(Chapter 1) **Introduction : Matter & Measurement**

(1.1) The study of Chemistry (not included just read it)

- **Matter** : anything has a mass & occupies space .
- **Atoms** : the infinitesimally small building blocks of matter .
- **Molecules** : Combination of two or more atoms , behave as a single object .

(1.2) Classification of Matter (not included just read it)

A- According to Matter **states** : **Gas – Liquid – Solid**

(Example : water vapor ,liquid water & ice)

B- According to Matter **Forms** : **Element – Compound – Mixture**

- **The Element** : the simplest Pure form of matter , composed of billions of one kind of atom . (like each one of the following : H,O,N,C,Na,Cl,F,Al,,etc) .
- **Compounds** : Combination between two or more element atoms ,each compound is composed of billions of one kind of molecule . (like H_2O compound composed of billions of H_2O molecules)
- **Mixtures** : the most complex & Popular form of matter , composed of two or more different kinds of elements or compounds , they are two types :

  1- Heterogeneous Mixtures : mixtures that don't have a uniform composition . (like the mixture between water & oils)

  2- Homogeneous Mixtures : mixtures that have a uniform composition ,they are called **Solutions** (solutions can be liquids , solids & gases)

  (like the solutions of salts & sugars in water)

(1.4) **Units of Measurement**

- The numbers without units can't represent anything .

- Because of the importance of the units , England & France each developed its own unit system , the English system & the metric system (the French one) .

- most of the countries in the world agreed to particular metric units for the use in scientific measurements , & they Called this new unit system : **SI System** , & its units : **SI Units** & they are :

1- **Kg** [kilogram] for mass. 2- **m** [meter]- distance. 3- **s** [second]- time
4- **K** [Kelvin] - temperature. 5- **mol** [mole] for amount of substance .
6- **A** [Ampere] for electric current . 7- **cd** [Candela] for luminous intensity .

- the previous 7 units are called **the basic SI Units** , in which all other units are derived (which are called **Derived SI Units**) , like **m^3** for the volume , **Kg/m^3** for density , **m/s** for speed ,,, etc .

- In England they still use the English system , for example they use :
1- **lb** (pound) for **mass .** 2- **mi** (mile) **& yard** for **distance .**
3- **gal** (gallon) & **qt** (quart) for **volume .**
Of course , these units are not SI units .

- Sometimes the basic SI Units are too large or too small to measure , so we use prefixes (that refer to certain decimal fractions) with these units to easier description of the measurement , these prefixes are :
1- **G** [Giga] means 10^9 2- **M** [Mega] means 10^6 3- **k** [Kilo] means 10^3
4- **d** [Deci] means 10^{-1} 5- **c** [Centi] means 10^{-2} 6- **m** [Milli] means 10^{-3}
7- **μ** [Micro] means 10^{-6} 8- **n** [Nano] means 10^{-9} 9- **p** [Pico] means 10^{-12}
10- **f** [Femto] means 10^{-15}

Examples: **cm** = 10^{-2} meter , **fg** = 10^{-15} gram , **μs** = 10^{-6} second , **Mm** = 10^6 meter
Question : what decimal fraction of a second is a Picosecond ? → 10^{-12}

- The SI units that preceded with prefixes are **non-basic SI units** .(like **μs** , **Mm**)

Temperature scales

- we have 3 temperature scales :

1- Kelvin Scale (the SI Scale)
2- Celsius Scale (English Scale)
3- Fahrenheit Scale (USA Scale)

- the lowest attainable temperature for gases is 0 K = -273 C (it's called the absolute Zero) , below it all gases are in liquid or solid states .
- In Celsius Scale , **water freezes at 0 ^oC & boils at 100 ^oC** (at sea level) & the normal human body temperature is 37 ^oC .

- In Kelvin Scale , **water freezes at 273 K & boils at 373 K** ,& the normal human temperature is 310 K .
- In Fahrenheit Scale , **water freezes at 32 ^oF & boils at 212 ^oF** , & the normal human temperature is 98.5 ^oF

- To Inter Converting between the temperature scales :

$\boxed{K = {}^oC + 273 \quad \leftrightarrow \quad {}^oC = K - 273}$ (hint to remember the sign in the law : K > ^oC)

$\boxed{{}^oC = ({}^oF - 32) * 5/9 \quad \leftrightarrow \quad {}^oF = 9/5\,{}^oC + 32}$

Questions : Convert 31 ^oC to Kelvin & Fahrenheit . → = 304 K = 87.8 ^oF

- Convert -11.5 ^oC to Kelvin & Fahrenheit . → = 261.5 K = 11.3 ^oF

Derived Units

Volume

- The derived SI unit for the Volume is : m^3 . We usually use the prefixes to measure small volumes (like dm^3 , cm^3 , mm^3 ,,,)
- The metric unit(**L -liter**) wasn't chosen to be SI unit for volume (**<u>it isn't SI unit</u>**)

m^3 = 1000 L , **L = dm^3** , **mL = cm^3**

Density

- It's the mass per unit volume → Density = $\frac{mass}{Volume}$
- The derived SI unit for the Density is : **Kg/m^3**

But we <u>commonly</u> use **g/cm^3** which is equal to **g/mL** to express the density

- the density for water is **1 g/mL or 1 g/cm^3** = 1000 Kg/m^3 or 1000 g/L .

(1.5) **Uncertainty In Measurement**

- we have two types of numbers in the scientific works :

1- **Exact Numbers** : we find them theoretically & mathematically . Examples : the # of our fingers (we count them mathematically) , the # of people in a place , & the # which has a defined value , like : the # of grams in each kilogram (=1000) , the # of cm in each m (=100),,,,, etc .

2- **Inexact Numbers** : they are numbers which **<u>aren't exactly</u> true** , we find them through experimental measurements , this is because of :

- the equipment errors (the equipments are not 100% ideal)
- human errors (false using of the equipment or false reading of the result)

Examples : the mass , density , length , temperature ,,,etc of anything .

Precision & Accuracy

- Precision : a measurement of how are measurements close to each other .
- Accuracy : a measurement of how are measurements close to the correct .

So u can be precise without being correct , but to be accurate ur measurements must be close to the correct . (look at figure 1.24 page 21)

Significant Figures

- It's clear that the inexact numbers coming from experimental measurements are very close to the correct , so they have part that is exactly true but the other part may be false , for example : I looked to my temperature thermometer to know what is the temperature , I found the liquid close to 27 , then I said its 27 ^{o}C , the first part of this # (2) is exactly true , but the second part (7) may be false (maybe 7.1 or 6.9) So :
  Inexact # = Exact part at the left + inexact part at the right

- The Scientists suggested a method to treat these inexact numbers , by rounding the inexact part to only 1 number .
  This method gives u what is called the **Significant figures** . so :
  significant figures = Exact part at the left + 1 inexact digit at the right

- Now to determine the number of significant figures in a reported measurement , look at the number & see if it has a decimal point :
  1**- If it doesn't have** , remove the zeros at the left & right , & count the rest.
  Example : 001020304000 → 1020304 → 7 numbers → 7 significant figures.
  2- **If it has** , remove only the zeros at the left & count the rest .
  Examples :
  00102030.4000 → 102030.4000 → 10 numbers → 10 significant figures .
  0.00020 → 20 → two numbers → 2 Sig. Fig.

- Note : exponential term (10^X) doesn't affect the # of Significant Figures
  Example : $4.00 * 10^{12}$ → 3 sig. fig.

Significant Figures in Calculations

- All the numbers in Chemistry (ex. The masses, Volume … etc) are coming from experimental measurements → they are significant figures , which means they are treated in a way that all of the number is exact except the last number at the right , so when we use the numbers in laws (like using the measured mass & volume in the Density law) the last number in the mass or in the Volume may give more than 3 inexact numbers in the result , & that's bad , so in order to have always all the numbers are exact except the last one , we must round the result of calculations as the following :

1- Addition & Subtraction

- We look to the measurement with the **fewest number of decimal places** , then we round the result to this fewest decimal places .
  Example : 20.45 + 1.322 + 8.1 = 29.872 will be rounded to 29.9
  How ? the fewest measurement in decimal places is 8.1 , with 1 decimal places only , so we round 29.872 to the first decimal place .

2- Multiplication & Division

- we look to the measurement with the **fewest number of significant figures** , then we round the result to this fewest significant figures .
  Example : 6.221 * 5.2 = 32.3492 will be rounded to 32
  How ? the fewest measurement in the number of significant figures is 5.2 , which have 2 sig. fig. , so we round the result to 2 sig. fig. → 32 .
  Question : write the answer of this division in significant figures : **12.53/5.7** .
  the result of division = 2.1982 , the fewest # of sig fig = 2 →answer = 2.2

- If in the same Law u have an (+) or (-) **&** (*) or (/) , what to do ?
  round the result of Multiplications & Divisions
  then round the result of additions & subtractions ,
  as the following example :

- what is the answer of 16.3521 – 1.448 / 7.08 in the correct sig. fig.??
  Answer : we will round the result of Division then the subtraction :
  1.448 / 7.08 = 0.20451 (the fewest sig. fig. are 3)→ we round it to 0.205
  so it becomes :
  16.3521 – 0.205 = 16.1471 (the fewest decimal places are 3) → the answer is 16.147

- **If all the processes in the law are multiplications & divisions then round the final answer only.** example :
  1.09 *(302 / 39)/(6 * 1.2) = 1.1722 (the fewest sig. fig. = 1) →answer = 1
  But if the choices in the exam aren't rounded (have decimal places) then the Dr who wrote the question must be forgotten to round , so choose the Choice that represents the unrounded answer, like 1.2 or 1.17 .

U must be excellent in doing this , because the Choices in the exam will be rounded according to these corrections for all the chapter's laws in this chemistry 103 (not just this chapter) .

More Examples :
- How many sig. fig. in the result of : 15.5 * 2.312 * 5.4 –2.49 ?
  Answer : after doing the multiplication & rounding the answer , it becomes
  72 –2.49 = 69.51 → 70 (because the fewest decimal places = 0 , so we round the result to the first number at the left of the decimal point)
- Solve SAMPLE EXERCISE 1.8 page 24 & its PRACTISE EXERCISE .

Important note : Doctors of this subject (Chem 103) in our beloved university have different thoughts about the method to solve the multi-processes questions that involve both multiplication or division & addition or subtraction, some say just find the result in calculator then round it , some say other ways , each way will result in a different # of sig. fig. :/ , so if ur Dr. didn't give u his method , then just follow the previous one because it's the right one :) .

- To calculate the # of significant figures in the term : **X ± Y** ,
  1- take the first # of Y with its position (example if Y = 123 then we take 100 & if it was 0.00123 then we take 0.001) & put it instead of the corresponding position in X , then count the sig fig of X , it equals the sig fig of the term **X ± Y .** As the following :

- **Example :**

  How many significant figures in the measurement 1000 ± 100 ??

  Answer : we take the first number of 100 with its position → 100 (nothing changed) , then put this 100 instead of the corresponding position in 1000 , which means instead of the 3 zeros , so it becomes 1100 , then we count the sig. fig. in it → has 2 significant figure → 2 sig fig in the term 1000 ± 100

- **Example :**

  How many significant figures in the measurement 1111 ± 22 ?

  Answer : we take 20 from 22 (the first digit & its position) & put it in 1111 → 1120 , then we count the sig fig which is 3 sig. fig.

- **Example :**

  How many significant figures in the measurement 12345 ± 75

  Answer : we take 70 from 75 & put it in 12345 → 12370 → 4 sig. fig.

- **Example (important to be understood) :**

  How many significant figures in the term 12.34567 ± 0.033

  Answer : we take 0.03 (the first digit at left with its position) from 0.033 & put it instead of the corresponding position in 12.34567 , **(deleting the excess numbers at the right**) , it becomes 12.33 , count the number of sig fig → 4 sig fig in that term .

- **Example :**

  How many significant figures in the term 0.03412533 ± 0.00099

  Answer : we take 0.0009 from 0.00099 , put it in 0.03412533 , it becomes 0.0349 , & count the sig fig → 3 sig fig in that term .

  *** this method is for fast solving , I can show u how these questions are solved , but it's not important , just save this way , it's the fastest & most accurate without confusing or errors .

(1.6) **Dimensional Analysis**

- It's a method we can use for the conversion between units , by using **The Conversion factors** , which are fractions whose numerator equals denominator but are expressed in different units . Example :
  Kg = 1000 g → the conversion factor : $\frac{1\ Kg}{1000\ g}$ OR $\frac{1000\ g}{1\ Kg}$

Example : if we have 20 Kg & want to convert it to grams , we choose $\frac{1000\ g}{1\ Kg}$ to let Kg with Kg be cancelled : 20 Kg * $\frac{1000\ g}{1\ Kg}$ = 20000 g .

In the previous example , there is something called **Desired & Given units** , the Desired one is the final unit we want to convert to , while the Given unit is that given in the Qs (the desired at numerator & the given at denominator).

- **Examples** :

- How many meters are there in 20 nanometers ?
  1 nm = 10^{-9} m → 20 nm * (10^{-9} m/1 nm) = 2 * 10^{-8} m

- How many cm are there in 8.5 inch ?? (given that : 1 Inch = 2.56 cm)
  8.5 inch * (2.56 cm/1 inch) = 21.76 cm

  Note : we don't round in such Qs , because we are dealing with exact numbers (defined values)

- **Using two or more Conversion Factors**

- How many mL are there in 154 pL ??
  mL =10^{-3} L , pL = 10^{-12}L , then :
  154 pL * (10^{-12} L/1 pL) * (1 mL/10^{-3}L) = 154 * 10^{-9} mL

- convert 515 m/s to mi/hr . (given that 1 mi = 1.61 Km)
  firstly , find how much **m** in **mi → 1mi = 1.61 Km * 1000m/1Km = 1610 m**
  then , find how much **s** in **hr → 1 hr = 60 min * 60s/1 min = 3600 s**
  finally , 515 m/s *(1mi /1610m) * (3600s/1hr) = **1151.55 mi/hr**

- In converting between metric units of volume , we put the same conversion factor but exponential to 3 , as the following example :

  $$\mathbf{1m/10^2\,cm \rightarrow (1m/10^2\,cm)^3 \rightarrow 1\,m^3/10^6cm^3}$$

- convert 10.6 Kg/m^3 to g/cm^3 . (use the previous $1m^3=10^6cm^3$)
  10.6 Kg/m^3 * (1000g/1Kg) * ($1m^3/10^6cm^3$) = 1.06 * 10^{-2} g/cm^3 .
  <u>Save that to convert $Kg/m^3 \rightarrow g/cm^3$ we divide by 1000</u>

- convert 1.36 * 10^9 Km^3 to L .
  Km = 10^3 m $\rightarrow$ **$Km^3 = 10^9\,m^3$** , m^3 = 1000 L , then :
  1.36 * 10^9 Km^3 * ($10^9m^3/1Km^3$) * ($1000L/m^3$) = 1.36 * 10^{21} L

- **Using the density as a Conversion Factor**

- How many grams are there in 1 qt of benzene ?
  given that the density of benzene = 0.879 g/mL , & 1L =1.057 qt .
  Answer : the density is a conversion factor : **0.879 g/1mL** OR **1mL/0.879 g**
  So : **1 qt * (1L/1.057 qt) * (mL/10 $^{-3}$L) * (0.879 g/mL) = 832 g**
  we can use the density law to solve it but we need the dimensional analysis to convert the units . (mass = density * Volume)

- What is the mass of 10 gal of water . (1 gal = 4 qt , 1L =1.057 qt) **.**
  Answer : 10 gal * (4 qt/1 gal) * (1L /1.057 qt) * (1000 g/L) = 3.78 * 10^4 g
  Notice : I chose the density (1000 g/L) not (1g/mL) to direct cancellation of (L) without using another conversion factor which is (10^3 mL/1L)

- This Subject (Dimensional analysis) helps u solving Qs without a chance for errors , but it is time consuming , & the Exam time is very limited , so u will see in the next chapters to the end of my Chemistry 103 summaries (the second & the final) , that I will give u the final form of the law which u will use for faster solving . So as a result , U will use this method only in this chapter's Qs (converting units).

End Of Chapter 1

(Chapter 2) **Atoms , Molecules & Ions**

(2.5) The Periodic Table

- Atomic Number : the number equals to **proton or electron** number in the element atoms , we put it **above** the element symbol in the periodic table.

- Atomic weight (mass number) : the Number equals to the **sum of Protons & neutrons** in the element atoms , we put it **below** the element symbol in the periodic table.

The periodic table is divided to **columns called Groups** & **rows called periods** as the following :

- **The Groups** : contain elements with <u>similar physical & chemical properties</u> , the group# represents the # of electrons at the outer shell . the groups are of two types A & B :

1A → **Alkali metals** (Li ,Na ,K …) : loss the 1 electron at the outer shell to reach the <u>nearest noble gas configuration</u> .
2A → **Alkaline earth metals** (Be ,Ba ,Mg ,Ca): loss the 2 e^- at the outer shell
3A → (Al): loss the 3 e^- . 4A→(C , Si) : share the 4 e^-
5A → (N,P …) : share or accept 3 e^-
6A → Chalcogens (O, S, Se) : share or accept 2 e^- to reach noble gas config.
7A → Halogens (F ,Cl ,Br ,I ,At) : share or accept 1 e^- .
8A → Noble/Rare/Inert gases (He ,Ne ,Ar) : they don't react at all .
1B → Coinage (Cu ,Ag , Au) they are used to make coins .

- The Periods :

1^{st} period contains→ 2 elements . 2^{nd}& 3^{rd} contain→ 8 elements .
4^{th}& 5^{th} contain → 18 elements . 6^{th} contains → 32 elements .

The elements are divided to 3 categories :

1- **Metals** : most of the elements are metals → All the left of the periodic table except H_2 which is nonmetal ,,, they All are <u>solids</u> except for Hg which is liquid , they have <u>luster</u> & high <u>electrical & heat conductivity</u> .
2- **Metalloids** : diagonal steplike line that runs from (B) to (At) , separating the Metals at the left from the Nonmetals at the right .
They have characteristics fall between metals & nonmetals .
They are : B , Te , Ge , Sb , As , Si . (a mnemonic for it : Bteejy Spicy , :P)
3- Nonmetals : they can be solids , Gases or liquids . They differ from metals by appearance & physical properties (like electrical & heat conductivity) .

(2.6) **Molecules & Molecular Compounds** (not included)

This section <u>isn't included</u> , but know that **Chemical Formulas** are two types **:**

1. **Molecular Formula :** describe the actual numbers & types of atoms in each molecule of a compound . ($\mathbf{C_6H_{12}O_6}$ is Glucose molecular formula)**.**
2. **Empirical Formula :** it gives only the **<u>simplest relative numbers</u>** of atoms in molecule **.** (like $\mathbf{CH_2O}$ is the Empirical Formula of Glucose).

(2.7) **Ions & Ionic Compounds**

- **The ion :** charged particle formed from atoms that gain or lose electrons , they are of two types :

1- Cations : positively charged Ions . **formed by** losing one electron (+1) , two (+2) or three (+3) .
2- Anions : negatively charged Ions . **formed by** accepting one electron (-1), two (-2) or three (-3) .
Examples :
Na^+ : cation , formed by losing one electron .
N^{-3} : anion formed by accepting 3 electrons .

- The size of the Cation is smaller than it's atom , while the anion is bigger .

Example , Li^+ is smaller than Li , O^{-2} is bigger than oxygen atom in O_2 .

- **Qs** : give the mass number , atomic number & the **charge** on the ion with : 22 proton , 26 neutrons & 19 electron . (the answer next page)

  Answer : mass number : 22+26 = 48 , **atomic number = protons** = 22 , **Charge = protons – electrons** = +3

- The 1A , 2A & 3A groups form +1 , +2 & +3 Cations , while 5A , 6A & 7A groups form -3 , -2 & -1 Anions , the Exceptions to this rule are : **Be , B , P** which share electrons only .

  Examples :

  Predict the charges on the most stable ion of : Ba , O , Al , F , N , K .

  Ba = +2 , O = -2 , Al = +3 , F = -1 , N = -3 , K = +1

- The Ions are either Simple ions (like Na^+) & polyatomic ions (like NH_4^+,SO_4^{-2})

- **Ionic Compounds** : Combination between metal & non-metal atoms , through complete transfer of electrons , to form Cations & Anions (instead of molecules as in molecular compounds) . the Ions are arranged in three dimensional structures . there is only empirical formulas for the Ionic compounds because there isn't a singular molecules of them .

- we can directly write the empirical formula for an ionic compound by putting value of the charge of one atom to the other & the opposite .

- **Examples :**

  - write the empirical formula for the ionic compound formed by Ca & Cl .

  Answer : the charge of Ca=+2 , & for Cl = -1 → we give 2 to Cl & 1 to Ca →the formula : $CaCl_2$

  - write the empirical formula for the ionic compound formed by Al & S .

  Answer : the charge of Al=+3 , & for S =-2 → we give 3 for S & 2 for Al → the formula : Al_2S_3

End Of Chapter 2

(Chapter 3) **Stoichiometry : Calculations with Chemical Formulas & Equations**

(3.1) **Chemical Equation** (not included , few terms to know)

In this equation : $2\ H_2 + O_2 \text{--->} 2\ H_2O$:

- The substances before the arrow are called **reactants** , & the ones after it are called **products ,** The numbers before each substance are called the **coefficients** , they indicate the relative numbers of molecules involved in the reaction .
  ** we read the previous equation as :
  every 2 molecules of H_2 (hydrogen) react with 1 molecule of O_2 (oxygen) to form 2 molecules of H_2O (water)

  We add (g) after the gaseous substances in the Chemical equation , (s) after the solid ones , (l) after the liquids & (aq) after the aqueous solutions (solved in water) .

(3.4) **Avogadro's Number & The Mole**

- **Avogadro's Number (A#)** : it's a number used in Chemistry to describe the **$6.02 * 10^{23}$** molecules of a matter . (this number is taken from the # of molecules in 12 g of pure ^{12}C) .
- **The Mole** : it's the amount (in grams) of matter that contains Avogadro's Number of molecules . But it's common to use the mole to describe the Avogadro's number of something .
  Example : 3 moles of O_2 ----means----> 3 times Avogadro's number of O_2 .
- **the moles number (n)** : the number of moles in a sample of matter .
  Example : If (n) for O_2 exists in a balloon = 3 then we have 3 moles of O_2 molecules in it .

Molar Mass

Molar Mass (M.M) : The mass of 1 mole of molecules , & it is equal to the mass of each atom in the Chemical Formula (the difference is in the units only).
* Save the mass of each atom of the following : H=1 , O=16 , C=12 , N=14

Example: M.M for O_2 = 2*16 = 32 g , M.M for N_2 = 2*14 = 28 g ,
M.M for CH_4 = 12+(4*1) =16 g
M.M for glucose ($C_6H_{12}O_6$) = 6*12 + 12*1 + 6*16 = **180 g**

The **Questions** on **converting** between **Moles , grams & molecules number**

- Be sure that u know these abbreviations (that I use) :
  n = number of moles , A# = Avogadro's number ($6.02 * 10^{23}$ mol^{-1})
  Molecules# = the number of molecules in a sample of matter .
  M.M = the molar mass (in g/mol) , Mass = the grams of something

- Now save these 4 main Equations :

---

n = mass / M.M -- (A)
mass = M.M * n -- (B)

n = molecules# / A# -- (C)
molecules# = A# * n -- (D)

- If the Question gives the mass & demands the molecules# Then :
  Find mass/M.M → n , then use (D) .

- And If the Question gives the molecules# & demands the mass Then :
  Find molecules#/A# → n , then use(B)

Examples :

- How many grams are in 10 moles of $C_6H_{12}O_6$ (M.M = 180) ??
  Answer : (use equation B) result = 1800 g

- Calculate the mass of 1 O_2 molecule .
  Answer : molecules#/A# = $1/(6.02*10^{23}) = 1.66 *10^{-24}$ = n , (use B)
  result = $5.3 * 10^{-23}$

- How many <u>molecules</u> of Oxygen are there in 36 grams of it ??
  Answer : mass/M.M= 36/32 = 1.125 = n , (use D) result = $6.8*10^{23}$

- How many <u>atoms</u> of Oxygen are there in 36 grams of it ??
  Answer : find molecules# of O_2 like the previous Qs , then multiply by 2 because O_2 have 2 Oxygen atoms result = $1.35*10^{24}$

- Calculate the number of Na^+ Ions in 14.2 g Na_2SO_4 . (M.M for Na_2SO_4 = 142)
  Answer : mass/M.M=0.1 = n , use D , multiply by 2 , result = $1.2 * 10^{23}$

- How many grams of hydrogen are in 65 g $C_2H_2O_2$ (M.M=58) ??
  Answer : we will find the number of hydrogen moles in $C_2H_2O_2$, then find its mass ,,, $n(H) = 2*n(C_2H_2O_2)$ --- because we have 2 H in each $C_2H_2O_2$
  $= 2 * (mass/M.M)_{C2H2O2} = 2*65/58 = 2.24$ moles H
  then $mass(H) = M.M(H) * n(H) = 1 * 2.24 = 2.24$ g H

(3.6) **Quantitative Information From Balanced Equation**

<u>The coefficients</u> of any balanced equation (for example : $2 H_2 + O_2 \text{--->} 2 H_2O$)
<u>Represent</u> :

1- the **relative # of <u>moles</u> or molecules** in the reaction .
here each 2 moles (or molecules) of hydrogen gas react completely with 1 mole (or molecule) of oxygen gas to produce 2 moles (or molecules) of water

2- the **relative volumes <u>of the gases if they present in the equation</u>** .
*** this is true at constant **pressure & temperature .** (in the Second exam)
*In this Example , each 2 liters of hydrogen gas react completely with 1 liter of oxygen gas to form water at the same pressure & temperature.

The Questions on the Balanced Equations

In such questions , u have a balanced equation (Ex : $2 H_2 + O_2 \rightarrow 2 H_2O$)
& information about 1 reactant/product , then u are asked to measure something about another reactant/product after a complete reaction with the given substance … To solve it , make this step :

- calculate the **ratio** = coefficient of the needed/coefficient of the given .

(in the mentioned example, If the needed info is about (H_2O) & the given is about (O_2) then → **ratio** = coefficient of H_2O/that of O_2 = 2/1 = **2**)

- (I will abbreviate the words : needed → ndd , & given → gvn)

Solving Questions on the Balanced Equations

The whole Idea of such questions related to the fact that :

n (ndd) = ratio * n (gvn) ------------- **#1**

molecules#(ndd) = ratio * molecules#(gvn) ------------- **#2**

- Now , If they give u (n) for a reactant or a product & ask for :
- the mass of another :

  mass (ndd) = M.M(ndd) * [ratio * n(gvn)] ------------- **#3**
- molecules# of another :

  molecules# = A# * [ratio * n(gvn)] ---------------- **#4**

- & If they give u molecules# for a reactant or a product :

  firstly , Calculate **molecules#(gvn)/A#** = n(gvn)

  then If they are asking for :
- the **moles of another** : **use #1**
- the **mass of another** : **use #3**

- But if they are asking for **molecules#** of another then directly **use #2**

- If they give u the mass for a reactant or a product :

  firstly , calculate **Mass(gvn)/ M.M(gvn)** = n(gvn)

  Then if they are asking for :
- the **moles of another** : **use #1**
- the **mass of another** : **use #3**
- **molecules# of another** : **use #4**

- **Note :** U will use the equation #3 in this Chapter & the next Chapters .

- **Examples** in the next page **:**

- Calculate the mass of O_2 consumed in combustion reaction with 22 g of C_3H_8 (M.M = 44) **,** If u know that the reaction is $C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$.

  Answer : ratio = coefficient of(O_2)/that of (C_3H_8) = 5/1= **5**

  Mass/M.M = 22/44 = 0.5 mol = n(gvn) → **use #3** :

  mass(O_2) = M.M(O_2) * [ratio *n(C_3H_8)] →→→ **result = 80 g .**

- Add to the previous Qs : Calculate the number of CO_2 molecules produced

  Answer : the ratio = coefficient of(CO_2)/ (C_3H_8)= **3** , n(gvn)=0.5 , (use **#4**)

  result = 9 * 10^{23}

- Calculate the mass of $KClO_3$ decomposed if there were 3.01 *10^{24} O_2 molecules produced , according to this equation : $2\ KClO_3 \dashrightarrow 2\ KCl + 3\ O_2$. (given that M.M for $KClO_3$ = 122.5 g/mol)

  Answer : the ratio = 2/3 , molecules#/A# = 3.01 *10^{24}/(6.02 *10^{23})= 5 = n(gvn) , (use **#3**) **, → mass($KClO_3$) = 408.33 g .**

(3.7) **Limiting Reactant**

- it's the reactant that consumed first & limit the quantity of the product .
- at the end of the reaction , there are two things **: the products & the excess of reactants** .
- we study this subject because in reality , we usually measure the grams of 1 reactant only which can give the needed amount of product if reacted with enough other reactants , & let it react with other unmeasured enough reactants , & calculate everything according to the grams of measured reactant & the ratios in Balanced Equation .

Mathematically

- To find L.R → L.R is the <u>lowest reactant in its fraction</u> : **(n/coefficient)** .
- if they give the masses of reactants (rather than n) & ask to determine the L.R , then L.R is the lowest reactant in its fraction **(mass/M.M)/coefficient**
- I will call the lowest fraction (which is for the L.R) → L.R frac.

- Use these to find the :

1- **moles consumed or produced** of a substance = [its coefficient * L.R frac.]
2- **mass consumed or produced = M.M * [its coefficient * L.R frac.]**
3- **moles of excess** of a reactant = its coefficient * (its fraction – L.R frac.)
4- **mass of excess = M.M *[its coefficient * (its fraction – L.R frac.)]**

- **Percent yield = (actual yield/theoretical yield) * 100 %**
  it's always less than 100%

- **Actual yield :** the product **mass** actually obtained from the reaction (given in the Qs)
- **theoretical yield :** the product **mass** theoretically calculated when all the L.R reacts . we calculate it according to the second equation above .

  * If the Question gives u information about L.R (like its mass or anything) then u can directly use any of the previous equations especially #3 (in the second previous page) . Look at the last example in the next page to understand what I'm talking about ^_^ .

- **Examples** :

- Consider the following reaction : $2\ Al + 3\ Cl_2 \text{---->} 2\ AlCl_3$, If 1.5 mol of Al & 3 mol of Cl2 are Allowed to react , (M.M of Cl_2 = 71) , Find the following :
  1- the L.R . 2- mass of $AlCl_3$ formed .(M.M=133.5)
  3- mass of excess left at the end .
  4- if 175 g of $AlCl_3$ obtained actually , calculate the percentage yield .

  Solution : Al Cl_2 ,,, Al Cl_2
  1- the fractions : n/coefficient : 1.5/2 3/3 → 0.75 1 → **Al is the L.R**
  →(L.R frac. = 0.75)
  2- mass formed = M.M * Coefficient * L.R frac. …… ($AlCl_3$ Coefficient =2)
  answer = 200.25 g
  3- mass of excess = M.M * Coefficient * (its fraction – L.R frac.) …… (the excess is Cl_2) so mass of excess Cl_2 = 71 * 3(1 - 0.75) = 71 * 0.75 = 53.25 g
  4- theoretical yield = mass formed (solved in part 2)= 200.25 → so :
  % yield = actual/theoretical *100 ,,,,, answer = 87.4 %

- If 2 g of Zn (M.M = 65.4) & 2.5 g of $AgNO_3$ (M.M=169.9) were allowed to react completely according to this equation :
  $Zn + 2\ AgNO_3$ -----> $2\ Ag + Zn(NO_3)_2$, (M.M for Ag =107.9) , Find :
  1- the L.R . 2- mass of $Zn(NO_3)_2$ formed (M.M=189.4) .
  3- mass of Ag formed .
  4- the amount (mass) of Zn reacted . (hint :mass reacted = mass consumed)

  Solution : Zn $AgNO_3$,,, Zn $AgNO_3$
  1- the fractions: (2/65.4)/1 (2.5/169.9)/2 → 0.0306 0.007 →
  $AgNO_3$ is the L.R (L.R frac. =0.007)
  2- mass of $Zn(NO_3)_2$ formed = 1.33 g 3- mass of Ag formed = 1.51 g
  4- mass of Zn reacted = 0.46 g .

- The Iron ore (Fe_2O_3) converts to Iron according to this equation :
  $Fe_2O_3 + 3\ CO \rightarrow 2\ Fe + 3\ CO_2$, If u start with 150 g of Fe_2O_3 (M.M=159.7) as a limiting reagent , what is the theoretical yield of Fe (M.M=55.85) ? & what is the percent yield if u know that the actual yield of Fe was 87.9 g ??

  Solution :
  the given info is about the L.R , so we solve as usual :
  mass/M.M (for Fe) = 150/159.7 = 0.94 = n(gvn)
  the ratio = Coeff. of (Fe)/(Fe_2O_3) **= 2**
  theoretical yield = mass (Fe) = M.M(Fe) * [ratio * n(gvn)] = 55.85 * 2 * 0.94
  = 104.9 g
  percent yield = actual / theoretical * 100 = 87.9/104.9 * 100 = 83.8 % .

End Of Chapter 3

:D

(Chapter 4) **Aqueous reactions & Solution Stoichiometry**

- The **aqueous solutions** are made by dissolving substances in water to make a homogeneous mixture where the water is the solvent (the larger part of the solution) & the substances are the solutes (minor part) .

(4.1) **General Properties of Aqueous Solutions** (not included , quick read)

- The solutes are of three categories according to their **electrical conductivity** :

1- **Strong Electrolytes** : the solutes that conduct a large quantity of electricity because they exist as Ions Only in the solution . they are :

A- the **Ionic Compounds** (salts like NaCl ,KBr & NH_4Cl) .

B- the **Strong acids** (they are : HCl, HBr, HI, HNO_3, $HClO_4$, $HClO_3$, H_2SO_4).

C- the **Strong Bases** (they are : LiOH , KOH , NaOH ,$Ca(OH)_2$, $Ba(OH)_2$).

2- **Weak electrolytes** : the solutes that conduct little electricity because they exist mainly in their molecular form in the solution & a small fraction as Ions . they are :

A- **Weak acids** (any compound has H able to loss , like : CH_3COOH , HF) .

B- **Weak bases** (any compound is able to gain H^+ , like NH_3) .

3- **Non-electrolytes** : the solutes that don't conduct electricity because they exist as their molecular form only . they are : All molecular Compounds except the strong acids & Bases (like sugars such as glucose –$C_6H_{12}O_6$-).

- **Molecular Compound** : is a group of two or more atoms held together by covalent bonds .
- **Dissolve** : make a homogeneous mixture . (glucose dissolve in water)
- **Dissociation** : separation to Ions . (the glucose doesn't dissociate in water)
- **Solubility** : the **ability of substance to dissolve** in solvent making a solution

(4.5) **Concentrations of solutions**

Molarity

M = n/V -------- (1a) M : molarity in (M - molar)=mol/L

M = (mass/M.M)/V -------- (1b) V : volume in (L)

n = MV -------- (2a)

mass = M.M*MV -------- (2b)

V = n/M ------- (3a)

V = (mass/M.M)/M ------- (3b)

- In strong electrolytes :

- **n(for ion) = i * n(electrolyte) & M(for ion) = i * M(electrolyte)** where (i) is the # of that ion in the formula of the salt , acid or base .
- **Examples** :
- Calculate the molarity of a solution made by dissolving 23 g Na_2SO_4 (M.M= 142) in enough water to make a solution of 125 mL .
  Answer : convert mL to L → 0.125 L , then use 1b , result = 1.3 M

- What is the molarity of K^+ ions in a 0.015 M solution of K_2SO_4 .
  Answer : i = 2 , → M(K^+) = i*M(salt) = 2 * 0.015 = 0.03 M

- How many grams of Na_2SO_4 are required to make 0.350 L of 0.5 M solution of it , given that M.M for Na_2SO_4 = 142 .
  Answer : use 2b →→ result = 24.85 g .

- How many <u>milliliters</u> of 0.5 M Na_2SO_4 solution are needed to provide 0.038 mol of this salt .
  Answer : use 3a , then convert L → mL by multiplying with 1000
  result = 76 mL

Dilution

- we add water to a solution , so V & M is changed , but n remains constant , so :

MV(before) = MV(after)

- be aware of that V (after) = V (before) + V (added water)

- **Examples :**

- What volume of 14.7 M H_3PO_4 is needed to prepare 25 L of 3 M H_3PO_4 solution .
  Answer : M(before) = MV(after)/V(before) … continue → result = 5.1 L

- How many mL of 5 M $K_2Cr_2O_7$ solution must be diluted to 250 mL of 0.1 M (result = 5 mL)
  Note : there is no need to convert mL to L here , because the two volumes have the same unit .

- If 10 mL of a 10 M NaOH solution is diluted **with** 250 mL H_2O .
  what is the new molarity ?
  Answer : the additional water is 250 mL , so the new volume = 260 mL
  result = 0.385 M

(4.6) **Solution Stoichiometry & Chemical analysis**

In these questions , u have a balanced equation & information about one or two reactants , & asked to calculate → n , M or V for one of them

U can identify these Questions by "complete reaction" , "titrated with" or "neutralize"….. So we have 2 reactants as the following :

reactant 1	reactant 2
n	n
(M - V)	(M - V)

Because the case here is a balanced equation → n (ndd) = ratio *n (gvn)
where : **ratio = coefficient of the needed / coefficient of the given** .

- Now If the Qs gives (n) for reactant 1 (or both M&V for it) , & asks for :

1- (n) for reactant 2 , then :
n (ndd) = ratio *n (gvn) -----------or---------- **n (ndd) = ratio *MV (gvn)**

2- Mass for reactant 2 , then :
mass (ndd) = M.M * [ratio*n (gvn)] -----------or----------
mass (ndd) = M.M * [ratio*MV (gvn)]

3- Volume for reactant 2 , after giving the M for it , then :
V(ndd) = [ratio*n(gvn)] / M(ndd)
V(ndd) = [ratio*MV(gvn)] / M(ndd) -------- mostly used in neutralizing & titration

- Molarity for reactant 2 , after giving V for it , then :
  M(ndd) = [ratio*n(gvn)] / V(ndd)
  M(ndd) = [ratio*MV(gvn)] / V(ndd) -------- mostly used in neutralizing & titration

- **Titration :** The process of reacting two solutions with each other , one of them has a known concentration (it's called a **Standard solution**) while the other has unknown concentration (we want to calculate it) .
  End Point : the point at which the color of the indicator changes .
  (it occurs at the equivalent point) .
  Equivalent Point : the point at which all the unknown solution reacts with known moles of the known solution . At this point , we use the balanced equation to determine the ratio & solve as we solve normal balanced equation , using the suitable law from the listed above , to determine the unknown concentration or any unknown quantity .

  - **Note :** In neutralizing & acid-base titration , the Qs usually don't give u the balanced equation so u can't determine the ratio unless u write the equation, & To write it :
    put the number of H in the acid as a coefficient for the base & the number of OH in the base as a coefficient for the acid , (there is no need to write the products) , for example :
    $Ba(OH)_2$ neutralize HCl : $Ba(OH)_2 + 2\ HCl \rightarrow$ etc.
    H_3PO_4 titrate $Ca(OH)_2$: $2\ H_3PO_4 + 3\ Ca(OH)_2 \rightarrow$ etc.

Examples :

- How many grams of $Ca(OH)_2$ (M.M=74) are needed to neutralize 25 mL of 0.1 M HNO_3 .
  Answer : by using the previous note, the equation is : $Ca(OH)_2$ + 2 HNO3 so the ratio =1/2=0.5 ,then continue using :
  mass (ndd) = M.M * [ratio*MV (gvn)] ... answer = 0.0925 g

- If a sample of Fe^{+2} reacted completely with 47.2 mL of 0.0224 M MnO_4^- according to this equation : $MnO_4^- + 5\ Fe^{+2} + 8H^+ \rightarrow Mn^{+3} + 5\ F^{+3} + 4\ H_2O$
  Calculate the moles of (Fe^{+2}) reacted .
  Answer : ratio = 5/1=5 ,,, then : **n (ndd) = ratio *MV (gvn)**
  result = $5.3 * 10^{-3}$ mol

- If 45.7 mL of 0.5 M H_3PO_4 is required to neutralize a 20 mL sample of $Ba(OH)_2$, what is the molarity of $Ba(OH)_2$ solution ?
  Answer : the equation is : 2 H_3PO_4 + 3 $Ba(OH)_2$,,, so the ratio = 3/2
  then **M = ratio * MV(gvn)/V(ndd)** →→→ M($Ba(OH)_2$) = 1.71 M

- What is the concentration of 30 mL HCl are needed to react with 20 mL of 0.15 M NaOH ?
  Answer : ratio = 1 , **M(ndd) = [ratio*MV(gvn)] / V(ndd)** →result = 0.1 M

- Calculate the mass of NaOH (M.M = 40) that is required to titrate 50 mL of 0.22 M H_2SO_4 .
  Answer : 2 NaOH + H_2SO_4 ,,,, so the ratio = 2/1=2 , then :
  mass (ndd) = M.M * [ratio*MV (gvn)] →→→mass(NaOH) = 0.88 g

- If we have a sample of Chloride (Cl^-) (M.M= 35.5) being titrated with Ag^+ according to this equation : Ag^+ + Cl^- -----> $AgCl_{(S)}$.
  Find the mass of Cl^- at the equivalent point , if 20.2 mL of 0.1 M Ag^+ is needed to react with the sample of Cl^- .
  Answer : ratio = 1 ,,, then : **mass (ndd) = M.M * [ratio*MV (gvn)]**
  result = 0.0717 g

- **Note :** Mixing substances with each other in which these substances have a common Ion (like H^+ in acids or OH^- in bases or any Ion in salts) , leads to combination in the concentration of this common ion , If you want to calculate its concentration use this law :

  M(for the common ion) after mixing = $\frac{(iMV)1+(iMV)2}{V1+V2}$

  i = the number of the common ion in the substance formula .
  1 : the first mixed substance , 2 : the second mixed substance

- In mixing acids or bases , they must be strong to use this law , the weak acids or bases when mixed have a different issue .

- **Examples :**

- Calculate the molar concentration of H^+ in a solution made by mixing 50 mL of 0.1 M HNO_3 & 50 mL of 0.4 M HCl .

  Answer : note that the acids are strong so we can use the previous law :

  M(for H^+) after mixing = $\frac{\mathbf{(iMV)1+(iMV)2}}{\mathbf{V1+V2}} = \frac{(1*0.1*0.05)+(1*0.4*0.05)}{0.05+0.05}$

  = 0.25 M

- Calculate the molar concentration of OH^- in a solution made by mixing 50 mL of 0.1 M NaOH & 100 mL of 0.4 M $Ca(OH)_2$.

  Answer : M(OH^-) after mixing strong bases = $\frac{\mathbf{(iMV)1+(iMV)2}}{\mathbf{V1+V2}}$

  M(OH^-) = $\frac{(1*0.1*0.05)+\ (2*0.4*0.1)}{0.1+0.05}$ = (0.005 + 0.08)/(0.15) = 0.56666 M

- Calculate the molar concentration of the solution made by mixing 30 mL of 0.2 M $NaNO_3$ & 70 mL of 0.3 M $Ca(NO_3)_2$.

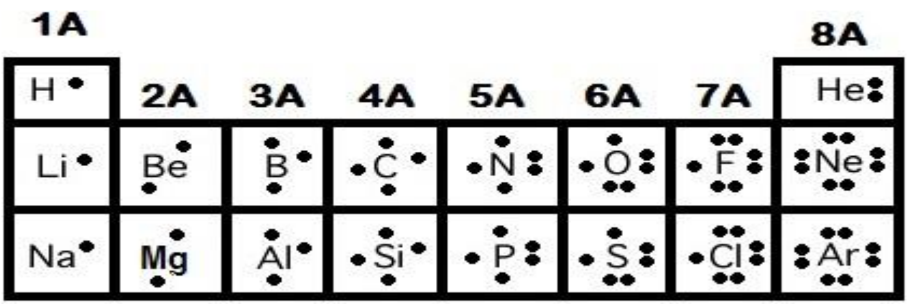
  Answer : these are salts , so use the same law :

  M(NO_3^-) after mixing = $\frac{\mathbf{(iMV)1+(iMV)2}}{\mathbf{V1+V2}}$ = … continue by the same way :)

End Of Chapter 4

(CH.8) **Chemical Bonding** – (Ionic/Covalent)

- **Lewis symbol** : chemical symbol surrounded by dots equal to the # of valence electrons .



  - Valence electrons : the electrons in the outer most shell (equal Group #).

Examples :

Li → 1A Group → 1 valence electron , Ca → 2A Group → 2 valence electrons

- **Octet Rule** : elements tends to gain , loss or share electrons in order to be surrounded by 8 e^- in the outer most shell . (octet means 8)
  **the elements that don't follow the octet rule are :
  1- H : which loss 1 electron to be a proton (H^+) , or accept 1 electron to reach He configuration (2 e^- in the outer most shell) .
  2- Li : which loss 1 electron to reach He configuration (2 e^- in outer shell) .
  3- Be & B ,,,, they share their valence electrons & they become stable in different compounds where the outer shell is not 8 electrons .

- The groups 1A , 2A & 3A loss 1,2,3 electrons to reach the configuration of the previous period's Noble Gas , (Li tend to reach He not Ne or Ar) , **While** groups 4A,5A,6A & 7A share or accept electrons to reach the configuration of the same period's Noble Gas (F accepts 1 e^- to reach Ne) .

- What do we mean by element configuration ?
  it is the distribution of electrons of an atom in its orbits , as the following :
  configuration of Na (11 electrons) : $1s^2\ 2s^2\ 2p^6\ 3s^1$
  what is that ?, elements have orbits in which they fill them by their electrons one after another , in sequence : 1s 2s 2p 3s 3p 4s 3d 4p 5s ... (اس اس بس بس دبس) , in which "s" can be filled up to 2 electrons , "p" up to 6 electrons , "d" up to 10 electrons .
  we just put the electrons of each orbital above it , like $1s^2$,,, now we often use the noble gas symbol between brackets like [He] to abbreviate its configuration when it's a part of the configuration of a another element .
  examples :
  [He] : $1s^2$ → [Li] : $1s^2\ 2s^1$ = [He] $2s^1$
  [Ne] : $1s^2\ 2s^2\ 2p^6$ → [Na] : $1s^2\ 2s^2\ 2p^6\ 3s^1$ = [Ne] $3s^1$

(8.2) **Ionic Bonding**

- It is an **electrostatic attraction** results from complete transfer of electron(s) from a **metal** to a **nonmetal** .
  - **The formation of the Ionic Bonding is highly exothermic ,** this is because the attraction between unlike ion charges arranges the ions oppositely in a solid array (**Lattice**) , releasing too much energy & making it highly stabilized (the releasing energy → exothermic reaction).

- The loss of e^- from atoms is always endothermic , The gain is exothermic .

Lattice Energy (L.E) : the amount of energy required to separate a mole of a solid ionic compound into gaseous ions .

- Lattice Energy is connected to the strength of the ionic bond & so to its Melting Point(m.p) .

the higher L.E → stronger ionic bond & higher melting point **(m.p) .**

- Lattice Energy **depends on** :

1. Magnitude of the **charges** (the multiplication of Ions Charges) in the ionic compound . **the higher charge magnitude → higher L.E .**
2. If there are two compounds with the same charge magnitude , then it is determined by : the **length** of the bond (size of the molecule) .
   the shorter bond (smaller size) → higher L.E .

- So the Priority is for the Charge magnitude , then to the bond length .

- To Compare **molecules in their sizes** (to determine the higher L.E/m.p) :
  Look at the periodic table , then draw an imaginary line (in ur mind) between the two atoms which make each molecule u are comparing , then the line which is **above** the others is the line of the smallest molecule which means the line of the highest L.E & m.p . (this is for fast solving) look at the following examples to get the point ;) .

Example : which one of the following has the highest L.E (or the strongest ionic bond or the highest m.p) Or : Rank the following according to … etc .

KCl $AlCl_3$ KBr MgO

Answer : Firstly , we compare their charges (the multiplication of the charges)

KCl → K= +1 ,Cl=-1→1*1= **1** $AlCl_3$ → Al=+3 ,Cl=-1→3*1= **3**

KBr → K=+1 ,Br=-1→1*1= **1** MgO → Mg=+2 ,O=-2→2*2= **4**

so --- **MgO > $AlCl_3$ > KBr,KCl**

now we compare KBr & KCl according to the size . use the part of periodic table in the next pages & make imaginary lines between K&Br , & between K&Cl . Now which line is above the other ??? it's the KCl line → KCl is higher L.E & m.p than KBr . finally , **MgO > AlCl3 > KCl > KBr** .

Example : Rank the following according to their strength (or L.E or m.p) :

Al_2O_3 , Al_2S_3 , K_3N , Rb_3N

Answer : Charge: 3*2=6 3*2=6 1*3=3 1*3=3

So : Al_2O_3 & Al_2S_3 > K_3N & Rb_3N

then we compare Al_2O_3 with Al_2S_3 , & K_3N with Rb_3N using the imaginary lines :

Al_2O_3 line is above Al_2S_3 line → $Al_2O_3 > Al_2S_3$

K_3N line is above Rb_3N → $K_3N > Rb_3N$

Finally : **$Al_2O_3 > Al_2S_3 > K_3N > Rb_3N$**

Example : Rank the following according to m.p (or strength or L.E) .

CsI NaF CaO LiS_2

Answer : Charges: 1*1=1 1*1=1 2*2=4 1*2=2

so : CaO > LiS_2 > CsI & NaF , we compare CsI & NaF , make lines , then :

NaF line is above CsI → NaF > CsI → Finally , **CaO > LiS_2 > NaF > CsI**

(8.3) **Covalent Bonding**

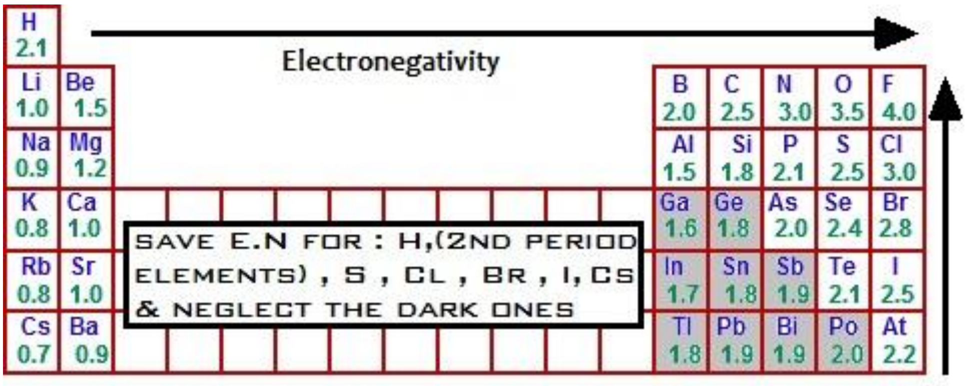
- It's a Chemical bonding formed by **sharing** a pair(s) of electrons between two atoms .
  - This sharing might be equal or unequal between the atoms , so we have :
    1- **Polar** Covalent Bond → unequal sharing of the electrons (like in HF)
    2- **Non-Polar** Covalent Bond → equal sharing of the electrons
    (this happen in the diatomic molecules → H_2 , O_2 , N_2 , F_2 , Cl_2 , Br_2 , I_2)

- In the multiple Covalent Bonding (double & triple bonds) , the more # of bonds between the two atoms → the stronger the bond & less the length .
  So for example : the relative strength & length of C-O bonds in :
  CO (triple bond) , CO_2 (double bond) & H_3C-OH (single bond) is :
  strength : CO > CO_2 > H_3COH , length : H_3COH > CO_2 > CO

(8.4) **Bond Polarity & Electronegativity**

- **Bond Polarity** : a characteristic describes how **bad is the electron sharing** between the atoms of the bond .
  - So higher Polarity is less sharing & more close to the complete transfer , while Lower Polarity is more equally sharing of the electrons .

- **Electronegativity** (E.N) : the ability of atom in a molecule to attract electrons toward itself .



- **Electronegativity increases by going up & right** (as shown)

* Note that the 2nd period starts with 1 E.N & increases .5 for each square to the right , until reaching F . While 3rd period starts with .9 & increases .3 until reaching P then increase .4 then .5 .

- **Electronegativity Difference (ΔE.N)** in a molecule :
  the difference in E.N between the atoms which form the molecule .
  - ΔE.N measures How much one atom in a molecule attracts the electrons toward itself more than the other (more close to it) , so the higher E.N has a partial -ve charge , while the less E.N has a partial +ve charge .

- We can refer to the polarity by the E.N difference in a molecule :
  higher ΔE.N → one atom attracts more e^- → less sharing → **higher polarity** & vise versa .
  Example : Look at the following molecules :

	F_2	HF	LiF
ΔE.N :	4–4 = **0**	4–2.1 = **1.9**	4–1 = **3**
Type:	Non-polar Covalent	Polar Covalent	Ionic

- the scientists found that **when the E.N difference in a molecule is equal or more than 2** , the electron is transferred rather than shared → the bond is more likely to be **Ionic** <u>(the Ionic character is more than Covalent Character)</u> **,but when ΔE.N is less than 2** → the bond is more likely to be **Covalent** ,<u>(the Covalent character is more than Ionic Character)</u> .

- In General **: More ΔE.N → more Polarity , more Ionic character & less Covalent Character , & Vise Versa .**

Dipole Moment

- It is a quantitative measure of the polarity of dipole molecules ; (molecules with equal & opposite charges at their ends - like HF , HCl) .
- So we can refer to the polarity of a molecule by both : the Dipole moment & it's ΔE.N .
- Because of the E.N difference , two equal opposite charges appear on the dipoles , **the more ΔE.N , the higher Charges appear & the bond length becomes <u>slightly</u> shorter** .

- **Dipole Moment (μ) = Qr**
  (Q : charge magnitude in coulomb (C) , r: Bond length in meter (m))

  - I don't Know if the calculations on the dipole moments are included because my Dr when he gave me this course , he didn't mention it ,,, so I will leave it to u , If it's included , go to page 311 in the book & solve the sample exercise & its practice exercise .

- When ΔE.N is higher , the Charge is higher which means that the Dipole moment is higher , so **More ΔE.N → more Dipole moment .**
  - we neglect the bond length effect on the dipole moment because it changes slightly comparing to Q for different ΔE.N .

Using the periodic table to solve Questions about :
ΔE.N , the Polarity , Ionic/Covalent Characters , Dipole moments
<u>& sizes of Particles (atoms & Ions)</u>

H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn

- In these Questions , they give u formulas of some compounds & demand the highest or the lowest or the ranking of them according to their ΔE.N , Polarity , Ionic Character , covalent character or dipole moments , By using the periodic table which is put in isolated page in the exam .

- U can only solve these Questions by finding the ΔE.N , then :
More ΔE.N → more Polarity , more dipole moment , more Ionic character & less Covalent Character . (& vise versa)

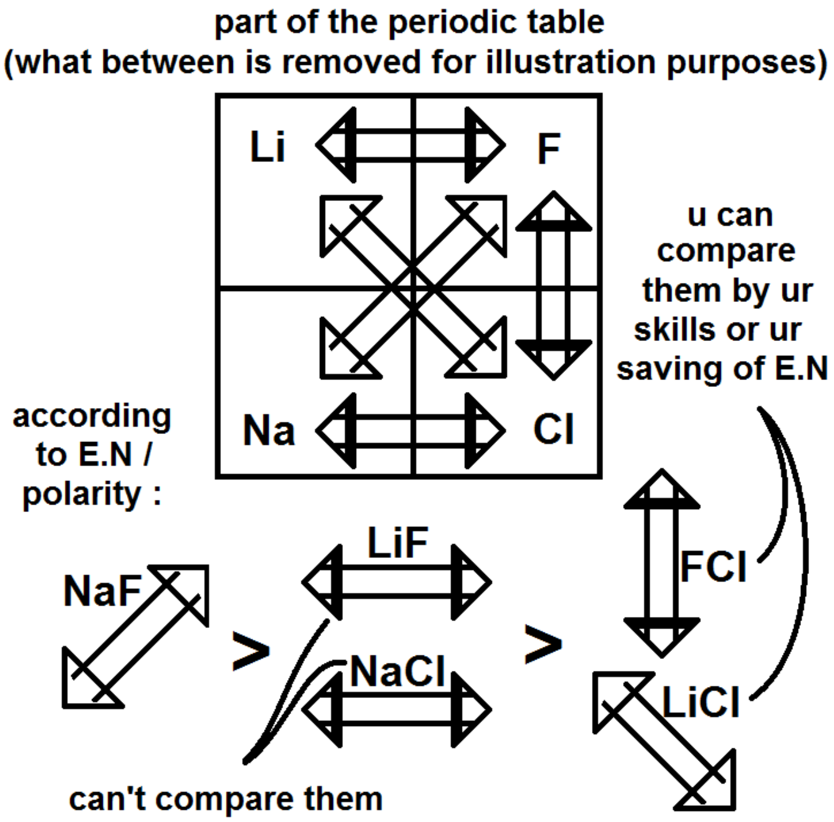
- You can use the following method (which I developed) to achieve that , u can find the relative ΔE.N (→ rest of variables) , for given compounds .
it's a very fast method , as the following :
make an imaginary line on the periodic table between the 2 atoms which make each molecule .
then the line will be one of the following :
1- the line is exactly horizontal (the molecule atoms at the same period)
(the line is like this — ,& I will refer to the horizontal by this line "—")
2- the line is oblique from inferior left to superior right .
(the line appears like this / , & I will refer to this oblique by this line "/")
3- the line is oblique from superior left to inferior right .
(the line appears like this \ ) .
4- the line is exactly vertical (the molecule atoms are at the same Group) .
(the line is like this |) .

- now we can determine which molecule has more ΔE.N :

1- take this general rule :
(Oblique /) is higher ΔE.N than **(horizontal —)** which in turn is higher ΔE.N than **(vertical |)** & **(Oblique \)** .
So : **(/) > (—) > (|) ~ (\)**
look at the next picture to get the point .

- The exception to this rule (unlikely to be asked in the exam) is when u are comparing between horizontal & vertical lines, the rule which says (—) is higher ΔE.N than (|) , is true except when :

the vertical line starts at one of (B to F) atoms then extend downward , & the horizontal extends within the right of periodic table (groups 4,5,6,7A) . at this case the opposite occurs , **(|)** is higher ΔE.N than **(—)** .
Examples : BrF & BF_3 → BrF > BF_3 ,,,, ClF & OF_2 → ClF > OF_2 .



2- When comparing between the (/) themselves or Horizontals or Verticals:
   * the longer line → higher ΔE.N
3- When comparing between the **(Oblique \)** molecules :
   the line which is less oblique & more horizontal → higher ΔE.N .
   except if these oblique lines **start at one of (B to F)** atoms , then the more oblique & **more vertical is the higher .**

The Following examples will explain what I'm talking about , it's easy & fast .
* After ranking the molecules according to ΔE.N, **this ranking is the same ranking according to the Polarity / Ionic Character / Dipole moment** , but if u want the ranking according to **Covalent Character** , just reflect the ranking .

Examples

Which of the following is the most polar ? SCl_2 SBr_2 $SeCl_2$ $SeBr_2$

Answer : from the periodic table , the lines : — \ / —

so : <u>$SeCl_2$</u> > SCl_2 ~ $SeBr_2$ > SBr_2 .

Which of the following is the highest in its Ionic Characters :

RbBr RbCl $CaBr_2$ $CaCl_2$

Answer : the lines : / / — /

firstly , RbBr , RbCl & $CaCl_2$ > $CaBr_2$

then , RbCl > RbBr & $CaCl_2$ because it's the longest

so : **RbCl** is the highest ionic character

Which of the following is the lowest Dipole moment ? :

HCl , HBr , LiCl , LiBr

Answer : the lines : \ \ \ \

firstly , because we are comparing between (\) , the less oblique & more horizontal is higher E.N (higher dipole moment) → the most oblique has the lowest dipole moment → it's HBr .

Rank these according to their <u>Covalent Characters</u> : AlN , Al_2O_3 , BN , B_2O_3

Answer : the lines : / / — —

firstly , AlN & Al_2O_3 > BN & B_2O_3

secondly , Al_2O_3 line is longer than AlN ,,, & B_2O_3 line is longer than BN

so : Al_2O_3 > AlN > B_2O_3 > BN . but **according to covalent character** (the opposite) : **BN > B_2O_3 > AlN > Al_2O_3**

Which of the following has the lowest covalent character ??

FeO CuO FeF_2 CuF_2

Answer ---------------------------- the lines : / / / /

all of them are / so compare the lengths : medium, shortest , longest , medium

the longest is the highest E.N → the lowest covalent character → FeF_2

*** this method is correct but has limitations , why ??**
because the rules are applicable only when u are comparing between :

related compounds (have atoms in common & the other atoms are in the same group or period) ,which means if u compare between LiI & AlN for example , it isn't possible to determine which is more E.N difference except if u save each atom's E.N , because there isn't any common atom , so we can't compare , if u use the rules here , u will obtain false answer which is AlN is higher than LiI , while the true is the opposite … fortunately in the exam , they give u related compounds . what I mean by "the other atoms are in the same group or period" , is as following , suppose that u face in the exam a question asking to compare between $SeCl_2$ & BCl_3 for example , here there is a common atom (Cl) , but the other atoms (Se,B) aren't in the same group or period , here u must answer according to what u save of E.N. ,,,, so as a general conclusion : **if u answer according to ur saving of E.N , this is the best answer , & better than this method . this method is for related compounds only , & to help u in the exam when u get confused & want a quick answer .** (for me I used it)

- **Now to solve the Questions about the sizes of particles (atoms & Ions)**

Save these Rules :

1. The atoms become smaller by going up & right . So :
   * (8A noble gases < 7A < 6A < 5A < 4A < 3A < 2A < 1A) for each period .
   * the atoms of each group are smaller than the atoms below it .

2. **Cations** (+1,+2 ,+3) are :
   - **smaller** than **their neutral atoms** & all the **atoms , cations & anions in the same period & periods below** .
   - **smaller** than the **Anions** (-3 , -2 , -1) at the **period above** .
   *** Regarding cations in the same period , the more + , the smaller :

   +3 < +2 < +1

   (unimportant note , unusual to ask about : +1 is bigger than the atoms in (5A,6A,7A,8A) of the same period & the noble gas in the period above) .

3. **Anions** (-3,-2,-1) are :
   - **bigger** than **their neutral atoms** & All the **atoms , cations & anions** in the $\boxed{\textbf{same}}$ **period & periods** $\boxed{\textbf{above}}$.
   - **bigger** than the **Cations** (+1,+2,+3) at the **period below** .

   *** Regarding Anions in the same period , the more – , the larger :

   $$\boxed{\mathbf{-3 > -2 > -1}}$$

   (unimportant note , unusual to ask about : anions are bigger than the (5A,6A,7A,8A) atoms at the period directly below it , but surely smaller than their anions) .

Examples

- **Which** of the following is the **biggest** ?? S^{-2} , S , O^{-2} , O

  Answer : according to rule 3 :

  Anions are bigger than their neutral atoms → $S^{-2} > S$, $O^{-2} > O$

  S^{-2} is bigger than all atoms & anions in the periods above → $S^{-2} > O \& O^{-2}$

  So → **S^{-2} is the biggest** ,,,

  the ranking ?? (unusual to be asked in the exam) → $S^{-2} > O^{-2} > S > O$

  why ?? because S > O (rule 1) & $O^{-2} > S$ (rule 3) (Anions are bigger than the (5,6,7,8A) atoms of the period below).

- Which of the following is the smallest ?? Rb^{+} , Sr^{+2} , Y^{+3}

  Answer : according to rule 2 : for the Cations in the same period →

  +3 < +2 < +1 → **Y^{+3} is the smallest**

- Find the size ranking of the following : N^{-3} , O^{-2} , F^{-1} , F , Ne

  Answer : according to rule 3 :

  N^{-3} , O^{-2} , F^{-1} are in the same period → $N^{-3} > O^{-2} > F^{-1}$

  N^{-3} , O^{-2} , F^{-1} are bigger than the atoms in the same period

  → $N^{-3} > O^{-2} > F^{-1} > F \& Ne$

  F > Ne --- rule 1 so → **$N^{-3} > O^{-2} > F^{-1} > F > Ne$**

- What is the biggest of these particles : Na^{+} , Li^{+} , Mg , Al

  Answer : $Na^{+} > Li^{+}$ -------- rule 2 , Mg > Al ----- rule 1

  $Mg > Na^{+}$ -------- rule 2 → so Mg is the biggest

  what is the smallest ?? → Li^{+} :)

End Of Chapter 8

(Chapter 10) **Gases**

(10.2) **Pressure**

- **Pressure** : Force per unit area → P = F/A
- Pressure **units** :

  1- **SI** unit → **Pa** (pascal)

  2- **Common units** → **atm** (atmosphere), **bar** , **torr** (Torricelli) , **mm Hg** (millimeter of mercury)

  1 atm = 1 bar = 10^5 Pa = 100 KPa (KPa : kilo pascal)

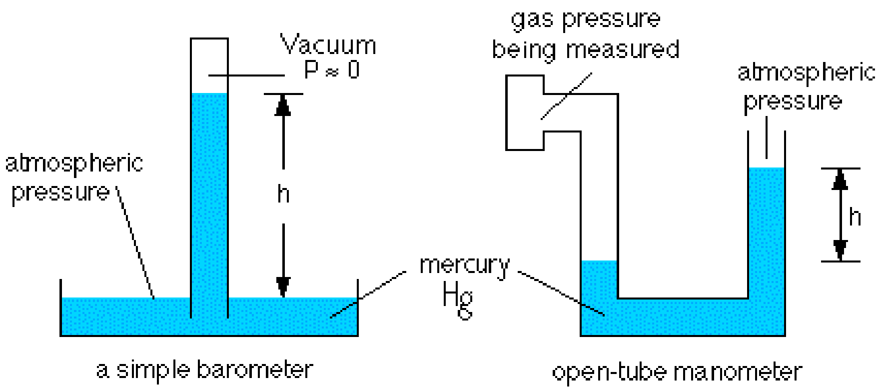
  1 torr = 1 mm Hg

  760 mm Hg (torr) = 1 atm (bar) = 10^5 Pa = 100 KPa

  - **Atmospheric pressure** : the pressure exerted by the air
  - **Standard atmospheric pressure** = Pressure at sea level = **1 atm** (= 1 bar) .
    - the devices which are used to measure the pressure , are :

      1- Barometer 2- Manometer

  - Scientists calculated the length of mercury in the Barometer tube at sea level (in which the pressure is 1 atm) , & they found it a constant # = 0.76 m = 760 mm , and they invented the common pressure unit , the (mm mercury) , to easily measure pressure (by height of the mercury). the application for this unit is the Manometer



- the Manometer measures the pressure of enclosed gases .
  (look at the manometer picture above) , the atmosphere press the mercury in the tube at the opened end & the enclosed gas do the same at the closed end , the higher pressure will move the Hg to the lower side .

- Because , as we said , 1 atm makes movement of 760 mm Hg (as in barometer) , so every 760 mm Hg moved from one side to the other means that this side is higher Pressure than the other by 1 atm :
  P(gas) = P(atmosphere**)+h** (this is according to the picture above because P(gas) > P(atm))

  the unit of P(atmosphere) is in mm mercury (which is length unit) , so after calculating P(gas) by mm mercury , we can convert it to (atm) or (Pa) to determine what is the gas pressure , this is the benefit of inventing mm mercury unit , it makes the measurement of gases pressure done directly by the mercury length & without using any advanced technological devices .

- If the mercury is moved toward the gas { P(atm) > P(gas) } then the equation becomes : **P(atm) = P(gas) + h**

(10.3) The Gas Laws

- u should know that the volume of a gas is affected by several variables (like pressure , temperature & # of gas particles) , because of that , there were many efforts spent to connect them & study their effect on gases volume .

  The following laws are the results of those efforts :

1- **Boyle's** Law (Pressure-Volume)

V relate with 1/P ----- (means if P increases , the Volume decreases)
VP = constant ----------- at constant Temperature (T) & gas moles (n)
$V_1P_1 = V_2P_2$ -------- for the same sample (same n) of gas at constant T

2- **Charles's** Law (Temperature-Volume)

V relate with T ----------------------- (means if T increases , V increases)
V=constant * T --- at constant P & n
$V_1/V_2 = T_1/T_2$ ----------------- for the same sample of gas at constant P

3- **Avogadro's** Law (moles-Volume)

V relate with n ----------------------- (means if n increases , V increases)
V = constant * n -- at constant T & P
$V_1/V_2 = n_1/n_2$ ----- for different samples of the same gas at same T & P

4- (10.4) Ideal Gas Law

PV =RnT

(R : Ideal gas constant = **0.0821** atm.L/mol.K = 8.314 J/mol.K)

(P → atm , n → moles , **T → Kelvin** not in C^o ,,, K = C^o + 273**)**

5- Combined Gas Law

here we have two conditions of the same gas sample (n1 = n2) :

(PV/T)1 = (PV/T)2

6- If we have two different samples for the same gas (n1 doesn't equal n2) :

(PV/nT)1 = (PV/nT)2

- **Avogadro's Hypothesis : equal Volumes of Gases contain the same # of molecules at constant T & P .** (n ←→ V for gases under constant T & P)
  In other words , if the equation of a reaction contains gases then u can read the gases by Volumes as well as by moles .
  Example : $2H_2(g) + O_2(g) \rightarrow H_2O(g)$ ----------------- if T & P are constant , then we can read the equation of these gases as :
  every 2 ((Liters)) of hydrogen gas react with 1 ((Liter)) of Oxygen gas to give 1 ((Liter)) of vapor water , if T & P aren't changed .
  (If u recall my first exam summary - section (3.6) **,** I mentioned this) .

- the Standard Temperature & Pressure **(STP)** is the condition when :

T = 0 C^o & **P = 1 atm** .

Examples

1- If the temperature of 10 L gas sample increases from 100 K to 400 K at 1 atm , what is the final Volume ??
Answer : firstly , the sample isn't changed ($n_1 = n_2$) & we have two conditions → use Combined Gas Law & cancel the P (because P isn't changed) or use Charles law directly , so :
(V/T)1 = (V/T)2 → V2 = V1/T1 * T2 → V2 = 10/100 * 400 = 40 L

2- Calculate the Volume of 1 mol ideal gas at 0 C^o & 1 atm (or at STP) .
Answer : VP = RnT → V=RnT/P →→→→ V=22.4 L
* this Volume is the **Molar Volume** which means it's the volume of 1 mol gas at STP (T=273 K , P=1 atm) & it is **always for all gases = 22.4** L at STP (supposing that gases are ideal , in fact the gases aren't ideal & differ slightly from this number , but in this course ,we suppose they are ideal).

3- What is the P exerted by 82.5 g of CH_4 (M.M = 16) in a 75 L container at 35 C^o ?
Answer : PV = RnT → P = RnT/V → P = R*(mass/M.M)*T/V
Continue , but make sure to convert T → P = 1.74

4- Calculate the Volume occupied by 8 g of O_2(g) at STP .
at STP → 1 mol occupy 22.4 L ,,,, lets calculate the moles
n = mass/M.M = 8/32 = .25
if 1 mol occupy 22.4 L , 0.25 mol occupy V (1 → 22.4 , 0.25 → V)
cross multiplication → V =22.4 * 0.25 = 5.6 L .

5- If 7 L of N_2 gas is reacted completely with H_2 gas at STP according to this equation : $N_2(g) + 3\ H_2\ (g)$ ---------> $2\ NH_3(g)$
calculate :
A- volume of NH_3 produced . B- Volume of H_2 required .
C- mass of NH_3 formed (M.M=17) .
Answer :
STP → constant P & T → we can read the equation as Liters other than as moles because the molecules are Gases .
A- from the equation , NH_3 Volume = ratio * N_2 volume = (2/1) * 7 = 14 L
remember from the first summary , ratio = coefficient (needed/given)
B- H_2 volume required = ratio * N_2 volume = 3 * 7 = 21 L
C- using Ideal gas Law → PV = RnT → n=PV/RT → mass/M.M = PV/RT
→ mass = M.M*PV/RT ……… continue . mass = 10.62 g

6- If 2 L of CO(g) are mixed with 2 L of O_2(g) reacting according to this equation : $2\ CO(g) + O_2(g)$ ------> $2\ CO_2$
what is the resulting total volume of gases after the reaction end ?

Answer : (read the substances by volumes because they are gases)
here we are facing a Qs of limiting reactant .
Because every 2 liter CO react with 1 liter O2 (according to the equation) , 1 liter O_2 will remain without reacting to the end of the reaction .
now the limiting reactant is CO(g) , so the volume of CO_2(g) produced is as following :
2L CO ---give---> 2L CO_2 (according to equation)
we have 2L of CO so **2L** of CO2(g) **produced** at the end of reaction .
but we end with **1L** of O_2 also , which didn't react ,
→ the total Volume is = 3 Liters :D

7- If this reaction : **I_2(g) + 5 F_2(g) → 2 IF_5(g)** take place , calculate the volume of F_2 measured at 37 C^o & 705 torr that is needed to react completely with 350 g I_2 , (M.M for I_2 = 253.8) .
Answer :
we need to find the volume of I_2 at 37 C^o & 705 torr , because if we find it we can answer the Qs directly according to the equation by reading it by volumes (because the substances are gases) .
Now to find the volume of I_2 , we use the Ideal gas law :
$V(I_2) = RnT/P$ → $V(I_2) = R*mass*T/(M.M*P)$ →
$V(I_2) = 0.0821 * 350 * (37+273)/(253.8*705/760) = 37.8$ L
* notice that we converted C^o to K by adding 273 to it , & torr to atm through division by 760 .
then from the equation , every liter of I_2 react with 5 liters of F_2 , so :
volume of F_2 = 5 * $V(I_2)$ = 5 * 37.8 = 189 L .

8- If a Temp. of a certain gas at 1 atm & 1 L volume increased from 100 K to 1000 K . what is the final volume if the pressure & moles remain constant .
Answer : (use the combined gas law , after removing P because it is constant , or use Charles law)
$(V/T)1 = (V/T)2$ → $V2 = V1 *T2/T1$ → → → V2 = 10 L

9- If a Volume of He sample originally at 25 C^o & 740 torr is reduced to 6.75L at 42 atm & 85 C^o , what is the original Volume ??

answer :   (use combined gas law)
(PV/T)1 = (PV/T)2   →→   **V1 = V2* T1/T2 * P2/P1** →→ V2 = 336.3 L

10- If 5.42 L of air measured at 7.35 torr & 23 C^o is heated to 35 C^o in the same container , what is the new pressure ??
Answer :   hints (the same container means that V isn't changed) , (use the combined gas law without V because it is constant)
(P/T)1 = (P/T)2   →→   P2 = P1 * T2/T1 →→→ P2 = 7.648 torr = 0.01 atm
* notice here that we don't have to change the unit of pressure , the rule here is that the new pressure unit will be the same as the unit used in the Law whether it's torr or atm or any other unit .

(10.5) **Ideal Gas Law Applications**

1- Relating the Gas **Density** to Gas **Molar Mass**

$\boxed{\textbf{d = M.M*P/RT}}$ (this is another form of the Ideal Gas Law)
$\boxed{\textbf{M.M = RdT/P}}$ (like V=RnT/P),(the original law → **M.M = RmassT/PV**)
*(the unit of d : **g/L** or **Kg/m^3** , if the question gives u another unit , u must convert to these units before starting solving it) .

Examples :
- what is the density of UF6(g) (M.M = 352) at 779 mm Hg , 62 C^o ??
Answer : direct use of the first law .   **d =** 13.12 g/L (or Kg/m^3)

- what is the highest gas density among the following gases at the same P & T :

	O_2	N_2	H_2	CO_2
Answer : M.M :	32	28	2	44

according to the law previously , **the more M.M , the more density** so :
$\underline{CO_2}$ > O_2 > N_2 > H_2

- what is the formula of a gas exerts a pressure of 1.4 atm at 27 C^o & has a density of 1.82 g/L at 27 C^o ?? u have these Choices :
CO_2 CO CH_4 O_2
(hint : if u want to determine the formula of substance , one way is to find its M.M then compare its M.M with the choices M.M)

Answer : after using the second law previously → the M.M = 32 , which is the M.M of O_2 .

- solve the SAMPLE EXERCISE 10.8 page 408 :D .

2- Relating the **Gas Volume** to the Amount(**mass or moles**) **of another substance in a reaction** .

* in the reactions that have gases & solid/liquid components , if we know info about the gases we can determine the moles or even the mass of the solid/liquid components that will be reacted or produced .

Example :

- If this reaction happened $NaN_3(s) \rightarrow 2\ Na(s) + 3\ N_2(g)$, & the resulted nitrogen has a volume of 36 L at 1.15 atm & 26 C^o , find the mass of NaN_3 that decomposed (M.M for NaN_3 = 65) .

Answer : (hint : find the N_2 moles firstly by the ideal gas law , then use equation #3 in my Ch3 of first exam summary) which is :

mass (needed) = M.M(needed) * [ratio * n(given)] ,,,

here the needed substance is NaN_3 , the given one is N_2 & the ratio = 1/3 (<u>ratio = coefficient of the needed substance/given</u>)

so it becomes : mass (NaN_3) = M.M(NaN_3) * 1/3 * n(N_2)

→ mass(NaN_3) = 65 * (1/3) * n(N_2)

now u find n(N_2) from ideal gas law , then continue the solution .

(10.6) Gas Mixtures & Partial Pressures (**Dalton's Law**)

All what we considered previously took 1 gas alone & did calculations on it , here we will face questions telling us about mixtures of gases in the same container , each have its pressure come from its particles .

Dalton's Law : the Total Pressure of a mixture of gases equal the sum of the pressures of each gas if it was alone . (these pressures are called Partial Pressures) .

* the pressure of any gas sample don't affected by adding another gas to it , it is the total pressure which is affected not the partial pressures .

P(total) = P1 + P2 + …..etc
P(total) = RT/V * (n1 + n2 + …etc)
P(total) = RT/V * n(total)

- If u focus on the third form of Dalton's Law , u will clearly see that the difference in gases type don't affect the pressures they exert , it is their moles (# of particles & not the M.M or the weight of gases) which produce the pressure , so 1 mol H_2 (the lightest gas) will give the same pressure of 1 mol Wf6 (the heaviest known gas) at the same T & P. (O.o) ;)

Example : a gaseous mixture made from (6 g O_2 + 9 g CH_4) is placed in a 15L vessel at 0 C^o .
find the partial pressures of O_2 & CH_4 ,,,,, & the total pressure .
Answer : (hint : partial pressure is the normal pressure calculated by the ideal gas law) ,,, → P(O_2) = RnT/V → P(O_2) = R*mass*T/(M.M*V)
→ → → P(O_2) = 0.28 atm
P(CH_4) = R*mass*T/(M.M*V) → → → P(CH_4) = 0.84 atm
P(total) = RT/V * (n1 + n2) →
P(total) = RT/V * [(mass/M.M)**O2** + (mass/M.M)**CH4**] → P(total) = 1.12 atm
Or
P(total) = P(O_2) + P(CH_4) = 1.12 atm

Partial Pressures & Mole Fractions

The idea here is that because all the gases are located in the same container , they have the same volume so :

P1 / P(total) = n1 / n(total)

n1/n(total) = mole fraction for substance 1 = X1

P1 = X1 * P(total) ←→ **X1 = P1 /P(total)**

X1 + X2 + ….. = 1 (all the fractions of anything make 100% =1)

P1/P2 = n1/n2 (when T isn't changed , V already constant)

Example : A gaseous mixture contains N_2 , Ne , He . If the partial pressure for N_2 = 0.4 atm , Ne = 1.2 atm & He = 0.4 atm . Calculate the mole fraction for each gas .

Solution : P(total) = P1 + P2 + P3 = 2 atm → $X(N_2) = P(N_2)$ /P(total) … Continue to the rest.

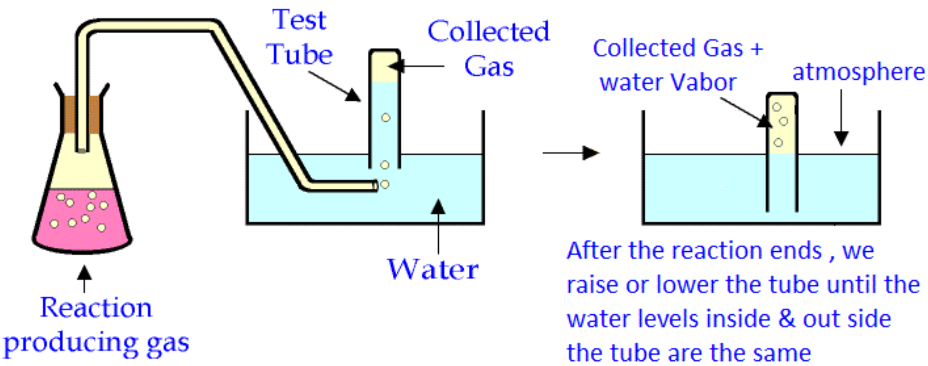
Example : U have a mixture of H_2 & He in a flask at 25 C°, $P(H_2)$ = 25 torr , P(He) = 150 torr . If u have 0.56 g He , How many grams of H_2 are present ?

Solution : P1/P2 = n1/n2 →

$P(H_2)/P(He) = mass(H_2)/M.M(H_2) * M.M(He)/mass(He)$

→ $25/150 = mass(H_2)/2 * 2/0.56$ →→→ $mass(H_2) = 0.093$

Collecting gases over water



- Look at the picture , u see a reaction produces a gas , this gas then go through a tube into a glass of water , then collected in the top of upside down inserted test tube inside the water , after the end of reaction , we remove the tube & the gas producing reaction , & try to equalize the levels of water inside & outside in order to what ???

  When the levels of water are the same then :

  $P(atm) = P(gas) + P(H_2O_{vapor})$ → $P(gas) = P(atm) – P(H_2O)$

  Now apply ideal gas law on the gas :

  P(gas) V = RnT → n(gas) = PV/RT

  → $\boxed{\mathbf{n(gas) = [P(atm)-P(H_2O)]*V/RT}}$

What is the purpose , & why all these calculations ??!!!

using this method , we can find the mass of a solid or liquid substances that are reacted & gave a gas . this is done by determining the moles of the gas collected (using this equation) , then use this # of moles to know the moles of the solid or the liquid reacted → finally their masses .

Example : $2\ KClO_3$ ------> $2\ KCl(s) + 3\ O_2(g)$, If the volume of O_2 collected over H_2O is 0.25 L at 26 C^o & 765 torr , How many grams of $KClO_3$ (M.M = 122.6) were decomposed ??

Given that the partial pressure of H_2O at 26 C^o = 25 torr .

Solution : (hint : we will find $n(O_2)$ then use equation #3 in CH3 summary)

$n(O_2) = [P(atm) - P(H_2O)] *V/RT$ →

$n(O_2) = [(765-25)/760] *0.25/(0.0821*299) = 0.00992$ mol

mass(needed) = M.M * [ratio*n(given)]

mass($KClO_3$) = M.M * [ratio*n(O_2)] →

mass(KClO3) = 122.6 * (2/3 * 0.00992) = 0.81 g

(10.8) Molecular Effusion & Diffusion

- **Diffusion :** The spread of a substance through space .

  we refer to the speed of molecules in space by the average speed (**U_{rms}**)

  $\boxed{U_{rms} = \sqrt{\frac{3RT}{M.M}}}$ **,,,** R : rate constant (**8.314 J/K.mol**) , T : temp in K

  Unit of U_{rms} is J/Kg , we can convert it to m/s as the following example .

  Example : Calculate the average speed (m/s) of NH_3 molecules at 25 °C given that (M.M of NH_3 =17 g/mol) , (1 J = 1 kg . m^2/s^2)

  Answer :

  in such questions , u must convert the unit of U from J/g to m/s

  $$U = \sqrt{\frac{3RT}{M.M}} = \sqrt{\frac{3*\left(8.314\frac{J}{K.mol}\right)*(298\ K)}{17\ g/mol}} = \sqrt{437.22\,\frac{J}{g}} = \sqrt{437.22\,\frac{1000g.m2/s2}{g}}$$

  = 661.2 m/s

- from the previous law , **the less gas M.M , the higher its rate or speed** .

  Example : which of the following gases effuses fastest ?

	CO_2	NH_3	N_2	Cl_2	
Solution : M.M :	44	17	28	71	so → <u>NH_3</u> > N_2 > CO_2 > Cl_2

- **Effusion** : escaping of gas particles through atiny holes into a lower pressure area .

Graham's Law of Effusion

$$\boxed{\frac{r1}{r2} = \sqrt{\frac{M.M2}{M.M1}}}\quad \text{(r : rate)}$$

$$\boxed{\mathbf{M.M2 = M.M1 * (r1/r2)^2}}$$

Example : the Effusion rate of gas-A is 1.14 times that of SO_2 (M.M=64) . Find M.M(A) . (hint : this info means $r(A)/r(SO_2) = 1.14$)

Solution : $M.M(A) = M.M(SO_2) * [r(SO_2)/ r(A)]^2$ →→→ M.M(A) = 49 g/mol

rate = Length travelled/time ,,, L :length t :time ,,, **then :**

$$\boxed{\frac{r1}{r2} = \frac{(L/t)1}{(L/t)2} = \sqrt{\frac{M.M2}{M.M1}}}$$

Or **rate = volume escaped / time** ,,, V :volume t :time ,,, **then** :

$$\boxed{\frac{r1}{r2} = \frac{(V/t)1}{(V/t)2} = \sqrt{\frac{M.M2}{M.M1}}}$$

Example : which of the following gases travels 2 m in 2 seconds through a tube , if u know that H_2 travels 46.9 m in 10 sec ?? N_2 NH_3 Cl_2 CO_2 (the M.M for them mentioned in a previous example)

Solution : $\frac{(L/t)1}{(L/t)2} = \sqrt{\frac{M.M2}{M.M1}}$ → $\frac{(46.9/10)}{(2/2)} = \sqrt{(M.M_2/2)}$

→→→ M.M will be 44 g/mol , so the gas is **CO_2** :D

Example : it takes 192 sec for 1.4 L of unknown gas to effuse through a porous wall & 84 sec for the same volume of N_2 to effuse through the same wall at the same T & P . what is the molar mass of the unknown gas ?

Answer : $\frac{(V/t)1}{(V/t)2} = \sqrt{\frac{M.M2}{M.M1}}$,,,,,, (V1 = V2)

→ $\frac{t2}{t1} = \sqrt{\frac{M.M2}{M.M1}}$ → $192/84 = \sqrt{(M.M_2/28)}$

→→→ M.M will be 146 g/mol

Example : If methane (CH_4) effuses 3.3 times faster than $Ni(CO)_X$.
What is the value of x ??
(M.M of CH_4 = 16) , (M.M of Ni = 58.7) , (M.M of CO = 28)

Answer : one way to determine x is by knowing the M.M of $Ni(CO)_X$, we achieve that using Graham's Law :

M.M2 = M.M1 * (r1/r2)² → M.M2 = 16 * $(3.3/1)^2$ = 174.24 g

then : M.M of $Ni(CO)_X$ = 174.24 = (M.M of Ni) + X*(M.M of CO)

174.24 = 58.7 + X*28 → **X = 4.1 ~ 4**

End Of Chapter 10

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Message me on Facebook ((Faris Alshboul)) for any clarification

العِلْمُ يُحْيِي قُلُوبَ المَيِّتِينَ كَما
يُحْيِي البِلَادَ إِذَا ما ماتَتِ المَطَرُ
والعِلْمُ يَجْلُو العَمَى عَنْ قَلْبِ صاحِبِه
كَما يُجَلِّي سَوَادَ الظُّلْمَةِ القَمَرُ

سامحوني على اي اخطاء لغوية او حسابية ويا ريت تبلغوني برسائل عنهم حتى اعدلهم

والسلام عليكم ورحمة الله وبركاته 0:)