Chapter 9: Stoichiometry

9.1: Calculating Quantities in Reactions

Avogadro's Number: - a group of (6.022×10^{23}) molecules = 1 mole

Stoichiometry: - the calculation of quantities in a chemical reaction.
- the coefficients of various reactants and /or products form **mole ratios**.
- these mole ratios hold for moles, molecules or atoms.

Mole Ratio: - a ratio form between the coefficient of the required chemical amount to the given chemical amount. $\left(\dfrac{\textbf{require coefficient}}{\textbf{given coefficient}}\right)$

Example 1: Interpret the chemical equation $4\ NH_{3\,(g)} + 7\ O_{2\,(g)} \rightarrow 4\ NO_{2\,(g)} + 6\ H_2O_{(g)}$ in terms of
a. moles. b. molecules. c. masses.

	$4\ NH_{3\,(g)}$	$7\ O_{2\,(g)}$	$4\ NO_{2\,(g)}$	$6\ H_2O_{(g)}$
a.	**4 moles of NH_3**	**7 moles of O_2**	**4 moles of NO_2**	**6 moles of H_2O**
b.	**4 molecules of NH_3**	**7 molecules of O_2**	**4 molecules of NO_2**	**6 molecules of H_2O**
c.	$m = nM$ $m = (4\ \cancel{\text{mol}})(17.04\ \text{g}/\cancel{\text{mol}})$ $\boldsymbol{m = 68.16\ \text{g}}$	$m = nM$ $m = (7\ \cancel{\text{mol}})(32.00\ \text{g}/\cancel{\text{mol}})$ $\boldsymbol{m = 224.0\ \text{g}}$	$m = nM$ $m = (4\ \cancel{\text{mol}})(46.01\ \text{g}/\cancel{\text{mol}})$ $\boldsymbol{m = 184.0\ \text{g}}$	$m = nM$ $m = (6\ \cancel{\text{mol}})(18.02\ \text{g}/\cancel{\text{mol}})$ $\boldsymbol{m = 108.1\ \text{g}}$

***Note the Law of Conservation of Mass holds after converting mole of each chemical to its mass.**

Example 2: 4.35 mol of $PCl_{5\,(g)}$ is decomposed into its elements. Write a balance equation and determined the amount of chlorine produced.

$$\underline{4}\ PCl_{5\,(g)} \rightarrow P_{4\,(s)} + \underline{10}\ Cl_{2\,(g)}$$
4.35 mol ? mol

$$n_{Cl_2} = 4.35\ \cancel{\text{mol PCl}_5} \times \frac{\textbf{10 mol Cl}_2}{\textbf{4}\ \cancel{\textbf{mol PCl}_5}} = 10.875\ \text{mol Cl}_2 \qquad \boldsymbol{n_{Cl_2} = 10.9\ \text{mol}}$$

Gravimetric Stoichiometry: - stoichiometry that involves quantities of masses.

Density-Volume Stoichiometry: - stoichiometry that involves quantities of volumes with densities given.

Particles Stoichiometry: - stoichiometry that involves quantities of particles such as atoms and molecules.

General Stoichiometry Procedure:

1. Predict the products and balance the chemical equation.
2. Put all the information given under the appropriate chemicals. If necessary, determine the molar mass of any chemical involved (i.e. if the chemical given or asked involves mass).
3. Find the moles of the given chemical by using the proper conversion factor. This can be $\left(mass \times \frac{1\ \text{mol}}{\text{Molar Mass (g)}}\right)$ if mass is given. It can be $\left(particles\ given \times \frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{particles}}\right)$ if number of atoms or molecules are given. If a Volume is given with Density, then it can be $\left(Volume\ given \times \frac{\text{Density (g)}}{1\ \text{mL or } 1\ \text{L}} \times \frac{1\ \text{mol}}{\text{Molar Mass (g)}}\right)$
4. Continue to find the mole of the required chemical by using mole ratio. $\left(\times \frac{\text{require coefficient}}{\text{given coefficient}}\right)$
5. Convert mole of the required chemical to the type of quantity asks. If the question requires mass, then use conversion factor $\left(\times \frac{\text{Molar Mass (g)}}{1\ \text{mol}}\right)$. If the questions asks for the number of particles, then use conversion factor $\left(\times \frac{6.022 \times 10^{23}\ \text{particles}}{1\ \text{mol}}\right)$. If it wants the volume and density is given, then we can use $\left(\times \frac{\text{Molar Mass (g)}}{1\ \text{mol}} \times \frac{1\ \text{mL or } 1\ \text{L}}{\text{Density (g)}}\right)$

Example 3: Determine the mass of carbon dioxide formed when 50.0 kg of butane ($C_4H_{10\,(l)}$) is burned.

$2\ C_4H_{10\,(g)}$	+	$13\ O_{2\,(g)}$	$\rightarrow$	$8\ CO_{2\,(g)}$	+	$10\ H_2O_{(g)}$
50.0 kg				? g		
58.14 g/mol				M = 44.01 g/mol		

$$50.0\ \text{kg}\ \cancel{C_4H_{10}} \times \frac{1\ \cancel{\text{mol}}\ \cancel{C_4H_{10}}}{58.14\ \cancel{\text{g}}\ \cancel{C_4H_{10}}} \times \frac{8\ \cancel{\text{mol}}\ \cancel{CO_2}}{2\ \cancel{\text{mol}}\ \cancel{C_4H_{10}}} \times \frac{44.01\ \text{g}\ CO_2}{1\ \cancel{\text{mol}}\ \cancel{CO_2}} = 151.3931189\ \text{kg}\ CO_2$$

$$\boldsymbol{m_{CO_2} = 151\ \text{kg}}$$

Example 4: Barium bromide solution was mixed with an excess sodium phosphate solution. What was the mass of barium bromide solid needed in the original solution to form 3.21 g of precipitate?

$3\ BaBr_{2\,(aq)}$	+	$2\ Na_3PO_{4\,(aq)}$	$\rightarrow$	$Ba_3(PO_4)_{2\,(s)}$	+	$6\ NaBr_{(aq)}$
? g				3.21 g		
297.13 g/mol				M = 601.93 g/mol		

$$3.21\ \cancel{\text{g}}\ \cancel{Ba_3(PO_4)_2} \times \frac{1\ \cancel{\text{mol}}\ \cancel{Ba_3(PO_4)_2}}{601.93\ \cancel{\text{g}}\ \cancel{Ba_3(PO_4)_2}} \times \frac{3\ \cancel{\text{mol}}\ \cancel{BaCl_2}}{1\ \cancel{\text{mol}}\ \cancel{Ba_3(PO_4)_2}} \times \frac{297.13\ \text{g}\ BaCl_2}{1\ \cancel{\text{mol}}\ \cancel{BaCl_2}} = 4.753645607\ \text{g}\ BaBr_2$$

$$\boldsymbol{m_{BaBr_2} = 4.75\ \text{g}}$$

Example 5: Aluminium metal combines with chlorine gas to form a white solid. What is the mass of this product when 3.46×10^{25} molecules of chlorine gas is used for this reaction?

$2\ Al_{(s)}$	+	$3\ Cl_{2\ (g)}$	$\rightarrow$	$2\ AlCl_{3\ (g)}$
		3.46×10^{25} molecules		? g
		6.022×10^{23} molecules = 1 mol		M = 133.33 g/mol

$$3.46 \times 10^{25}\ \cancel{\text{molecules } Cl_2} \times \frac{1\ \cancel{\text{mol } Cl_2}}{6.022 \times 10^{23}\ \cancel{\text{molecules } Cl_2}} \times \frac{2\ \cancel{\text{mol } AlCl_3}}{3\ \cancel{\text{mol } Cl_2}} \times \frac{133.33\text{ g } AlCl_3}{1\ \cancel{\text{mol } AlCl_3}} = 5107.1\text{ g } AlCl_3$$

$$\boldsymbol{m_{AlCl_3} = 5.11 \times 10^3\text{ g or } 5.11\text{ kg}}$$

Example 6: At room temperature and normal pressure, nitrogen and ammonia have densities of 1.130 g/L and 0.687 g/L. Suppose 400. mL of nitrogen gas is reacted with excess hydrogen under the same conditions. What is the volume of ammonia gas formed from this reaction?

$N_{2\ (g)}$	+	$3\ H_{2\ (g)}$	$\rightarrow$	$2\ NH_{3\ (g)}$
400 mL = 0.400 L				? mL
D = 1.130 g/L				D = 0.687 g/L
M = 28.02 g/mol				M = 17.04 g/mol

$$0.400\ \cancel{\text{L } N_2} \times \frac{1.130\ \cancel{\text{g } N_2}}{1\ \cancel{\text{L } N_2}} \times \frac{1\ \cancel{\text{mol } N_2}}{28.02\ \cancel{\text{g } N_2}} \times \frac{2\ \cancel{\text{mol } NH_3}}{1\ \cancel{\text{mol } N_2}} \times \frac{17.04\ \cancel{\text{g } NH_3}}{1\ \cancel{\text{mol } NH_3}} \times \frac{1\text{ L } NH_3}{0.697\ \cancel{\text{g } NH_3}} = 0.8007985028\text{ L } NH_3$$

$$\boldsymbol{V_{NH_3} = 0.801\text{ L or } 801\text{ mL}}$$

Assignment

9.1 pg. 304 #1 and 2 (Practice); pg. 307 #1 to 4 (Practice); pg. 309 #1 to 4 (Practice); pg. 311 #1 and 2 (Practice); pg. 311 # 1 to 7

9.2: Limiting Reactants and Percentage Yield

Excess Reactant: - the reactant that is not completely used up in the reaction.

Limiting Reactant: - the reactant with the smaller amount (accounting for the mole ratio of the two reactants).
- a limiting reactant will always be completely used up in the reaction.
- if the mass of the product is calculated using the excess reactant, it will always be more than the mass determined using the limiting reactant.
- ***just because the reactant has a smaller mass initially does not mean it is a limiting reactant***.
- the mass of the product calculated from the limiting reactant is referred to as the ***theoretical yield*** (because it is the amount that the reaction ***should*** produce).
- the theoretical yield can be determined by calculation without the need of experimentation.

Note: A limiting reagent question will always have enough information to find the moles of both reactants.

Actual Yield: - the mass of the product that was actually produced in the lab. It is also referred to as the ***Experimental Yield***.

Steps to deal with Limiting Reagent Problems: (the quantities of both reactants are given in the question)

1. Calculate the mass of the product using one reactant using the stoichiometric method.
2. Calculate the mass of the product using the other reactant using the stoichiometric method.
3. **The reactant that generates the smaller product mass is the limiting reactant**. That product mass calculated is the ***theoretical yield***.
4. The ***other reactant that gives a larger product mass*** is the ***excess reactant***.
5. If the question provides the actual yield, then determine the percentage yield and / or percentage error.

$$\boxed{\%\ \text{Yield} = \frac{\text{Actual}}{\text{Theoretical}} \times 100\% \qquad \%\ \text{Error} = \frac{|\text{Theoretical} - \text{Actual}|}{\text{Theoretical}} \times 100\%}$$

Example 1: 5.00 g of phosphorus is reacted with 15.00 g of chlorine gas to produce phosphorus trichloride.

a. Determine the theoretical yield of the product produced and identify the limiting and excess reactant.

$P_{4\,(s)}$	+	$6\,Cl_{2\,(g)}$	$\rightarrow$	$4\,PCl_{3\,(s)}$
5.00 g		15.00 g		? g
M = 123.88 g/mol		M = 70.90 g/mol		M = 137.32 g/mol

Since there is enough information to determine the moles of two reactants (quantities of both reactants are given), we need to find the mass of the product from each of these reactant before labelling which reactant is limiting.

① $5.00\ \cancel{g}\ \cancel{P_4} \times \frac{1\ \cancel{mol}\ \cancel{P_4}}{123.88\ \cancel{g}\ \cancel{P_4}} \times \frac{4\ \cancel{mol}\ \cancel{PCl_3}}{1\ \cancel{mol}\ \cancel{P_4}} \times \frac{137.32\ g\ PCl_3}{1\ \cancel{mol}\ \cancel{PCl_3}} = 22.2\ g\ \ PCl_3$

② $15.00\ g\ Cl_2 \times \frac{1\ \cancel{mol}\ \cancel{Cl_2}}{70.90\ g\ \cancel{Cl_2}} \times \frac{4\ \cancel{mol}\ \cancel{PCl_3}}{6\ \cancel{mol}\ \cancel{Cl_2}} \times \frac{137.32\ g\ PCl_3}{1\ \cancel{mol}\ \cancel{PCl_3}} =$ **19.4 g PCl_3** (lesser product mass)

Since Cl_2 gives a smaller calculated product mass, **Cl_2 is the limiting reactant**; **P_4 is the excess reactant**.

b. The actual yield of the product was measured at 17.51 g in an experiment. What are the percentage yield and the percentage error from the lab?

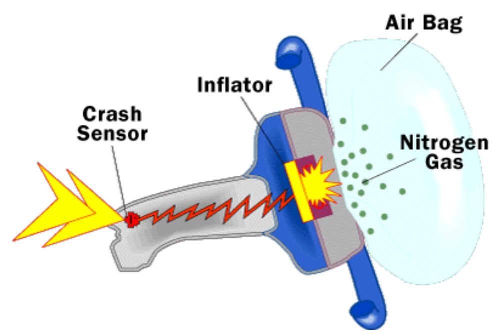
$$\% \text{ Yield} = \frac{17.51\text{ g}}{19.4\text{ g}} \times 100\% \qquad \% \text{ Error} = \frac{|19.4\text{ g} - 17.51\text{ g}|}{19.4\text{ g}} \times 100\%$$

% Yield = 90.3% **% Error = 9.74%**

Assignment

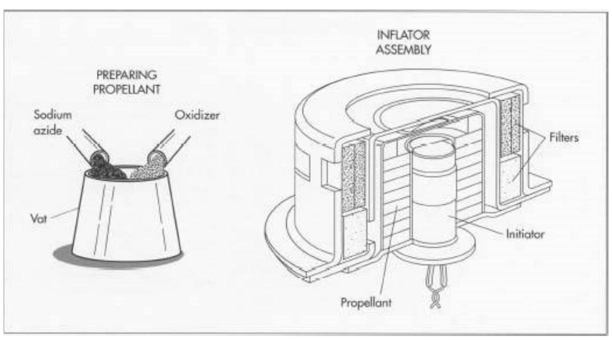
9.2 pg. 314 #1 to 3 (Practice); pg. 317 #1 to 3 (Practice);
pg. 318 #1 to 3 (Practice); pg. 319 #1 and 14

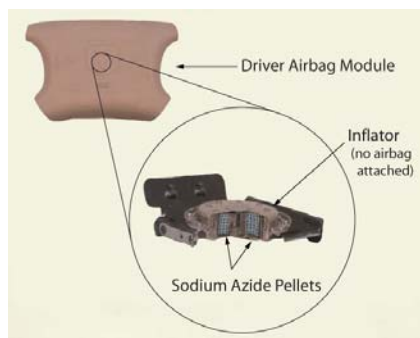
9.3: Stoichiometry of Cars



Air-Bag Chemistry:
- solid sodium azide (NaN_3), an unstable substance, can decompose given enough initial energy into sodium metal and nitrogen gas. Hence allowing the airbag to inflate quickly.

$$2\ NaN_{3\,(s)} \rightarrow 2\ Na_{(s)} + 3\ N_{2\,(g)}$$





Example 1: Suppose 55.8 L of nitrogen gas is needed to inflate an airbag. What mass of sodium azide must be placed inside the airbag column if the density of nitrogen is 0.92 g/L?

$2\ NaN_{3\,(s)}$	$\rightarrow$	$2\ Na_{(s)}$	+	$3\ N_{2\,(g)}$
? g				55.8 L and 0.92 g/L
M = 65.02 g/mol				M = 28.02 g/mol

$$55.85\ \cancel{\text{L}}\ \cancel{N_2} \times \frac{0.92\ \cancel{\text{g}}\ \cancel{N_2}}{1\ \cancel{\text{L}}\ \cancel{N_2}} \times \frac{1\ \cancel{\text{mol}}\ \cancel{N_2}}{28.02\ \cancel{\text{g}}\ \cancel{N_2}} \times \frac{2\ \cancel{\text{mol}}\ \cancel{NaN_3}}{3\ \cancel{\text{mol}}\ \cancel{N_2}} \times \frac{65.02\text{ g } NaN_3}{1\ \cancel{\text{mol}}\ \cancel{NaN_3}} = 79.44475584\text{ g } NaN_3$$

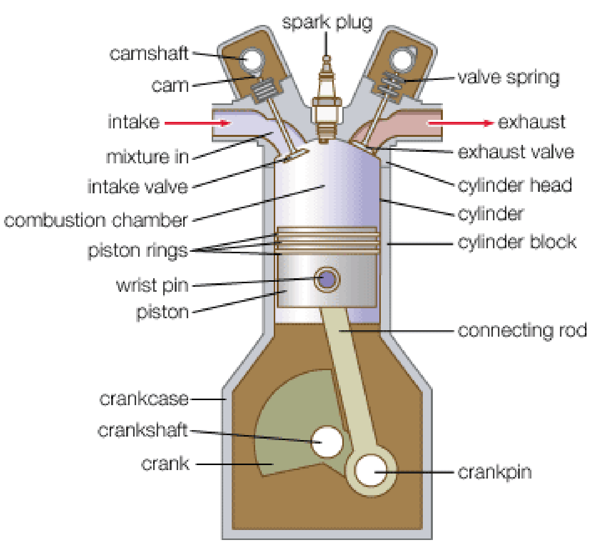
$$m_{NaN_3} = 79\text{ g}$$

Fuel-Oxygen Ratio: - the amount of fuel to amount of oxygen to ensure complete combustion in an engine.
- the optimal fuel ratio for a car engine depends on what the car is doing (see below).

Engine Activity	Fuel Oxygen Mole Ratio	Results
Starting	1: 1.7	Very low engine efficiency Most Harmful Pollutants (NO and CO) produce - incomplete combustion
Idling	1: 7.4	Still low efficiency and Some Other Pollutants (NO and CO) produce
Running at Normal Speed	1: 13.2	Optimal efficiency and Main Exhausts (NO_2, CO_2 and H_2O) produce

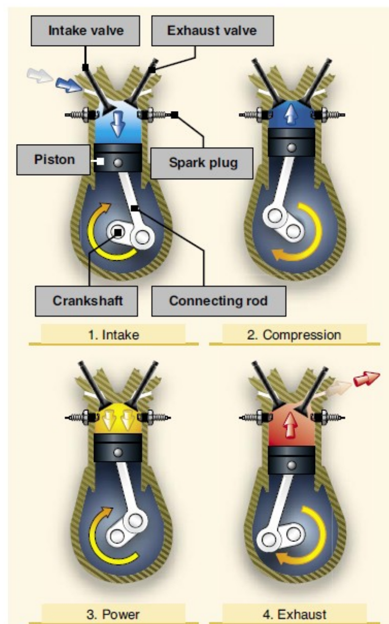
- automobile gasoline is a mixture; the main component of gasoline is isooctane – C_8H_{18} (the other component is heptane – C_7H_{16}).
- a rating of 87 octane means 87/13 isooctane to heptane ratio in the fuel. It also means how much the fuel can be compressed. High performance engine requires a higher compression and hence higher octane rating.
- assuming gasoline is pure octane, the equation for the hydrocarbon combustion is

$$2\ C_8H_{18\,(l)} + 25\ O_{2\,(g)} \rightarrow 16\ CO_{2\,(g)} + 18\ H_2O_{\,(g)}$$



(left) a typical automobile engine cylinder.

(right) the four-stroke cycle of igniting fuel and ejecting exhaust in a cylinder; the crankshaft turns in the direction of the arrow; the spark-plug ignites the fuel during the compression part of the cycle.



Engine Cycle: - a complete turn of events when the cylinder intakes the fuel, then compresses and ignites the fuel. This follows by the power output where the exhaust products (CO_2 and H_2O) push down on the piston. Finally, the exhaust gases are pushed out of the cylinder (see diagram above).
- each four-stroke cycle as described above ***uses up a volume of air equal to that of the size of a cylinder***.
- a typical automobile engine can come in four, six (V-6), or eight (V-8) cylinders. Some very high power engine has twelve cylinders (V-12).

Flooding the Engine: - when there is too much fuel in the cylinder and not enough engine – the engine will stop, and it will not start again until the air intake is reset. Need to open the carburetor and tip the air intake up).

Engine Stall: - when there is no enough fuel in the cylinder and too much air – the engine will act like it is out of gas.

Example 2: A V-8 (eight cylinder) engine burned for 30.0 cycles. Determine the volume of gasoline (isooactane) burned if each cylinder has a 600. mL air capacity.
(Density of isooctane = 0.692 g/mL; Density of oxygen = 1.33 g/L. Air is 21.0% oxygen)

$2\ C_8H_{18\ (l)}$	+	$25\ O_{2\ (g)}$	→	$16\ CO_{2\ (g)}$	+	$18\ H_2O_{(g)}$
? mL		30.0 cycles; 8 cylinders; 600 mL air/cylinder				
0.692 g/mL		1.33 g/L; air = 21.0% O_2				
M = 114.26 g/mol		M = 32.00 g/mol				

First, we have to calculate the volume of oxygen in the cylinder (Volume of Air ≠ Volume of Oxygen)
Note the density of oxygen is given in g/L → we need to convert the volume of oxygen to Litre.

$$30.0\ \cancel{\text{cycles}} \times \frac{8\ \cancel{\text{cylinders}}}{1\ \cancel{\text{cycle}}} \times \frac{600\ \cancel{\text{mL air}}}{1\ \cancel{\text{cylinder}}} \times \frac{21.0\ \text{mL}\ O_2}{100\ \cancel{\text{mL air}}} \times \frac{1\ \text{L}\ O_2}{1000\ \cancel{\text{mL}\ O_2}} = 30.24\ \text{L}\ O_2 \text{ used}$$

Next, we can determine the volume of C_8H_{18} (gasoline) used with regular stoichiometry.

$$30.24\ \cancel{\text{L}\ O_2} \times \frac{1.33\ \cancel{\text{g}\ O_2}}{1\ \cancel{\text{L}\ O_2}} \times \frac{1\ \cancel{\text{mol}\ O_2}}{32.00\ \cancel{\text{g}\ O_2}} \times \frac{2\ \cancel{\text{mol}\ C_8H_{18}}}{25\ \cancel{\text{mol}\ O_2}} \times \frac{114.26\ \cancel{\text{g}\ C_8H_{18}}}{1\ \cancel{\text{mol}\ C_8H_{18}}} \times \frac{1\ \text{mL}\ C_8H_{18}}{0.692\ \cancel{\text{g}\ C_8H_{18}}} = 16.60204405\ \text{mL}$$

$$\boldsymbol{V_{C_8H_{18}} = 16.6\ \text{mL}}$$

Example 3: A V-6 (six cylinder) engine has a individual cylinder air capacity of 500. mL. How many cycles are needed to burn 1.00 gal (3.75 L) of isooctane?
(Density of isooctane = 0.692 g/mL; Density of oxygen = 1.33 g/L. Air is 21.0% oxygen)

$2\ C_8H_{18\ (l)}$	+	$25\ O_{2\ (g)}$	→	$16\ CO_{2\ (g)}$	+	$18\ H_2O_{(g)}$
3.75 L = 3750 mL		?? cycles; 6 cylinders				
0.692 g/mL		500 mL air/cylinder; air = 21.0% O_2				
M = 114.26 g/mol		1.33 g/L; M = 32.00 g/mol				

First, we have to calculate the volume of oxygen used with regular stoichiometry. Density of C_8H_{18} was given in g/mL → need to convert 1.00 L C_8H_{18} to mL.

$$3750\ \cancel{\text{mL}\ C_8H_{18}} \times \frac{0.692\ \cancel{\text{g}\ C_8H_{18}}}{1\ \cancel{\text{mL}\ C_8H_{18}}} \times \frac{1\ \cancel{\text{mol}\ C_8H_{18}}}{114.26\ \cancel{\text{g}\ C_8H_{18}}} \times \frac{25\ \cancel{\text{mol}\ O_2}}{2\ \cancel{\text{mol}\ C_8H_{18}}} \times \frac{32.00\ \cancel{\text{g}\ O_2}}{1\ \cancel{\text{mol}\ O_2}} \times \frac{1\ \cancel{\text{L}\ O_2}}{1.33\ \cancel{\text{g}\ O_2}} = 6830.484227\ \text{L}\ O_2$$

Next, we can determine the number of cycles after converting volume of oxygen to volume of air.
(Volume of Air ≠ Volume of Oxygen)

$$6830.484227\ \cancel{\text{L}\ O_2} \times \frac{100\ \cancel{\text{L air}}}{21.0\ \cancel{\text{L}\ O_2}} \times \frac{1000\ \cancel{\text{mL air}}}{1\ \cancel{\text{L air}}} \times \frac{1\ \cancel{\text{cylinder}}}{500\ \cancel{\text{mL air}}} \times \frac{1\ \text{cycle}}{6\ \cancel{\text{cylinders}}} = 10842.03846\ \text{cycles}$$

1.08×10^4 cycles ≈ 11 thousands cycles

Incomplete Combustion: - when the fuel-oxygen ratio is low (oxygen is the limiting reactant), the product of the combustion can be a mixture $CO_{(g)}$, $C_{(s)}$ with $H_2O_{(g)}$. Carbon monoxide is a colourless and odourless gas. It is highly toxic because it inhibits haemoglobin in the blood to deliver oxygen to cells.

$$2\ C_8H_{18\ (l)} + 13\ O_{2\ (g)} \rightarrow 8\ CO_{(g)} + 8\ C_{(s)} + 18\ H_2O_{(g)}$$

Photochemical Smog: - under the high temperature in the engine, nitrogen in the air can combine with oxygen to form $NO_{(g)}$ and $NO_{2\,(g)}$ (photochemical smog).

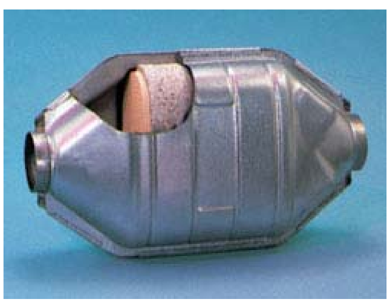
$$N_{2\,(g)} + O_{2\,(g)} \rightarrow NO_{(g)}$$
$$2\,NO_{(g)} + O_{2\,(g)} \rightarrow 2\,NO_{2\,(g)}$$

- nitrogen dioxide can further react with oxygen under sunlight to form ground-level ozone (when inhale, it can react with cholesterol to form plague, thereby increasing the chance of a heart-attack). Ozone is also considered as a constituent of photochemical smog.

$$NO_{2\,(g)} + O_{2\,(g)} \rightarrow NO_{(g)} + O_{3\,(g)}$$

(above) Smog over the Los Angeles skyline

Catalytic Converter: - converts $NO_{(g)}$ (result of burning nitrogen at high temperature) to $N_{2\,(g)}$ and $O_{2\,(g)}$ in a relatively short amount of time. Oxygen gas along with the catalytic converter is also used to produce $CO_{2\,(g)}$ from $CO_{(g)}$.



$$2\,NO_{(g)} \xrightarrow{\text{catalytic converter}} N_{2\,(g)} + O_{2\,(g)}$$
$$2\,CO_{(g)} + O_{2\,(g)} \xrightarrow{\text{catalytic converter}} 2\,CO_{2\,(g)}$$

A catalytic converter for most modern vehicle. Leaded gasoline deactivates catalytic converter. Therefore, they are not legally used in vehicles

Example 4: Calculate the mass of carbon monoxide formed during an incomplete combustion of 1 gallon (3.75 L) of isooctane, $C_8H_{18\,(l)}$. (Density of isooctane = 0.692 g/mL)

$$2\,C_8H_{18\,(l)} + 13\,O_{2\,(g)} \rightarrow 8\,CO_{(g)} + 8\,C_{(s)} + 18\,H_2O_{(g)}$$

C_8H_{18}: 3.75 L; 0.692 g/mL; M = 114.26 g/mol

CO: ? g; M = 28.01 g/mol

$$3.75\ \cancel{\text{L}\ C_8H_{18}} \times \frac{1000\ \cancel{\text{mL}}\ \cancel{C_8H_{18}}}{1\ \cancel{\text{L}}\ \cancel{C_8H_{18}}} \times \frac{0.692\ \cancel{\text{g}}\ \cancel{C_8H_{18}}}{1\ \cancel{\text{mL}}\ \cancel{C_8H_{18}}} \times \frac{1\ \cancel{\text{mol}}\ \cancel{C_8H_{18}}}{114.26\ \cancel{\text{g}}\ \cancel{C_8H_{18}}} \times \frac{8\ \cancel{\text{mol}}\ \cancel{CO}}{2\ \cancel{\text{mol}}\ \cancel{C_8H_{18}}} \times \frac{28.01\ \text{g CO}}{1\ \cancel{\text{mol}}\ \cancel{CO}}$$

$$= 2544.580781\ \text{g CO}$$

$$\boldsymbol{m_{CO} = 2.54 \times 10^3\ \text{g or } 2.54\ \text{kg}}$$

Assignment

9.3 pg. 322 #1 to 4 (Practice); pg. 324 #1 to 3 (Practice);
pg. 327 #1 (Practice); pg. 327 #1 to 9
Chapter 9 Review pg. 329−332 #21 to 49 (odd); Optional: #22 to 50 (even)