General and Organic Chemistry

Lecture 4

L4 : CH_3 Continuation
Stoichiometry

23/oct/2024

* General and Organic Chemistry, Ch3: Stoichiometry
Continuation
L4

☆ Chemical Equations:

- A chemical equation consists of: Reactants and Products

on the left side of the arrow ← Reactants and Products → on the right side of the arrow

Ex: $\underbrace{A + B}_{\text{Reactants}} \longrightarrow \underbrace{C + D}_{\text{Products}}$

A representation of a chemical reaction:

$$\overbrace{C_2H_5OH + \underline{3}O_2}^{\text{reactants}} \longrightarrow \overbrace{\underline{2}CO_2 + \underline{3}H_2O}^{\text{Products}}$$

These are moles number that were added to balance the equation

☆ Note: $\underline{3}O_2$ → subscript NOT changeable
we can change the # on the side but not the subscript in order to balance the equation

- The Equation is balanced
[there are the same number of atoms on each side]

reactants:	Products:
2C	2C
6H	6H
7O	7O

☆ In this Equation: 1 mole of Ethanol [C_2H_5OH] reacts with 3 moles of oxygen [O_2] to produce 2 moles of Carbon dioxide [CO_2] and 3 moles of water [H_2O]

- This (the balanced equation) represents an overall ratio of reactants and products, not what "actually" happens during a reaction

- we use the coefficients in the balanced equation to decide the amount of each reactant that is used and the amount of product that is formed
↳ The numbers on the side Ex: $\underline{3}O_2$

* writing and balancing the Equation for a Chemical reaction :

1. Determine what reaction is occurring
   - what are the reactants
   - what are the products
   - what physical states are involved [Solid, Gas, Liquid - -]

2. write the unbalanced equation as a whole then balance it by changing the coefficients (numbers on side) so that there are equal number of atoms of each type on both sides

Ex: $\underbrace{CaO + C}_{\text{reactants}} \longrightarrow \underbrace{CaC_2 + CO_2}_{\text{Products}}$ — unbalanced :

	r		p
	C = 1	$\neq$	C = 3
	Ca = 1	=	Ca = 1
	O = 1	$\neq$	O = 2

↓ balancing

$2CaO + 5C \longrightarrow 2CaC_2 + CO_2$

	r		p
Now:	C = 5	=	C = 5
	Ca = 2	=	Ca = 2
	O = 2	=	O = 2

*Note : doing CaO_2 to balance is incorrect, we can't change the subscript, that will change the whole chemical compound

- In balanced equations, the coefficients (ex: 2CaO) can be fractions but usually they are given as lowest integer multiplies, bc they tell us ratios

↳ ex: $2CaO + 5C \rightarrow 2CaC_2 + CO_2$

2 Moles of CaO and 5 of C give us 2 Moles CaC_2 and 1 CO_2

↓

All Correct:

This is the same as :

$4CaO + 10C \rightarrow 4CaC_2 + 2CO_2$ ↳ 4 Moles of CaO and 10 of C give 4 Moles CaC_2 and 2 CO_2 [we just multiply by 2]

↓

$CaO + 2.5C \rightarrow CaC_2 + 0.5\,CO_2$ same as : 1 Mole CaO and 2.5 Mole C give 1 Mole CaC_2 and 0.5 CO_2 [divided by 2]

★ Important Concepts :
For Balanced chemical equations :

- The number of Atoms is Conserved but not the number of molecules, thats because a whole molecule will react to make another molecule or compound and not Just stay the same

- The Coefficients tell you how much of each substance you have in Moles as stated in Concept II
  - however they do $\boxed{\text{not}}$ indicate mass ratio of the Substance, they indicate Mole ratios

- Coefficients are there to balance the equation so the sum of them on each side doesn't have to be equal

- Atoms are nither created nor destroyed, rather they are Conserved in a chemical reaction as the law of Conservation of mass states

- Chemical equations can be used to relate the masses of the reacting Chemicals

* To calculate Masses of reactants and Products in reaction:

1. Balance the equation
2. convert the given mass of reactant or product to Moles
3. use the balanced equation to set up the Mole ratios bc thats what an equation deals with (Moles)
4. use the ratios to calculate the # of Moles of the desired reactant or Product
5. if required, convert from Moles to grams

~~Ex: If 6.25g of Phosphorus is burned what Mass of oxygen does it combine with?~~

~~Mr P = 30.97~~
~~Mr O = 16.00~~

$$\cancel{P_4 + 5O_2 \rightarrow 2P_2O_5}$$

~~6.25 g ?.g~~

$$\cancel{6.25\,g\,P \times \frac{1\,Mol\,P}{30.97\,g\,P} = 0.20\,mol}$$

$$\cancel{= 0.20\,Mol\,P}$$

$$\cancel{0.20\,Mol\,P_4 \times \frac{5\,Mol\,O_2}{1\,Mol\,P_4} = 1\,Mol\,O_2}$$

Ex: Considering the reaction:

$Mr\ P = 30.97g$
$Mn\ O = 16.00g$

$Mr\ P_4 = 123.188\ g$
$Mr\ O_2 = 32.00g$

$$\overset{6.25g}{P_4} + \overset{??g}{5O_2} \longrightarrow 2P_2O_5$$

if 6.25g of Phosphorus is burned, what $\boxed{\text{mass}}$ of oxygen does it Combine with

→ we want to convert from grams to Mols so we multiply: 1 Mole of P_4 = 123.19 g P_4

① $6.25\cancel{g P_4} \cdot \dfrac{1\ Mol\ P_4}{123.19\cancel{g P_4}} = 0.05\ Mol\ P_4$

→ we want to convert using the ratio from the Equation, we see that we have a ratio of 5 Mols of oxygen for every 1 Mole of P_4

② $0.05\ \cancel{Mol\ P_4} \cdot \dfrac{5\ Mol\ O_2}{1\ \cancel{Mol\ P_4}} = 0.25\ Mol\ O_2$

③ $0.25\ \cancel{Mol\ O_2} \cdot \dfrac{32.00g\ O_2}{1\ \cancel{Mol\ O_2}} = \boxed{8.07\ g\ O_2}$

↳ we convert back to mass since the question asked for it

Balancing an Equation:

CH_4 reacts with oxygen Producing Carbon dioxide and water

unbalanced (we can see the # of atoms doesn't match)

$$CH_4 + O_2 \longrightarrow CO_2 + H_2O$$

r		p
C=1	=	C=1
H=4	≠	H=2
O=2	≠	O=3

Balanced

$$CH_4 + \underline{2}O_2 \longrightarrow CO_2 + \underline{2}H_2O$$

Now:

r		p
C	=	C = 1
H	=	H = 4
O	=	O = 4

Ex: -Methane (CH_4) reacts with Oxygen to make carbon dioxide and water

-Ammonia (NH_3) reacts with oxygen to make nitrogen monoxide and water

Mr O = 16.00
Mr C = 12.01
Mr N = 14.01
Mr H = 1.008

Mr O_2 = 32
Mr CH_4 = 16.04
Mr NH_3 = 17.034
Mr H_2O = 18.016

$$\overset{1.00g}{CH_4} + 2O_2 \rightarrow CO_2 + \underline{2H_2O}$$

$$\overset{??g}{2NH_3} + \frac{5}{2}O_2 \rightarrow 2NO + \overset{\boxed{0.124\ Mol}}{\underset{\boxed{2.23\ g\ H_2O}}{3H_2O}}$$

what Mass of ammonia would Produce the same amount of water as 1.00g of methane reacting with excess oxygen?

* first we need to know how much water 1.00g Methane Makes:

$$1.00g\ \cancel{CH_4} \cdot \frac{1\ Mole\ CH_4}{16.04\ \cancel{g\ CH_4}} = 0.062\ \cancel{Mol\ CH_4} \cdot \frac{2\ Mol\ H_2O}{1\ \cancel{Mol\ CH_4}} \boxed{= 0.124\ Mol\ H_2O}$$

This means: For every 1 Mol of CH_4 2 Mols of H_2O are produced

$$* \quad 0.124\ \cancel{Mol\ H_2O} \cdot \frac{18.016 g H_2O}{1\ \cancel{Mol\ H_2O}} = \boxed{2.23\ g\ H_2O}$$

$$0.124\ \cancel{Mol\ H_2O} \cdot \frac{2\ Mol\ NH_3}{3\ \cancel{Mol\ H_2O}} = \boxed{0.082\ Mol\ NH_3}$$

$$0.082\ \cancel{Mol\ NH_3} \cdot \frac{17.034 g NH_3}{1\ \cancel{Mol\ NH_3}} = \boxed{1.396\ g\ NH_3}$$

* Limiting Reactants : Its the reactant that runs out first and thus limits the amounts of Products that can be formed

- The goal is to determine which reactant is limiting to calculate the amounts of Products that will be formed

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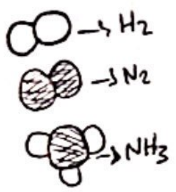
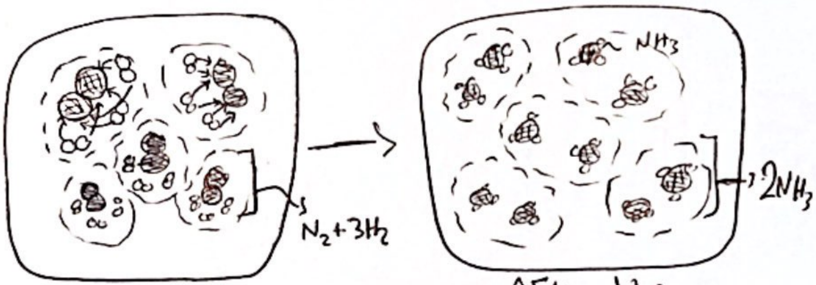
★ To get the Concept : For the reaction

$$[\; N_2 + 3H_2 \rightarrow 2NH_3 \;]$$

"For every Mole of N_2 we need 3 Moles of H_2 to Make 2 Moles of NH_3"

① if we had 5 Moles of N_2 and 15 Moles of H_2 :





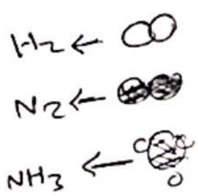
Before the reaction

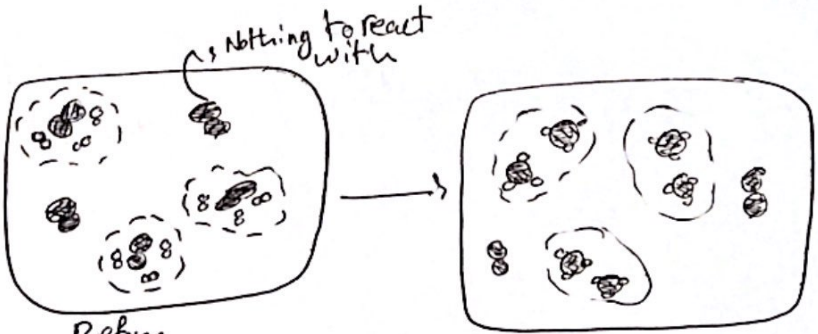
After the reaction

★ we needed 1 Mole N_2 and 3 Moles H_2 to make 2 Mole NH_3
but since we had 5 Moles N_2 that means we need 15 Moles H_2 and make 5 of 2 Mole NH_3 in order to cover the amounts needed and not have leftovers)
(we needed to use all reactants and not have leftovers)

② For that, if we didn't have enough resources :

if we only had 5 Moles N_2 and 9 Moles H_2





Before

After

- Here, bc we didn't have enough H_2 to react with all the N_2 we have some excess N_2 unreacted which makes H_2 a limiting reactant bc it limited and "stopped" the reaction from using all the reactants by not being enough in number

So : [H_2 ran out before all the N_2 reacts]

Ex: which mixture could produce the greatest amount of product?

$[2H_2 + O_2 \rightarrow 2H_2O] \rightarrow$ we need 2 moles of H_2 for every 1 mole of O_2 to make 2 moles of H_2O

① 2 moles H_2 and 2 moles O_2
↳ would make 2 moles H_2O and will be 1 mole O_2 leftover so H_2 is the limiting reactant (ran out first

② 2 moles H_2 and 3 O_2
↳ would make 2 H_2O and would have 2 O_2 left H_2 still limiting

③ 2 H_2 and 1 O_2
↳ would make 2 H_2O with no leftover

④ 3 H_2 and 1 O_2
↳ would make 2 H_2O but with 1 H_2 left so O_2 is the limiting reactant

So here all of them have the same product

* we must think with the ratios and what we have

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* Concept check: A reacts with B, if you react 10.0 g of A with 10.0 g of B
- what is the info you need to determine the <u>mass</u> of product that will be produced

$$\underset{10g}{?A} + \underset{10g}{?B} \longrightarrow \underset{??g}{?X}$$

* we need to know the mole ratio between A and B to see how much we need of both to produce desired amount of the product

* we also need molar masses for A, B and the product

since we are converting:

$$10g\ A \longrightarrow \underbrace{\text{moles}}_{\text{need mole ratio of } A \to X} \longrightarrow \underbrace{\text{mass}}_{\text{finding the product mass}}$$

* So for the Past Concept check, if we had all the numbers:

- You react 10.0 g A with 10.0 g B, what is the mass of the Product if you know that:

M_r of A = 10 g/mol
M_r of B = 20 g/mol
M_r of C = 25 g/mol

and the equation is: $\underbrace{A + 3B}$ $\rightarrow 2C$

↳ we need 1 Mol of A for every 3 Moles of B to Produce 2 Moles of C

$10.0 g A \xrightarrow{\text{to Moles}}$: $\cancel{10.0 g A} \cdot \frac{1 \text{ Mol A}}{\cancel{10 g A}} = \boxed{1 \text{ Mol A}}$

$1 \text{ Mol A} \xrightarrow[\text{of B}]{\text{to Moles}}$ $\cancel{1 \text{ Mol A}} \cdot \frac{3 \text{ Mol B}}{\cancel{1 \text{ Mol A}}} = \boxed{3 \text{ Mol B}}$ ↳ we need this much

the amount of B Moles we have : $\cancel{10.0 g B} \cdot \frac{1 \text{ Mole B}}{\cancel{20 g B}} = \boxed{0.5 \text{ Mol B}}$ ↳ we only have this much

* From the values we have, we can see that B is the limiting reactant bc it runs out first so, the amount of B is what determines how much Product we have so we use the limiting reactant to calculate the rest

$\frac{2 \text{ Mol C}}{3 \cancel{\text{ Mol B}}} \cdot 0.5 \cancel{\text{ Mol B}} = \boxed{\frac{1}{3} \text{ Mol C}}$

$\frac{1}{3} \text{ Mol C} \xrightarrow{\text{grams C}}$ $\frac{1}{3} \cancel{\text{ Mol C}} \cdot \frac{25 g C}{1 \cancel{\text{ Mol C}}} = \boxed{8.33 g C}$

* to know how much of A didn't react:

$\frac{1 \text{ Mol A}}{3 \cancel{\text{ Mol B}}} \cdot 0.5 \cancel{\text{ Mol B}} = \frac{1}{6} \text{ Mol A}$ ↳ reacted

so $\frac{5}{6}$ Mol A is unreacted

Ex:

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

- we need:

1 Mol CH_4 and 2 Mol O_2

to make:

1 Mol CO_2 + 2 Mol H_2O

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answer thes scinarios:

if we are given:

only 1 Mol CH_4 and 1 Mol O_2 : how much Products would we have

	CH_4	$+2O_2$	$\rightarrow CO_2$	$+2H_2O$
Moles before reaction	1	1	$\rightarrow$ 0.00	0.00
Moles After rection	0.5	0.0	$\rightarrow$ 0.5	1

the limiting reactant is Oxygen bc it runs out first

So we use it to determine how much stuff or Prodct we can make

$$\frac{1\,\text{Mol } CH_4}{2\,\cancel{\text{Mol } O_2}} \cdot 1\,\cancel{\text{Mol } O_2} = \frac{1}{2}\,\text{Mol } CH_4$$

$$\frac{1\,\text{Mol } CO_2}{2\,\cancel{\text{Mol } O_2}} \cdot 1\,\cancel{\text{Mol } O_2} = \frac{1}{2}\,\text{Mol } CO_2$$

$$\frac{2\,\text{Mol } H_2O}{2\,\cancel{\text{Mol } O_2}} \cdot 1\,\cancel{\text{Mol } O_2} = 1\,\text{Mol } H_2O$$

here we put the ratio and multiply with what we have:

for CH_4 as example:

for every 1 Mol of CH_4 we should have 2 Mol of O_2 (we get this from the equation)

$$\left(\frac{1\,\text{Mol } CH_4}{2\,\cancel{\text{Mol } O_2}}\right) \cdot 1\,\cancel{\text{Mol } O_2}$$

↳ we multiply with what we have to cacel out factors

$$\boxed{= \frac{1}{2}\,\text{Mol } CH_4}$$

★ The Percent Yield :

- an important indicator of the efficiency of a particular laboratory or industrial reaction

↳ what we got from reaction actually by doing

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

↳ what we are supposed to get based on calculations

★ The actual yield will be very different from the theoretical yield because of a lot of factors:

Ex: $A + B \longrightarrow C + D$

Some of the factors are:

- Having byproducts (side products) so those will take a portion
- also because most reactions are at an equillibrium that means its not just in one direction ($\longrightarrow$)
  [redacted], that means that <u>not</u> <u>all</u> the reactants will turn to products completely, there will be a portion unreacted.

Ex: $P_4 + 6F_2 \rightarrow 4PF_3$

what <u>mass</u> of P_4 is needed to produce <u>85.0 g</u> PF_3, if the reaction has 64.9% yield?

↳ actual yield

$Mr\ P = 30.97$

$Mr\ F = 19.00$

$Mr\ PF_3 = 87.97$

$$\frac{64.9}{100} = \frac{85.0\,g\,PF_3}{\text{Theo}\,g\,PF_3} \cdot \frac{100}{100}$$

$$\frac{64.9}{100} = \frac{85.0}{\text{Theo}}$$

$64.9\ \text{Theo} = 8500$

$\boxed{\text{Theo} = 130.9\,g\ PF_3}$

$$130.9\,g\,PF_3 \cdot \frac{1\,mol\,PF_3}{87.97}$$

$= \boxed{1.48\ mol\ PF_3}$

$$\frac{1\,mol\,P_4}{4\,mol\,\cancel{PF_3}} \cdot 1.48\,mol\,\cancel{PF_3}$$

$= 0.37\ mol\ P_4$

$$0.37\,\cancel{mol\,P_4} \cdot \frac{123.88\,g\,P_4}{1\,\cancel{mol\,P_4}}$$

$\boxed{= 45.83\ g\ P_4}$