



Review

Avogadro's # = $N_A = 6.022 \times 10^{23}$ particles/mol = 1 mole

Ex. In 2 mol of CH_4 , there are (is)...

... 6.022×10^{23} molecules CH_4

... 2 mole C = $1 \times (6.022 \times 10^{23})$ atoms C = 6.022×10^{23} atoms C

... 4 mole H = $4 \times (6.022 \times 10^{23})$ atoms H = 2.4088×10^{24} atoms H

Ex. In 3.5 mol SF_6 , there are (is)...

... $3.5 \times (6.022 \times 10^{23})$ molecules $\text{SF}_6 = 2.1077 \times 10^{24}$ molecules SF_6

... 3.5 mole S = $3.5 \times (6.022 \times 10^{23})$ atoms S = 2.1077×10^{24} atoms S

... $3.5 \times 6 = 21$ mol F = $21 \times (6.022 \times 10^{23})$ atoms F = 1.26462×10^{25} atoms F

Ex. In 2 mol of the reaction



there are (is)...

$$\dots 2 \text{ mol rxn} \times \frac{1 \text{ mol C}_6\text{H}_{12}\text{O}_6}{1 \text{ mol rxn}} = 2 \text{ mol C}_6\text{H}_{12}\text{O}_6$$

$$\dots 2 \text{ mol rxn} \times \frac{6 \text{ mol O}_2}{1 \text{ mol rxn}} = 12 \text{ mol O}_2$$

$$\dots 2 \text{ mol rxn} \times \frac{6 \text{ mol H}_2\text{O}}{1 \text{ mol rxn}} = 12 \text{ mol H}_2\text{O}$$

$$\dots 2 \text{ mol rxn} \times \left[\frac{1 \text{ mol C}_6\text{H}_{12}\text{O}_6}{1 \text{ mol rxn}} \times \frac{6 \text{ mol O}}{1 \text{ mol C}_6\text{H}_{12}\text{O}_6} + \frac{6 \text{ mol O}_2}{1 \text{ mol rxn}} \times \frac{2 \text{ mol O}}{1 \text{ mol O}_2} + \frac{6 \text{ mol CO}_2}{1 \text{ mol rxn}} \times \frac{2 \text{ mol O}}{1 \text{ mol CO}_2} + \frac{6 \text{ mol H}_2\text{O}}{1 \text{ mol rxn}} \times \frac{1 \text{ mol O}}{1 \text{ mol H}_2\text{O}} \right]$$

$$= 2 \text{ mol rxn} \times \left[\frac{6 \text{ mol O} + 12 \text{ mol O} + 12 \text{ mol O} + 6 \text{ mol O}}{1 \text{ mol rxn}} \right]$$

$$= 72 \text{ mol O} / \text{mol rxn}$$



Molar Mass = mass of one mole of a substance; usually has units of g/mol

$$\begin{aligned}\text{Ex. molar mass (PCl}_5\text{)} &= 1 \times \text{molar mass (P)} + 5 \times \text{molar mass (Cl)} \\ &= 1(30.974) + 5(35.453) \\ &= 208.239 \text{ g/mol}\end{aligned}$$

Percent Composition: % of total mass contributed by each element in a molecule

$$\% \text{ composition (x)} = \frac{\# \text{ atoms (x) in molecule} \times \text{molar mass (x)}}{\text{molar mass (molecule)}} = \frac{\text{mass (x) in molecule}}{\text{mass (molecule)}}$$

Ex. % composition of CO_2

$$\% (\text{C}) = \frac{1 \text{ atom C}}{1 \text{ molecule CO}_2} = \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \times \frac{12.011 \text{ g C}}{1 \text{ mol C}} \times \frac{1 \text{ mol CO}_2}{44.011 \text{ g CO}_2} = \frac{12.011 \text{ g C}}{44.011 \text{ g CO}_2} = 0.2729$$

$$\% (\text{O}) = \frac{2 \text{ atom O}}{1 \text{ molecule CO}_2} = \frac{2 \text{ mol O}}{1 \text{ mol CO}_2} \times \frac{16 \text{ g O}}{1 \text{ mol O}} \times \frac{1 \text{ mol CO}_2}{44.011 \text{ g CO}_2} = \frac{32 \text{ g O}}{44.011 \text{ g CO}_2} = 0.7271$$

$$= 1 - \% (\text{C}) = 1 - 0.2729 = 0.7271$$

Ex. % composition CaCO_3

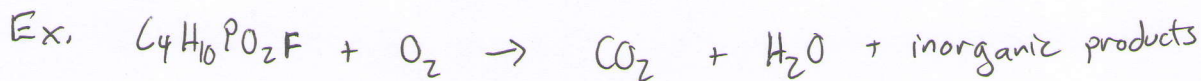
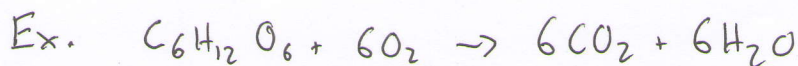
$$\% (\text{Ca}) = \frac{40.08 \text{ g Ca}}{(40.08 + 12.011 + 3 \times 16) \text{ g CaCO}_3} = \frac{40.08 \text{ g Ca}}{100.091 \text{ g CaCO}_3} = 0.4004$$

$$\% (\text{O}) = \frac{(3 \times 16) \text{ g O}}{100.091 \text{ g CaCO}_3} = 0.4796$$

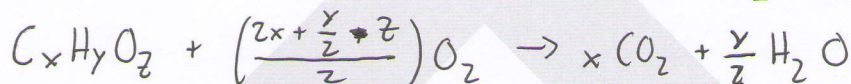
$$\% (\text{C}) = 1 - (0.4004 + 0.4796) = 0.12$$



Combustion Reactions: Burning a compound in O_2 to produce CO_2 , H_2O , and inorganic products

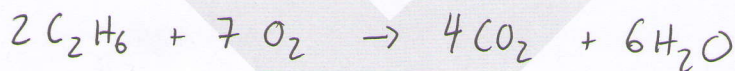
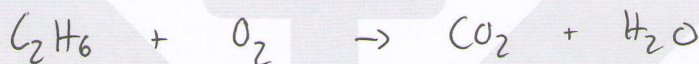


If compound burned only contains C, H, O, we can write and balance a combustion reaction in a general way:

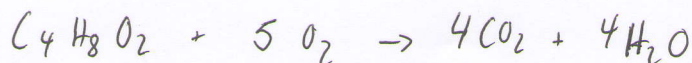


Compounds that only contain C, H, O are called hydrocarbons.

Ex. Write and balance the chemical equation for the combustion of ethane, C_2H_6



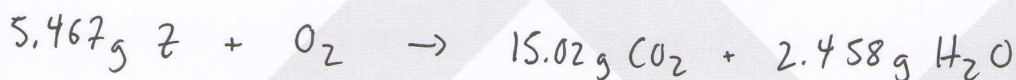
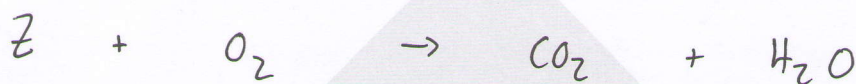
Ex. Write and balance the chemical equation for the combustion of butyric acid, $C_4H_8O_2$





Combustion Analysis: determining the % composition of an unknown substance by combusting it (burning it in O_2)

Ex. 5.467 g of unknown compound Z are burned in excess O_2 , producing 15.02 g of CO_2 and 2.458 g of H_2O . What is the percent composition of Z? Z contains only C, H, and O (it is a hydrocarbon).



$$15.02 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.011 \text{ g } CO_2} \times \frac{1 \text{ mol } C}{1 \text{ mol } CO_2} = \boxed{0.34128 \text{ mol } C \text{ in } Z} \times \frac{12.011 \text{ g } C}{1 \text{ mol } C} = \boxed{4.0991 \text{ g } C \text{ in } Z}$$

$$2.458 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.0158 \text{ g } H_2O} \times \frac{2 \text{ mol } H}{1 \text{ mol } H_2O} = \boxed{0.27287 \text{ mol } H \text{ in } Z} \times \frac{1.0079 \text{ g } H}{1 \text{ mol } H} = \boxed{0.2750 \text{ g } H \text{ in } Z}$$

$$\text{mass O in } Z = 5.467 \text{ g} - (4.0991 \text{ g} + 0.2750 \text{ g}) = \boxed{1.0929 \text{ g } O \text{ in } Z}$$

\downarrow total mass Z \downarrow mass C in Z \downarrow mass H in Z

$$\underline{1.0929 \text{ g } O} \times \frac{1 \text{ mol } O}{16 \text{ g } O} = \boxed{0.0683 \text{ mol } O \text{ in } Z}$$

$$\% (C) = 4.0991 \text{ g} / 5.467 \text{ g} = 0.7498 \bullet$$

$$\% (H) = 0.2750 \text{ g} / 5.467 \text{ g} = 0.0503 \bullet$$

$$\% (O) = 1.0929 \text{ g} / 5.467 \text{ g} = 0.1999 \bullet$$

-or-

$$C_{0.34128} H_{0.27287} O_{0.0683} \Rightarrow C_{4.997} H_{3.995} O_1 \Rightarrow C_5 H_4 O$$



Do we know for sure that the unknown Z has the formula C_5H_4O ?

What is the % composition of $C_{10}H_8O_2$?

$$\% (C) = (10 \times 12.011) / (10 \times 12.011 + 8 \times 1.0079 + 2 \times 16) = 0.7455 \approx 0.7498$$

$$\% (H) = (8 \times 1.0079) / 161.122069 = 0.05004 \approx 0.0503$$

$$\% (O) = 1 - (0.7455 + 0.05004) = 0.20446 \approx 0.1999$$

remember, we rounded
 $C_{4.997}H_{3.995}O_1$
to get
 $C_5H_4O_1$

How do we know whether Z was C_5H_4O or $C_{10}H_8O_2$ or ... $C_{5n}H_{4n}O_n$?

Does it matter?

* $CH_2O \Rightarrow$ formaldehyde \Rightarrow gas, colorless, smelly, highly toxic

* $C_2H_4O_2 \Rightarrow$ acetic acid \Rightarrow vinegar, (liquid), clear, tasty

* $C_3H_6O_2 \Rightarrow$ lactic acid \Rightarrow solid, important in anaerobic exercise

* $C_6H_{12}O_2 \Rightarrow$ glucose \Rightarrow sugar, sweet, safe, solid, white

Empirical formula: the formula of a compound with the smallest whole-number ratio of elements in that compound

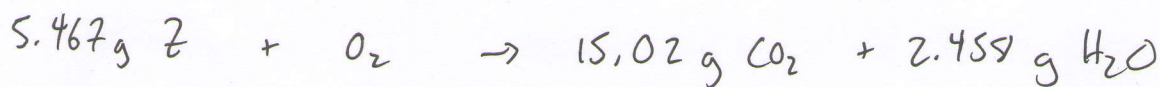
~~Formula~~

Molecular Formula: the actual formula of a compound, giving the number of atoms of each element present in one molecule of that compound

Ex.	<u>Molecular Formula</u>		<u>Empirical Formula</u>
	CH_2O	\longrightarrow	$(CH_2O)_1 \longrightarrow CH_2O$
	$C_2H_4O_2$	\longrightarrow	$(CH_2O)_2 \longrightarrow CH_2O$
	$C_3H_6O_2$	\longrightarrow	$(CH_2O)_3 \longrightarrow CH_2O$
	$C_6H_{12}O_2$	\longrightarrow	$(CH_2O)_6 \longrightarrow CH_2O$



Back to our combustion analysis...



What is n ? To ~~know~~ find n , we must know either

- 1) the molar mass of Z , or
- 2) the number of moles of Z combusted (which gives us the molar mass of Z)

1) Assume the molar mass of $Z = 241.683 \text{ g/mol}$. What is the molecular formula of Z ?

$$\frac{\text{molar mass } Z}{\text{molar mass } C_5H_4O} = \frac{241.683 \text{ g/mol}}{80.561 \text{ g/mol}} \approx 3 = n$$

The molecular formula of Z is $(C_5H_4O)_3 \Rightarrow C_{15}H_{12}O_3$

2) Assume ~~0.034~~ 0.034 mol of Z were combusted in the reaction above. What is the molecular formula of Z ?

$$\frac{5.467 \text{ g } Z}{0.034 \text{ mol } Z} = 160.794 \text{ g/mol of } Z = \text{molar mass of } Z$$

$$\frac{\text{molar mass } Z}{\text{molar mass } C_5H_4O} = \frac{160.794 \text{ g/mol}}{80.561 \text{ g/mol}} \approx 2 = n$$

The molecular formula of Z is $(C_5H_4O)_2 \Rightarrow C_{10}H_8O_2$