General and Organic Chemistry

Lecture 3

Ch3: Stoichiometry

22/10/2024
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General and Organic Chemistry : Ch3 : Stoichiometry
L3

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* Stoichiometry : The study of quantities of materials consumed and produced in chemical reactions
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- objects behave as though they were all identical, just like counting how many objects you have using weight, you assume their weight to be the same

- Atoms are too small to count
- to ~~count~~ count, we need the average mass of an object

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* Atomic Masses :

* ^{12}C is the standard for atomic mass, having a mass of exactly 12 atomic mass units
  - we used it bc its the most common, and have high abundance naturally -

- the masses of all other atoms are given relative to this standard

- Element occur in nature as a mixture of isotopes
  ex: Carbon = 98.89% ^{12}C
  1.11% ^{13}C
  < 0.01% ^{14}C

So for carbon we can't just say its atomic mass is 12 bc its a mixture of ^{12}C, ^{13}C, ^{14}C, so we calculate the average atomic mass :

formula :

$$\left(\underset{\text{Percentage, turn into decimal}}{\text{Isotope}_1 \text{ natural abundance}}\right) \cdot \left(\text{atomic mass of Iso}_1\right) + \left(\text{Isotope}_2 \text{ natural abundance}\right) \cdot \left(\text{atomic mass of Iso}_n\right) \ldots$$

so for carbon :

$$\overset{^{12}C}{(0.9889)}(12u) + \overset{^{13}C}{(0.0111)}(13.0034u) = \underline{\underline{12.01}}\ u$$

^{14}C is so small its negligible

- Even though natural Carbon does <u>NOT</u> Contain a single atom with mass 12.01, but for stoichiometric purposes we can <u>Consider</u> Carbon to be <u>composed of only one type of atom</u> with 12.01 mass

↳ this enables us to count atoms of natural Carbon by weighing a Sample of carbon
[Just like we are working with one type of atom instead of multiple [Isotopes]]

* and for determining the number or Count of isotopes we use the $\boxed{\text{Mass Spectrometer}}$ device

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* The Mole: It's the number equal to the number of carbon <u>atoms</u> in exactly <u>12</u> grams of Pure ^{12}C

So: 1 mole of something was found to Consist of $\underbrace{6.022 \times 10^{23}}_{\text{Avogadro's number}}$ units of that substance

ex: 1 mole (mol) of Eggs = 6.022×10^{23} Eggs

* yes 1 mole of anything is equal to 6.022×10^{23} of units of that thing but that doesn't mean it has the Same mass to another 1 mole of a different thing *

* Just like how a dozen is = 12 So
a dozen Eggs = 12 Eggs
a dozen calculators = 12 calculators
a Mole = 6.022×10^{23} of something

★ Molar Mass: Mass in grams of 1 Mole of the Substance

$$\underset{\text{moles}}{n} = \frac{m \rightarrow \text{mass}}{Mr \rightarrow \text{Molar mass}}$$

$$\boxed{Mr = \frac{m}{n}}$$

ex: - Mr of N = 14.01 mol

- Mr of H_2O is Mr of H_2 + Mr of O

$(2 \times \underbrace{1.008g}_{Mr\,H}) + 16.00g$

$\boxed{= 18.02\ g/mol}$

Solve Exercises

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*Problem Solving ~~s~~ stratigies:

1- Read the Problem and see whats the goal -

2- work/thing from the final goal backwards to know where to start

3- Solve .... Does the answer make sense?

Ex: 100g samples of each [H_2O, CO_2, ~~H_2O~~ $C_3H_6O_2$, N_2O]

Rank from greatest to least number of O atoms

↳ ① goal

② we start by finding how many atoms of O are in each:

100g H_2O → How many Moles of H_2O are in 100g?? → Moles to atoms

$100g\cancel{H_2O} \cdot \frac{1\,Mol\,H_2O}{18.02\,g\,\cancel{H_2O}} = \boxed{5\,Moles\ of\ H_2O}$ → $5\,\cancel{mole\,H_2O} \cdot \frac{6.022\times10^{23}\,molecules}{1\,\cancel{mole\,H_2O}}$

$= 3.011 \times 10^{24}$ Molecules → atoms → $3.011\times10^{24}\,\cancel{\text{molecule } H_2O} \cdot \frac{1\,atom\,O}{1\,\cancel{H_2O\,molecule}}$

↳ This is how many molecules of H_2O are in 5 mol and not how many atoms of H or O

So there is 3.011×10^{24} O atoms in 100g of H_2O

③

- Now we do it for N_2O :

$Mr(N) = 14.01$ g/mol

$Mr(O) = 16.00$ g/mol

$Mr\ N_2O = 2(14.01) + 16.00$

$= \underline{\underline{44.02}}$ g/mol

100g N_2O $\xrightarrow{1}$ How many moles does it have? $\xrightarrow{2}$ How many N_2O molecules are in 100g of it

$\xrightarrow{3}$ How many O atoms are in the molecules?

to convert to mole we multiply with it

① $100g\ \cancel{N_2O} \cdot \boxed{\dfrac{1\ \text{Mole}\ N_2O}{44.02\ g\ \cancel{N_2O}}} \simeq \boxed{2\ \text{mol}\ N_2O}$

② $2\ \cancel{\text{mol}\ N_2O} \cdot \dfrac{6.022 \times 10^{23}\ \text{molecules}\ N_2O}{1\ \cancel{\text{mol}\ N_2O}} = \boxed{1.2 \times 10^{24}\ \text{Molecules}\ N_2O}$

③ $\dfrac{1.2 \times 10^{24}}{\cancel{\text{molecule}}\ N_2O} \cdot \dfrac{1\ O\ \text{atom}^{*}}{1\ \cancel{N_2O\ \text{molecule}}} = \boxed{1.2 \times 10^{24}\ O\ \text{atoms}}$

* $\dfrac{1\ O\ \text{atom}}{1\ N_2O\ \text{molecule}}$ ⇒ means that there is 1 atom of O in every 1 N_2O molecule

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we do the same for the rest

Solution : H_2O

CO_2

$C_3H_6O_2$

N_2O

No of O atoms ↑

★ Percent Composition of Compounds:

- Mass percent of an element:

$$\text{Mass\%} = \frac{\text{mass of element in compound}}{\text{Mass of whole compound}} \cdot 100\%$$

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ex: 100 g samples [H_2O, CO_2, N_2O, $C_3H_6O_2$]
rank from greatest to least Percent oxygen by mass
↳ goal

① H_2O ~~[illegible]~~ Find O

$$O\% = \frac{16.00}{16.00 + 2(1.008)} \cdot 100\% = 88.8\% \text{ oxygen in } H_2O$$

because its a percent that means its value is fixed no matter the volume of ~~sides~~ ~~[illegible]~~ the compound
So the anser is the same still weather its 10 g H_2O
- 20 g -- 1000000 g H_2O — always 88.8% will be oxygen in (H_2O)

- we do the same for the rest:

H_2O
CO_2 ↑ O%
$C_3H_6O_2$
N_2O

★ Determining the formula of a Compound:

Formulas

Empirical Formula

↓

It only shows:

- Type of atoms in the Compound
- The simplest (whole number) mole ratio

Ex: C_2H_4

↓

the empirical formula of this compound is:

CH_2

shows type of atoms C, H

- the simplest mole ratio

$\frac{C_2H_4}{2 \;\; 2} = CH_2$

we can't divide more bc we need whole numbers

Molecular Formula

↓

It shows:

- Type of atoms in Compound
- Their actual mol ratio
- arrangement of atoms in 3D space

Ex: C_6H_6

C_2H_4

⋮

★ So:

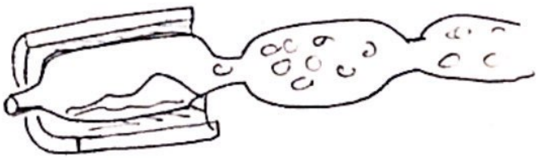
$$\text{Molecular Formula} = (\text{Empirical Formula})n$$

n: an integer [1, 2, 3 ...]

M.F = E.F when $\boxed{n=1}$

Ex: $\boxed{CH_4}$ this is both Molecular and Empirical

- we use a Combustion احتراق Device to:
- Determine the Mass Percent of each Element in a Compound



* Excersise on formulas :

- Compound with 146 g/mol Molar mass
  49.3% is C
  6.9% is H
  43.8% is O
  by mass percentage

Find Empirical Formula (E.F) and Molecular Formula (M.F) :

Mr of :
$C = 12.01$ g/mol
$H = 1.008$ g/mol
$O = 16.00$ g/mol

$$\text{mass\%} = \frac{\text{mass of element}}{\text{Mass of compound}} \cdot 100\%$$

$$C : 49.3\% \times 146 = \frac{49.3}{100} \times 146 = \boxed{71.978 \text{ g/mol}}$$

$(\text{Mr of } C)n = 71.978$
$(12.01)n = 71.978$
$\boxed{n = 6}$ so C_6

Same for the rest :

$$H : \frac{6.9}{100} \times 146 = 10.074 \text{ g/mol}$$

$(1.008)n = 10.074$
$\boxed{n = 10}$ C_6H_{10}

Same
:
:

Formula : $\boxed{C_6H_{10}O_4}$ → this is M.F bc we can still divide to get E.F

$$C_{\frac{6}{2}}H_{\frac{10}{2}}O_{\frac{4}{2}}$$

$\boxed{C_3H_5O_2}$ → this is now the E.F