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Class : XI-Bio/CS ***Chemistry***

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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 : CO_2, $C_6H_{12}O_3$, NaCl

4. Define atomic mass.

* One twelth 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 atonilearass unit.

5. Define relative atomic mass.

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

6. Define molecular mass

* Ratio of the mass of a moleculated 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 contain many elementary particles.

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

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

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.

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

10. Define gram equivalent mass.

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

Equivalence factors Basicity of acids (or) acidity of base (or) $e^{-}gaid\ (or) e^{-}loss$.

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 symbel.

Ex: The molecular formula acetic acid $C_2H_4O_2$ (CH_3COOH).

13. Define whole number.

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

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 quantines 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.

Cu → Cu^2 SO_4ze

18. Define Reduction.

Reduction is the addition of H_2 removal of O_2.

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

$Fe^{3+} + e^- \rightarrow Fe^{2+}$

19. Redox reaction.

The oxidation reaction are accompanied by reduction reaction and vice versa called redox reaction $H_2 + S_2 +$ → H_2S

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 redom reactions the oxidation number of element changes.

* Oxidation number increase is called oxidation.

* Oxidation number decrease is called reduction.

$2KM_nO_2 + 10FeSO_4 + 8H_2SO_4 \rightarrow K_2SO_4 + M_nSO_4 + 5Fe_2(SO_4) + 8H_2O.$

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.

$2KM_nO_4 + 10FeSO_4 + 8H_2SO_4 \rightarrow K_2SO_4 + M_nSO_4 + 5Fe_2(SO_4)_2 + 8H_2O$

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.

$2KClO_3$ → $2KCl + 3O_2$.

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.

Oxidation

(i) $CuSO_4(aq) + Zn_{(s)} \rightarrow Cu_{(s)} + ZnSO_4(aq)$

Reduction

Here Zinc metal replaced on copper.

Oxidation

(ii) $Z^{\circ}n + 2HCl \rightarrow znCl_2 + H^{\circ}{}_2$

Reduction

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

$2H_2O_2 \rightarrow 2H_2O + O_2$

Reduction

28. Define competitive electron transfer reaction.

Oxidation

$Cu_{(s)} + 2AgNo_3(aq) \rightarrow Cu^{+2}(NO_3)2aq + 2Ag(S)$

Reduction

* 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.

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

33. Equivalent mass of Base.

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

34. Equivalent mass of oxidising agent:

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

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2. QUANTUM MECANICAL MODEL OF ATOM

1. How many protons and neutrons are present in $^{18}_{8}$ 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 ?

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)

$1s^2\ 2s^2\ 3s^2\ 3p^6\ 4s^1\ 3d^2$

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 $3d^{10}$ $4s^1$ instcad 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 coper 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 spilt 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. -NV-

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.

Hortizontal 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 transion 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 Zeff = Z – S

* Between nucleus and outermost electron.

12. Define – shielding effect.

* The force of attraction between nucleas 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.

Mg + (energy) $E_1 \rightarrow M^+ + e^-$

Increase in period.

Decrease in down a group

14. Define electron affinity.

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

$M_g + e^- \rightarrow M^- +$ energy KJ / Mole – unit.

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.

* Compaining 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 → easily lose their electron and less Ionisation energy.

Right side elements in periodic table → 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
Third period	Na	Mg	Al	Si

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.

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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 :- ${}_1H^1$ (99.985%) no neutrons.

Deutrium:- ${}_1H^2$ (0.015%) heavy hydrogen.

Tritium :- ${}_1H^3$ (~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 aqu NaOH (or) KOH

* Using Nickel anode and Iron cathode

At Anode : $2OH^- \rightarrow H_2O + \frac{1}{2}O_2 + 2e^-$

At cathode : $2H_2O + 2e^- \rightarrow 2OH^- + H_2$

Overall the reaction : $H_2O \rightarrow H_2 + \frac{1}{2}O_2$

6. Give the laboratory preparation of Hydrogen

* $Zn + 2HCl \rightarrow ZnCl_2 + H_2\uparrow$

7. Give the industrial production of H_2.

* Hydrocarbon (methane) mixed with steam

* In the presence of Ni catalyst

* 800 - 900°C and 35 atm

$CH_4 + H_2O \rightarrow CO + 3H_2$.

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

* Steam passed over red – hot coke

* Produce CO and H_2O.

* Used to prepare organic compound methanol.

$C + H_2O \xrightarrow{1000°C} CO + H_2$

Waster gas / syngas

9. How will you remove CO from in water gas or how will you convert the CO into CO_2 in water gas.

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

* Catalyst Fe/Cu and potassium carbonate used to absorbed CO_2.

* $CO + H_2O \rightarrow CO_2 + H_2$ water gas

* $CO_2 + K_2CO_3 + H_2O \rightarrow 2K\,H\,CO_3$

10. Preparation of Deutrium.

* Normal water contain $1.6x10^{-4}$% of Heavy Water

* Dissociation = $H_2O > D_2O$

(protium water > Heavy water)

$2D_2O \xrightarrow{\text{Electrolysis}} 2D_2 + O_2$

11. Transmutation reaction.

$$^{6}_{3}Li + ^{1}_{0}n \rightarrow ^{4}_{2}He + ^{3}_{1}T$$

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

$2H_2 + O_2 \rightarrow 2H_2O$

With halogen

$2H_2 + X_2 \rightarrow 2HX\ (x - Cl, Br)$

With metal

$2Li + H_2 \rightarrow 2LiH$

$2Na + H_2 \rightarrow 2NaH$

15. Why H_2 is reducing agent.

$HC \equiv CH \xrightarrow{Ni / H_2} H_2C = CH_2 \xrightarrow{Ni / H_2} CH_3 - CH_3$

Unsaturated Saturated.

16. How will you prepare metal Hydrides.

$4LiH + AlCl_3 \rightarrow Li\ (AlH_4) + 3LiCl$

$4NaH + B(OH_3)_3 \rightarrow Na\ (BH_4) + 3\ CH_3\ ONa$

17. Exchange (or) substitution reaction of Deuterium.

Deuterium replace Hydrogen in a compound.

$CH_4 + 2D_2 \rightarrow CD_4 + 2H_2$

$2NH_3 + 3D_2 \rightarrow 2ND_3 + 3H_2$

18. Properties of tritium

$${}^{3}_{1}T \rightarrow {}^{3}_{2}He + {}^{0}_{1}e$$

Half life period 12.3 year,

β emitter

19. Uses of Hydrogen

Haber Process

$N_2 + 3H_2 \xrightarrow{\frac{380-450°C}{200atm/Fe}} 2NH_3$

* Used to prepare ammonia.

* Ammonia used for preparation of HNO_3 fertilizer and explosives.

* It is used to prepare industrial solvent $CO + 2H_2 \xrightarrow{Cu} CH_3OH$

* Unsaturated fatty oils → saturated fat by Pt / H_2

* Metal oxides are converted into metals at high temperature.

$CuO + H_2 \rightarrow Cu + H_2O$

$WO_3 + 3H_2 \rightarrow W + 3H_2O$

* 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

$2Na + 2H_2O \rightarrow 2NaOH + H_2$ (metal)

$Ba + 2H_2O \rightarrow Bo(OH)_2 + H_2$

$3Fe + 4H_2O \rightarrow Fe_3O_4 + H_2$

21. Reaction of water with F_2

$2F_2 + H_2O \rightarrow 4HF + O_2$

22. Write the reaction of water with non metal.

$CO_3 + H_2O \rightarrow HCO_3^- + OH^-$

23. Why water is amphoteric oxide?

* Water is act as acid as well as base

by accept proton (acid)

by donate proton (Base)

$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$ (donate proton)

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$ (Accept proton)

24. Write reaction of water with covalent compounds.

$SiCl_4 + 2H_2O \rightarrow SiO_2 + 4HCl$

$P_4 O_{10} + 6H_2O \rightarrow 4H_3 PO_4$

25. Define Hard water.

* Water containing soluble salts of Bicarobonates, 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 harness (Bicarbonates salts of (Ca, mg)

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

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

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

* Ex: Bicarbonates into insoluble carbonates.

$Ca(HCO_3)_2 \rightarrow CaCO_3 + H_2O + CO_2$

$Mg(HCO_3)_2 \rightarrow MgCO_3 + H_2O + CO_2$

$MgCO_3 + H_2O \rightarrow Mg(OH)_2 + CO_2 + H_2O$

29. Another methods (Clark's method)

$Ca(HCO_3)_2 + Ca(OH)_2 \rightarrow 2CaCO_3 + 2H_2O$

$Mg(HCO_3)_2 + 2Ca(OH)_2 \rightarrow 2CaCO_3 + Mg(OH)_2$

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Unit – 5

ALKALI AND ALIKALINE EARTH METALS

Alkalin 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 enthalphy decrease in group**

   * Increase size

   * Screening effect

3. **Why the seond ionization enthalpies of alkali metal is high compare to Ist lonisation enthalpies of alkali metal.**

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

   Ex Na → (Ne electronic configuration).

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

   * $LiClO_4$ is upto 12 time more soluble than the other alkali metal salts Ex $NaClO_4$, $KClO_4$ et

   * Due to strong solvation of small size.

5. **How the alkali metals are shows coloured flame in flames?**

   When heating the alkali metal salt.

   * The valence electron is 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 polarisizing 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 emmoniated to give ammoniated cation and ammoniated electron.

M + (x + y) NH_3 → (M (NH_3)x) $\left[e^{(NH_3)}{}_y\right]$

Why the alkali metal form blue colour solution with alkali.

M + (x + y) NH_3 → $[M(NH_3)_x]^+ + \left[e^{(NH_3)}{}_y\right]^-$ Blue colour.

* 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.

$M_{(s)} \rightarrow M^+_{(g)} + e^-$

10. What is reaction of alkali metal with carbon .

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

$2Li + 2C \rightarrow Li_2C_2$

* 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 → air craft part

Li + Mg→ armour plates, thermonuclear reaction.

* Li (CO_3) → medicine.

* Na is used to make Na / Pb alloy needed to make Pb (Et_4) and Pb $(Me)_4$

* These organolead compounds were earlier used as anti – knock additives to petrol, but nowdays 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 → 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 enthalphy.

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

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

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

* Atomiradius and density.

Alkaline earth metal < alkalimetal.

* Alkaline metals 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 similary 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

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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})

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

2. Boyles law

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

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

3. Charless law

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

$V = kT$ $\frac{V}{T}$ = constant $P \alpha T$ (or) $\frac{P}{T}$ 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 $V \alpha n$.

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 + total = $P_1 + P_2 + P_3$ ….

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

If 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 diffusion = $\alpha \frac{1}{\sqrt{m}}$

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 P_c

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

16. Define critical Volume V_c

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

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UNIT 7. Thermodynamics

1. Explain the terminology used in thermodynamics.

system

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

Surrounding:-

* Everthing 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 independ 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) Iso choric process (v) Cyclic process (vi) Spontaneus process
(vii) Non spontaneous process (viii) Reversible process
(ix) Irreversible

6. Differentiate about exothermic and enothermic process

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

7. Define path 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.
* Algebaric 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 it characteristics

H is defind 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 to another form.

11. Define adibatic 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, dH = 0 dP = 0dU = 0 dT = 0

16. Define Ieroth 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 (for perfect crystalline substance)

29. What are criteria for spontaneity of a process

$-\Delta H$ = negative → exothermic spontaneous

$-\Delta S$ = Positive → spontaneous

$\Delta G = \Delta H - T\Delta S(-\Delta G)$ → spontaneous

$\Delta H - T\Delta S < O$ → 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.

$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$

* Heterogeneous equilibrium:

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

$H_2O_{(l)} \rightleftharpoons H_2O_{(g)}$

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".

Rate $\propto [Reactant]^x$

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 = Kc → Equilibrium reaction

Q > Kc → Backward reaction

Q < Kc → 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)

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

2. Define molarity (m)

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

3. Define Normality (N)

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

4. Define formality (F)

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

5. Molefraction (X)

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

6. Mass percentage (% w/w) :

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

7. Volume percentage (% v/v) :

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

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

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

9. Parts per million (ppm) :

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

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

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 solute $\propto$ X solute in solution

P solute = K_H X 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.

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

* 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.

$P_A \propto X_A$

$P_A = KX_A$

$P_A = P^\circ_A X_A$

16. Define ideal solution with example.

* There is no change in volume on mixing two components (solute and solvent) $(\Delta \boldsymbol{Vmixing} = \mathbf{0})$

* There is no exchange of heat when the solute is dissolved in solvent $(\Delta \boldsymbol{Hmixing} = \mathbf{0})$

* 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 donot obeys Raoults law over entire range of concentration.

$\Delta Hmixing \neq 0$ & $\Delta V\ mixing \neq 0$ Beenzene $\propto$ acetone.

18. Factors responsible for deviation from Raoults law

* Solute – solvent interaction
* Dissociation of solute
* Association of solute (Dimer)
* Pressure, Temperature, Concentration

19. Define colligative properties

* The properties only doesnot depends 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 semipremeable membrane in the opposite direction of osmosis. (before you write definition of osmosis)

27. Define van't Hoff factor.

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

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 forme – (side wise 1 (less overlap)

30. What is dipole moment.

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

μ – Dipole moment

q - Change 2d – distance between two change.

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.

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

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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.

$2SO_2 + O_2 + 2H_2O \rightarrow 2H_2SO_4$

$4NO_2 + O_2 + 2H_2O \rightarrow 4HNO_3$

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 earths 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.

$CaCO_3 + H_2SO_4 \rightarrow CaSO_4 + H_2O + CO_2.\uparrow$

10. CLASSICAL (OR) LONDON SMOG:

* It consist of coal smoke and fog.

* It accurs in cool humid climate.

* Chemical combination SO_2 and $SO_3 \propto$ 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.

$$N_2 + O_2 \rightarrow 2NO$$

$$2NO + O_2 \rightarrow 2NO_2$$

$$NO_2 \xrightarrow{*} NO + (O)$$

$$(O) + O_2 \rightarrow O_3 + NO \rightarrow NO_2 + O_2$$

$$NO_2 \xrightarrow{*} NO + O$$

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

$$CF_2Cl_2 \xrightarrow{h\gamma} CF_2Cl + Cl^{\circ}$$

$$CFCl_3 \xrightarrow{h\gamma} CFCl_2 + Cl^{\circ}$$

$$Cl + O_3 \rightarrow ClO^{\circ} + O_2$$

$$ClO^{\circ} + O \rightarrow Cl^{\circ} + O_2$$

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)

$$Benzene + Ethylene \rightarrow Ethyl\ Benzene \xrightarrow[(dehydrogenation)]{Fe_2O_3/Al_2O_3} styrene$$

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 caranogenic.
* Green chemistry faccous is used the liquefied CO_2 with sutiable detergent. Not harmful not caranagenic, not pollute ground water.
* Now a days H_2O_2 used to bleaching clothes – less water utilizes.

(ii) Bleaching of paper:

Normally bleaching by using chlorine.

Green bleaching by H_2O_2 used.

(iii) Synthesis of chemicals:

$$CH_2 = CH_2 + O \xrightarrow[pd(II)/cu(II)]{Catalyst} CH_3CHO$$

$CH_3CHO(90\%\ yield)$ 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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