(ACTC)   ADVANCED CHEMISTRY TUITION CENTRE, PWD ROAD, NAGERCOIL 9940847892.

Book Back - Self Evaluation Objectives
1. Chemical Calculations
A. Choose the best answer :
1. The volume occupied by 16g of oxygen at S.T.P.
a) 22.4 L   b) 44.8 L   **c) 11.2 L**   d) 5.6 L
2. Avogadro's number represents the number of atoms in
a) 12g of C^{12}   b) 320g of S   c) 32g of Oxygen   d) 12.7g of iodine.
3. The value of gram molecular volume of ozone at S.T.P is
a) 22.4 L   b) 2.24 L   c) 11.2 L   d) 67.2 L
4. The number of atoms present in 0.5 gram- atoms of Nitrogen is same as the atoms in
a) 12g of C   b) 32g of S   **c) 8g of the oxygen**   d) 24g of magnesium.
5. The number of gram-atoms of oxygen in 128g of oxygen is
a) 4   **b) 8**   c) 128   d) $8 \times 6.02 \times 10^{23}$
6. The total number of moles present in 111 g of $CaCl_2$ is
a) One mole   b) Two moles   c) Three moles   d) Four moles
7. Which of the following weighs the most?
a) One gram-atom of nitrogen   b) One mole of water
c) One mole of Sodium   d) One molecule of H_2SO_4
8. Which of the following contains same number of carbon atoms as are in 6.0g of carbon (C-12)?
a) 6.0g ethane   **b) 8.0g methane**   c) 21.0g Propane   d) 28.0g CO
9. Which of the following contains maximum number of atoms?
a) 2.0g hydrogen   b) 2.0g oxygen   c) 2.0g nitrogen   d) 2.0g methane
10. Which one among the following is the standard for atomic mass?
a) H   **b) $_6C^{12}$**   c) $_6C^{14}$ d) $_8O^{16}$
11. Which of the following pair of species have same number of atoms under similar conditions ?
a) 1L each of SO_2 and CO_2   b) 2L each of O_3 and O_2
c) 1L each of NH_3 and Cl_2 d) 1L each of NH_3 and 2L of SO_2
12. 2.0 g of oxygen contains number of atoms same as in
a) 4 g of S   b) 7 g of nitrogen   c) 0.5 g of H_2 d) 12.3 g of Na
13. The number of gm-molecules of oxygen in 6.02×10^{24} CO molecules is
a) 1 gm-molecule   b) 0.5 gm-molecule   c) 5 gm-molecule   **d) 10 gm-molecule**
14. Hydrogen phosphate of certain metal has a formula MHPO4, the formula of metal chloride is
a) MCl   b) MCl_3 **c) MCl_2**   d) MCl_4
15. A compound contains 50% of X (atomic mass 10) and 50% Y (at. mass20). Which formulate
pertain to above date ?
a) XY   **b) X_2Y**   c) X_4Y_3 d) $(X_2)_3 Y_3$
16. Which of the following compound has / have percentage of carbon same as that in ethylene
(C_2H_4) ?
a) propene   b) Cyclohexane   c) Ethyne   d) Benzene
17. 5L of 0.1 M solution of sodium Carbonate contains

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a) 53 g of Na_2CO_3   b) 106 g of Na_2CO_3 c) 10.6 of Na_2CO_3

d) 5 x 102 millimoles of Na_2CO_3

3.Atomic Structure – I

1. Atomic mass of an element is not necessarily a whole number because :

(a) It contains electrons, protons and neutrons

(b) It contains allotropic forms

(c) Atoms are no longer considered indivisible

(d) It contains isotopes

 (e) None of these.

2. No two electrons in an atom will have all four quantum numbers equal. The statement is known as   **(a) Exclusion principle**   (b) Uncertainty principle   (c) Hund's rule

(d) Aufbau principle   (e) Newlands law.

3. When the 3d orbital is complete, the new electron will enter the

(a) 4p orbital   (b) 4f orbital   (c) 4s orbital   (d) 4d orbital   (e) 5s orbital.

4. The preference of three unpaired electrons in the nitrogen atom canbe explained by :

(a) Pauling's exclusion principle   (b) Aufbauprinciple

(c) Uncertainty principle   **(d) Hund's rule**   (e) None of these.

5. The number of orbitals in a p-sub-shell is

(a) 1   (b) 2   **(c) 3**   (d) 6   (e) 5.

6. The nucleus of an atom contains :

(a) Electrons and protons   **(b) Neutrons and protons**

(c) Electrons, protons and neutrons   (d) Neutrons and electrons   (e) None of these

7. Which is the lightest among the following?

(a) An atom of hydrogen   **(b) An electron**   (c) A neutron

(d) A proton   (e) An alpha particle.

8. Which of the following has no neutrons in the nucleus?

(a) Deuterium   (b) Helium   **(c) Hydrogen**   (d) Tritium   (e) An alpha particle.

9. When the value of the azimuthal quantum number is 3, the magnetic quantum number can have values :   (a) +1,-1   (b) +1,0, 1   (c) +2,+1,0, -1,-2

(d) +3,+2, +1,0, -1,-2,-3   (e) +3,-3.

10. 2p orbitals have :

(a) n = 1,l = 2   (b) n=1, l = 0   (c) n = 2, l = 0   **(d) n = 2, l =1**   (e) n =1, l =1.

11. The atomic number of an element is 17 and its mass number is 37. The number of protons, electrons and neutrons present in the neutral atom are :

(a) 17, 37,20   (b) 20,17,37   **(c) 17, 17, 20**   (d) 17, 20,17   (e) 37, 20, 17.

12. The maximum number of electrons that can be accommodated in the nth level is :

(a) n2   (b) n+1   (c) n-1   **(d) 2n2**   (e) 2 + n.

13. The magnetic quantum number decides :

(a) The distance of the orbital from the nucleus

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(b) The shape of the orbital

(c) The orientation of the orbital in space

(d) The spin of the electron

 (e) None of these.

4.Periodic Classification – I

1. The elements with atomic numbers 31 belongs to :

(a) d-block   (b) f-block   **(c) p-block**   (d) s-block

2. Representative elements are those which belong to :

(a) s and d-blocks   **(b) s and p-blocks**   (c) p and d-blocks   (d) d and f-blocks

3. The most electronegative element of the periodic table is :

(a) Iodine   **(b) Flourine**   (c) Chlorine   (d) Oxygen

4. Which of the following forms stable gaseous negative ion.

(a) F   **(b) Cl**   (c) Br   (d) I

5. The elements having highest ionization energies within their periods are called :

(a) Halogens   **(b) Noble gases**   (c) Alkali metals   (d) Transition elements

6. A property which progressively increases down a group in the periods table is :

(a) Ionization enthalpy   (b) Electronegativity

(c) Electron gain enthalpy   **(d) Strength as a reducing agent**.

7. Elements whose atoms have their s and p-sub-levels complete are the:

(a) Normal elements   (b) Transition elements   (c) Halogens   **(d) Inert gases.**

8. The law of triad is applicable to :

(a) Chlorine, bromine and iodine   (b) Hydrogen, oxygen and nitrogen

(c) Sodium, neon and calcium   (d) All of the above

9. The law of octaves was stated by :

(a) Dobereiner   (b) Mendeleev   (c) Moseley   **(d) Newland**

10. Which of the following property decreases down a group :

(a) Ionization enthalpy   (b) Atomic radii   (c) Valency   (d) All the above properties

11. Which of the following has the lowest melting point ?

(a) CsCl   (b) RbCl   (c) KCl   (d) NaCl   **(e) LiCl.**

12. Which of the following hydroxide is most basic ?

(a) Mg (OH) 2   **(b) Ba (OH) 2**   (c) Ca(OH) 2   (d) Be (OH)$_2$

13. Excluding hydrogen and helium, the smallest element in the periodic table is :

(a) lithium   (b) Oxygen   **(c) Fluorine**   (d) Chlorine

14. Which one among the following species has the largest atomic radius:

(a) Na   (b) Mg   (c) Al   (d) Si

15. Which of the following is the lightest metal ?

(a) Calcium   **(b) Lithium**   (c) Magnesium   (d) Sodium

16. Which of the following has highest ionization potential?

(a) Sodium   (b) Magnesium   (c) Carbon   **(d) Fluorine**

17. With respect to chlorine, hydrogen will be

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(a) Electropositive   (b) Electronegative   (c) Neutral   (d) None of these.

18. Which element has the greatest tendency to lose electrons?

(a) Chlorine   (b) Sulphur   **(c) Francium**   (d) Berylium.

19. Halogens belong to the :

(a) s-block   **(b) p-block**   (c) d-block   (d) f- block

(e) Zero group of the periodic table.

20. Compared to first ionization enthalpy of an atom, the second is :

(a) Greater   (b) Less   (c) Same   (d) Negligible

21. Which arrangement of the following set of atoms is in order of increasing atomic radius:

Na, Rb, K and Mg ;

(a) Na, Mg, K, Rb   (b) Na, K, Mg, Rb   **(c) Mg, Na, K, Rb**   (d) Na, Mg, Rb, K

22. The first attempt to classify the elements was made by :

(a) Mendeleev   (b) Newland   (c) Lother Meyer   **(d) Dobereiner**

23. Characteristic of transition elements is incomplete in :

(a) d-orbitals   (b)f-orbitals   (c) p-orbitals   (d) s-orbitals

24. Which of the following will have lowest first ionization enthalpy ?

(a) Na   (b) Al   (c) Mg   (d) Si

25. Which of the following atoms is likely to give off more energy on gaining an electron ?

(a) Na   (b) Mg   (c) Al   **(d) Cl**

26. Transition metals have the electronic configuration :

(a) $ns^2 nd^{1-10}$ (b)$ns^2 np^6(n-1)d^{1-10}$ **(c) $ns^2 (n-1)d^{1-10}$**   (d) $ns^2 np^6(n-1)d^{1-10}$

27. In the first transition series the incoming electron enters the :

(a) 4d-orbital   **(b) 3d-orbital**   (c) 5d-orbital   (d) 6d-orbital

5. Group 1S -Block Elements

1. Atoms of the same element having same atomic number but different mass number are called

(a) isotopes   (b) isobars   (c) isotones   (d) isomerism

2. Deuterium nucleus consists of

(a) 2 protons only   (b) one neutron

(c) one proton and one neutron   (d) 2 protons and one neutron

3. Deuterium with oxygen gives

(a) oxydeuterium   (b) water   **(c) heavy water**   (d) all the above

4. Tritium is prepared by bombarding lithium with

(a) deuterons   (b) mesons   **(c) slow neutrons**   (d) all helium nucleus

5. At room temperature ordinary hydrogen consists of about

(a) 25% para and 75% ortho   (b) 75% para and 25% ortho

(c) 99% para and 1% ortho   (d) 1% para and 99% ortho

6. D_2O reacts with P_2O_5 and gives

(a) DPO_4 (b) D_2PO_4 (c) D_3PO_3 **(d) D_3PO_4**

7._____ is used for the preparation of deuterium

(a) deuterium oxide   (b) heavy water   **(c) both a and b** (d) deuterium peroxide

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8. H_2O_2 is a powerful agent
(a) dehydrating   **(b) oxidising**   (c) reducing   (d) desulphurising

9. is used as a propellant in nucleus
(a) H_2O_2   (b) D_2O (c) ND_3 (d) $CH_2 = CH_2$

10. The oxidation state of alkali metals is
(a) +2   (b) 0   **(c) +1**   (d) +3

11. When heated in bunsen flame, lithium gives colour
(a) yellow   (b) blue   (c) lilac   **(d) crimson red**

12. On moving down the group, density of the alkali metals
(a) increases   (b) decreases   (c) increases and then decreases
(d) decreases and then increases

13. If the element can lose an electron readily, they are said to be
(a) electronegative   **(b) electropositive**   (c) electronegative   (d) electrovalent

6. Group 2S -Block elements

1. Among the following, which is known as `alkaline earth metal'.
(a) Sodium   **(b) Calcium**   (c) Lithium   (d) Potassium

2. Alkaline earth metals are
(a) monovalent   (b) trivalent   **(c) divalent**   (d) zero valent

3. Among alkaline earth metals _____ is having the highest ionization energy.
(a) Beryllium   (b) magnesium   (c) Calcium   (d) Barium

4. The colour given by barium in flame is
(a) Brick red   **(b) Apple Green**   (c) Red   (d) Blue

5. The third most abundant dissolved ion in the ocean is
(a) Beryllium   (b) Barium   (c) Calcium   **(d) Magnesium**

6. Quick lime is
(a) Calcium oxide   (b) Calcium hydroxide
(c) Calcium nitrate   (d) calcium sulphate

7. The formula of bleaching powder is
(a) $CaCl_2. H_2O$ **(b) $CaOCl_2. H_2O$**   (c) $CaSO_4. 2H_2O$ (d) $CaSO_4. ½H_2O$

8. Plaster of paris is
(a) $CaSO_4. 2H_2O$ (b) $CaCl_2$ (c) $CaSO_4$ **(d) $CaSO_4. H_2O$**

9. The compound used in making moulds for statues is
(a) Epsom salt   (b) Calcium sulphide   **(c) Plaster of paris**   (d) Gypsum

10. The element used in pyrotechnics is
(a) Magnesium   (b) Barium   (c) Calcium   (d) Beryllium

7. P – Block Elements

1) The elements of group 13 to 18 of the periodic table are known as
a) s - block elements   **b) p - block elements**   c) d - block elements   d) f - block elements

2) The general electronic configuration of group 18 elements is
a) ns^2 b) $ns^2 np^1$ c) $ns^2 np^{1-5}$ **d) $ns^2 np^6$**

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3) The basic oxide among the following
a) Bi_2O_3   b) SnO_2 c) HNO_3 d) SO_3

4) The most stable hydride of the following
a) NH_3   b) PH_3 c) AsH_3 d) BiH_3

5) The formula of Borax is
          a) $NaBO_2$ **b) $Na_2B_4O_7$**   c) H_3BO_3 d) None of the above

6) The general electronic configuration of carbon group elements is
a) ns^2np^6 b) ns^2 c) ns^2np^1 **d) ns^2np^2**

7) The process used for the manufacture of ammonia is
a) Contact process   b) Ostwald process   **c) Haber's process**   d) Linde's process

8. The oxides of non-metals are usually
a) ionic   b) coordinate   **c) covalent**   d) none of the above

9. Metallic oxides are generally
a) acidic   **b) basic**   c) amphoteric   d) neutral

10. Fixation of nitrogen is a source for
a) Various oxygen compounds   b) Various phosphorus compounds
c) Various nitrogen compounds   d) Various sulphur compounds

11. The oxyacid of nitrogen which is used in the manufacture of azo dyes.
a) Nitrous acid   b) Nitric acid   c) Hyponitrous acid   d) Pernitric acid

12. The hydride of V group element which is used in the manufacture of artificial silk
a) ammonia   b) stibine   c) phosphine   d) bismuthine

13. Anaesthetic used for minor operation dentistry
a) nitrous oxide   b) nitric oxide   **c) nitrous oxide + oxygen**   d) nitrogen dioxide

14. An allotrope of carbon discovered by Richard Smalley *et al.*
a) graphite   b) diamond   **c) fullerene**   d) carbon black

8. Solid State – I

1. The structure of sodium chloride crystal is:
 (a) body centred cubic lattice
(b) face centred cubic lattice
(c) octahedral
(d) square planar

2. The number of atoms in a face centred cubic unit cell is:
(a) 4   (b) 3   (c) 2   (d) 1

3. The 8:8 type of packing is present in:
(a) CsCl   (b) KCl   (c) NaCl   (d) MgF2

4. In a simple cubic cell, each point on a corner is shared by
(a) 2 unit cells   (b) 1 unit cells   **(c) 8 unit cells**   (d) 4 unitcells

5. An amorphous solid is :
(a) NaCl   (b) CaF_2 **(c) glass**   (d) CsCl

6. Each unit cell of NaCl consists of 4 chlorine ions and:

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(a) 13 Na atoms   **(b) 4 Na atoms**   (c) 6 Na atoms   (d) 8 Na atoms

7. In a body centred cubic cell, an atom at the body of centre is shared by:

(a) 1 unit cell   (b) 2 unit cells   (c) 3 unit cells   (d) 4 unit cells

8. In the sodium chloride structure, formula per unit cell is equal to

(a) 2   (b) 8   (c) 3   **(d) 4**

9. In a face centred cubic cell, an atom at the face centre is shared by:

(a) 4 unit cell   **(b) 2 unit cells**   (c) 1 unit cells   (d) 6 unit cells

9. Gaseous State – I

1. A curve drawn at constant temperature is called an isotherm. This shows relationship between

(a) P and 1/V   (b) PV and V   **(c) P and V**   (d) V and 1/P

2. The critical temperature of a gas is that temperature

(a) Above which it can no longer remain in the gaseous state

(b) Above which it can not be liquified by pressure

(c) At which it solidifies   (d) At which volume of gas becomes zero.

3. A gas deviates from ideal behavior

(a) high T   (b) low P   (c) high T and low P   **(d) low T and high P**

4. If a gas expands at constant temperature.

(a) Number of molecules of the gas decreases

(b) The kinetic energy of the molecules decreases

(c) The kinetic energy of the molecules decreases

(d) The kinetic energy of the molecules increases

5. The molecules of a gas A travel four times faster than the molecules of gas B at the same temperature. The ratio of molecular weight (MA/MB) will be

(a) 1/16   (b) 4   (c) 1/4   **(d) 16**

10. Chemical Bonding

1. The crystal lattice of electrovalent compounds is composed of

(a) Atoms   (b) Molecules   **(c) Oppositely charged ions**   (d) Both molecules and ions

2. The compound which contains both ionic and covalent is

(a) CH_4 (b) H_2 **(c) KCN**   (d) KCl

11. Colligative Properties

1. Properties which depend only on number of particles present in solution are called

(a) Additive   (b) Constitutive   **(c) Colligative**   (d) None

2. Which solution would possess the lowest boiling point

(a) 1% NaCl solution   (b) 1% Urea solution   (c) 1% glucose solution

(d) 1% sucrose solution

3. In cold countries, ethylene glycol is added to water in the radiators of cars during winters. It results in : (a) Lowering boiling point   (b) Reducing viscosity

(c) Reducing specific heat   **(d) Lowering freezing point**

4. Which of the following 0.1M aqueous solutions will have the lowest freezing point?

(a) Potassium sulphate   (b) Sodium chloride   (c) Urea   (d) Glucose

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5. The Van't Hoff factor of 0.005M aqueous solution of KCl is 1.95. The degree of ionisation
of KCl is (a) 0.94   **(b) 0.95**   (c)0.96   (d) 0.59

12. Thermodynamics – I

1. Which of the following is not a state functions?

(a) q   (b) q + w   (c) ΔH (d) V +PV

2. Which of the following is an extensive property?

(a) volume   (b) density   (c) refractive index   (d) molar volume

3. Which of the following is an exothermic reaction?

(a) melting of ice   **(b) combustion reactions**   (c) hydrolysis   (d) boiling of water

4. Which of the following is reversible process?

(a) Diffusion   **(b) melting**   (c) neutralization   (d) combustion

5. In which process, work is maximum?

(a) reversible   (b) irreversible   (c) exothermic   (d) cyclic

13. Chemical Equilibrium – I

1. In which equilibrium pressure has no effect

(a) $PCl_{5(g)} \leftrightarrow PCl_3(g) + Cl_2(g)$ **(b) $H_2(g) + I_2(g) \leftrightarrow 2HI(g)$**

(c) $2SO_2(g) + O_2(g) \leftrightarrow 2SO_3(g)$ (d) $NH_4Cl(g) \leftrightarrow NH_3(g) + HCl(g)$

2. For the equilibrium $N_2O_4(g) \leftrightarrow 2NO_2(g)$, the Kp and Kc values are related as

(a) Kp = Kc(RT)   (b) $Kp = Kc (RT)^2$ (c) $Kp=Kc (RT)^{-1}$ (d) $Kp=Kc (RT)^{-2}$

3. For endothermic equilibrium, increase in temperature changes the Keq value as

(a) No change   **(b) Increases**   (c) Decreases   (d) None

4. In the heterogeneous equilibrium $CaCO_3(s) \leftrightarrow CaO(s) + CO_2(g)$ the Keq value is given by

(a) partial pressure of CO_2   (b) activity CaO

(c) activities of $CaCO_3$ (d) $[CaO]/[CaCO_3]$

5. For the equilibrium reaction $H_2(g) + I_2(g) \leftrightarrow 2HI(g)$

(a) Kp = Kc   (b) Kp > Kc   (c) Kp < Kc   (d) Kp = 1/Kc

14. Chemical Kinetics – I

1. $mol.dm^{-3} sec^{-1}$ is the unit of

(i) rate   (ii) rate constant   (iii) order   (iv)   active
mass

2. The elementary step with slow rate represents

(i) rate determining step   (ii) maximum rate step   (iii) third order rate

(iv) overall order

3. Molecularity is determined for

(i) an elementary reaction   (ii) an overall reaction

(iii) an over all stoichiometric reaction   (iv) a fractional order reaction

16. Purification of Organic Compounds

1. Organic compounds are soluble in

a) Non-polar Solvents   b) Polar solvents   c) Water   d) HCl

2. Decolourisation of coloured compounds can be effected by using

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a) Animal charcoal   b) Carbon   c) Coke   d) Infra-red rays

3. Compounds having boiling points widely apart 40 K and above can be purified by

a) Crystallisation   **b) Simple distillation**   c) Fractional distillation   d) Sublimation

4. Nitrobenzene and benzene can be separated by the method of

a) Simple distillation b) Crystallisation   c) Fractional crystallisation   d) Chromatography

5. Purification of two miscible liquids possessing very close boiling points can be separated using

a) Fractional distillation   b) Sublimation   c) Simple distillation d) Steam distillation

6. Purification of mixture of compounds can be done by steam distillation only if the impurities are

a) Non-volatile   b) Volatile   c) Insoluble in Water   **d) both a & c**

7. When the stationary phase is solid, then the compounds can be separated on the basis of

a) Adsorption   b) Partition   c) Both partition and adsorption d) Either

8. Column Chromatography is based on the principle of

a) Adsorption   b) Partition   c) Absorption   d) Distribution

9. In Ascending paper Chromatography, the solvent moves

a) Upwards   b) Downwards   c) Horizontally   d) None

10. The existence of wide range of organic compounds is due to their, property of

a) Extensive catenation   b) Lower boiling points   c) Polymerisation   d) Isomerism

18. Hydrocarbons

1) Alkanes can be represented by the formula

a) C_nH_{2n+2}   b) C_nH_{2n-2} c) C_nH_{2n} d) C_nH_{2n-3}

2) Alkenes are represented by the formula

a) C_nH_{2n+2} b) C_nH_{2n-2} **c) C_nH_{2n}**   d) C_nH_{2n-3}

3) Alkynes are represented by the formula

a) C_nH_{2n+2} **b) C_nH_{2n-2}**   c) C_nH_{2n} d) C_nH_{2n-3}

4) The type of substitution reaction that takes place when methane is treated with Cl_2 in presence of light   a) ionic   b) nucleophilic   c) electrophilic   **d) radial**

5) When n-hexane is passed over hot alumina supported chromium, vanadium or molybdenum oxide the compound formed is

a)cyclopentane   b)toluene   **c)cyclohexane**   d)benzene

6) When the identical groups are on the same or opposite sides of the bonds in alkenes the isomerism is called as

 a) chain isomerism   **b) geometrical isomerism**

c) position isomerism   d) optical isomerism

7) Diels-Alder reaction is the reaction between

a) diene and dienophile   b) electrophile and nucleophile   c) oxidant and reductant

d) none.

8) Unsaturated compounds with two double bonds are called as

a) diene   b) olefins   **c) alkadiene**   d) paraffins.

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9) The hybridization of carbons in ethylene is

a) sp^2   b) sp^3 c) sp   d) dsp^2

10) Alcohols can be dehydrated to olefins using

a) H_2SO_4   b) Pd   c) $SOCl_2$ d) Zn/Hg

11) When alkyl halides are treated with alcoholic KOH, the products are

a) olefins   b) alcohols   c) alkanes   d) aldehydes

12) Witting reaction is used to prepare

a) an alkene   b) an alkane   c) an alkyne   **d) none of the above**.

13) Electrolysis of potassium succinate gives

a) ethylene   b) acetylene   c) ethane   d) none of the above.

19. Aromatic Hydrocarbons

1. Aromatic compounds are

a) benzenoid compounds   b) non-benzenoid compounds

c) aliphatic compounds   d) alicyclic compounds

2. Benzene was first isolated by   a) Huckel   **b) Faraday**   c) Hofmann   d) Barthelot

3. Benzene undergoes   a) addition reactions   b) oxidation reactions

c) polymerisation reactions   **d) electrophilic substitution reactions**

4. The modern theory of aromaticity was introduced by

a) Faraday   b) Hofmann   **c) Huckel**   d) Berthelot

5. Any compound can be aromatic if they have ____ delocalized electrons.

a) $4n + 2$ b) $4n + 1$ c) $4n$ d) $4n - 2$

6. The function of $FeCl_3$ in chlorination of benzene is to produce

a) Cl   **b) Cl^+**   c) Cl^- d) C

7. The ortho and para directing groups are

a) activating group   b) deactivating group   c) both   d) none

8. The purpose of adding conc. H_2SO_4 in nitration of benzene is to produce

a) NO_2 b) NO_2^- **c) $NO_2^\pm$**   d) NO_3^-

9. An example of polycyclic aromatic hydrocarbon

a) pyridine   b) pyrole   **c) naphthalene**   d) cyclohexane

10. The compound which is used as a solvent for the extraction of fats and oils

a) naphthalene   **b) benzene**   c) cyclohexane   d) butane

20. Organic Halogen Compounds   CH3 .CH - CH2 - CH-CH3 is

                                                          Br   Cl

a. 2-Bromo-3-chloro-4-methylpentane   b. 2-Methyl-3-chloro-4-bromopentane

c. 2-Bromo-3-chloro-3-isopropyl propane   d. 2,4-Dimethyl-4-Bromo-3-chlorobutane.

2. For reacting with HCl, the alcohol which does not require ZnCl2 is

a. $CH_3 CH_2 OH$ b. $CH_3CH_2CH_2OH$ c. $CH_3CH (CH_3)OH$ **d. $C (CH_3)_3C-OH$**.

3. For converting alcohols into alkyl halides, the best reagent is

a. PCl_3 **b. PCl_5**   c. $SOCl_2$ d. None of the above

4. The olefin, which is not important for Markovni Koff's addition of HCl, is

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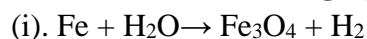**(ACTC)   ADVANCED CHEMISTRY TUITION CENTRE, PWD ROAD, NAGERCOIL 9940847892.**

a. Propene   b. But-1-ene   c. 2-Methyl-propene   **d. Ethylene**

5. The SN1 reaction of alkyl halides is not affected by the nature of the

a. alkyl group   b. the halogen   c. medium   **d. nucleophile**

LESSON 1

B. Fill in the blanks

1. One mole of a triatomic gas contains **$3 \times 6.023 \times 10^{23}$** atoms.

2. One mole of Sulphuric acid contains **$4 \times 6.023 \times 10^{23}$** Oxygen atoms.

3. 11.2 L of carbon dioxide at S.T.P contains **6.023×10^{23}** oxygen atoms.

4. Equal volumes of different gases under similar conditions of temperature and pressure contain equal number of **molecules**

5. A decimolar solution of NaOH contains **4g** of NaOH per litre of the solution.

6. 7 g of CO contains **1.505×10^{23}** O atoms.

7. The mass of 1 x 1022 formula units of CuSO4 5H2O is **4.142g**

C. Match the following

Column A	Column B
1. CaC_2	Calcium carbide
2. Law of multiple proportions	John Dalton
3. Hydrargyrum	Liquid element
4. 2 gm-equivalents of Na_2CO_3	106 g
5. 22.4 L at S.T.P	6.02×10^{23} Ne atoms
6. Number of gm-molecules per litre of solution	Molarity of solution
7. 1 gm-atom of rhombic sulphur	1/8 gm-molecules
8. Centimolar solution	0.01 moles of solute in one L of solution
9. Mohr's Salt	$(NH_4)_2SO_4.Fe(SO_4).6H_2O$

D. Answer the following

1. Can two different compounds have same molecular formula ? Illustrate your answer with two examples.

Two different compounds can have the same molecular formula. They are said to be isomers.

(e.g.,)   (i)   **C_2H_6O**. This molecular formula is possessed by ethanol

C_2H_5OH, Dimethyl ether **$CH_3– O – CH_3$**

(ii) C_3H_6O. This is common molecular formula for Acetone

$$CH_3 – C – CH_3$$
$$||$$ and for propionaldehyde **$CH_3 – CH_2 – CHO$**
$$O$$

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2. What are the essentials of a chemical equation ?

A chemical equation is the short scientific representation of a chemical reaction. In order to write the chemical equation correctly, we must know the reacting substances, all the products formed and their chemical formulae.

3. What are the informations conveyed by a chemical equation ?

(i)   A chemical equation is the stoichiometric equation which is a short scientific representation of a chemical reaction.

(ii)   The chemical equation explains the relationship between the number of mole of the reactants and products of a chemical reaction.

(iii)   The chemical equation explains the conditions at which the reaction take place such as temperature, pressure, catalyst, etc.

4. Balance the following equations

(i). $Fe + H_2O \rightarrow Fe_3O_4 + H_2$

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

(ii). $Fe_2 (SO_4)_3 + NH_3 + H_2O \rightarrow Fe(OH)_3 + (NH_4)_2 SO_4$

   $Fe_2 (SO_4)_3 + 6 NH_3 + 6 H_2O \rightarrow 2Fe (OH)_3 + 3 (NH_4)_2SO_4$

(iii). $KMnO_4 + H_2SO_4 \rightarrow K_2SO_4 + MnSO_4 + H_2O + O_2$

   $2 KMnO_4 + 3 H_2SO_4 \rightarrow K_2SO_4 + 2 MnSO_4 + 3 H_2O + 5/2 O_2$
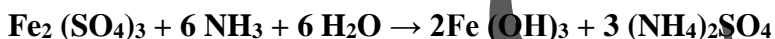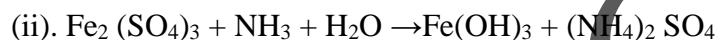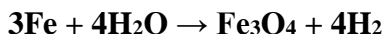
   This equation is to be multiplied by 2 since we cannot have fractional molecule as $5/2 O_2$

   $4 KMnO_4 + 6 H_2SO_4 \rightarrow 2 K_2SO_4 + 4 MnSO_4 + 6 H_2O + 5 O_2$

(iv). $K_2Cr_2O_7 + H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + H_2O + O_2$

   $2 K_2Cr_2O_7 + 8 H_2SO_4 \rightarrow 2 K_2SO_4 + 2 Cr_2 (SO_4)_3 + 8 H_2O + 3 O_2$

2. GENERAL INTRODUCTION TO METALLURGY

1. **Difference between mineral and ore**.

No	**Mineral**	**Ore**
i)	The natural material in which the metal or their compounds occur in the earth is known as **mineral**.	A mineral from which a metal can be profitably extracted is called an **ore.**
ii)	**All ores are minerals**	**All minerals are not ores**
iii)	clay (Al_2O3 $2SiO_2$ $2H_2O$)	bauxite (Al_2O_3 $2H_2O$)

2. What is matrix?

The ore is generally associated with rock impurities like clay, sand etc. called **'**gangue or matrix.

3. What is mineral?

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The natural material in which the metal or their compounds occur in the earth is known as **mineral**. .

4. What is mining?

The biggest source of metal is the earth's crust and the process of taking out the ores from the earth crust is called mining.

5. What are the mineral Source from sea.

Four elements such as Na, Mg, Cl_2 and Br_2 can be extracted from the oceans or salt brines, where they are present as monoatomic ions (Na^+, Mg^{2+}, Cl^-, Br^-).

6. Explain Gravity separation process or hydraulic washing

This method is especially suitable for heavy 'oxide' ores like haematite, tinstone, etc. In this, the powdered ore is placed on a sloping floor (or platform) and washed by directing on it a strong current of water. The lighter sandy, and earthy impurities are washed away; while the heavier ore particles are left behind.

7. Explain Froth flotation process

- This method is especially suitable for sulphide ores like zinc blende (ZnS), and copper pyrites (CuFeS2).
- This process is based on the fact that the sulphide ore particles are only moistened by **oil;**
- oxide, and gangue particles are moistened only by **water.**
- In this process, the powdered ore is mixed with water and a little pine oil (a foaming agent) and the whole mixture is then stirred vigorously by blowing compressed air.
  The oil forms a foam (or froth) with air. The ore particles stick to the froth, which rises to the surface; while the rocky, and earthy impurities (gangue) are left in water
- The froth is skimmed off, collected, and allowed to subside to get concentrated ore.

8. How will you separate Magnetic ore from non-Magnetic ore

- This method is meant for separating **magnetic impurities from non- magnetic ore particles,** e.g., tinstone (a tin ore) in which tinstone is non- magnetic;
- Theimpurities iron, manganese and tungstates are magnetic.
- The powdered ore (containing the associated magnetic impurities) is made to fall (from a hopper) on a belt moving over electromagnetic roller.
- The magnetic impurities fall from the belt in a heap near the magnet, due to attraction;
  the non-magnetic concentrated ore falls in separate heap, away from the magnet, due to the influence of centrifugal force.
  (Diagram refer book page no41)

9. Explain Chemical method

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- This method is employed in case where the ore is to be in a very pureform, e.g., aluminium extraction. Bauxite (Al_2O_3), an ore of aluminium, contains SiO_2 and Fe_2O_3 as impurities.
- A bauxite ore is treated with NaOH, the Al_2O_3 goes into solution as sodium meta- aluminate leaving behind the undissolved impurities [Fe_2O_3, SiO_2,$Fe(OH)_3$,etc.], which are then filtered off.

    $Al_2O_3+ 2NaOH \rightarrow 2NaAlO_2 +H_2O$
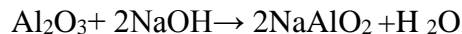
    The filtrate (containing sodium meta-aluminate) on dilution, and stirring gives a precipitate of aluminium hydroxide, which is filtered, and ignited to get pure alumina.

    $NaAlO_2 +2H_2O \rightarrow Al(OH)_3+NaOH$

    $2Al(OH)_3 \rightarrow Al_2O_3+3H_2O$
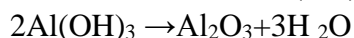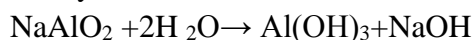
10. What are the Metallurgical processes

Metallurgy is a branch of chemistry which deals with,

(i) Extraction of metals from ores

(ii) Refining of crude metal

(iii) Producing alloys and the study of their constitution, structure and properties.

(iv) The relationship of physical and mechanical treatment of metals to alloys.

11. What are the metal are extracted from electrolysis method

The noble metals such as Au, Ag, etc are usually extracted by electrolysis of their chlorides.

12. What are the ore is separated by Roasting method

Oxides or hydroxides. Heavy metals, e.g. Cu, Zn, Fe, Pb, Sn, etc., are extracted by making use of roasting and smelting methods.

13. Define Roasting.

Roasting one of the oxidation method where ore is converted into metal oxide.

(a) Volatile impurities like S, As, Sb etc. get oxidized.

b) The sulphide ores decompose to SO_2

(c) The moisture is removed.

$2ZnS + 3O_2 \rightarrow 2ZnO + 2SO_2$

$2PbS + 3O_2 \rightarrow 2PbO + 2SO_2$
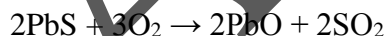
14. What is Calcination

Another method of conversion of ore into metal oxide (oxidation) is called calcination. (absence of air).

$CaCO_3$ (limestone) $\rightarrow CaO + CO_2$

$MgCO_3$ (Magnesite) $\rightarrow MgO + CO_2$
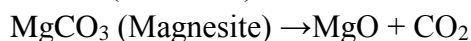
15. What is Smelting – Reduction

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Smelting is one of reduction method where the metal oxide is converted into metal is called as Smelting.

$$Fe_2O_3 + 3C \rightarrow 2Fe + 3CO$$
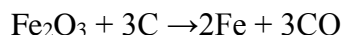
16. Explain Bessemerisation process

Principle:

The principle involved in this process is that cold air blowed through refractory lined vessel known as converter containing molten pig iron at about 2 atmospheric pressure, oxidizing the impurities and simultaneously converting pig iron to steel.

Procedure:

The molten pig iron is mixed in mixers and then charged into converter. About 15-16 tonnes of iron can be charged at a time. The converter is first set in the horizontal position and after charging the converter is adjusted in vertical position.

After charging a blast of cold air is admitted through the hole provided at the bottom at a pressure of about 2-3 kg/cm3. The blast is continued for about 15 minutes during which the impurities are oxidized. Mn is oxidized to MnO and Si is oxidized to SiO_2. Carbon is also oxidized to CO. The resulting oxides of Mn and Si (MnO and SiO_2) combine together to form slag of manganese silicate:

Diagram:

17. What is matte?

A mixture containing sulphide of copper and iron, called **matte.**

18. What is mean by Anode mud?

The insoluble impurities either dissolve in the electrolyte or fall at the bottom and collect as **anode mud**.

19. Explain electrolyting refining of copper?

Anode: Impure copper metal

Cathode:Pure copper

Electrolyte: Copper sulphate and Sulphuric Acid On passing electric current a pure copper is deposited on the cathode side

Reaction at Anode and cathode:

1) Cu^{2+} ions (from copper sulphate solution) go to the cathode (negative electrode), where they are reduced to copper, which gets deposited on the cathode.

$Cu^{2+} + 2e- \rightarrow Cu$
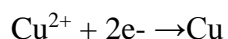
2) Copper (of impure anode) forms copper ions, and these go into solution of electrolyte.

$Cu \rightarrow Cu^{2+} + 2e-$
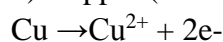
20. Explain Zone refining

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Principle:
Melting point of a substance is lowered by the presence of impurities. Consequently, when an impure molten metal is cooled, crystals of the pure metal are solidified, and the impurities remain behind the remaining metal.

Diagram:
Procedure:
- This method is employed for preparing highly pure metal (such as silicon, tellurium, germanium), which are used as semiconductors
- The process consists in casting the impure metal in the form of a bar.
- A circular heater fitted around this bar is slowly moved longitudinally from one end to the other.
- At the heated zone, the bar melts, and as the heater moves on, pure metal crystallizes, while the impurities pass into the adjacent molten part.
- In this way, the impurities are swept from one end of the bar to the other. By repeating the process, ultra pure metal can be obtained.

21. Explain Mond's process
- Thermal methods include methods as carbonyl method, decomposition of hydrides etc.
- The carbonyl method is used for the refining of metals like Ni and Fe.
- For example, in case of nickel, the impure metal is heated with CO.
- The nickel carbonyl thus formed is then decomposed (after distilling off the impurities) to get pure nickel metal and CO. The process is known as **Mond's process.**

    $Ni + 4CO \rightarrow Ni(CO)4 \rightarrow Ni + 4CO$

Based on the following facts:
(a) Only nickel (and not Cu, Fe, etc.) forms a volatile carbonyl, $Ni(CO)_4$, when CO is passed over it at $50°C$.
(b) the nickel carbonyl decomposes at $180°C$ to yield pure nickel.

22. What is Acid Bessemer process
The impurities present in the pig iron are basic, e.g., manganese, a lining of silica brick is used and the process is known as **acid Bessemer process.**

23. What is basic Bessemer process.
If impurities are acidic, e.g., sulphur, phosphorus etc., a basic lining of lime (CaO) or magnesia (MgO) is used in the converter and process is then known as **basic Bessemer process.**

A. Fillup the blanks
1. The earthy impurities associated with ores are **gangue (or) matrix**
2. Froth flotation process is suitable for concentrating **Sulphide** ores.
3. Highly pure metals are obtained by **zone refining** process.
4. Gangue + flux → **Slag**

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5. A mineral from which metal can be profitably extracted is called **ore**

6. A mixture containing sulphides of copper and iron is called **matte**

7. **Pine oil** is used as a foaming agent.

B. Write in one or two sentence

1. Distinguish between ore and mineral with suitable example ?

No	Mineral	Ore
(i)	It is single compound or a complex mixture of various compounds	It is a minerals containing sufficient amount of the metal from which a metal can be profitably and readily extracted.
(ii)	Minerals of aluminium are clay (Al_2O_3. $2SiO_2$. $2H_2O$) and Bauxite (Al_2O_3. $2H_2O$)	The ore of aluminium is Bauxite because from Bauxite only Al can be extracted profitably.

2. What are the elements obtained from sea water source?

Chlorine, Bromine, Iodine, Sodium, Potassium , Calcium, Magnesium are obtained from sea water source.

3. What are the different methods of concentration of ores?

  (i) Hand picking

  (ii) Gravity separation process or hydraulic washing

  (iii) Forth floatation

  (iv) Electromagnetic separation process

  (v) Chemical method

4. What is gravity separation ?

Gravity separation process is the process of concentration of heavy oxide ores like haematite, tinstone, etc. In this process, the powdered ore is washed by a strong current of water. The lighter sandy and earthy impurities are washed away while the heavier ore particles are left behind.

5. Name the ores which are concentrated by froth floatation process.

Sulphide ores like zinc blende (ZnS), copper pyrites ($CuFeS_2$) and Galena (PbS) are concentrated by forth floatation process.

6. Define Metallurgy.

Metallurgy is the process of separation or extraction of the metal from its ore by applying various steps such as roasting, smelting and electrolytic refining, etc.

7. What are the major steps involved in the metallurgical process ?

(i) Concentration, (ii) roasting or calcination, (iii) smelting, (iv) reduction, (v) Purification by chemical or physical method.

8. What is calcination? Give example.

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Calcination is another method of connecting the ore into metal oxide in which the ore is subjected to the action of heat at high temperature in the absence of air below its melting point. The process of calcination is carried out in the case carbonate and hydrated ore. As a result of calcination, moisture is removed, gaseous impurities are removed the mass becomes porous and thermal decomposition of the ore takes place

(e.g.,)   CaCO₃   →   CaO + CO₂ ↑
**   Lime stone**
**   2Fe₂O₃. 3H₂O   →   2Fe₂O₃ + 3H₂O**
**   Limonite**

9. What is the principle involved in Bessemer process ?

The principle involved I Bessemer process is that cold air blowed through refractory lined vessel known as converter containing molten big iron at about 2 atmospheric pressure oxidising the impurities and simultaneously converting big iron to steel.

10. What is meant by electrolytic refining? Give example.

Electrolytic refining is the method of refining and gives a metal of high purity by applying electric current through the metallic salt solution (electrolyte) in which impure metal act as anode and pure metal act as cathode.

(e.g.,) electrolytic refining is used for Cu, Au, Ag, Pb, Zn and Al.

In the electrolytic refining of copper, copper sulphate and dilute H2SO4 is used as electrolyte. Impure copper is taken as anode and pure copper is taken as cathode. By passing electric current through the electrolyte, pure metal is deposits on cathode.

    At Anode: **Cu   →   Cu²⁺ + 2e⁻**
    At Cathode: **Cu²⁺ + 2e⁻   →   Cu**
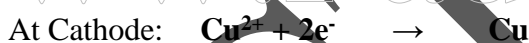
The net result is the transfer of pure copper from anode to cathode.

11. What is anode mud?

In the process of electrolytic refining, metal ions undergo reduction and pure metal is deposited at the cathode. The insoluble impurities either dissolve in the electrolyte or fall at the bottom and collect as anode mud. For example, in the refining of copper, impurities like Fe and Zn dissolve in the electrolyte, white Au, Ag and Pt are left behind as anode mud.

12. What do you understand by the following terms (i) roasting (ii) smelting

(i)Roasting : It is one of the oxidation method, where the ore is converted into metal oxide. In the process of roasting, the ore either alone or with the addition of suitable material is subjected to the action of heat in excess of air at a temperature below its melting point.

(ii)Smelting : It is one of the reduction method, where the metal oxide is converted into metal. The process of smelting is that in which the ore is melted with a flux and with a reducing agent and it involves calcination, roasting and reduction.

C. Explain briefly on the following

1. Write short note on source of element in living system.
2. Explain froth flotation process with neat diagram.

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3. How electrolytic separation process is useful in the separation of magnetic impurities from nonmagnetic ores? Draw the diagram.

4. How the impurities of ore are removed by chemical method?

5. What is roasting ? Explain different types of roasting with suitable example.

6. What is smelting ? Explain the process with example.

7. What is Zone refining ? Describe the principle involved in the purification of the metal by this method.

8. How nickel is extracted by Mond's process? Write the various reactions involved in the process.

9. Write short note on mineral wealth of India.

10. Give a brief account of the mineral wealth of Tamil Nadu.

ATOMIC STRUCTURE

Explain the four types of quantum numbers in detail.

The numbers which designate and distinguish various atomic orbitals and electrons present in an atom are called quantum numbers.

In an atom, the state of each electron is different with respect to the nucleus. In order to define the state of the electron completely, four quantum numbers are used.

They are - 1. Principal quantum number (n)

2. Azimuthal quantum number (l)

3. Magnetic quantum number (m)

4. Spin quantum number (s).

1. Principal Quantum Number (n):

1. It determines the energy shell in which the electron is revolving around the nucleus. It is also known as major energy level.

2. It is denoted by the symbol n and may have any integral value except zero. *i.e.,* it can have the value $n =1,2,3, \ldots.$ etc.

3. The value $n = 1$ denotes that the electron is in the first shell (K shell).

The value $n = 2$ denotes that the electron is in the second shell (L shell).

The value $n = 3$ denotes that the electron is in the third shell (M shell).

The value $n = 4$ denotes that the electron is in the fourth shell (N shell).

4. As the distance of the electron from the nucleus increases, its energy becomes higher and higher.

5. The maximum number of electrons in a major energy level is given by $2n^2$.

Principal quantum number 'n'	Designation	Maximum number of electrons ($2n^2$)

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1	K	2
2	L	8
3	M	18
4	N	32

2. Azimuthal Quantum Number or Orbital Quantum Number (l):

1. It represents the **sub shell** to which the electron belongs.

2. It is denoted by the symbol l. Its value depends on the principal quantum number n. It may have any value ranging from 0 to $(n-1)$.

3. The value $l = 0$ denotes that the electron is in the s sub shell or s orbital.

The value $l = 1$ denotes that the electron is in the p sub shell or p orbital.

The value $l = 2$ denotes that the electron is in the d sub shell or d orbital.

The value $l = 3$ denotes that the electron is in the f sub shell or f orbital.

Principal quantum number 'n'	l Value $l = (n\text{-}1)$	Name of the sub shells or orbital
1	0	1s
2	0	2s
	1	2p
3	0	3s
	1	3p
	2	3d
4	0	4s
	1	4p
	2	4d
	3	4f

3. Magnetic Quantum Number (m):

1. It represents the **orientation** of an atomic orbital in space.

2. It is denoted by the symbol m. The possible value which m can have depends upon the value of l. The values are -'l' through zero to +'l' and thus there are $(2l+1)$ values.

3. Its value tells the orientations of orbital in space. The value of $m = 0$ denotes that the orbital has **no orientation**. The value of $m = 1$ denotes that it has three orbital with **three types** of orientations. The value of $m = 2$ denotes that it has five orbital with **five types** of orientations.

Principal quantum number 'n'	l Value $l = (n\text{-}1)$	m value $(\text{-} l \ ... \ O \ .... + l \ )$	Name of the sub shells or orbital with orientation
1	0	0	1s
2	0	0	2s
	1	-1 0 +1	$2p_x$, $2p_y$, $2p_z$
3	0	0	3s

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	1	-1 0 +1	$3p_x$, $3p_y$, $3p_z$
	2	-2 -1 0 +1 +2	$3d_{xy}$, $3d_{xz}$, $3d_{yz}$, $3dz^2$, $3dx^2 - y^2$

4. Spin Quantum Number (*s*):

1. It represents the direction of the spin of the electrons.

2. It is denoted by the symbol s.

3. The electron may spin in the clockwise - direction or anticlockwise direction. And hence it can have only two values namely either + ½ or –½.

4. Two electrons with the same sign of spin are said to have parallel spins and are represented by (or) --while those having opposite spins are said to have anti parallel spins - and are known as paired up electrons.

	ORBIT	ORBITAL	
1.	It is a definite circular path in which the electron is supposed to revolve around the nucleus.	It is the three dimensional region around the nucleus in which there is maximum probability of finding the electron.	
2.	It is circular in shape.	It has different three dimensional shapes. *Eg.* '*s*'-orbitals are spherical, *p*-orbitals are dump bell shaped etc.	
3.	An orbit can contain a maximum $2n^2$ electrons where *n* represents the order of the orbit from the nucleus.	An orbital can contain a maximum of only 2 electrons.	
4.	The position and momentum of the electron can be calculated at the same time.	It is not possible to find the exact position and momentum of the electron at the same time.	
5.	They are designated as K, L, M, N etc	They are designated as s, p, d, f etc	

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1.Define Atom

All matter is composed of very small particles called atoms

2.Define Orbit.

The nucleus is surrounded by electrons that move around the nucleus with very high speed in circular paths called **orbits**

3.Define Mass Number

Protons and neutrons present in the nucleus are collectively also known as nucleons. The total number of nucleons is termed as **mass number(A)** of the atom.

4.Define Atomic Number

A number of electron or The number of protons in an atom is called its atomic number **(Z).**

5.State Defects of Rutherford's model
•According toRutherford'smodel,an atom consist of a positive nucleus with the elctrons moving around it in circular orbits.
- However it had been shown by J. C. Maxwell that whenever an electron is subjected to acceleration, it emits radiation and loses energy.
- As a result of this, its orbit should become smaller and smaller and finally it should drop into the nucleus by following a spiral path.
- This means that atom would collapse and thus Rutherford's model failed to explain stability of atoms.
- Another drawback of the Rutherford's model is that it says nothing about the electronic structure of the atoms i.e., how the electrons are distributed around the nucleus and what are the energies of these electrons.
- Therefore, this model failed to explain the existence of certain definite lines in the hydrogen spectrum.
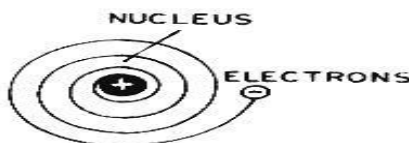
6.Explain Postulates of Bohr's model of an atom

- The electrons revolve round the nucleus only in certain selected circular paths called orbits. These orbits are associated with definite energies and are called energy shells or energy levels or quantum levels. These are numbered as 1, 2, 3, 4 ….. etc. (starting from the nucleus) are designated as K, L, M, N ….etc.

- As long as an electron remains in a particular orbit, it does not lose or gain energy. This means that energy of an electron in a particular path remains constant. Therefore, these orbits are also called stationary states.

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- Only those orbits are permitted in which angular momentum of the electron isa whole number multiple of h/2 ,Where 'h' is planck's constant. An electron moving in a circular orbit has an angular momentum equal to m r where m is the mass of the electron and , the angular momentum, m r is a Whole number multiplicity of h/2

mvr = nh/2   where n=1,2,3,…..

In other words, angular velocity of electrons in an atom is quantised.
- If an electron jumps from one stationary state to another, it will absorb or emit radiation Definite frequency giving a spectrl line of tht frequency whichdependsuponhe initial and final levels. When n electron jumps b ck to the lower energy level, it radia es
same amount of energy in the form of radi tion.
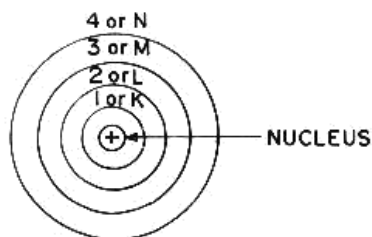
7.ExplainLimitation of Bohr's Theory

- Bohr Theory had explained the existence of various lines in H- spectrum, but it predicted that only a series of lines exist. At that time this was exactly what had been observed. However, as better instruments and techniques were developed, it was realized that the spectral line that had been thought to be a single line was actually a collection of several lines very close together (known as fine spectrum). Thus for example, the single H - spectral line of Balmer series consists of many lines very close to each other.
- Thus the appearance of the several lines implies that there are several energy levels, which are close together for each quantum number n. This would require the existence of new quantum numbers.
- Bohr's theory has successfully explained the observed spectra for hydrogen atom and hydrogen like ions (e.g. He^+, Li^{2+}, Be^{3+} etc.), it can not explain the spectral series for the atoms having a large number of electrons.
- There was no satisfactory justification for the assumption that the electron can rotate only in those orbits in which the angular momentum of the electron i.e. he could not give any explanation for using the principle of quantisation of angular momentum and it was introduced by him arbitrarily.
- Bohr assumes that an electron in an atom is located at a definite distance from the nucleus and is revolving round it with definite velocity, i.e. it is associated with a fixed value of momentum. This is against the Heisenberg's Uncertainty Principle according to which it is impossible to determine simultaneously with certainty the position and the momentum of a particle
- Does not explain Stark And Zeeman effect

8.State Zeeman Efeect:

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If a substance which gives line emission spectrum, is placed in a magnetic field, the lines of the spectrum get split up into   number of closely spaced lines. This phenomenon is known as Zeeman effect. Bohr's theory has no explanation for this effect.

9.State Stark 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. Bohr's theory is not able to explain this observation as well.

10.Define Quantum Numbers

The quantum numbers are nothing but the details that are required to locate an electron in an atom. In an atom a large number of electron orbitals are permissible. An orbital of smaller size means there is more chance of finding the electron near the nucleus. These orbitals are designated by a set of numbers known as quantum numbers.

11.Explain principal quantum number (n)

The electrons inside an atom are arranged in different energy levels called electron shells or orbits. Each shell is characterized by a quantum number called principal quantum number. This is represented by the letter 'n' and 'n' can have values, 1,2,3,4 etc.

12.Explain subsidiary or azimuthal quantum number (*l*)

According to Sommerfield, the electron in any particular energy level could have circular path or a variety of elliptical paths about the nucleus resulting in slight differences in orbital shapes with slightly differing energies due to the differences in the attraction exerted by the nucleus on the electron. This concept gave rise to the idea of the existence of sub- energy levels in each of the principal energy levels of the atom. This is denoted by the letter '*l*' and have values from 0 to n-1.
if n=1, *l*=0 only one value (one level only) *s* level.
n=2, *l*=0 and 1 (2 values or 2 sub- levels) *s* and *p* level.

13.Magnetic quantum number (m)

This explains the appearance of additional lines in atomic spectra produced when atoms emit light in magnetic field. Each orbitals is designated by a magnetic quantum number m and its values depends on the value of '*l*' . The values are -'*l*' through zero to +'*l*' and thus there are (2*l*+1) values.
Thus when *l*=0, m= 0 (only one value or one orbital)
l=1, m= -1, 0, +1 (3 values or 3 orbitals)

14.Spin quantum number(*s*)

The electron in the tom rotates not only around the nucleus but also around its own

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axis and two opposite directions of rotation are possible (clock wise and anticlock wise). Therefore the spin quantum number can have only two values $+1/2$ or $-1/2$. For each values of m including zero, there will be two values for *s*.

15.State Pauli's exclusion principle

"it is impossible for any two electrons in a given atom to have all the four quantum numbers identical

16.State Hund's rule of maximum multiplicity

No pairing occurs until all orbitals of a given sub- level are half filled. This is known as Hund's rule of maximum multiplicity.

17.State Aufbau Principle
In the ground state of the atoms, the orbitals are filled in order of their increasing energie**s**

(OR)
The lower the value of (n+1) for an orbital, the lower is its energy. If two orbitals have the same (n+1) value, the orbital with lower value of n has the lower energy.

18.Explain why the electronic configuration of Cr andCu are written as $3d^5$, $4s^1$ and $3d^{10}$ $4s^1$ instead of $3d^4$ $4s^2$ and $3d^9$ $4s^2$
Chromium
Expected configuration : $1s^2,2s^2,2p^6,3s^2,3p^6,3d^4,4s^2$
Actual configuration   : $1s^2,2s^2,2p^6,3s^2,3p^6,3d^5,4s^1$
Copper
Expected configuration : $1s^2,2s^2,2p^6,3s^2,3p^6,3d^9,4s^2$ Actual
configuration   : $1s^2,2s^2,2p^6,3s^2,3p^6,3d^{10},4s^1$
19.What are the charge and mass of an electron?
Charge of electron is Negative($1.602x10^{-19}$ C) and mass is $9.11x10^{-31}$ kg.
20.What is the charge of an electron,proton and neutron?
•   Electron - Neg tive. •   Proton - Positive, Neutron - Neutral
21.Explain Shape of Orbital

S-ORBITALS

• l = 0 ,m = 0. This means that the probability of finding the electron in s-orbital is the same in all directions at a particular distance. In other words s-orbitals are spherically symmetrical.
- The electron cloud picture of 1s-orbital is spherical.
- The *s*-orbitals of higher energy levels are also spherically symmetrical.
- They are more diffused and have spherical shells within them where probability of finding the electron is zero.
- These are called nodes. In 2*s*-orbital there is one spherical node.

p-orbital:

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- l = 1 m=+1, 0, -1. This means that there are three possible orientations of electron cloud in a *p*-sub-shell.
- The three orbitals of a *p*-sub-shell are designated as *px, py* and *pz* respectively along x-axis, y-axis and z-axis respectively. Each *p*-orbital has two lobes, which are separated by a point of zero probability called node.
- *p*-orbital is thus dumb bell shaped.
  ### *d*-orbitals:
- l = 2 , m = 0, ±1, ±2 indicating that d- orbitals have five orientations, i.e.,

- there are five *d*-orbitals which are as *dxy, dyz, dzx, dz^2* and dx^2-y^2. All these five orbitals, in the absence

  of magnetic field, are equivalent in energy and are, therefore, said to be five-fold degenerate

- The three orbitals namely *dxy, dyz* and *dzx* have their lobes lying symmetrically between the coordinate axes indicated in the subscript to d,
  e.g. the lobes of *dxy* orbital are lying between the x-and y-axes.
- lobes along the axes (i.e. along the axial directions), e.g., the lobes of d orbital lie along the x and y-axes, while those of *dz*
  the z-axis. This set is known as eg set.
  ### 22.What is stability of atom
  The extraordinary stability of half-filled and completely filled electron configuration can be explained in terms of symmetry and exchange energy. The half-filled and completely filled electron configurations have symmetrical distribution of electrons and this symmetry leads to stability

  ### 23.Mention the Uses of Pauli's exclusion principle
- The greatest use of the principle is that it is helpful in determining the maximum number of electrons that a main energy level can have.
- Let us illustrate this point by considering K and L shells.
  K-shell: For this shell n = 1. For n = 1, l = 0 and m = 0.
  Hence *s* can have a value either +1/2 or −1/2. The different values
 (i) n = 1, l = 0, m = 0 *s* = +1/2 (1st electron)
 (ii) n = 1, l = 0, m = 0, *s* = -1/2 (2nd electron)

- K shell there is only one sub- shell corresponding to l = 0 value (*s*-sub-shell) contains only two electrons with opposite spins.

- L-shell: For this shell n = 2. For n = 2 the different values of l, m and *s* give the following eight combinations of four quantum numbers.
  n= 2, l = 0, m = 0, *s* = +1/2
- Eight combinations given above show that L shell is divided into two sub-shells corresponding to l = 0 (*s* sub-shell) and l = 1 (p sub-shell) and this shell cannot contain more than 8 electrons, i.e., its maximum capacity for keeping the electrons is eight.

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25.Explain Rutherford's nuclear model of atom.

- An atom consists of a tiny positively charged nucleus at its centre.
- The positive charge of the nucleus is due to protons. The mass of the nucleus, on the other hand, is due to protons and some neutral particles each having mass nearly equal to the mass of proton.
- This neutral particle, called neutron, was discovered later on by Chadwick in 1932. Protons and neutrons present in the nucleus are collectively also known as nucleons. The total number of nucleons is termed as **mass number(A)** of the atom.
- The nucleus is surrounded by electrons that move around the nucleus with very high speed in circular paths called **orbits**.
- Thus, Rutherford's model of atom resembles the solar system in which the sun plays the role of the nucleus and the planets that of revolving electrons.

- The number of electrons in an atom is equal to the number of protons in it. Thus, the total positive charge of the nucleus exactly balances the total negative charge in the atom making it electrically neutral. The number of protons in an atom is called its atomic number**(Z).**
- Electrons and the nucleus are held together by electrostatic forces of attraction.

B. Fill up the blanks

1. The decomposition of an electrolyte by passage of electricity is known as **electrolysis.**

2. When cathode rays are focused on thin metal foil, it gets heated up to **incandescence**

3. Cathode rays produce **glow (or) fluorescence** on the walls of the discharge tube.

4. The radiations which were not influenced by a magnet were called **electro magnetic radiations**

5. Neutrons are discovered by **Chadwick**

C. Write in one or two sentence

1. What is the charge of an electron, proton and a neutron ?

The charge of an electron is negative.The charge of a proton is positive. The charge of a neutron is neutral. (no change)

2. What is atomic number?

Atomic number is defined as number of unit positive charges on the number of protons. (i.e.,) equal to number of protons. It is denoted by Z which is also equal to the number of electrons in an atom.

3. What is the maximum number of electrons that an orbital can have?

An atomic orbital can have a maximum number of two electrons.

4. How many orbitals are there in the second orbit? How are they designated?

There are 4 orbitals in the second orbit. Second orbit n = 2.

For n = 2, the possible values of l = 0,1

l = 0, there is one 2s orbital.

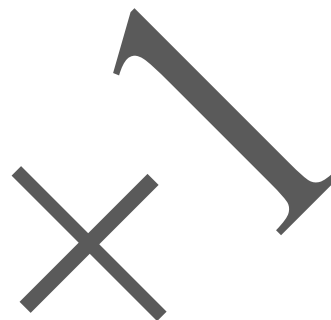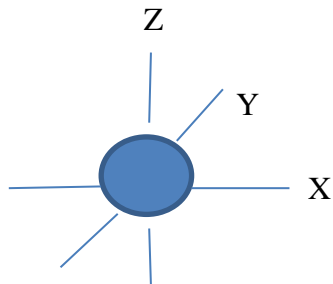l = 1, indicates three p orbitals (m = -1, 0, +1)

Total number of orbitals = 1 + 3 = 4

They are 2s 2p$_x$, 2p$_y$, 2p$_z$.

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5. Sketch the shape of s and p-orbital indicating the angular distribution of electrons.

6. What are the charge and mass of an electron?

The charge of an electron = 1.6022×10^{-19} coulomb

The mass of an electron = 9.10939×10^{-31} kg

7. What is an orbital?

An orbital is a three dimensional boundary of space where there is maximum probability of finding electron.

8. Give the order of filling of electrons in the following orbitals 3p, 3d,4p, 3d and 6s.

According to Aufbau principle, the order of filling of electrons in the following orbitals is 3p, 4s, 3d, 4p, 5s, 4d.

9. What is meant by principal quantum number?

Each shell is characterized by a quantum number called principle quantum number. It specifies the location and energy of an electron. It is represented by the letter ˈnˈ and ˈnˈ can have values 1, 2, 3, 4, …. For the levels K, L, M, N, … etc.

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

Oxygen: $^{18}_{8}O$

Atomic number = 8 = Number of protons = Number of electrons

Mass number   = 18 = Number of protons + Number of neutrons

Number of neutrons = 18 − 8 = 10

The number of protons = 8

The number of electrons = 8

The number of neutrons = 10

11. What are the particles generally present in the nuclei of atoms?

The particles generally present in the nuclei of atoms are protons and neutrons which are collectively known as nucleons.

12. The atomic mass of an element is 24 and its atomic number is 12. Show how the atom of the element is constituted?

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The atomic number of the element is 12. It indicates the number of protons present in the atom of the element is 12 and the number of electrons is 12.

Mass number (or) Atomic mass = 24

Number of protons + Number of neutrons = 24

Number of neutrons = 24 − 12 = 12

So the atom constitutes 12 protons, 12 electrons and 12 neutrons.

13. How will you experimentally distinguish between a ray of neutron and ray of proton?

The ray of neutron and ray of proton are allowed to pass through an electric field. The ray of proton is deflected towards the negative electrodes indicates that the ray contains positive charged particles whereas the ray of neutron does not deflect and it indicates that the ray of neutron contains neutral particles.

14. What is the principal defect of Bohr atom model?

Bohr's atomic model could not explain spectrum of multi electron atoms. (i.e.,) this theory cannot explain the spectral series for the atoms having a large number of electrons. Bohr's atomic model could not explain Zeeman and Stark effect.

15. Write the complete symbol for : (a) The nucleus with atomic number 56 and mass number 138 ; (b) The nucleus with atomic number 26 and mass number 55 ; (c) The nucleus with atomic number 4 and mass number 9.

(a)   Atomic number   56

$^{138}_{56}Ba$

Mass number   138

Barium Atomic number 56, mass number is 138

So (a) is $^{138}_{56}Ba$

(b) $^{55}_{26}X$ The atomic number is 26 and mass number is 55.

The element is Iron.

So (b) is $^{55}_{26}Fe$

(c) $^{9}_{4}X$ The atomic number is 4 and mass number   is 9.

This element is Beryllium .So (c) is $^{9}_{4}Be$

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

An atomic orbital has n = 3, the possible values of l are 0, 1, 2.

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

An atomic orbital has l= 3, the possible values of m are -3, -2, -1 ,0, +1, +2, +3.

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

Chromium $_{24}$Cr. Its electronic configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$.

19. Which energy level does not have p-orbital?

The energy level K (n = 1) does not have p orbital. It contains only 1s orbital.

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

An atom of an element has 19 electrons. The element is potassium $_{19}$K.

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Electronic configuration of K is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$. It contains three 2p orbitals and three 3p orbitals. So total number of p-orbitals are six.

21. How many electrons can have s+ ½ in a d-sub-shell?

In a d-sub shell, there are five orbitals are present. Each orbital can have maximum of 2 electrons in five d orbitals, 5 electrons can have s + ½. and 5 electrons can have s – ½.

22. Write the values of *l* and m for p-orbitals.

When l = 1 sub shell = p

l = 1, m = -1, 0, +1, the orbitals are p_x, p_y and p_z.

23. Which quantum accounts for the orientation of the electron orbital?

Magnetic quantum number (m) accounts for the orientation of the electron orbital.

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

When n = 2 and l = 0 the orbital is 2s.

The shape of 2s orbital is symmertrical sphere. When n = 2 and l = 1 the orbital is 2p. The shape of 2p orbital is dump-bell shape.

25. Give the values for all quantum numbers for 2p electrons in nitrogen (Z = 7).

The atomic number Z = 7 means the element is nitrogen.

$_7N$ Electronic configuration is $1s^2\ 2s^2\ 2p^3$ (or) $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$.

The quantum number for 2p electrons are n = 2, l = 1,m = -1, 0, +1. S = + 1/2 or -1/2.

26. Give the electronic configuration of Mn2+ and Cu. Atomic number of Cu = 29 and Mn = 25.

$$_{25}Mn\ \rightarrow\ Mn^{2+}\ +\ 2e^-$$

Mn – | 25e | Mn^{2+} – | 23e |
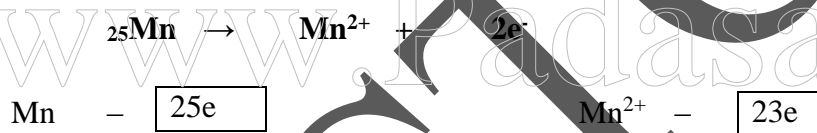
Mn^{2+} Electronic configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^3$

$_{29}Cu$ Electronic configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^{10}$

27. 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 and copper have 5 and 10 electrons in 3d orbitals rather than 4 and 9 electrons respectively as expected. To acquire more stability one of the 4s electron goes into 3d orbital so that 3d orbitals get half-filled or completely filled I chromium and copper respectively. A half filled or completely   filled d level is more stable than a partially filled level.

| $_{24}$ Cr | **Chromium:**

Expected configuration : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^4\ 4s^2$
Actual configuration   : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$

Copper: | $_{29}$Cu |

Expected configuration : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^9\ 4s^2$
Actual configuration   : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$

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D. Explain brieflyon the following

1. Describe Aufbau principle. Explain its significance in the electronic build up of atoms.

2. Using the s, p, d, notation, describe the orbital with the following quantum numbers? (a) n = 1, l=0; (b) n = 2,l = 0; (c) n = 3, l = 1; (d) n = 4, l =3.

3. Using the a Aufbau principle, write the electronic configuration in the ground state of the following atoms : Boron (Z = 5) Neon (Z = 10) and Aluminium (Z = 13).

4. What is Rutherford's α- ray scattering experiment? What are its conclusions?

5. What are the postulates of Bohr theory of atom?

6. Explain the various quantum numbers which completely specify the electron of an atom.

Ln4

Fill in the blanks

1.Mendeleev's periodic law states that the properties of the elements are the periodic functions of the **atomic mass**

2. The Modern periodic law states that the physical and chemical properties of the elements are periodic functions of their **atomic number**

3. The long form of the periodic table is constructed on the basis of repeating electronic **configuration** of the atoms when they are arranged in the order of increasing atomic numbers.

4. The first three periods containing 2, 8 and 8 elements respectively are called **short periods**

5. The valency of representative elements is given by the number of electrons in the outermost orbital and/or equal to **eight** Minus the number of outermost electrons

Write in one or two sentences

1.Arrange F, Cl, Br and I in the order of increasing electronic gain enthalpy.

   The increase order of electronic gain enthalphy is I, Br, F, Cl.

2.Write electronic configurations for the elements of atomic numbers 6 and 14 and from this find out of which group in the periodic table each elements belongs.

Atomic number	Element	Electronic configuration
6	Carbon	$1s^2, 2s^2, 2p^2$
14	Silicon	$1S^2, 2s^2, 2p^6, 3s^2, 3p^2$

3. Which of the following electronic configurations has the lowest ionization enthalpy ? (a) $1s^2, 2s^2, 2p^6$; (b) $1s^2, 2s^2, sp^6$; (c) $1s^2, 2s^2, 2p^6 , 3s^2$.

   The electronic configuration $1s^2, 2s^2, 2p^6$ has the lowest ionization enthalpy.

4. State Modern Periodic Law.

   The modern periodic law states that "the physical and chemical properties of the elements are the periodic function of their atomic number".

5. Why Noble gases have zero electron gain enthalpy?

   In the case of Noble gases , the outer **s** and **p** orbitals are completely filled. No more electrons can be accommodated in these orbitals. Noble gases therefore, show no tendency to accept electrons. Their electron gain enthalpies are zero.

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6. Which of the following pairs of elements would you expect to have lower first ionization enthalpy ? (a) Cl or F ; (b) Cl or S ; (c) K or Ar ; (d)Kr or Xe.

(a) Cl, (b) S, (c) K, (d) Xe are the elements that have lower first ionization enthalpy.

7. Why do elements in the same group have generally similar properties?

Elements in the same vertical column or group have similar electronic configuration, have the same number of electrons in the outer orbitals and therefore similar properties. Group 1 is alkali metals ith ns1 configuration have similar chemical properties.

8. Name any two transition elements and any two inner transition elements.

Two transition elements are Ag, Au (silver, gold)

Two inner transition elements are La, Ac (lanthanum, actinium)

9. Arrange the order of increasing atomic volumes in : (a) Li, Na and K; (b) C, N and O ; (c) Ca, Sr and Ba.

The increasing order of atomic volumes in (b) C, N and O, (c) Ca, Sr and Ba, (a) Li, Na and K.

10. Name the different blocks of elements in periodic table. Give the general electronic configuration of each block.

The different blocks of the elements in the periodic table are, Electronic configuration

(i)	s-block elements	ns^{1-2}
(ii)	p-block elements	$ns^2 np^{1-5}$
(iii)	d-block elements	$ns^2(n-1) d^{1-10}$
(iv)	f-block elements	$ns^2(n-1)d^{1-10} (n-2)f^{1-14}$
(v)	zero group elements	$ns^2 np^6$

11. To which block does the element with configuration $3d^{10}4s^2$ belongs

The element with configuration $3d^{10} 4s^2$ belongs to d-block in the periodic table.

12. Why nitrogen has higher I.E. value than oxygen?

Nitrogen $_7N$ $1s^2$ $_2s^2$ $_2px^1$ $_2py^1$ $_2pz^1$ is half filled stable electronic configuration.

Oxygen $_8O$ $1s^2$ $_2s^2$ $_2p^4$ $_2px^2$ $_2py^1$ $_2pz^1$ It has less stable configuration. From the stable half-filled electronic configuration of nitrogen, the removal of electron becomes difficult. It requires more I.E. so, Nitrogen has higher I.E. value than oxygen. But oxygen by losing the last electron can acquire a stable half filled configuration.

13. Out of fluorine and chlorine, which has greater electron gain enthalpy?

Out of fluorine and chlorine, chlorine has greater electron gain enthalpy. This is because, the fluorine atom has a very compact electronic shell due to its small size. The compactness of fluorine shell results in electron repulsion whenever an electron is introduced into its 2p orbital.

14. Why are d-block elements called transition elements?

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d-block elements form a bridge between the chemically active s-block elements an less reactive metals of groups 13 and 14 and thus take the name as transition elements. Actually the properties of d-block elements are in between characters of s-block and p-block and so they are named as transition elements.

15.What property did Mendeleev use to classify elements in his periodic table?

Mendeleev used the increasing order of atomic weights of the elements to classify them in his periodic table.

16. Among the elements Li, K, Ca, S and Kr which one has the lowest first ionization enthalpy? Which has the highest first ionization enthalpy?

K has the lowest first ionisation energy.

Kr has the highest first ionisation energy.

D. Explain briefly the following

1. Why does the first ionization enthalpy would have higher electron gain enthalpy?

Group 1 element have the electronic configuration as ns1 .By losing one electron they attain the nearest inert gas configuration. So to attain the stable octet configuration , always group, elements are ready to loose one electron from the outer most ns orbital and hence they have higher electron gain enthalpy.

For example,

$_3$Li, $1s^2$ $2s^1$

Li—>Li$^+$ + e$^-$

By losing one electron lithium acquires stable [He] configuration .So it is ready to loose one electron. Hence group 1 elements have higher electron gain enthalpy.

2. Which of the following pairs of elements would have higher electron gain enthalpy? (a) N or O ; (b) F or Cl . Explain.

(a)Among N (or) O, oxygen has higher electron gain enthalpy .Because oxygen due to its higher nuclear charge than nitrogen it attracts electron Gain enthalpy

(b)Among F(or) Cl, Chlorine has higher electron gain enthalpy. Because the incoming electron is more readily accepted by the chlorine atom because of weaker electron repulsion. The electron gain enthalpy of chlorine is therefore higher than that of fluorine.

3. Lanthanides and actinides are placed in separate rows at the bottom of the periodic table. Explain the reason for this arrangement ?

The 4f series (Lanthanides) and 5f series (Actinides) are placed separately the periodic table at the bottom to maintain, its structure, to give importance to the periodicity and to preserve the principle of classification by keeping elements with similar properties in a single column. The two rows of elements at the bottom of the periodic table called lanthanides $_{58}$Ce -$_{71}$Lu and

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Actinoids $_{90}Th$ - $_{103}Lr$ are characterized by similar electronic configuration $(n-2)f^{1-14}$ $(n-1)_{0-10}$ ns^2. The properties of the elements are quite similer and they are placed together . In order to avoid, undue extension of the periodic table, they are placed separately.

4. What do you mean by representative elements? Name the groups of the periodic table, which contain representative elements.

s-block and p-block elements are together called representative elements .Group number 1 and 2 which have the electronic configuration ns^1 and ns^2 are named as s-block elements .Group numbers from 13 and 17 which have the electronic configuration ns^2 np^{1-5} are named as p-block elements . So groups 1, 2, 13, 14, 15, 16 and 17 contain representative elements.

5. Define transition elements. Name the different transition series.

Transition elements are the d-block elements form a bridge between the chemically active metals of s-block elements and less reactive metals of group 13 and 14 and thus named as transition elements . They are placed in between s-block and p-block in the periodic table as 4 transition series between the 3rd and the 12th group.

3d series – first transition series   - Sc to Zn
4d series – second transition series - Y to Cd
5d series – third transition series   - La to Hg
6d series – fourth transition series   - From Ac and is incomplete.

6. Which element of the following pairs have smaller ionization enthalpy? (a) Ca (or) Be ; (b) Ca (or) K ; (c) Cl (or) I ? Justify your answer.

(a) Ca (or) Be. Ca has smaller ionisation enthalpy. In the group   2, as we move down Be,Mg,Ca,Sr,Ba,Ra, the periodic property ionisation enthalpy decreases due to atomic size. So Ca has smaller I.E. then Be.

(b) Ca (or) K. K has smaller   I.E. than Ca. Because in $_{19}K$ $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^1$ electron is easily removed in order to attain the nearest insert gas configuration . So I.E. of K is smaller than Ca and along the period the I.E. increases so Ca has a higher I.E.

(c)   Cl (or) I. I has smaller I.E than Cl. Because we move down the 17th group (halogens), the periodic property I.E. gradually decreases due to increases in atomic size. I.E. of I is less than that of Cl.

7. Why is Na atom bigger than the atoms of both lithium and magnesium ?

Atomic radii is the periodic property. As we move down the group, Atomic radii increases due to the addition of electrons in the new electronic level. As we move across the period, atomic radii decreases due to the addition of electrons in the same electronic level but with move nuclear attraction.

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| Atomic radii | Li | |
| Increases ↓ | Na | Mg |

Atomic radii Decreases →

So comparing with Li, Na has greater atomic size. Comparing with Mg also Na has greater atomic size. Na atom is bigger than both Lithium and Magnesium.

8. What do you mean by the term electron gain enthalpy? How does electron gain enthalpy change along a period and in a group?

Electron gain enthalpy is the amount of energy released when an isolated gaseous atom accepts an electron to form a monovalent gaseous anion.

Electron gain enthalpies generally decreases on moving down the group. This is expected on the account of the increase in size of the atoms, the effective nuclear attraction for electrons decreases. As a result, there is less tendency to attract additional electron with an increase in atomic number down the group.

Electron gain enthalpies generally increase as we move across the period from left to right . This is due to the increase in the nuclear charge, which results in greater attraction for electrons.

9. Explain how the elements are arranged in the form of the periodic table.

In the long form of periodic table elements are arranged in horizontal rows called periods and vertical columns called groups.

Periods. There are seven periods and each period starts with a different principal quantum number.

(i) The first corresponds to first energy level (n=1). This period contains 2 elements H ($1s^1$), He ($1s^2$).

(ii) The second period starts with the electron beginning to the second energy level (n = 2). It can accommodate 8 electrons in 2s , 2p orbitals. It starts with Li ($1s^2\ 2s^1$) and ends with Ne ($1s^2$, $2s^2$, $2p^6$). Total number of elements in this period are 8.

(iii) The third period starts with electrons entering the third energy shell (n=3). It starts with Na ($1s^2,2s^2, 2p^6, 3s^1$) and ends with Ar ($1s^2,2s^2,2p^6,3s^2,3p^6$).It contains 8 elements.

(iv) The fourth period starts with electrons entering the fourth energy shell (n=4). It contains 18 elements starting from potassium K ($1s^2,2s^2,2p^6,\ 3s^2,3p^6\ ,4s^1$) and ends with Kr ($1s^2,2s^2,2p^6,3s^2,3p^6,4s^2,3d^{10},4p^6$). In this period, 3d transition series Sc-Zn are also included between the 4s and 4p block elements.

(v) The fifth period begins with 5s orbital (n=5). It contains of 18 elements including 4d transition series Y-Cd. It starts with Rb ($1s^2,2s^2,2p^6,3s^2,3p^6,4s^2,\ 3d^{10},4p^6,5d^1$) and ends with Xe ($1s^2,2s^2,2p^6,3s^2,3p^6,4s^2,3d^{10},4p^6,5s^2,5p^6$).

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(vi)   The sixth period starts with filling of 6s orbital (n=6). It contains 32 elements including 14 inner transition elements. (Lanthanides) It starts from Caesium (Z=55) and ending with radon (X=86).The different energy levels filling up in these elements is 6s, 4f, 5d, 6p.

(vii)   The seventh period begins with 7s orbital (n=7) and followed by 5f and 6d.It would also have 32 elements, but it is still incomplete with 23 elements in it.

Groups. A vertical column in the periodic table is known as group. A group consists of elements having similar configuration of outer energy shell. There are 18 groups in the long form of periodic table. According to IUPAC, these groups are numbered from 1 to 18. The elements belonging to the same group are said to constitute a family.

10. What are normal, transition and inner- transition elements?

(i)   The normal elements are the s and p block elements comprise those belonging to groups 1d and 2d 13 to 18. They also named as representative elements or main group elements. The outer electronic configuration varies from ns^{1-2} and $ns^2 np^{1-2}$

(ii)   The transition elements are d-block elements comprise those belonging to groups 3 to 12. The outer electronic configuration is $(n-1) d^{1-10} ns^2$.They form a bridge between the chemically active metals of s-block elements and less active metals of groups 13 and 14 and thus they are named as transition elements.

Inner transition elements are the two of elements at the bottom of the periodic table, called Lanthanoids and Actinoids. Their electronic configuration is $(n-2)f^{1-14} (n-1)d^{1-10} ns^2$. The last electron added to each element is f electron and hence they are called f-block elements.

11. What are the differences between normal and transition elements?

Normal elements	Transition elements
They comprise of group 1 and 2 and 13 to 18 with general outer electronic configuration ns^{1-2} and $ns^2 np^1$ to $ns^2 np^6$	They consist of group 3 to12 with general electronic Configuration ns2 (n-1) d1-10. They are called d-block elements because d-orbital is progressively filled.
They consists of soilds, liquid and gaseous elements.	All of them are soild except mercury which is liquid.
They consist of metals, non-metals and metalloid.	They are all metals.
Some of the elements have variable oxidation states.	Most of the elements have variable oxidation states.

12. Explain why radii of positive ions are always smaller than the radii of corresponding neutral atoms and why negative ions have larger radii than the corresponding neutral atom.

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The radii of positive ions or cations are smaller than the corresponding atomic radius because the number of nuclear charges in the positive ion are more than the electrons and the electrons are attracted by the nucleus effiectively.
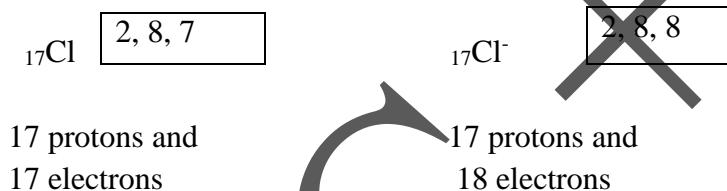
(e.g.,)

$_{11}$Na $_{11}$Na$^+$ 2, 8, 1   2, 8

       11 protons and   11 protons and
       11 electrons   10 electrons
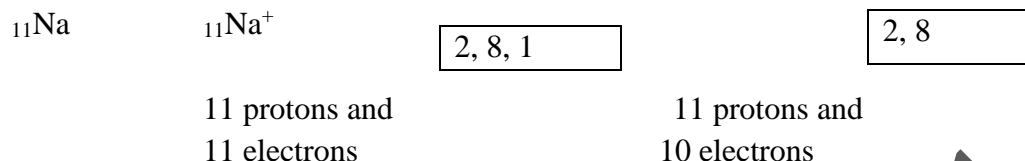
The 10 electrons in Na$^+$ are attracted move by 11 protons or positive charges.

Where as in the negative ion the electrons are move than the nuclear charge and are not effectively attracted so move away from the nucleus.

(e.g.,)

       $_{17}$Cl   2, 8, 7 $_{17}$Cl$^-$ 2, 8, 8

       17 protons and   17 protons and
       17 electrons   18 electrons

13. Explain the size of the group Cl$^-$ >Na$^+$.

Na$^+$ ion is formed by the loss of one electron from Na neutral atom.

             Na $\longrightarrow$ **Na$^+$** + **e$^-$**
$(1s^2, 2s^2, 2p^6, 3s^1)$ $(1s^2, 2s^2, 2p^6)$
       11 electrons   10 electrons
       11 protons   11 protons
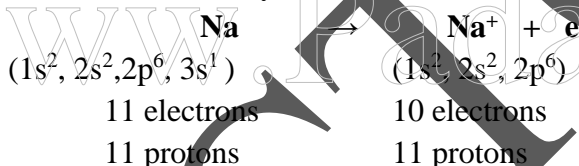
In Na$^+$ 11 protons attract the 10 electrons more effectively than Na atom. So Na$^+$ atomic size is less than Na. In Na$^+$, only K, L shells are filled.

But Cl$^-$ it is formed by the gain of one electron. In Cl$^-$, K, L, M shells are filled. In Cl$^-$,18 electrons are present which are not effectively attracted by 17 protons.

             Cl + **e$^-$** $\longrightarrow$ **Cl$^-$**
$(1s^2, 2s^2, 2p^6, 3s^2, 3p^5)$ $(1s^2, 2s^2, 2p^6, 3s^2, 3p^6)$
       17 protons   17 protons
       17 electrons   18 electrons
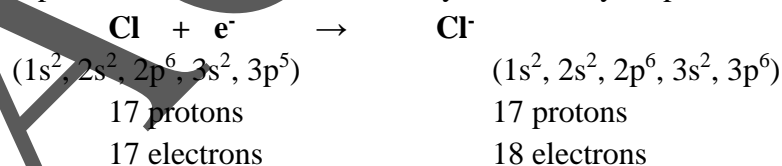
14. What is electron gain enthalpy? On what factors does it depend?

Electron gain enthalpy is the amount of energy released when an isolated gaseous atom accepts an electron to form a monovalent gaseous anion.

       Cl$_{(g)}$ + e$^-$ $\longrightarrow$ Cl$^-_{(g)}$ EA (349 kj mol^{-1})
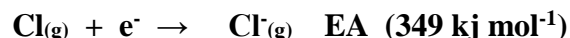
The magnitude of electron affinity is influenced by a number of factors such as (i) Atomic size, (ii) Effective nuclear charge, (iii) Screening effect by inner electrons.

(i)   When the atomic size increases, the value of electron affinity decreases.

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(ii)   When the nuclear charge increases, the value of electron affinity also increases.

(iii)   When the screening effect by inner electrons increases, the value of electron affinity decreases.

15. Give the general variation of electron gain enthalpies in the periodic table.

16. Define the term ionic radius. Justify that the radius of anion is larger than the parent atom.

17. What do you mean by ionization enthalpy? How does it vary across a period and down a group?

18. What is meant by electronegativity? On what factors does it depend?

19. What are the essential features of the periodic table of Mendeleev? Discuss how his table has been modified subsequently.

Ln 5

B. Fill in the blanks

1. The first element in the periodic table is .**Hydrogen**

2. **Protium** is the common formd of hydrogen.

3. The half-life of tritium is . **12.3 years**

4. Deuterium reacts with ammonia to form. **Deuteroammonia**

5. The rare isotope of hydrogen is. **Tritium $1H^3$**

6. **Heavy water (or) D_2O** is employed in nuclear reactor to slow down the speed of fast moving neutrons.

7. The magnetic moment of para hydrogen is . **Zero**

8. Deuterium with salt and other compounds forms . **deutrates**

9. Hydrogen peroxide was first prepared by **L.J. Thenard** in . **1813**

10. Pure H_2O_2 is **. unstable**

11. The Arabic word `Alquili' means . **plant ash**

12. The electronic configuration of potassium is **. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$**

13. All alkali metals have **low** melting and boiling points.

14. On moving down the group of alkali metals, ionization energy . **decreases**

15. **Lithium** is the lightest of all solid elements.

C. Write in one or two sentences

1.What are isotopes? Mention the isotopes of hydrogen.

        Atoms of the same element having same atomic number but different mass number are called
isotopes.

Isotopes of hydrogen- Protium, Deuterium, Tritium.

2.Write a short note on tritium.

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It occurs in the upper atmosphere only where it is continuously formed by nuclear reactions induced by cosmic rays. Unlike deuterium, it is radioactive, with a half-life of ~ 12.3 years. It's nucleus consists of one proton and two neutrons.

3. How does deuterium react with nitrogen?

Hydrogen combines with nitrogen in the presence of a catalyst to form nitrogen deuteride which are also known as heavy ammonia or deutero ammonia.

$$3D_2 + N_2 \rightarrow 2ND_3$$
Deutero ammonia

4. How does deuterium react with metals?

Deuterium reacts with alkali metals at high temperatures (633 K) to form deuterides.

$$2Na + D_2 \rightarrow 2NaD$$
Sodium deuteride

5. Mention the uses of deuterium.

(i)   It is used as tracers in the study of mechanism of chemical reaction.

(ii)   High speed deuterons are used in artificial radioactivity.

(iii) Deuterium oxide known as heavy water is employed as moderator in nuclear reactor to slow down the speed of fast moving neutrons.

6. How is tritium prepared?

(i)   By bombarding lithium with slow neutrons.
$$_3Li^6 + _0n^1 \rightarrow _1T^3 + _2He^4$$
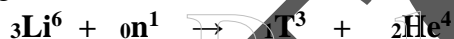
(ii)   By bombarding Beryllium with deuterons.
$$_4Be^9 + _1D^2 \rightarrow _1T^3 + _2He^4$$
$$_4Be^9 + _1D^2 \rightarrow _1T^3 + 2_2He^4$$
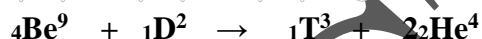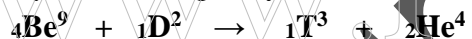
7. How do you convert para hydrogen to ortho hydrogen?

(i)   By treatment with catalyst like Pt or Fe.

(ii)   By passing an electric discharge.

(iii) By heating to 800˚c or more.

(iv)   By mixing with paramagnetic molecules like O_2, NO, NO_2.

(v)   By mixing with nascent hydrogen atomic hydrogen.

8. How does heavy water react with metals?

Heavy water reacts slowly with alkaline earth metals liberating heavy hydrogen.

$$_2Na + 2D_2O \rightarrow 2NaOD + D_2$$
Sodium deuteroxide
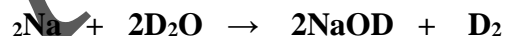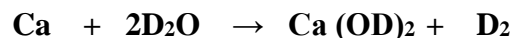
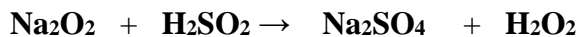$$Ca + 2D_2O \rightarrow Ca(OD)_2 + D_2$$
Calcium deuteroxide

9. How is hydrogen peroxide prepared in the laboratory?

Hydrogen peroxide is prepared in the laboratory by the action of dilute sulphuric acid on sodium peroxide. Calculated quantity of **Na_2O_2** is added in small proportions to a **20%** ice cold solution of sulphuric acid.

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Na_2O_2 + $H_2SO_2 \rightarrow$ Na_2SO_4 + H_2O_2

**30%** solution of H_2O_2 is obtained by this process.

10. How is hydrogen peroxide solution concentrated?

Hydrogen peroxide solution contains the impurities like organic material or metallic ions which may catalyse its explosive decomposition. The solution of H_2O_2 is concentrated

(i) by careful evaporation of the solution obtained on a water bath preferably under reduced pressure using fractionating column.

(ii) by distillation under reduced pressure at temperature below 330 K, the concentration upto **90%** solution is used.

11. Write about the reducing property of hydrogen peroxide.

H_2O_2 acts as a reducing agent with powerful oxidising agents. Moist silver oxide, acidified **KMnO₄,** ozone, chlorine and alkaline solutions of ferricyanides are reduced.

Ag_2O + $H_2O_2 \rightarrow$ $2Ag$ + H_2O + O_2
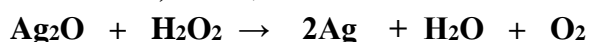
Silver oxide   Silver

12. Mention two important uses of H2O2.

(i) H_2O_{22} destroys bacteria and hence it is used as an antiseptic and germicide for washing wounds, teeth and ears.

(ii) It is also used as propellant in rockets.

13. Why alkali metals have low melting and boiling points?

All the alkali metals have low melting and boiling point due to the weak bonding in the in the crystal lattice. The weak inter ionic bonds are attributed to their larger atomic radii and to the presence of just one valence electron.

14. Why alkali metals have strong electropositive character?

All the alkali metals have just one electron in addition to the stable octet configuration which they can readily loose.

M $\rightarrow$ M^+ + e^-

So alkali metals   have low ionisation energies and have a greater tendency to lose electrons forming unipositive ion. Therefore they have strong electropositive character. The alkali metals are so highly electropositive that they emit electrons when irradiated with light.

Ln6

B. Fill in the Blanks

1. The general electronic configuration of alkaline earth metals is

$\underline{ns^2}$.

2. The ionic radius **increases** on moving down the group 2.

3. In flame, calcium gives **brick red** colour.

4. Beryllium resembles more with an element in 13th group **aluminium**.

5. Magnesium comes from the name of the mineral **magnesite**

6**. Mg^{2+}** ion is present in chlorophyll

7. Magnesium is prepared by the electrolysis of fused   **Magesia (or) Magnesium chloride**.

8. With air, Magnesium forms

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Magnesium oxide and **magnesium nitride**.

9. The formula of epsom salt is **$MgSO_4.7H_2O$**

10. Epsom salt is used as **purgative**.

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C. Match the following

1. Magnetite   **$MgCO_3$**
2. Dolamite   **$MgCO_3 . CaCO_3$**
3. Epsom salt   **$MgSO_4 . 7H_2O$**
4. Carnallite   **$Mg Cl_2 . KCl . 6H_2O$**
5. Gypsum   **$CaSO_4 . 2H_2O$**

D. Write in one or two sentence

1. Why the oxides of Group 2 metals have high melting points?

The oxide of group 2 metals have high melting point because the number of bonding electrons are twice as **(i.e.,)** group one elements higher. Their atoms hold the valence electrons more strongly and due to this more attraction they have high melting point

2. Why there is increase in the ionisation potential for forming M^{3+} ion for group 2 metals?

The electrons of group 2 have two s-electrons in the outermost orbital. By losing 2 electron, they attain the nearest inert gas configuration with the absorbtion of low ionisation of low ionisation potential. After the removal of 2 electrons, it is very difficult to remove one more electron to get M^{3+} ion as it will disturb the outer eight electron configuration which is very stable. So there is an increase in the ionisation potential for forming M^{3+} ion.

3. Why the ionization potential of M^{2+} is not very much greater than M^+?

In alkaline earth metals, the ionisation potential of M^{2+} is not very much greater than M^+. This is due to high lattice energies in the crystalline compounds and high hydrogen energies of M^{2+} ions in solution. These energies more than counter balance the higher values of the second ionisation potentials with the results that M^{2+} ions are formed in preference to the M^+ ions .

4. Why a precipitate of $Mg(OH)_2$ is not formed when aqueous ammonia,
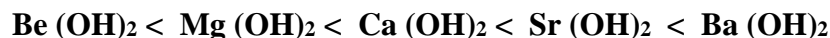**NH_4OH is added to a solution of $MgCl_2$?**

When $MgCl_2$ solution is treated with aqueous ammonia (NH_4OH), magnesium hydroxide is a product.

The ammonium ions present in the solution in equilibrium with undissolved $Mg (OH)_2$ producing the highly soluble but very slightly ionized ammonium hydroxide.

$$MgCl_2 + 2NH_4OH \leftrightarrow 2NH_4Cl + Mg (OH)_2$$
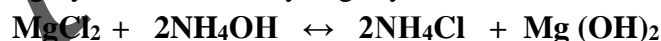
5. List the carbonates and hydroxide of alkaline earth metals in order of
their increasing stability and their solution.

The stability of the carbonates and hydroxides of alkaline earth metals icreases from Be to Ba.

$$BeCO_3 < MgCO_3 < CaCO_3 < SrCO_3 < BaCO_3$$
$$Be (OH)_2 < Mg (OH)_2 < Ca (OH)_2 < Sr (OH)_2 < Ba (OH)_2$$

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This is because, the ease with which the atoms of alkaline earth metals lose electrons increases with the rise in atomic number.

6. Why do beryllium halides fume in air?

Beryllium halide fumes in air due to its hydrolysis because they are covalent and form HCL on hydrolysis.

$$BeCl_2 + 2H_2O \rightarrow Be(OH)_2 + 2HCL$$

7. Why group 2 elements are harder than alkali metals?

Alkaline earth metals (or) group **2** elements are harder than alkali metals because of smaller atomic size, they have strong metallic bond. The atomic radius becomes smaller due to higher nuclear charge of these atoms which tends to draw the orbital electrons inwards. Because of the smaller atomic radius,these elements are harder than alkali metals.

8. Beryllium halides are covalent whereas magnesium halides are ionic. Why?

Beryllium halides are covalent due to smaller atomic size, higher ionisation energy of **Be** whereas magnesium halides are ionic due to the larger atomic size and lesser ionisation energy of **Mg.**

9. Why are monoxides of alkaline earth metals are very stable?

The monoxides of alkaline earth metals are very stable due to the strong force of attraction between M^{2+} and O^{2-} ions. Because of strong forces of attraction , the monoxides have high lattice energy and they are more stable.

10. The basic strength of the oxides of group 2 elements increases from Be to Ba. Why?

As we move from **Be** and **Ba**, the ease with which the atoms of alkaline earth metals lose electrons increases with the increase in atomic number. This results in the increased chemical reactivity of the elements and increased ionic character in the compounds (oxide of group 2 elements). So there is an increase in the basic strength of the oxides of alkaline earth metals from Be to Ba due to increase in the electropositive character.

D. Explain briefly on the following

1. What are alkaline earth metals? Why are they called so?

2. In what respects Be and Mg differ from all the other metals of group 2.

3. How can you explain the anomalous behaviour of beryllium.

4. How does magnesium occur in nature? How is the metal extracted from its Ore?

5. In the light of metallic bonding account for the following properties of group 2 elements.

a. These are harder than alkali metals

b. These are good conductors of heat and electricity.

6. Why the first ionization energy of alkaline earth metals higher than that of Ist group.

7. Mention the uses of plaster of Paris.

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8. How is plaster of paris prepared?

9. How is MgSO4 prepared?

10. Mention the uses of Magnesium?

Ln7

B. Fill in the blanks

1. The general electronic configuration of Boron group elements is $\underline{ns^2\ np^1}$.

2. Boron combines with nitrogen to form **BN (or) Boron nitride**.

3. **Borox**is used to identify the metallic radicals in the qualitative analysis.

4.**Borozine (or) Borozole $B_3N_3H_6$** is known as `inorganic benzene'.

5. In diamond, every carbon atom is bonded with the other by **covalent** bond.

6. C60 Buckminister fullerene was nicknamed as **bucky ball** .

7. Carbon tetrachloridoes **not undergo** hydrolysis.

8. Nitrogen was discovered by **Daniel Rutherfort** .

9. Nitric acid means **aqua tortis** .

10. Oxidising power of nitric acid **decreases** with dilution.

11. Dioxygen is also called as **molecular oxygen**.

12. Atomic oxygen combines with molecular oxygen to give **ozone**.

13. The ozoniser commonly used in the preparation of ozone are **Sieman's ozoniser** and **Brodies ozoniser**.

14. Ozone can liberate a **nascent** oxygen easily .

15. **Ozone** is used in the manufacture of synthetic camphor.

C. Match the following

a.

1. Borax	**$Na_2B_4O_7$**
2. Graphite	**Allotrope of carbon**
3. ZnO	Neutral oxide
4. CFCs	Ozone
5. NH3	Fertilizer
	Fixation of nitrogen

b.

1. Inert pair effect	Stabilisation of lower oxidation state
2. Oxyacid	Nitric acid
3. Liquid nitrogen	Refrigerant
4. Ostwald process	Platinum gauze
5. Molecular oxygen	Cell fuel

c. Borax bead test

1. Copper	Red
2. Iron	Bottle green

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3. Manganese	Colorless
4. Cobalt	Blue
5. Chromium	Green

D. Write in one or two sentence

1. Mention the reasons for the stabilisation of lower oxidation state of p-block element.

(i)   The strength of M-X bond decreases.

(ii)   The lattice energies of the compounds containing M^{4+} ion decreases.

(iii)   The electrons in the ns sub shell do not prefer to form bonds.

2. Show the electron accepting property of boron trifluoride by giving an example.

Boron trifluoride behaves as Lewis acid (or) acceptor of electrons. Because in BF_3, the valence shell of boron is short of the octet by two electrons, gets its octet completed by accepting alone pair of electrons from the nitrogen or oxygen atom of a Lewis base.

3.Give an example of monovalent and trivalent element in group III

Monovalent element in group III- Thallium Tl^{+1}

Trivalent element in group III-Aluminium Al^{3+}.

4.Why diamond is hard compared with graphite?

In diamond, every carbon atom is bonded with other in a tetrahedral manner by strong covalent links resulting in the formation of giant molecule. The C-C bonds are very strong. The combined strength of many carbon-carbon bonds within the structure of diamond give the great hardness to diamond. But in graphite, the carbon atoms are arranged in regular hexagons in flat parallel layers. There is no strong bonding between the layers. Because of the weak Vander waals forces between the layers, graphite is soft.

5. Why Boron family has a tendency to form hydrides?

Boron family has a tendency to form hydrides because they have small atomic size and high ionisation energy. Boron forms covalent compounds due to highest ionisation energy.The hydrides of non-metals are more stable.So boron has a tendency to form hydrides by the formation of covalent bond with hydrogen.
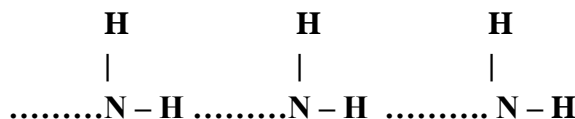
6. Boron does not form B3+ ion. Why?

$_5B\ 1s^2\ 2p^2\ 2p^1$

Boron has 3valence electrons .Due to the small atomic size and the highest ionisation potential ,the removal of 3e is impossible .So B^{3+} ion is not formed.

7. Why NH₃ has high boiling point than PH₃?

Ammonia has higher boiling point due to the presence of intermolecular hydrogen bonding because of its high electronegativity whereas in phosphine ,there is no intermolecular hydrogen bonding and it has low boiling point.

```
        H           H           H
        |           |           |
………N – H ………N – H  ………. N – H
```

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    | | |
   H   H   H

8. NH_3 is soluble in water whereas other hydrides of group 15 elements are insoluble in water. Why?

     Ammonia is soluble in water because it is a strong base and dissolves to give NH_4OH. Ammonia contains intermolecular hydrogen bonding and it is more soluble in water. But the other hydrides of group 15 elements are less basic and insoluble in water.

9. Which is considered to be "earth's protective umbrella"?

     The ozone layer in the upper atmosphere is considered to be 'earth's protective umberlla'.

     The ozone in the upper atmosphere is important in shielding us from the intense ultraviolet radiation coming from the sun. Ozone shield is a shell about 30 km altitude which contains enough ozone to absorb short wavelength UV radiation. Hence, ozone is considered to be earth's protective umberlla'.

10. Mention any 3 uses of ozone.

        (i)   As germicide and disinfectant

        (ii)   For bleaching oils, ivory, flour, starch, etc.

 (iv)   In the manufacture of artificial silk and synthetic camphor.

.

11. What are CFC's? Mention its environmental action.

     CFC's are chlorofluorocarbons.

     CFC's react with O_3 and cause hole in the ozone layer.CFC's are long lived molecules and diffuse into the stratosphere where are decomposed by UV radiation to produce chlorine. The chlorine atom react with ozone, this cause a decrease in the concentration of ozone at a faster rate than its formation from O_2. CFC's are one of the causes for environmental pollution.

12. What are compound oxides? Give an example.

     Compound oxide are oxides that behave as if they contain two different oxides.

(e.g.,) Ferrous ferric oxide are Fe_3O_4. This is considered to be the mixture of FeO and Fe_2O_3

13. Mention the metal ions present in haemoglobin and myoglobin and state its function.

     Haemoglobin contains Fe^{++}

     Haemoglobin is an iron containing co-ordinates compound in red blood cells, responsible for the transport of oxygen from the lungs to varies parts of the body. It contains of heme, a complex of Fe (II) bonded to a porphyrin ligand and globin protein.

     Myoglobin is a substance in muscle tissue acting as a reservoir for the storage of oxygen and as a transport of oxygen within muscle cells. It consists of heme, a complex of Fe (II) and has a single polypeptide chain.

14. What happens when ozone reacts with

a) lead sulphide b) potassium manganate

E. Explain briefly on the following

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1. Explain inert pair effect with suitable example.

2. Give an account of nature of hydrides of 15th group elements.

3. How is boron extracted from borax?

4. What happens when boron reacts with

a) conc.H2SO4 b) conc.HNO3 c) SiO2

5. How is borax prepared from colemanite?

6. How borax bead test is helpful in identifying basic radicals in

qualitative analysis?

7. Discuss the structural difference between diamond and graphite.

8. Write a short note on fixation of nitrogen.

9. How nitric acid is prepared by ostwald process.

10. Why silicon carbide is used as an abrasive?

11. How molecular oxygen is important for all oxygenated animals?

12. How ozone reacts with the following (a) PbS (b) KmnO4

Ln8

B. Fill in the Blanks :

1. In NaCl ionic crystal each Na+ ion is surrounded by **6** Cl⁻ ions

and each Cl⁻ ion is surrounded by **6** Na⁺ ions.

2. The coordination number of Cs+ in CsCl crystal is **8**

3. **Amorphous** solids do not possess sharp melting points and can be considered as **super cooled** liquids.

4. A body centred unit cell has an atom at the each vertex and at **centre** of the unit cell.

5. The three types of cubic unit cells are **simple cubic, fcc** and **bcc**

6. A crystal may have a number of planes or axes of symmetry but it possesses only one **centre** of symmetry.

7. Amorphous solids that exhibit same physical properties in all the directions are called **isotropic**.

8. Crystalline solids that exhibit different physical properties in all directions are called **anisotropic**.

9. The number of atoms in a single unit cell of cubic close packedsphere is **4**

10. In a bcc, an atom of the body centre is shared by **1** unit cell.

11. The Weiss indices of a plane are 1/2, 1/2, 1/2. Its miller indices will Be **2, 2, 2** and the plane is designated as **(222) plane**.

12. A plane is parallel to x & z axes and makes unit intercepts along y-axis. Its Weiss indices are ∞**, 1,** ∞**.** Its Miller indices are **0, 1, 0**. The

plane is designated as **(010) plane**.

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SNo	Crystalline solid	Amorphous solid

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1	The substance in which the atoms or molecules are arranged in a very **regular** and orderly fashion in **three dimensional** pattern.	The substance in which there is **no regular** arrangement of atoms.
2	The ultimate particles are arranged in a definite pattern throughout the entire three dimensional network of a crystal.	The molecules do not have definite geometric pressure.
3	They have **sharp melting point.**	They **do not have sharp melting point.**
4	They have **well –defined faces** and shapes.	They **do not have well-defined faces** and shapes.
5	They are **anisotropic**. i.e. they exhibit different physical properties in all the directions.	They are **isotropic**. i.e. they exhibit same physical properties in all the directions.
6.	They have characteristic **heats of fusion.**	They **do not have** characteristic **heats of fusion.**
7.	They are generally in compressible.	They may be compressed to an extent.
8.	They give a **regular cut** when cut with a sharp edged knife.	They give **irregular cut**.
9.	They are regarded as **true solids**.	They are regarded as **super cooled liquids** or pseudo solids.
10	Eg. **Sugar, sodium chloride, sulphur**	Eg: **glass, plastic, rubber**

Difference between crystalline and amorphous solid.

C. Write in one or two sentence :
1. What governs the packing of particles in crystals?

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The smallest structure of which the crystalline soild is built by its repetition in three dimensions called unit cell.

2.What is meant by 'unit cell' in crystallography?

A unit cell is the fundamental elementary pattern of a crystalline soild. The charaterisation of the crystal involves the identification of its unit cell. The smallest structure of which the crystalline soild is built by its repetition in three dimensions is called unit cell.

3.How many types of cubic unit cell exits?

There are three types of cubic units cells. They are (i) simple cubic unit cell, (ii) body centred cubic cell (bcc), (iii) face centred cubic cell (fcc).

4.What are Miller Indices?

When the reciprocal of Weiss indices are multiplied throughout by the smallest number in order to make all reciprocals as integers, we obtain the Miller indices of a plane. The Miller indices for a particular family of planes are usually written (h, k, l) where h, k and l are positive integers or zero.

5. Mention the number of sodium and chloride ions in each unit cell of NaCl

The unit cell of sodium chloride consists of 14 chloride ions and 13 sodium ions. Each chloride ion is surrounded by 6 sodium ions and similarly each sodium ion is surrounded by 6 chloride ions.

The particles at the corners, edges and faces do not lie wholly within the unit cell but shared by other unit cells. A particle ions at a cornor is shared by eight unit cells, one at the centre of the face is shared by two and one at the edge is shared by four. The unit cell of sodium chloride has 4 sodium ions and 4 chloride ions.

No. of sodium ions = 12 (at edge centres $\times$ ¼) + 1(body centre) $\times$ 1

$\qquad$ = (12 $\times$ ¼) + 1 = 3 + 1 = **4**

No. of chloride ions = 8 (corners) $\times$ ⅛ + 6 (face centre) $\times$ ½

$\qquad$ = 1 + 3 = **4**

6. Mention the number of cesium and chloride ions in each unit cell of CsCl

The unit cell of CsCl has one Cs^+ ion and one Cl^- ion.

In CsCl, each Cs^+ ion is connected to 8 Cl^- ions and each Cl^- ion to 8 Cs^+ ions. Thus , each atom is at the centre of a cube of atoms of opposite end.

No. of caesium ions = 1(body centre) $\times$ 1 = 1

No. of chloride ions = 8 (corners $\times$ ⅛) = 8$\times$ ⅛ = 1

D. Explain briefly on the following :

1. Define and explain the following terms

a) Crystalline solids b) Amorphous solids c) Unit cell

2. Give the distinguishing features of crystalline solids and amorphous solids.

3. Explain the terms Isotropy and Anisotropy.

4. What is the difference between body centred cubic and face centred cubic?

5. Draw a neat diagram for sodium chloride structure and describe it accordingly.

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6. Draw a neat diagram for Cesium chloride structure and describe it accordingly.

Ln9

B. Fill in the blanks

1. The correction term for pressure deviation is $\underline{\mathbf{P + n^2a /v^2}}$ in the Vanderwaal equation of state.

2. The relation between inversion temperature and Vanderwaal's constants `a' and `b' is $\underline{\mathbf{T_i = 2a/Rb}}$

3. To liquefy Helium **adiabatic demagnetization** method is exclusively used.

4. The adiabatic expansion of a real gas results in **cooling**

5. The rate of diffusion of gas is **inversely proportional** to square root of both and molecular mass.

C. Match the following

A	**B**
1. Ideal gas behavior	Low pressure and high temperature
2. Adiabatic demagnetization	Liquid Helium
3. CO_2 at 31.1°C	Critical temperature
4. Joule Thomson Experiment	Liquid oxygen
5. Ratio of the partial pressure to the total pressur e	Mole fraction of the gas

D. Write in one or two sentence

1. Write the mathematical expression for Boyle's law.

$P \propto 1/V$ at constant temperature.

$PV = $ constant at constant temperature.

$P_1 V_1 = P_2 V_2 = $ constant.

2. Compare the partial pressures of gases A and B when 3 moles of A and moles of B mixed in constant volume, and 25oC and 1 atm pressure.

Solution:

Partial pressure of A = Total   pressure $\times$ Mole fraction

Mole fraction of   A= 3/3+5 =3/8

Total pressure = 1 atmosphere

Total number of moles = 8

Partial pressure of   A = $1 \times 3/8 \times P$

Number of moles of   B = 5

Mole fraction of B = 5/8

Partial pressure of   B = $1 \times 5/8 \times P$

Partial pressure of A : B = $3/8 \times P : 5/8 \times P$

3.Give the correction factors for the volume and pressure deviation for a Vanderwaal's gas.

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Volume correction. The corrected volume of the real gas = V − b where b is excluded volume for 1 molecule, $b = 4V_m$

V − b = Volume correction.

Pressure correction. The corrected pressure of 1 mole of the real gas = **P + a / V²**

Where a / V^2 is the cohesion pressure where a is the Vanderwaal's costant.

P + a / V² = pressure correction.

4.A sample of an ideal gas escapes into an evacuated container, there is no change in the kinetic energy of the gas. Why?

An ideal gas has no forces of attraction between the gas molecules and therefore when they expand there is no change in the kinetic energy as they behave as independent molecules.

5. What is the change in temperature when a compressed real gas is allowed to expand adiabatically through a porous plug.

When a compressed real gas is allowed to expand adiabatically a porous plug, the temperature of the gas decreases. When the gas is allowed to escape into a region of low pressure, the molecules move apart rapidly against the intermolecular attractives forces. In this case, work is done by the gas molecules at the expense of internal energy of the gas and cooling occurs. This reduction in temperature is referred as Joule Thomson effect.

6. Define Boyle's law and Charle's law.

Boyle's law. For a given mass of a gas at constant temperature, the pressure (P) is inversely proportional to its volume (V).

P α 1 / V at constant temperature.

PV = constant.

$\boxed{P_1V_1 = P_2V_2}$

Charle's law. For a given mass of a gas at constant pressure, its volume (V) varies directly as its absolute temperature (T).

$\boxed{V \; \alpha \; T}$ (or) V / T = constant.

For a given mass of a gas, at constant volume, the pressure varies directly as its absolute temperature.

$\boxed{P \; \alpha \; T}$ (or)   P   /   T   =   constant.

7. What are measurable properties of gases?

The measurable properties of gases are volume (V), pressure (P), temperature (T) and number of moles (n) of the gas in the container.

8. What is the molar volume of nitrogen at 500K and 600 atm according to ideal gas law?

For an ideal gas, the equation of state P_1 = 1atm. ; T_1 = 273 K

$P_1V_1 / T_1 = P_2V_2 / T_2$ V_1 = 22.4 lit.

$V_2 = P_1V_1 / T_1 \times T_2/P_2$ P_2 = 600 atm. ; T_2 = 500 K

$V_2 = 1 \times 22.4 / 273 \times 600 \times ?/$ 500 V_2 = ?

$V_2 = 1 \times 22.4 \times 500 / 600 \times 273$

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$V_2 = 22.4 \times 5/6 \times 273$

Molar volume of nitrogen = **0.06837 litres.**

9. Define Graham's law of diffusion.

Graham's law of diffusion states that, "under the same conditions of temperature and pressure, the rates of diffusion of different gases are inversely proportional to the square roots of their molecular masses ". Mathematically, the law can be expressed as,

$r_1/r_2 = \sqrt{M_2/M_1}$ where r_1 and r_2 are the rates of diffusion of gases 1 and 2 while M_1 and M_2 are their molecular masses respectively.

10. Give the values of R-gas constant in calories and Joules.

The value of R-gas constant in calories R = 1.987 cals k^{-1} mol^{-1}

The values of R-gas constant in Joules R = 8.314 joule k^{-1} mol^{-1}

11. What are the units of Vanderwaal's constants `a' and `b' ?

Unit of vanderwaal's constants

ʾaʾ = $litre^2$ atm mol^{-2} (or) dm^6 atm mol^{-2}

'bʾ = litre mol^{-1} (or) dm^3 mol^{-1}

12. Write the significance of Vanderwaal's constants.

Significance of Vanderwaal's constant `a' and `b',

(i)   The term a $/V^2$ is the measure of the attractive forces of the molecules. It is also called the cohesion pressure (or) internal pressure .

(ii)   The inversion temperature of a gas can be expressed in terms of `a' and `b'.
$T_i = 2a/Rb$

(iii)   The vanderwaal's constants `a' and `b' are used to calculate the critical constants of a gas.

13. Write the limitations of vanderwaal equation of state.

(i)   It could not explain the quantitative aspect of deviation satisfactorily as it could explain the qualitative aspects of P and V deviations .

(ii)   The values of a and b are also found to vary with P and T such variations are not considered are not considered in the derivation of Vanderwaal's equation.

(iii)   Critical constants calculated from Vanderwaal's equation deviate from the original values determined by other experiments.

14. Define Joule-Thomson effect.

The phenomenon of producing lowering of temperature when a gas is made to expand adiabatically from a region of high pressure into a region of low pressure is known as Joule – Thomson effect.

15. What is meant by inversion temperature ?

The characteristic temperature below which a gas expands adiabatically into a region of low pressure through a porous plug with a fall in temperature is called inversion temperature (T_i)

$T_i = 2a/Rb$ where **R** = gas constant a and b = Vanderwaal's constant.

E. Explain briefly on the following

1. At 27^oC, H_2 is leaked through a tiny hole into a vessel for 20 minutes. Another unknown gas at the same T and P as that of H_2 is leaked through the same hole for 20 minutes. After effusion of

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the gas, the mixture exerts a pressure of 6 atm. The H_2 content of the mixture is 0.7 moles. If volume of the container is 3 litres what is the molecular weight of unknown gas ?

2. Calculate the pressure exerted by 5 moles of CO_2 in one litre vessel at 47^oC using Vanderwaal's equation. Also report the pressure of gas if it behaves ideally in nature. Given that a=3.592 atm lit2 mol^{-2}. b = 0.0427 lit mol^{-1}Ans.: P real = 77.2 atm P ideal = 131.36 atm

3. Calculate the total pressure in a 10 L cylinder which contains 0.4 g of helium, 1.6 g of oxygen and 1.4 g of nitrogen at 27^oC. Also calculate the partial pressures of He gas in the cylinder. Assume Ideal behaviour for gases. R = 0.082 L atm k-1 mol^{-1} Ans. Ptotal = 0.4926 atm, pHe = 0.2463 atm. Po2 = 0.1231 atm, pN2 = 0.123 atm

4. The critical constants for water are 374^oC, 218 atm and 0.0566 litre mol^{-1}. Calculate `a' and `b' of water. Ans. a = 2.095 lit2 atm mol^{-1} b = 0.0189 lit mol^{-1}

5. Vanderwaal's constant in litre atmosphere per mole for carbon dioxide are a = 3.6 and b = 4.28 x 10^{-2}. Calculate the critical temperature and critical volume of the gas. R = 0.0820 lit atm K^{-1}. mol^{-1}

6. Explain the causes for deviation for real gases from ideal behaviour.

7. Deduce the relationship between critical constants and Vanderwaal's constants.

8. Describe Linde's process of liquefaction of gases with neat diagram.

9. Describe Claude's process of liquefaction of gases with neat diagram.

10. What is meant by adiabatic demagnetisation? Explain its use in liquefaction of gases.

Ln10

B. Fill in the blanks

1. In NaCl, Na+ ion has **Neon** and Cl$^-$ ion has **Argon** electron configurations.

2. Linear overlap of two atomic p-orbitals leads to **covalent bond formation** .

3. Born-Haber cycle is related with **Lattice enthalpy determination**.

4. Two atoms of similar electronegativity are expected to form **covalent** compounds.

5. Repulsion between bond pair-bond pair is **less** than in between lonepair- lonepair.

C. Match the following

1. Electrovalent bonding   Electron transfer
2. Covalent bonding   Electron sharing
3. Valence Bond theory   Heitler and London
 4. Polarised Bond   Fajan's theory
5. Resonance   Benzene

D. Write in one or two sentence

1. Arrange NaCl, MgCl₂ and AlCl₃ in the increasing order of covalent character.

    The increasing order of covalent character is $AlCl_3 > MgCl_2 > NaCl$

2.Find σ and π bonds in the following : CH₃-CH₃, CH₂=CH₂, CH≡CH

    $CH_3 – CH_3$ - seven sigma (σ) bonds.

    $CH_2 = CH_2$ - 5 σ bonds 1π bond.

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$CH \equiv CH$ - 3 σ bonds and 2π bonds.

3.Among Na+, Ca+2, Mg+2, Al+3 which has high polarising power ?

Al^{3+} has high polarizing power.

4.What is the structure of BeCl2 ?

$BeCl_2$ structure is linear.

Cl – Be – Cl

5. Write the differences between electrovalent and covalent bonds.

No	Electrovalent bonds	Covalent bonds
1.	It is formed by the transfer of electrons from a metal to a non-metal atom.	It is formed by sharing of electrons between non – metal atoms.
2.	It consists of electrostatic force between cations and anions .	It consists of shared pair of electrons between atoms.
3.	They are non-rigid and non-directional.	They are rigid and directional.
4.	They are strong bonds	They are weak bonds

6. Give reason : CCl4 is insoluble in H2O while NaCl is soluble.

CCl4 is a non-polar covalent compound and it is insoluble in water. This is because of the fact that these compounds due to their big size, are not able to interact with water and so they are insoluble in water . But NaCl is soluble in water. NaCl is a polar electrovalent compound. In water, due to solvation of ions by the solvent molecules, the strong interionic attractions are weakened and exist as separated ions.

7. sp³ hybridisation is involved in CH₄, H₂O and NH₃. Why are the bond angles different in three cases?

In CH_4, H_2O and NH_3, sp^3 hybridisation is involved. But the bond angles are different due to bond pair-lone pair electron repulsion.

In CH_4, four bond pair of electrons are present and the bond angle is $109°28'$.

In H_2O, two bonding pairs and two lone pairs of electrons are present. The lone pair-lone pair repulsion is greater than lp-bp repulsion and the bond angle is reduced to $104.5°$.

In NH_3, three bonding pairs and one lone pair of electrons are present. The lone pair of electrons occupy more space and due to the greater repulsion between bond pair and lone pair result in the reduced angle to about $107°$.

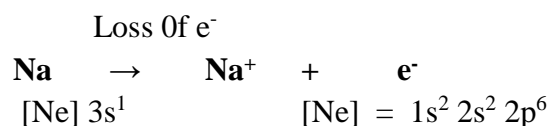**8. Explain the co-ordinate bond formation between BF₃ & NH₃.**

A co- ordinate covalent bond is formed between BF_3 and NH_3.

NH_3 molecule (Donar) gives a pair of electrons (lone pair) to BF_3 molecule which is electron deficient (acceptor) and has an empty orbital to accommodate the pair of electrons. Thus a co-ordinate bond is formed and the molecule as a whole is represented as $H_3N \rightarrow BF_3$

9. What is octet rule? Explain with an example.

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The tendency for atoms to have eight electrons in their outershell by interacting with other atoms through electron sharing or electron transfer is known as the octet rule of chemical bonding.

Loss 0f e⁻

$$Na \quad \rightarrow \quad Na^+ \quad + \quad e^-$$
[Ne] $3s^1$ [Ne] = $1s^2\,2s^2\,2p^6$

$$Cl \quad + \quad e^- \quad \rightarrow \quad Cl^-$$
[Ne] $3s^2\,3p^5$ [Ar] = $1s^2\,2s^2\,2p^6\,3s^2\,3p^6$

Na atom is ready to lose one electron to attain the nearest inert gas configuration [Ne] and Cl atom is ready to accept one electron to attain the nearest inert gas configuration [Ar].

10. What are the different types of bonds?

  (i) ionic (or) electrovalent bond

  (ii) covalent bond

  (iii) co-ordinate covalent (or) dative bond.

11. What is meant by electrovalent bond. Explain the bond formation in AlBr₃ and CaO.
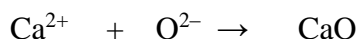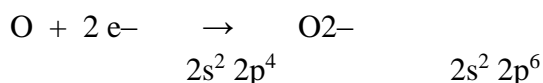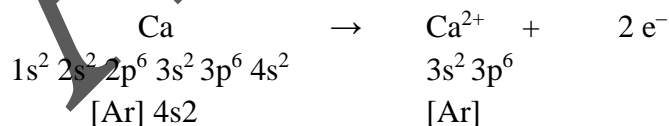
The binding forces existing as a result of electrostatic attraction between the positive and negative ions is termed as electrovalent or ionic bond.

Aluminium has 3 is in addition to the eight electron configuration and therefore gives away the three electrons and forms Al^{3+} [Bromine needs le⁻ to attain an eight e⁻ configuration so gains le⁻ and 3 bromine atoms take up 3 electrons and form 3 Br⁻]

In aluminium bromide (AlBr₃), aluminium ion has three positive charges and therefore it bonds with three bromide ions to form AlBr₃ which is a neutral ionic molecule.
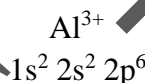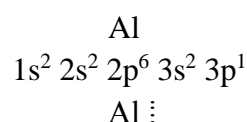
$$Al \quad \rightarrow \quad Al^{3+} \quad + \quad 3e^-$$
$1s^2\,2s^2\,2p^6\,3s^2\,3p^1$ $1s^2\,2s^2 2p^6$

$$3\,Br \quad + \quad 3\,e^- \quad \rightarrow \quad 3\,Br^-$$
$$Al^{3+} \quad + \quad 3\,Br^- \quad \rightarrow \quad AlBr_3 \text{ (Ionic bond)}$$
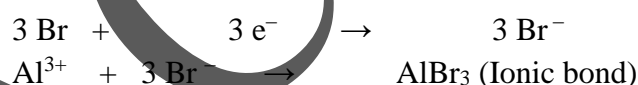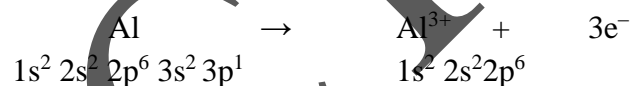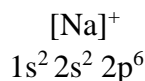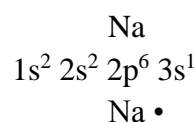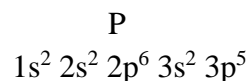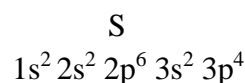
In CaO, the formation of the ionic bond involves two electron transfers from calcium to oxygen atom. Thus, doubly charged positive and negative ions are formed. An ionic bond is formed between Ca^{2+} and O^{2-} ions.

$$Ca \quad \rightarrow \quad Ca^{2+} \quad + \quad 2\,e^-$$
$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2$ $3s^2\,3p^6$
[Ar] 4s2   [Ar]

$$O \; + \; 2\,e- \quad \rightarrow \quad O2-$$
$2s^2\,2p^4$ $2s^2\,2p^6$

$$Ca^{2+} \quad + \quad O^{2-} \rightarrow \quad CaO$$

12. Give the electron dot representation for PH3 and ethane.

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13. Write the Lewis dot structures for the following. S, S²⁻, P, P³⁻, Na, Na⁺, Al and Al³⁺.

S S^{2-}

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

P P^{3-}

$1s^2\,2s^2\,2p^6\,3s^2\,3p^5$

Na   Na⁺

$1s^2\,2s^2\,2p^6\,3s^1$

Na • $[Na]^+$

                                            $1s^2\,2s^2\,2p^6$

Al Al^{3+}

$1s^2\,2s^2\,2p^6\,3s^2\,3p^1$

Al ⋮ Al^{3+}

                                            $1s^2\,2s^2\,2p^6$

14. What are the important features of valence bond theory?

(i)   A covalent bond is formed when the orbital of one atom is situated in such a way that it overlaps with the orbital of another atom, each of them contributing one unpaired electron.

(ii)   The two atomic orbitals merge to form a single bond orbital which is occupied by both the electrons.

15. What is meant by hybridisation?

Dissimilar orbitals like s, p, d with nearly the same energy on the same atom may combine or mix completely to form an equal number of equivalent energy new orbitals with properties of their on. This is known as hybridization of orbitals.

16. Define resonance. Give the various resonance structures of CO₂ and CO₃ ²⁻ ion.

Whenever a single Lewis structure cannot describe a molecular structure accurately, a number of structures with similar energy, positions of nuclei, bonding and non-bonding pairs of electrons are considered to represent the structure. Such structure is called canonical structure. This phenomenon is known as resonace.

E. Explain briefly on the following

1. Discuss the important properties of electrovalent compounds.

2. Calculate the lattice energy of NaCl using Born-Haber cycle.

3. Explain the important properties of covalent compounds.

4. Discuss the partial covalent character in ionic compounds using Fajan's rule.

5. Explain the polarity of covalent bonds in H2O and HCl.

6. Discuss the shapes of following molecules : NH3, H2O, CH4, PCl5 and SO2.

7. Discuss VSEPR model applied for linear, trigonal planar, tetrahedral and octahedral geometries of molecules.

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8. Explain the formation and difference between a sigma bond and a pibond. Which has more bond strength?

9. Calculate the lattice enthalpy of CaCl2 given that the enthalpy of :

i) Sublimation of Ca in 121 kJ mol-1

ii) Dissociation of Cl2 to 2Cl is 242.8 kJ mol-1

iii) Ionisation of Ca to Ca2+ is 2422 kJ mol-1

iv) Electron gain for Cl to Cl- is -355 kJ mol-1

v) û+f

(o) overall is -795 kJ mol-1

(Ans : 2870.8 kJ mol-1)

Ln11

B. Fill in the blanks

1. Relative lowering of vapour pressure is equal to **mole fraction of the solute (X_2)** in solution.

2. A liquid having high vapour pressure has **low** boiling point.

3. The least count of Beckmann's thermometer is **0.01 K**.

4. Molal elevation constant is a characteristic constant for a given **solvent**.

5. Semipermeable membrane allows the passage of **solvent** through it.

6. For a deliquescence to occur, the vapour pressure of water in the air must be **greater** than that of the saturated solution.

7. Depression in freezing point is **more** pronounced if camphor is used as a solvent in place of water for same amount of solute and solvent.

8. Every solution behaves as ideal solution **when the obey Raoult's law** .

9. The osmotic pressures of 0.1M glucose and 0.1M NaCl solutions are **different** .

10. Solutions that have same osmotic pressure are called **isotonic** solutions.

C. Answer the following in one (or) two sentences

1. What are colligative properties?

The colligative properties are the properties of solution that depend on the number of solute particles dissolved in it and independent on the nature of the particles.

(e.g.,) lowering of vapour pressure $-\Delta P$, osmotic pressure $-\pi$.

2. Define relative lowering of vapour pressure.

The relative lowering of vapour pressure is defined as the ratio of the lowering of vapour pressure to the vapour pressure of the pure solvent.

$P^0 - P / P^0$ = Relative lowering of vapour pressure

Where p^0 = vapour pressure of pure solvent

P = Vapour pressure of the solution.

3. What do you understand by molal elevation of boiling point? What are abnormal solutes?

Molal elevation of boiling point is defind as the boiling point elevation produced when 1 mole of the solute is dissolved in one kg (1000 g) of the solvent.

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Abnormal solutes are the solutes which associate to form dimers (or) dissociate completely to give more than one ions. **(e.g.,)** Nacl, Kcl, CH_3COOH, C_6H_5COOH.

4. Addition of non-volatile solute always increases the boiling point of the solution. Why?

When a non-volatile solute is added to the solvent, the vapour pressure of the solution decreases. Because of the interaction between the solute and solvent, the escaping tendency of the solvent to the vapour state decreases . Since the vapour pressure of the solution is lower than that of pure solvent, the boiling point of a solution will be higher than that of pure solvent.

5. Volatile hydrocarbons are not used in the brakes of automobile as lubricant, but non-volatile hydrocarbon are used as lubricants. Why?

When volatile hydrocarbons are used in the brakes of automobile as lubricant, they are readily converted to vapour and their lubricating nature is reduced in the failure of brake applications. So always a non-volatile (non-vaporisable) hydrocarbon must be used as lubricants.

6. Prove that the depression in freezing point is a colligative property.

Freezing point of solvent T^0. Freezing point of solution T.

$T < T^0$

$\Delta T_f = T^0 - T$

ΔT_f is depression in freezing point. It is directly proportional to the molarity of the solution.

$\Delta T_f \, \alpha \, m$

$\Delta T_f = K_f \cdot m$

Where K_f cryoscopic constant which contains a definite number of solute particles K_f is defined as the depression in freezing point produced when one mole of solute is dissolved in 1 kg solvent. Freezing point depression (ΔT_f) of a dilute solution is found to be directly proportional to the number of moles of the solute dissolved in a given amount of the solvent which contains a definite number of solute particles. ΔT_f is independent of nature of the solute as long as it is non-volatile. Hence , depression in freezing point is considered as a colligative property.

7. Explain the terms osmosis and osmotic pressure.

Spontaneous movement of solvent particles from a dilute solution orfrom a pure solvent towards the concentrated solution through asemipermeable membrane is known as **osmosis** (Greek word : 'Osmos' = to push). The flow of the solvent from its side (a) to solution side (b) separated by semipermeable membrane (c) can be stopped if some definite extra pressure is applied on the solution risen to height (h). This pressure that just stops the flow of solvent is called **osmotic pressure** of the solution. This pressure ($\Box$) has been found to depend on the concentration of the solution.

Fig. 11.8

8. What are isotonic solutions?

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Two solutions of different substaces having the same osmotic pressure at same temperature are said to be isotonic to each other. They are known as isotonic solutions.

9. What are the advantages of Berkley-Hartley method?

1. The osmotic pressure is recored directly and the method is quick.

2. There is no change in the concentration of the solution during the measurement of osmotic pressure.

3. The osmotic pressure is balanced by the external pressure and there is minimum

10. Explain how the degree of dissociation of an electrolyte may be determined from the measurement of a colligative property.

The solutes which dissociates in solvent are called electrolytes. They show an increase in number of particles present in solution. This effect results in an increase in colligative properties obtained experimentally.

The Van't Hoff factor(i) =experimental colligative property

Noramal colligative property

So we can calculate the degree of dissociation (α) using the following equation.

α = i -1/ n - 1

Where 'n' is the total number of particles furnished by one molecule of the solute.

Problems:

1. The vapour pressure of pure benzene at a certain temperature is 640 mm of Hg. A non-volatile non-electrolyte solid weighing 2.175 g is added to 39 g of benzene. The vapour pressure of the solution is 600 mm of Hg. What is molecular weight of solid substance? [69.6]

2. Calculate the freezing point of an aqueous solution of a non-electrolyte having an osmotic pressure 2.0 atm at 300 K. K_f = 1.86 k.kg.mol-1. R = 0.0821 lit.atm.k-1 mol-1 [-0.151□C]

3. What weight of non-volatile solute (urea) NH2 CO NH2 needs to be dissolved in 100 g of water in order to decrease the vapour pressure of water by 25%. What will be the molality of solution?   [13.88 m]

4. 20 g of sucrose solution in one litre is isotonic with a solution of boric acid containing 1.63 g of boric acid in 450 ml. Find the molecular weight of boric acid. [61.94]

5. A solution containing 6 gm of a solute dissolved in 250 ml of water gave an osmotic pressure of 4.5 atmosphere at 27□C. Calculate the boiling point of the solution. The molal elevation constant for water is 0.52 [373.095]

D. Explain briefly on the following

1. Explain the determination of relative lowering of vapour pressure by Ostwald- Walker method?

2. Describe about Beckmann thermometer.

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3. Explain the determination of depression in freezing point by Beckmann method.

4. What is elevation of boiling point? Explain its determination by Cottrell's method.

5. Explain the laws of osmotic pressure? Explain its determination by Berkley-Hartley method.

6. What are abnormal colligative properties? Explain with example and write its determination using Van't Hoff factor.

Ln:12

B. Fill in the blanks

1. Translational energy of molecules is a part of **Total internal** energy of the system.

2. Specific heat of a liquid system is **intensive** property.

3. Work done in the reversible expansion is **greater than other process** .

4. Combustion is an **exothermic** process .

5. Heat of neutralisation of a strong acid is **greater** than that of a weak acid.

C. Write in one or two sentence:

1. Name the equipment using which heat of combustion of compounds are determined?

Enthalpy changes of combustion of chemical substances are experimentally determined using a bomb calorimeter.

2. Energy can be created and be destroyed. State whether this is true or false.

False. Energy can neither be created nor be destroyed.

3. Define zeroth law of thermodynamics.

Zeroth law of thermodynamics states that "if two systems at different temperatures are separately in thermal equilibrium with a third one, then they tend to be in thermal equilibrium with themselves". (or)

"When two objects are in the thermal equilibrium with the third object, then there is thermal equilibrium between the two objects itself".

4. Give the relation between Δu and ΔH.

$$\Delta H = \Delta u + p \Delta v$$

(or)

$$\Delta H = \Delta u + \Delta ng \, RT$$

5. Define an adiabatic process.

Adiabatic process is defined as that one which does not exchange heat with its surroundings during the change from initial to final states of the system.

6. Write the differences between an exothermic and an endothermic process.

Exothermic process	Endothermic process
A process when transformed from initial to final states by evolution of heat.	A process when transformed from initial to final states by absorption of heat.

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The final state of the system possess lower energy than the initial state. The excess energy is evolved as heat.	The final state of the system possess higher energy than the initial state. The excess energy needed is absorbed as heat by the system from the surroundings.
If the physical transformation is exothermic, heat is removed to bring about the initial to final state.	Generally in a physical transformation which is endothermic heat is supplied to bring about the initial to final state.
(e.g.,) Freezing of a liquid at its freezing point is an endothermic process	e.g.,) Melting of a solid by supplying heat is an exothermic process.

7. What are intensive and extensive properties?.

Intensive properties are the properties that are independent of the mass or size of the system . **(e.g.,)** Refractive index, surface tension, density, temperature, boiling point, freezing point.

Extensive properties are the properties that depend on the mass or size of the system. **(e.g.,)** Volume, number of moles, mass energy, interal energy, etc.

8. Define first law of thermodynamics.

First law of thermodynamics states that "energy may be converted from one form one form to another, but cannot be created or be destroyed".

Energy of an isolated system must remain constant although it may be changed from one to another.

For an isothermal process, $\Delta U = 0$ q = -W

9. Explain thermal and mechanical equilibrium processes.

A system which satisfies the conditions of thermal, mechanical and chemical equilibria and contains the macroscopic properties which are independent of time is said to be in thermodynamic equilibrium.

Thermal equilibrium	**Mechanical equilibrium**
It sets the condition that there should be no flow of heat from one portion or part of the system to another part of the system.	It implies that there is no work done by one portion or part of the system over another portion or part of the same system
(i.e.,) temperature of the system remaining constant at every point of the system.	**(i.e.,)** pressure of the system being constant at all its points.

D. Explain briefly on the following

10. Describe a bomb calorimeter and explain how heat of formation of an organic compound is determined.

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11. Compare the enthalpy changes that occur between the neutralization of a strong acid and a weak acid by sodium hydroxide. Explain the differences seen

Ln 13 Chemical Equilibrium-I

B. Fill in the blanks

1. In endothermic equilibrium reaction the increase in temperature **increase the K_{eq}** of the reaction.

2. When the reactant is a liquid which decomposes to gaseous products. Then the equilibrium is called as **liquid-vapour equilibrium**

3. When reactants and products are in gaseous state, the equilibrium constant can be expressed in terms of **partial pressure**

4. Value of the equilibrium constant is **independent** of the initial concentration of reactants.

5. According to law of mass action, the rate of a chemical reaction is proportional to **product of active masses** of reactants.

C. Match the following

1. Kp $K_C (RT)^{\Delta n}$
2. $CaCO_3 \leftrightarrow CaO(s) + CO_2(g)$ Heterogeneous equilibrium
3. Rate of reaction   active mass of reactants
4. $H_2(g) + I_2(g) \leftrightarrow 2HI(g)$ homogeneous equilibrium
5. $C(s) + O_2(g) \rightarrow CO_2(g)$ irreversible reaction

D. Write in one or two sentence

1. Define law of mass action.

Law of mass action states that " the rate of a chemical reaction is directly proportional to the product of active masses of the reactants".

2.Write the Kp expression for $PCl_{5(g)} \leftrightarrow PCl_{3(g)} + Cl_{2(g)}$

$$PCl_{5(g)} \leftrightarrow PCl_{3(g)} + Cl_{2(g)}$$

$$K_p = P_{pcl_3} . P_{cl_2} / P_{pcl_5}$$

In terms of partial pressure values

$$K_p = x^2 p/ 1-x^2$$

3.Relate Kp and Kc when $\Delta n = 0$, $\Delta n = 1$, $\Delta n = 2.0$.

$$K_p = K_c. RT^{\Delta n}$$

When $\Delta n = 0$

$$K_p = K_c \qquad [RT^0 = 1]$$

When $\Delta n = 1$

$$K_p = K_c. RT$$

When $\Delta n = 2$

$$K_p = K_c (RT)^2$$

4.Give an example of irreversible reaction.

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Reactions which go to completion and never proceed in the reverse direction are called irreversible reaction.

(e.g.,)   2 Na + 2 H$_2$O → 2 NaOH + H$_2$

5. Reason out why equilibrium concentrations remain constant.

The equilibrium concentrations of reactants and products remains constant. This is because, since the forward reaction rate equals the backward reaction rate,as and when the products are formed, they react back to form the reactants in equal capacity.

E. Explain briefly on the following

21. Differentiate irreversible and reversible reactions.

22. Explain the characteristics of a chemical equilibrium.

23. Write a note on heterogeneous equilibrium reaction.

24. Two moles of H2 and three moles of I2 are taken in 2 dm3 vessel and heated. If the equilibrium mixture contains 0.8 moles of HI, calculate Kp and Kc for the reaction H2(g) + I2(g) 2HI(g) Ans : (0.036)

25. At 25̊C, Kc for the reaction 3C2H2(g) C6H6(g) is 4.0. If the equilibrium concentration of C2H2 is 0.5 mol. lit-1. What is the concentration of C6H6?

(Ans) [C6H6] = 0.5 mol.lit-1

Ln 14 Chemical Kinetics-I

B. Fill up the blanks

1. Decomposition of aqueous NH$_4$NO$_2$ proceeds by **first order** reaction.

2. Fractional orders are found in **polymerization** reaction.

3. In a **zero order** reaction rate does not depend on the reactant concentration.

C. Match the following

1. slow step   rate determining step

2. order   experimentally determined

3. molecularity   theoretical concept concentration

4. unit of first order `k`   sec-1

5. rate is independent of reactant   zero order

detail

Describe the factors on which the rate of a reaction depends.

i)Effect of nature of reactant and product:

Changing the chemical nature of any reacting species will change the rate of the reaction. For example, in halogenation reactions, the reactions involving iodine is found to be slower than those involving chlorine.

ii) Effect of reacting species (Effect of concentration): Increasing the concentration of the reactant increases the rate of reaction.

(iii) **Effect of temperature :** Increasing the temperature, increases the rate of reaction.

(iv) **Effect of a catalyst :** A catalyst alters the rate of reaction. A positive catalyst increases the rate of reaction and a negative catalyst decreases the rate of reaction.

(v) **Effect of surface area of the reactant :** Greater the surface area of the reactants, greater is the rate of reaction. Thus the reactants in a powdered form reacts fastly than present as larger particles. eg., Powdered zinc reacts with dil. H$_2$SO$_4$ faster than a zinc rod.

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(vi) **Effect of radiation :** Photochemical reactions are influenced by light radiation. The rate of photochemical reactions depend on the intensity of incident radiation.

D. Write very short answers
1. Define half life period.

Half life period ($t_{1/2}$) is defined as the time requird for half the iitial concentration of the reactant to react. (or) It is defined as the time required for 50% completion of the reaction.

t $_{1/2}$ = 0.6932/ k

2.Name the factors that affect the rate of reaction.
(i)　Nature of the reactants and products
(ii)　Concentration of the reacting species
(iii)　Temperature of the system
(iv)　Presence of catalyst
(v)　Surface area of the reactants
(vi)　Exposure to radiation

3.What is molecularity?

Molecularity is defined as the number of atoms or molecules taking part in an elementary step leading to a chemical reaction .

4.What is a rate determining step?

In a multi step reaction the step which has the lowest rate value among the other steps of the reaction is called the rate determining step or rate limiting step.

5.List the factors on which an order of the reaction depend.

Order of the reaction depends upon pressure, temperature and concentration.
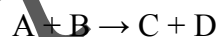
6. Write the rate law of p A + q B ⟶ l C + m D reaction.

p A + q B ⟶ l C + m D

The rate law is given by the expression rate α $[A]^p . [B]^q$

Rate = k $[A]^p . [B]^q$ where k is rate constant.

7. Define the rate of a reaction.

The rate of the reaction is defined as the change in the concentration of any one of the reactant or product in the reaction per unit time.

For a general reaction,

A + B → C + D

Rate = -d[A] / dt = -d[B] / dt = +d[C] / dt = +d[D] / dt

E. Explain briefly on the following
1. Compare and contrast the terms, order and molecularity of a reaction.
2. Describe the factors on which the rate of a reaction depends.

Ln15. BASIC CONCEPTS OF ORGANIC CHEMISTRY

A. Write IUPAC name of the following

(a) Ans : 2-propanol

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(b) Ans : 2- methyl-2-propanol

c) **Ans : 3-methyl-2-butanol**

d) **Ans : 2-methyl-2-butanol**

(e) **Ans : pent-1-ene-3-one**

(f) **Ans : propanal**

(g) **Ans : 3-methyl butanoic acid**

(h) **Ans : ethoxy ethane**

(i) **Ans : 1-methoxy propane**

(j) **Ans : 2-methoxy propane**

(k) **Ans : 1-amino butane**

B. Explain briefly on the following

1. Homolytic and heterolytic fission.

2. Substitution reaction.

3. Addition reaction.

4. Elimination reaction.

5. Polymerisation reaction.

6. Condensation reaction.

7. Hydrolysis.

8. Reduction and oxidation reactions.

9. Electrophilic and Nucleophilic reagents.

10. Carbonium ions and carbanions.

11. Free radicals.

12. Inductive effect.

13. Resonance effect.

Ln 16. Purification of Organic Compounds

B. Fill in the blanks

1. The Compounds separated and purified by crystallisation can be dried **by keeping in sunlight (or) by using infra-red light** over

2. Camphor can be purified by the process of sublimation.

3. In simple distillation the compounds should not decompose at **ordinary pressure**

4. Water insoluble compounds can be purified by **steam distillation** .

5. In T.L.C the stationary phase is a **thin layer of silica gel or alumina on a glass plate.**

6. Chromatographic technique was first introduced by **M.S.Tswett** .

7. In paper chromatography, the mobile phase travels by **capillary** action through the paper.

8. The adsorbent used in column Chromatography method is **alumina (or) MgO (or) silica gel (or) starch.**

9. In Chromatographic technique, the separation of compounds are brought about by **differential** movement of the compounds.

10. Paper Chromatography is **partition** Chromatography.

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C. Write in one or two sentence

1. What are the different stages followed during Crystallisation?

Crystallisation is carried out in four stages. (a) preparation of the solution of the substance I a suitable solvent, (b) filtration of the hot solution, (c) crystallization by cooling the hot filtrate, (d) isolation and drying of the purified substance.

2.Define steam distillation.

The process of separation and purification of organic compounds (solid as well as liquids) which are immiscible in water and not decomposable at steam temperature by the passage of steam through them.

3.What are different types of distillation?

(i) Simple distillation

(ii) fractional distillation

(iii) Steam distillation

(iv) distillation under reduced pressure.

4. Give the advantages of distillation under reduced pressure

(i)   The compounds which decompose o heating to their boiling points under normal pressure can be purified by distillation under reduced pressure. This is because under reduced pressure, a liquid would boil at temperature much below its normal boiling point.

(ii)   In distillation under reduced pressure, a liquid boils at temperature well below the normal boiling point. So, the distillation under reduced pressure is more fuel economical.

5. What are the types of paper chromatography?

(i)   Ascending paper chromatography. The mobile phase moves upwards on the paper strip in this case

(ii)   Descending paper chromatography. The mobile phase moves downwards on the paper strip.

(iii)   Radial or circular paper chromatography. The mobile phase moves horizontally along a circular sheet of paper

D. Explain briefly on the following

1. Explain the method of purifying a solid organic compound.

2. Write short notes on a) Fractional crystallisation b) Solvent extraction

3. Explain the purification of compounds by using thin layer chromatography.

4. What are the various principles used in chromatographic separation?

5. Write down the general characteristics of organic compounds.

Ln17. Detection and Estimation of Element

Ln 18.Hydrocarbons

B. Fill up the blanks

1) In alkanes, the carbon atoms are connected by **sigma (σ)** bonds.

2) Treatment of 1,2-dibromopropane with zinc and ethanol gives **propylene .**

3) C is But-2-ene is an **Z or geometrical** isomer.

4) Addition of HCl to an olefin follows **Markovnikov's** rule.

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5) An alkene reacts with ozone to form **carbonyl compound**

6) CaC2 on hydrolysis gives **acetylene**

7) Ethylenedibromide on treatment with KOH gives **Acetylene**

8) Electrolysis of sodium maleate gives **Acetylene**

C. Explain briefly on the following

1) Mention any five chemical properties of alkanes.

2) Discuss the general methods of preparing alkanes.

3) What is hydroboration?

4) What is ozonolysis?

5) What is witting reaction?

6) What is polymerisation?

7) How is ethylene hydrated?

8) What is the action of ozone on acetylene.

9) What happens when acetylene is passed through red-hot tube?

Ln 19. Aromatic Hyrocarbons

B. Fill in the blanks

1. Many synthetic drugs used are **aromatic** in part.

2. The **coal tar** forms the source of many organic compounds.

3. The modern theory of aromaticity was introduced by **Huckel**.

4. Ortho and para directing groups are called as **activating** groups.

5. Meta directing groups are called as **deactivating** groups.

6. Alkyl substituted benzenes are prepared by **Friedel- Craft's alkylation (or) Wurtz-fitting** reaction.

7. Naphtha obtained by fractional distillation of **petroleum** is passed over platinum.

8. Aromatic compounds readily undergo **electrophilic** substitution reactions.

9. **Fluorine** reacts vigorously with aromatic hydrocarbons even in the absence of catalyst.

10. In the presence of **platinum** benzene reacts with hydrogen to give cyclohexane.

C. Explain briefly on the following

1. How is benzene is prepared commercially?

2. Explain the term aromaticity.

3. Write a note an activating groups in benzene.

4. How would you convert the following?

a) sodium benzoate to benzene b) phenol to benzene c) benzene to toluene

5. Write briefly on resonance in benzene.

Ln 20. Organic Halogen Compounds

B. Fill in the blanks

1. Markonikoff's rule is followed for the addition of HCl to **alkenes**

2. In Swarts reaction metallic fluorides are added to **alkyl chloride (or) bromide**

3. Hoffman's rule is applicable to elimination of **alkene with more than two β hydrogen atom**

4. Chloropicrin is prepared by adding nitric acid to **$CHCl_3$**

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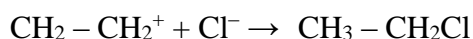
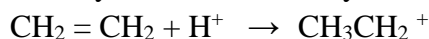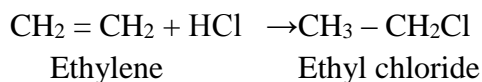### C. Write in one or two sentence

1. What are Lewis acids?

Lewis acids are the mixture of concentrated hydrochloric acid and anhydrous $ZnCl_2$. $ZnCl_2$. They accept a pair of electrons.

**(e.g.,) BF_3, $AlCl_3$, Ag^+, Na^+**

2. What is an electrophilic addition?

Alkenes react with hydrogen halide to form alkyl halide. These addition reactions are initiated by electrophile therefore called electrophilic addition reactions.

$$CH_2 = CH_2 + HCl \rightarrow CH_3 - CH_2Cl$$

Ethylene   Ethyl chloride

$$CH_2 = CH_2 + H^+ \rightarrow CH_3CH_2{}^+$$

$$CH_2 - CH_2{}^+ + Cl^- \rightarrow CH_3 - CH_2Cl$$

3. What is Hunsdiecker reaction?

Silver carboxyiates in carbon tetrachloride are decomposed by chlorine or bromine to form alkyl halide.

$$CH_3CH_2COOAg + Br_2 \rightarrow CH_3CH_2Br + CO_2 + AgBr$$

Silver propionate   Ethyl bromide
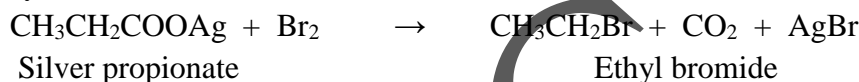
4. What is Finkelstein reaction?

Alkyl iodides are prepared by treating the corresponding chloride or bromide with a solution of sodium iodide in acetone. The exchange of halogen between alkyl halide and sodium iodide occurs.

$$RCl + NaI \xrightarrow{Acetone} RI + NaCl$$

$$RBr + NaI \xrightarrow{Acetone} RI + NaBr$$

Alkyl iodide

5. What is Swarts reaction?

Alkyl fluorides are obtained by treating alkyl chloride or bromide with metallic fluorides such as AgF or SbF_3. It is Swarts reaction.

$$CH_3 - CH_2Br + AgF \rightarrow CH_3 - CH_2F + AgBr$$

Ethyl bromide   Ethyl fluoride
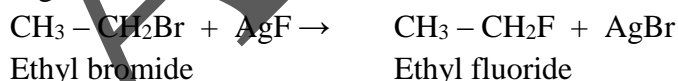
D. Explain briefly on the following

1. Discuss SN1 mechanism
2. Discuss SN2 mechanism
3. Discuss E1 elimination
4. Discuss E2 elimination
5. What are the uses of alkyl halides?
6. What are the general reactions of aryl halides.
7. What are aralkyl halides ? How are they prepared?

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8. What are Grignard reagents? Discuss its synthetic uses.
9. Discuss the general methods of preparation of alkyl halides.

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