Moles, reactions and stoichiometry.

1. Balance these following reactions:
   a. $NaMnO_4$ (s) → Mn_2O_7 (s) + Na_2O (s)
   b. $Fe_2(CO_3)_3$ (aq) + NaOH (aq) → FeO(OH) (s) + Na_2CO_3 (aq) + H_2O (l)
   c. Aqueous sodium sulfate reacts with solid copper (I) bromide in a double displacement reaction.
   d. The combustion of $C_9H_{15}O$.

2. How many nitrogen atoms are in 42.86 moles?

3. Explain why 1 mole of CO_2 has a different mass than 1 moles of carbon monoxide.

4. What is the mass of 1.86×10^{20} CO_2 molecules?

5. Elemental analysis determines that a sample of $NaMnO_4$ contains 90.1853 grams of oxygen. What is the mass of the sample?

6. Consider the synthesis of water from H_2 and O_2 (2 H_2 (g) + O_2 (g) → 2 H_2O (l).
   a. If you start with 5 grams of H_2 and 50 grams of O_2, how much water will be formed (assume 100% yield)?
   b. How much of each reactant is left over when the reaction completes?
   c. Determine the mass of water produced if this reaction proceeds with a 73.86 % yield.

Quinine was the first effective "Western" treatment for malaria. It was first used in the 17^{th} century and remained in use until the 1940s when other effective drugs were discovered that have fewer unpleasant side effects.

7. Elemental analysis of quinine indicates that is 74.03% carbon, 7.471% hydrogen, 8.636% nitrogen, and 9.863 % oxygen (these are all mass percent values) and has a molecular weight of 324.46 g mol^{-1}. Determine the molecular formula of quinine.

8. Combustion of quinine yields the toxic by-product NO_2 (g).
   a. Determine the mass of NO_2 produced if 100 grams of quinine is combusted in excess oxygen.
   b. How much of each reactant is needed to produce 250 grams of water?
   c. If 250 grams of CO_2 is produced when 250 grams of each reactant are combined, determine the % yield.
   d. Determine the mass of **all compounds** left in a flask after the combustion of 25 grams of quinine in 50 grams of oxygen.

① a. $2NaMnO_4(s) \rightarrow Mn_2O_7(s) + Na_2O(s)$

b. $Fe_2(CO_3)_3(aq) + 6\,NaOH(aq) \rightarrow 2\,FeO(OH)(s) + 3Na_2CO_3(aq) + 2H_2O(l)$

c. $Na_2SO_4(aq) + 2CuBr(s) \rightarrow Cu_2SO_4(s) + 2NaBr(aq)$

d. $4\,C_9H_{15}O + 49O_2(g) \rightarrow 36\,CO_2 + 30\,H_2O$

② $\dfrac{42.86\text{ mol} \mid 6.022\times10^{23}\text{ atom}}{1\text{ mol}} = 2.58\times10^{25}$

③ CO_2 contains more atoms than CO. For this reason, 1 mole of CO_2 has more mass than one mol of CO

④ $\dfrac{1.86\times10^{20}\text{ molecules}}{} \Big| \dfrac{1\text{ mol}}{6.022\times10^{23}\text{ molecules}} \Big| \dfrac{44.01\text{ g}}{1\text{ mol}} = 1.359\times10^{-2}\text{ g}$

⑤ $NaMnO_4 \rightarrow MW = 141.93\text{ g/mol}$

$\dfrac{90.1853\text{ g O}}{} \Big| \dfrac{1\text{ mol O}}{16\text{ g O}} \Big| \dfrac{1\text{ mol }NaMnO_4}{4\text{ mol O}} \Big| \dfrac{141.93\text{ g }NaMnO_4}{1\text{ mol }NaMnO_4} = \boxed{200\text{ g}}$

⑥ a. $\dfrac{5\text{ g }H_2}{} \Big| \dfrac{\text{mol}}{2.02\text{ g}} \Big| \dfrac{2\text{ mol }H_2O}{2\text{ mol }H_2} \Big| \dfrac{18.02\text{ g }H_2O}{1\text{ mol }H_2O} = 44.6\text{ g }H_2O$

$\dfrac{50\text{ g }O_2}{} \Big| \dfrac{1\text{ mol}}{32\text{ g }O_2} \Big| \dfrac{2\text{ mol }H_2O}{1\text{ mol }O_2} \Big| \dfrac{18.02\text{ g }H_2O}{1\text{ mol }H_2O} = 56.3\text{ g }H_2O$

only 44.6 g is made (H_2 runs out)

ⓑ H_2 is Limiting Reactant, so NONE left

O_2: 50 g @ beginning. How much used to make 44.6 g H_2O?

$\dfrac{44.6\text{ g }H_2O}{} \Big| \dfrac{1\text{ mol}}{18.02\text{ g}} \Big| \dfrac{1\text{ mol }O_2}{2\text{ mol }H_2O} \Big| \dfrac{32\text{ g}}{\text{mol}} = 39.6\text{ g}$

left $= 50\text{ g} - 39.6\text{ g} = \boxed{10.396\text{ g }O_2}$

(c) Theoretical yield = 44.6 g

$$73.86 = \frac{\text{Actual yield}}{44.6\ g} \times 100$$

Yield = 32.94 g

7. Assume 100 g total → this way mass % = mass

$$\frac{74.03\ g\ C}{} \Big| \frac{1\ mol}{12.01\ g} = 6.16\ mol\ C \div 0.616 = 10$$

$$\frac{7.471\ g\ H}{} \Big| \frac{mol}{1.01\ g} = 7.397\ mol\ H \div 0.616 = 12$$

$$\frac{8.636\ g\ N}{} \Big| \frac{1\ mol}{14.01\ g} = 0.616\ mol\ N \div 0.616 = 1$$

$$\frac{9.863\ g\ O}{} \Big| \frac{1\ mol}{16\ g} = 0.616\ mol\ O \div 0.616 = 1$$

Empirical Formula

$C_{10}H_{12}NO$

E.F.W. = 10(12.01) + 12(1.01) + 14.01 + 16 = 162.23

M.F.W. = 324.46

$$\frac{324.46}{162.23} = 2$$

M.F.

$C_{20}H_{24}N_2O_2$

8. $C_{20}H_{24}N_2O_2 + 27O_2 \longrightarrow 20\ CO_2 + 12\ H_2O + 2NO_2$

a) Excess O_2 ... so quinine is L.R.

$$\frac{100\ g\ Q}{} \Big| \frac{mol}{324.46\ g} \Big| \frac{2\ mol\ NO_2}{1\ mol\ Q} \Big| \frac{46.01\ g}{1\ mol} = 28.36\ g\ NO_2$$

b) Q: $$\frac{250\ g\ H_2O}{} \Big| \frac{1\ mol}{18.02\ g} \Big| \frac{1\ mol\ Q}{12\ mol\ H_2O} \Big| \frac{324.46\ g}{1\ mol} = 375.12\ g\ \text{Quinine}$$

O_2: $$\frac{250\ g\ H_2O}{} \Big| \frac{1\ mol}{18.02\ g} \Big| \frac{27\ mol\ O_2}{12\ mol\ H_2O} \Big| \frac{32\ g}{1\ mol} = 998.89\ g\ O_2$$

c) 1 determine L.R. → this will establish the theoretical yield

2 determine % yield

$$\frac{250\text{ g } O_2}{} \left| \frac{1\text{ mol}}{32\text{ g}} \right| \frac{20\text{ mol } CO_2}{27\text{ mol } O_2} \left| \frac{44.01\text{ g}}{1\text{ mol}} \right. = 254.688\text{ g } CO_2$$ ← theoretical yield

$$\frac{250\text{ g Q}}{} \left| \frac{1\text{ mol}}{324.46\text{ g}} \right| \frac{20\text{ mol } CO_2}{1\text{ mol Q}} \left| \frac{44.01\text{ g}}{1\text{ mol}} \right. = 678.2\text{ g } CO_2$$

$$\%\text{ yield} = \frac{250}{254.688} \times 100 = 98.2\%$$

(d)

$$\frac{25\text{ g Q}}{} \left| \frac{1\text{ mol}}{324.46\text{ g}} \right| \frac{20\text{ mol } CO_2}{1\text{ mol Q}} \left| \frac{44.01\text{ g}}{1\text{ mol}} \right. = 67.8\text{ g } CO_2$$

$$\frac{50\text{ g } O_2}{} \left| \frac{1\text{ mol}}{32\text{ g}} \right| \frac{20\text{ mol } CO_2}{27\text{ mol } O_2} \left| \frac{44.01\text{ g}}{1\text{ mol}} \right. = \boxed{50.94\text{ g } CO_2}$$

O_2 is L.R.

$\boxed{\emptyset\ O_2\text{ left}}$

$$\frac{50\text{ g } O_2}{} \left| \frac{1\text{ mol}}{32\text{ g}} \right| \frac{24\text{ mol } H_2O}{27\text{ mol } O_2} \left| \frac{18.01\text{ g}}{1\text{ mol}} \right. = \boxed{25.01\text{ g } H_2O}$$

$$\frac{50\text{ g } O_2}{} \left| \frac{1\text{ mol}}{32\text{ g}} \right| \frac{2\text{ mol } NO_2}{27\text{ mol } O_2} \left| \frac{46.01\text{ g}}{1\text{ mol}} \right. = \boxed{5.33\text{ g } NO_2}$$

$$\frac{50\text{ g } O_2}{} \left| \frac{1\text{ mol}}{32\text{ g}} \right| \frac{1\text{ mol Q}}{27\text{ mol } O_2} \left| \frac{324.46\text{ g}}{\text{mol}} \right. = 18.78\text{ g Q used}$$

$$25 - 18.78 = \boxed{6.22\text{ g quinine left}}$$