

Sri Venkateshwara Vidhyalayaa Higher Secondary School,**Thasampalayam, Gobi – 638 476.****Class : XI-Bio/CS****Chemistry****Unit – I****BASIC CONCEPT OF CHEMISTRY AND CHEMICAL CALCULATIONS****1. Define Matter.**

The matter is defined as anything that has mass and occupies space.

2. Define Element (or) atom.

We know that an atom is smallest electrical neutral particle.

* Contains fundamental particles.

* Electrons

* Protons

* Neutrons

3. Define compound.

* Made up of molecules

* Contains two (or) more atoms of different element.

Ex : CO₂, C₆H₁₂O₃, NaCl

4. Define atomic mass.

* One twelfth of the mass of carbon – 12 atom in its ground state.

* 1amu (or) 1u $\approx 1.6605 \times 10^{-27}$ kg

* Ratio of the average atomic mass factor to unified atomic mass unit.

5. Define relative atomic mass.

* Relation atomic mass = $\frac{\text{Average mass of the atom}}{\text{Unified atomic mass}}$.

6. Define molecular mass

* Ratio of the mass of a molecule to the unified atomic mass unit.

7. Define mole.

* Collection of 6.022×10^{23} particles (or) ions (or) atoms (or) molecules.

* The amount of substance contains many elementary particles.

* As there are atoms in 12g of carbon 12 isotope.

$$\text{mole} = \frac{\text{mass}}{\text{molar mass}}$$

8. Define Avogadro number.

* The total number of ions or molecules or atoms in one mole of any substance equal to 6.00×10^{23} .

9. Define molar mass.

$$\text{Molar mass} = \frac{\text{mass}}{\text{mole}}.$$

10. Define gram equivalent mass.

$$\text{Gram equivalent mass} = \frac{\text{Molar Mass (gmol}^{-1}\text{)}}{\text{Equivalence Factor (eqmol}^{-1}\text{)}}$$

Equivalence factors Basicity of acids (or) acidity of base (or) e^- gain (or) e^- loss.

11. Define empirical formula.

The simplest ratio of the number of different atoms present in one molecule of the compound as subscript to atomic symbol.

Ex: The ratio of C : H : O = 1 : 2 : 1 and hence empirical formula CH_2O .

12. Define molecular formula.

The actual number of different atoms present in one molecule as subscript to atomic symbol.

Ex: The molecular formula acetic acid $\text{C}_2\text{H}_4\text{O}_2$ (CH_3COOH).

13. Define whole number.

$$\text{Whole number (n)} = \frac{\text{Molar mass of compound}}{\text{Calculated empirical formula mass}}.$$

14. Define stoichiometry.

It is the numerical relationship between chemical reactants and products is called stoichiometry.

* quantitative relationship.

15. Define limiting reagent and Define excess reagent.

* When the reaction is carried out using non – stoichiometric quantities of reactant and products.

* The product yield will be determined by reactant that is completely consumed.

* It limits the further reaction from taking place.

The other reagent which are excess is called excess reagent.

16. Gram equivalent mass:-

* Gram equivalent mass of an element, compound ion is mass that combines (or) displace 1.00 (or) 8g O_2 (or) 35.5g Cl_2 .

17. Define oxidation.

Oxidation is addition of oxygen (or) removal of Hydrogen.

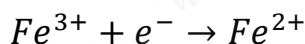
Electron concept.

Loss one (or) more electrons.

**18. Define Reduction.**

Reduction is the addition of H_2 removal of O_2 .

Electron concept gain of one (or) more e^- .

**19. Redox reaction.**

The oxidation reaction are accompanied by reduction reaction and vice versa called redox reaction $\text{H}_2 + \text{S}_2 \rightarrow \text{H}_2\text{S}$

20. Define oxidation number.

* It is imaginary charge.

* It caused by when all the atom are removed from the molecule removed as ions.

* May it be negative, positive, zero.

21. Define redox reaction in terms of oxidation number.

* During redox reactions the oxidation number of element changes.

* Oxidation number increase is called oxidation.

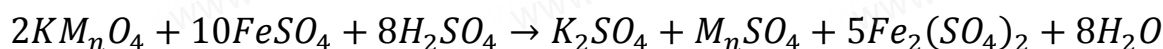
* Oxidation number decrease is called reduction.

**22. Define Reducing agent.**

The species which undergo the loss of electrons during the reaction are called reducing agent.

23. Define oxidizing agent.

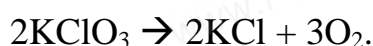
The species which undergo the gain of electrons during the reaction are called oxidizing agent oxidation number increase is called oxidation.

**24. Define decomposition reaction.**

* Redox reaction in which compound breaks down into two (or) more compound.

* These reaction are opposite to combination reactions.

* Oxidation number of elements of same substance changed.

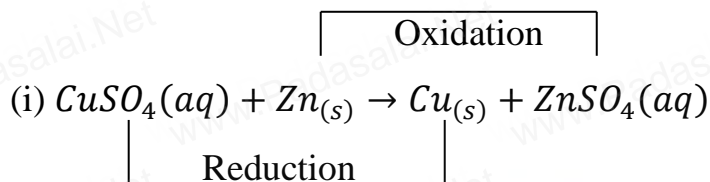


25. Define displacement reactions.

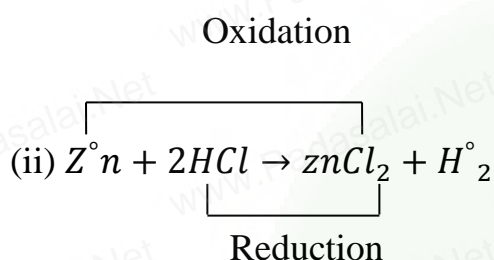
Redox reaction in which an ion (or atom) in compound is replaced by an ion (or atom) of another element called displacement reaction.

26. Types of displacement reactions.

- * Metal displacement reactions.
- * Non – metal displacement reaction.



Here Zinc metal replaced on copper.

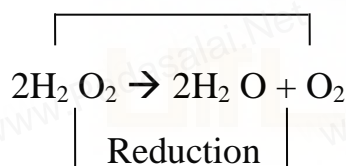
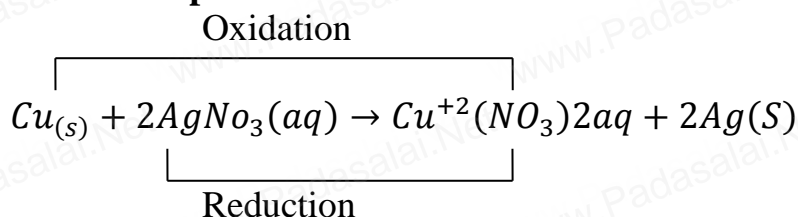
**27. Define disproportionation reaction.**

(or)

Define Auto redox reaction.

- * In some redox reactions.
- * The same compound can undergo both oxidation & reduction.
- * The oxidation state of one and same element is both increased and

Decreased oxidation

**28. Define competitive electron transfer reaction.**

- * In the above reaction between CU & Ag takes place.
- * CU has tendency to release the electrons.
- * Ag has tendency to accept the electrons.

Definition:-

* One atom electrons and another atom accept the electrons called compatibly electrons transfer reaction.

* Write the order of e^- releasing order of metal.

Zine > copper > silver

29. Define Balanced equation.

* The number & kinds of molecules present on both sides are qual in chemical reaction called balanced equation.

30. Define Atomic mass.

* Mass of single atom.

* Collective mass of electron, proton neuron.

31. Molar volume:

* The volume occupied by any one mole of by substance in gaseous state at given T.P.

32. Equivalent mass of Acid.

$$E = \frac{\text{Molar mass of acid}}{\text{Bosidty of acid}}$$

33. Equivalent mass of Base.

$$E = \frac{\text{Molar mass of reducing agent}}{\text{No.of mole of electrons lossed by one mole of oxidising agent}}$$

34. Equivalent mass of oxidising agent:

$$E = \frac{\text{Molar mass of oxidising agent}}{\text{No.of mole of electrons gained by one mole of oxidising agent}}$$

2. QUANTUM MECHANICAL MODEL OF ATOM

1. How many protons and neutrons are present in $^{18}_8\text{O}$?

Atomic Number	-	8
Mass number	-	18
Number of protons	-	Atomic number = 8
Number of protons + Number of neutrons	-	Mass number
8 + Number of neutrons	-	18
Number of neutrons	-	$18 - 8 = 10$

2. An atomic orbital has $n = 3$. What are the possible values of l ?

Atomic orbital ' n ' value is 3.

For a given ' n ' the possible ' l ' value are 0 to $(n-1)$

For $n = 3$, the possible l values are 0, 1 and 2.

3. An atomic orbital has $l = 3$. What are the possible values of m ?

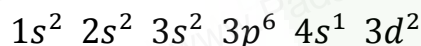
Atomic orbital ' l ' value is 3.

For a given ' l ', the possible ' m ' value are $-l$ through 0 to $+l$

For $l = 3$, the possible m values are $-3, -2, -1, 0, +1, +2, +3$.

4. Give the electronic configuration of chromium. ($Z = 24$).

Electronic configuration of chromium is ($Z = 24$)



5. An atom of an element has 19 electrons. What is the total number of p-orbital?

The number of electrons in an atom of an element is 19.

The electronic configuration is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

The total number of P -orbitals for the element is in $2P$ level 3 and $3P$ level 3.

i.e. '6' orbitals.

6. What is shape of the orbital with (i) $n = 2$ and $l = 0$; (ii) $n = 2$ and $l = 1$?

(i) $n = 2$ and $l = 0$, The orbital is $2s$. Its shape is symmetrical sphere.

(ii) $n = 2$ and $l = 1$, The orbital is $2p$. Its shape is dumb bell.

7. Give the electronic configuration of Mn^{2+} and Cu. Atomic number of Cu = 29 and Mn = 25.

Copper : Atomic number is 29.

Electronic configuration is $1s^2 2s^2 2p^6 3s^2 4s^2 3d^5$

Mn^{2+} - Electronic configuration is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^0 3d^5$

8. **Explain why the electronic configuration of Cr and Cu are written as $3d^5 4s^1$ and $3d^{10} 4s^1$ instead of $3d^4 4s^2$ and $3d^9 4s^2$?**

Chromium electronic configuration is $3d^5 4s^1$ instead of $3d^4 4s^2$ and copper electronic configuration is $3d^{10} 4s^1$ instead of $3d^9 4s^2$. Because half filled and completely filled electron configuration have symmetrical distribution of electrons and this symmetry leads to stability, configuration have symmetrical distribution of electrons and this symmetry leads to stability.

9. **What are the drawbacks of Thomson model of an atoms?**

Thomson's model of atom could account the electrical neutrality of atom, but it could not explain the results of gold foil scattering experiment carried out by Rutherford.

10. **State Heisenberg's uncertainty principle.**

Heisenberg's Uncertainty Principle states that it is impossible to determine simultaneously with certainty the position and the momentum of a particle.

11. **What is Zeeman effect?**

If a substance which gives a line emission spectrum is placed in an external electric field, its lines get split into a number of closely spaced lines. This phenomenon is known as Stark effect.

12. **What is the total number of orbitals associated with the principal quantum number $n = 3$?**

For $n = 3$,

For $n = 3$, the possible values of l are 0, 1 and 2. Thus, there is one 3s orbital ($n = 3, l = 0$ and $m, = 0$); there are three P orbitals ($n = 3, l = 1$ and $m, = -1, 0, 1$) there are five 3d orbitals ($n = 3, l = 2, m, = -2, -1, 0, 1, 2$).

Therefore, the total number of orbitals is $1 + 3 + 5 = 9$.

13. **Using s, p, d, f notations, describe the orbital with the following quantum numbers**

(a) $n = 2, l = 1$ (b) $n = 4, l = 0$ (c) $n = 5, l = 3$ (d) $n = 3, l = 2$.

	n	l	orbital
(a)	2	1	2p
(b)	4	0	4s
(c)	5	3	5f
(d)	3	2	3d

14. **What is $(n + l)$ rule?**

The lower the value of $(n + l)$ for an orbital, the lower is its energy. If two orbitals have the same $(n + l)$ value, the orbital with lower value of n has the lower energy.

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UNIT 3. PERIODIC CLASSIFICATION OF ELEMENTS

1. Define triads.

The atomic weight of middle element nearly equal to the arithmetic mean of atomic weights of the remain two elements.

Li, Na, K – Atomic weight 23. $\frac{Li(7)+K(39)}{2} = 23.$

2. Law of octaves.

On arranging the elements increasing order of atomic weights, he observed that the properties of every eighth elements are similar to the properties of first element.

3. Define periodic law.

“The properties of the elements are the periodic function of their atomic weights”.

4. Define the Modern periodic law.

“The properties of the elements are the periodic function of their atomic number”

5. Define group and columns.

Horizontal rows in the periodic table called – periods vertical columns in the periodic table called – groups.

6. Define Lanthanides.

* The filling up of 4f orbitals. / Ce-58 and ends with Lu – 71 - / first inner transition series are called Lanthanides.

7. Define actinides.

* The filling 5f orbitals
 * AC (89) to Lr (103)
 * Second inner transition series

8. Define Atomic radius.

* Distance between the centre of its nucleus and outer most shell containing valance electron.

9. Define Covalent radius.

* One – half of the internuclear distance between two identical atoms linked together by single covalent bond.

* $r_{cl} = \frac{d_{A-A}}{2}$

10. Define Metallic radius.

* One – half of the distance between two adjacent metal atoms in the closely packed metallic crystal Lattice.

11. Define effective nuclear charge.

* The force of attraction $Z_{\text{eff}} = Z - S$

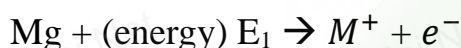
* Between nucleus and outermost electron.

12. Define – shielding effect.

* The force of attraction between nucleus and outer most electron is shielded by inner electron called shielding effect.

13. Define Ionisation energy.

The amount of energy is required to remove electron from an Isolated gaseous atom.



Increase in period.

Decrease in down a group

14. Define electron affinity.

The amount of energy released when adding electron to gaseous atom.



Increase in period

Decrease in the group.

15. Why Be and N electro affinity zero

Be – 4 ($1S^2 2S^2$) – completely filled

N – 7 ($1S^2 2S^2 2P^3$) – half filled

16. Why the Noble gas zero electron affinity.

* Electronic configuration $ns^2 np^6$

* Completely filled so electron affinity is zero.

17. Why the halogen has high electron affinity.

* Electronic configuration $ns^2 np^5$.

* It require one e^- to get stable electronic configuration.

18. Define electro negativity.

The relative tendency of atom attract the shared pair of electrons towards itself

Increased in period

Decreased in group (no unit)

19. Define oxidation state.

- * Comparing capacity relative to H atom.
- * Equal to the valance shell electron.
- * Equal to eight minus the number of valance electron.

20. Define periodicity.

- * The repetition of physical and chemical properties at regular intervals called periodicity.

21. Define Iso electronic.

The Number of electrons are same for elements by loss or gain of electrons.

Example : Na^+ and F^- both have same electrons (10). Na^+ loss 1 electron F^- gain 1 electron.

22. What are factors affecting electron affinity.

- * Size of atom
- * Nuclear charge
- * Electronic configuration
- * Shielding effect

23. What are factors affecting Ionisation energy.

- * Size of atom
- * Nuclear charge
- * Electronic configuration
- * Screening effect
- * Shape of orbital

24. Why the left side of element and right side of elements are more reactive compared to middle element periodic table?

Left side elements in periodic table \rightarrow easily lose their electron and less Ionisation energy.

Right side elements in periodic table \rightarrow easily accept electron to get stable (electron affinity high)

25. Why the Noble gas are inert?

- * It has completely filled electronic configuration.
- * So it neither accept nor lose e^-

26. Correlate the ionization energy and metallic character

Less ionization energy – metallic character

High Ionisation energy → non – metallic character

27. Define “Diagonal relationship”.

The similarity in properties existing between diagonally placed element is called “Diagonal relation ship.

Second period	Li	Be	B	C
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Third period	Na	Mg	Al	Si
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Li & Mg has similar properties.

28. Which are orbitals are more stable.

Completely filled and half filled orbitals are stable.

Order:-

Completely filled > Half filled > Partially filled.

29. Affbau principle.

In the ground state of atom, the orbitals are filled in order of their increasing energies.

LITLIT

4. Hydrogen

1. Explain why hydrogen is not placed with the halogen in the periodic table.

- * Hydrogen has less electron affinity.
- * Hence it is less reactive than Halogen.
- * Hydrogen has +1 oxidation state.

2. What are similarity of Hydrogen with alkali metals?

- * Both form unipositive ions (Na^+ , Li^+) (4+)
- * Both form halides (HX) NaX)
- * Both form oxides (H_2O_2) (H_2O) (Na_2O_2 , Na_2O)
- * Both form sulphides (H_2S) Na_2S)
- * Reducing Agent.

3. Define Isotopes of Hydrogen.

- * Same atomic number different mass number called isotopes.

Protium :- ${}_1\text{H}^1$ (99.985%) no neutrons.

Deutrium:- ${}_1\text{H}^2$ (0.015%) heavy hydrogen.

Tritium :- ${}_1\text{H}^3$ (~1 atom per 10^{18} H atom) radio active

4. Define ortho and para hydrogen.

Ortho Hydrogen:- In molecular hydrogen spin of hydrogen nuclei in same direction.

Para Hydrogen:- In molecular hydrogen spin of hydrogen nuclei in different direction.

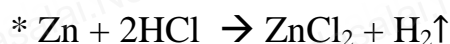
5. How will you prepare the H_2 by electrolysis.

- * Electrolysis of aqua NaOH (or) KOH
- * Using Nickel anode and Iron cathode

At Anode : $2\text{OH}^- \rightarrow \text{H}_2\text{O} + \frac{1}{2}\text{O}_2 + 2e^-$

At cathode : $2\text{H}_2\text{O} + 2e^- \rightarrow 2\text{OH}^- + \text{H}_2$

Overall the reaction : $\text{H}_2\text{O} \rightarrow \text{H}_2 + \frac{1}{2}\text{O}_2$

6. Give the laboratory preparation of Hydrogen**7. Give the industrial production of H₂.**

* Hydrocarbon (methane) mixed with steam

* In the presence of Ni catalyst

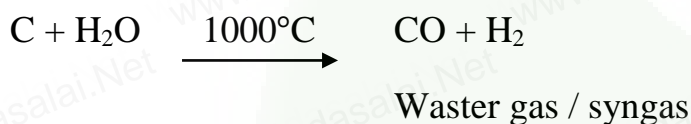
* 800 - 900°C and 35 atm

**8. How will produce water gas (or) syngas (or) synthetic gas**

* Steam passed over red – hot coke

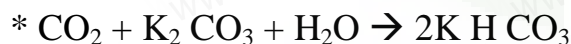
* Produce CO and H₂O.

* Used to prepare organic compound methanol.

**9. How will you remove CO from in water gas or how will you convert the CO into CO₂ in water gas.**

* Mixing of gas mixture with steam at 400°C. (shift Converter)

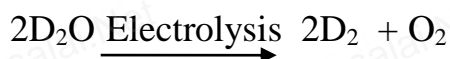
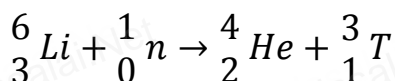
* Catalyst Fe/Cu and potassium carbonate used to absorbed CO₂.

**10. Preparation of Deutrium.**

* Normal water contain 1.6x10⁻⁴% of Heavy Water

* Dissociation = H₂O > D₂O

(protium water > Heavy water)

**11. Transmutation reaction.**

12. Properties of Hydrogen

* Colour less, odorless, tasteless

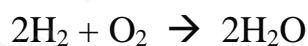
* Non polar * Highly flammable. Reducing Agent.

13. How will you convert the Hydrogen gas to liquid Hydrogen

* By applying high pressure and low temperature

14. Chemical properties of Hydrogen.

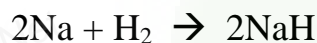
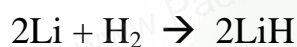
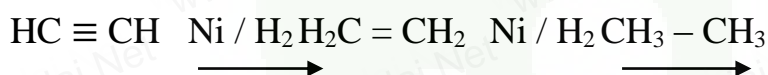
With oxygen



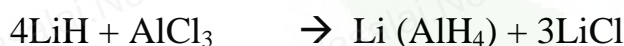
With halogen



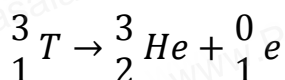
With metal

**15. Why H₂ is reducing agent.**

Unsaturated Saturated.

16. How will you prepare metal Hydrides.**17. Exchange (or) substitution reaction of Deuterium.**

Deuterium replace Hydrogen in a compound.

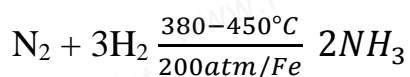
**18. Properties of tritium**

Half life period 12.3 year,

β emitter

19. Uses of Hydrogen

Haber Process



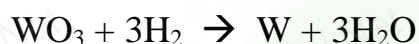
* Used to prepare ammonia.

* Ammonia used for preparation of HNO_3 fertilizer and explosives.

* It is used to prepare industrial solvent $\text{CO} + 2\text{H}_2 \xrightarrow{\text{Cu}} \text{CH}_3\text{OH}$

* Unsaturated fatty oils \rightarrow saturated fat by Pt / H_2

* Metal oxides are converted into metals at high temperature.



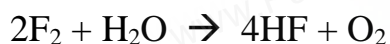
* Rocket fuel, used in rechargeable battery. Used to Generating Electrical Energy

* Atomic Hydrogen and oxy hydrogen torches used for cutting and welding.

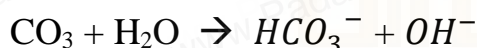
20. Chemical properties of water



21. Reaction of water with F_2



22. Write the reaction of water with non metal.

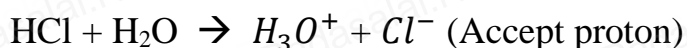
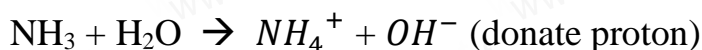


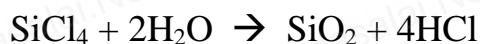
23. Why water is amphoteric oxide?

* Water is act as acid as well as base

by accept proton (acid)

by donate proton (Base)



24. Write reaction of water with covalent compounds.**25. Define Hard water.**

* Water containing soluble salts of Bicarbonates, chlorides and sulphate it is called hard water.

* This Properties known as Hardness.

26. Define soft water.

* Water free from soluble salts of Ca, Mg types of Hardness it is called soft water.

27. Types of Hardness

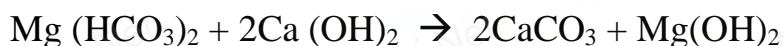
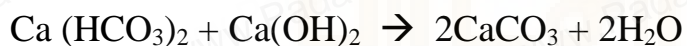
* Temporary hardness (Bicarbonates salts of (Ca, mg)

* Permanent hardness (chlorides, sulphates of ca, Mg)

28. How will you remove the temporary hardness from water.

* When heated the water the temporary hardness producing salt are settled as precipitated

* Ex: Bicarbonates into insoluble carbonates.

**29. Another methods (Clark's method)**

Unit – 5

ALKALI AND ALIKALINE EARTH METALS

Alkaline and Alkaline Earth Metals.

1. Why the atomic and ionic radii of alkali, metal increases in moving down the group.

* Increases in the number of shell in group.

2. Why the ionization enthalpy decrease in group

* Increase size

* Screening effect

3. Why the second ionization enthalpies of alkali metal is high compare to 1st ionisation enthalpies of alkali metal.

* The removal of one e^- from alkali metal they attain noble gas electronic configuration.

Ex Na \rightarrow (Ne electronic configuration).

4. Why lithium salts are more soluble than the salt of the first group.

* LiClO_4 is upto 12 times more soluble than the other alkali metal salts Ex NaClO_4 , KClO_4 etc

* Due to strong solvation of small size.

5. How the alkali metals show coloured flame in flames?

When heating the alkali metal salt.

* The valence electron goes to higher energy level.

* After some time back into its actual energy level.

* The excess energy is emitted as light.

6. All the alkali metals are ionic crystal But why the Li shows the covalent character.

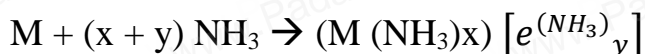
* Due to its small size so high polarizing on Iodide ion.

* Iodide ion being the largest can be polarised to a greater extent by Li ion.

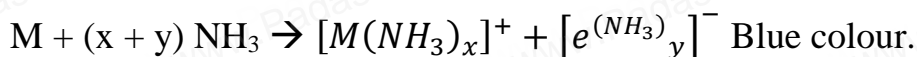
7. Why the alkali metal shows the conducting in Nature.

* It readily loses its valence electron in ammonia solution.

* Both cation and electron are ammoniated to give ammoniated cation and ammoniated electron.



Why the alkali metal form blue colour solution with alkali.



* Blue colour is due to ammoniate electron absorbs energy in the visible region of light.

* It give the blue colour to solution.

* In concentration blue become bronze colour.

9. Why the alkali metals are good reducing agent.

* It lose their valance electron easily

* So it is good reducing agents.



10. What is reaction of alkali metal with carbon .

* In alkal metals li only react with carbon gives lithium carbides.



* But other alkali metals do not react with carbon directly.

11. Uses of Alkalimetals.

* Used for making alloys.

For example “White metal” [Pb + Li]

Motor engines.

Li + Al \rightarrow air craft part

Li + Mg \rightarrow armour plates, thermonuclear reaction.

* Li (CO₃) \rightarrow medicine.

* Na is used to make Na / Pb alloy needed to make Pb (Et₄) and Pb (Me)₄

* These organolead compounds were earlier used as anti – knock additives to petrol, but nowadays lead – free petrol.

* Liquid Na metal – coolant in fast breed on nuclear reactor.

* K has a vital role in biological system.

* KCl is used as fertilizer.

KOH \rightarrow is used for manufacture of soft soap.

* It is also used as excellent absorbent of CO_2 .

* CS is used in devising photoelectric cell.

13. Why the LiF is insoluble in water?

* All halides are soluble in water except LiF due to its high lattice enthalpy.

* (Small size of Li^+ and F^-)

* Due to the presence of covalent nature both LiBr and LiI are soluble in organic solvent.

14. Why the alkaline earth metal are harder than the alkali metal?

* Atomic radius and density.

Alkaline earth metal < alkali metal.

* Alkaline earth metals are soft yet less than that of the alkali metals.

* It is due to metallic bonding stronger in alkaline earth metal.

15. What is diagonal relation, How is it so.....?

* The similarity in properties of the first element of each group with lower right of next period.

* Diagonally opposite element is known as diagonal relationship.

* Due size of ions, Polarizing power, Electropositive character

UNIT 6. GASEOUS STATE

1. Define the pressure

✱ Pressure is defined as force divided by area to which the force is applied

✱ SI unit – Pascal (Nm⁻²)

$$\text{Pressure: } \frac{\text{force}}{\text{Area}}$$

2. Boyles law

At given temperature the volume occupied by a fixed mass of gas inversely proportional to its pressure

$$V \propto \frac{1}{P}$$

3. Charless law

The volume is directly proportional to temperature at constant P and n for fixed mass of gas

$$V = kT \quad \frac{V}{T} = \text{constant} \quad P \propto T \quad (\text{or}) \quad \frac{P}{T} = \text{constant}$$

4. Isobars

The volume of gas linearly increases with temperature at given pressure such lines are called 'isobars'

5. Avagadros law

Equal volumes of all gases under the same condition of T&P contain equal number of molecule $V \propto n$.

6. Dalton law of partial pressure

"The total pressure of non – reacting gas is sum of partial pressure of gases present in the mixture" $P_{\text{total}} = P_1 + P_2 + P_3 \dots$

7. Different diffusion

The property of gas which involves the movement of gas molecule through the another gas is called.

8. Define Gay – lussac's law

at constant volume pressure of fixed mass of gas is directly proportional to temperature $P \propto T$.

9. Define effusion

It is process in which a gas escapes from a container through small hole

10. Define Graham law of Diffusion

“Rate of effusion (or) diffusion is inversely proportional to the square root of molar mass
 $\text{diffusion} \propto \frac{1}{\sqrt{m}}$

11. Define real gas and ideal gas

Ideal gas $PV = nRT$

Individual gas molecule occupy negligible volume when compared to the total volume of gas no attractive force between the gas molecules.

The gases doesn't obey above condition are called real gas

12. Define compressibility factor Z

* The deviation of real gases from ideal behavior

* IT is measure in term of a ratio of PV to nRT $Z = \frac{PV}{nRT}$

Ideal gas $PV = nRT$ So $Z = 1$.

13. Define Boyle point

The temperature at which a real gas obeys Ideal gas law over an appreciable range of pressure is called Boyle temperature (or) Boyle point

14. Define critical temperature TC

The temperature above which can not be liquefied even at high pressure denoted by TC.

15. Define critical pressure Pc

The minimum pressure is required to liquefy 1 mole of gas at Tc

16. Define critical Volume Vc

The volume occupied by 1 mole gases at its Tc and Pc

17. Define Joule Thomson effect

The Phenomenon of lowering of temperature when gas made to expand adiabatically from region of high pressure into region of low pressure

18. Define inversion temperature Ti

At which the temperature below which a gas expands adiabatically into region of low pressure through porous plug with fall in temperature $T_i = \frac{2a}{Rb}$

19. What are the methods used to liquefaction of gases

* Linde's method

* Claudes process

* Adiabatic process

UNIT 7. Thermodynamics

1. Explain the terminology used in thermodynamics.

system

- * Any portion of matter under consideration.
- * Which is separated from the rest of universe

Surrounding:-

- * Everything in the universe that is not part of the system and interact with it is called as surrounding.

2. Types of Systems:-

1. Isolated system 2. Closed system 3. Open system 4. Homogeneous system
5. Heterogeneous system

3. Define extensive and Intensive property.

- * The properties that depend on the mass (or) size of system.
- * Example: Volume, mass etc.
- * The properties that independent on the mass (or) size of system.
- * Example:- Density, temperature etc.

4. What are the state function.

- * The variable used to describe the state of system called state function.
- * Example: P, V, T

5. What are the thermodynamic process.

- (i) Isothermal process (ii) Adiabatic process (iii) Isobaric process
- (iv) Isochoric process (v) Cyclic process (vi) Spontaneous process
- (vii) Non spontaneous process (viii) Reversible process
- (ix) Irreversible

6. Differentiate about exothermic and endothermic process

Endothermic Process	Exothermic Process
* Absorption of heat	Evolution of heat
* Final state of system higher energy initial state of system lower energy	Final state of system lower energy Initial state of system higher energy
* Physical transformation take place	Physical transformation take place.
Example: Melting of ice.	Example: Forming ice.

7. Define path function?

- * Thermodynamic property of the system whose value depends on the path.

Example: Heat (q) work (w)

8. Define work, heat, energy terms.**Work**

Force multiplied by distance of displacement (S)

$$W = F.S.$$

Heat :

- * It is form of energy.
- * Algebraic quantity
- * It is path function and is not a state function

Energy is

- * Capacity to do the work
- * State function
- * Extensive property
- * Unit 'J' (or) KJ

9. Define enthalpy "H" and Explain its characteristics

H is defined as. Sum of internal energy "u" of the system and product of pressure and volume of the system.

$$H = U + PV$$

Characteristics

- * It depends on state function
- * But it is path function
- * Known by the term "heat content"

10. Define first law of thermodynamics.

- * Energy can neither be created nor destroyed but may be converted from one form to another form.

11. Define adiabatic process.

- * It is defined as one in which there is no exchange of heat (q) between system and surrounding during process $q = 0$

12. Define Isothermal process.

- * The system exchanges heat with its surrounding and the temperature of system remains constant $dT = 0$ during changes from its Initial to final state.

13. Define Isobaric process.

* It is defined as one in which the pressure of the system remains constant during change from initial to final state $dP = 0$

14. Define Isochoric process.

* It is defined as one in which the volume of the system remains constant during change from initial to final state $dv = 0$.

15. Define cyclic process.

* When a system returns to its original state after completing a series of changes.

* Then it is said to be cycle is completed.

* This process is known as cyclic process.

$$dU = 0, \quad dH = 0 \quad dP = 0 \quad dU = 0 \quad dT = 0$$

16. Define Zeroth law of thermodynamics.

* "If the two system are separately in thermal equilibrium with third one then.

They tend to be thermal equilibrium with them selves"

17. Define heat of combustion.

"The change in enthalpy of a system when one mole of substance is completely burnt in excess of air (or) oxygen.

Denoted by ΔH_c

18. Define specific heat capacity.

"The heat is absorbed by one kilogram of substance to raise its temperature by one Kelvin at specified temperature.

19. Define molar heat capacity. (cm)

"The amount of heat is absorbed by one mole of the substance to raise its temperature by 1 kelvin.

20. Application of Bomb calorimeter.

* Used to determine the amount of heat released in combustion reaction.

* Used to determine the calorific value of food.

* Used in many industries (food processing, explosive testing etc)

21. Heat of neutralisation.

"The change in enthalpy when one gram equivalent of an acid is completely neutralised by one gram equivalent of Base (or) vice versa in dilute solution.

22. Define Hess's Law.

“The enthalpy change of reaction either at constant volume (or) constant pressure is same whether it takes place in a single (or) multiple step.

$$\Delta H_r = \Delta H_1 + \Delta H_2 + \Delta H_3$$

23. Define lattice energy. (or) Lattica enthalpy

* The amount of energy required to completely remove the constituents ions from its crystal lattice to an infinite distance.

24. Define Entropy.

* It is a measure of molecular disorder (randomness) of system.

* unit of entropy Jk^{-1} .

25. Define entropy statement.

* The entropy of an isolated system increase during a spontaneous process.

26. Define Kelvin planck statement.

“It is impossible to construct a machine that absorbs heat from a hot source and converts it completely into work by cyclic process without transferring a part of heat to a cold sink.

27. Define – Clausius statement.

* It is impossible to transfer heat from a cold reservoir without doing some work.

28. Third law of thermodynamics.

* The entropy of pure crystalline substance at absolute zero is zero.

$$S = 0 \text{ (for perfect crystalline substance)}$$

29. What are criteria for spontaneity of a process

$-\Delta H = \text{negative} \rightarrow \text{exothermic spontaneous}$

$-\Delta S = \text{Positive} \rightarrow \text{spontaneous}$

$\Delta G = \Delta H - T\Delta S (-\Delta G) \rightarrow \text{spontaneous}$

$\Delta H - T\Delta S < 0 \rightarrow \text{spontaneous}$

30. Difference between Reversible and Irreversible

Reversible	Irreversible
The process in which system can be restored the initial state from final state	The process in which system can't restored the initial state from final state
The system and surrounding must be in equilibrium.	The system and surrounding must be not in equilibrium.

8. PHYSICAL AND CHEMICAL EQUILIBRIUM

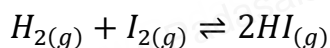
1. Why the chemical equilibrium is called dynamic equilibrium?

- * At equilibrium the forward and backward reaction proceeding at the same rate.
- * There is no macroscopic change is occurs.

2. Types of equilibrium with examples.

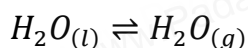
- * Homogeneous equilibrium.

In this equilibrium all the reactant and products are in same phase.



- * Heterogeneous equilibrium:

In this equilibrium all the reactant and products are different phase.



3. Law of mass action.

“At any instant the rate of a chemical reaction at given temperature is directly proportional to the product of active masses of the reactants at that instant”.

$$\text{Rate} \propto [\text{Reactant}]^x$$

4. Define equilibrium constant.

At a given temperature the ratio of the product of active mass of reactions products raised to the respective stoichiometric coefficient in the balanced chemical equation to that of reactants is constant.

5. Write application of equilibrium constant.

- * Predict the direction in which the net reaction will takes place.
- * Predict the extents of reaction.
- * Calculate the equilibrium concentration of reactants and products.

6. Write the correlation btw Q and Kc.

$Q = K_c \rightarrow$ Equilibrium reaction

$Q > K_c \rightarrow$ Backward reaction

$Q < K_c \rightarrow$ Forward reaction

7. le – chatelier principle.

“If the system at equilibrium is distributed then the system shift itself in a direction that nullifies the effect of that disturbance”.

8. Define (Q) reaction quotient.

- * The ratio of the products of active mass of reaction products raised to respective

stichiometric coefficients in the balanced chemical equation to that of reactants. $Q = \frac{[C]^l[D]^m}{[A]^x[B]^y}$ -NV-

9. SOLUTIONS

1. Define molality (m)

$$= \frac{\text{Number of moles of solute}}{\text{Mass of the solvent (in kg)}}$$

2. Define molarity (M)

$$= \frac{\text{Number of moles of solute}}{\text{Volume of solution (in lit)}}$$

3. Define Normality (N)

$$= \frac{\text{Number of grams equivalents of solute}}{\text{Volume of solution (in lit)}}$$

4. Define formality (F)

$$= \frac{\text{Number of formula weight of solute}}{\text{Volume of solution (in lit)}}$$

5. Molefraction (X)

$$= \frac{\text{Number of moles of components}}{\text{Total number of moles of all the components present in solution}}$$

6. Mass percentage (% w/w) :

$$= \frac{\text{Mass of solute (in g)}}{\text{Mass of solution (in g)}} \times 100$$

7. Volume percentage (% v/v) :

$$= \frac{\text{Volume of solute (in ml)}}{\text{Volume of solution (in ml)}} \times 100$$

8. Mass by volume percentage (%w/v):

$$= \frac{\text{Mass of solute (in g)}}{\text{Volume of solution (in ml)}} \times 100$$

9. Parts per million (ppm) :

$$= \frac{\text{Number of parts of components}}{\text{Total number of parts of all components}} \times 10^6$$

$$= \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 10^6$$

10. Advantages of using standard solution:

- * The error due to weighing the solute can be minimised.
- * We can prepare working standard of different concentrations by using standard solution.
- * More stable less microbial growth.

11. What are the factors influencing solubility?

- * Nature of solute and solvent
- * Effect of temperature
- * Effect of pressure of solution

12. Henry's law:

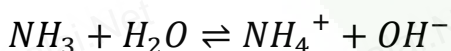
The partial pressure of gas in vapour phase (vapour pressure of solute) is directly proportional to the mole fraction (x) of the gaseous solute in the solution at low concentration”

$$P_{\text{solute}} \propto X_{\text{solute in solution}}$$

$$P_{\text{solute}} = K_H X_{\text{solute in solution}}$$

13. What are the limitations of Henry's law?

- * Applicable at moderate temperature and pressure only.
- * Obey only less soluble gases.
- * The gases reacting with solvent do not obey Henry's law.



- * The gases obeying Henry's law should not associate or dissociate while dissolving in the solvent.

14. Define vapour pressure of liquid.

- * The vapour pressure of liquid is at equilibrium is established between liquid and its vapour. The pressure of vapour in equilibrium with its liquid is called vapour pressure of liquid.

15. Define Raoult's law.

The solution of volatile liquids, the partial vapour pressure of each components ($A \propto B$) of solution is directly proportional to its mole fraction.

$$P_A \propto X_A$$

$$P_A = KX_A$$

$$P_A = P_A^\circ X_A$$

16. Define ideal solution with example.

* There is no change in volume on mixing two components (solute and solvent)

$$(\Delta V_{\text{mixing}} = 0)$$

* There is no exchange of heat when the solute is dissolved in solvent

$$(\Delta H_{\text{mixing}} = 0)$$

* Escaping tendency of solute and solvent present in it should be same of pure liquids

Eg: benzene & Toluene

17. Define Non-Ideal solution with example.

* The solutions which do not obey Raoult's law over entire range of concentration.

$\Delta H_{\text{mixing}} \neq 0$ & $\Delta V_{\text{mixing}} \neq 0$ Benzene & acetone.

18. Factors responsible for deviation from Raoult's law

* Solute – solvent interaction

* Dissociation of solute

* Association of solute (Dimer)

* Pressure, Temperature, Concentration

19. Define colligative properties

* The properties only do not depend upon the chemical nature of solute particle

depends on the number of the solute particles called colligative properties.

Ex : ΔP , ΔT , ΔT_b

20. Elevation of Boiling point.

* Solution boils at higher temperature T_b than the boiling point of the pure solvent (T_b)

this increase in the boiling point is known as elevation of boiling point.

21. Depression in Freezing point.

* The temperature at which the solid and the liquid states of the substance have same vapour pressure.

22. Define osmosis.

* Spontaneous process by which solvent molecules pass through a semi permeable membrane

from a solution of lower concentration to a solution of higher concentration.

23. Osmotic pressure:

“The pressure must be applied to the solution to stop the influx of the solvent through the semipermeable membrane.

24. Isotonic solution:

* Two solution having same osmotic pressure at given temperature are called isotonic solutions.

25. Define hemolysis:

Solvent from outside of the cell will flow into the cell to normalise the osmotic pressure and this process is called hemolysis.

26. Reverse osmosis:

A process in which a solvent passes through a semipermable membrane in the opposite direction of osmosis. (before you write definition of osmosis)

27. Define van't Hoff factor.

$$\frac{\text{Normal (actual) molar mass}}{\text{Observed (abnormal) molar mass}} = \frac{\text{Observed colligative property}}{\text{Calculated colligative property}}$$

28. Define Vant's hoff equation. ((Based on osmotic pressure)

For Dilute solutions, the osmotic pressure is directly proportional to the molar concentration of solute and Temperature of solution $\pi = cRT$

C – Concentration of solution in molarity

T – Temperature

R – Gas constant

29. Which bond (or) stronger $\sigma \propto \pi$ why

* σ bond is stronger

* σ bond formed – head on overlap (maximum overlap)

* π bond formed – (side wise 1 (less overlap)

30. What is dipole moment.

The polarity of covalent bond is measure by Dipole moment $\mu = q \times 2d$.

μ – Dipole moment

q - Charge $2d$ – distance between two charge.

UNIT 10. CHEMICAL BONDING

1. Define octet rule.

*The atom transfer (or) share electrons so that all the atoms involved in chemical bonding obtain 8 electrons in their out shell. (Valence shall)

2. Define covalent bond.

* Mutual sharing of one (or) more pair of electrons between two combining atoms form chemical bond called covalent bond.

3. Define Ionic (or) electrovalent bond.

* The complete transfer of electrons leads to the formation of cation and anion.

*Both these Ions are heldtogether by electrostatic attraction force which is known as Ionic bond.

4. Define Co-ordinate covalent bond.

* Two electrons which are necessary for covalent bond formation.

* These two electrons are shared by both combining atom.

5. Define bond length.

* The distance between the nuclei of two covalently bonded atom is called Bond length.

6. Define Bond angle.

* The Directional nature of covalent bond creates the fixed angle between two covalent bonds in the molecule.

7. Bond enthalphy.

* The minimum amount of energy required to break one mole of particular Bond in molecule in their gaseous state.

8. Define sigma (σ) and Pi (π) bond.

* When two atomic orbitals overlap linearly along the axis, the resultant bond is called sigma σ bond. (or) axial overlap (or) head on overlap

* When the two atomic orbitals overlap sideway the resultant covalent bond is called π bond.

9. Define Hybridisation.

* Process of intermixing of atomic orbitals of same atom with comparable energy to form equal number of new orbital with same energy.

10. Define Bond order.

* The half of the Difference between the number of electron in Bonding molecular orbital and antibonding molecular orbital is called bond order.

$$\text{Bond order} = \frac{N_b - N_a}{2}$$

UNIT – 15 ENVIRONMENTAL CHEMISTRY

1. DEFINE ENVIRONMENTAL POLLUTION:

Any undesirable change in our environment that has harmful effects on plant, animals and human beings is called environmental pollution.

2. DEFINE POLLUTANTS?

Which substance cause the pollution to environment is called pollutants.

3. TYPES OF POLLUTANTS:

(i) Biodegradable (ii) Non biodegradable

4. DEFINE BIODEGRADABLE POLLUTANTS WITH EXAMPLE?

Pollutants are easily decomposed by Natural Biological process.

Eg: Animal wastes and plant wastes.

5. DEFINE NON – BIODEGRADABLE POLLUTANTS WITH EXAMPLE?

Pollutants are not easily decomposed by Natural biological process.

Eg: DDT, plastics.

6. HOW ACID RAIN IS FORMED?

The SO_2 react with O_2 form SO_3 as like N_2 combine with O_2 form NO_2 it react with rain to form acid rain.



7. DEFINE GREEN HOUSE EFFECT?

* Heating up of earth surface due to infrared radiation reflected by CO_2 layer in atmosphere.

* The radiation reflected by earth's surface by CO_2 layer in the atmosphere.

8. DEFINE GLOBAL WARMING:

The heating up of earth through the green house effect is called global warming.

9. STONE LEPROSY :

* Acid rain cause extensive damage to building and structural materials of marbles.

* This attack on marbles is termed as stone leprosy.



10. CLASSICAL (OR) LONDON SMOG:

- * It consist of coal smoke and fog.
- * It accurs in cool humid climate.
- * Chemical combination SO_2 and $\text{SO}_3 \propto$ humidity.
- * SO_2 oxidised SO_3 .

11. REDUCING SMOG:

- * Chemically reducing in nature.
- * High concentration of SO_2 so it is called reducing smog.

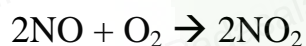
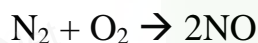
12. PROTOCHEMICAL SMOG (OR) LOS ANGEL SMOG.

- * Consist of smoke, dust, fog and air pollutants.
- * Occurs in sun shine.
- * Chemical combination oxides of N and hydro carbons.

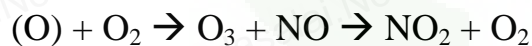
13. DEFINE OXIDISING SMOG.

- * Chemically oxidizing in nature.
- * High concentration of NO_2 and O_3 .

So it is also called as oxidising smog.



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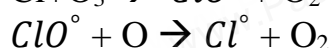
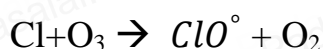
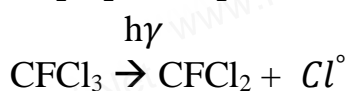
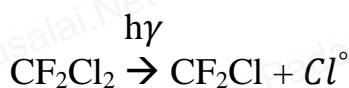
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**14. WHY OZONE LAYER IS CALLED EARTH PROJECTIVE UMBERLLA?**

- * Ozone layer shield the harmful UV radiation from sunlight.
- * It behave like a umbrella.
- * It prevent skin cancer.

15. OZONE HOLE? HOW OZONE DEPETION TAKES PLACE.

- * The loss of ozone molecule in the atmosphere



16. DEFINE POINT SOURCE AND NON POINT SOURCE:

- * Easily identified source of place of pollution – Municipal - point source.
- * Can't easily identified source of place of pollution, acid rain, mining waste non – point source.

17. DEFINE EUTROPHICATION:

Water Bodies receive excess nutrients that stimulates excessive plant growth (algae). Kill animals by depriving it of oxygen. Loss of Biodiversity.

18. ALAGE BLOOM:

The enchanced plant growth in water bodies is called algae bloom.

19. BOD (BIOLOGICAL OXYGEN DEMAND)

- * The total amount of oxygen in mg consumed by micro organisms.
- * The micro organism is decomposing the waste in one litre of water at 20°C for a period of 5 days.
- * Expressed in ppm.
 - > 5ppm – clean water
 - < 17ppm – polluted water

20. COD (CHEMICAL OXYGEN DEMAND)

Amount of oxygen required by organic matter in a sample of water. For it oxidation by strong oxidising agent $K_2Cr_2O_7$ in acidic medium for 2 a hrs.

21. TDS (TOTAL DISSOLVED SOLIDS).

- * Most of the salts are soluble in water Ca, Na, K etc.,
- * Drinking water ids 500 ppm> cause irritation in stomach and intestine.

22. GREEN CHEMISTRY:

- * It is branch of science.
- * It encouraging the design producing Harzardous substance free to the environment.
- * Reduce and use of Harzardous substance.
- * Make eco – friendly compound.
- * Example : styrene produced by traditional and greener routes.

Traditional route:

- * This method involved two step (carainagenic)



Green route:

- * By this method avoid carcinogenic benzene so we
- * Starts with cheaper and environmentally safer xylene.

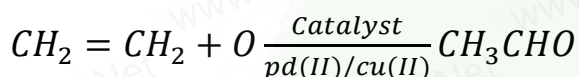
(i) Dry cleaning of clothes:

- * Now a days tetrachloroethylene not used in dry cleaning clothes.
- * Pollute the ground water and carcinogenic.
- * Green chemistry process is used the liquefied CO₂ with suitable detergent. Not harmful not carcinogenic, not pollute ground water.
- * Now a days H₂O₂ used to bleaching clothes – less water utilizes.

(ii) Bleaching of paper:

Normally bleaching by using chlorine.

Green bleaching by H₂O₂ used.

(iii) Synthesis of chemicals:

CH₃CHO (90% yield) it is a one step process.

(iv) Instead of petrol methanol used as fuel in auto mobiles.

(v) Neem based pesticide have been synthesized which are more safer than chlorinated hydrocarbon.

(If you find any correction consult your Chemistry Staff)

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