships between chemicals in a balanced reaction.

Stoichiometry is a fancy name to represent the math and conversions used in converting mass to moles, molecules to moles, moles of one compound to moles of another in a balanced chemical reaction, and many other math conversions related to chemicals amounts.

**Coefficients:**

The *coefficients* in any balanced equation represent the relative ratio between the number of moles or amount of particles of each substance. *The coefficients DO NOT represent mass.*

Moles to mole ratios in a balanced chemical reaction are equivalents.

\[ 3 \text{ H}_2 (g) + \text{ N}_2 (g) \rightarrow 2 \text{ NH}_3 (g) \]

**Conversions:**

**Mole to Mole Conversions**

*Try this 1:*

3 \text{ H}_2 (g) + \text{ N}_2 (g) \rightarrow 2 \text{ NH}_3 (g) tells us that 1 mole of nitrogen will stoichiometrically react with 3 moles of hydrogen to theoretically (ideally) produce 2 moles of ammonia.

a) If we have 3.00 moles of \text{ N}_2, and more than enough \text{ H}_2, how much \text{ NH}_3 can we make?
b) How many moles of \( \text{H}_2 \) are needed to stoichiometrically react with 3.75 moles of \( \text{N}_2 \) gas?

c) How many moles of \( \text{NH}_3 \) gas will theoretically be produced from 3.75 moles of \( \text{N}_2 \) and the necessary amount of \( \text{H}_2 \) gas?

d) How many moles of \( \text{N}_2 \) gas and \( \text{H}_2 \) gas will theoretically be required to produce 0.258 moles of \( \text{NH}_3 \) (g)

**Mass to Mass Conversions**
- Solve for necessary molar masses for \( A \) and \( B \)
- Convert grams \( A \) to moles \( A \)
- Convert moles \( A \) to moles \( B \)
- Convert moles \( B \) to grams \( B \)

**Example:**
What mass of \( \text{CO}_2 \) is emitted by a car when \( 5.0 \times 10^2 \) g of pure octane are used?

\[
2 \text{C}_8\text{H}_{18}(l) + 25 \text{O}_2(g) \rightarrow 16 \text{CO}_2(g) + 18 \text{H}_2\text{O}(g)
\]

\[
5.0 \times 10^2 \text{ g C}_8\text{H}_{18} \times \frac{1 \text{ mol C}_8\text{H}_{18}}{114.3 \text{ g C}_8\text{H}_{18}} \times \frac{16 \text{ mol CO}_2}{2 \text{ mol C}_8\text{H}_{18}} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 1.5 \times 10^3 \text{ g CO}_2
\]
Stoichiometry in Action

Is there enough Oxygen to burn Octane completely?

• The balanced equation shows that 2 moles of octane require 25 moles of oxygen to burn completely:

\[ 2 \text{C}_8\text{H}_{18}(l) + 25 \text{O}_2(g) \rightarrow 16 \text{CO}_2(g) + 18 \text{H}_2\text{O}(g) \]

• A shortage of O\(_2\) creates pollutants such as carbon monoxide (CO) or carbon soot.

• 1990 amendments to the Clean Air Act require oil companies to put additives in gasoline that increased its oxygen content.

• MTBE (methyl tertiary butyl ether, CH\(_3\)OC(CH\(_3\))\(_3\)), a compound that does not readily biodegrade, was the additive of choice by the oil companies.

• MTBE is a polar molecule that dissolves in polar water. It made its way into drinking water through gasoline spills at gas stations, from boat motors, and from leaking storage tanks. MTBE gives water an unpleasant taste at very low concentrations, and can make groundwater non-potable.

• Ethanol (C\(_2\)H\(_5\)OH), made from the fermentation of grains, has replaced MTBE to increase oxygen content in gasoline. Ethanol was not used originally because of its higher cost.

Try this 2:

\[ 2 \text{C}_4\text{H}_{10}(g) + 13 \text{O}_2(g) \rightarrow 8 \text{CO}_2(g) + 10 \text{H}_2\text{O}(l) \]

a) Solve for the molar masses of each species in the above balanced reaction.

b) Determine the mass in grams of oxygen required to stoichiometrically react with 248 grams of butane (C\(_4\)H\(_{10}\))

c) Solve for the mass in grams of each product that will theoretically be formed in the reaction with 248 grams of butane.
Other Conversions:
Using the same concepts, relate molecules to molecules in a balanced reaction. Later we will convert volumes of solutions (Molarity: mol/L) to moles and convert volume of gases to moles (PV = nRT). For gases at standard temperature and pressure (STP) which is 0°C and 1 atm the molar volume is 22.4 L.

Limiting Reactant:
Many times there will be one reactant that will run out before others. The limiting reactant will be the first to run out and makes the least amount of product. The limiting reactant is completely consumed in the reaction.

Theoretical Yield:
The calculated (paper) value completely using up the limiting reactant, assuming perfect conditions, is considered a theoretical yield.

Try this 3:
2 H₂ (g) + O₂ (g) → 2H₂O (l)
a) In an experiment, 3.00 moles of H₂ react with 3.00 mole of O₂. Which is the limiting reactant? Solve for the theoretical yield of H₂O produced in moles.

b) An experiment has 3.00 grams of H₂ react with 3.00 grams of O₂. Which is the limiting reactant? Solve for the theoretical yield of H₂O produced in grams.
Experimental Yield (also known as Actual Yield):
In life, experiments are not perfect (imagine that!). It is possible that some of the reactants fail to react, side reactions may occur producing an unexpected product, the lab technician may fail to collect and measure all the product as some splatters or is lost in a transfer, or the measured product is not pure and has extra contaminants adding mass, etc. The amount recovered is generally less than the theoretical yield. The actual amount of product measured from the experiment is called the *experimental yield (actual yield)*.
Experimental yields are given in a question or determined experimentally in lab.

Percent Yield:
\[ \frac{\text{Experimental Yield}}{\text{Theoretical Yield}} \times 100 = \text{percent yield} \]

*Try this 4:* Solve for the percent yield if the calculated theoretical yield is 24.5 grams and the experimental yield recovered is 21.8 grams.

Enthalpy, \(\Delta H\):
Enthalpy is a quantity of thermal energy (heat in kJ per mole of reaction) associated with chemical and physical changes under constant pressure.

a) Exothermic reactions release thermal energy (-\(\Delta H\))
b) Endothermic reactions absorb thermal energy (+\(\Delta H\))
Stoichiometry of Enthalpy, $\Delta H$:

Amounts of energy required or produced in a reaction can be calculated using enthalpy to mole ratios, similar to mole to mole ratios in a balanced chemical equation. Often in problem solving the absolute value of energy is requested and words are used to convey the sign of heat absorbed or given off in the reaction.

Try this 5:

$$4 \text{NH}_3 (g) + 5 \text{O}_2 (g) \rightarrow 4 \text{NO} (g) + 6 \text{H}_2\text{O} (l) \quad \Delta H = -906 \text{kJ}$$

a) Determine the heat in kJ associated with the complete reaction of 5.05 grams of NH$_3$.

b) How many grams of NH$_3$ must react to produce 125 x 10$^5$ kJ of energy?

c) Determine the limiting reactant associated with the complete reaction of 72.0g of NH$_3$ reacting with 100.0g of O$_2$ gas. Determine the heat in kJ associated with the complete reaction.
Everyday Chemistry: Bunsen Burners
• Most Bunsen burners allow the user to adjust the amount of air/oxygen mixed with the fuel methane.
• Bunsen burners with the air closed off will give a yellow, smoky flame that is not very hot and creates soot (carbon) and CO.
• As the amount of air going into the burner increases, the flame becomes bluer, less smoky, and hotter.
• At the optimum adjustment, the flame has a sharp, inner blue cone, gives off no smoke, and is very hot.
• Too much air causes the flame to become cooler again and it may blow out.

Practice Problems:
1. For the decomposition reaction… \( \text{N}_2\text{H}_4 \text{(l)} \rightarrow \text{NH}_3 \text{(g)} + \text{N}_2 \text{(g)} \)
   a) Balance the reaction
   b) Calculate the moles of each product that is theoretically produced from
      4.72 moles of \( \text{N}_2\text{H}_4 \)

2. For the single replacement reaction…
   \( \text{Al} \text{(s)} + \text{H}_2\text{SO}_4 \text{(aq)} \rightarrow \text{Al}_2\text{(SO}_4)_3 \text{(aq)} + \text{H}_2 \text{(g)} \)
   a) Balance the reaction
   b) Calculate the grams of \( \text{H}_2 \) that is theoretically produced from
      12.0 grams of \( \text{Al} \) and excess \( \text{H}_2\text{SO}_4 \text{(aq)} \).
3. For the single replacement reaction...

\[
\text{HCl (aq)} + \text{O}_2 (g) \rightarrow \text{Cl}_2 (g) + \text{H}_2\text{O (l)}
\]

a) Balance the reaction

b) When 62.3 grams of HCl react with 16.8 grams of O\(_2\), 48.1 grams of Cl\(_2\) gas are collected. Determine the limiting reactant, theoretical yield of Cl\(_2\), and percent yield of Cl\(_2\) for the reaction.

4. For the synthesis reaction...

\[
\text{P}_4 (s) + \text{F}_2 (g) \rightarrow \text{PF}_5 (g)
\]

a) Balance the reaction

b) When 80.0 grams of P\(_4\) react with 100.0 grams of F\(_2\), 128.1 grams of PF\(_5\) gas are collected. Determine the limiting reactant, theoretical yield of PF\(_5\), and percent yield of PF\(_5\) for the reaction.

5. The evaporation of water...

\[
\text{H}_2\text{O (l)} \rightarrow \text{H}_2\text{O (g)} \quad \Delta H_{\text{rxn}} = +44.01 \text{ kJ}
\]

a) Is this an endothermic or exothermic reaction?

b) Exercise is usually accompanied by sweat. The evaporation of sweat from your skin is a way your body cools itself. What mass of water in grams has to evaporate to absorb 255 kJ of heat?