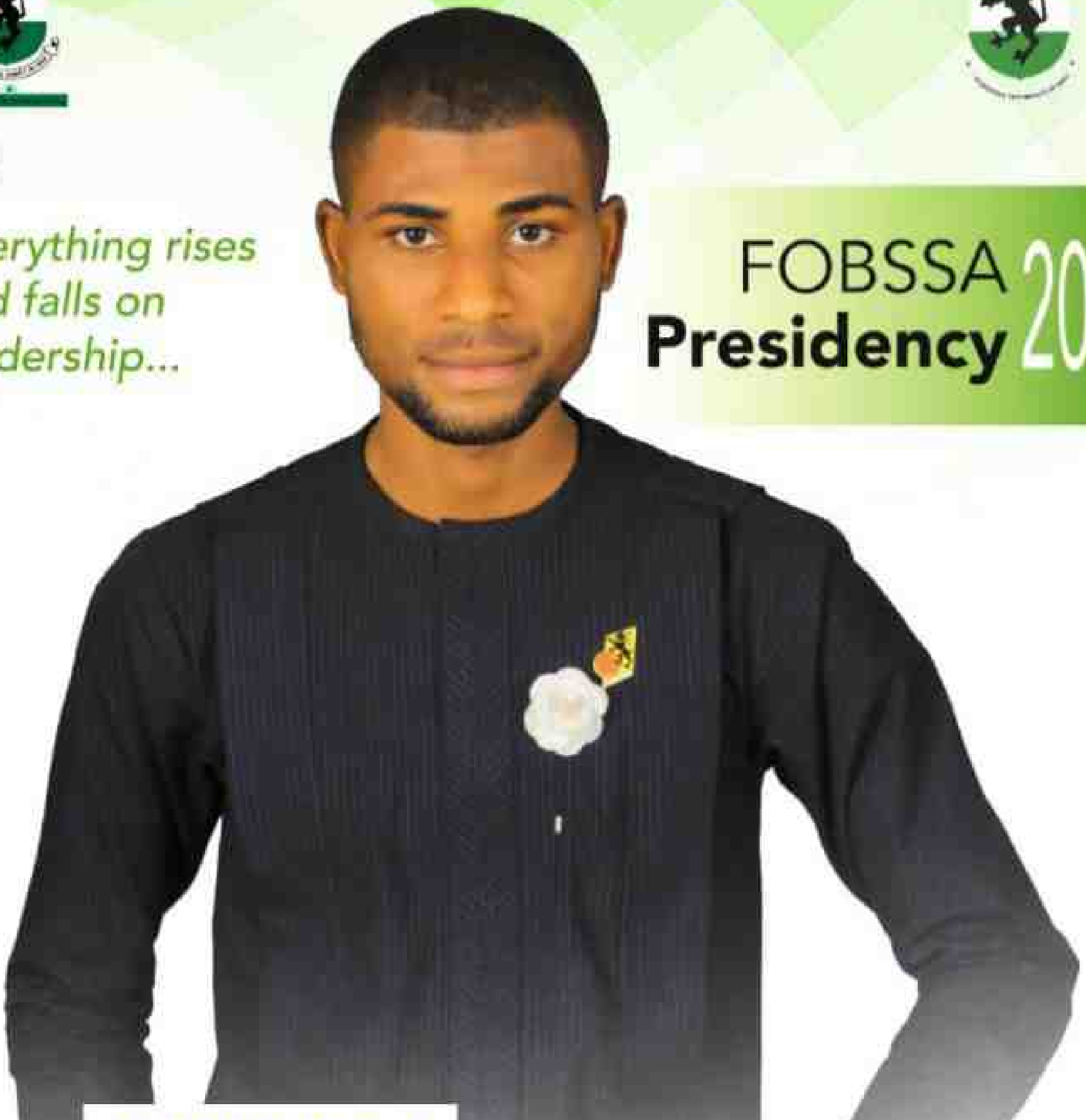
Everything rises and falls on leadership...

FOBSSA **Presidency** 2020



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FACULTY OF BIOLOGICAL SCIENCE STUDENTS ASSOCIATION **(FOBSSA)**

*The Man with the **FOBSSA**ites needs at heart...*

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CHEM 101

INTRODUCTORY CHEMISTRY

Chemistry is the part of natural science that deals with the nature and Composition of matter as well as the changes matter undergoes along with the energy assocuated with these changes.

SCIENTIFIC METHOD

The emergence of a Chemistry Concept is usually not an abrupt process. It is a steady process that passes through some developmental Stages which is well organised and referred to as "SCIENTIFIC METHOD".

The Scientific method is an empirical method of knowledge acquisition which has characterised the development of natural for several decades. It Involves Careful observation which may even lead to doubt about what is observed, this is because Cognitive assumptions about how the world works influences how one interpretes what is perceived. The Scientific method is not necessarily Stepwise but rather a gradual process of Creative thinking and testing which leads to objective Verifiable discoveries about nature.

The basic Concept of Scientific method Includes:

① FACTS:
These are basically observed phenomenon that occur naturally and are induced by experimentation. They are realities which is always the Same under the Same Condition.

② CONCEPTS: These are ideas that are Suggested Continually and once Valid, they are retained based on quantitative experimentation. If they Can't be Validated, they Can be rejected.

③ OBSERVATIONS:
These are basic to Scientific thinking and are the facts that existing ideas must explain. Useful observation must be quantitative and give room for Comparison. It should be able to Supply piece of Information that Can be Compared with other facts.

④ HYPOTHESIS
It is a proposal made to explain an observation. The inquisitive nature of human beings makes them Come up with questions about things that they See or hear which may lead to the development of ideas

or hypothesis. The most Conclusive testing Comes from reasoning based on Carefully Controlled experimental data.

Consequently, hypothesis is derived from the result of experiment and it has to be Consistent with the experimental result before it Can be acceptable depending on how well additional test match the predictions. A hypothesis may require refinement alterations, expansion or even rejection, however, if a parti-Cular hypothesis becomes Very well Supported, a general theory may be developed.

NOTE: Avogadros hypothesis, Gay-lussac law are not yet modified.

⑤ EXPERIMENT:

This is a Clear Set of procedural Steps that test hypothesis. It helps to determine whether observation agree or Conflict with the prediction derived by a hypothesis. For a hypothesis to Stand the test of time, it must Confirm the experimental results.

NOTE: Usually, hypothesis are altered to Conform to the experimental results Conducted to Verify them, but the results Cant be altered

Experiment is the key to Scientific method. In the actual Sense, when an experiment is Conducted using the Same procedural steps, the results obtained from a particular experiment should be reproduceable. Scientific method is a method of Systematic observation measurement and experiment and the formulation, testing and modification of hypothesis. This method is Characterized as natural Science over the decades.

6 LAW / PRINCIPLE (MODEL)

These are usually a Statement of generalized that relate to established fact.

Anthoine Lavoisier was the first person to talk about the matter, that matter Can neither be Created nor destroyed in 1774.

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MOLE $ STOICHIOMETRY

Mole is the unit of measurement of Substance. It is one of the Seven (7) based S.I. unit and it is a measure of amount of substance. It is defined as the amount of Substance that Contains the same number of entities as there are atoms in exactly 12g of Carbon (12). The number is Called "Avogadro's number". Therefore, we Can Say that one mole Contains 6.022×10^{23} entities. Entities Could be ions, molecules, atoms, or formular unit. One Mole of Carbon will Contain; 6.022×10^{23} Carbon (12) atoms. It follows that the mole is a Counting unit just like dozen but its different from dozen, in the sense that it Specifies the number of a Substance in a fixed mass of Substance, therefore, One (1) mole of a Substance represents a fixed number of Chemical entities and has a fixed mass. Atoms have fixed mass, therefore, the mole unit allows us to determine the number of atoms, molecules, formular unit in a Sample by weighing it.

One mole of Iron (Fe) atoms has an atomic mass of 58.85 atomic mass unit (A.M.U.) and for Sulphur (S), it has 32.07 (A.M.U.).

Mole Concept is Introduced to be able to quantify how much a mass of an element, Compound(s) or molecule is, by Avogadro in the year 1811.

1 mole $= 6.022 \times 10^{23}$ particles/molecule

Elements Combine in mole ratio.

The mass of one (1) mole of a Substance is its MOLAR MASS.

$$n = \frac{\text{Number of Particles of a Substance}}{\text{Number of Avogadro } (N_A)}$$

where n => mole

∴ Amount is measured in mole

1 mole of X = Molar mass of X

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molar mass}}$$

Examples:

① How many moles of Na^+ ion is produced by dissolving 7.5g of Na_2CO_3 in water?

Solution

$$Na_2CO_3 + H_2O \longrightarrow 2Na^+$$

$$Na_2CO_3 \Rightarrow (23 \times 2) + (12) + (16 \times 3)$$
$$\Rightarrow 106 g/mole$$

Number of mole of Na_2CO_3

$= \frac{\text{mass of } Na_2CO_3}{\text{Molar mass}}$

$\Rightarrow \frac{7.5g}{106 g/mole}$

$\Rightarrow 0.07$ mole

But from the equation

1mole of Na_2CO_3 gives $2Na^+$

0.07 mole of $Na_2CO_3 \Rightarrow (2 \times 0.07)$

$\Rightarrow 0.14$ mole

OR

Say, 106g of $Na_2CO_3 \Rightarrow$ 2 mole of Na^+

7.5g of $Na_2CO_3 \Rightarrow \left(\frac{2 \times 7.5}{106}\right)$

$= 0.14$ mole

② In the electrolysis of molten aluminium oxide. How many electrons is required to increase the mass of the Cathode by 2.7g
($Al = 27$, $O = 16$)

Solution

	Cathode	Anode
Al_2O_3	Al	O

$Al_2O_3 \rightleftharpoons 2Al^{3+} + 3O^{2-}$

$Al^{3+} + 3e^- \rightarrow Al_{(s)}$

1mole of Aluminium $\Rightarrow$ 3 mole of electrons

27g of Aluminium $\Rightarrow 3 \times 6.02 \times 10^{23}$ electrons

2.7g of Aluminium will be

[Cross multiply]

$\Rightarrow \frac{3 \times 6.02 \times 10^{23} \times 2.7}{27}$

$\Rightarrow 180.6 \times 10^{22}$ electrons.

OR

$nAl \Rightarrow \frac{2.7}{27} \Rightarrow 0.1$mole

$ne^- \Rightarrow 3 \times$ number of mole

$\Rightarrow 3 \times 0.1$

$\Rightarrow 0.3$ mole

Number of electrons/particles/atom

= mole X avogadro Con tant.

$\Rightarrow 0.3 \times 6.02 \times 10^{23}$

$\Rightarrow 180.6 \times 10^{22}$ electrons

N:B number of moles, n, = Concentration $\left(\frac{mole}{dm^3}\right)$ X Volume (dm^3)

MOLAR VOLUME

Molar Volume is the Volume of 1mole of a gas at S.t.p.

It is given as : $22.4dm^3 = 22400cm^3$

$n = \frac{V}{\bar{V}} (dm^3)$ where

V = molar Volume of given gas

$\bar{V}$ = molar volume of a gas at S.t.p. $(22.4dm^3)$

If you are not asked to Calculate Volume in S.t.p, you Can use the general gas equation

QUESTIONS

① What Volume of 0.5M HCl will liberate $100cm^3$ of CO_2 at S.t.p when reacted with Na_2CO_3.

Solution

$$2HCl + Na_2CO_3 \downarrow 2NaCl + CO_2 + H_2O$$

If $2HCl \rightarrow 1CO_2$ (from equation)

Remember $0.5M HCl \Rightarrow 0.5mol/dm^3$

$0.5mol/dm^3$ of HCl $\Rightarrow 0.5mol \longrightarrow 1000cm^3$

$\frac{0.5x}{1000}mol \longleftarrow x\,cm^3$ of HCl

1mole of $CO_2 \rightarrow 22400cm^3$ at $\boxed{S.t.p}$

But from equation

2moles of HCl $\rightarrow$ 1mole of CO_2

$\frac{0.5x}{1000}$ mole $\rightarrow \frac{0.5x}{1000} \times \frac{1}{2}$

$= 0.00025x$ mole of CO_2

Since 1mole of $CO_2 = 22400cm^3$ at S.t.p

$0.00025x$ mole of $CO_2 = 100cm^3$

Cross multiply to find x ie the Volume of HCl required.

$\Rightarrow 0.00025x \times 22400 = 100$

$\Rightarrow x = \frac{100}{(0.00025 \times 22400)}$

$x \Rightarrow 17.86cm^3$ at S.t.p

Therefore Volume of 0.5M HCl required to liberate $100cm^3$ of CO_2

$\Rightarrow 17.86cm^3$ at S.t.p.

OR

number of moles of CO_2 $= \frac{V}{\bar{V}} = \frac{100cm^3}{22400cm^3}$

$= 0.0045$

number of moles of HCl $= 2 \times$ no of moles of CO_2 (ie from equation)

HCl : CO_2

2 : 1

no of moles of HCl

$= 0.0089$ moles

Recall $n = CV'$ where C = Molar Conc. of HCl

V' = its molar Volume.

Since Its the Volume we are to find

$$\frac{n}{C} \Rightarrow V' \Rightarrow V' = \frac{0.0089}{0.5}$$

$$= 0.017857 dm^3$$

Recall $0.017857 dm^3 = 1000 Cm^3 \times 0.017857$

$$= \underline{\underline{17.86 Cm^3}}$$

② A manganese is a transition metal element essential for the growth of bones. What is the mass in grams of 3.22×10^{20} manganese atoms?

Solution

1mole of Manganese $= 55g = 6.02 \times 10^{23}$ atoms

$X = 3.22 \times 10^{20}$ atoms

$$X = \frac{3.22 \times 10^{20} \times 55}{6.02 \times 10^{23}}$$

$$= 29.419 \times 10^{-3} g$$

$\Rightarrow 0.0294g$ of Manganese.

③ Using anhydrous metaphosphoric acid, what mass of metaphosphoric acid will be used to prepare 0.75M metaphosphoric acid in Standard quantity? (Molar mass of metaphosphoric acid $= 79.979 g/mol$)

Solution

Molarity (Concentration) $\Rightarrow 0.75M$

Recall

$$\text{molarity} = \frac{\text{mass}}{\text{molarmass}} \times \frac{1000}{\text{Volume}(Cm^3)}$$

Since it is in Standard quantity its Volume will be $1000 Cm^3$

$$\Rightarrow 0.75 = \frac{m}{79.979} \times \frac{1000}{1000}$$

$$\Rightarrow m = 0.75 \times 79.979$$

$$= 57.98g$$

OR

Mass = molarity x molarmass (Since it is in standard quantity)

$= 0.75 \times 79.979$

$\Rightarrow 59.98g$

SUMMARY

N.B $\% \text{ by mass} = \frac{\text{mass given}}{\text{molarmass}} \times \frac{100\%}{1}$

MOLAR MASS

Molar mass of a Substance is the mass per mole of its entity and its unit is in gram per mole (g/mole).

NATURE OF ELEMENTS

There are two types of elements

Ⓘ Monoatomic element

Ⓘⓘ Molecular elements

① Monoatomic elements : These are elements that occur as individual elements. Examples are Na, Cu, Ne, etc.
(Sodium) (Copper) (Neon)

The molar mass of monoatomic elements are numerical value for the periodic table expressed in g/mol.

Neon (Ne) = 20.18 ; Gold (Au) = 197.0
Iron (Fe) = 55.85

② Molecular elements:
These are elements that occur as molecules. You must know the molecular formular to determine the molecular mass.
Examples are: Oxygen molecule (O_2) Hydrogen molecule (H_2), Nitrogen molecule (N_2).

MOLAR MASS OF COMPOUNDS

The molar mass of a compound is the sum of the molar masses of the atoms of the elements in a formular. Examples are:

Molar mass of SO_2 = 64.07g/mol,
Oxygen molecule molar mass = 32g/mole
Glucose molar mass ⇒ $C_6H_{12}O_6$
$(12\times6)+(1\times2)+(16\times6)$
= 180g/mol.

PERCENTAGE COMPOSITION OF COMPOUNDS

Examples

① What mass of Copper(II) Sulphate pentahydrate ($CuSO_4\cdot5H_2O$) must you dissolve in 200g of H_2O such that each gram of the resultant solution contains 0.05g of Cu^{2+}?

Solution

Solute ⇒ $CuSO_4\cdot5H_2O$
Molar mass of $CuSO_4\cdot5H_2O$
$\Rightarrow (63.5+32+64+(5\times18))$
$\Rightarrow 249.5$g/mol

63.5g of Cu is found in 249.5g of $CuSO_4\cdot5H_2O$
0.05g of Cu is found in ?
$$\Rightarrow \frac{249.5\times0.05}{63.5}$$
⇒ 0.196g of $CuSO_4\cdot5H_2O$

Mass of Solvent + Mass of Solute
= Mass of Solution.
Mass of Solution was given as 1g (ie from each gram)

Mass of Solvent = 1 − 0.196g
⇒ 0.804g of H_2O
If 0.804g of H_2O dissolves 0.196g $CuSO_4\cdot5H_2O$
200g of H_2O dissolves $\frac{200\times0.196}{0.804}$
= 48.76g $CuSO_4\cdot5H_2O$

∴ The mass of $CuSO_4\cdot5H_2O$ ⇒ 48.76g

② An iron Complex with molecular formula $C_{34}H_{32}FeN_4O_4$ Constitutes 6% of haemoglobin by mass. If a blood Sample Contains 0.75g of haemoglobin, what mass of Iron is Contributed by the Complex?
(C=12, H=1, Fe=56, N=14, O=16)

Solution

$C_{34}H_{32}FeN_4O_4$ by mass is formular for Heme.

Molar mass = ((12x34)+(1x32)+(56)+(14x4)+(16x4))
= 616g/mol

formular mass = 616g //

Mass of haemoglobin = 6% of 616

$= \frac{6 \times 616}{100} = 36.96g$

616g of Heme Contains 36.96g of Haemoglobin

$\frac{616 \times 0.75}{36.96} \leftarrow$ 0.75g of Hemoglobin

= 12.5g of $C_{34}H_{32}FeN_4O_4$

But Iron = 56g in 616g of $C_{34}H_{32}FeN_4O_4$

⇒ Iron Content in 12.5g of $C_{34}H_{32}FeN_4O_4$

$\Rightarrow \frac{56}{616} \times 12.5g$

= $\underline{\underline{1.14g \text{ of}}}$ Fe (Iron)

③ A 550mg tablet Containing Vitamin C $C_6H_8O_6$ was dissolved in water and reacted with bromine and the reaction is

$C_6H_8O_6 + Br_2 \rightarrow C_6H_6O_6 + 2HBr$.

The resulting Solution was treated with 400cm³ of 0.145M Solution of NaOH

$HBr + NaOH \rightarrow NaBr + H_2O$

(C=12, Na=23, O=16, Br=79, H=1)

① Calculate the percentage of Vit C in the tablet

(ii) Calculate the percentage of Carbon in the Vitamin C Content of the tablet.

Solution

Molar Conc. of NaOH ⇒ 0.145M

Amount in 400cm³ ⇒ $0.145 \times \frac{400}{1000}$

= 0.0058 moles.

From equation 1mole HBr = 1mole NaOH

0.0058mole HBr = 0.0058mole NaOH

From equation 1mole of $C_6H_8O_6$ = 2mole HBr

$\frac{0.0058}{2}$ mole of " = 0.0058mole HBr

⇒ 0.0029mole of $C_6H_8O_6$ (Vitamin C)

Molar mass of Vit. C = (12x6) + (8x1) + (16x6) =
⇒ 176g/mol.

Mass of Vit. C = molar mass x mole of Vit. C

⇒ $\underline{\underline{176 \times 0.0029}}$

$\Rightarrow$ 0.510g of Vit. C

$\Rightarrow$ 510mg of Vit. C

① percentage of Vitamin C in the tablet

$$\Rightarrow \frac{\text{mass of Vit.C}}{\text{Total mass of tablet}} \times \frac{100}{1}$$

$$\Rightarrow \underline{\underline{92.8\%}}$$

⑪ To get the mass of Carbon you have to use the gravimetric mass of Carbon with respect to Vitamin C.

Mass of Carbon in Vitamin C

$$= \frac{(12\times6)}{((12\times6)+(1\times8)+(12\times6))} \times 510.4$$

$$\text{ie } \frac{\text{formular mass of } C_6}{\text{Formularmass of } C_6H_8O_6} \times \frac{510.4}{1}$$

$$\Rightarrow \frac{72}{176} \times 510.4 = 208.8\text{mg of Carbon}$$

Therefore Percentage of Carbon in the Vitamin C Content of the tablet

$$\Rightarrow \frac{\text{mass of Carbon}}{\text{Mass of Vit.C}} \times \frac{100}{1}$$

$$= \frac{208.8}{510.4} \times \frac{100}{1} = \underline{40.91\%}$$

④ Narceine is a narcotic in opium. It Crystallizes from water Solution as a hydrate that Contains 10.8% water. If the molarmass of Narceine hydrate is 499.52g/mol. Determine X in Narceine. XH_2O.

Solution

Molar mass of Narceine $\cdot XH_2O = 499.52$ g/mol.

molar mass of $XH_2O = 18X$ g/mole

% of water $= 10.8\%$

% of Narceine $\cdot XH_2O \Rightarrow 100\%$

$$\text{Formular } \frac{\text{molar mass of } XH_2O}{\%\text{ of water}} = \frac{\text{Total molarmass (Narceine}\cdot XH_2O)}{100\%}$$

$$\Rightarrow \frac{18x}{10.8} = \frac{499.52}{100}$$

$$\Rightarrow 2.997 = x$$

$$X \simeq 3$$

ie its Narceine trihydrate.

STOICHIOMETRY

Stoichiometry is the study of mass–Volume relationship chemical reaction. It is the study of quantitative aspect of Chemical relations.

To Solve any problem in Stoichiometry there must be a balanced reaction equation. It shows how reactants react in Stoichiometric quantities to form products.

In Stoichiometry, the following are possible;

① mass – mass relationship
(ii) Mass – Volume
(iii) Mass – no of mole
(iv) mass – no of molecule (particles)

In Stoichiometry, 1st thing to do is to balance the equation

Examples

① 5g of Fe was reacted with Sulphur to form FeS. What is the mass of FeS produced?

Solution

molar mass of FeS = 88g

$Fe + S \longrightarrow FeS$

$\Rightarrow$ 56g of Fe $\longrightarrow$ 88g of FeS

5g of Fe $\Rightarrow \left(\frac{5 \times 88}{56}\right)$ of FeS

$\Rightarrow$ 7.85g of FeS was produced.

This is an expression of Simple mass – mass relationship.

② 5g of Fe was reacted with 5g of Sulphur to form FeS. What mass of FeS is produced?

Solution

This question is Very tricky. It is not same with question no 1 we Solved. To Solve this we must understand what limiting reactant and excess reactant are all about.

Limiting reactant is the reactant that limit the amount of Product to be formed. It's the reactant that will form the least amount of Products. It is the reactant which once consumed the reaction ends.

Once two amounts of the reactant are given, you must find the limiting reactant. If you know your limiting reactant, the other reactant will be excess reactant.

A trial Calculation must be done.

lets See how

Trial Calculation

$Fe + S \rightarrow FeS$ — — ①

56g 32g 88g

Since the question said 5g of Fe reacted with 5g of Sulphur, we want to Check if the reaction is ok!!! or if any of the reactant is in excess. To do this, we must make use of the equation (well balanced equation).

N.B our ① is Correct balanced!!!

To know the limiting reactant.

56g of Fe reacted with 32g of Sulphur [from equation]

5g of Fe " "

N:B
If after Cross multiplication and it give us 5g of Sulphur, it means the reaction is ok!!! ①
But if it give us less than 5g of Sulphur, it means Sulphur is in excess & Fe is limiting ②
If it gives us greater than 5g of Sulphur, it means Sulphur is limiting & Fe is excess ③

Cross multiply

$$\Rightarrow \frac{5 \times 32}{56} = 2 \cdot 86g \text{ of Sulphur}$$

From the answer above (ie 2·86g of Sulphur) it Satisfies ② ie Sulphur is in excess Since only 2·86g of Sulphur out of 5g of Sulphur reacted.

Excess Sulphur = 5g − 2·86 = 2·14g

So Fe (iron) is limiting reactant ok!!!
Sulphur excess reactant ok!!!

N:B it is only the limiting reactant that you must use to get the mass or Volumes of the product formed. Please, dont use the excess reactant.

Since our limiting reactant is Fe

Say 56g of Fe → 88g of FeS

$$5g \text{ of Fe} \rightarrow \frac{5 \times 88}{56}$$

$$= 7 \cdot 86g \text{ of FeS}$$

Mass of FeS produced = 7·86g

③ Strontium hydride (5·70g) was made to react with 4·75g of H_2O according to the following reaction equation

$$SrH_2 + 2H_2O \rightarrow Sr(OH)_2 + 2H_2$$

Sr = 87·62g/mol

ⓐ How many gram of hydrogen can form from given mass of SrH_2 hence what is its mole?

ⓑ How many gram of hydrogen can be formed from a given mass of H_2O hence what is the mole of hydrogen formed?

ⓒ Which is the limiting reactant and excess reactant?

ⓓ What Volume of H_2 is produced at s.t.p?

ⓔ How many gram of H_2 can be formed and what is its mole?

Solution

Check if the equation is balanced.

$$\underset{89.62g}{SrH_2} + \underset{36g}{2H_2O} \rightarrow \underset{121.62g}{Sr(OH)_2} + \underset{4g}{2H_2}$$

Formular mass of $SrH_2 \Rightarrow (87 \cdot 62 + (2 \times 1)) = 89 \cdot 62 g/mol \times 1 mole$

$= 89 \cdot 62g$

Formular mass of $2H_2O \Rightarrow 2(16+2) = 36g$

Formularmass of $Sr(OH)_2$ = 1mole × 121.62g/mol = 121.62g

Formular mass of $2H_2 \Rightarrow 2(2\times1) = 4g$

ⓐ

To get the gram $ mole of H_2 formed from SrH_2. N.B [it is specifed that H_2 is formed from SrH_2]

So Say 89.62g $SrH_2 \longrightarrow$ 4g of H_2

5.70g $SrH_2 \longrightarrow \frac{4\times5.7}{89.62}$

= 0.254g of H_2

To Convert it to mole, divide with molar mass of H_2 ie $\frac{0.254g}{2g/mol} \Rightarrow$ 0.127mole

ⓑ

To get the gram $ mole of H_2 formed from H_2O. [Note: it is also specified that H_2 is formed from H_2O]

So Say 36g of $H_2O \longrightarrow$ 4g of H_2

4.75g of $H_2O \longrightarrow \left[\frac{4\times4.75}{36}\right]$

= 0.53g of H_2

To Convert to mole $\Rightarrow \frac{0.53}{2}$

= 0.26 mole of H_2

ⓒ

To get the limiting and excess reactant Compare the 2 reactants in the equation.

Say 89.62g of $SrH_2 \longrightarrow$ 36g of H_2O

5.7g of $SrH_2 \longrightarrow$ [(if it give 4.75g of H_2O, the reaction is OK!!! (no excess, no limiting)
if it give < 4.75g of H_2O, the water is excess, then SrH_2 Limiting)
if it give > 4.75g of H_2O, the H_2O is limiting, then SrH_2 excess)]

Cross multiply

$\Rightarrow (5.7\times36) \div 89.62 \Rightarrow$ 2.29g of H_2O ie it is less than 4.75g of water given.
Hence SrH_2 is limiting and water (H_2O) is excess.

ⓓ To get the volume of H_2 produced at s.t.p, please make use of the limiting reactant only.

89.62g of $SrH_2 \longrightarrow$ 4g of H_2

5.7g of $SrH_2 \rightarrow$ 0.254g of H_2

$\Rightarrow$ 0.127mole

1mole of a gas = 22400 cm^3 at s.t.p

0.127mole of a gas $\Rightarrow$ 2844.8 cm^3 of H_2

$\Rightarrow$ 2.845 dm^3

OR

1 mole of $SrH_2 \rightarrow$ 2moles of H_2

$\left(\frac{5.7}{89.62}\right)$ mole of $SrH_2 \rightarrow \frac{2\times5.7}{89.62}$

$\Rightarrow$ 0.127 mole

∴ 1mole of a gas $\Rightarrow$ 22.4dm^3 at s.t.p

0.127 mole " $\Rightarrow$ 0.127×22.4

$\Rightarrow$ 2.85 dm^3

② A Solution of $MgSO_4 \cdot 7H_2O$ Contains 0.35g of Mg^{2+} ions.

ⓐ What is mass of water of Crystallization in this Solution.

ⓑ What is the mass of SO_4^{2-} ions in this Solution

ⓒ What is the mass of the Solute in this Solution.

Solution

ⓐ Molar mass ($MgSO_4 \cdot 7H_2O$

$\Rightarrow 24 + 32 + 64 + (7 \times 18)$

$\Rightarrow$ 246 g/mole

$MgSO_4 \Rightarrow$ 120g/mol

$Mg^{2+} \Rightarrow$ 24 g/mol

$SO_4^{2-} \Rightarrow$ 96 g/mol

To get the mass of water of Crystallization in this Solution,

24g of $Mg^{2+} \longrightarrow$ 126g of $7H_2O$

0.35g of $Mg^{2+} \longrightarrow \frac{126 \times 0.36}{24}$

$\Longrightarrow$ 1.838g of $7H_2O$

ⓑ To get the mass of SO_4^{2-} ions in this Solution

24g of $Mg^{2+} \longrightarrow$ 96g of SO_4^{2-}

0.35g of $Mg^{2+} \longrightarrow \frac{96 \times 0.35}{24}$

$\Longrightarrow$ 1.40g of SO_4^{2-} ions

ⓒ To get the mass of the Solute in this Solution.

24g of $Mg^{2+} \Rightarrow$ 120g of $MgSO_4$

0.35g of $Mg^{2+} \Rightarrow \frac{0.35 \times 120}{24}$

$\Rightarrow$ 1.75g of $MgSO_4$

③ 1.14g of PbS was reacted with 1g of Oxygen to form 1g of PbO according to the reaction

$$2PbS + 3O_2 \longrightarrow 2PbO + 2SO_2$$

(i) Which is the limiting reactant?

(ii) What mass in gram of excess reactant remain unreacted?

(iii) What is the % mass of PbO

(Pb = 207, O = 16, S = 32)

Solution

$$2PbS + 3O_2 \longrightarrow 2PbO + 2SO_2$$

mass of PbS $\Rightarrow 2(207+32) = 478g$

mass of $O_2 \Rightarrow 3 \times 32 = 96g$

mass of PbO $\Rightarrow (207+16) \times 2 \Rightarrow 446$ g

① If 478g of PbS react with 96g of O_2, 1.14g of PbS will react with what?

Cross multiply

$\frac{1.14 \times 96}{478} \Rightarrow$ 0.23g of O_2

Since 0.23g of O_2 we got is less than 1g of Oxygen, it means the Oxygen that reacted in the question given is the excess reactant (i.e it is not 1g that should react, its only 0.23g that should go into reaction.

Excess Oxygen in the reaction = 1 − 0.23

= 0.77g

Hence, the limiting reactant is PbS.

(ii) Mass of excess reactant ⇒ $\boxed{\text{Mass that reacted in the question ie 1g}}$ − $\boxed{\text{Mass that suppose to react}}$ ie 0.23g

⇒ 1g − 0.23g

⇒ 0.77g

(iii) To get the % mass of PbO, please make use of the limiting reactant -

478g of PbS → 446g of PbO

1.14g of PbS → $\frac{446 \times 1.14}{478}$

⇒ 1.06g of PbO

Since from the question 1g of PbO was formed ⇒ Actual yield ⇒ 1g of PbO

theoretical yield ⇒ 1.06g of PbO

$\Rightarrow \%\text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical}} \times \frac{100}{1}$

$\Rightarrow \frac{1}{1.06} \times \frac{100}{1} \Rightarrow 94.3\%$ of PbO

(4) Consider $Mg(OH)_2 + 2H^+ \rightarrow Mg^{2+} + 2H_2O$

If $5dm^3$ of 6M containing H^+ is reacted with 17moles of $Mg(OH)_2$

(i) Which is the limiting reactant?

(ii) What mole of excess reactant remains unreacted?

(iii) What mole of Mg^{2+} was produced

Solution

Molarity of H^+ ⇒ 6M

Volume of H^+ ⇒ $5dm^3$

no of moles = C × V

⇒ 6 × 5 = 30 moles of H^+

$2H^+$ reacts with $Mg(OH)_2$ from equation

1 mole of H^+ reacts with ½ mole of $Mg(OH)_2$

30 mole of H^+ needs $(\frac{1}{2} \times 30)$ moles of $Mg(OH)_2$

= 15 moles of $Mg(OH)_2$.

But 17 moles of $Mg(OH)_2$ reacted in the question instead of 15 moles.

Hence limiting reactants is H^+, but $Mg(OH)_2$ is excess reactant.

(ii) Mole of excess reactant ⇒ 17moles − 15moles

⇒ 2 moles of $Mg(OH)_2$

(iii) Since H^+ is the limiting reactant

If $2H^+ \longrightarrow 2H_2O$

30 moles H^+ ⟶ 30 moles of H_2O produced

If $2H^+ \longrightarrow Mg^{2+}$

30 moles H^+ ⟶ 15 moles of Mg^{2+} produced.

⑤ 6 moles of Nitrogen reacts with 14moles of hydrogen to form NH_3

(i) find the limiting reagent

(ii) find the excess reactant and the mass of excess reactant that remains unreacted.

(iii) find the residual gases.

Solution

Equation of the reaction

$$N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_3$$

Comparing we have

1 mole of N_2 ⟶ 3 moles of H_2

6 moles of N_2 ⟶ 18 mole of H_2

But 14 moles of H_2 is what reacted in the question not 18 moles of H_2. Since the 14 moles that reacted is lower than what supposed to go, into reaction, Hydrogen is the limiting reactant, hence Nitrogen is the excess rectant.

(ii) To get mass of excess of excess N_2 that remains unreacted.

1 mole of N_2 ⟶ 3 moles of H_2

4.67 moles of N_2 ⟵ 14 moles of H_2

Excess $N_2 \Rightarrow 6 - 4.67 = 1.33$ mole

∴ Mass of excess $N_2 \Rightarrow 1.33 \times 28$

$\Rightarrow 37.3g$

(iii) To find the residual gases.

from equation

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

1 mole 3 moles 2 moles

Since the limiting reactant is Hydrogen, please that is what you should use.

$$N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$$

1 mole 3 mole 2 mole

$\frac{14}{3}$ mole 14 mole $\frac{14}{3} \times 2$ mole

$= 9.33$ mole

∴ Residual gases ⇒ Gas produced + excess gas that is unreacted

⇒ 9.33 mole of NH_3 + 1.33 mole of N_2

⑥ Aspirin is made by reacting Salicylic acid with acetic anhydride according to the following reaction equation.

$$C_7H_6O_3 + (CH_3CO)_2O \rightarrow C_9H_8O_4 + CH_3COOH$$

In a particular reaction 3.077g of Salicylic acid and 5.50cm^3 of acetic anhydride reacted to form 3.281g of aspirin.

(a) Which is the limiting reactant

Take

(Density of acetic anhydride = 1.08g/cm^3)

Solution

Remember density $= \frac{\text{mass}}{\text{Volume}} \Rightarrow$ Mass = density x Volume

$= 1.08 \times 5.5$

$= 5.94g$

Formularmass of Salicylic acid ($C_7H_6O_3$) = 138g

" " " aceticanhydride = 102g

$(CH_3CO)_2O$

To get the limiting reactant,

Say 138g of Salicylicacid ⟶ 102g Acetic anhydride

3.077g of " ⟶ $\frac{3.077 \times 102}{138}$

= 2.27g of A.A

Since, it is 5.94g was used in the reaction according to the question instead of 2.27g of Acetic anhydride, the acetic anhydride is the excess reactant, where as, the Salicylic acid is the limiting reactant.

⑦ Calculate the mass percent of nitrogen in ammonium nitrate, hence, how many grams of nitrogen are in 35kg of ammonium nitrate (NH_4NO_3)

Solution

$NH_4NO_3 \Rightarrow 14 + 4 + 14 + 48 = 80g/mol$

① % Nitrogen in $NH_4NO_3 \Rightarrow \frac{[illegible]}{\text{molarmass } NH_4NO_3} \times \frac{100}{1}$

$\Rightarrow \frac{28}{80} \times 100 \Rightarrow 35\%$

If 1kg = 1000g of NH_4NO_3

35.8kg ⇒ 35800g of NH_4NO_3

So if 80g of NH_4NO_3 Contains 28g of N

35800g " " " $\left\{ \frac{28 \times 35800}{80} \right.$

⇒ 12,530g of N_2

⑧ Calculate the mass of C in 16.55g of Glucose.

Solution $C_6H_{12}O_6$ (glucose)

mass of C $= \frac{(12 \times 6)}{180} \times 16.55$

$C_6H_{12}O_6$ molecular mass (ie $(12 \times 6) + (12 \times 1) + (16 \times 6)$)

= 180g

mass of C $= \frac{12 \times 6}{180} \times 16.55 = 6.62g$

MASS PERCENTAGE FROM THE CHEMICAL FORMULAR.

Example: $C_6H_{12}O_6$

find (i) %C (ii) %H (iii) %O

Solution

(i) $\%C = \frac{12 \times 6}{180} \times 100 \Rightarrow 40\%$

(ii) $\%H = \frac{12 \times 1}{180} \times 100 \Rightarrow 6.67\%$

(iii) $\%O = \frac{16 \times 6}{180} \times \frac{100}{1} \Rightarrow 53.33\%$

CHYDONN'S QUOTES

A good way to test your understanding then learn more is by trying to explain it to another person (Classmates)

ACTUAL YIELD & THEORETICAL YIELD

When reactions are conducted in the lab. It is assumed that 100% of the limiting reactant becomes product and that ideal separation and purification method exist for isolating the product, and that we use perfect laboratory technique to collect all the product formed.

In other words, we assume that we obtain the theoretical yield.

However, in reality, the theoretical yield is never obtained. Rather, when experiments are conducted, the amount of products obtained is often less than the theoretical yield and it is referred to as the actual yield.

Theoretical yield of a chemical reaction is the amount of product expected based on the stoichiometrically equivalent molar ratio in the balanced equation.

The truth of the matter is, theoretical yield is never obtained as a result of a number of uncontrollable factors which may include side reactions, inadequacy in weighing and purity of the reactant.

The actual yield is the actual amount of products obtained under experimental conditions

$$\% \text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times \frac{100}{1}$$

$$\text{Yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}}$$

Note that Percentage yield must be < 100%

(ii) that actual yied < theoretical yield

(iii) Pyrolusite is an ore that contains large deposits of MnO_2.

Questions

① 100kg of Sand is made to react with Carbon to produce Silicon Carbide. If the molar mass of Silicon = 28. The mass of Silicon Carbide formed is 51.4kg. What is the percentage yield of Silicon Carbide?

Solution

$$SiO_2 + 3C \rightarrow SiC + 2CO$$

$$SiO_2 \rightarrow SiC$$

$$(28+32)g \rightarrow (28+12)g$$

$$\text{If } 60g \longrightarrow 40g$$

$$100g \longrightarrow \frac{40 \times 100}{60} \Rightarrow 66.67kg$$

Note: Actual yield ⇒ 51.4kg

$$\therefore \% \text{ yield} \Rightarrow \frac{51.4}{66.7} \times \frac{100\%}{1} = 77.09\%$$

② Aspirin (acetylsalicylic acid) is made by reacting Salicylic acid with acetic anhydride

$$C_7H_6O_3 + (CH_3CO)_2O \rightarrow C_9H_8O_4 + CH_3COOH$$

In one reaction, 3.077g of Salicylic acid and 5.50ml of acetic anhydride reacting to form 3.281g of aspirin.

ⓐ What is the percentage yield of the reaction? (density of acetic anhydride = 1.08g/cm³)

Solution

Mass = Volume × density

$= 5.50 cm^3 \times 1.08 g/cm^3$

$= 5.94g$ of acetic anhydride.

Formular mass of Salicylic acid = 138g

" " " aceticanhydride = 102g

To get the limiting reactant, Say

138g of Salicylicacid → 102g aceticanhydride

3.077g " " " → $\frac{3.077 \times 102}{138}$

$\Rightarrow 2.27g$ of acetic anhydride

Since, it is 5.94g was used in the reaction according to the question Instead of 2.27g of acetic anhydride, the aceticanhydride is the excess reactant whereas the Salicylic acid is the limiting reactant.

To get the Aspirin that will be produced theoretically,

138g of Salicylicylcacid → 180g of Aspirin

3.077g of Salicylicacid → $\frac{180 \times 3.077}{138}$

Theoretical yield is $\Rightarrow 4.01g$ of aspirin

$\therefore$ % yield = $\frac{\text{Actual yield}}{\text{Theoretical}} \times 100\%$

N.B Actual yield will be given to you in the question In the question Aspirin formed (actual yield)

$= 3.281g$ of aspirin

$\therefore \% \text{ yield} = \frac{3.281g}{4.01g} \times \frac{100}{1} \Rightarrow 81.82\%$

PRECISION, ACCURACY, ABSOLUTE ERROR, RELATIVE ERROR, PERCENTAGE ERROR.

PRECISION: This is determined by the average mean deviation (d) for the set of results

$$\text{Mean deviation} = \frac{\sum |(x - x_i)|}{n}$$

The Smaller the mean deviation the higher the precision

Accuracy: A measurement is said to be accurate if it is close to the actual Value. The accuracy measurement is given by the percentage error. The Smaller the percentage error, the higher the accuracy.

Absolute error: This is another way or method one Can determine the precision of the values gotten. This is used to Comment on precision, If actual Value is given

$$\text{Absolute error} = |(\text{Average Values} - \text{Actual value})|$$

Relative error: This is Simply the ratio of absolute error to actual Value.

Percentage error: This is the product of relative error and 100%.

That is $\% \text{ error} = \frac{\text{absolute error}}{\text{True/actual value}} \times \frac{100\%}{1}$

The Smaller the % error, the higher the accuracy.

① Supposed we have 5 repeated measurement for mass of a Substance.

0.38, 0.36, 0.40, 0.44, 0.42g

if the actual mass is 0.42g, find:

ⓐ Absolute error

ⓑ Relative error.

ⓒ Percentage error

ⓓ Average error

ⓔ Comment on its accuracy and precision.

Solution

Absolute error ⇒ |average Value − Actual Value|

$$\text{Average Value} = \frac{0.38 + 0.36 + 0.40 + 0.44 + 0.42}{5}$$

$$= 0.40g$$

Since the actual value is given as 0.42g

Absolute error = $|0.40 - 0.42| = 0.02g$

This is Canbe use to Comment on its precision, in the sense that the Smaller the mean deviation or absolute error, the higher its precision.

ⓑ Relative error $\Rightarrow \frac{\text{Absolute error}}{\text{Actual Value}} \times \frac{1}{1}$

$$\Rightarrow \frac{0.02}{0.42} \times \frac{1}{1}$$

$$\Rightarrow 0.0476//$$

ⓒ Percentage error / % relative error

$$\Rightarrow \frac{\text{Absolute error}}{\text{Actual Value}} \times \frac{100}{1}$$

$$\Rightarrow \frac{0.02}{0.42} \times \frac{100}{1} = 4.76\%$$

This means that it is (100 − 4.76)% accurate ie 95.24% accurate.

ⓓ Average error

⇒ Get the following

$$|0.42 - 0.38| + |0.42 - 0.36| + |0.42 - 0.40| + |0.42 - 0.44| + |0.42 - 0.42|.$$

$$\Rightarrow \frac{0.04 + 0.06 + 0.02 + 0.02 + 0}{5}$$

$$\Rightarrow 0.028g$$

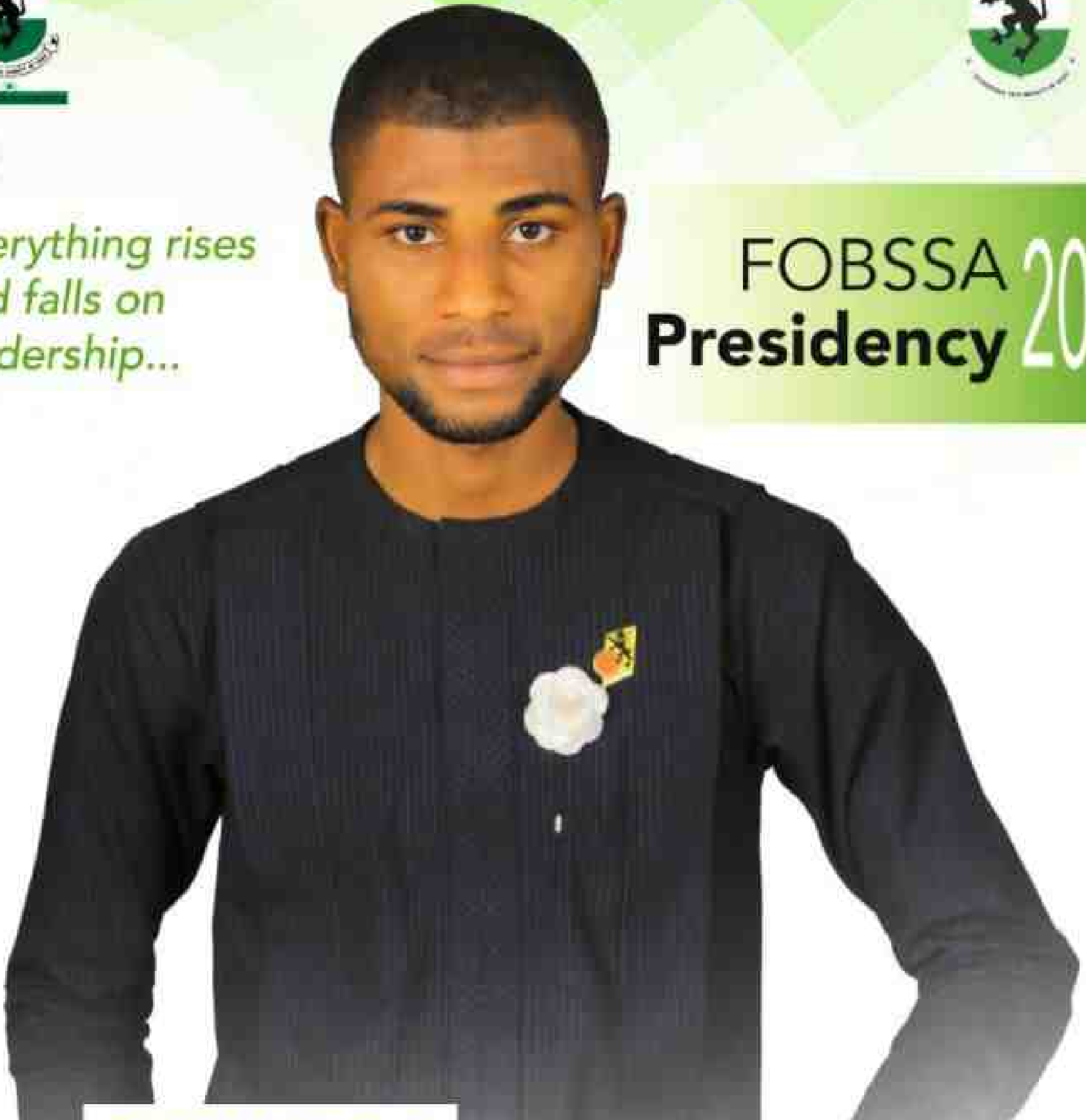
NOTE

The Concentration of a solute is a measure of its relative amount in a given Solution. A Solution is made up of Solute dissolved in Solvent. Most often the Solvent is of larger amount Compare to the Solute. Concentration of a Solution is usually expressed as the amount of Solute dissolved in 1dm³ of Solution.

Just like density and temperature, Concentration of a given solution is independent on the Volume. Example: 50 liter tank of a given Solution has the Same Concentration as 50ml beaker of the Solution.

But when a Concentrated Solution (high molarity) is diluted by adding Solvent (eg water) to it; The Solution will Increase, while the number of moles of Solute remains the Same. Hence it will Contain few Solute particles and will have a lower Concentration than the initial. That takes us to Solution Stoichiometry.

Everything rises and falls on leadership...



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FACULTY OF BIOLOGICAL SCIENCE STUDENTS ASSOCIATION **(FOBSSA)**

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SOLUTION STOICHIOMETRY

Solution terms one needs to get accustomed to in Solution Stoichiometry are

① Molarity $\Rightarrow \frac{\text{Number of moles of Solute}}{1dm^3 \text{ of Solution}}$
(mol/dm³)

OR

$\Rightarrow \frac{\text{No of millimoles of Solute}}{1cm^3 \text{ of Solution}}$

OR

$\Rightarrow \frac{\text{mass of Solute}}{\text{Molar mass}} \div \left[\frac{\text{Volume}(cm^3)}{1000}\right]$

$\Rightarrow \frac{\text{Mass of Solute}}{\text{molar mass}} \times \left(\frac{1000}{\text{Volume}(cm^3)}\right)$

(ii) Gram Concentration $\Rightarrow$ Molarity × Molar mass
(g/dm³) (mol/dm³) (g/mol)

(iii) Molality $\Rightarrow$ This is not Same with Molarity
(mol/kg)
$\Rightarrow \frac{\text{no of moles of Solute}}{1000g \text{ (1kg) of Solvent}}$

(iv) Normality $\Rightarrow \frac{\text{mass of Solute}}{\text{gram equivalent weight}}$

gram equivalent weight = $\frac{\text{molecular weight}}{\text{Charge gain/lost}}$

STANDARD SOLUTION

Some Standard Solutions are prepared by Serial dilution. Some Standard Solution are prepared from Solid material or Substances.

Primary Standard: This is a Substance that is weighed directly to prepare Standard Solution.

A primary Standard Should be Stable, Should have relatively high molar mass to minimise error due to weighing.

An example of a Primary Standard is Anhydrous Sodium Carbonate (Na_2CO_3)
Potassium Permanganate ($KMnO_4$)
Potassium Iodate

QUESTIONS

How may you prepare 0.25M Na_2CO_3 in 250cm³?

Solution

Using,
Molarity $\Rightarrow \frac{\text{no of mole of Solute}}{1dm^3 \text{ or Volume}}$

Molarity = 0.25M
Volume = 250 cm³

$0.25M = \frac{\text{no of moles}}{\left(\frac{250}{1000}\right)}$

no of moles $= \frac{250 \times 0.25}{1000} = \frac{\text{mass}}{\text{Molar mass}}$

$0.25 \times 0.25 = \frac{\text{mass}}{106 g/mole}$

= 6.625g of Na_2CO_3

So weigh out 6.625g of Na_2CO_3 in a weighing balance. Transfer the Content into 250cm³ Volumetric flask, dissolve it using distilled water, shake Vigorously, then make up to $250cm^3$ mark with distilled water.

N°.B 0.05M of Na_2CO_3 is more dilute than 0.25M Na_2CO_3

② How may you produce $250cm^3$ 0.05M Na_2CO_3 Solution from a 0.25M solution.

Solution

Use dilution formular

$$C_1V_1 = C_2V_2$$

$$250 \times 0.05 = 0.25 \times V_2$$

$$V_2 = \frac{250 \times 0.05}{0.25}$$

$$\Rightarrow 50cm^3$$

Add $50cm^3$ of 0.25M Solution. ie pipette out $50cm^3$ of 0.25M solution then add water to make it up to $250cm^3$ mark, what you will get will be 0.05M ($250cm^3$ of Na_2CO_3)

③ What Volume of distilled water must be added to $50cm^3$ of 0.25M Na_2CO_3 to make a 0.05M Solution?

Solution

Use $C_1V_1 = C_2V_2$

$C_1 = 50$, $V_1 = 0.25$; $C_2 = 0.05$, $V_2 = ?$

$$V_2 = \frac{50 \times 0.25}{0.05}$$

$$V_2 = 250cm^3$$

Volume of water to be added

$$= 250 - 50 \quad (\text{ie } V_2 - V_1)$$

$$= 200cm^3$$

④ How many moles of potassium iodide are in 84ml of 0.56 Molar.

Solution

$$\text{no of moles} = \text{Concentration} \times \frac{\text{Volume}}{dm^3}$$

$$= 0.56 \times \frac{84}{1000}$$

$$= 0.047 \text{ moles}$$

⑤ What Volume of distilled water must be added to $25cm^3$ of 10molar HCl to make a (1Molar Solution)

Solution

Use dilution formular

$$C_1V_1 = C_2V_2$$

$$V_2 = \frac{C_1 \times V_1}{C_2}$$

$$= \frac{25 \times 10}{1}$$

$$= 250cm^3$$

∴ Volume of water to be added

$$\Rightarrow 250 - 25 \quad (V_2 - V_1)$$

$$= 225cm^3$$

⑥ How many litres of 0.10M HCL will be needed to react with an antacid tablet Containing 0.10g of Magnesium hydroxide.

Balanced equation

$$Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + H_2O$$

Solution

Molar mass of $Mg(OH)_2$ ⇒

$24 + (16\times2) + (1\times2)$

$= 58g/mol$

No of moles of $Mg(OH)_2 = \frac{\text{Mass (g)}}{\text{molarmass (g/mol)}}$

$= \frac{0.10}{58}$

$= 0.00172$ mole

From the equation above $Mg(OH)_2 : HCl$

1 : 2

no of moles of HCl ⇒ 2×0.00172

⇒ 0.0034 mole

Volume in litre = $\frac{\text{mole}}{\text{molarity}}$

⇒ $\frac{0.00345}{0.10}$

⇒ 3.25×10^{-2} Litre

⑦ How much Sodium oxalate ($Na_2C_2O_4$) would a 500cm³ $\frac{1}{50}$ M $Na_2C_2O_4$ Contain?

Solution

Molarity = $\frac{\text{mass}}{\text{molarmass}} \times \frac{1000}{\text{Volume}}$

$\frac{1}{50} = \frac{x}{\text{Molarmass}} \times \frac{1000}{500}$

molarmass of $Na_2C_2O_4$ ⇒

$(23\times2) + (12\times2) + (16\times4) = 134g/mol$

$\frac{1}{50} = \frac{x}{134} \times \frac{1000}{500}$

$x = \frac{134\times5\cancel{0}\cancel{0}}{500\cancel{0}\cancel{0}}$

$x = \frac{670}{50} = 1.34g$

Secondary Standard: It is a Solution that has to be Standardized before it Can be referred to as Standard.

Examples:

Commercial HCl is about 11.96M

Commercial H_2SO_4 is about 18M

Questions

How may you prepare 500cm³ of 2M HCl from Commercial Solution with the following information? (% weight by weight ⇒ 37% Molecular weight = 36.5g/mol; density = 1.18g/cm³).

Solution

Molarity ⇒ $\frac{\text{\% weight by weight} \times \text{density (g/cm}^3) \times 1000}{\text{Molecular weight}}$

⇒ $\frac{37\% \times 1.18 \times 1000}{36.5}$

⇒ $\frac{37}{1\cancel{0}\cancel{0}} \times \frac{1.18}{1} \times \frac{10\cancel{0}\cancel{0}}{36.5}$

⇒ 11.96M, Now using $C_1V_1 = C_2V_2$

$V_2 = \frac{C_1 \times V_1}{C_2}$ ⇒ $\frac{500\times2}{11.96} = 83.6cm^3$

⇒ Pipette 83.6cm³ into 500cm³ Volumetric flask from 11.96M Solution and make it up to the 500cm³ with distilled water.

ONYEUKA BRIGHT CHIMEZIE

Introduction to Atomic Structure

- Introduction to historical development of atomic theory'
- Gravimetric law and their limitations
- Cathode raytube experiment.
- Charge to mass ratio of Cathode ray (e/m)
- Milikan experiment (Charge of an electron)
- Discovery of Radioactivity with brief description
- Rutherford's experiment
- Discovery of neutrons
- Substances incapable of decomposition by any mean element.

NOTE THE FOLLOWING

Democritus (460 – 370 BC)

- Matter Consisting of Simple indivisible particles Called atoms is made to float around in an empty Vessel.

- Aristotle (384 – 322 BC)

Matter is Continous and therefore infinitely divisible. He also states that all Substances were built from 4 elements ie earth, air, fire and water.

- Robert boyle (1661)

- Element is a Substance incapable of deComposing by any means with which we are at present acquanted.

- John Dalton (1766 – 1844)

① Each element is Composed of tiny indivisible particles Called atom

② Atoms can neither be Created nor destroyed

③ Atoms of different element Combine in a Simple whole number ratio by mass to form Compounds.

④ Atoms of Same element are alike in all respect having the Same weight but differed from atoms of other elements ie they have Identical properties.

⑤ All chemical Changes or reaction Involves the Separation and Combination of atoms.

- Atoms of different elements have different properties

MODIFICATIONS OF THE ATOMIC THEORY (INVALIDATIONS)

(ie Invalidation of John Dalton's Atomic theory stated above)

- The discovery of Subatomic particles: Proton, neutron, electron invalidates the first postulate ie number ①
- The discovery of radioactivity invalidates the second postulate ie number ②
- The discovery of Isotopy Invalidates the 4th postulate.
- The discovery of non-stoichiometric Compounds eg $FeSO_4$ and large macro molecules like Protein, Vitamin, Carbohydrate Invalidates the 3rd Postulates,

GRAVIMETRIC LAWS AND THEIR LIMITATIONS

These laws are also referred to as "LAWS OF CHEMICAL COMBINATION BY MASS".

These are laws that involve measurement by mass and were established by means of Scientific Investigation in support of the atomic hypothesis.

There are 4 gravimetric laws namely;

(i) LAW OF CONSERVATION OF MASS
(ii) LAW OF DEFINITE PROPORTIONS (CONSTANT COMPOSITION)
(iii) LAW OF MULTIPLE PROPORTION
(iv) LAW OF RECIPROCAL PROPORTION

Lets take it one by one.

① LAW OF CONSERVATION OF MASS

This law states "Matter is neither Created nor destroyed during a Chemical reaction but Changes from one form to another." This laws was proposed by "ANTHOINE LAVOISIER" in "1774" and Simply Implies that in the Cause of a chemical reaction, the sum of mass of the reactant before a chemical Change, equals the sum of mass of the Product after a Change. This law is also referred to as "Law of Indestructibility of matter".

LIMITATION OF CONSERVATION OF MASS.

This law is Invalidated by interConvertibility of mass and energy as given by the Einstein equation where;

$$E = MC^2$$

E = energy
m = mass ; C = speed of light.

This is usually observed during nuclear reaction especially nuclear fission. The total mass of all the fragments and neutron released, differs from the total mass of the original atom of the element and the masses of the bombarding neutrons. The difference in mass arises due to the Conversion of mass to energy which is liberated during fission.

② LAW OF DEFINITE PROPORTION (CONSTANT COMPOSITION)

This law states that "In any pure Chemical Composition regardless of the method of Preparation, the Constituent elements are always Combined in the Same proportion by mass." This law was put forward by "Joseph Proust" in "1799".

From this law, we know that two sample of CuO obtained as shown in the equation below Contains the same mass of Cu and O Combined in the Same ratio.

Decomposition of Copper nitrate ($Cu(NO_3)_2$)

$$2Cu(NO_3)_2 \longrightarrow CuO + O_2 + 2NO_2$$

Decomposition of Copper Carbonate

$$CuCO_{3(s)} \longrightarrow CuO_{(s)} + 2CO_{2(g)}$$

Examples

$Zn = 10.00g$; Mass of residue $= 8.11g$

$Zn = 13.20g$; mass of residue $= 10.70g$

What is the ~~rat~~ ratio of Zn residue to Oxygen

Solution

Mass of Compound (g)	Mass of Zn residue (g)	Mass of Oxygen residue (g)
A 10.00	8.11	1.80
B 13.20	10.70	2.50

$$\% \text{ Zinc in A} = \frac{\text{Mass of Zn residue}}{\text{Mass of Zn.O Compound}} = \frac{8.11}{10.00} \times \frac{100}{1}$$

$$= 81.1\%$$

$$\% \text{ Zinc in B} = \frac{\text{Mass of Zinc residue}}{\text{Mass of ZnO Compound}} = \frac{10.7}{13.20} \times \frac{100}{1}$$

% of Oxygen in A and B $= 81.1\%$

$= 100 - 81.1 \Rightarrow 18.9\%$

Ratio of Zinc residue to Oxygen in A and B

$\Rightarrow 81.1 : 18.9$

LIMITATION OF LAW OF DEFINITE PROPORTION. (CONSTANT COMPOSITION)

This law is invalidated by the phenomenon of Isotopy. Elements that exhibit Isotopy have atoms with different atomic masses eg $^{35}_{17}Cl$ & $^{37}_{17}Cl$ - Hence the composition by weight of any Compound that Contains Isotopic Compounds will vary depending on the particular Isotope present in the Compound.

Also the discovery of Isotopes invalidated the claim that all atoms of a particular element have identical Properties. For Instance Isotopes of Carbon have different properties eg $^{12}_{6}C$ is not used in radioactive dating while $^{14}_{6}C$ Is applicable in radio active dating due to its radio active nature.

(3) LAW OF MULTIPLE PROPORTION

It states that "If 2 elements A and B Combine to form more than one chemical Compound, then the various masses A which Combines Separately with a fixed element B are in a Simple whole number ratios.

Examples

① Consider Compound lead and Sulphur (PbS & PbS_2). The mass of Sulphur which Combine with 206g of lead in the 2 Compounds are 32g and 64g respectively and are in the ratio of 1:2

	PbS	PbS_2
Amt of S	32g	64g
Ratio with Molar masses	$\left(\frac{32}{32}\right)$	$(64 \div 32)$
	1 :	2

ie they are in Simple whole number ratios.

LIMITATIONS OF LAW OF MULTIPLE PROPORTIONS

This law is invalidated in Carbon Compounds and non Stoichiometric Compounds where the ratio of the Combination is not Simple, with lattice defects

e.g. For Non Stoichiometric Compounds example is

FeS Iron Sulphide

$FeS_{1.4}$ (Iron Sulphide with lattice defects)

1.14

1 ; 1.14

ie ratio of Combinations of non stoichiometric Compounds with lattice defects are not in Simple ratio.

Also Carbon Compds eg Hexane ie C_6H_{14} and decane ($C_{10}H_{22}$) are not also in Simple ratio.

④ LAW OF RECIPROCAL PROPORTIONS

It is a law Stated by "Berzelus in 1812"

It States that the masses of elements A,B,C which Combine with a fixed mass of another element "D" Separately are the Same as or Simple multiples of masses in which A,B,C themselves Combine with one another.

Consider a Compound formed between H, C, Na with a fixed mass of O.

Hydrogen + Oxygen → H_2O → 2g of H_2 + 16g of O

Carbon + Oxygen → CO → 12g of Carbon + 16g of O

Na (Sodium) + Oxygen ⇒ Na_2O → 46g of Sodium + 16g of O

You Can See that mass of Oxygen that Combines with H, C, Na is only 12g ie it is Same to H, C, Na.

LIMITATIONS

Phenomenon of Isotopy also.

N.B Rutherford discovered atom
Thompson discovered electron
Chadwick discovered Neutron

QUESTIONS

① Two different Samples A & B of Zinc Oxide ZnO were obtained from different Sources when heated in a Stream of H_2, they were reduced to yield the following results.

ZnO	mass of Sample	Mass of residue
Sample A	10.00	8.11
" B	13.20	10.70

which gravimetric law will this obey?

Solution

Using percentages of Zinc ;

Percentage Composition of Zinc in A

$$\Rightarrow \frac{8.11}{10.00} \times \frac{100}{1} = 81.1\%$$

Percentage Composition of Zinc in B.

$$\frac{10.70}{13.20} \times \frac{100}{1} \Rightarrow 81.1\%$$

Zinc A : Zinc B

$$\frac{81.1}{81.1} : \frac{81.1}{81.1}$$

1 : 1

(The law of Constant Composition) = Answer

② In $AgNO_3 + HCl \rightarrow AgCl_2 + HNO_3$ which laws are being disobeyed.

ANSWER

Since the equation is not balance, law of Constant Composition is disobeyed.

ONYEUKA BRIGHT CHIMEZIE

Also

Law of Conservation of matter is also disobeyed when an equation is unbalanced.

③ Two Samples of Copper oxide were reduced by passing a stream of hydrogen gas. Given the following data below, which law is it in accordance with and what is the formular of the oxide.

	Oxide	Cu	O
A	5.20g	4.62	0.58
B	3.10g	2.48	0.68

Solution

Ratio of Cu in A = $4.62/0.58 = 8$

Ratio of Cu in B = $2.48/0.68 = 4$

$8 : 4 \Rightarrow 2 : 1$

ie Cu_2O and CuO

It is in accordance with law of multiple proportion.

④ In an experiment to reduce two different oxides of Copper, it was discovered that 1g of O_2 Combined with 5.20g of Cu and 2.60g of Cu in the other oxides. Which gravimetric law does it illustrate and write its formula of the oxide.

Solution

1st Copper	2nd Copper
5.20g	2.60g

Ratio $\Rightarrow \frac{5.20}{2.60} \Rightarrow \frac{2}{1}$

Since they form oxide with oxygen we will have Cu_2O and CuO from 2 : 1 (ratio of Cu).

It obeys / illustrate law of multiple proportions.

⑤ If 35.5g of Iron (Fe) was made to react with Chlorine, The mass of the product obtained Could be either 80.51g or 103.01g depending on the amount of Chlorine Present. Which gravimetric law would this be in accordance with and what is the formular of its Chloride?

Solution

	mass of Product	Mass of Chlorine g
A	80.51	80.51 − 35.5
B	103.01	103.01 − 35.5

Mass of Chlorine A	Mass of Chlorine B
45.01	67.50

Ratio ; $\frac{67.50}{45.01} = \frac{1.5}{1} \times \frac{2}{2} \Rightarrow \frac{3}{2}$

Ratio ⇒ 3 : 2 ie $FeCl_3$ & $FeCl_2$

It obeys law of multiple proportions.

A Chemist Investigating a Compound decompose it into Simpler Substances, finds the mass of each element, Converts their mass into numbers of mole and arithmetically Converts their mole into whole number Integer. This procedure yields the empirical formular which is defined as the Simpleformula Showing the Simplest whole number ratio of moles of each element in the Compound.

By using a mass Spectrometer, a Chemist Can obtain the molecular mass, once the molecular mass is known, then the molecular formular Can be determined. Because molecular formular is a whole number multiple of the empirical formular. If the Compound is an organic Compound, the Compound Can be reacted with pure Oxygen (Combustion analysis). In order to determine the amount of Carbon and hydrogen. The Combustion reaction leads to the formation of CO_2 and H_2O. Once the masses of the products are obtained, the mass of Carbon and hydrogen Can be Calculated.

Molecular formular is thereby defined as the formular Showing the "Actual Number" of moles of "each element" in "One mole of the Compound".

QUESTIONS

① An analysis of an unknown Compound shows that the Compound Contains 0.21 mole of Zn, 0.14 mole of P, 0.56 mole of O. Determine the empirical formular of that Compound.

Solution

	Zn	P	O
No of mole	0.21	0.14	0.56
divide by lowest	$\frac{0.21}{0.14}$	$\frac{0.14}{0.14}$	$\frac{0.56}{0.14}$
	1.5	1	4
multiply by 2	3	2	8

Empirical formular ⇒ $Zn_3P_2O_8$

② Elemental analysis of a Sample of an ionic Compound showed 2.82g of Na, 4.35g of Cl, 7.83g of O. What is the empirical formular. If the empirical formular is Same as the molecular formular, Name the Compound.

Solution

	Na	Cl	O
masses	2.82	4.35	7.83
Divide by Molecular Mass	$\frac{2.83}{23}$	$\frac{4.35}{35.5}$	$\frac{7.83}{16}$
to get mole ⇒	0.123	0.123	0.489
Divide by the lowest	$\frac{0.123}{0.123}$	$\frac{0.123}{0.123}$	$\frac{0.489}{0.489}$
	1	1	4

Empirical formular $\Rightarrow NaClO_4$ = Molecular Formular
Name $\Rightarrow$ Sodium hypochlorate

③ An unknown Metal "M" reacts with Sulphur to form a Compound with the formular M_2S_3. If 3.12g of M reacts with 2.88g of Sulphur. What are the names of "M" and "M_2S_3" respectively?

Solution

M	S
$\frac{3.12}{x}$	$\frac{2.88}{32}$
$\frac{3.12}{x}$	0.09

(let the atomic mass of M be x)
atomic mass of Sulphur = 32g/mol

Since Sulphur S_3 has 3

$$\Rightarrow \frac{0.09}{\left(\frac{3.12}{x}\right)} = 3 \quad \Rightarrow 0.03 = \frac{3.12}{x}$$

$x = 104g/mol$

$M_2 \Rightarrow 104g/mol = x$

$2M = 104g/mol$

$\Rightarrow M = \frac{104}{2} = 52g/mol$

From Periodic table it is Chromium that has atomic mass of 52g/mol.

Name of $Cr_2S_3 \Rightarrow$ Chromium (III) Sulphide

④ What mass of $MgSO_4 \cdot 7H_2O$ must be dissolved in 300g of water such that 1g of the Solution Contains 0.075g of Mg^{2+}?

Solution

① 1g of the Solution is to Contain 0.075g of Mg^{2+}

Molar mass of $MgSO_4 \cdot 7H_2O \Rightarrow 24 + 32 + 64 + (7 \times 18)$

$= 246 g/mol$

From $MgSO_4 \cdot 7H_2O \Longrightarrow$

24g of Mg^{2+} Contained in 246 g/mol of Solute

0.075g of Mg^{2+} Contained $\frac{246 \times 0.075}{24}$

$= 0.768g$ of Solute

Mass of water $= 1 - 0.768g = 0.232g$

[Since 1g of Solution = Solute (mass) + mass of Solvent]

(dissolves)
$\Rightarrow$ 0.232g of H_2O Contains 0.768g of Solute
300g of H_2O dissolves X

$$\left[\frac{300 \times 0.768}{0.232}\right]$$

$= 993.1g$ of Solute

If you are asked to get mass of water of Crystallization

$$\Rightarrow \frac{(7 \times 18)}{246} \times 993 \Rightarrow 508.6g$$

Mass of Sulphur in the Solute

$$\frac{32.07}{246} \times 993 \Rightarrow 129.5g \text{ of } S$$

MY NAME IS CHYDONNS (A.K.A IgweOrmutaokuzie)

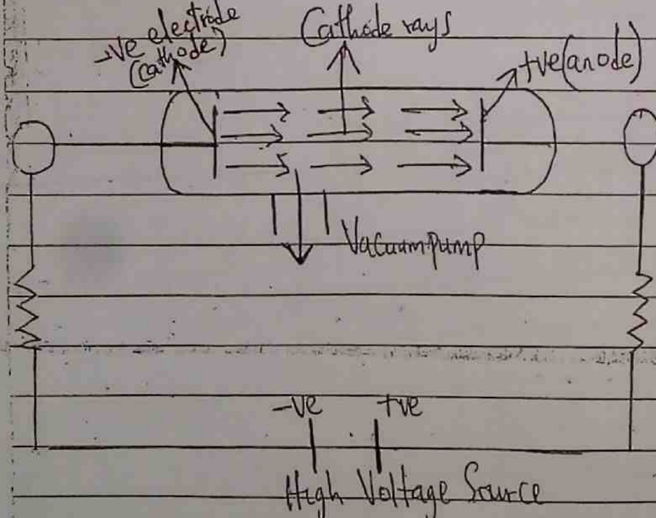
The Secret to Academic Success in U.N.N are

(i) Read lecturers note Very well and School Past questions, not, pass question paper formed by people because they ask only what they teach.

(ii) When you Come back from lecture, Please try even if it is for 30mins to go through, what you were thought for that day. Many student fails because they do Cummulative reading, ie they wait till exam time before they read their note. Please don't do like them. Thank You!!

INTRODUCTION TO CATHODE RAYS TUBE EXPERIMENT

CATHODE/DISCHARGE TUBE



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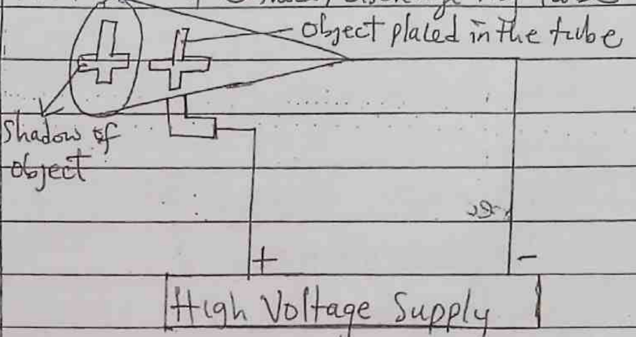
CATHODE RAY TUBE EXPERIMENT

A sketch of Cathode/dischage ray tube



This experiment was Carried out by Sir. William Crookes.

The dischage tube is made up of:

(i) A glass tube with 2 metal plates Sealed at both end.

(ii) Electrodes (Cathode and anode)

(iii) High Voltage Supply (10,000 Volts) under low pressure below 10mmHg.

What happened in the Cathode ray tube experiment?

The electrodes are Connected to a high Voltage (10,000 volts) and air is partially evaluated via the Vacuum pump. Ordinarily air is a good Insulator, but using dischage tube under low pressure below 10mmHg, a potential difference of about 10000 volts, the Insulating properties of air breakdown and air becomes a good Conductor. As the gas in the tube is being evacuated at a high Voltage across the electrodes under low pressure, a glowing of the residual gas in the tube will be observed.

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When the Pressure is reduced to about 0·01 mmHg, the gas will no longer emit light. The glass of the tube fluoresce as a result of the rays from the Cathode bombarding the glass. Since the rays are emanating from the Cathode, they are refered to as Cathode rays. J.J. Thomson (1897) showed that the negatively charged Cathode rays travelled in Straight lines from the Cathode to the anode and that in magnetic and electric fields, the direction in which they were deflected Correspond to that for Steam of negatively charged Particles.

Regardless of the gas in the chamber, they are Constant, So J. J. Thomson named them electrons. Other Scientists Involved are Julius Plucker, Eugene Goldstein, John Hittorf.

PROPERTIES OF CATHODE RAYS

(i) They travel in Straight lines emanating from the Cathode to anode, because of this, they are also Called electrons.

(ii) They Cast shadow of an object placed in its path but do not penetrate on the objects

(iii) They are deflected towards the positive electrode of an external electric field.

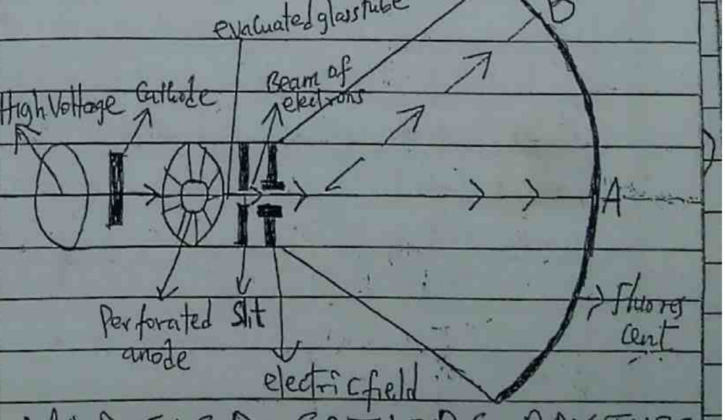
(iv) They are identical in nature and with the ratio of Charge to mass regardless of the nature, of the residual gas in the tube or the metals used as the electrodes

(v) They are deflected in the direction reflecting the pesence of the negatively Charged particles when placed in an external magnetic field.

(vi) They are generated in a discharge tube when a low pressure and a high Voltage is applied to the air/gas in the tube.

MEASUREMENT OF CHARGE TO MASS OF ELECTRON

J.J. Thomson used modified Cathode ray tube to determine the Charge to mass ratio of electron.



MODIFIED CATHODE RAY TUBE

Electrons produce a bright luminous Spot at A on the fluorescent Screen - Application of magnetic field leads to the deflection of the electron in a Circular path while the Spot is Shifted to B. A magnetic field of known strength is then applied to return the Spot to it's original position. The velocity of the electron can be Calculated from the Strength of the magnetic field. The slit allows the electron to pass through the discharge tube in a straight line - Zinc Sulfide (ZnS) in the fluorescent Substance used To Cause bright Spots by electrons for easy determination.

Charge to mass ratio Calculated by J.J. thomson, $e/m = 1.76 \times 10^8 C/g$

or 1.76×10^8 Coulomb/kg

Milikan used oil drop to determine the Charge & mass of an electron in 1909

Charge of electron $\Rightarrow 1.6 \times 10^{-19} C$

mass of electron $\Rightarrow 9.1 \times 10^{-28} g$

The Charge to mass ratio was determined experimentally to be $1.76 \times 10^8 C/g$, that is, it is Same with the Calculated Value.

Calculated Value was gotten by

$$\frac{e}{m} = \frac{1.6 \times 10^{-19} \text{ Coulombs}}{9.1 \times 10^{-28} \text{ (gram)}} \Rightarrow 1.76 \times 10^8 \boxed{C/g}$$

Note this formular for Calculation

$$\boxed{\frac{e}{m} = \frac{V}{Br}} \quad \text{or} \quad \boxed{\frac{e}{m} = \frac{E}{B^2 r}}$$

$$\frac{e}{m} = 1.76 \times 10^8 C/g \text{ or } 1.78 \times 10^8 C/g$$

[depending on the approximations]

$$= 1.76 \times 10^{11} C/kg \text{ or } 1.78 \times 10^{11} C/kg$$

where B = magnetic field strength

$$\overset{\text{velocity of electrons}}{V} = \frac{E \rightarrow \text{electric field strength}}{B \rightarrow \text{magnetic field strength}}$$

QUESTION:

If the $e/m = 1.76 \times 10^8 C/g$, the magnetic field strength is 2.5 teslas and the radius of Circular path described by the electron in the magnetic field is 0.05m. Calculate.

(i) the electric field strength

(ii) the velocity of the electron.

Solution

$$\frac{e}{m} = 1.76 \times 10^8 C/g$$

$B = 0.25 \times 10 \Rightarrow 2.5T$, $r = 0.05m$

Since e/m is in C/g, r should be in cm

$r = 5cm$

$$\frac{e}{m} = \frac{E}{B^2 r}$$

$$\Rightarrow E = B^2$$

$$E = \frac{2.5^2 \times 5 \times 1.76 \times 10^8}{1}$$

$$= 5.5 \times 10^9 \text{ volts}$$

(ii) Velocity of the electron

$$V = \frac{E}{B} = \frac{5.5 \times 10^9}{2.5}$$

$$\Rightarrow 2.2 \times 10^9 \text{ cm/s}$$

POSITIVE RAYS (PROTONS)

In 1866, Goldstein discovered positive rays. It is Important to recall that atoms are electrically neutral and Since electrons are negatively Charged it means that it must have a positive charge to Counter balance it. Goldstein observed that with a perforated Cathode, rays pass through the holes in the Cathode as luminals streams.

These rays are also deflected by electric and magnetic fields and they move towards the Cathode they Consist of streams of positive Charge particles and are therefore Called positive rays.

The ratio of the charge on a positive particle to its mass was found to be much Smaller than that for the electrons and they vary with the nature of the gas in the tube. The charge and mass of these particles Varied, but the highest of them has the Same mass as hydrogen atoms, and a charge equal in magnitude but of opposite Sign to that of the electrons.

This particle is a hydrogen atom that has lost an electron and from the Values of e and e/m, for the positive particles when hydrogen is used, the Value of m Can be Calculated; the mass is 1.0073 on the atomic weight Scale. This particle is Called the proton and is one of the fundamental Component of atom.

Mass of Proton = 1840 × mass of $\bar{e}$

$= 1840 \times 9.1 \times 10^{-28}$ g

$= 1.67 \times 10^{-24}$ g

Rutherford in 1919 proved that atoms Contain protons by bombarding high Velocity α particles from radium as projectiles on atoms Such as nitrogen and aluminium and found that protons were ejected as a result of these Collision.

THE NEUTRON (n^0)

In 1932, the English physicist Chadwick, discovered a third type of fundamental components of atom. He observed that when atoms of some elements are bombarded with high velocity alpha particles, uncharged particles (neutrons) are emitted. He called these neutral particles neutrons and they have a mass of 1.0087 on the atomic weight scale.

N.B the elements bombarded with α particle = Polonium foil

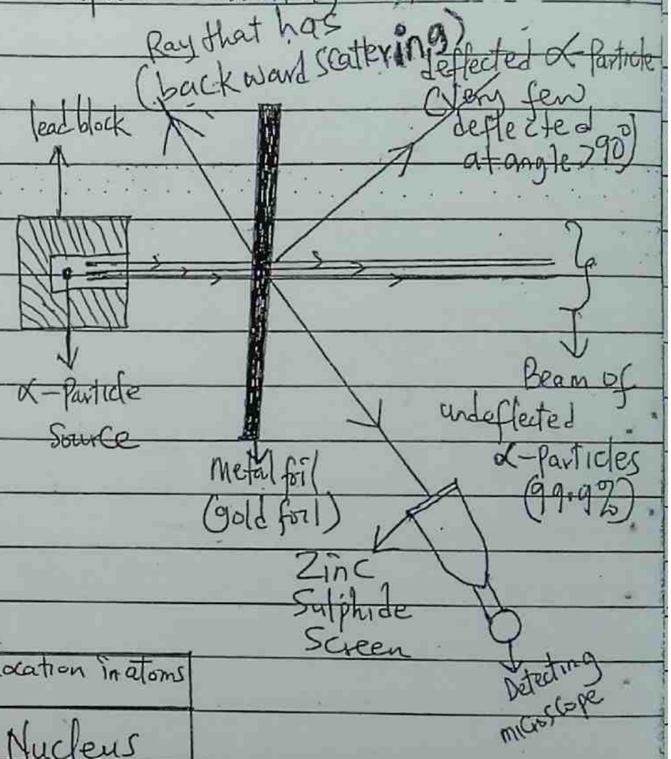
Summary of the 3 Components of Atom.

Name	Relative charge	Absolute Charge	Relative mass (A.M.U)	Absolute mass (g)	Location in atoms
Proton	H^{+1} (+1)	$+1.602 \times 10^{-19}$	1.00727	1.67252×10^{-24}	Nucleus
Neutron	0	0	1.00866	1.67493×10^{-24}	Nucleus
Electron	-1	-1.602×10^{-19}	0.000549	9.10939×10^{-28}	Orbitals

RUTHERFORD'S ATOMIC MODEL

The knowledge that atom is made up of electrons and a positive ion paved way for Rutherford to perform an experiment in a view to elucidating as to how and where the electrons and the positive ions are located in an atom. In 1909, Rutherland and Marden performed an experiment, Alpha Scattering Experiment. (Planetary model)

Alpha Scattering experiment

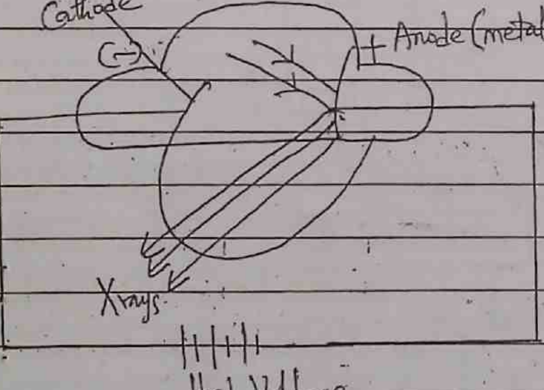


They directed a stream of very energetic α-particles from a radioactive source against a thin gold foil. They noticed that most of the α particles passed straight through the gold foil producing a flash on the screen behind it, indicating that gold atoms had a structure with plenty empty space. Also tiny flashes were observed on the other portions of the gold foil. This revealed that gold atoms deflected or scattered α-particle through large angles in which some even bounced back to the source.

Hence, the proposed model of atoms by Rutherford Can be stated as:

1. The electrons were moving in a close circular path (orbit) around the nucleus just like the planets around the sun.
2. The positive charge of the atom is located on the nucleus.
3. Atom has a tiny dense central core called the nucleus which contains the mass of the atom, leaving the rest of the atom almost empty.

"DISCOVERY OF XRAYS



Properties of Xrays

1. They have very high energy ie they penetrate materials which are impervious to light
2. They are undeflected by magnetic field ie they passed through because of high energy they have.
3. They have shorter wavelength ie is similar to light rays.

RADIOACTIVITY & ITS APPLICATIONS

Antoine Henri Becquerel discovered radioactivity in 1896. He observed that Uranium gave out radiation spontaneously.

Pierre and Marie Curie, in 1898 observed that Radium gave out the same radiation as uranium spontaneously.

Antoine Henri Becquerel was the supervisor, while pierre and Marie Curie were students

NUCLEAR REACTIONS

1. Radioactive decay (Natural radioactivity): This is a process in which a nucleus (ie unstable nucleus) spontaneously disintegrates giving off radiation. These emitted radiation Consist of one or more of the following:
   - (i) electrons
   - (ii) nuclear particles such as neutrons
   - (iii) Smaller nuclei such as helium nuclei.
   - (iv) Electromagnetic radiation such as gamma rays, x rays etc.

Example

$$\overset{\text{Radium}}{{}^{226}_{88}Ra} \longrightarrow \underset{\text{Radon}}{{}^{222}_{86}Rn} + \overset{\alpha\ \text{particle}}{{}^{4}_{2}He}$$

② Nuclei Bombardment Reaction (Artificial Radioactivity)

This is also known as Transmutation. It is a process in which a nucleus is bombarded by another nuclei or by a nuclear particle. The particle hitting the nucleus must have sufficient energy. If it has enough kinetic energy, it has a way of entering the nucleus causing instability and thereby leads to disintegration.

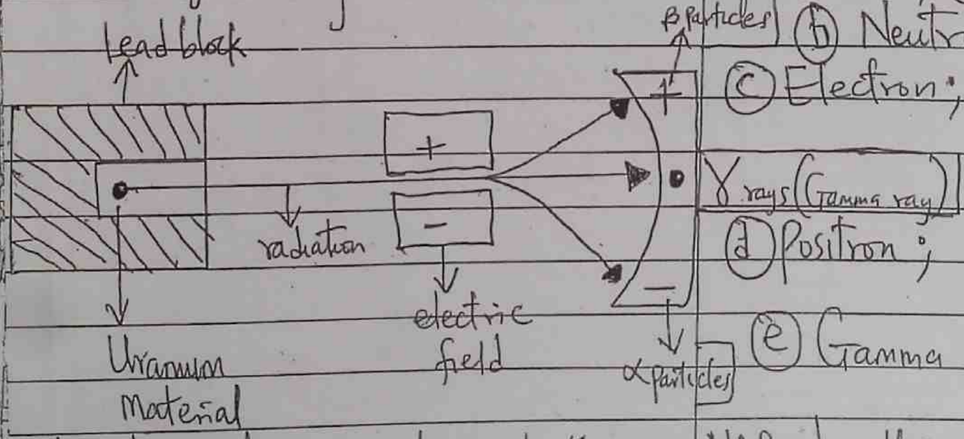
Causes of Radioactivity

Radioactivity Involves the disintegration of unstable atomic nuclei. Nuclear Instability may result from

(i) Excessive mutual repulsion of Protons

(ii) Proton-neutron Imbalance

(iii) Bombardment of the nucleus with high energetic substances.



N.B In nuclear equations both Charge and masses are Conserved.

THIS IS (1911)

ERNEST RUTHERFORD'S EXPERIMENT

NUCLEAR EQUATIONS

Nuclear equations unlike Chemical equations are represented using "Nuclei Symbols"

Example 1: A radioactive decay of $^{238}_{92}U$ is an α emission (alpha emission) decays to form ^{234}Th. Express as an equation.

Answer

$$^{238}_{92}U \longrightarrow {}^{234}_{90}Th + \underset{\text{alpha particle (Helium atom)}}{{}^{4}_{2}He}$$

SUB ATOMIC PARTICLES

These Includes:

ⓐ Proton; (^1_1H or ^1_1P)

ⓑ Neutron; ^1_0n

ⓒ Electron; β^- or $^0_{-1}e$ or $^0_{-1}\beta$

ⓓ Positron; β^+ or $^0_{+1}e$ or $^0_{+1}\beta$

ⓔ Gamma photon; $^0_0\gamma$

N.B In these Symbols

SubScript = atomic number eg ^1_0n has atomic number of zero

Super Script = Mass number ie (1) is its mass no

NB: In Radioactivity, it has been noted that element have a particular atomic and mass numbers that makes them stable enough to perform chemical reaction and they are called "MAGIC numbers".

- For protons, magic numbers are: 2, 8, 20, 28, 50, 82
- For neutron, magic numbers are: 2, 8, 20, 28, 50, 82, 126

Theoretically, it is observed that a system that has 114 protons should be stable ie 114 should be a magic number but experimentally it is not. Why? Assignment!!!

Answer

One of its answer is that the element that has 114 protons is Flerovium written as Fl. It's a superheavy artificial chemical element. Since it is formed by:

$$^{244}_{94}Pu + ^{48}_{20}Ca \rightarrow \underset{(unstable)}{^{292}_{114}Fl^*} \xrightarrow[decay]{hence} \boxed{^{290}_{114}Fl + 2(^1_0n)}$$

Since $^{292}_{114}Fl$ (unstable) further disintegrate (decay to its isotopes) with release of radiations (Gamma rays) to $^{290}_{114}Fl$, it is said to be unstable.

∴ Most natural radioactive decay gives a final stable product which contains a magic number. All radioactive decay series end up forming lead (Pb) as its stable final product.

Another Assignment

How many elements are in the periodic table? What is the last element in the periodic table and its symbol?

Answer

(I) We have 118 elements in the periodic table

(II) Unonoctium with symbol (Uuo) and with atomic number 118 is the last element

Most of the elements after Lawrencium (103) came about by transmutation, bombardment of different nuclei. University of California is where most (all) of new elements are discovered.

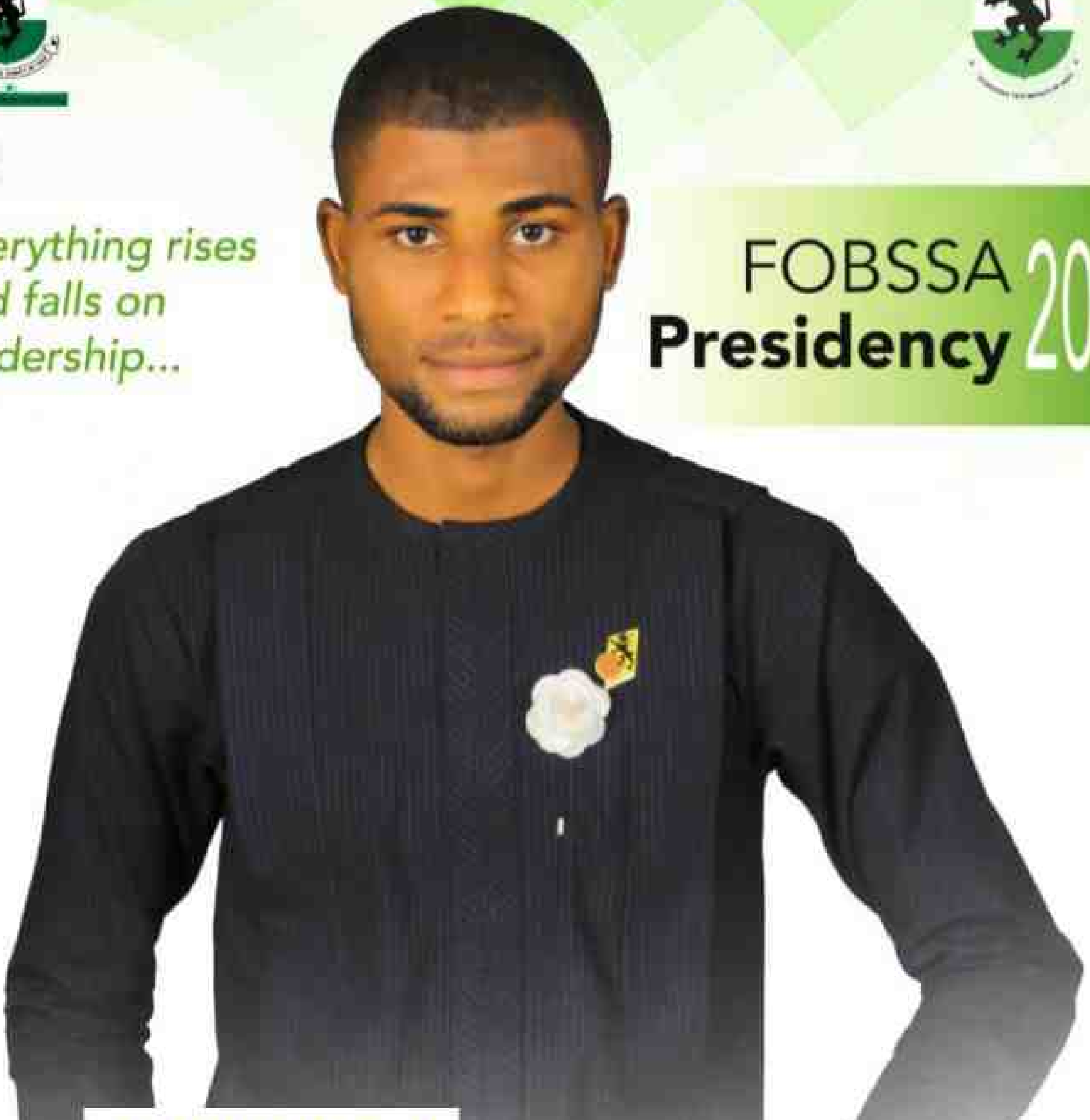
Stability also has to do with neutron proton ratio.

If a neutron proton ratio is ≥ 1 but < 1.5 it will form a stable nuclei. [ie n:p (1 – 1.5) b/w
This means that (n:p) very close to one or equal to one is very stable but n:p far from one is unstable ie too less or too large than one]

When n:p is too large [b/w 1.5 and above] or too small, Instability of the nuclei will be observed.

Everything rises and falls on leadership...

FOBSSA Presidency 2020



ONYEUKA

BRIGHT

CHIMEZIE

FACULTY OF BIOLOGICAL SCIENCE STUDENTS ASSOCIATION **(FOBSSA)**

*The Man with the **FOBSSA**ites needs at heart...*

All nuclei with Proton number greater than 83 are relatively unstable and for them to be Stable, they have to reduce their Proton number via disintegration or undergo radioactive decay.

Questions: Determine the n/p ratio of ⓐ $^{19}_{9}F$ ⓑ $^{64}_{29}Cu$ ⓒ $^{23}_{11}Na$

ⓓ $^{238}_{92}U$

Solution

ⓐ $^{19}_{9}F$ ⇒ neutron ⇒ 19−9=10 ; proton=9

n:p ⇒ 10:9 ie $\frac{10}{9}$

⇒ 1.11 ⇒ Stable

since it is in the range.

ⓑ $^{64}_{29}Cu$ ⇒ proton = 29 ; neutron = 64−29 = 35

n:p ⇒ 35:29 = 1.2 (stable)

ⓒ $^{23}_{11}Na$ ⇒ proton = 11 ; neutron = 12

n:p ⇒ 1.09 (stable)

ⓓ $^{238}_{92}U$ ⇒ proton = 92 ; neutron = 146

n:p ⇒ 1.6 (unstable since it is above ≥1.5)

Metastable is associated with gamma ray emission.

No stable nuclei are known with atomic number (83)

Assignment

Which element has atomic number 83

Answer ⇒ Bismuth (Bi)

All elements with atomic number of 83 or less have one or more stable nuclei with the exception of Technetium $_{43}Tc$.

TYPES OF RADIOACTIVE DECAY

The Common types of radioactive decay are

(i) Alpha emission: This is the emission of an alpha particle or Helium nuclei from an unstable nucleus.

$$^{226}_{88}Ra \xrightarrow{decay} {}^{222}_{86}Rn + {}^{4}_{2}He$$

↓
Radium

(ii) Beta Emission: This is the emission of light speed electron from an unstable nucleus. It is equivalent to the conversion of a neutron to a proton

eg $^{214}_{82}Pb \rightarrow {}^{214}_{83}Bi + {}^{0}_{-1}e$

$^{14}_{6}C \rightarrow ^{14}_{7}N + ^{0}_{-1}e$

(iii) POSITRON EMISSION: This is the emission of a positron from an unstable nucleus. This is equivalent to the Conversion of a proton to a neutron

Radioactive decay of Technetium (95)

$^{95}_{43}Tc \rightarrow ^{95}_{42}Mo + ^{0}_{1}e$

N.B When a positron and electron Collide, both particles Vanish (allenuated)

$^{0}_{1}e + ^{0}_{-1}e \rightarrow 2(^{0}_{0}\gamma)$

when they Vanish, they form 2 gamma photons that Carry away the energy ie they Vanish with the emission of 2 gamma photons.

(iv) ELECTRON CAPTURE:

This is the decay of an unstable nucleus by Capturing (pickingup) electrons from an inner orbital. In effect, a proton is Changed to a neutron in positron emission.

eg $^{40}_{19}K + ^{0}_{1}e \rightarrow ^{40}_{18}Ar$

It is believed that Argon in air is derived from Potassium which is usually found in rocks.

N.B When another orbital electron fills the Vacancy of the inner shell orbitals Created by electron Capture, an Xray photon is emitted.

⑤ Gamma emission: A gamma ray is a high energy electromagnetic radiation which is Simular to Xrays, and has neither Charge nor mass. They occur as a result of rearrangement of the neutrons and protons in a nucleus into lower energy states. It does not alter the neutron - proton ratio in the nucleus of atom.

The energy of the rays decreases in the order gamma > beta > alpha. In most Cases, gamma radiations go along alpha or Beta radiations.

$^{99}_{43}Tc \rightarrow ^{99}_{43}Tc + ^{0}_{0}\gamma$

$^{234}_{90}Th \rightarrow ^{234}_{91}Pa + ^{0}_{-1}e + ^{0}_{0}\gamma$

$^{234}_{90}Th \rightarrow ^{208}_{82}Pb + 6(^{4}_{2}He) + 4\,^{0}_{-1}e$

In this, it is the emission from an excited or meta stable nucleus. In many Cases radioactive decay results in a product nucleus that is in an excited state which on returning to the ground state emits γ particle.

⑥ Spontaneous fission : This is the spontaneous decay of an unstable nucleus in which a heavy nucleus of mass no >189 splits into lighter nuclei and energy is released.

HALF-LIFE OF A RADIOACTIVE NUCLEUS

Before we go into this, Remember

$$\frac{-dN}{dt} = \lambda N_0$$

$$\Rightarrow \frac{-dN}{N_0} = \lambda dt$$

Integrating both sides $\Rightarrow -\int \frac{1}{N_0} dN = \int \lambda dt$

$$\Rightarrow -\int_{N_0}^{N} \frac{1}{N_0} dN = \lambda \int dt$$

$$\Rightarrow -[\ln N - \ln N_0] = \lambda t$$

$$\Rightarrow \ln\left(\frac{N}{N}\right)^{-1} = \lambda t$$

OR

$$\Rightarrow \ln\left(\frac{N_0}{N}\right) = \lambda t$$

To Calculate halflife of a nucleus

$$\ln\left(\frac{N_0}{\left(\frac{N_0}{2}\right)}\right) \Rightarrow \ln\left(\frac{N_0}{\left(\frac{1}{2}N_0\right)}\right) = \lambda t_{1/2}$$

$$\ln 2 = \lambda t_{1/2}$$

$$\text{Half life } (t_{1/2}) = \frac{\ln 2}{\lambda} = \frac{0.693}{\lambda}$$

where λ = decay Constant

If you are asked to Calculate the number of disintegration of a radioactive nucleus

$$\frac{-dN}{dt} = \frac{\lambda \times \text{mass} \times \text{Avogadro number}}{\text{Molar mass}}$$

ie $\frac{\text{decay Constant} \times \text{mass} \times \text{Avogadro Number}}{\text{molar mass}}$.

APPLICATIONS OF NUCLEAR CHEMISTRY (RADIOACTIVITY)

Nuclear Chemistry (Radioactivity) has useful applications in the field of Archaelogy, Agriculture, Medicine, Power generation, Industry etc.

(1) Archaeological Applications : Radioactive disintegration is a useful phenomenon in the determination of the age of objects. The determination of age of ancient objects is based on the Characteristic half life of each radionuclide and the Constancy of the rate of decay with no external factor influence. This radioactive dating is possible because of the existence of the radioactive isotope called Carbon-14, ^{14}C, in the atmosphere. Carbon-14, is formed in the atmosphere via the action of Cosmic ray, neutrons on nitrogen. When energetic neutron bombards nitrogen, Carbon-14 is generated.

$$^{14}_{7}N + ^{1}_{0}n \rightarrow ^{14}_{6}C + ^{1}_{1}H$$

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Carbon-14 is oxidized in the atmosphere to Carbon (IV) oxide, which then mixes with the atmospheric CO_2.

Living organisms absorb this radioactive carbon in form of CO_2 through food chain when the living organism is dead the equilibrium is destroyed and the rate of decay reduces from that point since the Carbon-14 is no longer replenished. This process makes the estimation of how long ago the ancient object ceases to exist as a living thing by comparing the rate of disintegration of the object at the time with that for the living materials.

(ii) Agricultural Applications;

In agriculture, the growth and utilization of fertilizer by plants can be monitored using radioactive phosphate labelled P-32. The process involves feeding the plant with P-32 and the movement of the P-32 in the plant can be traced by Geiger Muller tube. This can be helpful to improve the yield of agricultural products. The phosphorous-32 can be used to detect the metabolic path.

Generally, radioactive emission is used to induce mutations in plants and animals in order to obtain improved varieties and desirable characteristics.

(iii) Medical Applications;

Iodine-131 which is both beta and gamma emitter with a halflife of 8 days is applicable in the diagnosis and treatment of thyroid diseases.

Thyroid Cancer can be treated by ingesting Iodine-131. The radioactive iodine-131, if concentrated on the location of the thyroid bombards the cancer cells with radiation which destroys the thyroid cancer.

Cobalt-60 and Caesium 137 is also useful in this area. Also $^{24}_{11}Na$ is used to monitor blood circulation, Technetium-99 used for brain and liver scans.

(iv) Nuclear power Generation application;

Nuclear fission can be monitored and controlled to generate nuclear fuel which serve as electricity source. The basic fuel for a nuclear reaction is Uranium 235.

$$^{234}_{92}U + ^{1}_{0}n \longrightarrow ^{90}_{38}Sr + ^{143}_{54}Xe + 3\,^{1}_{0}n$$

This is a typical example of artificial radioactivity

(v) Industrial applications: A radioactive material with a very short half life can be introduced into system of storage tanks and even underground pipelines inorder to locate any point of leakage using Geger Muller Counter detector. This can be achieved by introducing a known level or concentration of radioactive material into the tank or pipeline, if there is a leakage, the concentration of the radioactive material decreases.

Radioactivity is also applicable in forensic analysis and in environmental hazards investigations. Beta and Gamma radiations are used to control the thickness of sheet materials such as plastic, etc. Examples of radionuclide used for dating of rock materials are: Potassium-40, Uranium-235, Uranium-238, Thorium, Rubideum-87

NUCLEAR FISSION PROCESS

This is splitting of a heavy nucleus into 2 or more smaller nuclei. Fission of heavy nuclei is always accompanied by the ejection of 2 or more neutrons together with the release of gamma rays.

Fission processes take many different patterns:

$$^{235}_{92}U + ^{1}_{0}n \rightarrow ^{139}_{56}Ba + ^{94}_{36}Kr + 3(^{1}_{0}n) + ^{0}_{0}\gamma$$

$$^{235}_{92}U + ^{1}_{0}n \rightarrow ^{96}_{42}Mo + ^{138}_{50}Sn + 2^{1}_{0}n + ^{0}_{0}\gamma$$

FEATURES OF A FISSION PROCESS

(i) Heavy nucleus captures a neutron and split into 2 or more nuclei.

(ii) Splitting of a nucleus gives rise to two or more neutron production.

(iii) It leads to conversion of small masses into energy (gamma rays)

(iv) Most fission products are radioactive to give gamma radiations.

NUCLEAR CHAIN REACTION

Just like ordinary chemical chain reaction, nuclear chain reaction is a nuclear reaction that occurs in many steps. It is like a continous reaction. For instance, when U-235 nucleus is bombarded with a neutron, it produces 3 new neutrons. Each of the 3 neutrons produced in the reaction strikes another U-235 nucleus, thus causing 9 subsequent reactions. These 9 reactions can further give rise to 27 reactions. Hence, chain reaction is the process of propagation of the reaction by multiplication in 3's at each fission stages. Hence, fission chain reaction is described as a fission reaction where the neutrons from a previous step continue to propagate and replicate or repeat the reaction.

NUCLEAR FUSION PROCESS

At a very temperature, it is possible to fuse or join the nuclei of light isotopes to give those of heavier ones eg Deuterium and Tritium nuclei may fuse to give helium nuclei and neutrons

$${}_1^2H + {}_2^3H \xrightarrow{N} {}_2^4He + {}_0^1n$$

$$2\,{}_1^2H \longrightarrow {}_2^4He + {}_0^0\gamma$$

This how Hydrogen bomb operates. [ie a deadly bomb]
So Nuclear fusion is the process in which 2 light weight nuclei combines to form a single heavier nucleus.

The energy released by the sun results from a series of nuclear fusion process reactions. The overall reaction consists of fusion of four hydrogen nuclei (protons) to form helium nucleus

$${}_1^1H + {}_1^1H \rightarrow {}_1^2H + {}_1^0e$$

$${}_1^2H + {}_1^1H \rightarrow {}_1^3He + {}_1^0e$$

$${}_1^3H + {}_1^1H \rightarrow {}_2^4He + {}_1^0e$$

$$\text{Overall } 4({}_1^1H) \rightarrow {}_2^4He + 2({}_1^0e)$$

Differences between Nuclear fission and Nuclear fusion:-

Nuclear fusion	Nuclear fission
① Lighter nuclei combine together to form the heavier nucleus.	Heavier nucleus splits into lighter nuclei
② Fusion requires very high temperature	It does not require extreme high temperature before it can occur.
③ It does not lead to chain reaction	It leads to chain reaction
④ Fusion reaction is mostly unControllable	It can be monitored and Controlled.
⑤ The energy released cannot be put properly into use	The energy released can be channeled for peaceful usage.
⑥ Fusion gives rise to products which are non radioactive in nature.	Most product of fission reactions are radioactive in nature.
⑦ Fusion reactions does not left behind nuclear waste at the end	At the end of fission reaction nuclear waste is left behind.

NOTE :

RADIOACTIVE DECAY SERIES

All nuclides with $Z > 83$ (ie with atomic number > 83) are radioactive and decay most of the time by α emission.

There are 3 major type of naturally occuring radioactive decay series.

Radioactive decay series is a sequence in which one radioactive nucleus decays to a 2nd which then decays to a 3rd then to a 4th and so forth till it gets to a stable nucleus which is an isotope of lead.

① Radioactive decay series of U-238

$${}^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}He \quad (U\text{-}238 \xrightarrow{\alpha} Th\text{-}234)$$

$${}^{234}_{90}Th \rightarrow {}^{234}_{91}Pa + {}^{0}_{-1}e \quad (Th\text{-}234 \xrightarrow{\beta} Pa\text{-}234)$$

$${}^{234}_{91}Pa \rightarrow {}^{234}_{92}U + {}^{0}_{-1}e \quad (Pa\text{-}234 \xrightarrow{\beta} U\text{-}234)$$

$${}^{234}_{92}U \rightarrow {}^{230}_{90}Th + {}^{4}_{2}He \quad (U_{234} \xrightarrow{\alpha} Th\text{-}230)$$

② The second decay series of U-235 ends with Pb-207

${}^{235}_{92}U$ $\xrightarrow{\text{decays}}$ . . . ends with Pb-207

③ The 3rd decay series of Th-232 ends with Pb-208.

Nuclear Bombardment Reaction / Transmutation.

Ernest Rutherford discovered that it was possible to change one nucleus by nuclear particles.

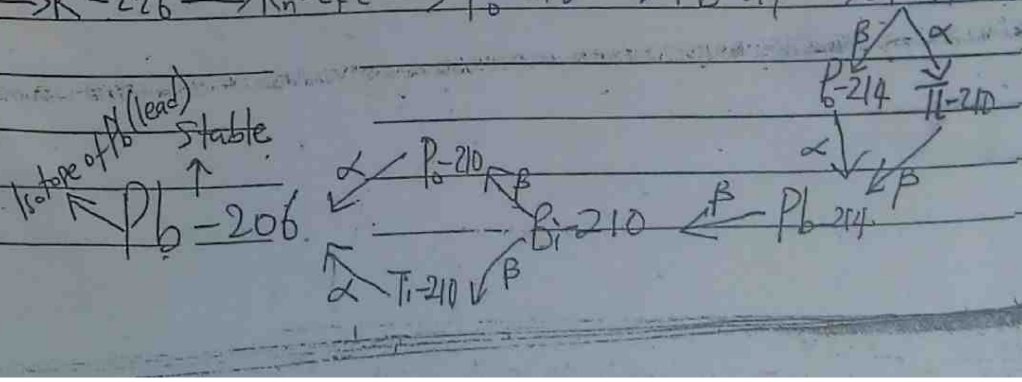
Transmutation is the change of one element to another by bombarding the nucleus of the elements with nuclear particles/nuclides.

Example ${}^{14}_{7}N + {}^{4}_{2}He \rightarrow {}^{17}_{8}O + {}^{1}_{1}H$

Chadwick in the discovery of neutrons also did (transmutation)

$${}^{9}_{4}Be + {}^{4}_{2}He \rightarrow {}^{12}_{6}C + {}^{1}_{0}n$$

$$U\text{-}234 \xrightarrow{\alpha} Th\text{-}230 \xrightarrow{\alpha} R\text{-}226 \xrightarrow{\alpha} Rn\text{-}222 \xrightarrow{\alpha} Po\text{-}218 \xrightarrow{\alpha} Pb\text{-}214 \xrightarrow{\beta} Bi\text{-}214$$



The first artifical radioactive isotope was produced in 1933 by bombardment of Al with α-particles.

$$^{27}_{13}Al + ^{4}_{2}He \rightarrow \underbrace{^{30}_{15}P}_{\text{first artificial radioactive isotope}} + ^{1}_{0}n$$

$$^{30}_{15}P \rightarrow ^{30}_{14}Si + ^{0}_{-1}e \quad \text{(further decay.)}$$

Notation for nuclear bombardment reaction. These are represented by abbreviated notations:

$$^{14}_{7}N(\alpha, p)^{17}_{8}O$$

where α = alpha particle

P = Proton $^{1}_{1}H$

Rule → In this notation, 1st write the nuclide symbol for the original nucleus, then in bracket, write the symbol of the projectile (incoming) particle, α-particle, then comma after which the symbol of the ejected particle (Proton), close the bracket, then write the product symbol.

PARTICLE SYMBOL

n = neutron

p = proton

$^{2}_{1}H$ = Deuterium, d

$^{4}_{2}He$ = α particle

Note: Most of the times when you are projecting the projectile particle, it must posses enough kinetic energy, capable to cause an impact or penetrate or else it will be drifted away.

The Cyclotron gives enough acceleration to the projectile particle so that it attains enough K.E capable to cause a nuclear reaction. A particle accelerator is used to give kinetic energy to the projectile particle to overcome repulsion force from the nucleus of the element to enable it penetrate and react. The device is known as a Cyclotron.

Tc (Technetium) was first prepared by directing the nucleus of Hydrogen – 2 atoms from a Cyclotron to a Mo target.

Condensed reaction → $^{96}_{42}Mo(d, n)^{97}_{43}Tc$

i.e $^{96}_{42}Mo + ^{2}_{1}H \rightarrow ^{97}_{43}Tc + ^{1}_{0}n$

Trans-Uranium element

The naturally existing element has atomic number, Z = 92, beyond which are those produced from transmutation. Trans-Uranium elements are with atomic number greater than 92

The first trans uranium element is $^{239}N_p$ Produced in 1940 by bombarding U-238 with neutrons.

$$^{238}_{92}U \longrightarrow {}^{238}N_p + {}^{0}_{-1}\beta$$

$$^{238}_{92}U + {}^{1}_{0}n \longrightarrow {}^{239}_{92}U$$

The second trans-Uranum element is Plutonium, P_u. This was produced by bombarding U-238 with a deuteron particle.

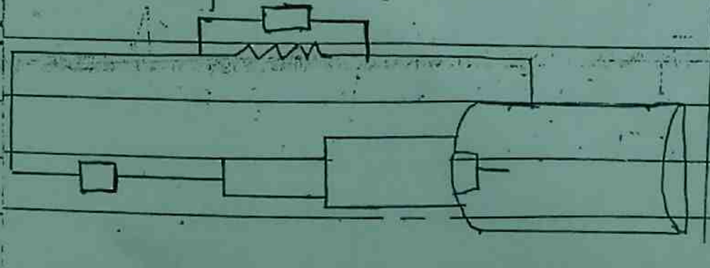
$$^{238}_{92}U + {}^{2}_{1}H \longrightarrow {}^{238}_{93}N_p + 2({}^{1}_{0}n)$$

$$^{238}_{93}N_p \longrightarrow {}^{238}_{94}P_u + {}^{0}_{-1}\beta$$

When radiations have contact with matter, it can ionize it, break its bonds or lead to effects in animals and plants.

DEVICES FOR COUNTING RADIATIONS

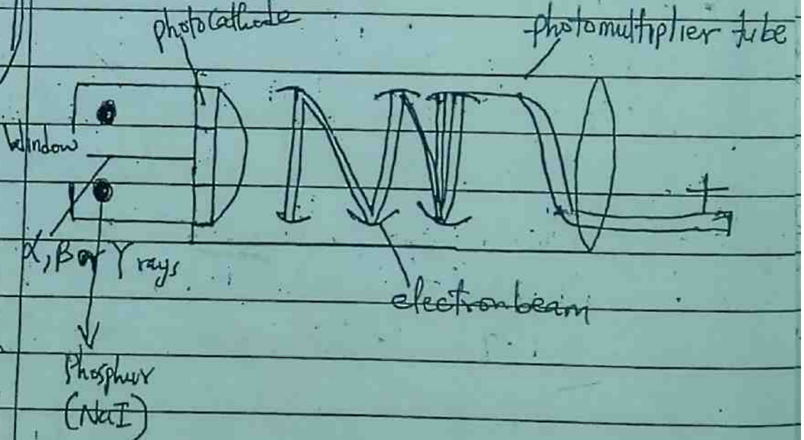
① Ionizations Counter eg Geiger muller Counter → detects the production of ions in matter.



The Geiger muller Counter Can be used to defect α and β particles directly but Cannot detect neutrons directly.

② Scintillation Counter: is a device that detects nuclear radiation from the flashes of light generated in a material by the radiation. It uses the phenomenon of photoelectric effect. It makes use of phosphur (NaI).

Phosphur is a substance that emits flashes of light when struck by radiation.



NaI Crystal Containing thalium I iodide is used as the phosphur to detect γ rays while a radiation Counter Can be used to measure the rate of nuclear disintegration in a radioactive material.

The activity of a radioactive Source is the no of nuclear disintegration per unit time occuring in a radioactive material.

Unit = Curie (Ci)
$= 3.7 \times 10^{10}$ disintegrations per Second.

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K-40

If P-40 has activity of 1.0×10^{-2} Ci, that means the material is disintegrating at this rate $= 1.0 \times 10^{-2} \times 3.7 \times 10^{10}$

$= 3.7 \times 10^{8}$ disintegrations/sec

KINETICS OF RADIOACTIVE DECAY

The rate of radioactive decay cannot be changed by varying temperature, pressure or the chemical environment of the chemical nuclei.

The rate of radioactive decay is the no of nuclei that disintegrates per unit time.

It is proportional to the no of radioactive nuclei in the sample.

$$Rate = K N_t$$

where N_t → no of radioactive nuclei at time, t

k ⇒ radioactive decay constant which is characteristic of a particular nuclei.

The rate of radioactive decay is a first order process.

HALF LIFE OF A RADIOACTIVE NUCLEUS ($T_{1/2}$)

The half life is the time it takes for one half of the nuclei in the sample to decay.

It is independent on the amount of sample. It can be obtained directly by observing how long it takes for one half of the sample of a radioactive nucleus to decay.

FOR EXAMPLE:

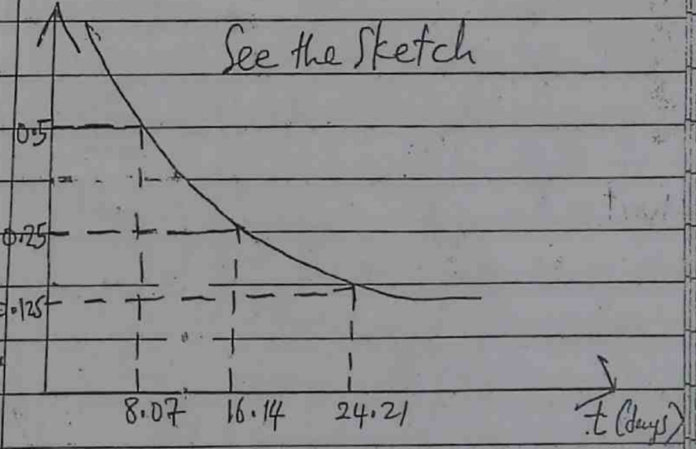
I-131

1g of I-131 decays to 0.5g which takes 8.07 days

0.5g of I-131 decays to 0.25g which takes 8.07 days.

Thus half life is 8.07 days

See the sketch



Because some radioactive nuclei decay quickly while some decay slowly, it's impossible to calculate the half life of a sample directly.

$$t_{1/2} = \frac{0.693}{K}$$

Formular Relating half life and decay constant

$t_{1/2}$ = half life

K = decay constant

To Calculate the fraction remaining after a given period of time Can be done using :

$$\ln \frac{N_t}{N_0} = -kt \quad \text{OR} \quad \log \frac{N_t}{N_0} = \frac{-kt}{2.303}$$

N_0 => no of nuclei in the original Sample at $t = 0$; t = period of time

N_t = no of nuclei after, the time, t

Fraction remaining = $\frac{N_t}{N_0}$

from $\ln \frac{N_t}{N_0} = -kt$

$$\Rightarrow \frac{N_t}{N_0} = e^{-kt}$$

No of nuclei remaining at a period of time given, N_t = $N_0 e^{-kt}$ where $e = \underline{\underline{2.718}}$

no of nuclei decayed at a period of time given = $N_0 - \left(N_0 e^{-kt}\right)$

OR

no of nuclei remaining at a period of time given, N_t = $N_0 \left(\frac{1}{2}\right)^{t/T_{1/2}}$

where t = time given

$T_{1/2}$ = half life

No of nuclei decayed at a period of time given = $N_0 - \left[N_0 \left(\frac{1}{2}\right)^{t/T_{1/2}}\right]$

Fraction remaining = $\left(\frac{1}{2}\right)^{t/T_{1/2}}$

Summary on the Applications of Radioactivity

① It is used in radioactive-dating with Carbon-14 to detect the age of a particular animal.

K-40 Can be used to detect the age of rocks.

② It is used to prepare trans-uranium elements (by transmutation)

③ It Can be used as chemical in Chemical analysis, which Can be used as a radioactive tracer and neutron captivation analysis

Meta-stable nucleus of Arsenic for neutron Captivation analysis.

④ It Can be used for medical therapy and diagnosis.

Cancer therapy – Ra-226 and Ru-222

recently Co-60 is also used.

⑤ To diagonise diseases by generating images to determine the parts of body it is in.

Eg Metastable Th-201.

⑥ To generate electricity.

⑦ To produce nuclear weapons eg Nitrogen bomb, hydrogen bomb and Atomic bomb.

MASS DEFECT

Although the nucleus of an atom is composed of protons and neutrons, it is a known fact that the masses of the nucleon is never the same as mass of the composite particle. For instance, the measured mass of the helium nucleus is 4.0002764 a.m.u. but the total mass of the constituent part is significantly more.

It is made up of 2 protons, 2 neutrons and 2 electrons.

The mass of helium nucleus can therefore be calculated thus:

2 protons = 2 x 1.007273
= 2.014552 a.m.u

2 neutrons = 2 x 1.008665 a.m.u
= 2.017330 a.m.u

where 1.007273 (is the relative mass of protons in a.m.u.
1.008665 (is the relative mass of neutron in a.m.u.

=> Total calculated mass = 2.014552 + 2.017330
= 4.030375 a.m.u.

Observed mass was given as 4.002764 a.m.u.

The difference between the observed and calculated mass ie (4.030375 — 4.002764 a.m.u) gave what is called the MASS DEFECT.

Mass defect represent the energy liberated whien the proton and neutron bind together to form the nucleus. It accounts for condition in which matter is converted to energy. Inter conversion of mass and energy this way predicted in Einstein's special theory of relativity can be expressed by the relationship

$$E = \Delta m c^2$$

where E = energy, Δm = change in mass, C = velocity of light in m/s.

The greater the Δm, the greater the stability of the nucleus.

Question: Calculate the mass defect of Neon ${}^{24}_{10}Ne$ given that the mass of neon from mass spectrometer measurement is 20.183. Take masses of proton, Neutron and electron as 1.0075, 1.0089, 0.006 respectively relative to Carbon 12.

Solution

Δm = Calculated mass – Observed mass

(Mass of Protons X no of Protons) + mass of neutrons X no of neutrons = Calculated mass

$(1.0075 \times 10) + (1.0089 \times 11) = 21.1729$

$\Delta m = 21.1729 - 20.183$

$= 0.9899$.

Hence the mass defect of Neon is 0.9899 g.

BINDING ENERGY

The actual mass per mole of a particular element is always less than the Calculated mass. The difference is the mass defects of the nucleus. The mass defect of a nucleus is equivalent to the energy that holds the nucleon together. Using Einstein equation, we can Calculate the energy equivalent of the mass defect.

The energy equivalent of a mass defect divided by the number of nucleon present in the nucleus is referred to as the binding energy per nucleon. The binding energy per nucleon is a measure of the magnitude of the force binding the nucleons. The unit of energy Can be electron Volt (eV) or mili electron volt

⇒ $1eV = 1.6 \times 10^{-19} J$

Binding energy is the energy that is required to break or Separate a nucleus into its Constituents nucleon. Hence the relative stabilities of Various nuclei Can be estimated from their respective binding energy.

$$\text{Binding energy per nucleons} = \frac{\Delta m}{\text{Mass number}}$$

$$= \frac{\text{Calculated mass} - \text{Observed mass}}{\text{mass number}}$$

Unit ⇒ g/mol or J/mol

NB Observed mass = actual mass.

Question: Calculate the binding energy in SI units per nucleon for $^{64}_{29}Cu$ given the following Parameters:

Observed or actual mass of Cu = 63.550

Proton = 1.0075 g/mol ; Neutron = 1.0089 g/mol

electron = 0.0006 g/mol

Take 1g = binding energy of 9×10^{13} J.

Solution

$^{64}_{29}Cu$ ⇒ mass no = 64

atomic no = 29 = Proton no

neutron = 35 (64–29)

Δm = Calculated mass – Actual mass

$= ((1.0075 \times 29) + (1.0089 \times 35)) - 63.550$

$= 64.528 - 63.550$

$= 0.979 g/mol$ ⇒

$$\text{Binding energy per nucleon} = \frac{0.979 g/mol}{64}$$

B.E per nucleon $= 1.53 \times 10^{-2}$ g/mol

$= 1.53 \times 9 \times 10^{13} \times 10^{-2}$

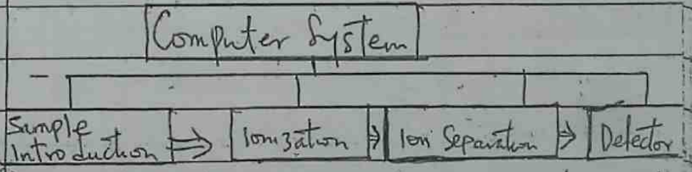
$= 1.37 \times 10^{12}$ J/mol

DETERMINATION OF MASSES

- Isotopes, mass number, atomic no
- Quantum theory
- Dual nature of matter (De Broglie)
- Planck's theory
- Photoelectric effect
- Atomic Spectrum (H₂)
- Bohr's atomic theory and its limitations.
- Uncertainty principle, Schrodinger ideas, Quantum no.

MASS SPECTROMETER

This is an Instrument used for the determination of masses or relative abundance of Isotopes of an atom. It is also used in Separating Isotopes of elements and estimation of atomic and molecular masses of Substances in chemical analysis.
It was Invented by J.J-Thomson and F-W Austin.



Block diagram of features of a typical mass.

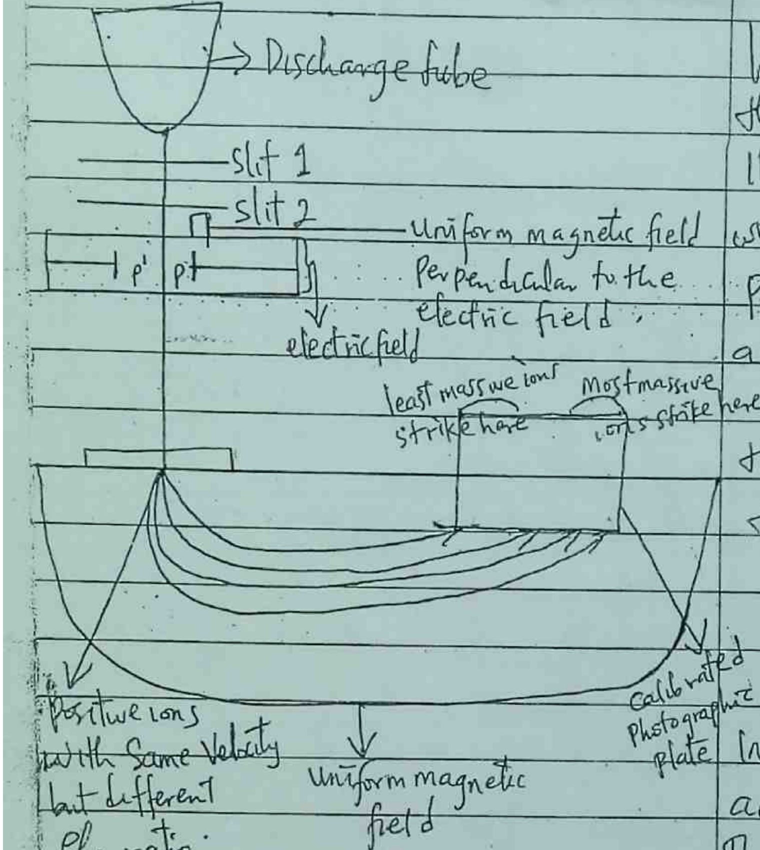


Diagram of a mass Spectrometer.

OPERATIONS OF MASS SPECTROMETER

At the ionization chamber, +ve ions are produced (electron Impact or chemical ionization method.). These ions are made to pass through an electric field and a uniform magnetic field perpendicular to the electric field,

At the ions Separation chamber, +ve ions of Same change and Velocity but different masses are Separated and allowed to enter a uniform magnetic field. The magnetic field Sucks out the ions by deflection with respect to their differences in masses, lighter ions are more strongly deflected than heavier ions - The detector is usuarlly a photographic plate. Here, +ve ions with the Same e/m ratio strike the photographic plate at one point which appears as a black spot. The relative intensity of the black Spot Indicater the relative no of ions that strike that particular point.

The modern type of Spectrometer are Interfaced with a Computer which Controls the operation of the Instrument, Collects and stores data and provides graphical output.
This graph is Called the mass Spectrum and is used in determining the atomic or molecular masses of the Sampler.

USES OF MASS SPECTROMETER

① It is used in the estimation of masses and Relative abundance of Isotopes in an atom.

② it is used in the discovery and Study of Isotopes of elements.

③ It is used in the Separation of Isotopes of atoms of an element

CHYDONN'S QUOTE

You only Say you know a particular Very well, if you are able to explain it well to your grandmother.

Albert Einstein ....

ISOTOPES

Isotopes are atoms of the same element that have the same atomic no but different mass no. The difference in mass no arises due to difference in the no of neutrons.

A specific atomic nuclei with given mass no, A, and atomic no, Z, is called a nuclide. The general symbol for designating a nuclide is

$$X^{A}_{Z} \quad \text{or} \quad {}^{A}_{Z}X$$

where X = Symbol of the element
Z = atomic no
A = mass no

Examples of elements that exhibit Isotopy are Chlorine, Oxygen, Carbon, Hydrogen, Nitrogen etc.

$Cl_2 \rightarrow {}^{35}_{17}Cl, {}^{37}_{17}Cl$

$O_2 \longrightarrow {}^{16}_{8}O, {}^{17}_{8}O, {}^{18}_{8}O$

$C \rightarrow {}^{12}_{6}C, {}^{13}_{6}C$

$H_2 \longrightarrow {}^{1}_{1}H, {}^{2}_{1}H, {}^{3}_{1}H$

$N_2 \longrightarrow {}^{14}_{7}N, {}^{15}_{7}N$

Isotopes of an element have identical chemical property but different physical property.

This is due to the fact that the chemical property of an atom depends on its electron. Since most naturally occuring elements including C have more than one Isotopes, average atomic mass becomes the more accurate expression of the atomic mass of the element. The various masses and relative abundance of Isotopes of an element can be used to calculate the average atomic mass of the element, using the following formular:

$$\text{Average atomic mass} = \frac{\begin{bmatrix}\%\text{abundance} \\ \times \\ \text{mass of 1st Isotope}\end{bmatrix} + \begin{bmatrix}\%\text{abundance} \\ \times \\ \text{mass of 2nd Isotope}\end{bmatrix}}{100}$$

To write it clearer

$$\text{Average atomic mass} = \begin{bmatrix}\%\text{abundance} \\ \times \\ \text{mass of 1st Isotope}\end{bmatrix} + \begin{bmatrix}\%\text{abundance} \\ \times \\ \text{mass of 2nd Isotope}\end{bmatrix}$$

Example: Calculate the average atomic mass of Chlorine given that:

Isotopes	% abundances
${}^{35}_{17}Cl$	77.2
${}^{37}_{17}Cl$	22.8

Solution

$$\text{Average atomic mass} = \frac{(77.2 \times 35) + (22.8 \times 37)}{100}$$

$$\Rightarrow \frac{2702 + 843.6}{100}$$

$$\Rightarrow 35.456 \text{ a.m.u.}$$

NOTE THE FOLLOWING

ⓐ Iso-electronic structure.

The existence of or occurrence of two or more species with the same electronic configuration (ie same number of electrons) is termed Isoelectronic.

for Instance :

O^{2-}, N^{3-}, F^-, Ne, Na^+, Mg^{2+}, Al^{3+}, etc are Iso electronic because they have 10 electrons

Ca^{2+}, Ar, P^{3-}, S^{2-}, Cl^-, K^+, Sc^{3+} are Iso electronic because they have 18 electrons

Some ionic compounds and neutral compounds also exhibit Iso electronic structure.

for Instance, 8 electron compounds are BH_4^-, NH_4^+, H_2O, CH_4 etc.

ⓑ

16 electron compounds are CO_2, NO_2^+, etc.

ISOBAR

They are Isotopes of different elements having the same nucleon number ie, they have same mass no but different atomic no.

examples

$^{40}_{18}Ar$, $^{40}_{19}K$, $^{40}_{20}Ca$

ISOTONES

These are elements with the same number of neutrons but different nucleon number ie, same neutron but different protons hence different nucleon (mass) number.

$^{30}_{14}Si$ $^{31}_{15}P$ $^{32}_{16}S$

(Their common neutron no = 16) but different proton & mass no

MASS NUMBER

The mass no of an atom is the Sum of the number of neutrons and Protons present in the nucleus of the atom.

ATOMIC NUMBER

This is the no of Protons in the nucleus of an atom. In a neutral atom, the atomic no = no of electrons.

MORE QUESTIONS ON STOICHIOMETRY.

① What mass of ZnS will be formed by 0.5g of Zn when Zn reacts with Sulphur.

Solution

$Zn + S \rightarrow ZnS$

1mole 1mole 1mole

65g of Zn produces 97g of ZnS

0.5g of Zn " $\frac{0.5 \times 97}{65}$ g of ZnS

$\Rightarrow 0.75$g of ZnS

② 0.5g of Zn was made to react with 0.5g of S, Calculate the mass of ZnS formed.

Solution

$Zn + S \rightarrow ZnS$

Zn = 65g S = 32g

Number of Zinc = $\frac{0.5}{65}$ = 0.00769

no of moles of Sulphur = $\frac{0.5}{32}$ ⇒ 0.056

Remember, the limiting reactant is the reactant that is Consumed Completely and determines the way the reaction goes.

Trial reaction Calculation

65g of Zn ≡ 32g of S

0.5g of Zn = $\frac{0.5 \times 32}{65}$ = 0.25g of S

Since Sulphur that Supposed to react is 0.25g instead of 0.5g, Sulphur is in excess ie excess reactant whereas Zn is the limiting reactant.

Thus 65g of Zn → 97g of ZnS,

0.5g of Zn → $\frac{0.5 \times 97}{65}$

= 0.746g of ZnS

③ 5g of Oxygen was made to react with 5g of Hydrogen in order to form Steam.

(i) Which of the reactant is in excess.

(ii) By what amount in gram?

(iii) What is the mass of Steam Produced?

Solution

Equation

$2H_2 + O_2 \rightarrow 2H_2O$

2 : 1 2

2moles of H_2 reacts with 1mole of O_2

2×2g H_2 will react with 32g of O_2

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Since

4g of H_2 reacts with 32g of O_2

5g of H_2 " " $\frac{5\times32}{4}$

$= 40g$ of O_2

Since the Oxygen that supposed to go into reaction is 40g instead of 5g that was given in the question, Oxygen is the limiting reactant where as Hydrogen is the excess reactant.

② Since Oxygen is the limiting reactant, it is the Oxygen that we will use to compare and determine by what amount in gram is the excess reactant.

32g of O_2 reacts with 4g of H_2

5g of O_2 reacts with $\frac{4\times5}{32}$

$\Rightarrow 0.625g$ of H_2

Excess reactant $\Rightarrow 5g - 0.625g$

$\Rightarrow 4.375g$ of H_2.

③ 1 mole of O_2 produces 2moles of H_2O

32g of O_2 produces $2\times18g$ of H_2O

5g of O_2 " $\frac{5\times36}{32}$

$= 5.625g$ of H_2O

OR

2moles of H_2 produces 2moles of H_2O

4g of H_2 " 36g of H_2O

0.625g of H_2 " $\frac{0.625\times36}{4}$

$= 5.625g$ of H_2O

ie 5.625g of H_2O

QUANTUM THEORY

Electromagnetic Radiation include: Light rays, Xrays, γrays, Radio waves, Micro waves etc. Electromagnetic radiations involves 2 models namely: wave model and particle model.

In terms of wave model, the disturbances caused by electric and magnetic fields throughout space setup electromagnetic waves.

The dispersion of electromagnetic rays is called the Electromagnetic Spectrum.

← decreasing wavelength (nm)

γ rays	X rays	UV rays	Visible rays	Infrared rays	Micro wave	Radio Frequency

→

~~decreasing frequency~~

Electromagnetic Spectrum.

An Electromagnetic radiation is characterised by certain parameters

① frequency, $v = \frac{c}{\lambda}$ ← speed of light

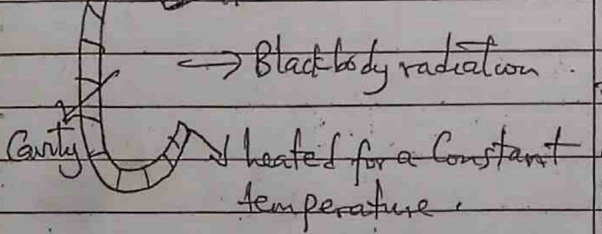
λ → wavelength (A° or (amstrong))

frequency = Hertz or S^{-1}

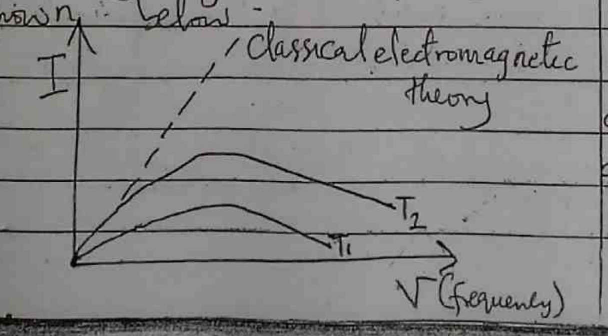
② Wave number = $\frac{1}{\lambda}$ ie reciprocal of wavelength.

BLACK BODY RADIATION

A black body is a substance/surface that will absorb and emit all wavelength of electromagnetic radiation incident on it. A black body radiation could emerge from a pinhole in the wall of a closed material kept at a fixed high temperature. It loses the energy absorbed by emitting radiation of all frequencies.



A plot of the graph of intensity of the emitted radiation from a heated black body against frequency is shown below.



where I = intensity of absorbed/emitted radiation.

Classical electromagnetic theory:

From classical electromagnetic theory, intensity is supposed to increase with frequency ie heat is absorbed in a continous manner, but this is not so, intensity increases to a max. value and decreases exponentially (according to T_1 and T_2).

All attempts to use classical electromagnetic theory to predict the wavelength of emitted light from a black body had failed except from the assumption made by Max planck.

Max planck's Proposal.

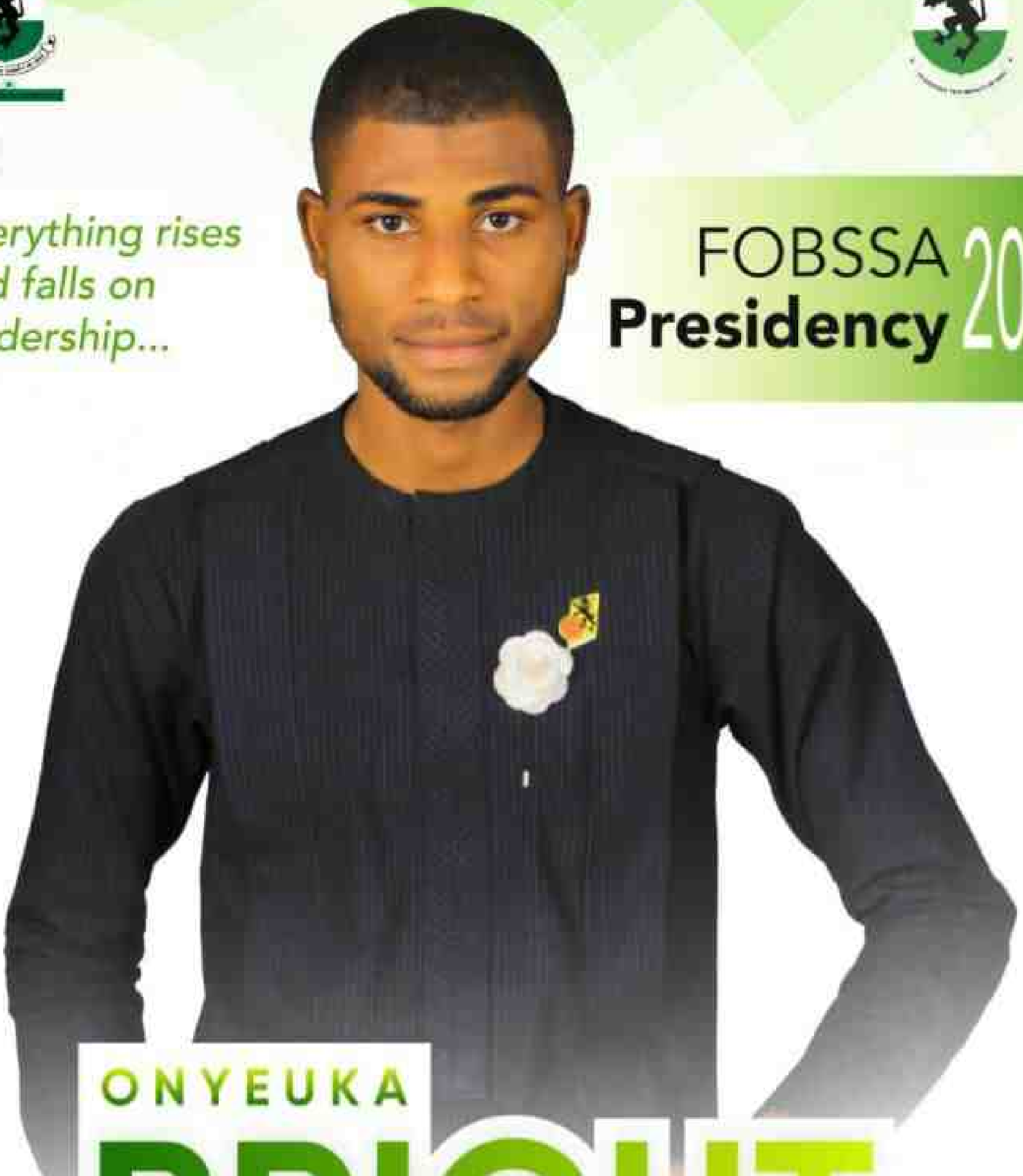
In 1900, Max plank a German, Physicist proposed the quantum theory in it's simplest form. He arrived at the theory while studying the radiation, emitted by a hot black body.

Planck proposed that the amount of energy absorbed or emitted by a body is not continous but it exist in discrete packets called Quanta.

He established that the energy absorbed/emitted in quanta is proportional to the frequency.

Everything rises and falls on leadership...

FOBSSA 2020
Presidency



ONYEUKA

BRIGHT

CHIMEZIE

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*The Man with the **FOBSSA**ites needs at heart...*

ie $E \propto V$

Thus $E = KV$ (where K = Planck's Constant)

$E = hV$ ....

E = energy, V = fequency

h = planck's Constant.

$E = hV$; Remember $V = \frac{C}{\lambda}$

$\Rightarrow E = \frac{hC}{\lambda}$; $h = 6.625 \times 10^{-34}$ J.S

$= 6.625 \times 10^{-27}$ erg.s

Planck's theory was able to bring together particle model of light with the wave model property of light, using the Einstein equation

$E = \frac{hc}{\lambda}$ [wave model equation]

$E = MC^2$ [Particle model equation]

where m = mass, C = Speed of light.

If you equate the 2 above equation

$M\overset{C}{\cancel{C^2}} = \frac{h\cancel{C}}{\lambda} \Rightarrow MC = \frac{h}{\lambda}$

when extended to a particle with finite velocity, V

$mV = \frac{h}{\lambda} \Rightarrow \lambda = \frac{h}{mV}$

This expression relates the momentum of a particle to its equivalent wavelength as suggested by De broglie

ie $\lambda = \frac{h}{mV}$ where MV = Momentum

λ = is a wave characteristic. and mV is a particle characteristic

λ = wavelength

Louis de Broglie used the expression $\lambda = \frac{h}{mV}$ to find the wavelength for a stream of electrons.

λ depends on Velocity of electrons.

Photoelectric effect

When U.V. light is incident, on some substances eg Zn plate, electrons of the Zn are ejected from the plate. This phenomenon is called Photoelectric effect while there substances are photoelectric substances

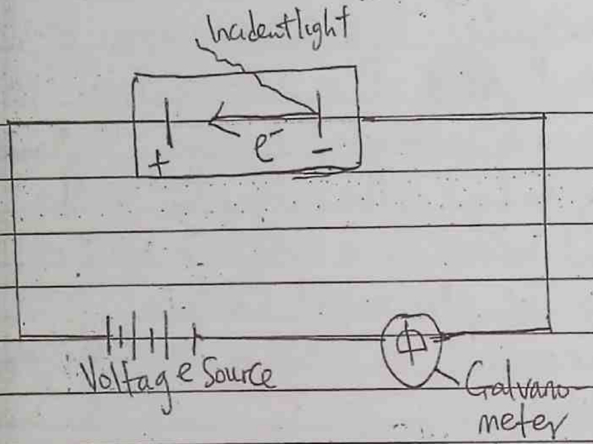


DIAGRAM SHOWING A PHOTOCELL

The photoelectric effect Could not be explained by Classical physics which had no quantum Concept. A detailed study of the photo-electric effect revealed, how the behaviour of the e^- is related to the xteristics of the incident light:

1. Below a xteristics threshold frequency, f_0, no electron is observed regardless of the light intensity.
2. Above the f_0, the electron is ejected and its K.E is independent on the light intensity.
3. Above the f_0, the no of emitted electron 'e^-' increases with the light intensity.
4. All metals exhibit the Same pattern, but each metal has different f_0.

Threshold Voltage, V_0

This is the minimum Voltage, requred to repel the most energetic electron, thereby Causing zero (0) plate Current.

Threshold frequency, f_0: This is the minimum frequency of light requred to liberate electrons from a photoelectric Substances

ALBERT EINSTEIN EXPLANATION

Einstein postulated that light Comes in packets Called photons. Each photon is equal in magnitude to 1 planck's quanta of energy and is a function of frequency.

$$E = hf$$

where f = fequency
h = planck's Constant
E = photon of light.

He explained that when a metal Surface absorbs quantum of energy, the energy is transferred to the electrons, Some of the energy is used to overcome the forces that bind the electrons to the metal while the remainder is used as kinetic energy of the ejected electrons.

K.E of electrons $= hf - W_0$

K.E

$hf = k.E + W_0$ where W_0 = work function/energy required to dissociate the electron.

The threshold frequency, f_0, correspond to the binding frequency of the electron. In other words, the energy of the photon at the threshold frequency equals the minimum energy needed to overcome the forces that bind the electrons to the metal.

$$hf = hf_0 + K.E$$

where $hf_0 = W_0$ (Work function)

hf = Photon energy

Einstein's explanation accounted for all the 4 properties of the photoelectric effect:

① When the energy of the photon is less than the hf_0 (Work function) then the energy is not enough to overcome the electrons binding energy. So no electrons can be ejected from the metal surface, no matter the intensity of the light.

② After the energy of the photon exceed the threshold value, electrons are ejected. The extra energy is used as K.E and this extra energy increases linearly with frequency. The intensity of the light beam is a measure of the no of photons it contains not the amount of energy each photon possess.

Each metal has its own Xteristics threshold frequency because electrons are bound more tightly to some metals than others.

Example

① The minimum energy needed to overcome the attractive force between the electron and the surface of silver metal is given as 7.52×10^{-12} erg. Calculate the maximum K.E of the electrons, ejected from silver which is being irradiated by U.V. light with wavelength of 360 A°. ($C = 3 \times 10^{10}$ cm/s)

Solution

$$hf = hf_0 + K.E_{max}$$

$c = 3 \times 10^{10}$ cm/s $= 3 \times 10^{8}$ m/s

$\lambda = 360 A^\circ = 360 \times 10^{-10}$ m

$= 3.6 \times 10^{-8}$ m

$K.E_{max} = hf - hf_o$

where $hf \Rightarrow \frac{hc}{\lambda}$

$K.E = \frac{hc}{\lambda} - hf_o$

$= \frac{6.625 \times 10^{-27} \times 3 \times 10^{8}}{3.6 \times 10^{-8}} - 7.52 \times 10^{-12}$

$= 5.521 \times 10^{-11} - 7.52 \times 10^{-12}$

$= 4.77 \times 10^{-11}$ erg.

② The work function of a particular metal is 0.83×10^{-18} J. What is the frequency of light that causes electrons to be emitted with a maximum K.E of 2.0×10^{-19} J

solution

$hf = W_o + K.E$

$f = \frac{W_o + KE}{h}$

$\Rightarrow \frac{(0.83 \times 10^{-18} + 2.0 \times 10^{-9})}{6.625 \times 10^{-34}}$

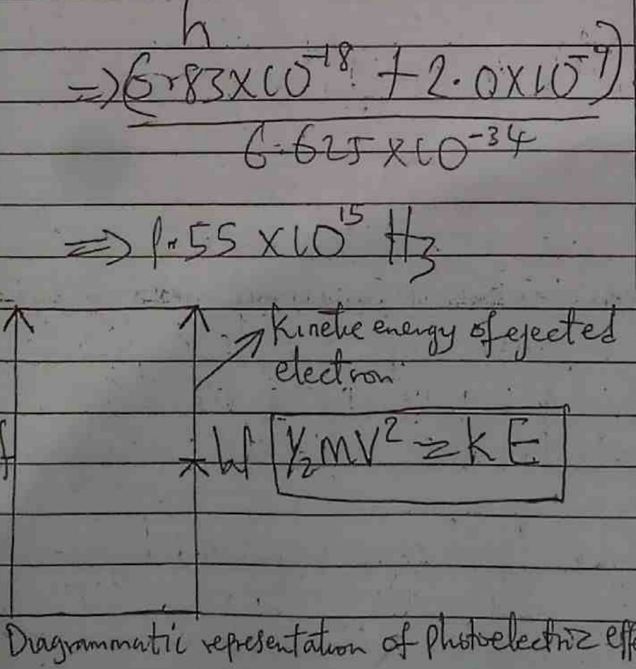
$\Rightarrow 1.55 \times 10^{15}$ Hz

ATOMIC SPECTRUM

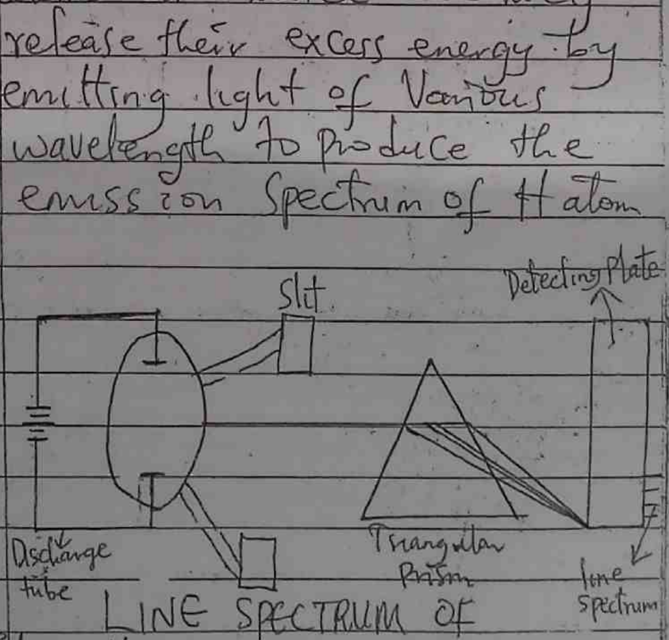
Atomic Spectra involves exchange of energy between an atom and electromagnetic radiation. The energy of an atom increases on absorbing a photon of energy. The atom moves into a higher energy state called the excited state. Atoms in excited state also give up their excess energy and return to their lowest energy state called ground state. The energy change is given as

$$\Delta E = E_{final} - E_{initial}$$

When an electric discharge is passed through hydrogen molecule in a discharge tube, the H_2 molecules gain energy and dissociate into H atoms. The resulting H atoms are excited and they release their excess energy by emitting light of various wavelength to produce the emission spectrum of H atom



Diagrammatic representation of photoelectric effect.



LINE SPECTRUM OF HYDROGEN ATOM

The prism dispasses the radiation into its various frequencies. Their frequencies appear as line at different positions in the photographic plate. This observation shows that the radiation is emitted at only certain discrete wavelength. Each line in the spectrum is related to a specific change in the energy of the electron

In the late 19th century, a Swiss Mathematician and Physicist John Balmer described atomic emission spectrum of hydrogen using an empirical formular that gave the wavelength of spectrum line observed in H atomic emission spectrum.

$$\frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

where $n_1 = 1, 2, 3, 4, \ldots$

$n_2 > n_1$

R = Empirical proportionality constant called Rydberg constant.

$= 109677.6 cm^{-1}$ from hydrogen

n = particular series of line spectrum.

BOHR'S ATOMIC THEORY

Bohr's atomic model explained the atomic line spectra using Rutherford's idea as well as Planck's and Einstein's idea of quantization of energy.

BOHR'S POSTULATES

ⓐ Steady state postulate —
This states that an atom has stationary non-radiating orbits, shells or energy levels along which electrons move and do not radiate energy irrespective of their acceleration.

ⓑ Orbit quantization rule:- This states that when an atom is in a steady state, an electron travelling in a circular orbit should have quantized value of the angular momentum which comply with the condition, so

$$MVr = \frac{nh}{2\pi}$$

r = distance of orbit from the nucleus
V = velocity
n = position of orbit / energy level
m = mass of electron
h => Planck's constant.

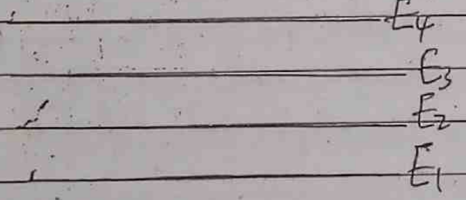
ⓒ Bohr's Frequency Condition —
This states that whenever energy

is absorbed or emitted by an atom, it does so in quanta of amount, $h\nu$, which which is the energy difference between 2 levels or orbits of electrons position

$$E = E_1 - E_2 = h\nu$$

If $E_1 > E_2$, energy is radiated (emitted) by the atom

If $E_1 < E_2$, energy is absorbed by the atom.



$$\nu = \frac{E_1 - E_2}{h}$$

Force of attraction (electrostatic) between the nucleus and the electron is given by Coulomb's law $\frac{Ze^2}{\gamma^2}$ and is opposed by the Centrifugal force $\frac{mv^2}{\gamma}$ which tends to make the electron escape from its orbit

$$\frac{Ze^2}{\gamma^2} = \frac{mv^2}{\gamma} \Rightarrow \gamma = \frac{Ze^2}{mv^2} \quad ①$$

From bohr's second postulate

$$mv\gamma = \frac{nh}{2\pi} \Rightarrow v = \frac{nh}{2\pi m\gamma} \quad (II)$$

$$V = \frac{nh}{2\pi m\gamma}$$

Substitute for V in $\gamma = \frac{Ze^2}{mv^2}$

$$\Rightarrow \gamma = \frac{Ze^2}{m\left(\frac{n^2h^2}{4\pi^2m^2\gamma^2}\right)} \Rightarrow \frac{Ze^2 \times 4\pi^2m^2\gamma^2}{mn^2h^2}$$

$$\cancel{\gamma}^{1} = \frac{4\pi^2 Ze^2 \cancel{m^2}^{m} \cancel{\gamma^2}^{\gamma}}{\cancel{m} n^2 h^2}$$

N.B $e = 4.8\times10^{-10}$ e.s.u

$$\gamma = \frac{n^2h^2}{4\pi^2 Ze^2 m} \quad (III)$$

$h = 6.63\times10^{-27}$

$\Rightarrow 6.63\times10^{-27}$

Since electrostatic force of attraction between the nucleus and the electron (Ze^2/γ^2) and is opposed by the Centrifugal force $\left(\frac{mv^2}{\gamma}\right)$ ~~which~~ tends to make the electron escape from its orbits

$$\Rightarrow E = \frac{1}{2}mv^2 - \frac{Ze^2}{\gamma} \quad (IV)$$

But $\gamma = \frac{Ze^2}{mv^2}$

$$\Rightarrow mv^2 = \frac{Ze^2}{\gamma}$$

$$\Rightarrow E = \frac{1}{2}\left(\frac{Ze^2}{\gamma}\right) - \frac{Ze^2}{\gamma} \text{ from (IV)}$$

$$E = \frac{-Ze^2}{2\gamma}$$

But $\gamma = \frac{n^2h^2}{4\pi^2 Ze^2 m}$ from (III)

the value of γ when $n=1$; $Z=1 \Rightarrow 0.529\,A^\circ$

Substitute for r

$$E = \frac{-Ze^2}{2\left(\frac{n^2h}{4\pi^2 mZe^2}\right)}$$

N:B hcR_H
⇒ $6.625\times10^{-34} \times 3\times10^{-10}$ cm/s
$\times 109677$
$= 2.18\times10^{-18}$ J

$$E = \frac{4\pi^2 mZe^2\,Ze^2}{2n^2h^2}$$

$$E = \frac{-2\pi^2 mZ^2e^4}{n^2h^2}$$

[As, n, Increases E, increases and the shell is farther away from nucleus.]

$$\text{But } E = E_2 - E_1 = h\nu \Longrightarrow$$

ie for when an electron drops from n_2 to n_1

$$h\nu = \frac{-2\pi^2 mZ^2e^4}{n_2^2h^2} - \left(-\frac{2\pi^2 mZ^2e^4}{n_1^2h^2}\right)$$

$$\text{where } E_2 = \frac{-2\pi^2 mZ^2e^4}{n_2^2h^2};\ E_1 = \frac{-2\pi^2 mZ^2e^4}{n_1^2h^2}$$

$$\nu = \frac{1}{h}\times\frac{2\pi^2 mZ^2e^4}{h^2}\left[\frac{1}{n_1^2}-\frac{1}{n_2^2}\right]$$

$$\frac{c}{\lambda} = \frac{2\pi^2 mZ^2e^4}{h^3}\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)$$

$$\frac{1}{\lambda} = \bar{\nu} = \frac{2\pi^2 mZ^2e^4}{ch^3}\left[\frac{1}{n_1^2}-\frac{1}{n_2^2}\right]$$

$$\frac{1}{\lambda} = \bar{\nu} = R_H Z^2\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right).$$

$$\text{where } R_H = \frac{2\pi^2 me^4}{ch^3} = \text{Rydberg Constant}$$

n = principal quantum no.; $h = 6.625\times10^{-34}$ Js
Z = atomic number
$e = \pm1.6\times10^{-19}$ C ; $m = 9.109\times10^{-28}$ g
Since 1 Coulomb = 3×10^9 e.s.u ⇒ $e = 4.8\times10^{-10}$ e.s.u [make use of this as e to get R_H]

N:B 1 Coulomb = 3×10^9 e.s.u
⇒ $1e^- = 4.8\times10^{-10}$ esu [This 4.8×10^{-10} esu is what you must use to get your R_H]

ATOMIC SPECTRUM OF HYDROGEN

White light is a combination of light of many different wavelength. When passed through a prism, white light is spread into its constituent wavelengths, resulting in a band spectrum. A band spectrum resembles a rainbow and contains many different wavelength of light.

When a light from a gas discharge tube is passed through a prism, the result is a line spectrum.

In contrast to a band spectrum, a line spectrum contains only certain discrete wavelength of light.

The wavelength of the lines of the hydrogen spectrum can be calculated from the equation

$$\frac{1}{\lambda} = \bar{\nu} = R_H Z^2\left[\frac{1}{n_1^2}-\frac{1}{n_2^2}\right]$$

R_H = Rydberg's Constant ; Z = atomic no

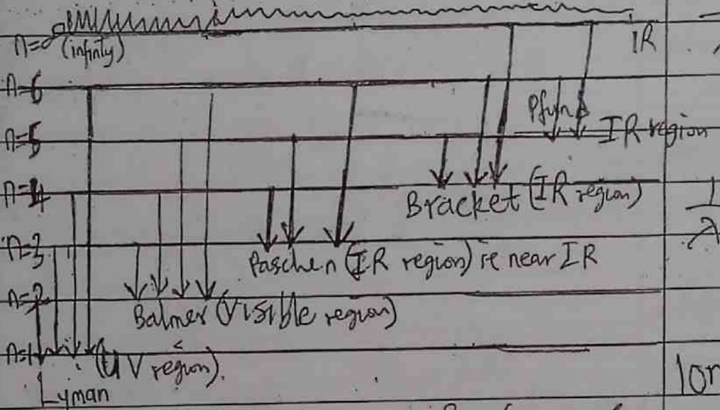
H, Z = 1 ; for He⁺, Z = 2
for Li⁺⁺, Z = 3 ; for Be⁺⁺⁺ = Z = 4
where Z = atomic number.

BOHR'S EXPLANATION of SPECTRUM OF HYDROGEN ATOM

N:B Lyman studied the ground state n = 1
Lyman Series occurs in the UV region.
Electrons at the lowest state have the lowest energy, but require the highest energy to remove it.

N.B $R = 1097 4376.42 m^{-1} = 1.097 \times 10^{5} cm^{-1}$ or $1.097 \times 10^{-4} nm^{-1}$

$\left(2\pi^2 me^4/Ch\right) = 2.18 \times 10^{-18}$



Bohr explanation of Spectrum of Hydrogen atom

There are 5 Series of the atomic Spectra in Various regions of the electromagnetic Spectrum and each is Called after the name of its discoverer.

Spectral Region		Spectral Series	n_1	n_2
UV	increasing order ↓	Lyman	1	2,3,4,5,6,7
Visible		Balmer	2	3,4,5,6,7
IR		Paschal	3	4,5,6,7
IR		Bracket	4	5,6,7
IR		Pfund	5	6,7,8

When the atom is ionized, the electron has only k.E and gives rise to a Continuum and not spectral lines. The energy of transition from ground state, n=1 to ionized state, n=∞ is called the ionization energy.

For hydrogen atom, this Corresponds to the Series limit of the lymans Series where n=1 and n_2=2

$$\frac{1}{\lambda} = 109678\left(\frac{1}{1^2} - \frac{1}{\infty^2}\right) cm^{-1}$$

$z=1$; $n_1=1$, $n_2=\infty$ & $\frac{1}{\infty}=0$

$$\frac{1}{\lambda} = 109678 cm^{-1} \Rightarrow \lambda = 912 A^\circ$$
$$= 9.12 \times 10^{-6} cm$$

Ionization Energy

$$E = h\nu = h\frac{c}{\lambda}$$

$$= \frac{6.63 \times 10^{-27} erg\text{-}s \times 3 \times 10^{10} cm}{9.12 \times 10^{-6} cm}$$

$= 2.18 \times 10^{-11} erg$ ~~or $2.18 \times 10^{-12} J$~~

But

$1 eV = 1.6 \times 10^{-12} erg$

Converting $2.18 \times 10^{-11} erg$ to eV (electron Volt)

$1 eV = 1.6 \times 10^{-12} erg$

$X = 2.18 \times 10^{-11} erg$

$$X = \frac{2.18 \times 10^{-11}}{1.6 \times 10^{-12}} eV$$

$\Rightarrow 13.6 eV$

$\Rightarrow$ Ionization of H-atom = 13.6eV

Bohr's theory is Successful in explaining.

① Stability
② Atomic Spectra
③ Ionization energy of hydrogen and hydrogen like ions having one electron.

① The theory can only explain the atomic spectra of hydrogen and hydrogen like atoms containing only one electron. It could not explain multi electron atoms.

② The spectral line intensities can not be calculated from the theory –

③ It could not explain the 5 spectral into which each of the H_2 atomic spectral lines are resolved using high resolution power spectrograph. This is called Zeeman effect

④ It could not explain how atoms bond to form molecules.

Calculations

① Calculate the wavelength, λ of the 1st spectral line in the Balmer series for He ion

Solution

Balmer, $n_1 = 2$, $n_2 = 2+1$ (ie 1st spectral line)

Helium has 2 electron (atomic no)

$$1/\lambda = Z^2R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$= 2^2 \times 1097.22\left(\frac{1}{2^2} - \frac{1}{3^2}\right)$$

$$\frac{1}{\lambda} = 60{,}956.67 \text{ cm}^{-1}$$

$$\lambda = 1.64 \times 10^{-5} \text{ cm}$$

② What is the wavelength, frequency and energy of Li^{2+} of second lines in visible region?

Solution

Since its Balmer that talk about visible region, $n_1 = 2$, $n_2 = 2+2$ (second line) atomic number $= 3$ (for Lithium)

①

$$\lambda = \frac{1}{RZ^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)}$$

$$= \frac{1}{R(9)\left(\frac{1}{4} - \frac{1}{16}\right)} = \frac{1}{9R(3/16)} = \frac{16}{27R}$$

②

$$\text{frequency}, \nu_1 = \frac{c}{\lambda} = \frac{c}{\left(\frac{16}{27R}\right)} = \frac{27Rc}{16}$$

③

$$E = \frac{hc}{\lambda} \Rightarrow E = \frac{hc}{\left(\frac{16}{27R}\right)} = \frac{27Rhc}{16}$$

③ A hydrogen atom absorbs a photon visible light and its electron enters the n=4 energy level. Calculate ⓐ The change in energy of the atom ⓑ The wavelength in nanometer of the photon

Solution

Visible, $n_1 = 2$, $n_2 = 2+2$ to give 4 (ie second energy level)

Using the formular $E = 2.18 \times 10^{-18}\left(\frac{1}{2^2} - \frac{1}{4^2}\right)$

Note $\frac{1}{\lambda} = R_H Z^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$

If you multiply both sides by hc

$$\Rightarrow \frac{hc}{\lambda} = hcR_H Z^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$E = 2.18 \times 10^{-18} Z^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

Since $\Rightarrow hcR_H$ ie $6.625 \times 10^{-34} \times 3 \times 10^{10} \times 109677.6$

$= 2.18 \times 10^{-18} J$

$$So\ E = 2.18 \times 10^{-18} \left(\frac{1}{4} - \frac{1}{16}\right) = 4.09 \times 10^{-19} J$$

$$E \Rightarrow 4.09 \times 10^{-19} J$$

(ii) $$\lambda = \frac{hc}{E} = \frac{6.625 \times 10^{-34} \times 3 \times 10^{-10}}{\left(\frac{2.18 \times 10^{-18} \times 3}{16}\right)}$$

$$\Rightarrow \frac{16 \times 6.625 \times 10^{-34} \times 3 \times 10^{-10}}{2.18 \times 10^{-18} \times 3}$$

$$= 4.86 \times 10^{-5} m$$

$$= 4.86 \times 10^{-5} \times 10^{9} nm$$

$$= 4.86 \times 10^{4} nm$$

$$\lambda \Rightarrow \underline{\underline{48600 nm}}$$

④ If an electron which is initially in the $n = 5$ level of a hydrogen atom emits a photon of wavelength $4340.4 A^o$ What will be its final energy level?

Solution

$$\frac{1}{\lambda} = R_H \left(\left(\frac{1}{n_1^2}\right) - \left(\frac{1}{n_2^2}\right)\right)$$

for absorption

$$\frac{1}{4340.47 \times 10^{10}} = 1.097 \times 10^{+7} \left(\frac{1}{n_1^2} - \frac{1}{5^2}\right)$$

For emission

$$\frac{1}{4340.47 \times 10^{10}} = -1.097 \times 10^{+7} \left(\frac{1}{5^2} - \frac{1}{n_2^2}\right)$$

Any of these ② Can be used.

$$\frac{1}{4340.47 \times 10^{10}} = 1.097 \times 10^{7} \left[\frac{1}{n_1^2} - \frac{1}{5^2}\right)$$

Dividing both sides by 1.097×10^{7}

$$\frac{\frac{1}{4340.47 \times 10^{-10}}}{1.097 \times 10^{7}} = \frac{1.097 \times 10^{7}}{1.097 \times 10^{7}} \left(\frac{1}{n_1^2} - \frac{1}{5^2}\right)$$

$$= \frac{1}{4340.47 \times 10^{-10}} \div 1.097 \times 10^{7} = \left[\frac{1}{n_1^2} - \frac{1}{5^2}\right]$$

$$= \frac{1}{4340.47 \times 10^{-10}} \times \frac{1}{1.097 \times 10^{7}} = \frac{1}{n_2^2} - \frac{1}{5^2}$$

$$\Rightarrow \frac{1}{4340.47 \times 1.097 \times 10^{-10+7}} = \frac{1}{n_2^2} - \frac{1}{5^2}$$

$$\Rightarrow \frac{1}{4761.5 \times 10^{-3}} = \frac{1}{n_2^2} - \frac{1}{5^2}$$

$$= \frac{1}{4.7615} + \frac{1}{25} = \frac{1}{n_2^2} \Rightarrow \frac{(25 + 4.7615)}{(4.7615 \times 25)} = \frac{1}{n_2^2}$$

$$\frac{29.7615}{119.04} = \frac{1}{n_2^2} \Rightarrow X$$

$$n_2^2 = \frac{119.04}{29.7615} \Rightarrow n_2^2 = 3.9997$$

$$n_2^2 \Rightarrow 3.9997 \Rightarrow n_2 = \sqrt{3.9997}$$

$$n_2 \simeq \underline{\underline{2}}$$

⑤ Calculate the radius of the first allowed Bohr's orbit

Solution

$$r = \frac{n^2 h^2}{4\pi^2 m e^2 z} \quad where\ n = 1,\ z = 1$$

$$r = \frac{1^2 \times 6.63 \times 10^{-27}}{4 \times \pi^2 \times 9.109 \times 10^{-28} (4.8 \times 10^{-10}\ e.s.u)^2}$$

$$= 0.529 \times 10^{-8}\ cm$$

$$\Rightarrow 0.529\ \mathring{A}$$

N:B e = Charge = 1.602×10^{-19} Coulomb

1 Coulomb = 3×10^9 e.s.u

$1.602 \times 10^{-19}\ e = 1.602 \times 10^{-19} \times 3 \times 10^9$

$1e^- \Rightarrow \underline{\underline{4.80 \times 10^{-10}}}$ e.su

WAVE MECHANICS

The most modern theory explaining the structure of an atom is called wave or quantum mechanics and it is based on 3 principles.

① The Dual-nature of matter (de-Broglie 1924)

De Broglie was the first person to suggest that all forms of matter including electrons can have both wave-like and particle properties. He combined Einstein's equation

$$E = MC^2$$

and Planck's equation $E = h\nu = \frac{hc}{\lambda}$ to obtain the equation given below

$$M\cancel{C}^{\overset{C}{2}} = \frac{h\cancel{C}}{\lambda}$$

$$MC = \frac{h}{\lambda} \Rightarrow \lambda = \frac{h}{MC}$$

where M = mass
C = velocity

From this equation, we deduce that matter can be in particle or in mass.

We can also calculate the wavelength of an electron moving with speed, C.

② UNCERTAINTY PRINCIPLE (HEISENBURG) in 1927

This states that "it is impossible to determine precisely the position and momentum of an electron simultaneously. i.e the more accurately we know the position, the more uncertain we are about velocity and vice versa.

This is because as the position is being measured, the velocity must be changing and therefore is uncertain

$$\underset{\substack{\downarrow \\ \text{Momentum}}}{\Delta P} \cdot \times \underset{\substack{\downarrow \\ \text{Position}}}{\Delta x} \geq \frac{h}{2\pi} \neq \underset{\substack{\searrow \\ \text{PSi}}}{\hbar}$$

(wave mechanics)

OR

$\Delta p \times \Delta x \geq \frac{h}{4\pi}$ or $\frac{1}{2}\hbar$; Recall $P = mv$

$\Delta x \cdot \Delta v \geq h/4m\pi$

ONYEUKA BRIGHT CHIMEZIE

③ Schrodinger's Ideas

The energy and position of an electron in an atom is descibed by the equation Involving a wave function, Psi ($\hbar$). The square of the wave function $\hbar^2$ gives the probability of the electrons position from the nucleus.

Schrodinger's equation

$$\frac{-h^2}{8\pi^2 m}\left(\frac{\partial^2\psi}{\partial x^2}+\frac{d^2\psi}{dy^2}+\frac{d^2\psi}{dz^2}\right)=(E-V)\psi_{(x,y,z)}$$

$$\Rightarrow \frac{-8\pi^2 m}{h^2}(E-V)\psi_{(x,y,z)}=\frac{\partial^2\psi}{dx^2}+\frac{d^2\psi}{dy^2}+\frac{d^2\psi}{dz^2}$$

$$\Rightarrow \frac{8\pi^2 m}{h^2}(E-V)\psi_{(x,y,z)}+\frac{\partial^2\psi}{dx^2}+\frac{d^2\psi}{dy^2}+\frac{d^2\psi}{d^2z}=0$$

~~where Ψ = Eigen function or wave function~~

$\pi = 3.142$; m = Mass of an electron

h = planck's Constant; V = potential energy of an electron

x, y, z are axis Coordinates;

E = kinetic energy

Name of Scientist Werner Heisenber. 1925

The Solution of Schrodinger equation defines the orbital which an electron Can occupy in terms of 4 quantum number. This will give the shape, energy and position of the orbital.

QUESTIONS

If a student measures the distance of a 20kg mass as 0.0162m but not 0.02m. find ΔVelocity? ie u.

Solution

Error (uncertainty) $\Rightarrow 0.02 - 0.0162$

in position $\Rightarrow 0.0038$m

It Can also be represented as percentage uncertainty,

Using the formular.

$$\%\text{error} \Rightarrow \frac{\text{error(absolute)}}{\text{relative}} \times 100\%$$

$$\Rightarrow \frac{0.0038}{0.02} \times \frac{100}{1} \Rightarrow 19\%$$

Therefore it is 19% inaccurate (uncertain), but the % Certainty X

$$\%\text{Certainty} \Rightarrow 100 - 19\%$$

$$\Rightarrow 81\%$$

This means that the student was 81% accurate (certain)

Using Heisenberg's formular

$$\Delta x \cdot \Delta V \geq \frac{h}{4m\pi}$$

$$81\% \times \Delta V \geq \frac{6.626 \times 10^{-34}}{4 \times 20 \times \pi}$$

$$\Delta V \geq \frac{6.626 \times 10^{-34}}{4 \times 20 \times 81 \times \pi} \times \frac{100}{1}$$

$$\Delta V \geq 3.254 \times 10^{-36}\%$$

② A particle of mass 1.72×10^{-17} g was accelerated to a speed of 29000 m/s. Calculate its wavelength.

Solution

$V \Rightarrow 29,000$; $m \Rightarrow 1.72 \times 10^{-17}$ g

$\Rightarrow 1.72 \times 10^{-20}$ kg

$$\lambda = \frac{h}{mc} \Rightarrow \frac{6.626 \times 10^{-34}}{1.72 \times 10^{-20} \times 2.9 \times 10^{4}}$$

$$\lambda = \frac{6.626 \times 10^{-34+20-4}}{4.988}$$

$\lambda = 1.328 \times 10^{-18}$ m

$\Rightarrow 1.328 \times 10^{-6}$ pm

QUANTUM NUMBER

Each electron in an atom is described by four different quantum numbers. The first 3 (n, L, m) specify the particular orbital of interest and the fourth (M_s) specifies how many electrons can occupy the orbital.

We have 4 quantum numbers.

① Principal Quantum number:

This is denoted by the letter (n). This refers to the shells or orbits; It describes the energy level of the shells. n can take values of 1, 2, 3, 4 corresponding to quantized energy level K, L, M, N.

② Azimuthal Quantum number, denoted by the letter, L; This can also be called angular momentum number or subsidiary quantum no. It gives the subshell or subshell that are found in a given energy level, n. it also gives us Information about the shape of the shell.

The subshells are quantized and can take (n-1) value starting from 0 such that L < n, ie 0, 1, 2, 3, ... and so on.

The no. of subshells in a given shell is equal to the value of principal quantum no (n) for that shell.

The name of the subshell is given by a small letter namely s, p, d, f and it corresponds to sharp (L=0, shell(1)), principal (L=1, shell(2)), diffuse (L=2, shell(3)) and fundamental (L=3, shell (4)).

Shell	n	No of Subshel	(L) n-1	Name of Subshells
K	1	1	0	1s
L	2	2	0,1	2s, 2p
M	3	3	0,1,2	3s, 3p, 3d
N	4	4	0,1,2,3	4s, 4p, 4d, 4f

Shapes of the Shell:
For s ie L=0 (spherical)
For p ie L=1 (dumb shell)
For d ie L=2 (double dumb shell)
f ie L=3 (complex geometry)

③ Magnetic quantum number: This is denoted by the letter, M_L. This gives the no of orbitals in a given subshell. This is calculated using (2L+1). for each value of L, M takes (2L+1) values, starting from -L to 0 and then to +L.

Each value of M represents one orbital which contains the maximum of 2 electrons.

The maximum no of electrons in a shell is $2n^2$.

$L \geq M_l$

Example		(n-1)	name of subshell	m	(2L+1)	2(2L+1) Total no of electrons
Shell	n	L		(-L, 0, +L)	No of orbitals (degenerate)	
K	1	0	1s	0	1	2
L	2	0	2s	0	1 (4)	2 (8)
		1	2p	-1, 0, +1	3	6
M	3	0	3s	0	1 (9)	2 (18)
		1	3p	-1, 0, +1	3	6
		2	3d	-2, -1, 0, +1, +2	5	10
N	4	0	4s	0	1 (16)	2 (32)
		1	4p	-1, 0, +1	3	6
		2	4d	-2, -1, 0, +1, +2	5	10
		3	4f	-3, -2, -1, 0, 1, 2, 3	7	14

④

ELECTRONIC SPIN QUANTUM NUMBER (M_s): This is denoted by "M_s" which can take the values of $+\frac{1}{2}$ or $-\frac{1}{2}$. These are called the spin states of the electron and are normally represented by the 2 arrows, ↑ (spin up) and ↓ (spin down).
An orbital can't hold more than 2 electrons and these 2 electrons have opposite spins (opposite directions) as stated by Pauli exclusive Principle.

$\boxed{\uparrow\downarrow}$ ie ↑ → Clockwise direction, $+\frac{1}{2}$
↓ → anticlockwise direction
($-\frac{1}{2}$)

N.B (In Summary)

① When $n=1$; $L=0 \Rightarrow S$
its shape is Spherical
when $n=2$; $L=(n-1) \Rightarrow 2-1=1 \Rightarrow P$
its shape is dumbbell.
When $n=3$; $L=(n-1) \Rightarrow 3-1=2 \Rightarrow d$
its shape is double dumbbell.
When $n=4$; $L=(n-1) \Rightarrow 4-1 \Rightarrow 3 \Rightarrow f$
its shape is Complex geometry
For $L=4 \Rightarrow g$ (it shape is Complex geometry)

② Can 4h exist?
NO, this is because, $n=4$
$L=4-1=3$
ie $L = 0, 1, 2, 3$
s p d f
ie no h, in the answer, so it can't exist.

③ For a particular sub energy level, it has different orientation of orbitals. So magnetic quantum number comes in. Magnetic quantum number is restricted to L values.
So
$m = -L, ---, 0, ... +L$
So orbitals of the same energy are degenerate.

If $L = 0$ (S) $\Rightarrow n=1$
$m=0$

If $L = 1$ (P) $\Rightarrow n=2$
$m = -1, 0, +1$

P orbitals has 3 degenerate orbitals

If $l = 2$ (d) $\Rightarrow n = 3$

$m \Rightarrow -2, -1, 0, 1, 2$

∴ d orbital has 5 degenerate orbitals.

④ $$\underbrace{S > P > d > f}_{\text{(Stability)}}$$

$$\underbrace{S < P < d < f}_{\text{(Energy)}}$$

This means that the higher the stability, the lower the energy ie the lower the $(n+l)$ value.

Also Spin quantum number tells us the direction it walks in, either in clockwise or anticlockwise direction.

$$S = +\frac{1}{2} \text{ or } -\frac{1}{2}$$

Spin quantum tells that there are 2 electrons that can occupy an orbital; for every l, there are 2 electrons.

Number of orbitals in an energy level $\Rightarrow n^2$

ie n^2 = Sum of all the orbitals in the Subenergy level

ie if $n=4 \Rightarrow n^2 = (4)^2 = 16$

From the table in the pervious Page, ie table on Magnetic Quantum no, we have that total number of degenerate orbitals

$= 1 + 3 + 5 + 7 = 16$ = number of orbitals in an energy level.

N.B Orbitals with lower $(n+l)$ value comes first before writing orbitals with higher $(n+l)$ while writing the electronic Configuration of an atom.

but if 2 orbitals have same or equal $n+l$ value, the orbital with lower n value must be written first OK!!!.

Maximum number of electrons in a particular energy level

$= 2n^2$

Number of orbitals in a Subshell

$\Rightarrow 2l+1$

number of elections in Subshells

$\Rightarrow 2(2l+1)$.

QUESTIONS

① Calculate the number of orbitals in a subshell when $n = 4$

Solution

$n=4 \Rightarrow l = 4-1 = 3$

no of orbitals in a subshells

$\Rightarrow 2l+1 = (2\times3)+1 = 7$ orbitals

② Calculate no of orbitals in an energy level when $n=4$

Solution

$n=4$

number of orbitals in an energy level

$\Rightarrow n^2 = 4^2 \Rightarrow \underline{16}$

OR

$n=4 \Rightarrow l = 4-1 = 3$

from $m \Rightarrow 0, 1, 2, 3 \Rightarrow l$

when $l=0 \Rightarrow (2l+1) = \underline{1}$

when $l=1 \Rightarrow (2l+1) = 3$

when $l=2 \Rightarrow (2l+1) = 5$

when $l=3 \Rightarrow (2l+1) = 7$

No of orbitals $= 1+3+5+7 = \underline{\underline{16}}$

③ Calculate the element if $n=3$, $l=2$
$m=0$, $S=\frac{1}{2}$

Solution

from $l=2$, it denotes d orbitals.
which have 5 degenerate orbitals

-2	-1	0	1	2
↑	↑	↑		

3electrons

Since $n=3$, $l=2$ $\Rightarrow 3d$
Since $m=0$ and $S=\frac{1}{2}$
it means wewill stop at $\overset{0}{\square}$ and
in clockwise direction ↑

$\Rightarrow 3d^3$

$\therefore 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^3$
$2+2+6+2+6+2+3 = 23$ electrons.
= Vanadium

④ Calculate the element, if $n=3$,
$l=2$, $m=0$, $S=-\frac{1}{2}$

Solution

from $l=2$
it denotes d orbitals which also
have 5 degenerate orbitals

-2	-1	0	1	2
↑↓	↑↓	↑↓	↑	↑

⟹ Total of 8 pms
ie 8 electrons

Since $n=3$, $l=2$ $\Rightarrow 3d$
Since $m=0$ and $S=-\frac{1}{2}$, it means
we will stop at $\overset{0}{\square}$ and in anti
clockwise (opposite) direction

$\Rightarrow 3d^8$

$\Rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^8$

$\Rightarrow$ 28 electrons ie Nickel

⑤ An electron is ejected from the Surface of an aluminum metal irradiated with a U.V light of wavelength 3645A°. Calculate the minimum energy required to Separate the electron from the aluminium metal.

Solution

$$①\ E = h\nu_0 = \frac{hc}{\lambda} = \frac{6.63\times10^{-27}\text{ erg.s}\times 3\times10^{-10}}{3645\times10^{-8}\text{ cm}}$$

$$\text{ie } \frac{6.63\times10^{-27}\text{ erg.s}\times 3\times10^{-10}\text{ cms}^{-1}}{3645\times10^{-8}\text{ cm}}$$

$$\Rightarrow 5.46\times10^{-12}\text{ erg.}$$

⑥ What is the kinetic energy required to Separate the electron ejected from the aluminum Surface irradiation with U.V light of wavelength 475A°

Solution

$$K.E = h\nu - h\nu_0$$

[$h\nu_0$ from question ⑤]

$$h\nu = \frac{hc}{\lambda} = \frac{6.63\times10^{-27}\text{ erg-s.}\times 3\times10^{10}\text{ cm/s}}{475\times10^{-8}\text{ cm}}$$

$$= 4.2\times10^{-11}\text{ erg}$$

$$K.E - 4.2\times10^{-11}\text{ erg} - 5.46\times10^{-12}$$

from question ⑤

$$= 3.654\times10^{-11}\text{ erg.}$$

⑦ An electron is moving with a velocity of 10^{-10} cms^{-1}. Assume that we can measure its position to 0.02A° of a typical radius. Compare the uncertainty in its momentum with the momentum of the electron itself.

Solution

The uncertainty in position, $\Delta x = 0.02 A^0$

The momentum of the electron itself ⇒ mv

$mv \Rightarrow 9 \times 10^{-28} g \times 10^{-10} cm/s$

$\Rightarrow 9 \times 10^{-18} gcms^{-1}$

The uncertainty in the momentum

$\Delta P \times \Delta x = \frac{h}{4\pi} \Rightarrow \Delta P = \frac{h}{4\pi} \div \Delta x$

$= \frac{\frac{6.63 \times 10^{-27}}{4 \times 3.142}}{(0.02 \times 10^{-8})}$

$\Rightarrow \frac{0.5 \times 10^{-27}}{0.02 \times 10^{-8}} \Rightarrow 0.25 \times 10^{-17}$

$\Rightarrow 2.5 \times 10^{-18} gcm s^{-1}$

$\Rightarrow \frac{\cancel{9 \times 10^{-18}}}{\underset{1}{\cancel{2.5 \times 10^{-18}}}} = 3.6$

∴ The uncertainty in the momentum is 3.6 times less than or as less than the momentum itself.

⑧ Calculate the energy of an electron in the first Bohr orbit of hydrogen.

Solution

$$E = \frac{-Ze^2}{2r}$$

where $r = 0.529 \times 10^{-8}$ cm ie 0.529A°

$Z = 1$

$$E = \frac{-(4.8 \times 10^{-12} e.s.u)^2}{2(0.529 \times 10^{-8} cm)}$$

$\Rightarrow -2.178 \times 10^{-11}$ erg.

⑨ Calculate the frequency of light with wavelength 5400A° and the energy in joule that correspond to the wavelength 5400A°

Solution

① frequency $= \frac{c}{\lambda} \Rightarrow \frac{3 \times 10^{10} cms^{-1}}{5400 \times 10^{-8} cm}$

$\Rightarrow 5.56 \times 10^{14} s^{-1}$

⑪ $E = h\nu = 6.63 \times 10^{-27} \times 5.56 \times 10^{14}$

$= 3.686 \times 10^{-12}$ erg

1 erg = 10^{-7} Joules or 1 Joule = 10^{7} erg

Energy ⇒ 3.686×10^{-19} Joules

⑨ 0.4 A° is the uncertainty in determining the position of an electron in an atomic orbital. What is the uncertainty in its velocity

Solution

$\Delta P \cdot \Delta x = \frac{h}{4\pi}$

$h = 6.63 \times 10^{-27}$ erg Sec

$= 6.63 \times 10^{-34}$ Joule sec.

$\Delta p = \frac{h}{4\pi \times \Delta x}$ but $\Delta x = 0.4 A^0$

$= 0.4 \times 10^{-10} m$

$\Rightarrow 4 \times 10^{-11} m$

$\Delta p = \frac{6.63 \times 10^{-34} Js}{4 \times 3.142 \times 0.4 \times 10^{-10} m}$

$\Delta p = 1.32 \times 10^{-24} kgms^{-1}$

$\Delta p = m \times \Delta$ velocity

mass $\Rightarrow 9.11 \times 10^{-31}$ kg

$\Delta V = \frac{\Delta P}{m} = \frac{1.32 \times 10^{-24}}{9.11 \times 10^{-31}}$

$\Delta V = 1448.96 m/s$

⑩ What is the halflife of an element That takes 15 days to decay to 1/4 of its original mass.

Solution

$\ln\left(\frac{N_o}{N}\right) = \frac{0.693t}{T_{1/2}}$ where $N = N_0 - \frac{1}{4}N_0$

$N = \frac{3}{4}N_0$

from $\frac{N_0}{N} = e^{kt} = \frac{N}{N_0} = e^{-kt} \Rightarrow$ where $k = \frac{0.693}{T_{1/2}}$

$\Rightarrow \frac{N_o}{N} = e^{\frac{0.693t}{T_{1/2}}}$ where $N = 3/4 N_0$

$t = 15$ days

$T_{1/2} = ?$

$e = 2.718$

$\frac{\cancel{N_0}}{3/4 \cancel{N_0}} = (2.718)^{\frac{0.693 \times 15}{T_{1/2}}}$

$\Rightarrow \frac{4}{3} = \frac{0.693 \times 15}{T_{1/2}} \ln 2.718$

$T_{1/2} = \frac{3 \times 0.693 \times 15 \ln 2.718}{4}$

$= 7.8$ days

⑪ Calculate the original mass of a radioactive Substance that is 7yrs old if its present mass is 4.5g and its halflife is 1600 yrs.

Solution

$\frac{N_o}{N} = e^{kt}$ where $kt = \frac{0.693}{1600} \times 7$

0.00303

$N_o = N(2.718)^{0.00303}$ $N = 4.5$

$N_o = 4.5(1.003)$

$N_o = 4.514g$ //

⑫ Calculate the number of disintegration per second of 1g of Uranium 232, $t_{1/2} = 5.77 \times 10^{9}$ years.

Solution

$\lambda = \frac{0.693}{t_{1/2}} \Rightarrow \frac{0.693}{(5.77 \times 10^{9} \times 365.25 \times 24 \times 60 \times 60)}$ seconds

$= 3.806 \times 10^{-18} s^{-1}$

$N \Rightarrow 232g$ of U $\Rightarrow 6.02 \times 10^{23}$ particle

1g of U $= \frac{6.02 \times 10^{23}}{232}$

Number of disintegration per Second $\Rightarrow \frac{-dN}{dt} = \lambda \times N$

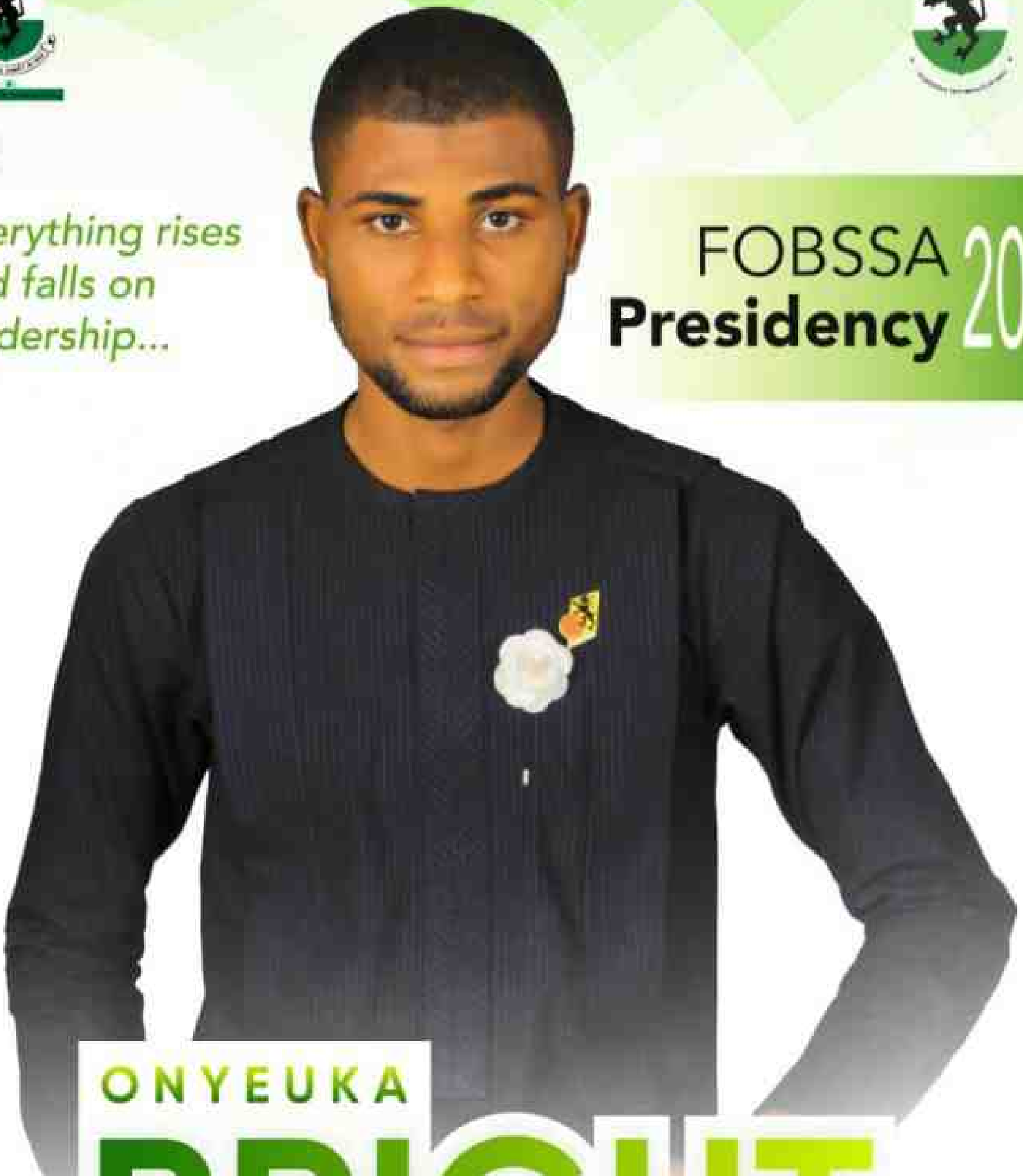
OR

$\frac{-dN}{dt} = \frac{\lambda \times \text{mass} \times \text{Avogadro No}}{\text{molar mass}}$

$\Rightarrow \frac{3.806 \times 10^{-18} \times 6.02 \times 10^{23}}{232} \Rightarrow 9876$ disintegration/secs

Everything rises and falls on leadership...

FOBSSA **Presidency** 2020



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FACULTY OF BIOLOGICAL SCIENCE STUDENTS ASSOCIATION **(FOBSSA)**

*The Man with the **FOBSSA**ites needs at heart...*

NB Higher the mass defect → Higher binding energy → more stability.

Element with longer half life are more dangerous than elements with short half life.

⑬ What are the frequency & wavelength of a photon emitted from a transition from $n=5$ to $n=2$ in the hydrogen atom.

ⓑ Calculate the energy associated with the first orbit of He^+, what is the radius of this orbit?

Solution

(ai) $\bar{V} = 3.29\times10^{15}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) s^{-1}$

from $\frac{C}{\lambda}$ = frequency $(\bar{V})$; $n_1 = 5$, $n_2 = 2$

$\bar{V} = 3.29\times10^{5}\left(\frac{1}{5^2} - \frac{1}{2^2}\right) s^{-1}$

$= 3.29\times10^{15}\left(\frac{1}{25} - \frac{1}{4}\right)$

$= 3.29\times10^{15}\left(\frac{4-25}{100}\right)$

$\Rightarrow 3.29\times10^{15}\left(\frac{-21}{100}\right)$

$\Rightarrow -6.909\times10^{14} s^{-1}$

(ii) $\lambda = 1/\bar{V}$

$\bar{V} = 1.097\times10^{7}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) m^{-1}$

$\lambda = \frac{1}{1.097\times10^{7}\left(\frac{1}{5^2} - \frac{1}{2^2}\right)}$

$\lambda = \frac{1}{1.097\times10^{7}\left(\frac{4-25}{100}\right)}$

$= \frac{1}{1.097\times10^{7}\left(\frac{-21}{100}\right)} \Rightarrow \frac{1}{\frac{-23.037\times10^{7}}{100}}$

$\lambda = \frac{100}{-23.037\times10^{7}} \Rightarrow \lambda = -43.4\times10^{-8} m$

$\lambda = -434 nm$

ⓑ $E = \frac{-2.18\times10^{-18} Z^2}{n^2}$

, $Z = 2$ (atomic no of He^+

$n = 1$

$E = \frac{-2.18\times10^{-18}(2)^2}{1^2}$

$E = -8.72\times10^{-18} J$

(ii) E_n = Radius of orbit

$E_n = \frac{0.529 A^{\circ} n^2}{Z}$ (m)

$= \frac{0.529\times10^{-10}\times1^2}{2}$

$\Rightarrow \frac{5.29\times10^{-11}}{2}$

$\Rightarrow 2.645\times10^{-11} m$ or $0.02645\times10^{-9} m$

$\Rightarrow 0.02645 nm$

(14) N:B Uncertainty Principles gives a Significant Value when applied to a Very small objects or a microscopic object (small mass)·

Uncertainty Principles gives a negligible Value when applied to a macroscopic object (large mass) =

Examples (a) $m = 10^{-6}$ kg

(b) $m = 9.11 \times 10^{-31}$ kg

Solution

(a) $\Delta x \cdot \Delta V x = \frac{h}{4\pi m}$

$m = 10^{-6}$ kg

$\Delta x \Delta V = \frac{6.626 \times 10^{34}}{4 \times 3.142 \times 10^{-6}}$

$\Rightarrow \frac{6.626 \times 10^{-34}}{12.571 \times 10^{-6}}$

$\Rightarrow 0.527 \times 10^{-28}$

$\Rightarrow 5.27 \times 10^{-29} m^2 s^{-1}$

Here we notice that a negligible Value was gotten when we used a larger or macroscopic Object·

(b) $\Delta x \cdot \Delta V = \frac{h}{4\pi m}$

$m = 9.11 \times 10^{-31}$ kg

$\Delta x \cdot \Delta V = \frac{6.626 \times 10^{-34}}{4 \times 3.142 \times 9 \times 11 \times 10^{-31}}$

$\Delta x \cdot \Delta V = \frac{6.626 \times 10^{-34}}{114.49 \times 10^{-31}}$

$\Rightarrow 0.0579 \times 10^{-3}$

$\Rightarrow 5.79 \times 10^{-5} m^2 s^{-1}$

Here, we notice that a more Significant Value was gotten when a Smaller object or microscopic object

(15) A microscope using suitable photons is employed to locate an electron when a distance of 0.1 Å m, What is the uncertainty Involved in the measurement of its Velocity

$m = 9.11 \times 10^{-31}$ kg.

Solution

① $\Delta x \Delta V = \frac{h}{4\pi m}$

multiply both Sides by $\frac{1}{\Delta x}$ to make ΔV = Subject of formula

$\frac{1}{\cancel{\Delta x}} \times \cancel{\Delta x} \cdot \Delta V = \frac{h}{4\pi m \Delta x}$

$\Delta V = \frac{h}{4\pi m \Delta x}$

$\pi = 3.142$

$m = 9.11 \times 10^{-31}$ kg

$\Delta x = 0.1 \overset{\circ}{A}_m \Rightarrow 0.1 \times 10^{-10}$ m

$\Delta V = \frac{6.626 \times 10^{-34}}{4 \times 3.142 \times 9.11 \times 10^{-31} \times 0.1 \times 10^{-10}}$

$\Delta V = \frac{6.626 \times 10^{-34}}{11.449 \times 10^{-41}}$

$\Delta V = 0.579 \times 10^{-34+41}$

$\Delta V = 0.579 \times 10^{7}$

$\Delta V = 5.79 \times 10^{6}$ m/s

(16) A golf ball has a mass of 40g and speed of 45m/s if the speed can be measured within accuracy of 20%, Calculate the uncertainty in position.

Solution

① $\Delta x . \Delta V = \frac{h}{4\pi m}$

$\Rightarrow \Delta V = \frac{h}{4\pi m \Delta x}$ $\pi = 3.142$

$m = \frac{40}{1000} = 0.04kg$

$\Delta V = \frac{2}{100} \times 45 \Rightarrow 0.9 m/s$

$\Delta x . \Delta V = \frac{h}{4\pi m} \Rightarrow \Delta x = \frac{h}{4\pi m \Delta V}$

$= \frac{6.626 \times 10^{-34}}{4 \times 3.142 \times 0.04 \times 0.9}$

$= 14.65 \times 10^{-34} m$

$\Rightarrow 1.465 \times 10^{-34} m$

THE ELECTRONIC THEORY OF VALENCY:

Dalton first put foward the idea in his atomic theory that Chemical Compounds are formed by the Union of atoms of element and since then. Chemists have been Concerned with the way in which atoms Combine and the Capacity possessed by any one atom for doing so.

Berzalius later proposed that all Compounds were electrical in nature Consisting of 2 parts, one positive and the other negative, the 2 being held together by electrostatic attraction. All chemical bonds are electrostatic in nature, bonds (linkages) between atoms resulting when the net attractive and repulsive forces is one of attraction.

In the electronic theory of Valency, Chemical Combination (Chemical bonding) is mainly accounted for in terms of the sharing or transferring of electrons b/w the atoms.

Also, In the electronic theory of Valency, It is observed that the Valence number (Valency) of an element is related to the number of electrons in the outermost shell of its atom.

N.B An atom transfer or Share electrons with another atom(s) in other to be stable by achieving either a duplet or octet structure.

The 2 electrons (duplet structure) is called the duet

The electrons which are involved in bond formation b/w atoms is the electrons in the outer most shell.

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BONDS
(CHEMICAL BONDING)

According to the electronic theory, there are different ways in which atoms can combine to give rise to various types of chemical bonds which canbe distinguished by the characteristics properties they confer upon the molecule containing them.

TYPES OF CHEMICAL BONDING

They are ⓐ Ionic/electrovalent bonding

ⓑ Covalent bonding which is further divided into ordinary covalent bonding and Dative/Coordinate bonding,

ⓒ Metallic bonding.

Other bonding are hydrogen bonding and Vander waals forces.

Some compounds are clearly ionic while some are clearly covalent, but many others possess both ionic and covalent properties.

N:B. Hydrogen bonding canbe
ⓐ Intermolecular H-bond
ⓑ Intramolecular H-bond,

We also have ① dipole-dipole interaction
(ii) Dispersion forces/London forces
(iii) Dipole-induced dipole.

N:B Kossel formulated the theory of ionic bond in 1916. He proposed that atoms of elements tends to acquire either duplet eg He or octet eg Ne/Ar.

IONIC BONDING

In this type of bondings, an atom transfers one or more of its Valence electrons ie (outermost electrons/electrons in the outermost shell) to the atoms with which it combines.

When this happens, the atom which loses electrons becomes positively charged, (cation) because it now contains fewer electrons than protons and the atom which accepts the donated electrons becomes negatively charged, (anion) because it has an excess of electrons over protons.

The cation or anion derived from a single atom is called a monoatomic ion. The bond between positively and negatively charged ions which is electrostatic bond is created by the transfer of one or more valence electrons from one atom to the other. This is the reason why it is called ionic bond or electrostatic bond and its force is called electrostatic force of attraction of the charged ions.

N:B This type of bonding takes place between highly electropositive element and highly electronegative element.

A good example of an ionic bonding is the reaction of a Sodium atom and a Chlorine atom to produce

Sodium Chloride

$$Na^+ + Cl^- \longrightarrow NaCl$$

N.B The following processes occur when Sodium and Chlorine combine together.

Firstly, Sodium loses its one electron in the outer most shell to become Isoelectronic with Neon (ie to have 10 electron like Neon) a very stable noble gas atom. This shows that Sodium atom donates an electron to form Sodium ion. Na^+.

That is $Na \xrightarrow{\text{lost electron}} Na^+ + e^-$

$1s^2 2s^2 2p^6 3s^1$ $\quad$ $1s^2 2s^2 2p^6$

$(2,8,1)$ $\quad$ $(2,8)$

Also, Chlorine gains one more electron to complete the octet structure (ie it has 8 outermost electrons) and be Iso-electronic with argon, a stable noble gas. This shows that Chlorine is a good electron acceptor forming a Chloride ion.

$$:\ddot{Cl}\cdot + e^- \xrightarrow{\text{gain electron}} :\underset{..}{\ddot{Cl}}\cdot^{\,-}$$

$1s^2 2s^2 2p^6 3s^2 3p^5$ $\quad$ $1s^2 2s^2 2p^6 3s^2 3p^6$

$(2,8,7)$ $\quad$ $(2,8,8)$

The electron released by Sodium (electron donor) is the electron received by Chlorine electron acceptor.

$$Na \longrightarrow Na^+ + e^-$$

$$Cl + e^- \longrightarrow Cl^-$$

The opposite charged ions Na^+ and Cl^- are held together by electrostatic force of ion pair; Na^+Cl^- $(NaCl)$.

Also as seen in MgO

ie $Mg \xrightarrow{\text{loses electron}} Mg^{2+} + 2e^-$

$1s^2 2s^2 2p^6 3s^2$ $\quad$ $1s^2 2s^2 2p^6$

$(2,8,2)$ $\quad$ $(2,8)$

$$:\underset{..}{O}: + 2e^- \xrightarrow{\text{gains electron}} O^{2-}$$

$1s^2 2s^2 2p^4$ $\quad$ $1s^2 2s^2 2p^6$

$(2,6)$ $\quad$ $(2,8)$

Magnesium atom donates 2 electrons to complete the octet structure and becomes Isoelectronic with neon, a noble gas and Oxygen accepts the 2 electrons to complete octet structure and also be Isoelectronic with neon.

The Solid oxide consists of aggregate of Mg^{2+} and O^{2-} ions, all being Isoelectronic with neon $1s^2 2s^2 2p^6$. These ions are not exactly atom of neon since their nuclei still contain either ($\underset{\text{Protons}}{12}$ for Mg)

and 8 for Oxygen protons whereas any atom of neon has 10 protons in its nucleus.

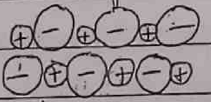
So $Mg^{2+} + O^{2} \rightarrow Mg.O$

Magnesium Oxide

Other examples are $MgCl_2$, $CaCl_2$ etc
The Information above shows 2 Important facts about atomic number

ⓐ It characterizes an element but not the number of electrons surrounding the nucleus at any particular time.

ⓑ It remains unchanged during chemical combination and is not affected by the transfer of electrons from one atom to another.



N.B ① Atoms of elements with low ionization energy and low electron affinity form positive ions.

(ii) Atoms of elements with high ionization energy and high electron affinity form negative ions.

(iii) The positive and negative ions are held together by electrostatic forces between ions of opposite charge to form ionic bond.

(iv) Reactions between representative metals and non-metals result in ionic bond.

But

Ionic Compound is formed during the formation of electrovalent bond and is favored by:
① Low ionization energy of the metal
② High electron affinity of the non-metal.
③ High lattice energy of the resulting Compound (the product).

It is not true that elements react in order to gain noble gas electronic configuration, rather they react because the compounds formed are more energetically stable than their elements.

PROPERTIES OF IONIC COMPOUNDS

① They have high melting and boiling point due to strong electrostatic force of attraction between the ions

② They are crystalline, strong, hard and found in the solid state at room temperature. They do not vaporize easily.

③ Because they dissociate to form ions, in water, they are good conductor of electricity. Hence, they conduct electricity in solutions, molten or fused state because of mobile ions.

④ Soluble in polar solvent due to the ion-dipole attraction. They dissociate to form ion in water, thus they do not contain molecules. When ionic compounds is placed in water, the force between the ions in the crystal weakens.

⑤ They occur mainly in inorganic compounds ie they do not dissolve in organic solvents such as toluene, ether, phenol, methane, Chloroform and benzene.

COVALENT BONDS

This is the overlap of atomic orbital to form molecular orbital. It involves the sharing of valence electrons of the non-metallic atoms of an element. A chemical bond through sharing of valence electrons is called covalent bond. Examples of compounds formed through covalent bond are H_2, F_2, CCl_4, CO_2, CO, HF, H_2O and HCl.

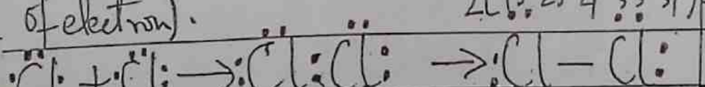
Here, two atoms share jointly one, two or three (or more) pairs of electrons giving rise to −, = and ≡ bonds respectively. The atoms of most molecules are firmly held by covalent bond. A typical covalent bond might be homopolar (the bond formed in diatomic molecules having only one type of atoms as in H_2, F_2, O_2, N_2 and Cl_2) or heteropolar (which is formed when the atoms involved in the bond formation are different. The electron density reside closer to the atoms with greater electron affinity eg HCl, HF. This the bond is polarized.

Examples of Homopolar Covalent bond

$H^{\cdot} + {}^{\cdot}H \rightarrow H{:}H$ ie $H-H$ (H_2)

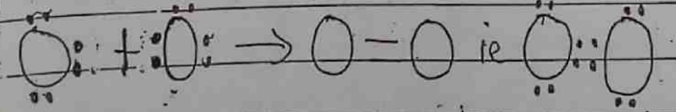
single homopolar covalent bond (1 shared pair of electron).

$2Cl(1s^2 2s^2 2p^6 3s^2 3p^5)$



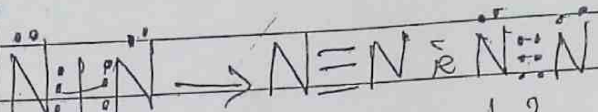
$:\underset{..}{\ddot{Cl}}\cdot + \cdot\underset{..}{\ddot{Cl}}: \rightarrow :\underset{..}{\ddot{Cl}}:\underset{..}{\ddot{Cl}}: \rightarrow :\underset{..}{\ddot{Cl}}-\underset{..}{\ddot{Cl}}:$

(1 shared pair of electron & 6 unshared pairs of electron)



$\underset{..}{\ddot{O}}: + :\underset{..}{\ddot{O}}: \rightarrow O=O$ ie $\underset{..}{\ddot{O}}::\underset{..}{\ddot{O}}$

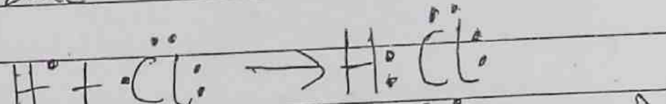
(2 shared pair of electrons & 4 unshared pair of electrons)



$\dot{\ddot{N}}\vdots + \vdots\dot{\ddot{N}} \rightarrow N\equiv N$ ie $\ddot{N}\vdots\vdots\ddot{N}$

3 shared pairs of electrons and 2 unshared pairs of electrons.

Examples of heteropolar Covalent bond.



$H^{\cdot} + \cdot\underset{..}{\ddot{Cl}}: \rightarrow H{:}\underset{..}{\ddot{Cl}}:$

(one shared pair of electrons and 3 unshared pairs of electrons)

The type of bond formed depends on temperature, pressure and medium. For example, $HCl_{(g)}$ at S.T.P is near covalent bond but on heating, electrons returns to H then the bond factually become covalent. In aqueous solution, HCl becomes acid forming H^+ (H_2O).

Transfer of electrons can only lead to electrovalent bond but the sharing of electrons leads to covalent or cordinate covalent bond. In the above examples involving Cl_2, only a single electron is required by such atom to complete its valence shell (octet structure). The covalent molecules formed from atoms requiring more than electron to complete the valence shell eg in case of tetrachloromethane (CCl_4) are also part of our discuss.

Other examples of covalent bond are CH_4, NH_3, CO_2, H_2O, CO, C_2H_2 etc.

PROPERTIES OF COVALENT COMPOUND)

① Covalent Compounds Consist of molecules which do not form ion in Solution -

② They do not Conduct electricity and so they Conduct electricity in Solution only if they react with Solvent molecule to form ions. Eg $HCl \rightarrow H^+ + Cl^-$. Graphite as well Conducts electricity due to its structure.

③ They occur mainly in organic Compounds So they are Soluble in organic Solvent like benzene which dissolve Camphor or Carbon disulphide which dissolves Sulphur.

④ They have low melting and boiling points as little energy are needed to break the weak intermolecular forces. This is why they are gases or Volatile liquid or Solid.

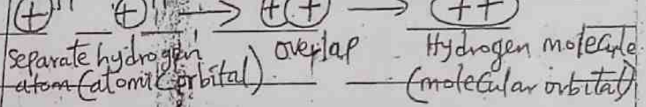
⑤ Some Covalent Compounds Such as diamond, graphite and Silicon (iv) Oxide have high melting and boiling point. This is because they form great/grant Crystalline Structure.

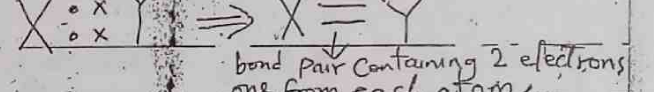
⑥ They are Soluble in non-polar Solvent like CCl_4, but Sugar, methyl alcohol, glycerol etc dissolves in H_2O and other Polar Solvents because of H-bonding.

N:B ① High charge and Small Size of Group 1, 2 $ 3 will form Covalent bond especially Li, Be and B. Also electronic theory of Valency, states that every atom shows a tendency in chemical Combination to adjust the number of electrons in its outermost shell So that it can adopt the stable electronic Configuration of one of the noble gases-

N:B The 2 main theories of Chemical bonding are
ⓐ Valence bond theory and
ⓑ molecular orbital theory.

Examples of molecular orbital eg





Also All Group, 1 and 2 elements are usually bound to form ionic Compounds with electronegative element except Li and Be. This is because they are the first Series which forms Covalent Compound because of their Small Size. Examples LiF, LiCl, $BeCl_2$

Also in Group 3, the first Series ie B, Al usually form Covalent Compounds instead of purely ionic Compounds. BCl_3, $AlCl_3$ are examples.

Note also that the 2 things that affect the electrostatic forces of ionic Compound are Atomic Size and Nuclear charge. That is, Ionic Compounds are favoured by higher nuclear charge and Small atomic Size.

Also $CaCO_3$ have higher boiling point than Na_2CO_3 meaning that Na_2CO_3 is more metallic.

Electronegativity difference determines the types of bond formed ie High electronegative difference of >0.9 or >0.8 = ionic

Low electronegative difference of <0.9 or <0.8 = Covalent

Electronegative difference of 0.9 or 0.8 = 50% ionic and 50% Covalent

Electronegativity Values Can be used to Calculate bond energy b/w atoms of element using the formular

$$D_{A-B} = \frac{1}{2}(D_{A-A} + D_{B-B}) + 23.06(X_A - X_B)^2$$

Provided that X_A and X_B are not largely different.

where D_{A-B} = Single bond energy b/w atoms A and B in kCal/mol

D_{A-A} = Single bond energy b/w 2 atoms of A in kCal/mol

D = Single bond energy b/w 2 atoms of B in kcal/mol

X_A = electronegativity Value of A

X_B = " " " " B

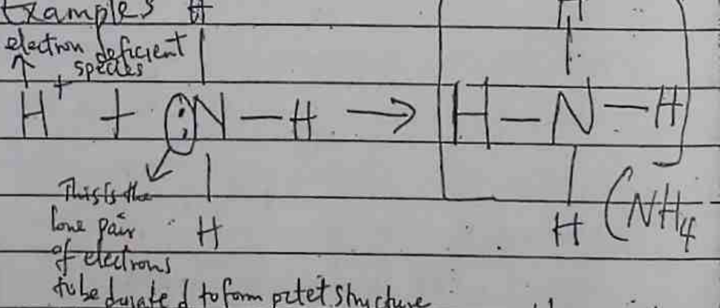
COORDINATE COVALENT BOND (DATIVE BOND)

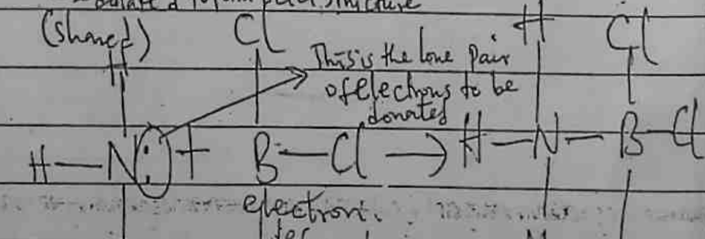
This is a bonding in which electrons are donated by only one of the elements. The element has electrons in excess. The element is called nucleophile and its no more electron loving but nucleus loving. If its electron loving it will be electrophile.

This means that electron deficient Species eg H^+, BCl_3, $AlCl_3$ etc react with electron rich species like NH_3, where the 2 electron pair been shared is donated by one of the participating atom giving rise to dative bond or Co-ordinate Covalent bond. One atom contributes in the 2 Valence electrons to be Shared in Contrast to the previous bonding Cases where there is equal Contribution.

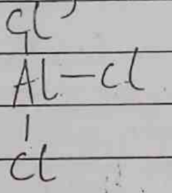
In this Case, one atom will have non-bonding electron or lone pair of electron eg NH_4Cl, Al_2Cl_2, CO, NH_3BCl_3 etc.

Examples



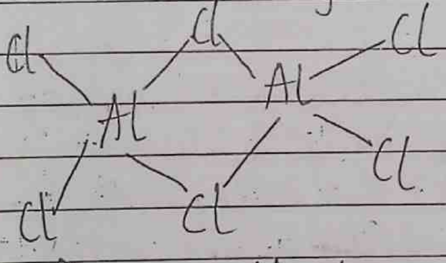


Dimenises of Aluminium Chlorides means that 2 of that Compounds must Combine to form a Compound that obeyed octet rule by Vacant orbital, eg as seen in $AlCl_3$



You can see that Aluminium have only 6 electrons (Sixtet) hence do not obey octet rule

For Aluminium to obey octet rule it has to dimenise this way -



Dimenised Al_2Cl_6.

N.B In term of atomic orbital, Dative bonding are (occurs) when empty atomic orbital are Combined with atomic orbital Containing lone pair.

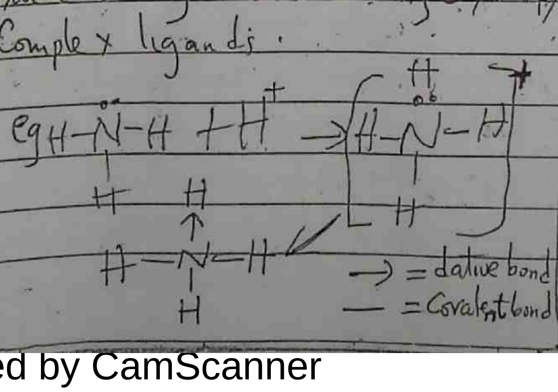
$$X + :Y \Rightarrow X:Y$$

$$\Rightarrow X \overset{\downarrow}{\leftarrow} Y$$

bond pair Containing 2 electrons both from one atoms

Dative bonding occurs in H_3O^+, NH_4^+, Complex ligands.

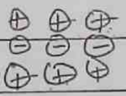
eg



→ = dative bond
— = Covalent bond

METALLIC BONDING

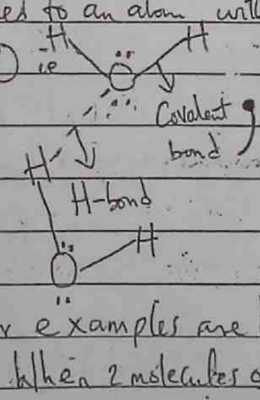
This is bonding between metallic ions and delocalised electrons. In this bonding there is a Strong electrostatic force in all directions.

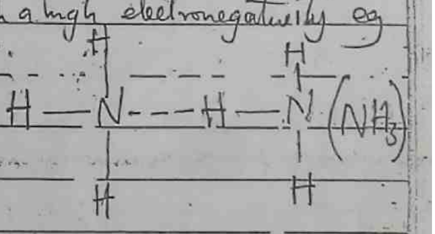


it Increases with Charge, increases with no of delocalised electrons and decreases with Size of ions.

Hydrogen bond is strong Intermolecular force of attraction between hydrogen atom attached to an electronegative atom (by Covalent bond) and the lone pair of another electronegative atom.

Water with molecular weight of 18 has boiling point of 100°C and high Surface tension which Could not be accounted for by bond types like ionic, Covalent, dative etc. Hydrogen bond which is formed between a hydrogen atom and a highly electronegative atom such as F, O and N accounted for this. The hydrogen atom must be bonded to an atom with a high electronegativity eg H_2O ie

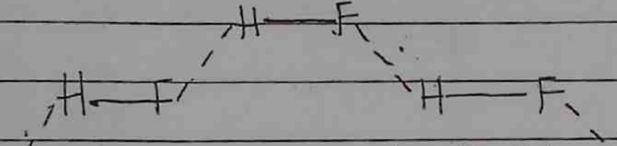




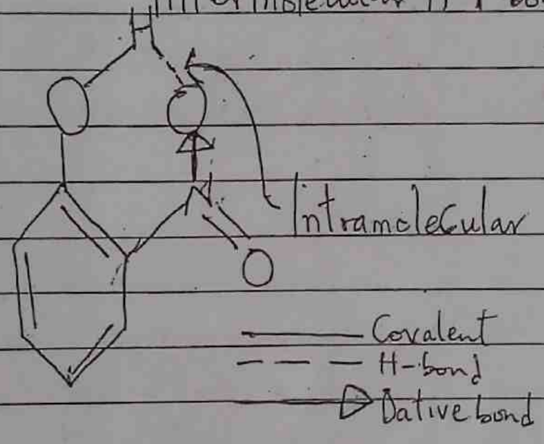
Other examples are HF, HCN etc

When 2 molecules of HF put close to one another, the hydrogen of one molecule will be attracted to the F atom of the other. This happen's because of the δ+ of the H and δ− of the F atom.

The attraction that holds the hydrogen atom and F atom of other molecule is called hydrogen bond. The more electronegative the atom is the stronger the H-bond. H-bond that occurs between 2 different molecules is called intermolecular Hydrogen bonds. While hydrogen bond that are form within the same molecule are called intra-molecular bond. These are shown below



Intermolecular H-F bond



Due to proximity of the O-H to oxygen of NO_2 intramolecular H-bond occurred

CONSEQUENCES/EFFECTS OF H-BOND

① it increases viscosity and dielectric constant of the liquid
② it increases melting point and boiling point and heat of vaporization and fusion higher than expected.
③ it increases hardness and rigidity of the crystalline solid.
④ it increases stability of molecules
⑤ it enhances miscibility eg ethanol mixes with water because of H-bonding

VANDER WAAL FORCES

The Vander Waals forces named after Dutch Scientist Johannes Didenik Vander Waals. This is the sum of the attractive or repulsive forces between molecules other than those due to covalent bond, H-bond or electrostatic interaction of ions with one another or with neutral molecules.

They are generally intermolecular forces of attraction. The forces that hold helium atoms or covalent substances like iodine together in a liquid or solid are called intermolecular force of attraction also known as Vander Waals forces. They occur as a result of:

① Forces between 2 permanent dipoles (dipole-dipole interaction also known as Keesom force) that result from the separation of negative and positive charges producing a dipole eg in HCl.

② Forces between a permanent dipole and a corresponding induced dipole (temporal dipole interaction also known as Debye force).

③ forces between 2 instantaneously induced dipoles (London or dispersion force). These 2 dipoles will briefly attract each other. eg CCl_4, C_4H_{10}.

N.B The force of attraction between 2 temporal dipole is known as

London forces or dispersion forces.

Thus, Intermolecular forces between non-polar molecules are referred to as London or dispersion forces eg as seen in Cl_2.

VanderWaals forces are relatively weak compared to Covalent bond but plays a fundamental role in life. It is a weak attractive force between atoms of non-polar molecules caused by a temporary change in dipole moment arising from a brief shift of orbital electrons to one side of one atom or molecule creating a similar shift in adjacent atoms or molecules. They arise from the random motion of electrons in an atom or molecule. It provides low binding energy leading to molecular crystals having low melting point and boiling point. This is responsible for Helium being liquid and Iodine solid at room temperature.

N:B London forces or dispersion force is one type of intermolecular force named after a German physicist Fritz London who first calculated it. It is the force of attraction b/w 2 temporary dipoles.
ie temporary → temporary induced dipole attraction of one another.
The overall strength of london forces depends on molecular shapes and size.

N:B Intermolecular forces affect physical properties

Intramolecular forces affect Chemical properties.

Intramolecular bond

Metallic bond > Ionic bond > ordinary Covalent > Dative bond

Intermolecular bond

Hydrogen Bond > Dipole dipole (Permanent bond) > Dispersion forces (London forces)

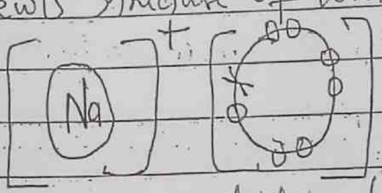
Hydrogen bond is the strongest intermolecular forces.

N:B: The higher the electronegativity of the atom bonded by hydrogen the greater the strength of the hydrogen bond. This is called Dipole–dipole attraction.
This bond is weaker than that of Covalent bond and electrovalent bond. This bond strength can affect the boiling point, Vapour pressure, and heat of Vapourisation of H bonded polar compounds as seen in NH_3 etc.

LEWIS STRUCTURE OF BONDED ATOMS / ELECTRON DISTRIBUTION.

ⓐ Lewis Structure of ionic Compound

eg



where x is the electron transfered from Na
o is the electron in Cl

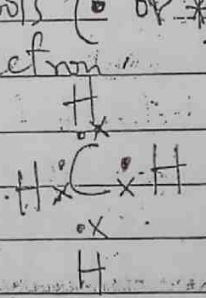
N:B The "+" & "−" shows an electrostatic force of attraction

Na_2O has higher melting point than NaCl because they are both ionic Compounds and because of higher nuclear charge and smaller size of O^{2-} than Cl^-. In other words Na_2O has higher melting point than NaCl.

ⓑ COVALENT BONDS

The Lewis structure of a Covalent Compound or polyatomic ion shows how the Valence electrons are arranged among the atoms in the molecule to show the Connectivity of the atoms. It makes use of symbols (• or *) to represent one electron.

eg CH_4 ⇒



Lewis Structure of Methane

When we have A:B denoted as A—B it means that one pair of electron is shared.

When we have A::B denoted as A=B, it means 2 pairs of electrons is shared.

When we have A:::B denoted as A≡B, it means 3 pairs of electrons is shared.

When we have A: , the 2 electrons are called the unshared pair of electron or lone pair.

Lone pairs do not contribute directly to the bonding but they influence the shape of the molecules and play an important role in its properties.

The octet rule (which states that each atoms shares electrons with neighbouring atoms to achieve a total of 8 electrons in the Valence shell (octet) provides a simple way of constructing a Lewis structure.

Steps in determining Lewis Structure

① Determine the no of electrons that are to be included in the structure by adding all the number of Valence electrons of the atoms in the molecules. eg

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0_2:- Each negative charge on an ion orresponding to an additional electrons nd each positive charge on an ion Corresponds one electron less.

rite the Chemical Symbols of the atoms arrangement that shows which atoms be bonded together, eg O — C — O ie e less electronegative is usually the Central om of a molecule, but there are many ell known exceptions (eg H_2O and NH_3)

istribute the electrons in pair So that here is one pair of atoms bonded to-ther and then Supply electron pairs (to rm lone pairs or multiple bond) until ch atom has an octet.

xamples

Draw the Lewis Structure of CH_3OH

Solution

Get the total Valence electrons

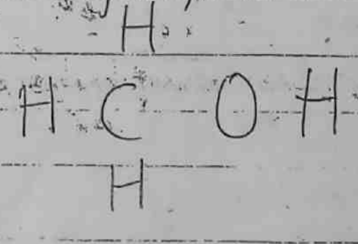
$1C \times 4e^- \Rightarrow 4$

$4H \times 1e^- = 4$

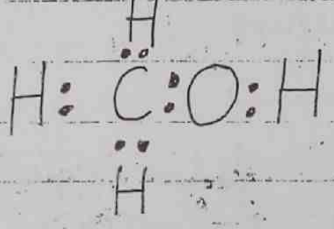
$1(O) \times 6e^- = \underline{6}$

$\underline{14e^-}$

e C is less electronegative, it will be he Centre

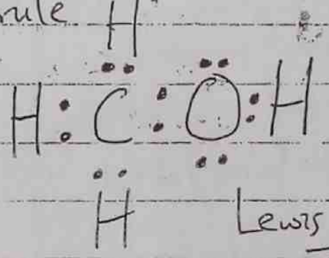


③ Assign a pair of electron in between the atoms



ie 10 electrons assigned

Then assign the remaining 4 electron to the Oxygen So that it will attain to octet rule



⇒ 2 lone pairs ie 4 lone electrons and 5 bond pairs ie 10 shared electrons

Lewis Structure of CH_3OH

② Draw the Lewis Structure of SF_4 and SF_6

Solution

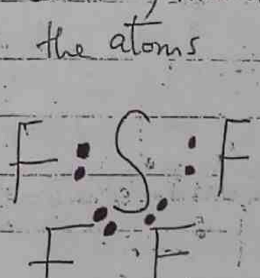
Get the total Valence electrons in SF_4

That is, $1S \times 6e^- \Rightarrow 6$

$4F \times 7e^- \Rightarrow \underline{28}$

$\underline{34e^-}$

Since S is less electronegative, it will be at the Centre, then a pair of electron in between the atoms



We have assigned 8 electrons remaining $26e^-$

Assign the remaining 26 electrons to the Flourine atoms and Sulfur atom to attain octet structure

We have assigned 24 electrons to the Fluorine atoms and they have attained octet structure, but it is remaining 2 electrons, so we assign it to the Sulfur atom as a lone pair.

Finally We have One lone pair on the Sulphur atom, that is 2 lone electrons. On the Sulfur atoms.

Also we have total of 26 lone electrons on both Sulfur and Flourine ie 13 lone pairs. We also have 3 bond pairs ie 6 shared electrons (b/w the Sulphur and Fluorine.

(b) Structure of SF_6

ⓐ Total Valence electron (ie total electrons in outermost shells) $\Rightarrow$ 1 S atom $\times 6e^- = 6$

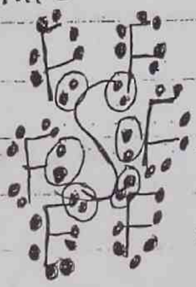
6 F atoms $\times 7e^- = 42$

$48e^-$

Since Sulfur is less electronegative it will be at the Centre then assign a pair of electron in between the atoms.

We have assigned $12e^-$ remaining $(48-12) = 36e^-$

Then assign $36e^-$ remaining to the Fluorine atoms as lone electrons to attain octet rule



Total electrons assigned $= 48e^-$

$\Rightarrow 6$ ⊙ (ie 4 bond pairs of e^-)

$\Rightarrow 12$ Shared electrons

Also we have 36 lone electrons.

③ Draw Lewis Structure of $BrCl_3$

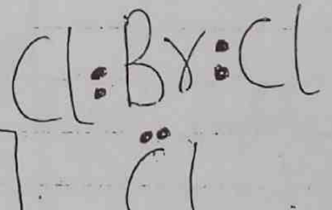
Get the total Valence electrons

1 Br $\times 7e^- = 7e^-$

3 Cl $\times 7e^- = 21e^-$

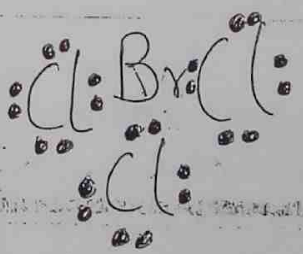
$28e^-$

Since Br is less electronegative, it will be the Central atom, then assign a pair of electron in between the atoms,



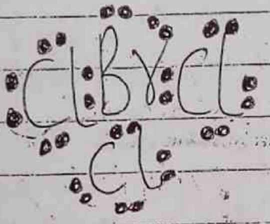
We have assigned 6 electrons remaining $22e^-$

Then assign the remaining $22e^-$ to the Chlorine atoms as lone electrons



We have assigned 18 lone electrons to chlorine atoms remaining 4 electrons

Finally assign the remaining 4 electrons as lone electrons to Bromines:



We have
2 lone pairs on Bγ
6 Shared electrons (ie 3 bond pairs)
22 lone electrons on Bγ and CL

④ Draw Lewis Structure of IO_3

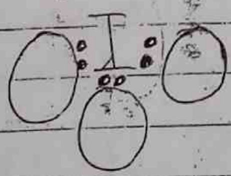
Get the total Valence electrons

$1I \times 7e^- = 7e^-$

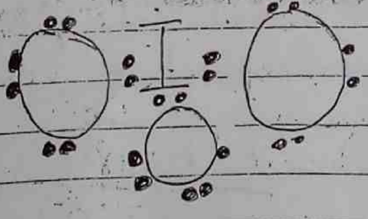
$3O \times 6e^- = 18e^-$

$\underline{25e^-}$

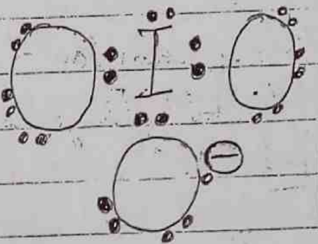
Since I is less electronegative, it will be the Central atom, then assign a pair of electron in between the atoms.



We have assigned 6 electrons, remaining 19 electrons then assign the remaining electrons as lone electrons.

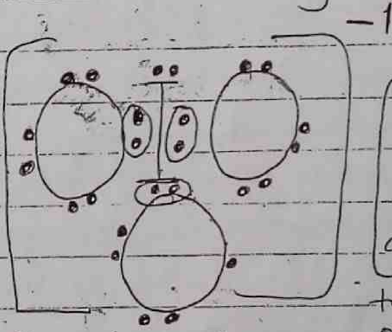


We have assign 7 electrons to the Chlorine atoms remaining 2 electrons assign it to the I atom (ie the 2 electron) So that it will remain one electron to Complete the Cl atom with 7 electrons.



Since the electrons in one of the C.L atoms is 7 Instead of 8 it means it need one electrons to attain octet structure:

ie Write it this way ie bring out ⊖ as −1



The Valency of −1 Shows it will receive an electron to attain to octet structure

We have one lone Pairs on I atom ie 2 lone electrons

" 3 bond pairs ie 6 shared electrons ⊙

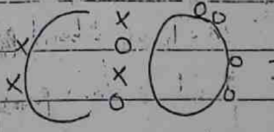
" 19 lone electrons with an extra ion ⊖ on Iodine and oxygen atoms:

⑤ Lewis Structure of CO

Total Valence electrons = 4+6=10

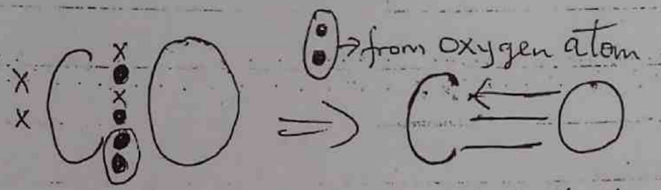
Assign a pair of electrons in between the atoms. Please lets use 2 different lewis Symbol eg ⊗ and • as Carbon electron & oxygen electron to understand this better since it has to obey octet rule

We have assigned 10 electrons but the C atoms have not attained octet structure

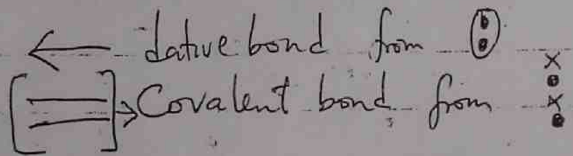


So Oxygen will donate 2 of its electron

Since this is a Special Case of Lewis Structure, hence the 2 electrons donated by Oxygen will form a dative bond



You Can See both have attain octet Structure



⑥ Lewis Structure of H_2O

Total Valence electron is $2H \times 1e^- = 2e^-$

$1(O) \times 6e^- = 6e^-$

$8e^-$

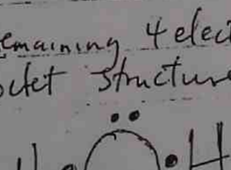
This is another Special Case, because Oxygen is more electronegative than hydrogen but it will be at the Centre as Central atom. [Check the step number 3 in writing Lewis Structure]

So we have,

Then assign a Pair of electron in between the atoms

[we have assigne 4 electrons remaining 4 electron]

Assign the remaining 4 electrons to the oxygen to attain octet structure, Since H is already duplet



[we have 2 bond pairs and 2 lone pairs]

⑦ Lewis Structure for CO_2

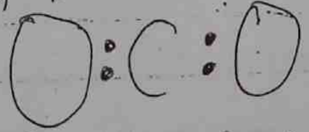
Solution

Total Valence electrons = $1C \times 4 = 4$

$2(O) \times 6 = 12$

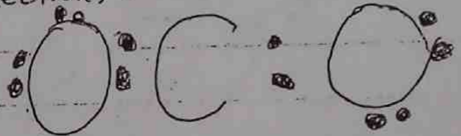
$16e^-$

Since Carbon is less electronegative it will be at the Centre, then assign a Pair of election in between the atoms.

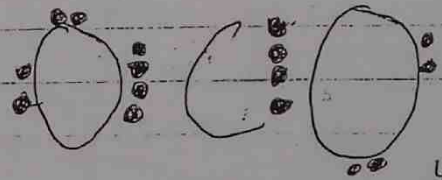


[we have assigned 4 electrons, remaining 12 electrons]

Since the remaining 12 electrons Cannot make the O & C atoms to have octet Structure, assigns 4 electrons to the 2 Oxygen atoms, so that it will remain 4 electrons.



Then finally assign the remaining 4 electrons in between the atoms



You Can See they have attained octet Structure.

⑧ Lewis Structure of NO_3

Total number of Valence electron

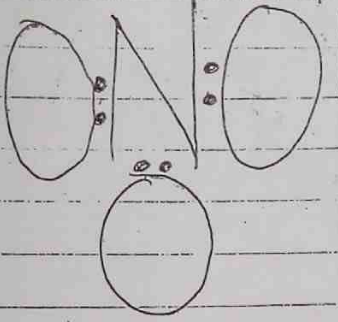
$\Rightarrow 1N \times 5e^- \Rightarrow 5e^-$

$3(O) \times 6e^- \Rightarrow 18e^-$

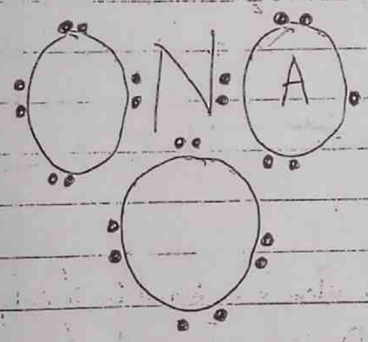
$23e^-$

Since N is less electronegative than O it will be the Central atom ie N is the Central atom

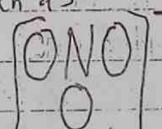
Assign a pair of electron in between the atoms



We have assign 6 electrons remaining (23−6) ie 17 electrons Then assign it to the O atoms



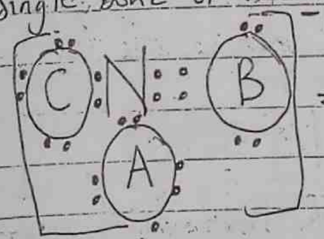
You Can see that it remains one electron for Oxygen A to attain octet Structure. So the Structure must be written as



[−1] means it will accept an electron

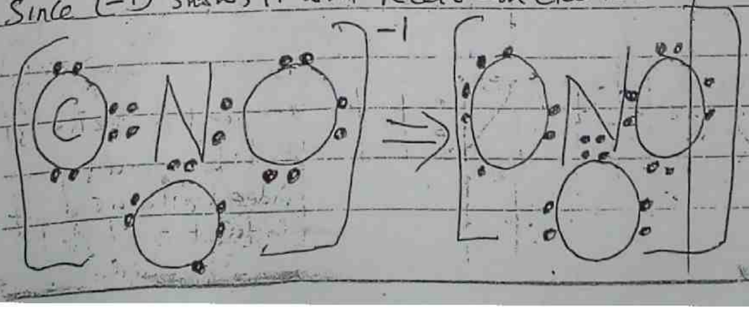
Why we did not worry about the electrons on Nitrogen atoms which is 6 in number instead of 8, is that it attain octet rule by exhibiting resonance (ie alternation of double and single bond or Vice Versa).

ie



$\Rightarrow$ O atom B donated its 2 electron to N for it to attain Octet structure

likewise O atom C can also donate its 2 electron and O atom A can also donate its 2 electron since (−1) shows it will receive an electron.



⑨ Lewis Structure of NH_3

Total Valence electrons

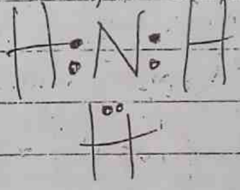
$\Rightarrow$ 1N × 5 electrons = $5e^-$

3H × 1 electron = $3e^-$

$8e^-$

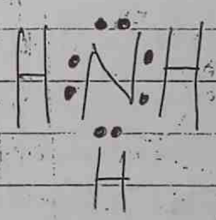
Nitrogen is more electronegative than Hydrogen but it will be the Central atom, because NH_3 & H_2O are to be Considered this way.

Assign a Pair of electron in between the atoms



we have assigned 6 electrons, remaining 2 electrons

Then assign the 2 electron to Nitrogen to attain octet Structure, Since H atoms have attained duplet Structure.

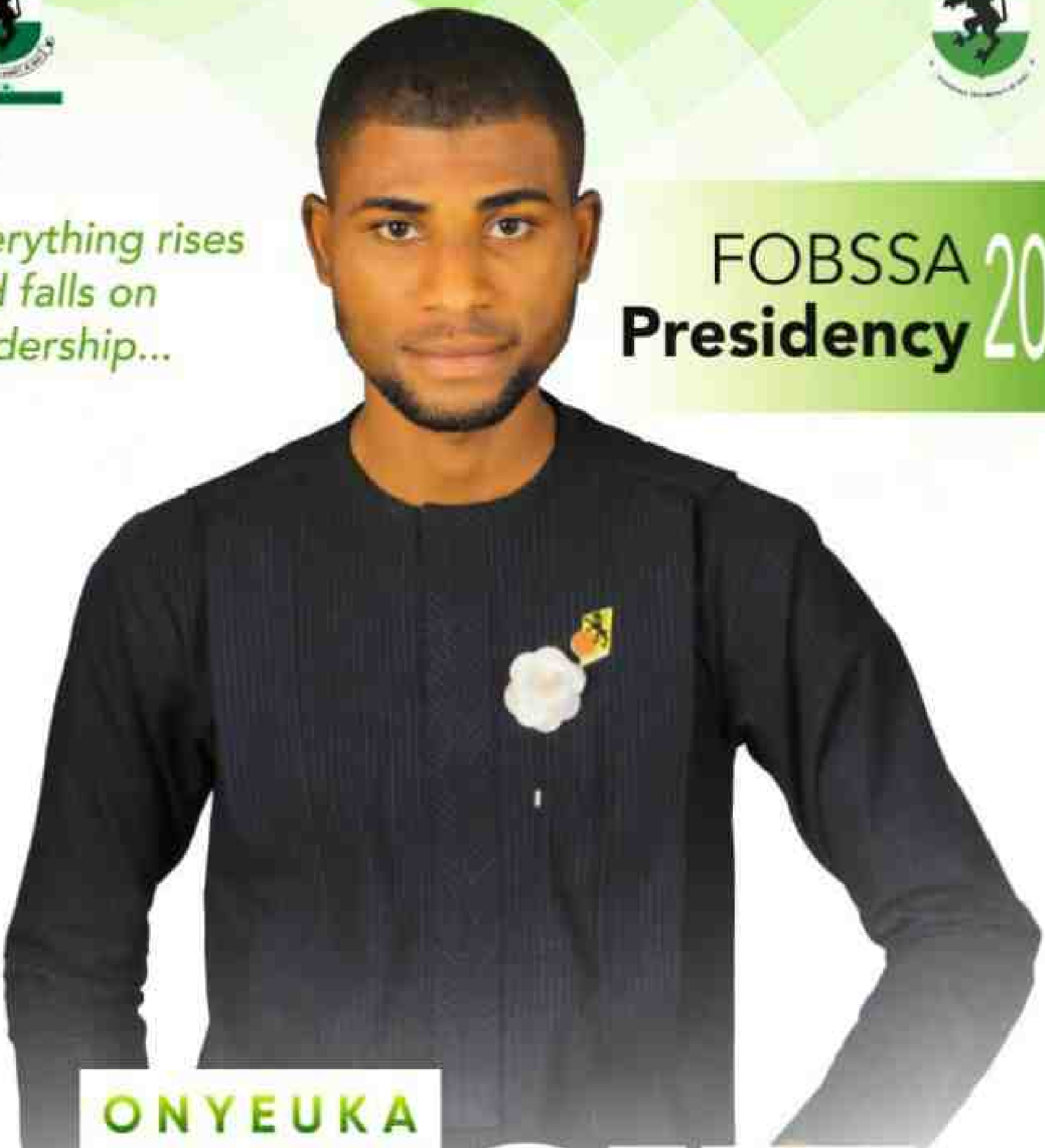


⑩ Draw lewis structures of this: ⓐ HCO_3^{-1}

ⓑ NH_4^{+1}

Everything rises
and falls on
leadership...

FOBSSA 2020 Presidency



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CHIMEZIE

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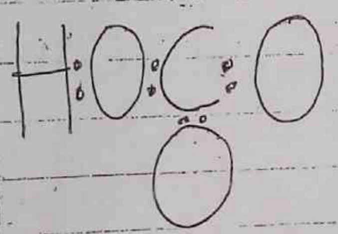
The Man with the **FOBSSA**ites needs at heart...

Since the charge is given as (-1) it
we must add one electron to the Valence electron.
Total Valence electrons in HCO_3

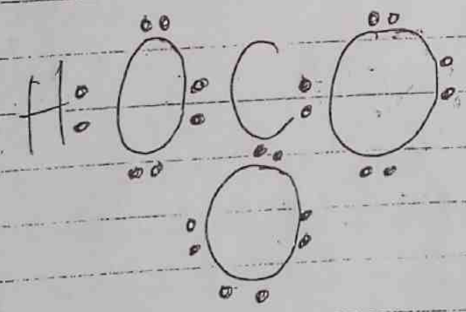
ⓐ ⇒ 1 H atom × $1e^-$ = $1e^-$
1 C atom × $4e^-$ = $4e^-$
3 O atom × $6e^-$ = $18e^-$
+ 1 electron from negative charge (-1) = $1e^-$
$\underline{24e^-}$

C atom is the less electronegative than Oxygen
C atom will be the Central atom
then assign a pair of electron b/w the atoms.



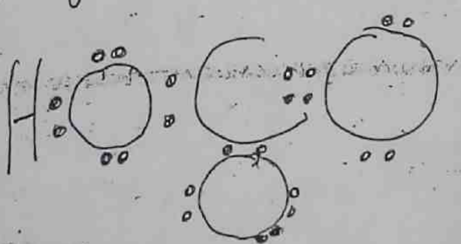
We have assigned $8e^-$ remaining 16 electrons Ok!!!

So assign the remaining electrons to the oxygen.

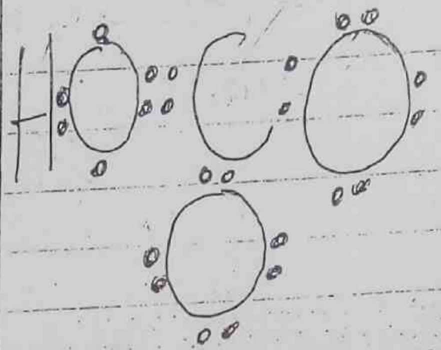


You can see that the Carbon is not in its octet structure

To make it to be in its octet structure the oxygen atom (ie one of the Oxygen atoms) should (will) donate 2 electrons, ie it has to exhibit resonance (alternation of Single and double bonds).



OR



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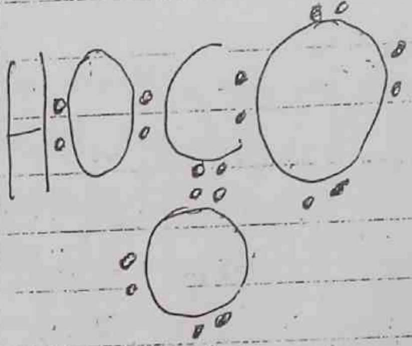
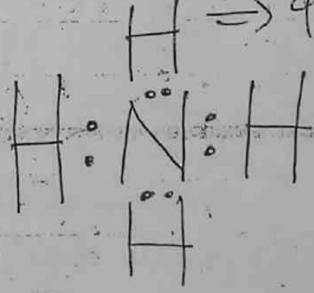


Note: The hydrogen atom can bond well only to either of the oxygen atoms that are Connected to the Carbon by a Single bond but not to the double bond.

ⓑ Lewis Structure of NH_4^+
Since the charge is given as +1, it means, we will subtract one electron from the total valence electron.

Total Valence electron ⇒
1N × $5e^-$ = $5e^-$
4H × $1e^-$ = $4e^-$
$\underline{9e^-}$
Subtracting the one electron ⇒ $9e^- - 1e^- = 8e^-$



We have Completely assigned 8 electrons and the atoms indeed attained octet and duplet structure

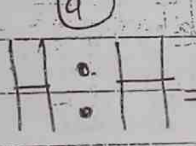
Draw the following Lewis Structure

ⓐ H_2 ⓑ N_2 ⓒ O_3 ⓓ PF_3

ⓔ SO_3

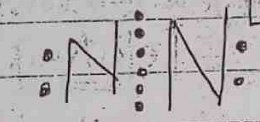
Solution

ⓐ

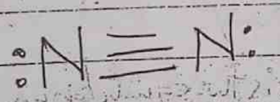


⇒ ie total of $2e^-$ [please follow the ~~sequence~~ (steps) explained above]

ⓑ

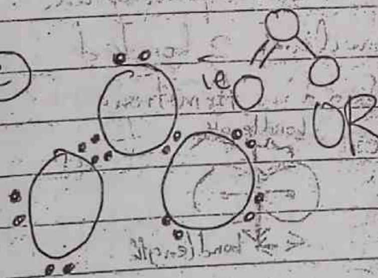


⇒ ie total of $10e^-$



(3 bond pairs and 2 lone pairs)

ⓒ



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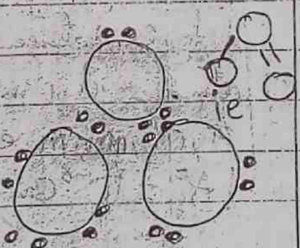
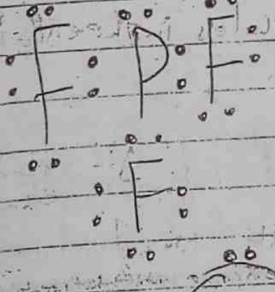
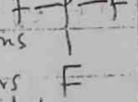


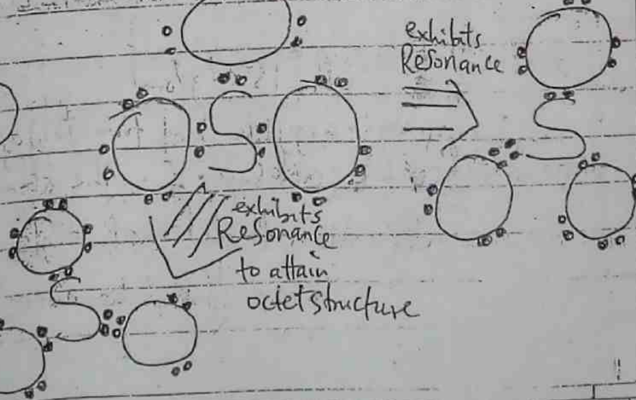
ⓓ



[3 bond pairs ie 3 shared electrons and 10 lone pairs ie 20 lone electrons]



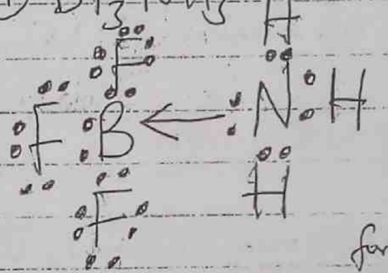
ⓔ



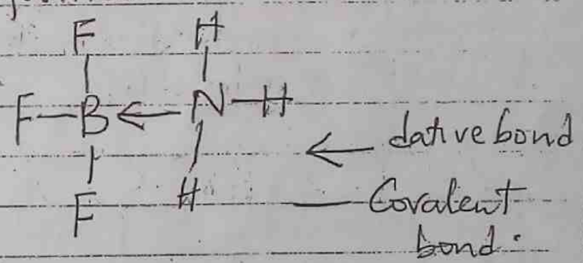
One Compound Can have 2 bonds

This occurs in formation of

① BF_3NH_3



[Nitrogen donated 2 electrons to Boron for Boron to attain octet structure. ie to have 8 electrons. Hence the 2 electron will form Dative bond whereas other bond formed are Covalent bond.



RESONANCE

This is pictured as blending of Structures.

It Can also be said to be the alternation of double and Single bond in a given Structure.

This is as seen in the Lewis Structure of O_3, SO_3, HCO_3^{-1}, NO_3 etc

See the Structures above.

Resonance has 2 main effects

ⓐ It averages the bond characteristics over the molecule.

ⓑ The energy of resonance hybrid structure is lower than that of any single contributing structure.

Two factors that affect the polarisability or deformability of anions are

① Radius of gyration (ie radius of the anion)

ⓑ Change of the anion.

As radius of anion increases, its deformability will increase.

$$F^{-1} < Cl^{-1} < Br^{-1} < I^{-}$$

That is Iodine ion has the highest deformability where as Flourine has the lowest deformability.

Also deformability or polarisability increases as anionic charge increases as seen in

$$F^{-1} < O^{2-} < N^{3-} < C^{4-}$$

C^{4-} has the highest deformability and F^{-1} has the lowest deformability among (in these given anions).

Molecular Parameters are

ⓐ Bond order ⓑ Bond length

ⓒ Bond energy ⓓ Bond angle.

ⓐ Bond order: This is simply defined as half of the number of electrons between the nucleus eg

$$F_2 \longrightarrow F-F$$

ie the single bond (—) has $2e^-$

$$\text{Bond order} = \frac{2e}{2} = 1 \text{ electrons}$$

$$= \underline{1}$$

ⓑ Bond length: This is the equilibrium distance between nuclei of 2 bonded atoms - it is measured in Armstrong A°, cm, nm, m etc.

$$1A^{\circ} = 10^{-8} cm$$

ⓒ Bond energy: This is the dissociation energy (A—B) of the bond between A and B of the molecules in the enthalpy change for the reaction eg.

$$A-B \longrightarrow A_{(g)} + B_{(g)} \quad \Delta H = \text{bond energy}$$

$$H_{2(g)} \longrightarrow 2H_{(g)} \quad \Delta H = 104 \text{ Gl/mol}$$

OR

$$2F \longrightarrow F_{2(g)} \quad \Delta H = 153 \text{ kg/mol}$$

Bond energy = energy (heat of formation) of diatomic atomic molecules from the atoms.

ⓓ Bond angle (θ) This is the internal angle drawn between lines drawn through the nuclei of the bonded atoms.

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There is a close relationship between bond order, bond length and bond energy. for a given pair of atoms, a lower bond order results in a higher bond length and also results in a lower bond energy.

NB A Single bond is longer and weaker while a triple bond is shorter and Stronger.

Since atomic radius Increases down the group, bond length Increases down the group and bond Strength decreases down the group.

Examples

F

Cl

Br $\downarrow$ down the group

Atomic radius Increases

Atomic radius Increases down the group this way F < Cl < Br

Hence bond length S–F < S–Cl < S–Br

But bond Strength S–F > S–Cl > S–Br

ie S–Br has the highest bond length and the lowest bond strength while S–F has the lowest bond length and the highest bond Strength.

⑪

Bond length C≡C < C=C < C–C

Bond Strength C–C < C=C < C≡C

ie Bond length decreases and bond Strength Increases hence bond order Increases.

LEWIS STRUCTURE AND MOLECULAR SHAPES

The shape of a molecule is Important in determining its properties and reactivities. To Construct Shapes of molecules from The lewis Structure, Valence Shell electron-pair repulsion (VSEPR) theory is employed.

The basic principle in this theory is that "each group of Valence electrons around a Central atom is located as far away from the others as possible in order to minimize repulsions".

A group of electrons may be defined as any number of electrons that occupy a localized region around an atom. So, one electron group may Consist of bonding electrons (an electron pair in a Single bond, the electron pair in a double bond and 3 electron pair in a triple bond) and non bonding electrons (a lone pair) or even a lone electron.

Each of these groups is a Separate group of Valence electrons that repels the other groups to minimize the angles between groups and occupy as much Space as possible around the Central atom. It is the 3-dimensional arrangement of nuclei joined by these groups Stated above that gives rise to the Shape of molecule.

The Shape of molecule is defined by the positions of the atomic nuclei.

Bond angle is the angle formed by the nuclei of 2 Surrounding atoms with the Nucleus of the Central atom at the Vertex.

The shapes of molecules when all the surrounding electrons groups are bonding differ from shapes when some are non bonding groups.

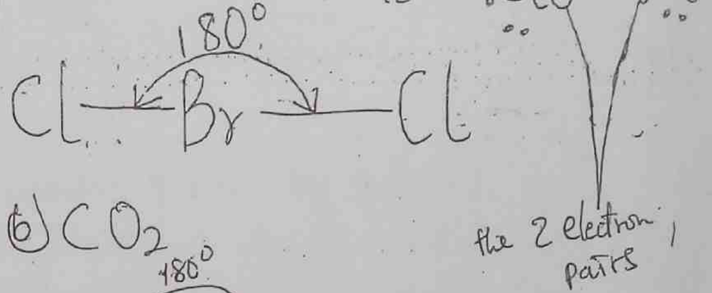
① Shapes and bond angle of molecule with bonding groups

Ⓐ Linear Arrangement:

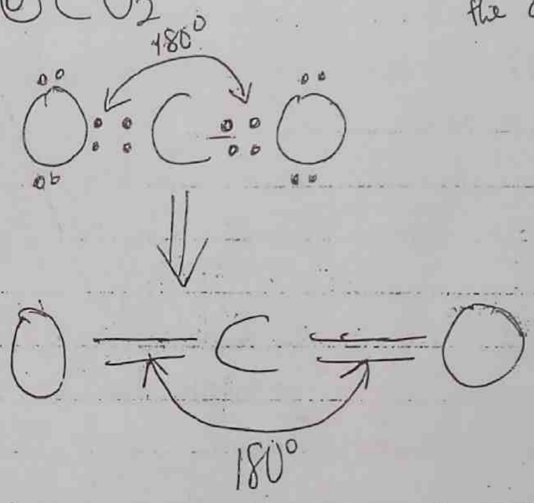
Two electrons groups attached to a Central atom are oriented as far apart as possible, they point in opposite directions.

This linear arrangements of electron groups results in a linear molecule with a bond angle of 180°.

N:B In linear molecules the electron pairs surrounds the central atoms. eg as seen in ⓐ $BeCl_2 \Rightarrow :\ddot{Cl}:Be:\ddot{Cl}:$



ⓑ CO_2



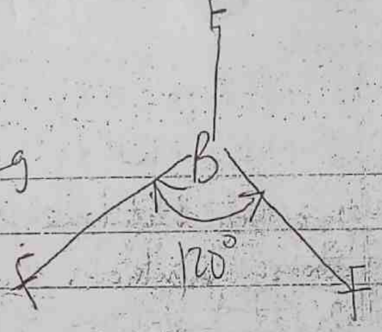
Also note that the lone pairs in Oxygen atoms of CO_2 and lone pairs in Cl atoms of $BeCl_2$ are not involved in the molecular shapes.

② Trigonal Planar arrangement

Trigonal Planar arrangement has an ideal bond angle of 120° and 2 possible molecular shapes.

ⓐ Trigonal planar: In this Shapes, the Central atom has 3 Surrounding atoms. The 3 electron groups are bonding groups, and the bond angle is 120°, example of molecules with trigonal planar shape are SO_3, BF_3

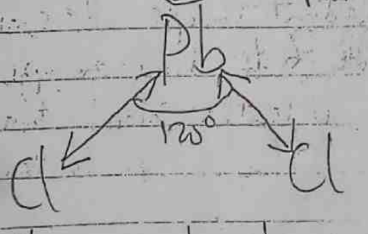
NO_3^- ; CO_3^{2-} eg



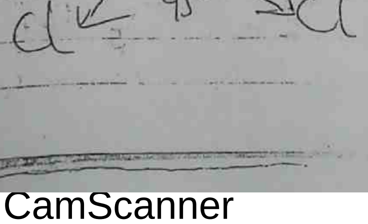
(b) BENT OR V-SHAPED :

V-shaped molecule occurs when one of the 3 electron group is a lone pair. Since a lone pair is held by only one nucleus, it is less Confined and exerts Stronger repulsions than a bonding pair. A lone pair repels bonding pairs more Strongly, than bonding pairs repel each other. The magnitude of electron pair repulsions diminish in the orders, lone pair/lone pair, lone pair/bond pair, bond pair/bond pair. The examples of V-shaped molecules are SO_2, O_3, $PbCl_2$, $SnBr_2$

eg



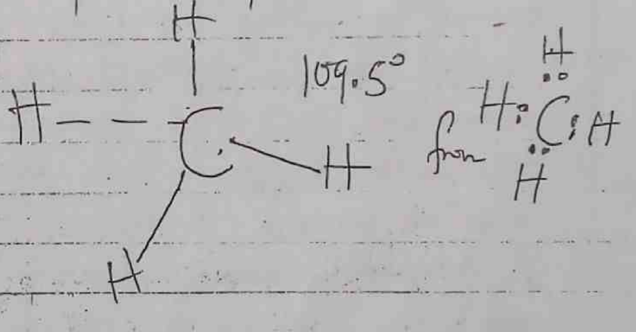
The Stronger repulsion decreases the angle between bonding pairs; an example is the large decrease from 120° to 95° as in $SnCl_2$.



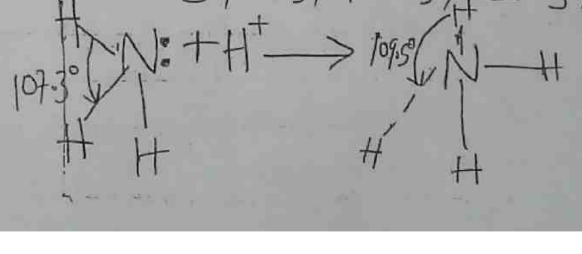
(3) TETRAHEDRAL ARRANGEMENT

All molecules or ions with four electron groups around a Central atom belong to this.

(a) Tetrahedral Shaped: When all pair electron groups are bonding groups as it is in methane CH_4, the shape of the molecule is tetrahedral. Other examples are CH_4, $SiCl_2$, SO_4^{2-}, ClO_4^-



(b) Trigonal pyramidal: This is when one of the 4 electron groups is a lone pair, the other 3 shared pair electrons: its molecule is trigonal pyramid. its bond angle is 109.5°. Measured bond is slightly less than the 109.5°. If there is a Stronger repulsions of a lone pair. In ammonium, for example the lone pair forces the N—H bonding pair closer, and the N—H—H bond angle is 107.3. Common examples of trigonal pyramidal (trigonal) shape is NH_3, PF_3, PCl_3, ClO_3, H_3O^+.



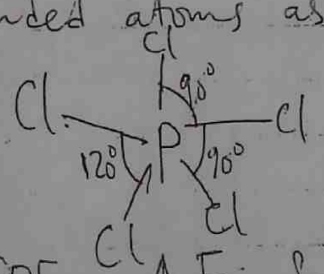
ⓒ Bent or V Shaped: When the 4 electrons groups around the Central atom is madeup of 2 bonding and 2 non bonding groups, the Shape of the molecule is bent or V Shaped. Some examples of molecules with V Shaped in the tetrahedral arrangement are H_2O, SCl_2

We might expect the repulsions from its 2 lone pair to have a greater effect on the bond angle than the repulsions from the single lone pair in NH_3 but the H-O-H bond angle is reduced to 104.5°.

④ Trigonal bipyramidal Arrangement

Molecules with 5 electron groups around the Central atoms have their Central atom from period 3 because only atoms in period 3 have d-orbitals which can expand their Valence electron beyond 8 electrons.

of the Trigonal bipyramidal arrangement is Trigonal bipyramidal shapes. Trigonal bipyramidal Shapes occurs or is gotten when all 5 positions are occupied by bonded atoms as Seen in PCl_5



Other examples are SOF_4, AsF_5. Some Shapes arise from molecules with lone pair at the equatorial Position.

Remember that lone pairs exerts Stronger repulsions than bonding Pairs.

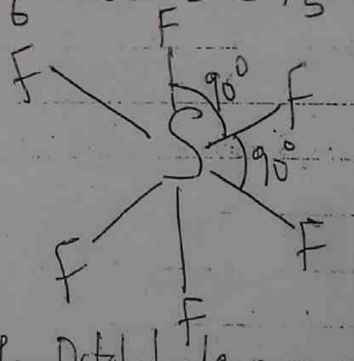
Other Trigonal Bipyramidal arrangements are SeeSaw Shape, T shape etc

⑤ Octahedral arrangement

The molecule or ion with this arrangement has six electron groups around the Central atom. Each of the six electron groups points to one of the 6 Vertices thereby giving rise to the bond angle of 90°.

Molecular Shapes with these arrangements are:

ⓐ Octahedral shape: Here, molecules with 6 bonding groups around the Central atom has octahedral shapes, for instance, SF_6 and IOF_5



Other Octahedral arrangement are

(ii) Square Pyramidal Shape eg IF_5, BrF_5 etc

(ii) Square planar Shape eg XeF_4 etc.

ONYEUKA BRIGHT CHIMEZIE

HYBRIDIZATION

Hybridization is defined as the mixing of different atomic orbitals during Covalent bonding to give equivalent hybrid orbitals.

When an atom is excited through absorption of energy, unpairing of its electrons may occur. For Instance, Berylium in its ground state has an electron Configuration of $1S^2 2S^2$. When excited, the Configurations transforms into $1S^2 2S^1 2P^1$; also the Carbon atomic element has the electronic Configuration of $1S^2, 2S^2, 2P_x^1$. $2P_y^1$, $2P_z^0$ in its ground state But when excited it becomes $1S^2, 2S^1, 2P_x^1, 2P_y^1, 2P_z^1$.

From these excited states of it

When $2S^1$ mixes with $2P^3 \Rightarrow Sp^3$ orbital

when $2S^1$ mixes with $2P^2 \Rightarrow Sp^2$ orbital

When $2S^1$ mixes with $2P^1 \Rightarrow Sp$ orbital

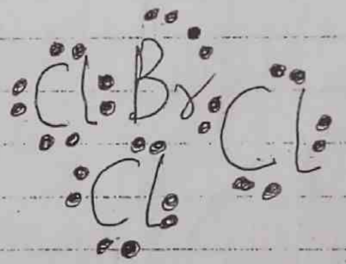
These Sp^3, Sp^2 and Sp^1 orbitals are the orbitals that exist in alkane, alkene and alkyne respectively.

Formular for Calculating Oxidation number ⇒

$$\text{No of Valence electrons} - \left[\text{No of unshared electrons} + \text{No of Shared electrons}\right]$$

N:B one must note that less electronegative element is assigned 0 as its no of shared electrons in the Calculation whereas the more electronegative element is assigned all the shared electrons.

Example: as seen in $BrCl_3$



N:B $Br =$ less electro-negative.

$Cl^- \Rightarrow$ high electro negative

Oxidation Number of Br^-

$$\Rightarrow 7 - [4 + 0] = \underline{\underline{3}}$$

no of Valence electrons in $Br = 7$

no of unshared electron $= 4$

no of shared electron $= 0$ (since it is less electro negative)

Finally

Oxidation number of $Br = 3$

Oxidation number of Cl (the more electronegative)

$$\Rightarrow 7 - (6+2) = -1$$

no of Valence electrons in $Cl = 7$

no of unshared electron in $Cl = 6$

no of Shared electron in $Cl = 2$

ie You Consider Br atom & Cl atom separately in the structure given.

NOTE THE FOLLOWING TABLE.

S/N	HYBRID	COMBINATION OF ORBITAL	NO OF HYBRID	SHAPE OF MOLECULES	EXAMPLE
1	Sp	$2S, 2P_x$	2	Linear	$BeCl_2$
2	Sp^2	$2S, 2P_x, 2P_y$	3	Planar (triangular)	BCl_3
3	Sp^3	$2S, 2P_x, 2P_y, 2P_z$	4	Tetrahedral	CH_4
4	dSp^2	$4S, 3P_x, 3P_y, 3d_{x^2-y^2}$	4	Square planar	XeF_4
5	dSp^3	$4S, 3P_x, 3P_y, 3P_z, 3d_z^2$	5	Trigonal Bipyramid	PCl_5
6	d^2sp^3	$4S, 3P_x, 3P_y, 3P_z,$ $3d_z^2, 3d_{x^2y^2}$	6	Octahedral	SF_6
7	d^2s	$4S, 3d_{xy}, 3d_{xz}$ $3d_y$	4	Tetrahedral	MnO_4^-

SERIES OF LINES IN HYDROGEN SPECTRUM

SERIES	Year found	Region In electromagnetic Spectrum	Value of PQN	Value of MQN
Lyman	1914	Ultraviolet	1	2, 3, 4, ...
Balmer	1885	Visible	2	3, 4, 5, ...
Paschen	1908	Near Infrared	3	4, 5, 6, ...
Bracket	1922	Far-Infrared	4	5, 6, 7, ...
Pfund	1924	Far-Infrared	5	6, 7, 8, ...

N:B PQN = Principal Quantum Number
MQN ⇒ Magnetic Quantum Number

Names	The Great Scientist Suggested that
John Dalton 1665	Atoms were indestructible. But experiment on electrolysis indicates Sub particles of atoms given rise to proton, electron, and neutron.
R. A. Millikan 1910	Electrons exist, which carries a unit electric charge equal to -1.6×10^{-19} C and mass of 9.1×10^{-31} kg.
Lord Rutherford 1911	Atoms Consist of a positive Core Called nucleus and electrons are revolving round just as other planets revolves round the Sun in Solar System.
Moseley 1914	Atomic number represent the number of proton in the nucleus. He did the experiment, by placing the element as anode of X-ray and he assigned the atomic no based on the frequency it emit.
Chadwick 1932	That the nucleus of an atom Contained a neutron when he bombarded a beryllium sheet with α particles.
JJ Thomson 1897	At low Pressure and high Voltage (P.d) a tube Containing gas began to glow. He Called it Cathode rays, he also measured Charge / mass ratio to be 1.76×10^{11} C/kg

Niels Bohr 1913 | Atoms Contain a minute nucleus Consisting of Proton and neutron which Contribute to the mass of the atom while electrons revolve round the nucleus in Circular orbits of definite quantum energy.

THE PERIODIC CLASSIFICATION OF ELEMENTS

The first attempt at classification was into 2: metals and non-metals. This Classification enabled Chemists to anticipate the properties of metals and non-metals yet undiscovered. However, this Classification is not so useful because there was no distinct boundary between the 2 groups.

In the first Classification of element was done by a German Chemist, Johann W. Dobereiner in 1829. He found that elements form themselves into sequence of 3, which he called triads. According to him, when the 3 elements were arranged in increasing atomic masses, the atomic mass of the middle element was approximately equal to the average of the first and the third element.

One of such triads is Chlorine, bromine and iodine with atomic masses of 35·5, 80 and 126·9 respectively. The average of the first and 3rd element is

$$35{\cdot}5 + 126{\cdot}9 = \frac{162{\cdot}4}{2} = 81{\cdot}2$$

This Value is close to the atomic mass of bromine 80. Other examples of such sequence are Calcium (80), Strontium (88) and Barium (137); Sulfur (32) Selenium (79) and Tellurium (128).

Another Classification was done by John Newlands in 1864. He found that if elements were arranged in increasing atomic masses, a particular element has physical and Chemical properties closely similar to those eight (8) places before and after it. e.g. He said that F has similar physical and Chemical properties with H (8 places before) and Cl (8 places after.

The first good attempt on Classification of elements was made in 1869 by 2 Scientist: Julius Lother Meyer a German, and Dmitri Mendeleeff a Russian, who worked independently and from different approaches to establish a more detailed relationship between atomic weight and properties.

Meyer plotted atomic weight of elements against the physical property of the element ie atomic Volume. He graphically and rightly showed a periodic Variation of atomic Volume with atomic weight.

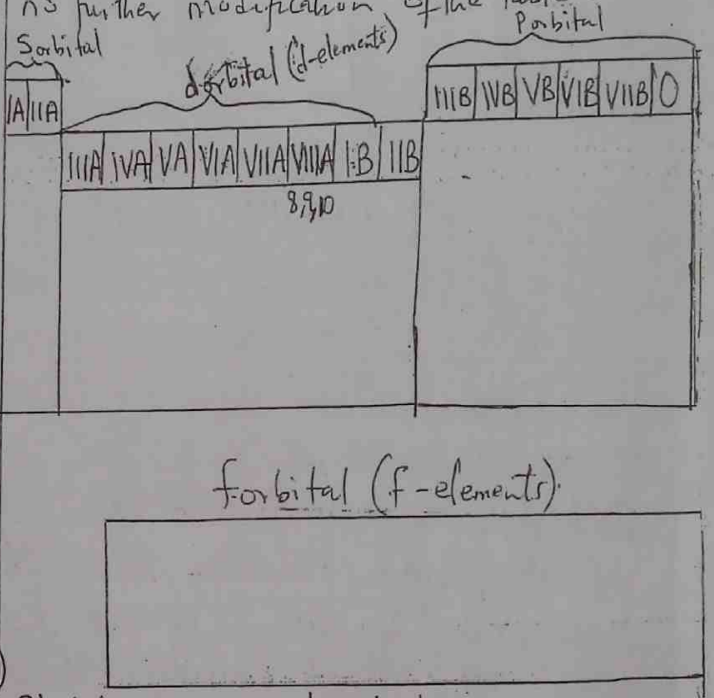
Mendeleef also looking at both physical and Chemical properties of elements arranged the elements into horizontal rows and Vertical Columns according to their Increasing atomic weight. He proposed the periodic law ie "that the properties of elements

Vary in relation to their mass". He was So Confident about his table that he left gaps for undiscovered elements, his table Composed (0-8) groups and 12 Series. Mendeleef was given Credit for establishing the periodic System of elements because using his table he was able to predit the properties of unknown element. Due to the fact that the Mendeleefs arrangement was based on Increasing atomic weight, Several elements appeared to be out of place in the table.

THE MODERN PERIODIC TABLE

This is the Classification of elements in the 20th Century.
As the knowledge of atomic Structure advanced, Moseley H.G.J discovered after his experiment in 1913, that "the Square root of the characteristics frequency of Xrays from an element is a linear function of its atomic number" This is known as Moseley's Law.
Based on this finding, elements are now Classified on atomic number basis, Such that the Periodicity of properties is associated with periodicity of election structure. This led to the modification of the Mendeleef's periodic law.
Also Mendeleev original periodic table was modified to the long form as proposed by Danish Chemist, Julius Thomson.

However, the discovery of noble gases in 1890—1900 by William Ramsay needed no further modification of the table.



Sketch of the periodic table

N:B S-orbital ≡ S-blocks
P-orbital ≡ P-blocks
d-orbital ≡ d-blocks
f-orbital ≡ f-blocks.

Under S-blocks: We have Alkalimetals and Alkaline earth metals.

Alkalimetals: This refers to group 1 (one) elements except H. They are good Conductors of heat and electricity; are soft to be cut with a knife, have low density and high electropositivity. They are easily oxidized and as Such not found free in nature.

Alkali earth metals: They are group 2 elements They are Silvery white, malleable ductile and harder than Alkali metals. The activity of these elements increases from top to bottom.

N.B Vertical Component of the Periodic table (ie Columns) are Called the groups. Horizontal Component of the periodic table (rows) are Called the periods.

Halogens: They are group VIIB. They are Salt formers - Examples are F, Cl, Br, I etc

Noble gases: This belong to group VIIIB or O. Until recently, the elements in this group were Called inert gases or unreactive element because no Chemical reactions Involving these elements were known. Their electronic Configuration may be represented $1s^2\ np^6$ ie Complete electronic Structure. However, Xenon, krypton, radon have been found to form Compounds and So the name Changed to noble elements Instead of inert gases/unreactive elements.

The representative elements:

They are elements that have incompletely filled outer S or P orbitals. These elements Show distinct and fair Variations in their Properties with atomic number. They are groups IA, IIA, IIIB, IVB, VB, VIB, VIIB, VIIIB or O ie Elements in S or P block (orbitals) are representative elements

The Coinage metals:

These are groups IB metal; Copper, Silver and gold. They are used as alloys in the Manufacture of Coins, they are therefore Called the Coinage metals. The gaseous atoms of these elements Cu, Ag and Au have an outer $(n-1)d^{10}\ ns^1$ electronic Configuration or $(ns^1\ (n-1)d^{10})$. The 3 metals resemble one another closely; they are relatively hard, extremely malleable and ductile and have high melting point close to 1000°C.

d-transition Element

All elements found in group III A to group IB are known as d-transition elements or transition metals. They are found between group II and Group III B of the periodic table drawn (sketch above). They are Considered as transition between the alkaline elements (base formers) on the left and the acid formers on the right. All of them are metals and are Characterized by last electrons being added to d-orbital energy level, from 8 to 18 electrons. They are:

1st transition Series: Sc through ${}^{29}Cu$
2nd transition Series: ${}^{21}_{39}Y$ through ${}_{41}Ag$
3rd transition Series: ${}_{59}La$ through ${}_{79}Au$

4th transition Series: ${}_{89}Ac$ through 111

The elements in group IIB as we said before are not d-transition element because their last electrons go into S orbitals, but they are listed as transition elements since their Chemical Properties are Simular to the transition metals.

Inner Transition Element (f-transition elements).

These are Called f-transition elements because their electrons are added to the f-orbital or the Second from the highest occupied energy level building from 18 to 32 electrons. They are found between groups $\underline{III}$A and $\underline{IV}$A.

1st Inner transition: $_{58}Cs$ to $_{71}Lu$
(Lanthanides)
2nd Inner transition: $_{90}th$ to $_{103}Lr$
(Actinides)

N:B The following Points

If We have $①S^{\boxed{1}} \rightarrow$ S block
It Implies that $\boxed{1}$ → tells you its group one
① → tells you its period one

If we have $nS^x nP^y$
(it Implies that n is the period
x+y+10 is the group
and finally it is in P block

eg $1S^2 \boxed{2S^2\ 2P^5}$ = Electronic Configuration of fluorine

n ie the period where the element is located is 2

The group will be 2+5+10 ⇒ group 17

and finally it is in P-block.

If we have $nS^x(n-1)d^z$
period = n
Block = d
group = x+z

eg $1S^2 2S^2 2P^6 3S^2 3P^6 \boxed{4S^2\ 3d^6}$
$nS^x\ (n-1)d^z$

Period = n = 4

block = d
group = 2+6 = 8 $\overset{VIIIA}{(8A)}$

Given the following elements in Period 2 and 3 (Representative elements).

Period 2 elements ⇒ Li, Be, B, C, N, O, F, Ne

Period 3 elements ⇒ Na, Mg, Al, Si, P, S, Cl, Ar

As you Can See the diagonal arrow, this shows that Li & Mg,
Be & Al,
B & Si, etc
have diagonal relationship.

This is because they have Similar electronegativity or electropositivity Value. Metallic nature increases down a group and non metallic nature increases across a period. Along a diagonal, there is both an increase in metallic and non-metallic nature. Each tends to Cancel the effect of the other.

Comparism of the elements of Lithium and Magnesium

Reactions	Diagonal Relationship
Trioxocarbonates(IV) CO_3^{2-}	Both $LiCO_3$ and $MgCO_3$ are decomposed by heat; other group 1 that form Carbonates with CO_3^{2-} donot deComposed by heat
Chlorides, Cl^-	Both Chlorides of Li^{+1} and Mg^{2+} are hydrated ie $LiCl \cdot 2H_2O$, $MgCl \cdot 6H_2O$ but the other group 1 Chlorides are not hydrated.
Oxides, O^{2-}	Both Li^{+1} and Mg^{2+} forms normal oxides when burn in oxygen The other group 1 metals forms peroxides eg Na_2O_2 Normal oxides form by Lithium and Magnesium are Li_2O, MgO
Hydration	Both Li^+ and Mg^{2+} ions are heavily hydrated in Solution, they are both dense Centres of Charge.

Comparism of Elements Beryllium and Aluminium

Reactions.	Diagonal Relationship.
With Trioxonitrate acid (HNO_3)	Both Be and Al are rendered Passive and will react
With alkali eg NaOH	Both dissolve in NaOH given off Hydrogen gas.
Chlorides, Cl^-	Both Chlorides $BeCl_2$ and $AlCl_3$ are electron deficient; both act as lewis acid.
Oxides, O^{2-}	Both oxides BeO & Al_2O_3 form Amphoteric oxides.

Comparism of Elements Boron and Silicon

Reactions	Diagonal Relationship
Hydrides	Both Boron and Silicon give a wide Variety of unstable hydrides, the Boranes and the Silanes eg BH_4, B_2H_6, SiH_4, Si_2H_6
Chlorides, Cl^-	Both Chlorides (BCl_3 and $SiCl_4$) are Volatile and will undergo hydrolysis to give an acidic Solution.
Oxides, O^{2-}	Both oxides are weakly acidic when dissolved in water $SiO_2 + H_2O \rightarrow H_2SiO_3$; $B_2O_3 + 3H_2O \rightarrow 2H_3BO_3$

These reasons discussed above in the table is why they have diagonal relationship.

Table for Periodicity of Properties

Properties	Across Period	Down the group
Ionic Radius	Decrease	Increase
Atomic Radius	Decrease	Increase
Ionization energy	Increase	decrease
Electron affinity	Decrease	Increase
Electrical property	Decrease	Increase
Electronegativity	Increase	Decrease.

Increase in atomic size down the group occurs because of increase in screening effect as electrons are added to new shells hence increase the n-values.

Assignment given was.

① Arrange the Isoelectronic ions in order of increasing atomic size.

$F^-, O^{2-}_8, Na^+_{11}, Al^{+3}_{13}, Ne_{10}, Mg^{2+}_{12}, N^{3-}_7$

Solution

$$Al^{3+}_{13} < Mg^{2+}_{12} < Na^+_{11} < Ne_{10} < F^-_9 < O^{2-}_8 < \boxed{N^{3-}_7}$$

(II) Between N^{3-} and K^+ which is bigger?

Solution

K^+ is bigger than N^{3-}

Ionization energy is also known as ionization potential.

Subsequent ionisation potential is usually higher than the first ionization energy.

Due to the addition of electron shells down the group ability to remove electrons increase, thereby ionisation energy decrease down the group but increases across the period.

There is an anomaly when removing electrons from Group 2 & Group 13 element.

Group 2	Group 13
ns^2	ns^2np^1

This is because it is easier to remove from a fully filled orbital (lower energy) than from a partially filled orbital (higher energy). Also from Group 15 to Group 16, because of repulsion it makes it to require less energy.

Lithium Can form Covalent bond. Also it is more Covalent than ionic.

All Carbonates in group 1 Cannot be decomposed by heat but Lithium Carbonates Can.

Lithium react with Carbon to form ionic Carbide ; none on Group one Can.

Lithium form nitride with nitrogen ie LiN ie Lithium donot behave like alkali metals.

Al_2O_3 is amphoteric, Al also form Covalent Compound like Lithium, Be etc.

$AlCl_3$ and $BeCl_2$ dimerise

Hydrogen belongs to neither the alkali metal nor with the halogens but has properties in Common with both, eg it forms +1 oxidation numbers as the alkali metals but H^+ has no Chemical existence because there is no ionically bonded hydrogen in Compound but Covalently bonded.

While fluorine is more reactive than its heavier group members, lithium is much less reactive than its group members.

Elements that shows properties of both metals and non metals are Called metalloids

Elements with most powerful metallic

Property are found in group IA. They are reducing agent (ie they Cause electron gain because they loss electron). The metallic Character Increases, down the group, Such that metals at the bottom of the group are most powerful reducing agents. This is due to the fact that their outermost electron, S electron is extremely well shielded from nuclear charge ie it Can easily be lost.

Periodic Table and Oxidation Number.

We Can Correlate the oxidation number of elements with their position in the periodic table. As a guide, the group number Corresponds to the oxidation no, for element in Group 1, 2, 3, 4.

eg Li = +1, Na = +1, Mg = +2, Be = +2 etc,

Examples of oxidation number of elements and electronic Configuration are:

① Hydrogen	+1, -1	$1S^1$
② Helium	0	$1S^2$
③ Lithium	+1	$1S^2 2S^1$
④ Beryllium	+2	$1S^2 2S^2$
⑤ Boron	+3	$1S^2 2S^2 2P^1$
⑥ Carbon	+2, +4, -4	$1S^2 2S^2 2P^2$
⑦ Nitrogen	+3, -3, +5	$1S^2 2S^2 2P^3$
⑧ Oxygen	-2	$1S^2 2S^2 2P^4$
⑨ Fluorine	-1	$1S^2 2S^2 2P^5$
⑩ Neon	0	$1S^2 2S^2 2P^6$
⑪ Sodium	+1	$1S^2 2S^2 2P^6 3S^1$
⑫ Magnesium	+2	$1S^2 2S^2 2P^6 3S^2$

㉑ Scandium	+3	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$
㉒ Titanium (Ti)	+4	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$
㉓ Vanadium	+2	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$
㉔ Chromium	+3	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$
㉕ Manganese	+2, +4	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$
㉖ Iron	+2, +3	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
㉗ Cobalt	+3	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$
㉘ Nickel	+2	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$
㉙ Copper	+1, +2	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$
㉚ Zinc	+2	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$

Oxidation no of non-metal is the Valency (power of the non-metal). Oxidation no of a Molecule; Compound eg H_2, O_2, NaCl is zero.

Please as a year one Student of Chemistry 101 You must know atleast the first 30 elements and their atomic number.

Please Memorize it this way

He → Hydrogen		Hello →
Has → Helium		Helen →
Little → Lithium		Listen →
Bright → Beryllium	OR	Broken →
Brain → Boron		Bottle →
Can → Carbon		Can →
Not → Nitrogen		Not →
Offer → Oxygen		Over →
full → Fluorine		flow →
Nine → Neon		Neon →

Subject → Sodium		So → Sodium
Many → Magnesium		May → Magnesium
Arts → Aluminium		All → Aluminium
Student → Silicon		Stupid → Silicon
Prefer → phosphorous		People → phosphorus
Such → Sulphur		Suffer → Sulphur
Combination → Chlorine		Chlorine → Chlorine
As → Argon	OR	Argon → Argon
Physics → Potassium		Potassium → Potassium
Chemistry → Calcium		Calcium → Calcium
Say → Scandium		Say → Scandium
To → Titanium		To → Titanium
Vero → Vanadium		Vero → Vanadium
Cry → Chromium		Cry → Chromium
More → Manganese		More → Manganese
For → Iron		For → Iron
Company → Cobalt		Company → Cobalt
Nigeria → Nickel		Nigeria → Nickel
Copper → Copper		Copper → Copper
Zinc → Zinc		Zinc → Zinc

They are the first 30 elements Ok!!!

ONYEUKA BRIGHT CHIMEZIE

The positions of hydrogen and helium Vary in different Versions of the table. Helium is usually placed above the inert gases (Group 0) because of its properties. However, it is not of course a p-block element ie (He), its electronic structure being $1s^2$.

Hydrogen is placed above Group 1 but Separated from it. it usually forms Singly Charged H^+ ions like the group 1 elements but otherwise is not Very Similar to them. It Can also form H^{-1} ions and is placed above the halogens in Some Versions of the Periodic table.

The periods are numbered starting from Period 1 Containing just Hydrogen and Helium.

Period 2 Contains the elements Lithium to Neon and So on.

The groups are usually numbered 1 to VII plus 0. However, it has recently been proposed that numbering 1 to 18 across the d-block.

Note: The electron Configuration Can be written in Noble gas notation, a Short hand option. This Can be done by using the Noble gas that Comes before the element, Putting it in brackets and Continuing the electron notation from there.

Detailed trends in physical and Chemical properties of particular elements in Chosen groups for the blocks are given hereunder!

Hydrogen, Sodium, Calcium, Nitrogen, phosphorous, Oxygen, Sulphur, Chlorine, Bromine, Iodine, Aluminium, Iron, Manganese, Copper, Zinc, Lanthanides and Actinides.

HYDROGEN

Hydrogen is one of the most Important elements in the world. It is all around us. It is a Component of water H_2O, fats, petroleum, table Sugar ($C_6H_{12}O_6$), ammonia (NH_3) and hydrogen peroxide (H_2O_2).

OCCURRENCE

There is little unCombined hydrogen on the earth but it is abundant, Combined in Compounds particularly water and hydroCarbons. (eg Crude oil and Coal). On Earth hydrogen is mostly found in association with Oxygen. Its most abundant for being water (H_2O).

Structure and Isotopes of Hydrogen.

1- Protium (^1_1H) is the most Common Isotope, Consisting of 99.98%, of naturally occurring hydrogen. It is a nucleus Containing a Single Proton. It is its most Common isotope Consisting of just one Proton and one electron.

— Deuterium (^2_1H) is another Isotope Containing a proton and neutron, and Consisting of only 0.156% of the naturally occurring hydrogen. Commonly indicated with symbol D.

— Tritum (^3_1H) with one proton and two neutrons is Called tritum (T) It is a radioactive isotope with a 12.3 year halflife, which is Continously formed in the upper atmosphere due to Cosmic rays.

Chemically, the 3 Isotopes have the Same Properties because each has one electron in the Valence Shell. They differ in physical Properties on account of differences in the number of neutrons.

SOME METHODS OF PREPARING HYDROGEN

① Action of water on Some Metals

Metals with high negative electrode potentials (Groups IA and II A) react with Cold Water e.g.

$$2Na(s) + 2H_2O_{(L)} \rightarrow 2Na^+OH^-_{(aq)} + H_2O$$

Mg powder reacts slowly with water but Vigorously with Steam.

$$Mg_{(s)} + 2H_2O_{(g)} \rightarrow MgO_{(s)} + H_{2(g)}$$

Heated Mg, Zn, Fe, Co, Ni, Pb, and Sn will decompose steam with the formation of the oxide of the metal and hydrogen.

② Action of dilute acids on metals.

All metals higher than hydrogen in the electrode potential series react with dilute hydrochloric and Sulphuric acids to give hydrogen eg.

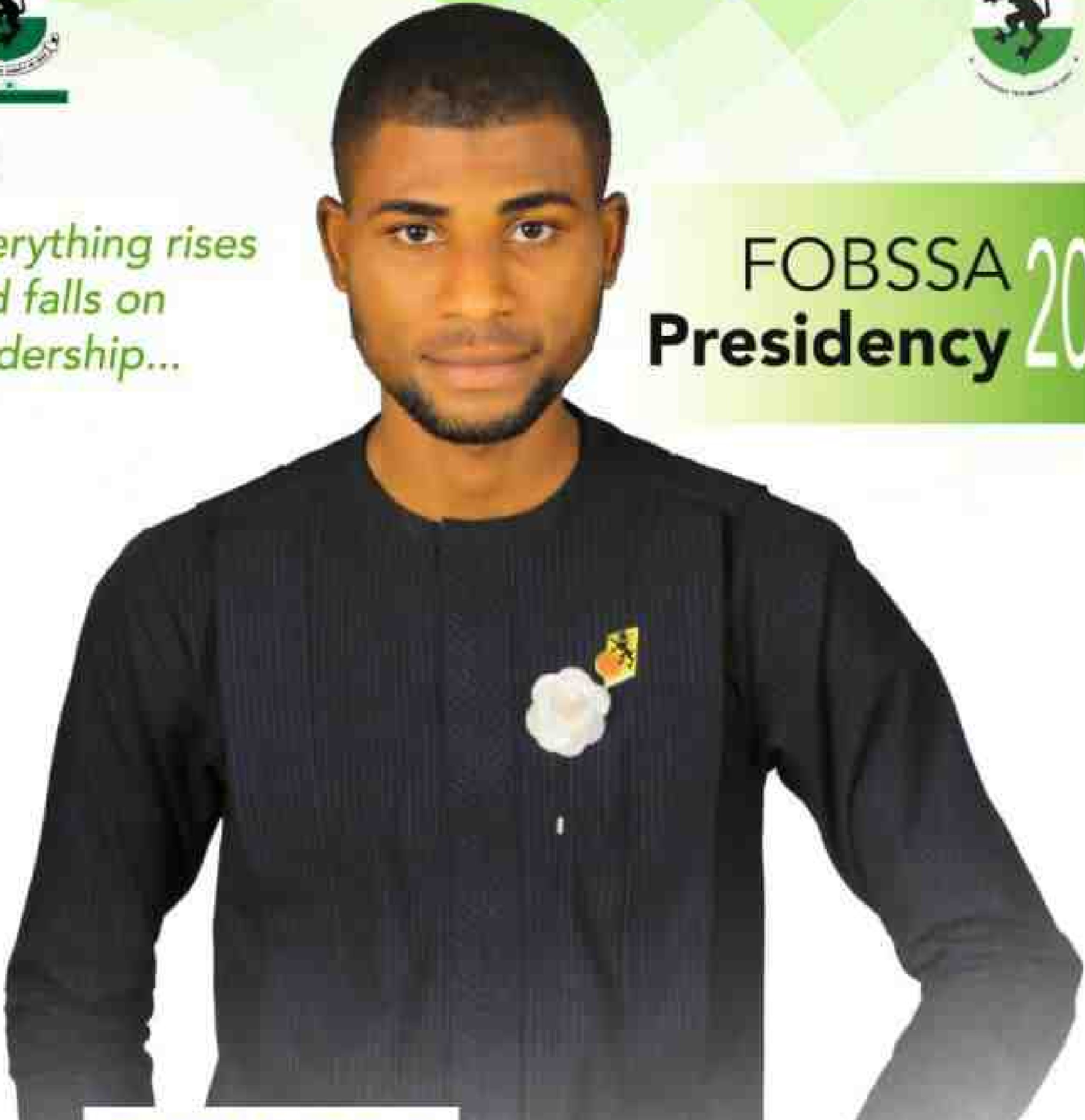
$$Zn_{(s)} + 2H^+_{(aq)} \rightarrow Zn^{2+}_{(aq)} + H_{2(g)}$$

Hydrogen gas Can be Prepared by reacting a dilute Strong acid like Hydrochloric acids with an active metal. The metal becomes oxides, while the H^+ (from the acid) gets reduced to hydrogen gas. This method is only practical for Producing Small amounts of hydrogen in the lab, but is much too Costly for industrial Production.

$$Zn_{(s)} + 2H^+_{(aq)} \rightarrow Zn^{2+}_{(aq)} + H_{2(g)}$$

Everything rises and falls on leadership...

FOBSSA
Presidency 2020



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BRIGHT
CHIMEZIE

FACULTY OF BIOLOGICAL SCIENCE STUDENTS ASSOCIATION **(FOBSSA)**

*The Man with the **FOBSSA**ites needs at heart...*

③ Action of Strong alkalis on Zinc and Aluminium:

$$Zn_{(s)} + 2OH^{-}_{(aq)} + 2H_2O_{(l)} \downarrow Zn(OH)^{2-}_{4\,(aq)} + H_{2(g)}$$

Zincate ion

$$2Al_{(s)} + 2OH^{-}_{aq} + 6H_2O \downarrow 2Al(OH)^{-1}_{4\,(aq)} + 3H_{2(g)}$$

Aluminate

⑭ The purest form of $H_{2(g)}$ Can Come from electrolysis of $H_2O_{(l)}$ using an amsde of Zinc amalgam

$$2H_2O_{(l)} \rightarrow 2H_{2(g)} + O_{2(g)} \quad \Delta H = 572kJ$$

⑤ H_2O is the most abundant form of hydrogen on the planet, So it Seems logical to try to extract hydrogen from water without electrolysis of water. To do so, we must reduce the hydrogen with +1 oxidation State to hydrogen with 0 oxidation State (in hydrogen gas). Three Commonly used reducing agents are Carbon (in Coke or Coal), Carbon monoxide, and methane.

These react with water Vapour form $H_{2(g)}$:

$$C_{(s)} + 2H_2O_{(g)} \rightarrow CO_{(g)} + H_{2(g)}$$

$$CO_{(g)} + 2H_2O_{(g)} \rightarrow CO_{2(g)} + H_{2(g)}$$

MANUFACTURE OF HYDROGEN

- From Methane →

$$CH_{4(g)} + H_2O_{(g)} \rightarrow CO_{(g)} + 3H_{2(g)}$$

- Cracking of Hydro. into Smaller molecules eg

$$C_6H_{12(g)} \rightarrow C_6H_{6(g)} + 3H_{2(g)}$$

- Electrolysis of water

- Action of Steam on Coke (Bosch reaction). The Process takes place in three stages.

Stages 1 ⇒ Steam is passed over white hot Coke

$$H_2O_{(g)} + C_{(s)} \xrightarrow{1700°C} CO_{2(g)} + H_{2(g)}$$

Stage 2 ⇒ The mixture of Carbon monoxide and hydrogen, known as water gas, is mixed with more Steam and passed over an Iron

Catalyst. Only the Carbon monoxide reacts.

$$H_2O_{(g)} + CO_{(g)} \rightarrow CO_{2(g)} + H_{2(g)}$$

Stage 3: Under pressure the Carbon dioxide is dissolved in water, thus leaving the hydrogen available for further use.

PROPERTIES OF HYDROGEN

PHYSICAL PROPERTIES

It is a Colourless, odourless gas without taste. The pure gas is not Poisonous, but does not Support life. H_2 is Sparingly Soluble in water. It is lightest known gas, its absolute densities being 0.0000899g
These are the Physical Properties of hydrogen.

CHEMICAL PROPERTIES

Hydrogen will Combine with all elements apart from the noble gases. The more electropositive (less electronegative) S-block metals are able to form Compounds with hydrogen in which the hydrogen accepts an electron to form the hydride ion H^+ in the majority of Cases hydrogen forms Covalent bonds with the other elements.

Reactions with Non-metals: Forms hydrides with many elements.
Hydrides of non-metals are Covalent eg CH_4, NH_3, H_2O, HCl.

Hydrides of reactive metals are Ionic.
$\overset{+}{Na}H^-$, $Ca^{+2}(H^{-1})_2$

Forms hydrogen bonds with highly electronegative atoms eg in liquid HF, H_2O.

Hydrogen + Halogen → Hydrogen halide
eg

$$H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$$

Hydrogen + Oxygen → Water

$$2H_{2}(g) + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

Hydrogen gas reacting with Oxygen to produce water and a large amount of heat.

- Hydrogen also has an ability to form Covalent bonds with a large Variety of Substances, Because it makes Strong O-H bonds.

- It is also Considered a good reducing agent for metal oxides.

Example: $CuO_{(s)} + H_{2(g)} \rightarrow Cu_{(s)} + H_2O_{(g)}$

ie $H_{2(g)}$ passes over $CuO_{(s)}$ to reduce the Cu^{2+} to Cu(s) while getting oxidized itself.

— Liberated from acids by many metals

$Zn_{(s)} + 2H^{+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + H_{2(g)}$

ie when $HCl_{(aq)}$ react with Zinc it produces $H_{2(g)}$

$Zn_{(s)} + 2HCl \rightarrow ZnCl_2 + H_{2(g)}$

— Hydrogen wanting to give up its single electron causes it to act like an alkali metal

$H_{2(g)} \rightarrow 2H^{+}_{(g)} + 2e^{-}$

— A half filled valence shell with one electron also causes hydrogen to act like a halogen because it wants to gain Noble gas configuration by adding an electron.

$H_{2(g)} + e^{-} \rightarrow 2H^{-}_{(g)}$

— Reactions with Active metals; Hydrogen accepts electron from an active metal to form ionic hydrides like LiH. By forming an ion with -1 charge, the hydrogen behaves like a halogen.

— $2M_{(s)} + H_{2(g)} \rightarrow 2MH_{(s)}$

M = group 1 metals. Example

$2K_{(s)} + H_{2(g)} \rightarrow 2KH_{(s)}$

$2K_{(s)} + Cl_{2(g)} \rightarrow 2KCl_{(s)}$

$M_{(s)} + H_{2(g)} \rightarrow MH_{2(S)}$ $\Rightarrow$ where

M = group 2 metals.

Example:

$Ca_{(s)} + H_{2(g)} \rightarrow CaH_{2(s)}$

$Ca_{(s)} + C_{2(g)} \rightarrow CaC_{2(s)}$

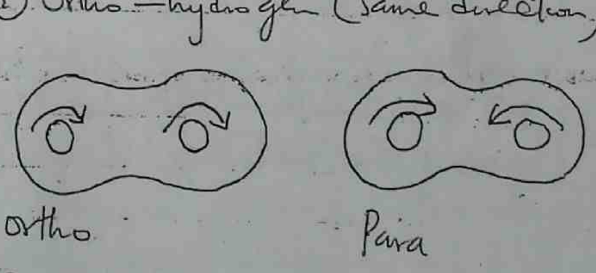
Reactions with Transition metals

- Reactions of hydrogen with transition metals to form metallic hydrides.

Allotropes of hydrogen

There are 2 allotropes of hydrogen

① Para-hydrogen (Opposite direction)
② Ortho-hydrogen (Same direction)



Directions of spin determines the allotropes of hydrogen.

USES OF HYDROGEN

- IN THE MANUFACTURE OF AMMONIA BY THE HABER PROCESS;

This is used in turn to manufacture nitric acid, which can then be converted into explosives, dyestuffs and nitrogenous fertilizers.

- EXTRACTION OF SOME METALS FROM THEIR ORES: Because hydrogen is a good reducing agent, it is used to produce metals like Iron, Copper nickel, Molybdenum, tungsten and Cobalt from their ores.
- Liquid hydrogen (Combined with Liquid Oxygen) is a major Component of rocket fuel (as mentioned above Combination of hydrogen and oxygen relapses a huge amount of energy)

- IN OXY-HYDROGEN BLOW LAMP FOR WELDING

- MANUFACTURE OF MARGARINE FROM VEGETABLES

Hydrogen is used for the hydrogenation of oils. Hydrogenation entails replacing double bonds in oils by hydrogen, Converting the double bonds into Single bonds.

This transformation of unSaturated fats to Saturated fats drastically increases the Shelf life of many foods.

- MANUFACTURE OF ORGANIC CHEMICALS eg CH_3OH, hydrogen Chloride and hydrochloric acid

- for filling balloons and airships

- For use as fuels in rocket.

Note: Electronic Configuration of a neutral hydrogen atom is $1s^1$

NITROGEN AND PHOSPHORUS

Nitrogen — 2,5
Phosphorus — 2,8,5

Similarities

(i) They belong to group 5 of the periodic table

(ii) They gain 3 electrons to attain an octet configuration.

(iii) Majority of Compounds formed are Covalent.

DIFFERENCES

(I) They belong to different period.
N — Period 2
P = Period 3

Occurrences of Nitrogen:

① Nitrogen is a major Component of Protein.
② They are also available in the atmosphere (78%)
③ The Soil Contains Nitrogen in the form of nitrates and nitrites

PREPARATION OF NITROGEN

① Action of Ammonia on Copper (II) oxide

$$2NH_3 + 3CuO \rightarrow N_2 + 3H_2O + 3Cu$$

② By heating dioxonitrate (III)

$$NH_4NO_3 \xrightarrow{heat} N_{2(g)} + 2H_2O + O_{2(g)}$$

③ Pure N_2 is obtained by heating NaN_3 (Sodium azide)

$$2NaN_3 \xrightarrow{\Delta} 2Na + 3N_2$$

Phosphorus

It occurs as apatite which is a Combination of Ca, F, P. ie Calcium, fluorine, Phosphorous. It occurs as Rock phosphate. Bones Contain alot of phosphorus. It Can be obtained by extraction by heating a mixture of phosphate rock, Silica and Coke.

ALLOTROPES OF PHOSPHORUS

① White phosphorus
② Red "
③ Black

OXYGEN AND SULPHUR

Oxygen — 2, 6
Sulphure, 2, 8, 6

SIMILARITIES

- They belong to the Same group ie Group 6
- They gain 2 electrons to form an octet Configuration

DIFFERENCES

① Oxygen exists as diatomic Molecule while Sulphure do not.

② Oxygen belongs to period 2 while Sulfur is in period 3

OCCURRENSCE OF OXYGEN

- Oxygen occurs in the atmosphere by a Volume of 21%
- It is Constant as a result of PhotoSynthesis but Varies due to Pollution.

PREPARATION OF O_2

① Thermal decomposition of oxides

$$2Ag_2O \xrightarrow{\Delta} 4Ag + O_2$$

$$2HgO \xrightarrow{\Delta} 2Hg + O_2$$

PROPERTIES OF OXYGEN

① Oxygen is Colourless, odourless and Tasteless

② it's a diatomic gas

③ It's very reactive forming Compounds with all elements except noble gases.

ALLOTROPES OF OXYGEN

The major allotropes of Oxygen is O_3, The ozone helps to shield the earth from U.V. rays.

SULPHUR

Allotropes

① Rhombic Sulphur

② Mono Clinic "

Transition temp:- This is the temp. of Conversion from allotrope to another.

CHEMICAL PROPERTIES OF SULPHUR

① Sulphur reacts with most metals.

$$Fe + S \rightarrow FeS$$

$$Mg + S \rightarrow MgS$$

USES OF SULPHUR

① To Manufacture acids like H_2SO_4 (Contact Process)

② Manufacture of $CaSO_3$

③ " " $MgSO_3$

CaSO, $MgSO_4$ and H_2SO_4 are used in the bleeding of wood pulp.

- Sulfur is used in Sulphide dyes, Carbon di Sulphide (CS_2) Matches, Fireworks, gun powders Vulcanization of rubber, SO_2

Chlorine Bromine and Iodine

Similarities are

$Cl_2 \rightarrow 2,8,7 \rightarrow$ gas period 3

$Br_2 \rightarrow 2,8,18,7 \rightarrow$ liquid, period $\boxed{4}$

$I_2 \rightarrow 2,8,18,18,7 \rightarrow$ Solid, period $\boxed{5}$

They all belong to group 7 of the Periodic table.

They all gain an electron to form an octet Configuration.

Their molecules are diatomic

Chlorine is the most electronegative of the three, then bromine and lastly iodine.

SODIUM, CALCIUM AND ALUMINIUM

SIMILARITIES AND DIFFERENCES

Sodium, Calcium and Aluminium are metals and are found in Group IA, IIA and IIIA respectively of the periodic Table. Some data of the elements are tabulated below:

Atomic Properties	Na	Ca	Al
Atomic no.	11	20	13
Electronic Configuration	2,8,1 $[Ne]3s^1$	2,8,8,2 $[Ar]4s^1$	2,8,3 $[Ne]3s^23p^1$
first ionization energy (kJ/mol)	496	1735	579
Electronegativity	0.9	1.0	1.5
Atomic radius (nm)	0.191	0.197	0.143
Ionic radius	0.102	0.100	0.053
Melting Point	371	1115	932
Boiling point	1156	1757	2720

Most of the difference in the Physical and Chemical Properties of the 3 metals can easily be explained in terms of the different number of protons in the nuclei (atomic numbers) and the different number of Valence electrons (ascertained from the electronic Configurations) eg by losing electrons, the 3 elements form Cations Na^+, Ca^{2+}, and Al^{3+} respectively, and so their Compounds, especially those of Na and Ca are Strongly electrovalent

SODIUM

Sodium is metallic element found in the first group of the Periodic table. Sodium is an element that is a member of the alkali metals group with a

Symbol Na : it is extremely reactive and electropositive (prone to lose the outermost electrons)

Pure Sodium is never found in the free State/pure in nature because it is a highly reactive metal and exists in Combination with other elements or radicals as positive ions. eg Na^+Cl^-. Chemically, Sodium is Commonly found Combined with Chloride to form NaCl known to us as table salt.

It has one electron in the outermost electron shell and thus wants to give up one electron to a highly electronegative element. In chemical Combination, this single electron is readily transferred, giving a unipositive metal ion with the) stable electronic Configuration of a noble gas. eg $Na^+ (1s^2\ 2s^2\ 2p^6)$ is Isoelectronic with Neon.

It is physically Silver Colored, tarnishes easily and has a low melting point and density.

Sodium reacts exothermically with water, releasing heat when in Contact with water

The Chemical equation:

Sodium + water → Sodium hydroxide + hydrogen gas

$$2Na_{(s)} + 2H_2O_{(l)} \longrightarrow 2NaOH_{(aq)} + H_{2(g)}$$

OCCURRENCE

The Sodium ion is abundantly found within the earth's oceans, bodies of water and minerals. Sodium occurs as Na_2CO_3 or as $Na.HCO_3$. $NaHCO_3$ which is found in East Africa, Australia and California. Na_2SO_4 and borate ($Na_2B_4O_7$) are also sources of the metal.

EXTRACTION AND USES

Sodium is extracted by the electrolysis of fused Sodium Chloride to which a little Calcium-Chloride has been added.

Sodium is employed in the manufacture of Sodamide, Sodium hydroxide, Sodium peroxide, Sodium Cyanide etc

The Sodium ion is vital to animal life because it maintains body fluid Volumes and maintains electric potentials in animal tissue. The most Common Compound of Sodium, Sodium Chloride, is used in food for Seasoning and preservation. In Soap, Sodium is used as Sodium salts of fatty acids because it is harder and higher melting than other Soaps.

In medicine, the salt form of a medication with high Sodium or potassium ingredient can improve bioavailability: it is used for chemical synthesis, analysis and heat transfer applications. Sodium is also a crucial element for animal and plant life by creating charge gradients and assisting in the development of energy.

Calcium (Ca)

Group 2 contains soft metals that are more metallic in character compared to Group 1 and are silver in color. All of the elements in Group 2 have two electrons in their valence shell giving them an oxidation state of +2. This enables the metals to easily lose electrons which increase their stability and allow them to form compounds via ionic bonds.

The alkaline earth metals are different in their reactions compared to Group one Metals.

Hydrogen: All of the alkaline earth metals react with hydrogen to create metallic hydrides. Here is an example of a reaction.

$$Ca_{(s)} + H_{2(g)} \longrightarrow CaH_{2(s)}$$

The four stable isotopes of Calcium are: ^{40}Ca, ^{42}Ca, ^{43}Ca and ^{44}Ca. The most abundant isotope, ^{40}Ca, makes up about 97% of naturally occuring Calcium. Calcium has a melting point of 1115 K and gives off a red flame when ignited. It is very reactive though not as reactive as Na and it is thus never found free in nature. The Group 2A metals have 2 electrons in the outer most shell preceded by a closed shell containing eight electrons (except Be which has a closed shell of 2).

In Chemical Combination these 2 outer electrons are transferred giving a dipositive metal ion with a stable electronic configuration of a noble gas, eg Ca^{2+} which is isoelectronic with Argon (2,8,8).

Compounds formed by Ca are therefore predominantly ionic and exist as high melting points solids.

The first ionization energy is considerably higher than that of Na (Group 1). The group IIA metals have higher melting and boiling points and densities than corresponding metals in Group IA.

OCCURRENCE

The element, Ca, is one of the

alkali earth metals. In nature, it is only found in Combination with other elements.

Calcium occurs mainly as the Carbonate, $CaCO_3$. Other Sources are dolomite $CaCO_3, MgCO_3$, gypsum.

EXTRACTION AND USES

Calcium is extracted by the electrolysis of Calcium Chloride to which a little fluorspar has been added to lower the melting point of $CaCl_2$.

It is essential for living organisms. Calcium, with the presence of Vit. D is well known for its role in building stronger, denser bones early in the lives of humans and other animals. Calcium Can be found in products such as milk, Cheese and other dairy products.

Calcium is an Important Component in Cement and mortars and thus is necessary for Construction

ALUMINIUM

Al has three Valence electrons. It is a light metal possessing Considerable strength, yet is malleable and ductile, with a silver or gray Color. It is a very reactive element so, it is found in nature Combined with other elements.

One would think aluminium will react with water but in reality Aluminium is protected by a layer of Al_2O_3, which is known as anodizing. The thickness of the layer Can vary through galvanic reactions and protect it from further oxidizing. Aluminium Can dissolve in both acids and bases, making it an amphoteric molecule.

For example in an aqueous OH^- Solution it will produce $Al(OH)_4^-$ and in an aqueous H_3O^+ solution it will produce $[Al(H_2O)_6]^{3+}$.

Another important feature of aluminium is that it is a good reducing agent, because of +3 oxidation state. It Can therefore react with acids to reduce $H^+_{(aq)}$ to $H_{2(g)}$.

For example

$$2Al_{(s)} + 6H^+_{(aq)} \rightarrow 2Al^{3+}_{(aq)} + 3H_{2(g)}$$

Aluminium Can also extract oxygen from any metal oxide. This reaction is known as the thermite reaction which is very exothermic.

$$Fe_2O_{3(s)} + 2Al_{(s)} \longrightarrow Al_2O_3 + 2Fe_{(s)}$$

OCCURRENSCE

Aluminium is extracted from bauxite $Al_2O_3 \cdot 2H_2O$.

USES

Aluminium has many uses. Many alloys are made of aluminium to prevent Corrosion. Its alloys, duralumin (Al/Mg/Cu) and Magnalium (Al/Mg) are light and strong and are therefore used in the Construction of air Craft, ships, buses, tube trains etc. Aluminium is a good heat and electrical Conductor.

Aluminium foil is used for wrapping food items (Chocolates) and for making milk and bottle tops.

DEVELOPMENT OF ELECTRONIC CONFIGURATION

INTRODUCTION

There are more than 220 Subatomic particles discovered so far as Components of atoms. Three of these electron, neutron and proton are of much Interest to the chemists. The number of electrons which is the Same as the number of protons in an atom determines the atomic number whereas the Sum of the number of protons and neutrons determines the mass number.

It Could be said that the number of electrons and their arrangement in an atom determines the reactivity of the atom. This is the Singular reasons why elements in the Same group in the Periodic table have the Same electron architecture. Elements with Similar electron arrangement react in the Same way. The number of electrons Could determine how slow or fast they will react.

The arrangement of electrons in an atom of an element is Called the electron Configuration. For any atom, the four quantum numbers, Principal, Subsidiary, Magnetic and Spin quantum number describe the position of each electron interms of the shell, Subshell, orbital or Spin. The principal quantum numbers, n, describes the energy, the Subsidiary quantum number, L, describes the Shape while the magnetic quantum number, m, describes the orientation of the Volume the electron occupies in an atom. This volume is Called an orbital. The fourth quantum number, Called the electron Spin quantum number, S, describes these Spin orientation of the electron in an orbital. It assumes only 2 Values $+\frac{1}{2}$, or $-\frac{1}{2}$

The orbitals occur in definite energy levels in the atoms. Therefore, it is possible to describe each electron based on its orbital occupation and the energy level.

The process of arranging the electrons is based on the Aufbau process.

THE AUFBAU PROCESS (German building up process).

The Aufbau process discussed how that electron fill the lowest energy orbital first, and then moveup to higher energy orbitals only after the lower energy orbitals are full. However, there is a problem with this rule. Certainly, 1s orbital should be filled before 2s orbitals, because the 1s orbital have a lower Value of n and thus a lower energy.

What about the 3 different 2p orbitals? In what order Should they be filled? The answer to this question Involves Hund's rule.

The Aufbau's principle employees a Shorthand notation for writing electronic Configuration. The formular

$$nX^y$$

where
n = principal quantum no
X = is the Subshell (orbital)
y = the number of electrons in the Subshell is adopted

For example $2P^2$ implies that the energy level is 2, Subshell (orbital) is P and number of electron in the orbital is 2.

This arrangement is based on three Concept (n+l) rule, Hund's rule and Pauli exclusion Principle

(N+L) RULE - ENERGY OF ORBITALS

This rule states that the electrons will enter the subshell or orbital of lowest energy first. What this implies is that in the building of atoms, electrons are fed into atomic orbitals and orbitals of lowest energy are filled first.

The orbitals of lowest energy is the one with lowest (n+l) values (n) = principal quantum number and l = Subsidary or Azimuthal quantum Number.

If 2 orbitals have the same (n+l) value, the one of lower energy will be orbital of lower n value. The relative energies of the orbital will be in the order:

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s etc.

The (n+l) Values of Some of these orbitals is given in the table below:

(n+l) Values of Some orbitals

Orbitals	n	l	(n+l)
3s	3	0	3
3P	3	1	4
4S	4	0	4
4P	4	1	5
5S	5	0	5
3d	3	2	5
4d	4	2	6
5P	5	1	6
6S	6	0	6

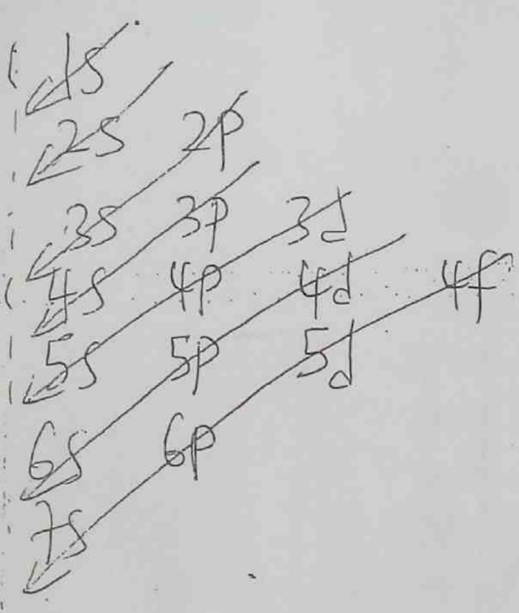
Note: that l values are 0, 1, 2, 3 for s, p, d, f orbitals respectively; and that from the table that 6S, 5P and 4d have the same (n+l) values.

However, to determine which orbitals will be filled first (we check out the orbitals that have lowest energy) using (n+l) rule. This means that we write the one with the lowest energy first (ie the one with lowest (n+l) value, then write the one that follows it ie in ascending order of their n+l values.

N:B If 2 orbitals have same (n+l) value, you must Write the one with lower value of n first Ok!!!

eg As seen in the table 3p and 4S have same (n+L) value ie (4), So since 3P has lower value of n ie 3, pls write it first before 4S.

From the table, we have the order 3S, 3P, 4S, 3d, 4P, 5S, 4d, 5p etc



This order also help to See how to write the Electronic Configuration

1S, 2S, 2P, 3S, 3P, 4S, 3d
4P, 5S, 4d, 5P, 6S, 4f
5d, 6P, 7S, etc

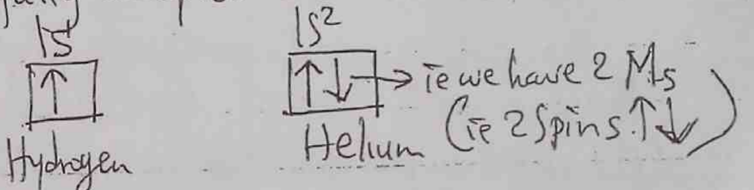
② PAULI EXCLUSION PRINCIPLE

P. W (ie Pauli extensively studied atomic Spectra. The Pauli's exclusion Principle states that, in an atom or molecule, no 2 electrons Can have the Same four quantum numbers, the Same

As an orbital Can Contain a maximum of only 2 electrons, The 2 electrons must have opposing Spins. This means that if one is assigned an up - Spin $(+\frac{1}{2})$, the other must be down Spin $(-\frac{1}{2})$ and they are Said to be paired.

Electrons in the Same orbital have the Same first 3 quantum numbers, eg $n=1$, $l=0$, $m=0$ for 1s Subshell, Only 2 electrons Can have these numbers So that their Spin moments must either be $+\frac{1}{2}$ or $-\frac{1}{2}$.

If 1S orbital Contain only one electron we have one MS (Spin moment) Value and the electronic Configuration written as $1S^1$ (Corresponding to Hydrogen). If $1S^2$ (Corresponding to helium, it is fully occupied we have 2 Ms ie



For He, $n=1$ for both e^-s, $l=0$ and $m=0$ for both electrons but $MS = +\frac{1}{2}$ and $-\frac{1}{2}$

As you Can See, 1s Subshell Can hold two electrons and when filled, the electrons have opposite Spin.

In any given shell, the total number of electrons is given by $2n^2$ where (n = Principal quantum number). All orbitals in each energy level have the Same energy and are Called degenerate. Since three quantum numbers, n, l and m are needed to define an orbital, Each orbital may hold up to two electrons

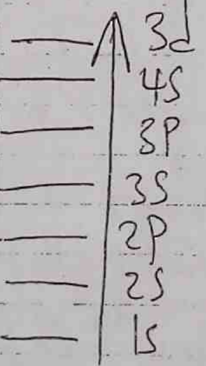
Provided they have opposite spins.

An extra quantum number is required to define the spin of an electron in an orbital. Thus four quantum numbers are needed to define the energy of an electron in an atom. The Pauli exclusion principle states that no 2 electrons in one atom can have all 4 quantum numbers the same

③ HUND'S RULE of MAXIMUM MULTIPLICITY

The rule was discovered by Friedrich Hund in 1925, is of important use in atomic chemistry, spectroscopy, and quantum chemistry. The rule is based on observation of atomic spectra, which is used to predict the ground state of an atom or molecule with one or more open electronic shells. The rule states that for a given electron configuration, the lowest energy term is the one with the greatest value of spin multiplicity. This implies that if 2 or more orbitals of equal energy are available, electrons will occupy them singly before filling them in pairs.

In atoms with more than one electron, the electrons, repel each other. The effective nuclear charge varies with the atomic number as the inner shell electrons screen the outer ones. The net result is only determined approximately. This approximation gives rise to similar orbital properties but different energies. As a result, the orbital energies are shifted as shown in fig 2.



Relative orbital energies

In line with these developments, Hund developed a rule that account for the fact that electrons are distributed as far as possible. The rule states that in a set of orbitals (subshells) of equal energy no 2 electrons will pair up until all the degenerate orbitals are single filled. In other words, suborbitals of identical energy are available, an atom tends to have as many unpaired electrons as possible. The electrons are arranged so as to give optimal number of unpaired electrons

For examples, when n is equal to 2, the 2 available orbitals (2S and 2P). The P orbital has 3 suborbitals (P_x, P_y, P_z) showing the orientation of the volume of the orbitals in an atom. Each of these orbital can accomodate only 2 electrons.

$\overset{P_z}{\boxed{\uparrow\downarrow}}$ $\overset{P_z}{\boxed{\uparrow\downarrow}}$ $\overset{P_z}{\boxed{\uparrow\downarrow}}$ Thus $2p^6$

However, for a $2p^3$ configurations hund's rule gives that the filling will be

$\boxed{\uparrow}$ $\boxed{\uparrow}$ $\boxed{\uparrow}$ $\Rightarrow P_x^1\ P_y^1\ P_z^1$

$2P = \boxed{\uparrow\downarrow}$ $\boxed{\uparrow}$ $\boxed{\uparrow}$ [Maximum multiplicity is maintained]

Rather than

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}$ $\boxed{\downarrow}$ (incorrect) or

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{}$ Incorrect

As a result of Hund's rule, constraints are placed on the way atomic orbitals are filled in the ground state using the Aufbau's principle. Before any 2 electrons occupy an orbital in a subshell, other orbitals in the same subshell must first each contain one electron. Also, the electrons filling a subshell will have parallel spin before the shell starts filling up with the opposite spin electrons (after the first orbital gains a second electrons). As a result, when filling up atomic orbitals, the maximium total spin state, assured.

In general, atoms acquires extra stability when degenerate orbitals are either half filled (i.all parallel spins) or completely filled spin paired.

Example as seen in noble gases - Ions with half filled configuration (p^3, d^5, f^7) or completely filled configuration (p^6, d^{10}, f^{14}) also show some degree enhanced stability.

example ① Nitrogen has atomic number of 7 ie $1s^2\ 2s^2\ 2p^3$

Chromium has atomic number $Z=24$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1,\ 3d^5$

Oxygen has atomic number of 8 $1s^2\ 2s^2\ 2p^4$

Remember that elemental Oxygen and Nitrogen are found in nature typically as dioxygen and dinitrogen respectively.

Summarily according to hund's rule of Maximum multiplicity:

The electrons tend to avoid

being in the same orbital thus as the electrons are successively added, a maximum number of electrons, will try to occupy orbitals singly.

When all the orbitals are singly occupied, only then the pairing of electrons commences.

In the ground state, the electrons occupying the orbitals singly will have their spin parallel.

Hund's rule indicates that electronic arrangement $\boxed{\uparrow}$ $\boxed{\uparrow}$ is more stable than $\boxed{\uparrow\downarrow}$ $\boxed{}$

$1s^2, 2s^2, 2p_x^1, 2p_y^1, 2p_z^1$ for Nitrogen ie atomic number = 7 (Correct)

$1s^2, 2s^2\ 2p_x^2, 2p_y^1, 2p_z^0$ (Incorrect)

THE ELECTRONIC CONFIGURATION OF IONS

The only difference between writing a normal electron configuration and writing, the electron configuration for an ion is that when writing it you have to remember to add or subtract electrons from your total electrons. That is, if the element is an anion (negative charge), you must add electrons to your total electrons.

$S + 2e \rightarrow S^{2-}$

$16e^-$ $18e^-$

Electronic Configuration

$S^{2-} = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$.

$S = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

If the element is a cation (ie positively charged), you must add electrons to your total electrons.

$Al \rightarrow Al^{3+} + 3e^-$

$1s^2, 2s^2, 2p^6$ for Al^{3+} (10electrons)

$1s^2\ 2s^2, 2p^6, 3s^2, 3p^1$ for Al (ie 13 electrons)

In determining the electronic configuration of ions, the (n+l) rule will still be followed. In filling the electrons, it should be noted that electrons are first lost from the orbital with the highest n value. Where many orbitals have the same n value, electron is lost first from the one with higher l, value.

5) Arrange the following orbitals in the order of increase stability:

6s, 5p, 4f, 7s, 4d.

Solution: the (n+l) rule is used to determine the relative energy of the orbitals.

ORBITALS	n	l	$n+l$
6s	6	0	6
5p	5	1	6
4f	4	3	7
7s	7	0	7
4d	4	2	6

N.B The orbital of lowest $(n+l)$ Value is the orbital of lowest energy and therefore the most stable.

Where 2 orbital have the same $(n+l)$ value, the more stable orbital is the orbital with the lower n value

② What is the maximum number of electronic orbitals with the following set of quantum number will possess and why?

ⓐ $n=2$, $l=1$

Answer

when $l=1$ and $n=2$; $m=-1, 0, +1$ (3 values of m. for each value of m there are two values of spin quantum number $(+\frac{1}{2}$ or $-\frac{1}{2})$

Therefore the orbital will occupy a maximum of 3×2 electrons $= \underline{6}$ (one electron for $\frac{1}{2}$ and another for $-\frac{1}{2}$)

N.B This was how the lecturer solved it.

Short Cut

Use the formula, maximum no of electrons $= 2(2l+1)$

Since $l=1$ $\Rightarrow 2(2(1)+1)$

$= \underline{\underline{6 \text{ electrons}}}$

ⓑ $n=3$, $l=3$

When $l=3$, This case is not possible since the value of l cannot be greater that $(n-1)$. This implies that l cannot be greater than n or equal to n.

ⓒ $n=3$, $l=2$

Here $m = -2, -1, 0, 1, 0, 2 \Rightarrow 5$ orbitals

Hence $5 \times 2 \Rightarrow 10$ electrons

N.B This 2 we multiply with was gotten from the fact that each □ must have maximum of 2 electrons.

OR $2(2l+1)$

$2((2 \times 2)+1) = 10$ electrons

ⓓ When $n=4$, $l=3$

Maximum no of electrons

$= 2(2(3)+1) = 14$ electrons.

Another question

Determine the maximum number of electrons in an orbital described by the following quantum numbers and write the appropriate orbital filling:

① $n=2$, $l=0$ ② $n=4$, $l=3$
③ $n=3$, $l=2$

Solution

① $n=2$, $l=0$: The Value of m is only one. This implies maximum of 2 electrons where $l=0$, it represents an S-orbitals since $n=2$ the orbital occupation is $2S^2$.

② $n=4$, $l=2$
Since $l=2$; $m=-2,-1,0,1,2$ (five values): This means maximum of ten electrons. Since $n=4$, orbital filling will be $4d^{10}$.

③ $n=3$, $l=2$ where $l=2$, $m=-2,-1,0,1,2$ (5 Values), hence maximum of 10 electrons. Since $n=3$, orbital filling will be $3d^{10}$

CONSEQUENCES OF ELECTRONIC CONFIGURATION

Electron Configuration of an atom of an element Could be in pointer to

1. its reactivity
2. Nature of bond it form with other element
3. The magnetic Properties,
4. Enhanced stability of the atom or ion
5. Colour of ions and Compounds.

DISCUSSION

① Reactivity

Free radicals that have unpaired electrons are known to be Very reactive. This is as a result of Instability in that State. When it reacts it becomes stable when the electrons are paired up in orbital

$$Cl^- + Cl^- \xrightarrow{fast} Cl_2 \text{---}$$

Group 1 elements have S^1 electron Configuration in their outermost shell. They will easily transfer this Single electron to another atom during reaction in order to attain outer P^6 which is more stable example:

Na = (atomic number = 11)
$1S^2, 2S^2, 2P^6, 3S^1$ will have S^1 to maintain more stable P^6.

TYPES OF BONDING

Highly electropositive elements have such electron Configuration that allows them lose electrons in order to achieve a measure of Stability.

When they lose electrons, they become positively charged. In this state, They Can form ionic bonds with other atoms that accepted the electrons

Na $(1s^2 2s^2 2p^6\ 2s^1)$

Cl $(1s^2 2s^2 2p^6 3s^2 3p^5)$

$$Na + \frac{1}{2}Cl_2 \longrightarrow Na^+Cl^- \rightarrow \text{Salt (ionic bond)}$$

Na^+: $1s^2 2s^2 2p^6$ Cl^-: $1s^2 2s^2 2p^6 3s^2$ $3p^6$

The reaction ② Shows that Cl^- with electronic Configuration of p^5 will easily accept electron to become Cl^- were to react, the most likely pathway is to Share the electrons in order to have $3s^2\ 3p^6$ (octet). This also ensures stability

If 2 atoms of Cl were to react, the most likely pathway is to share the electron in order to have $3s^2\ 3p^6$ (octet) Configuration.

MAGNETIC PROPERTIES.

Any Substance can exhibit any of three forms of magnetism; Paramagnetism diamagnetism and ferromagnetism, Presence of unpaired electrons in the electronic Configuration of an element leads to paramagnetism. This implies that it is attracted strongly in a magnetic field.

However, where paired electrons feature in the electronic Configuration of an element, it results in diamagnetism, This is a feature indicating that the substance is weakly repelled in a magnetic field.

Paramagnetism overshadows diamagnetism when both occur in an element.

All Compounds exhibit diamagnetism Properties because of presence of paired electrons in them. This is because the Process of Compounds formation results from bonding between atoms. Bonding is either by electron transfer or electron sharing.

Both Processes lead to electron sharing (having paired electrons in all orbitals). However, ions and atoms Could have paired electrons.

The extent of paramagnetism is given by the magnetic moment μ

$$\mu = \sqrt{n(n+2)}\ BM.$$

where BM = Bohr Magneton (the unit)

Example: find the Magnetic Moment of ① Cu^{2+} (ii) Be

Solution

Solution

Cu^{2+} has 1 unpair electron. ie Cu has 29 electrons, Cu^{2+} = 27 electrons Since it has lost 2 electrons.

For Cu

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$

$\boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow}\ \boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow}\ \boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow|\uparrow}$

ie it has one upair electron $\boxed{\uparrow}$ (the last orbital)

$\Rightarrow n = 1$ (no of unpaired electrons)

Magnetic Moment $\Rightarrow \sqrt{n(n+2)}$

$\Rightarrow \sqrt{1(1+2)}$

$= \underline{\sqrt{3}} = 1.73$ BM (paramagnetic)

(ii) For Be $\Rightarrow 1s^2 2s^2$

$\boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow}$ (no unpaired electrons)

$\Rightarrow n = 0$

$\mu = \sqrt{0(0+2)}$

$= 0$ (Diamagnetic)

Cooperative interations of Unpaired electrons of individual Paramagnetic atoms with one another gives rise to very strong attraction of some substances to a magnetic field, thus this phenomenon is known as Ferromagnetism - Examples are Fe, Nickel, Cobalt exhibit ferromagnetism.

Their magnetic Susceptibility is greatly enhanced Compared with what it would be if all the moments behaved independently.

Ferromagnetism features in many of the transition metals and their Compounds: Fe, Ni, Co Can form Permanent magnets.

Anti ferromagnetism results when the moments in adjacent atoms are paired so that they point in opposite directions. This is some sort of destructive interaction of unpaired electrons of Paramagnetic atoms. In this case, the substance is not attracted to a magnetic field as expected.

ENHANCED STABILITY OF ATOM OR ION:

Certain Configuration result in enhanced stability of the atom or ion. It had earlier been stated that electron configuration bearing halffilled or completely filled shell result in enhanced stability

$(4s^0, 3d^5)$

For example, Fe^{3+} has d^5 Configuration

($\frac{1}{2}$ filled) but Fe^{2+} (d^6) ie $4s^2 3d^6$ is not half filled. Therefore Fe(III) is more stable than Fe(II).

Where as Fe(II) is readily oxidized to Fe(III), Fe(III) is not reduced reduced to Fe(II)

Zn(II) is d^{10} configuration (Completely filled) and very stable it is difficult to oxidize it to Zn(II) ie d^9. The first ionization energies of the noble gases He, Ne, Ar, Xe and Rn are the highest in their respective periods This is because they have completely filled shells that give them enhanced stability that it is difficult to remove the electrons.

COLOR OF SUBSTANCE

Certain substances show different colours - for instance, transition metals and their compounds are coloured. Colours arise from absorption of visible light by substances. The colors that are seen are the colours seen are the emitted radiation.

Absorption of light leads to the promotion of electrons from a ground state to an excited state, when the absorbed radiation is emitted, the electrons return to the ground state.

Most substances that have unpaired electrons show colours because orbitals are available for electron transition.

Examples: Colours of transition metal ions.

Fe(II)	d^6	light blue
Fe(III)	d^5	brown
Cu(II)	d^9	blue
Co(II)	d^7	Pink
Zn(II)	d^{10}	Colourless
Cd(II)	d^{10}	Colourless
Mn(II)	d^5	Pink

PROPERTIES OF TRANSITION ELEMENTS

- They form coloured ions
- They have variable oxidation states
- form complex ions
- Acts as catalyst
- Are all metals
- Have similar ionization energies
- form alloys
- Exhibits paramagnetism

HUND'S RULE CONTD

When atoms are in their ground states, the electrons occupy the lowest possible energy levels.

The Simplest element, hydrogen has one electron, which occupies the (1s) level; this level has the principal quantum number $n=1$ and the Subsidiary number $l=0$.

Helium has 2 electrons, the second electron also occupies the 1s level. This is possible because the 2 electrons have opposite spins. This level is now full. The next atom lithium, has three electrons. The 3rd electron occupies the next lowest level. This is 2s level, which has the principal quantum number $n=2$ and Sub sidiary quantum number $l=0$.

The electronic structures of the first few atoms in the periodic table may be written as

$H = 1s^1$
$He = 1s^2$
$Li = 1s^2\ 2s^1$
$Be = 1s^2\ 2s^2$
$B = 1s^2, 2s^2, 2p^1$
$C = 1s^2, 2s^2, 2p^2$
$N = 1s^2, 2s^2, 2p^3$
$O = 1s^2, 2s^2, 2p^4$
$F = 1s^2, 2s^2, 2p^5$
$Ne = 1s^2, 2s^2, 2p^6$
$Na = 1s^2, 2s^2, 2p^6$

An alternative way of showing the electronic Configuration (structure) of an atom is to draw boxes for orbitals and arrows for electrons.

Eg.

ⓐ Electronic Structure of H atom in ground

$\Rightarrow 1s^1$

$\Rightarrow \boxed{\uparrow}$

ⓑ Electronic Structure of He atom in ground state

$\Rightarrow 1s^2$

$\Rightarrow \boxed{\uparrow\downarrow}$

ⓒ Electronic Structure of Chlorine in ground state

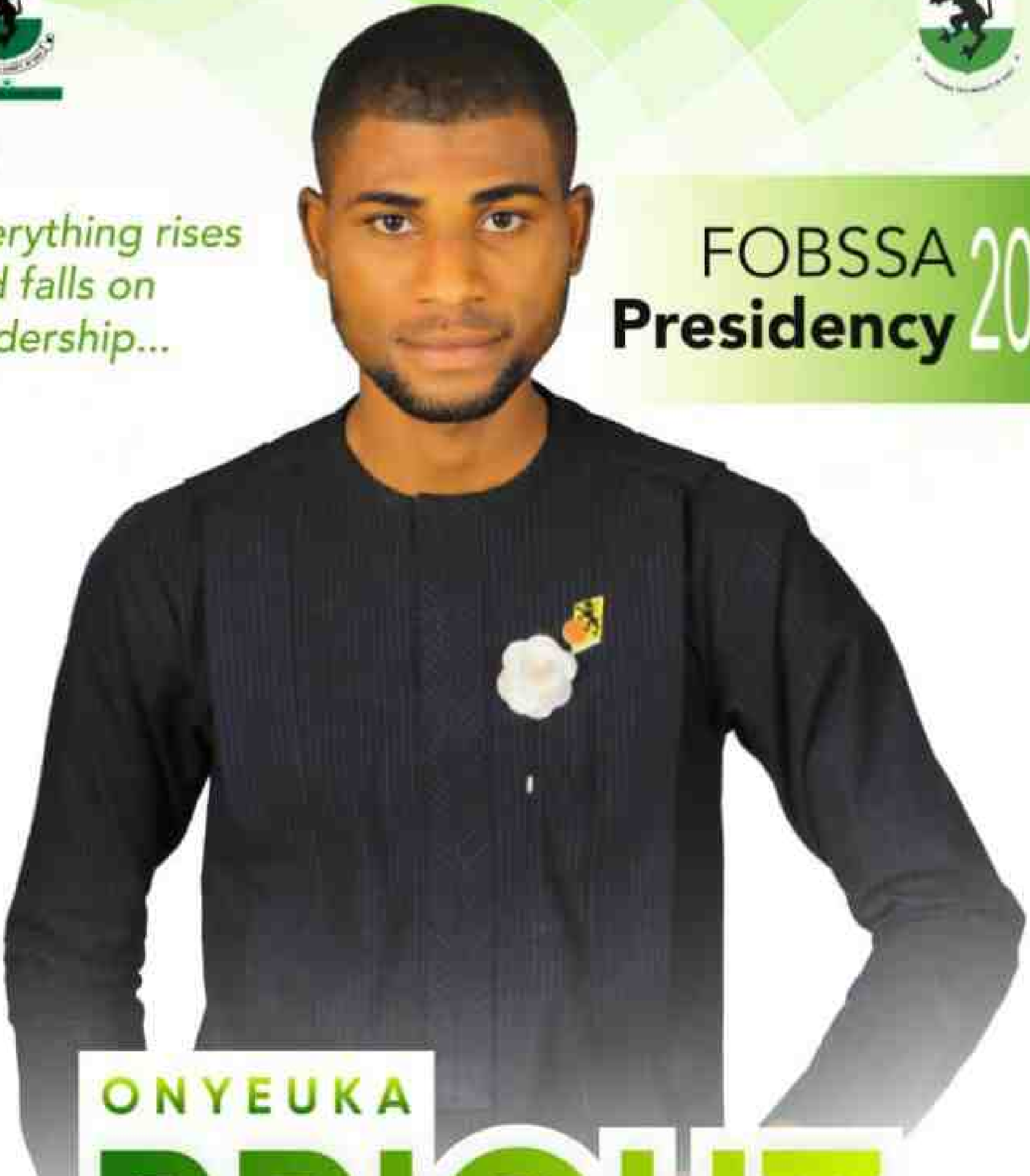
$\Rightarrow 1s^2 2s^2 2p^6 3s^2 3p^5$

$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^5$
$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}$	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}$

IMPORTANT POINT TO NOTE FOR PRACTICAL

Everything rises and falls on leadership...

FOBSSA Presidency 2020



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