

11th SOME BASIC CONCEPTS OF CHEMISTRY (SOLUTION)

01. (2) Gases occupy the volume of container. So they don't have definite volume or shape. Liquids don't have definite shape but have definite volume.
02. (3) Salt and sugar don't form homogeneous mixtures pulses and grains also don't form homogeneous mixtures.
03. (4) $1\text{L} = 1000\text{ cm}^3 = 1\text{ dm}^3 = 10^{-6}\text{ m}^3$
04. (4) 1.93 and 1.95 are close to each other but are far away than actual value. So they are precise but not accurate.
05. (1) 100 - one significant digit, zeroes on the right of non-zero digit should not be counted.

06. (3) One amu is equal to $\frac{1}{12}$ th of mass of one atom of C-12 isotope. 16 amu means 16 times heavier than $\frac{1}{12}$ th of mass of one atom of C-12 isotope.

07. (1) Molecular mass = mass of one molecule
Molar mass = mass of one mole molecule
Hence, molecular mass is lesser than molar mass.

08. (2) 72 g glucose

$$\text{moles} = \frac{72\text{ g}}{180\text{ g/mol}} = 0.4\text{ moles}$$

$$\text{no. of moles of H-atom} = 0.4 \times 12 = 4.8\text{ moles}$$

09. (4) One molecule of H_2O has mass of 18 amu which is also equal to $18 \times 1.66 \times 10^{-24}\text{ g}$.

10. (2) Moles = $\frac{4.07}{1}$ $\frac{24.27}{12}$ $\frac{71.65}{35.5}$
= 4.07 2.02 2.02
Dividing by least number 2.02
= 2 1 1

Hence empirical formula is H_2CCl empirical formula mass = 49.5

$$\text{Molar mass} = 98.9 \quad \text{Ratio} = 0.5$$

11. (3) Molecular formula = (empirical formula) $\times n$
 n is equal to or more than 1

$\frac{\text{empirical formula}}{\text{molecular formula}} = \frac{1}{n}$ which must be less than or equal to 1.

Hence it can never be more than 1.

12. (2) $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$
1 mole CH_4 reacts with 2 moles O_2
1 mL CH_4 reacts with 2 mL O_2
1 L CH_4 reacts with 2 L O_2
1 mole CH_4 gives 2 mole H_2O vapours.

13. (1) When 3g of N_2 reacts with 3g H_2

$$\text{Moles of } \text{N}_2 = \frac{3}{28}$$

$$\Rightarrow \text{Moles of } \text{H}_2 = \frac{3}{2}$$

$$\text{Since, } \frac{3}{28 \times 1} < \frac{3}{2 \times 3}$$

Hence N_2 is the limiting reactant.

14. (3) Moles of methane = 2 moles

For 2 moles methane, moles of water produced = 4 moles

Since only 85% methane reacted,

$$\text{Mass of water produced} = 0.85 \times 4 \times 18 = 61.2$$

15. (1) Avg. molar mass

$$\Rightarrow 30 = 25x_A + 40(1 - x_A)$$

$$\Rightarrow 30 = -15x_A + 40$$

$$\Rightarrow x_A = \frac{2}{3}$$

$$\text{Total moles} = \frac{\text{Given total mass}}{\text{Molar mass}}$$

$$= \frac{100}{30} = \frac{10}{3}$$

$$\text{Moles of A} = \frac{2}{3} \times \frac{10}{3} = \frac{20}{9} = 2.2$$

16. (2) When dilution is done by adding the solvent the concentration of solution decreases. Hence all types of concentration decreases, but amount of solute remains same.

17. (1) 1) $9 \times 3.5 N_A$

$$2) 2.4 \times 10^{24} = 24 \times 10^{23} = 4 N_A$$

$$3) 18 N_A$$

$$4) 21 N_A$$

Hence first option has highest number of O-atoms.

18. (2) $39.5 = 40x + 36 \times 0.1 + 38(0.9 - x)$

$$39.5 = 2x + 3.6 + 34.2$$

$$= 39.5 = 2x + 37.8$$

$$1.7 = 2x$$

$$\Rightarrow 0.85 = x$$

19. (2) Only second option satisfies the given condition of empirical formula. Here $n = 1$.

20. (4) $2.79 = \frac{1000 \times M}{1000 \times 1.25 - 58.5 M}$

$$\Rightarrow 3487.5 - 163.2 M = 1000 M$$

$$M = \frac{3487.5}{1163.2} = 3$$

21. (4) Moles of Na_2CO_3

$$= \text{Molarity} \times \text{Volume of solution(L)}$$

Since volume of solution is not given, we cannot find the no. of moles.

22. (2) i) $\frac{300}{1} > \frac{200}{1}$ ii) $\frac{2}{1} > \frac{1.5}{1}$

iii) $\frac{100}{1} > \frac{50}{1}$ iv) $\frac{3}{1} > \frac{2.5}{1}$

In all the cases B is the L.R.

23. (3) In 36 g glucose,

$$\text{Moles of glucose} = \frac{36}{180} = 0.2$$

$$\text{No. of moles of O atoms} = 6 \times 0.2 = 1.2$$

$$\text{No. of O atoms} = 6 \times 0.2 \times N_A$$

$$\text{No. of glucose molecules} = 0.2 N_A$$

$$\text{No. of hydrogen atoms} = 12 \times 0.2 N_A$$

24. (2) For compound A $\Rightarrow \frac{X}{Y} = \frac{28}{16}$

For compound C $\Rightarrow \frac{X}{Y} = \frac{28}{?}$

Ratio of Y in A to C is 1 : 3

$$\therefore Y \text{ in C is } 16 \times 3 = 48$$

25. (3) No. of O-atoms in 1 mol $\text{H}_2\text{SO}_4 = 4 N_A$

$$\text{No. of O-atoms in 1 mol } \text{CaCO}_3 = 3 N_A$$

Hence statement-1 is false.

26. (4) When different compounds are formed by combining similar elements we apply law of multiple proportions.

27. (2) The product is always formed according to amount of limiting reactant. So if limiting reactant is same, amount of product should not change.

Stoichiometric coefficient represent moles or molecules of a substance involved in a reaction.

Hence both statements are correct, but statement-2 is not giving proper explanation.

28. (3) Two molar solution of sulphuric acid has molarity equal to 2. But to calculate the mass of sulphuric acid present, we must know the volume of solution which is not given. Hence statement-1 is false.

Statement-2 is the definition of molarity.

29. (1) Molality is moles of solute per kg of solvent. Since both moles and mass do not change with temperature, so molality is independent of temperature.

30. (4) Both temperature and pressure can be changed to change matter states. Gases can be converted to liquids by increasing pressure. Liquids can be converted into vapour by increasing temperature.

31. (3) Milli = 10^{-3} , Nano = 10^{-9} ,
Deci = 10^{-1} , Pico = 10^{-12}

32. (3) $^{\circ}\text{F} = \frac{9}{5} \times ^{\circ}\text{C} + 32 = \frac{9}{5} (-40) + 32 = -40^{\circ}\text{F}$

33. (3) Number of moles of Fe atoms = $\frac{0.34 \times 3.3}{100 \times 56}$

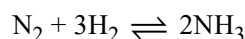
$$\begin{aligned} \text{Number of Fe atoms} &= \frac{0.34 \times 3.3}{56 \times 100} \times 6.02 \times 10^{23} \\ &= 1.21 \times 10^{20} \end{aligned}$$

34. (1) 20 ppm means, 20 g solute in 10^6 g solution

$$\text{mass \%} = \frac{20\text{g}}{10^6\text{g}} \times 100 = 2 \times 10^{-3}$$

$$\therefore 20 \text{ ppm} = 20 \times 10^{-4} \quad \% \frac{w}{w}$$

35. (1) In the formation of NH_3 ,



N_2 and H_2 can be present in any ratio, but they will react in the ratio 1 : 3. Hence statement-1 is true. However statement-2 is correct.

36. (2) Sulphur has atomic mass of 32.

Since compound contains 8% sulphur.

Let molar mass of compound is M

$$\Rightarrow 8\% \text{ of } M = 32$$

$$\Rightarrow M = \frac{32}{0.08} = 400$$

37. (1) Moles of $\text{CO}_2 = \frac{200 \times 10^{-3}}{44}$

$$\text{Initial molecules} = \frac{200 \times 10^{-3}}{44} \times 6.02 \times 10^{23}$$

$$\text{Final molecules} = \frac{200 \times 10^{-3}}{44} \times 6.02 \times 10^{23} - 10^{21}$$

$$\text{Final moles} = \frac{1.7 \times 10^{21}}{6.02 \times 10^{22}} = 2.85 \times 10^{-3}$$

38. (4) The S.I unit of volume is not litre. It is m^3 .

39. (4) At given temperature and pressure, all gases do not occupy same volume. It depends upon their moles.

According to ideal gas equation, $PV = nRT$,

Volume occupied by gas depend upon pressure and temperature.

Hence, equal amount of all gases, at same temperature and pressure occupy equal volumes.

40. (3) Mass of 1 molecule of water = $18 \times 1.66 \times 10^{-24} \text{ g}$

$$\begin{aligned} \text{Volume} &= \frac{\text{mass}}{\text{density}} = \frac{18 \times 1.66 \times 10^{-24}}{1} = 30 \times 10^{-24} \\ &= 3 \times 10^{-23} \text{ cm}^3 \end{aligned}$$

41. (3) $\text{C}_2\text{H}_4 + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 2\text{H}_2\text{O}$

2.8 kg will require 9.6 kg O_2 because

28 kg ethylene requires 96 kg O_2 .

42. (1) Mol is the unit of amount of substance while molarity is the unit of concentration.

43. (1) $M_{\text{Final}} \times V_{\text{Final}} = MV_1 + MV_2$

$$M_{\text{Final}} = \frac{M(V_1 + V_2)}{V_{\text{Final}}} = M$$

44. (1) Micro = 10^{-6} , Deca = 10,

Mega = 10^6 , Giga = 10^9 , Femto = 10^{-15}

45. (4) Check limiting reactants in all cases and H_2O will form according to limiting reactant.

(p) $\frac{10}{1} = \frac{5}{0.5}$ hence both are limiting reactant.

(q) $\frac{20}{1} > \frac{2.5}{0.5}$ hence O_2 is L.R. so 5 volume will form.

46. (3) Molarity = $\frac{\text{Moles of solute}}{\text{Vol. of solution (L)}}$

$$\begin{aligned} &= \frac{w_1/M_1 \times 1000}{\rho} \end{aligned}$$

47. (2) $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$

16 g methane gives 44 g CO_2

Hence 8 g methane gives 22 g CO_2

48. (4) 1 dm^3 is same as 1 litre. So, it is equal to 1000 ml or 1000 cm^3 . Also $1 \text{ m}^3 = 10^3 \text{ L}$ or $1 \text{ L} = 10^{-3} \text{ m}^3$.

49. (3) There is no change in total mass of reaction mixture during a chemical reaction.

50. (3) Molality = $\frac{\text{Moles of solute}}{\text{Mass of solvent (kg)}}$

STRUCTURE OF ATOM (SOLUTION)

01. (4) 4 is correct i.e. high voltage, low pressure)

At high voltage of around 10000 volt applied then beam of rays arise from cathode which consists of negatively charged particle reaches to the anode.

At low pressure, no hindrance generated to the beam of cathode rays.

02. (3) Cathode, anode, anode

03. (4) Both (1) and (2)

Cathode rays get deflected in electric as well as magnetic field because of the charge (i.e., negative charge present on it)

04. (1) As we increases the charge on the particle, more deflection will occur in the presence of electric field.

05. (2) Canal rays depends upon the nature of gas taken in discharge tube.

06. (3) Millikan's oil drop experiment responsible for the discovery of charge of electron.

07. (3) A very
- α
- particle (1 in 20000) were deflected by
- 180°
- because at the centre of an atom most of the mass is concentrated i.e., charge density is very high so, when it reaches near to the nucleus, it get back at an angle of
- 180°
- .

08. (4) Mass number can be whole number or a fraction. Mass number cannot be in fraction.

09. (1) Same mass number .

Isobar are there species which have same mass number and atomic number is different so chemical properties is different.

10. (3)
- ${}_{17}^{35}\text{Cl}^-$

As ${}_{17}^{35}\text{Cl}^- \rightarrow$ is ion \rightarrow Anion

$$Z = P \neq e, \quad Z = 17, \quad P = 17, \quad e = 18$$

$$A = P + n, \Rightarrow n = A - P, \Rightarrow n = 35 - 17 = 18$$

11. (1)

12. (4) Energy decreases,
- λ
- increases,
- ν
- decreases

$$\text{Since we know that : } E \propto \nu \propto \frac{1}{\lambda}$$

Decreasing energy order \rightarrow

Gamma rays β Rays α Rays X-Rays UV Rays Visible

Infrared Microwaves Radio wave

\leftarrow Decreasing wavelength

13. (3)
- $\nu = \frac{c}{\lambda}, \quad \lambda = \frac{c}{\nu} = \frac{3 \times 10^7 \times 5}{684}$

$$\nu = \frac{c}{\lambda}, \quad \lambda = \frac{c}{\nu} = \frac{3 \times 10^8}{1368} = 2.192 \times 10^5 \approx 219.3 \text{ m}$$

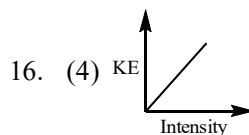
14. (2) Diffraction

Wave nature of light \rightarrow Diffraction and interference

Particle nature of light \rightarrow Black body radiation

\rightarrow Photoelectric effect

15. (3) Only statement-1 is correct, but statement-2 is incorrect.



Kinetic energy is independent of photo intensity.

17. (2)
- $E = W + KE$

$$\frac{124 \times 1.6 \times 10^{-19}}{300} \text{ J} = W + \frac{1.68 \times 10^5}{6.022 \times 10^{23}}$$

$$6.6 \times 10^{-19} \text{ J} = W + 0.278 \times 10^{-18}$$

$$0.66 \times 10^{-18} \text{ J} - 0.278 \times 10^{-18} = W$$

$$W = 0.39 \times 10^{-18} = 3.9 \times 10^{-19} \text{ Joule}$$

18. (4)
- $n^2, \frac{1}{n^2}, \quad r \propto \frac{n^2}{Z}, \quad E \propto -\frac{Z^2}{n^2}$

19. (4)
- $n_i < n_b$

20. (2)
- $mvr = \frac{h}{\pi}$
- For
- $n = 2, \quad mvr = \frac{2h}{2\pi}, \quad mvr = \frac{h}{\pi}$

21. (3) Heisenberg's uncertainty principle.

A/q to this principle, it states that it is impossible to find out position and velocity of electrons simultaneously.

22. (2) Total no. of orbitals in a subshell =
- $2l + 1$

$$\text{For } l = 5, \text{ Total no. of orbitals} = 2 \times 5 + 1 = 11$$

23. (2) Principal Q. Number
- \rightarrow
- Size of shell

Azimuthal Q. Number \rightarrow Shape of subshell

Magnetic Q. Number \rightarrow Orientation of orbital

Spin Q. Number \rightarrow Spin of electron

24. (1)

25. (3) Half filled 4s orbital and fully filled 3d orbital
-
- $\text{Cu}_{29} = 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$

26. (3)

27. (3) Different number of angular nodes but same number of total nodes.

$3s = \text{Radial node} = n - l - 1$	$\text{Angular node} = (l)$	Total node
$= 3 - 0 - 1 = 2$	0	$2 + 0 = 2$

$3p \quad 3 - 1 - 1 = 1$	1	$1 + 1 = 2$
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28. (1)

29. (3)
- $\lambda = 4000 \times 10^{-12} \text{ m}$
- $4 \times 10^3 \times 10^{-12} \text{ m}$
-
- $4 \times 10^{-9} \text{ m} = 4 \text{ nm}$

$$E = \frac{1240}{\lambda / \text{nm}} \text{ eV/atom}$$

$$1 \text{ Joule} = \frac{n!240 \times 1.6 \times 10^{-19} \text{ J}}{4}$$

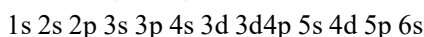
$$= \frac{1 \times 4}{1240 \times 1.6 \times 10^{-19}} = n = \frac{4 \times 10^{19}}{1984} = n$$

$$n = 2.0 \times 10^{16}$$

30. (2) Hund's rule of maximum multiplicity

31. (3) Can't be negative

32. (2) 4d > 4f Increasing energy



33. (2) Use the concept of (n + l) rule

34. (4) KE = q × V₀

$$\frac{1}{2}mv^2 = 1.6 \times 10^{-19} \times 0.35$$

$$v^2 = \frac{2 \times 1.6 \times 10^{-19} \times 0.35}{9.1 \times 10^{-31}} = 0.12 \times 10^{-19+31}$$

$$v = \sqrt{0.12 \times 10^{12}} \quad v = 3.5 \times 10^5$$

35. (2) Ist H-atom IInd H-atom

$$Z = 1 \quad Z = 1$$

$$n = 5 \quad n = 1$$

$$\frac{E_1}{E_2} = \left(\frac{Z_1}{Z_2}\right)^2 \times \left(\frac{n_2}{n_1}\right)^2 \quad \frac{E_{HI}}{E_{HII}} = \left(\frac{1}{1}\right)^2 \times \frac{1}{25}$$

$$\frac{E_I}{E_{II}} = \frac{1}{25} \text{ or } \frac{1}{(5)^2}$$

36. (3) Longest wavelength means energy, frequency and wave number is lowest.

$$n_2 = n_1 + 1$$

For Balmer series : $n_1 = 2 \Rightarrow n_2 = 2 + 1 = 3$

$$\Delta E = 13.6 \times (2)^2 \times \left[\frac{1}{(2)^2} - \frac{1}{(3)^2} \right] \text{ eV/atom}$$

37. (2) 15 line

Maximum no.

$$\text{of emission lines} = \frac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$$

$$= \frac{(6-1)(6-1+1)}{2} = \frac{5 \times 6}{2} = \frac{30}{2} = 15 \text{ lines}$$

38. (3) $E = -2.18 \times 10^{-18} \times \frac{Z^2}{n^2} \text{ J}$

$$\Delta E = 2.18 \times 10^{-18} \times \left[\frac{1}{1} - \frac{1}{\infty^2} \right] \text{ J}$$

$$\Delta E = 2.18 \times 10^{-18} \text{ Joule}$$

For the $\text{He}^+ \rightarrow \text{He}^{2+} + 1e^-$

$$\Delta E = 2.18 \times 10^{-18} \times \frac{4}{1} \text{ T} \quad \Delta E = 8.72 \times 10^{-18} \text{ Joule}$$

39. (1)

40. (4) Take the reference of solution 12.

41. (2) All d-orbitals do not have similar shape.

$$d_{x^2-y^2}, d_{xy}, d_{yz}, d_{xz} \Rightarrow \text{Clover leaf like shape.}$$

$$d_{z^2} \Rightarrow \text{Doubnut shape}$$

42. (4) i) Longest wavelength in Balmer series = 3 → 2

ii) Shortest energy of Lyman series = 2 → 1

iii) Longest frequency of Balmer series = 0 → 2

iv) 2nd line of Lyman series means = 3 → 1

43. (1) $\Delta x((m.\Delta V) \geq \frac{h}{4\pi} \quad 1.46 \times 1-33 \text{ J}$

$$\Delta x = \frac{6.62 \times 10^{-34}}{40 \times 45 \times 2 \times 10^{-3}} = \frac{6.6 \times 10^{-34}}{3600 \times 10^{-1}}$$

$$= \frac{6.6 \times 10^{-34}}{3.6} \approx 1.84 \times 10^{-34} \text{ m}$$

44. (3) (III) and (IV)

45. (4)

46. (4) 25

$$\text{Total no. of orbitals in a shell} = (n)^2 = (5)^2 = 25$$

47. (1) Zeeman effect ⇒ Splitting of spectral lines in the presence of magnetic field.

Stark effect ⇒ Splitting of spectral lines in the presence of electric field.

48. (3) $\bar{\nu} = 1.096 \times 10^7 \times (1)^2 \left[\frac{1}{1} - \frac{1}{\infty^2} \right]$

$$\lambda = \frac{10^{-7}}{1.0967} = 0.911 \times 10^{-7}$$

49. (3) $V = 3d^3 4s^2 \quad \mu = \sqrt{3(3+2)} = \sqrt{15} = 3.87$

$$\text{Cr} = 3d^5 4s^1 \quad \mu = \sqrt{6(6+2)} = \sqrt{48} = 6.92$$

$$\text{Mn} = 3d^5 4s^2 \quad \mu = \sqrt{5(5+2)} = \sqrt{35} = 5.9$$

50. (3) A) Orbital angular momentum = $\sqrt{l(l+1)} \frac{h}{2\pi}$

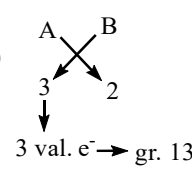
B) Angular momentum $mvr = \frac{nh}{2\pi}$

C) Spin Angular momentum SAM = $\sqrt{s(s+1)} \frac{h}{2\pi}$

D) Magnetic momentum = $\sqrt{n(n+2)} \text{ BM}$

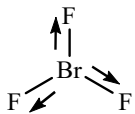
CLASSIFICATION OF ELEMENTS & PERIODICITY IN PROPERTIES (SOLUTION)

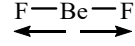
01. (1) According to IUPAC : 1 \Rightarrow Un, 8 \Rightarrow oct.
118 \Rightarrow ununoctium
02. (2) $\text{Fe}^{+2} \Rightarrow 26 - 2 = 24e^-$
 $\text{Mn}^{+2} \Rightarrow 25 - 2 = 23 e^-$ Not iso-electronic
03. (4) C \Rightarrow 12, N \Rightarrow 14, O \Rightarrow 16
Dobernier triad $\Rightarrow \frac{12+16}{2} = \frac{28}{2} = 14$
But, C, N and O does not show similar properties, so it is not dobernier triad.
04. (3) Eka aluminium is Ga.
05. (1) $\text{Ca}^{+2} < \text{K}^+ < \text{Ar}$
Ar \Rightarrow noble gas (High radii in period)
Ionic raddii of cation $\propto \frac{1}{\text{Cation charge}}$
06. (1) 108 \Rightarrow Un + nil oct + ium \Rightarrow Unniloctium
07. (4) Electronic configuration of P is $[\text{Ne}] 3s^2 3p^3$
08. (2) Ionic radius $\propto \frac{1}{\text{Atomic no.}}$ (in isoelectronic sp.)
09. (3) $\text{CN}^- \Rightarrow 6 + 7 + 1 \Rightarrow 14e^-$ [Isoelectronic]
(same no. of e^-)
 $\text{CO} \Rightarrow 6 + 8 \Rightarrow 14e^-$
10. (1) $\text{CO} \Rightarrow 6 + 8 \Rightarrow 14e^-$, $\text{CN}^- \Rightarrow 6+7+1 \Rightarrow 14e^-$
 $\text{NO}^+ \Rightarrow 7+8-1 \Rightarrow 14e^-$ $\text{C}_2^{2-} \Rightarrow 6+6+2 \Rightarrow 14e^-$
11. (3) Al has vaccant d-orbitals, so it can expand its valency to 6.
12. (1) $\text{As}_2\text{O}_3 \Rightarrow$ Amphoteric oxide
 $\text{N}_2\text{O} \Rightarrow$ Acidic
 $\text{Na}_2\text{O} \Rightarrow$ Basic
 $\text{Cl}_2\text{O}_7 \Rightarrow$ Acidic
13. (3) Period \Rightarrow no. of shell \Rightarrow principle quantum no.
14. (4) Down the group metallic character increases and left to right metallic character decreases
15. (3) Left to right in period \Rightarrow non-metallic character increases.
16. (2) Oxidising property in period increases from left to right and decreases down the group.
17. (3) O \longrightarrow O^+
 $2s^2 2p^4$ $2s^2 2p^3$ (Half filled)
B \longrightarrow B^+
 $2s^2 2p^1$ $2s^2$ (Full filled)
18. (3) 105 \Rightarrow Dubnium (Db)
19. (2) 4th period \Rightarrow includes s, p, d orbitals which are fille with e^- , $s + p + d \Rightarrow 18e^-$
 $2 + 6 + 10$
20. (2) I.E. increases in period from left to right
Be $\Rightarrow 1s^2 2s^2$ (full filled) \Rightarrow High I.E. compared to B.
21. (3) $\text{N}^{3-} \Rightarrow 7 + 3 \Rightarrow 10e^-$ [Same no. of e^-]
 $\text{Na}^+ \Rightarrow 11 - 1 \Rightarrow 10e^-$ [Same no. of e^-]
22. (3) Gr. (15) \Rightarrow General configuration of valence shell is $ns^2 np^3 \Rightarrow 3 + 2 \Rightarrow 5$ valence e^-
23. (3) Fact based question.
24. (1) Z = 114
Electronic configuration $\Rightarrow [\text{Rn}] 5f^{14} 6d^{10} 7s^2 7p^2$
No. of valence e^- in valence shell = $2 + 2 = 4e^-$
General valence e^- configuration of gr.(14) $\Rightarrow ns^2 np^2$
25. (1) CO \Rightarrow neutral
 $\text{BaO} \Rightarrow$ (Metal oxide) \Rightarrow Basic
 $\text{Al}_2\text{O}_3 \Rightarrow$ (Amphoteric)
 $\text{Cl}_2\text{O}_7 \Rightarrow$ (Non-metal oxide) \Rightarrow Acidic
26. (1) Z = 29 (cu) \Rightarrow 4th period \Rightarrow 3d series.
27. (2) General d-block electronic configuration is
 $(n-1)d^{1-10} ns^{0-2}$
28. (4) Higher the +ve charge smaller is ionic raddii.
29. (1) Higher the Z_{eff} \Rightarrow Higher attraction on e^- and size decreases.
30. (2) Na $\Rightarrow 1s^2 2s^2 2p^6 3s^1 \Rightarrow$ Na is alkali metal easily donates e^- therefore lowest I.E.
31. (3) F is more oxidising than Cl, as down the group oxidising property decreases.
32. (3) Cl belongs to 3rd period \Rightarrow have vaccant d-orbital, so can easily accomodate e^- and releases more energy.

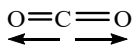
33. (4) $[\text{AlCl}(\text{H}_2\text{O})_5]^{+2}$
 $x + (-1) + 5(0) = +2$
 $x - 1 = +2 \Rightarrow x = +2 + 1 = +3$
34. (3) e^- gain enthalpy order in halogens is
 $\text{Cl} > \text{F} > \text{Br} > \text{I}$
35. (1) Due to small size of F atom \Rightarrow $l.p - l.p$.
 Repulsion between two F atoms is high and bond becomes weak and actual order is
 $\text{Cl}-\text{Cl} > \text{Br}-\text{Br} > \text{F}-\text{F} > \text{I}-\text{I}$
36. (3) Noble gases have high I.E due to stable electronic configuration.
37. (4) I.E_1 of N $>$ I.E_1 of O.
 (Half filled configuration)
 and across the period Z_{eff} increases.
38. (3) Be $\Rightarrow 1s^2 2s^2$ (Full filled configuration), so
 I.E_1 of Be $>$ I.E_1 of B
 and $2s \Rightarrow 2 + 0 \Rightarrow 2$ ($n + l$ rule)
 $2p \Rightarrow 2 + 1 \Rightarrow 3$
 Higher the $(n + l)$ value higher the energy.
39. (4) Radius of noble gas is highest in period due to van der waals force of attraction.
 He \Rightarrow small size noble gas \Rightarrow highest I.E
40. (1) Down the group no. of shells increases interelectronic repulsion increases and therefore shielding increases.
41. (3) Ionic radius $\propto \frac{1}{\text{Atomic no.}}$ (For isoelectronic sp)
 For anion \Rightarrow Higher the charge \Rightarrow Higher the radius
42. (2) Silver (Ag) and gold (Au) have nearly same atomic size due to lanthanoid contraction which is seen in 6th period.
43. (1) In isoelectronic species \Rightarrow for cation
 Higher the charge of cation \Rightarrow smaller will be size.
44. (1) e^- gain enthalpy and I.E have same value but have opposite sign.
45. (1) 
 3 val. $e^- \rightarrow$ gr. 13
46. (1) Higher the charge of cation \Rightarrow smaller will be the size of ion.
47. (1) He $\Rightarrow 1s^2$ (small size)
48. (3) d-block elements are not radioactive.
49. (4) Electron affinity \uparrow ses from left to right in period.
50. (2) NO \Rightarrow Neutral
 $\text{As}_2\text{O}_3 \Rightarrow$ Amphoteric
 $\text{Cl}_2\text{O}_7 \Rightarrow$ Acidic (Non-metal oxide)
 $\text{Na}_2\text{O} \Rightarrow$ Basic (Metal oxide)

MOLECULAR STRUCTURE & CHEMICAL BONDING (SOLUTION)

01. (1) $\text{PCl}_5 \rightarrow Z = \text{no. of } \sigma\text{-bonds} + l.p = 5 + 0 \rightarrow 5$
Hybridisation $\rightarrow sp^3d$

02. (3) $\text{BF}_3 \rightarrow$  $\mu_{net} = 0$

$\text{BeF}_2 \rightarrow$  $\mu_{net} = 0$

$\text{CO}_2 \rightarrow$  $\mu_{net} = 0$

03. (1) Bond length depends on atomic size of bonded atom.

04. (4) Bond length $\propto \frac{1}{\text{Bond order}}$

Bond order of $\text{O}_2 \rightarrow 2$, $\text{O}_3 \rightarrow 1.5$, $\text{O}_2^{2-} \rightarrow 1$

05. (4) $\text{O}_2^+ \rightarrow (2 \times 8) - 1 = 15e^-$ (odd e^-) \rightarrow Paramagnetic

06. (1) $\text{BF}_3 \rightarrow Z = \text{no. of } \sigma\text{-bonds} + l.p.$
 $= 3 + 0 \rightarrow 3 \rightarrow sp^2$

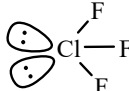
07. (4) $\text{NO}_2^+ \rightarrow \text{O} = \text{N}^+ = \text{O}$ (Hybridisation $\rightarrow sp$)
(Geometry \rightarrow linear)

08. (1)

In octahedral geometry three 180° angles are seen.

09. (1) If $l.p$ arranged on axial position $\rightarrow 6$ $l.p$ -B.P repulsion.

If $l.p$ arranged on equatorial position $\rightarrow 4$ $l.p$ -B.p repulsion.

10. (3) 

(Bent T-shape) due to $l.p$ -B.p repulsion.

11. (3) Orbitals along different axis can not overlap.

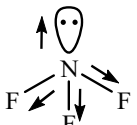
$P_x - P_x \rightarrow \checkmark$

$P_y - P_y \rightarrow \checkmark$

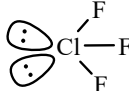
$P_z - P_z \rightarrow \checkmark$

$P_x - P_y \rightarrow \times$

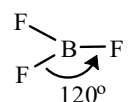
12. (3) 1 Debye = 3.33×10^{-30} coulomb metre.

13. (4)  $\mu_{net} \neq 0$ (Polar molecule)

14. (4) H-bonding is possible with more electronegative atoms like F, O, N

15. (2)  2 $l.p$ and 3-B.P.

16. (3) $\text{BF}_3 \rightarrow$ Hybridisation $(z) = 3 + 0 \Rightarrow 3 \Rightarrow sp^2$

Geometry \Rightarrow Trigonal planar  120°

17. (2) $\text{NF}_3 \Rightarrow Z = 3 + 1 \Rightarrow 4 \Rightarrow sp^3$

$\text{H}_2\text{O} \Rightarrow Z = 2 + 2 \Rightarrow 4 \Rightarrow sp^3$


($Z = \text{no. of } \sigma\text{ bonds} + l.p$)

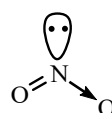
18. (4) $\text{NO}_2^+ \Rightarrow sp \Rightarrow$ Bond Angle $\Rightarrow 180^\circ$

$\text{NO}_2^- \Rightarrow sp^2 \Rightarrow$ Bond angle $\Rightarrow 115^\circ$

(but due to lone pair bond angle is least)

$\text{NO}_2 \Rightarrow$ odd e^- species \Rightarrow bond angle $\Rightarrow 134^\circ$

19. (3) $\text{BF}_3 \Rightarrow Z = 3 + 0 \Rightarrow 3 \Rightarrow sp^2 \Rightarrow$ 

$\text{NO}_2^- \Rightarrow Z = 2 + 1 \Rightarrow sp^2 \Rightarrow$ 

20. (1) $\text{N}_2^+ \Rightarrow \text{B.O.} = \text{Total no. of } e^- - (+ve \text{ charge})$

$= 14 - 1 = 13 e^- \Rightarrow \text{B.O.} \Rightarrow 2.5$

$\text{N}_2^- \Rightarrow \text{B.O.} = 14 + 1 \Rightarrow 15e^- \Rightarrow 2.5$

21. (1) $\text{O}_2 \Rightarrow 16 e^- \Rightarrow$ Even no. of e^- , still O_2 is paramagnetic according to MOT.

$\text{O}_2 \Rightarrow 16e^- \Rightarrow \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2$

$\pi 2p_x^2 = \pi 2p_y^2, \pi^* 2p_x^1 = \pi^* 2p_y^1$

(Unpaired $e^- \Rightarrow$ Paramagnetic)

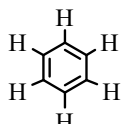
22. (3) $\text{NO} \Rightarrow 7 + 8 \Rightarrow 15e^- \Rightarrow \text{B.O.} \Rightarrow 2.5$

Odd $e^- \Rightarrow$ Paramagnetic

$\text{NO}^+ \Rightarrow 7 + 8 - 1 \Rightarrow 14e^- \Rightarrow \text{B.O.} \Rightarrow 3$

even $e^- \Rightarrow$ Diamagnetic

23. (4) $\text{NO}^+ \Rightarrow 7 + 8 - 1 = 14e^- \Rightarrow \text{B.O.} \Rightarrow 3$

24. (4) $\text{C}_6\text{H}_6 \Rightarrow$ 

25. (4) $\text{O}_2^- \Rightarrow 16 + 1 \Rightarrow 17e^- \Rightarrow$ odd $e^- \Rightarrow$ paramagnetic

26. (1) $\text{Bond length} \propto \frac{1}{\text{B.O.}}$

$\text{O}_2^+ \Rightarrow 16 - 1 \Rightarrow 15e^- \Rightarrow \text{B.O.} \Rightarrow 2.5$

$\text{O}_2^- \Rightarrow 16 + 1 \Rightarrow 17e^- \Rightarrow \text{B.O.} \Rightarrow 1.5$

$\text{O}_2^{2-} \Rightarrow 16 + 2 \Rightarrow 18e^- \Rightarrow \text{B.O.} \Rightarrow 1$

$\text{O}_2 \Rightarrow 16e^- \Rightarrow \text{B.O.} \Rightarrow 2$

27. (2) Refer Ans. no. 26

28. (1) $O_2^{2-} \Rightarrow 16 + 2 \Rightarrow 18e^- \Rightarrow B.O. \Rightarrow 1$

$B_2 \Rightarrow 10 e^- \Rightarrow B.O. \Rightarrow 1$

29. (3) $CO \Rightarrow 6 + 8 \Rightarrow 14e^- \Rightarrow$ (even e^- species)
 \Rightarrow Diamagnetic.

30. (3) $CO \Rightarrow 6 + 8 \Rightarrow 14e^- \Rightarrow B.O. \Rightarrow 3$

For $CO_3^{2-} \Rightarrow B.O. = 1.33$

For $CO_2 \Rightarrow B.O. = 2$

$$\boxed{\text{Bond length} \propto \frac{1}{B.O.}}$$

31. (2) $Be_2 \Rightarrow \sigma 1s^2 \sigma^* 1s^2$

$B.O. = \frac{2-2}{2} = 0$ (Molecule cannot exist if $B.O. = 0$)

32. (1) $N_2 \Rightarrow 14e^- \Rightarrow B.O. \Rightarrow 3$

$N_2^+ \Rightarrow 14 - 1 \Rightarrow 13e^- \Rightarrow B.O. \Rightarrow 2.5$

$$\boxed{\text{Bond energy} \propto B.O.}$$

33. (2) $N_2 \Rightarrow \cdot\ddot{N}=\ddot{O}:$ (Octet is incomplete)

34. (4) There is no equilibrium between canonical form's.

35. (4) As per MOT, antibonding molecular orbitals have more energy.

36. (1) $sp^2 \Rightarrow$ Trigonal planar

$sp \Rightarrow$ linear geometry

$sp^3 \Rightarrow$ Tetrahedral geometry

37. (2) $SF_4 \Rightarrow Z = \sigma\text{-bonds} + l.p = 4 + 1 \Rightarrow 5 \Rightarrow sp^3d$

$IF_5 \Rightarrow Z = 5 + 1 \Rightarrow 6 \Rightarrow sp^3d^2$

$NO_2^+ \Rightarrow Z = 2 + 0 \Rightarrow sp$

$NH_4^+ \Rightarrow Z = 4 + 0 \Rightarrow sp^3$

38. (1) $BH_3 \Rightarrow 3 B.P \Rightarrow \sigma e^-$ (Incomplete octet)

39. (2) $He_2 \Rightarrow \sigma 1s^2 \sigma^* 1s^2 \Rightarrow B.O. = \frac{2-2}{2} = 0$
 ($4e^-$)

$Li_2 \Rightarrow \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \Rightarrow B.O. = \frac{4-2}{2} = 1$
 ($6e^-$)

$C_2 \Rightarrow \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \pi 2p_x^2 = \pi 2p_y^2$
 ($12e^-$)

$B.O. \Rightarrow \frac{8-4}{2} = \frac{4}{2} = 2$

$O_2^+ \Rightarrow \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2p_z^2$
 ($15e^-$)

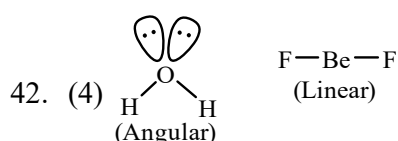
$\pi 2p_x^2 = \pi 2p_y^2 \quad \pi^* 2p_x^1 = \pi^* 2p_y^1$

$B.O. = \frac{10-5}{2} = \frac{5}{2} = 2.5$

40. (3) H_2O_2 , HCN and acetic acid contains intermolecular H-bonding, in cellulose.

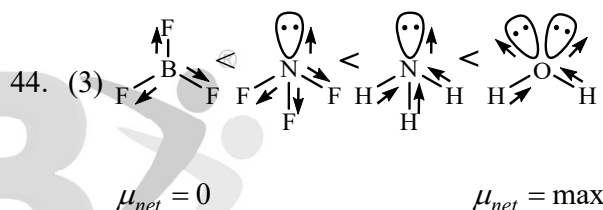
Intramolecular H-bonding is present.

41. (4) Dipole-dipole interaction is present in methanol molecules.



43. (3) $Z = 3 + 0 \Rightarrow 3 \Rightarrow sp^2$

Geometry \Rightarrow Trigonal planar $B.\text{angle} \Rightarrow 120^\circ$



45. (1) $He_2 \Rightarrow 4e^- \Rightarrow \sigma 1s^2 \sigma^* 1s^2$
 $\Rightarrow B.O. = \frac{2-2}{2} = 0$

46. (2) $\boxed{\text{Lattice enthalpy} \propto \frac{\text{Charge}}{\text{Size}}}$

47. (2) In BF_3 , B-F bonds are polar but molecule is non-polar.

48. (4) CCl_4 is non-polar and for $\mu_{net} \neq 0$

\Rightarrow Polar molecule.

49. (1) In $NH_3 \Rightarrow Z = 3 + 1 \Rightarrow sp^3 \Rightarrow l.p\text{-}B.p$ repulsion decreases bond angle.

50. (1) Refer ans. 49.

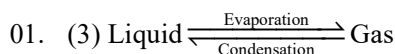
THERMODYNAMICS (SOLUTION)

01. (2) Definition of thermodynamics.
02. (4) Rate of reaction related to kinetics of reaction.
03. (2) In insulated vessel there is no exchange of heat and matter by system.
04. (3) Heat is a path function.
05. (1) Energy is a path function.
06. (3) For isolated system $\Delta U = 0$, hence energy remains constant.
07. (2) For adiabatic process $q = 0$
- $$\therefore \Delta U = q + w \Rightarrow \boxed{\Delta U = q}$$
08. (4) Characteristics of reversible process.
09. (3) For isothermal process,
- $$\Delta U = 0 \quad \therefore \quad q = -w$$
- For reversible process
- $$w_{\max} = -2.303nRT \log \frac{V_f}{V_i}$$
- $$\therefore \quad \boxed{q = 2.303nRT \log \frac{V_f}{V_i}}$$
10. (1) Work during free expansion of a gas is zero. The expansion of gas in vacuum ($P_{\text{ex}} = 0$) is called free expansion.
11. (4) At constant volume, $P\Delta V = 0 \quad \therefore \quad \Delta V = 0$
- $$\therefore \Delta U = q + w = q - P\Delta V$$
- $$\therefore \quad \boxed{\Delta U = 50 \text{ kJ}}$$
12. (3) Vapourisation of water at 100°C and 1 bar
- $$\text{H}_2\text{O}(l) \longrightarrow \text{H}_2\text{O}(g) \quad \Delta H_{\text{vap}} = 41 \text{ kJ mol}^{-1}$$
- $$\therefore \Delta H = \Delta U + RT\Delta n \quad \therefore \quad \Delta U = \Delta H - RT$$
- $$\Rightarrow 41 \text{ kJ} - (8.314 \times 373) \text{ J}$$
- $$\Rightarrow 41 \text{ kJ} - 3 \text{ kJ (app.)} \quad \Rightarrow \quad \boxed{\Delta U = 38 \text{ kJ}}$$
13. (2) Molar heat capacity is constant for 1 mole of a given system and does not depend upon amount of matter but depends upon condition of temp. pressure and volume.
14. (4) $q = c \times \Delta T \times m$
- Where, q = heat, c = heat capacity / specific heat
- $$\Delta T = \text{change in temperature}$$
- $$m = \text{mass}$$
15. (3) Find out the correct equation (correct instead of incorrect)
- $$\Delta H = \Delta U + \Delta nRT \quad \therefore \quad \Delta U = \Delta H - \Delta nRT$$
16. (4) Condition of standard enthalpy of reaction, $T = 298 \text{ K}$ and $P = 1 \text{ bar}$
17. (3) $\Delta H_{\text{fus}} = 6.00 \text{ kJ mol}^{-1}$
- $$\therefore n_{\text{H}_2\text{O}} = \frac{90}{18} = 5$$
- $$\therefore \Delta H_{\text{fus}} \text{ for } 90 \text{ gm (5 moles) of water is}$$
- $$\Delta H = 6.00 \text{ kJ} \times 5 = 30 \text{ kJ}$$
18. (1) $\text{H}_2\text{O}(s) \longrightarrow \text{H}_2\text{O}(l) \quad \Delta n = 0$
- 1 mole 1 mole
- (No gas involved $\therefore \Delta n = 0$)
- $$\therefore \Delta H = \Delta U + \Delta nRT \Rightarrow \Delta H = \Delta U + 0$$
- $$\Delta H = \Delta U \text{ but } \Delta H = 6000 \text{ J} = 6 \text{ kJ}$$
- $$\therefore \Delta U = 6 \text{ kJ}$$
19. (1) The enthalpy change of a reaction where 1 mole of a molecule is formed from its constituent elements (atoms) at standard condition is called standard enthalpy of formation.
20. (4) According to Hess's law
- $$\Delta H_1 = \Delta H_2 + \Delta H_3 + \Delta H_4$$
21. (2) a) $\text{C}_6\text{H}_6 + \frac{15}{2}\text{O}_2 \rightarrow 6\text{CO}_2 + 3\text{H}_2\text{O}$
- $$\Delta H_1 = -3267 \text{ kJ}$$
- b) $\text{C} + \text{O}_2 \rightarrow \text{CO}_2 \quad \Delta H_2 = -393.5 \text{ kJ}$
- c) $\text{H}_2 + 1/2\text{O}_2 \rightarrow \text{H}_2\text{O} \quad \Delta H_3 = -285.83 \text{ kJ}$
- d) $6\text{C} + 3\text{H}_2 \rightarrow \text{C}_6\text{H}_6 \quad \Delta H = ?$
- By using Hess law
- $$6\text{C} + 6\text{O}_2 \longrightarrow 6\text{CO}_2$$
- $$3\text{H}_2 + \frac{3}{2}\text{O}_2 \longrightarrow 3\text{H}_2\text{O}$$
- $$\text{C}_6\text{H}_6 - \frac{15}{2}\text{O}_2 \longrightarrow 6\text{CO}_2 - 3\text{H}_2\text{O}$$
-
- $$6\text{C} + 3\text{H}_2 \longrightarrow \text{C}_6\text{H}_6$$
- $$\therefore \Delta H = (6 \times \Delta H_2) + (3 \times \Delta H_3) - (\Delta H_1)$$
- $$= (6 \times -393.5) + (3 \times -285.83) - (-3267)$$
- $$= -2361 - 857.49 + 3267 = +48.51 \text{ kJ}$$

22. (3) Definition of spontaneous process.
23. (2) Entropy change is very high when system converts into gaseous phase from solid phase in a process.
24. (4) According to 2nd law of thermodynamics, the entropy increases in a spontaneous process.
i.e. $\Delta S_T = \Delta S_{\text{sys}} + \Delta S_{\text{sur}} > 0$ (+ve)
25. (1) For reversible process, entropy change in system is $\Delta S = \frac{q}{T}$.
26. (3) For endothermic spontaneous process,
 $\Delta H < T\Delta S$
 $\therefore \Delta G = \Delta H - T\Delta S$ If $\Delta H < T\Delta S$ then $\Delta G = -ve$
27. (2) Spontaneous process with no. cross over temperature in exothermic process with positive entropy change in system i.e.
 $\Delta H = -ve$ and $\Delta S = +ve$
28. (3) We have, $\Delta G = \Delta H - T\Delta S$
at eqm. $\Delta G = 0$ (Cross over temp. means eqm temp.)
 $\therefore T\Delta S = \Delta H \quad \therefore T = \frac{\Delta H}{\Delta S}$
29. (4) 3rd law of thermodynamics state that system has zero entropy at absolute zero temperature.
30. (2) We have, $\Delta G = \Delta G^\circ + 2.303RT \log K$
at eqm $\Delta G = 0 \therefore \Delta G^\circ = -2.303RT \log K$
31. (4) All are the condition of an endothermic reaction to be spontaneous.
32. (3) Atomisation of methane,
$$\begin{array}{c} \text{H} \\ | \\ \text{H}-\text{C}-\text{H} \longrightarrow 4\text{H} + \text{C} \\ | \\ \text{H} \end{array} \quad \begin{array}{c} \text{(g)} \\ \text{(g)} \end{array} \quad \begin{array}{c} \text{(g)} \\ \text{(g)} \end{array}$$

(4C-H) $\Delta H = 1665 \text{ kJ}$
 $\therefore 4 \times \Delta_b H_{\text{CH}} = 1665 \text{ kJ}$
 $\therefore \Delta_b H_{\text{CH}} = \frac{1665}{4} = \boxed{416 \text{ kJ}}$
33. (3) Heat constant pressure (isobaric process)
 $\boxed{q_p = \Delta U + P\Delta V}$
34. (1) For given reaction, $\Delta n = 0 \therefore \Delta H = \Delta U$
35. (1) $\Delta n = n_2 - n_1 = 0 - 1 = -1$
36. (4) Entropy is the measure of randomness in system.
37. (3) According to law of thermodynamics.
1) Isothermal process = $q = -w$
2) Adiabatic process = $\Delta U = w$
3) Isochoric process = $\Delta U = q$
4) Isobaric process = $q = \Delta U + P\Delta V$
38. (4) Free expansion $\Rightarrow P_{\text{ex}} = 0 \Rightarrow w = 0$
Reversible process $\Rightarrow W_{\text{max}} = \text{maximum work}$
For compression $\Rightarrow \Delta V = -ve \therefore W = -P\Delta V$ (+ve)
39. (2) At eqm the temperature, $T = \frac{\Delta H}{\Delta S}$
40. (3) As the product is form from its constituent elements/atoms.
41. (1) Spontaneous process may exothermic or endothermic depending upon the value of " ΔS " and " T ".
42. (2) For isolated system $\Delta U = 0$
For isothermal process $\Delta T = 0$
For spontaneous process $\Delta G = -ve$
43. (1) Statement is of 2nd law of thermodynamics.
44. (1) For free expansion : $P_{\text{ex}} = 0$
 $\therefore \text{work} = -P_{\text{ex}}\Delta V = 0$
45. (2) During melting of ice, entropy increases hence it is spontaneous.
46. (1) Combustion of hydrogen gives water,
 $\text{H}_2 + 1/2\text{O}_2 \rightarrow \text{H}_2\text{O}$.
47. (3) For Habers process
 $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3 \quad \Delta n = -2$
and $\Delta H = \Delta U - 2RT$
 $\therefore \Delta U > \Delta H$
48. (4) i) Temperature is increases
ii) C_V is molar heat capacity.
49. (1) Intensive property is the ratio of two extensive properties.
50. (2) For spontaneous process,
 $\Delta S_T > 0$ (+ve) $\Rightarrow \Delta G < 0$ (-ve)

EQUILIBRIUM (SOLUTION)



at equilibrium rate of evaporation = rate of condensation

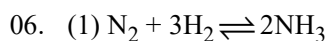
02. (3) Melting point or freezing point of a substance is the temperature at which the liquid phase and solid phase coexist.

03. (2) Higher the volatility lesser will be in the intermolecular forces and hence lesser will be the boiling point.

∴ Water will have greater boiling point than methyl alcohol

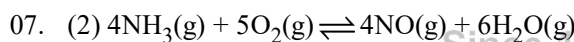
04. (1) Equilibrium can only be obtained when carried out in closed container.

05. (3) Solubility of CO₂ in water is exothermic change hence increasing temperature will decrease solubility.



[NH₃] will increase with respect to time as it is a product while that of [N₂] and [H₂] will decrease with respect to time

Assuming the reactants were taken in the ratio of their stoichiometric w-efficient the $n_{N_2} < n_{H_2}$ as $n_{N_2} : n_{H_2} = 1 : 3$ so initially [H₂] > [N₂]



$$K_c = \frac{[NO]^4 [H_2O]^6}{[NH_3]^4 [O_2]^5} = [NO]^4 [H_2O]^6 [NH_3]^{-4} [O_2]^{-5} \dots (1)$$

$$\text{Given } K_c = [NO]^a [H_2O]^b [NH_3]^c [O_2]^d \rightarrow \dots (2)$$

comparing (1) and (2)

$$a = 4 \quad b = 6 \quad c = -4 \quad d = -5$$

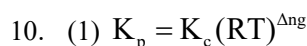
08. (3) An equilibrium reaction is homogenous when all the reactants and products are in same phase.

09. (4) When a factor is multiplied to an equilibrium then new equilibrium constant = (old equilibrium constant) multiplied factor

$$\text{Given multiplication factor} = \frac{1}{3}$$

$$\Rightarrow K_c^1 = (K_c)^{\frac{1}{3}} = \sqrt[3]{K_c}$$

$$\text{New } K_c = \sqrt[3]{(\text{old } K_c)}$$

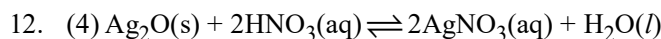


For K_p K_c

$$\Delta ng > 0$$

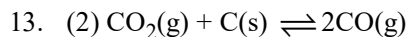
$$c + d > a + b$$

11. (4) [pure solid] or [pure liquids] almost remains constant and hence are replaced by 1 (unity) while writing K_c for a heterogenous equilibrium



$$K_c = \frac{[AgNO_3]^2}{[HNO_3]^2}$$

Changing the [AgNO₃] or [HNO₃] will not have any change in value of K_c as it depends only on temperature.



$$t = 0s \quad 0.33 \text{ bar} \quad - \quad 0$$

$$t = t_{eg} \quad 0.33 - x \quad \quad \quad 2x$$

$$\text{Given } K_p = 3$$

$$K_p = \frac{P_{CO}^2}{0.33 - x} \Rightarrow \frac{(2x)^2}{0.33 - x} = 3$$

$$\Rightarrow \frac{4x^2}{\frac{1}{3} - x} = 3 \Rightarrow \frac{4x^2 \times 3}{1 - 3x} = 3$$

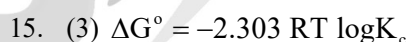
$$\Rightarrow 4x^2 = 1 - 3x$$

$$\Rightarrow 4x^2 + 3x - 1 = 0 \Rightarrow x = \frac{1}{4}, -1$$

x cannot be -1

$$\therefore x = \frac{1}{4} \Rightarrow P_{CO} = 2x = \frac{1}{2} \text{ bar}$$

14. (2) Q_c and K_c have same expression hence both will have same units.

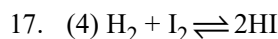


for reaction to be spontaneous $\Delta G^\circ < 0$

$$\Rightarrow \log K_c > 0$$

$$\therefore \log K_c > 1$$

16. (3) Value of equilibrium constant K_c only depends upon temperature, hence by adding product value of K_c will not change.



$$Q_c = \frac{[HI]^2}{[H_2][I_2]}$$

$$K_c = Q_c \text{ at equilibrium}$$

by adding same moles of H₂ of equilibrium will decrease the value of Q_c while K_c still remains the same as it's value only depends on temperature.

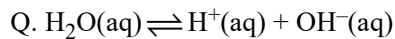
18. (3) As per the given graph there is a sudden jump in [H₂] at equilibrium and there after equilibrium is reestablished. From this we may conduct the diagram represents addition of H₂ at equilibrium

19. (4) Pressure donot have any effect on solid or aqueous phase homogenous equilibrium and also for gaseous phase reaction where total no. of moles on both reactant and product are equal
20. (1) On increasing the pressure of a reaction at equilibrium the reaction proceeds in a direction of decreasing no. of gaseous moles.
 (1) $\text{CO(g)} + 3\text{H}_2\text{(g)} \rightleftharpoons \text{CH}_4\text{(g)} + \text{H}_2\text{O(g)}$
 increasing P will shift the reaction in forwards direction
 (2) $\text{C(s)} + \text{CO}_2\text{(g)} \rightleftharpoons 2\text{CO(g)}$
 increasing P will shift the reaction in backward direction.
21. (3) Addition of inert gas at constant volume donot have any effect on equilibrium i.e this change will not bring any shift because keeping the volume constant the partial pressure or molar concentration donot change. Therefore statement-1 is incorrect and statemen-2 is correct.
22. (2) Correction in question
 $\text{N}_2\text{(g)} + 3\text{H}_2\text{(g)} \rightleftharpoons 2\text{NH}_3\text{(g)} \quad \Delta H = -x\text{kJ}$
- $$\log \frac{K_2}{K_1} = \frac{\Delta H}{2.303R} \left\{ \frac{1}{T_1} - \frac{1}{T_2} \right\}$$
- Since the reaction is exothermic hence increasing the temperature will decrease the equilibrium constant and decreasing the temperature will increase the equilibrium constant.
 no. of gaseous moles of product = 2 moles
 no. of gaseous moles of reactant = 4 moles
 $\therefore \Delta n_g = 2 - 4 = -2$
 \therefore Increasing the pressure will shift the reaction in forward direction and will result in higher yield.
23. (1) Melting of ice
 $\text{H}_2\text{O(s)} \rightleftharpoons \text{H}_2\text{O(l)} \quad \Delta H_r = +ve$
 Ice have hexagonal arrangement of H₂O molecules with void spaces hence increasing the pressure will cause the H-bonding to break there by breaking the hexagonal cage like structure and thus resulting in formation of water
 As melting is an endothermic process hence increasing the temperature will also cause melting of ice.
 \therefore melting of ice is favoured by high pressure and high temperature.
24. (1) The value of K_p or K_c just given us an idea of extent of reaction rather than the rate reaction.
25. (4) Equilibrium established involving a weak electrolyte is known as ionic equilibrium where the equilibrium involves ions in aqueous medium
 $\text{AB(aq)} \rightleftharpoons \text{A}_{(\text{aq})}^+ + \text{B}_{(\text{aq})}^-$
 wea electrolyte
26. (2) P. increasing temperature for exothermic reaction will shift the reactin in backward direction
 P \rightarrow (ii)
 Q. Increasing temperature for endothermic reaction will shift the reaction in forward direction
 Q \rightarrow (i)
 R. Haber's process
 $\text{N}_2\text{(g)} + 3\text{H}_2\text{(g)} \rightleftharpoons 2\text{NH}_3\text{(g)}$
 Increasing the pressure will cause a forward shift
 R \rightarrow (i)
 S. Contact process
 $2\text{SO}_2\text{(g)} + \text{O}_2\text{(g)} \rightleftharpoons 2\text{SO}_3\text{(g)}$
 Increasing the pressure will cause of forward shift.
 S \rightarrow (i)
27. (3) A proton acceptor is a base while a proton donor is an acid according Bronsted-Lowry theory
 $\text{HCl} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{Cl}^-$
 here H₂O is behaving as Bronsted base while in the reaction
 $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$
 H₂O is donating H⁺ \therefore it is behaving as bronsted aci.
 From the above two reaction we can conclude that H₂O can accept as well donate H⁺ ion.
28. (3) $\text{HCO}_3^- \rightarrow \text{H}_2\text{CO}_3$
 HF acid \rightarrow F⁻ conjugae base
 HSO₄⁻ acid \rightarrow SO₄²⁻ conjugae base
 NH₂⁻ Base \rightarrow NH₃ conjugate acid
29. (1) Stronger acids have weaker conjugate base as HClO₄ and HI are stronger acids than H₃O⁺ hence ClO₄⁻ and I⁻ would be weaker base compared to H₂O.
30. (2) Question correction
 Statement-2 : In dissociation of weak acid presence of common ion H⁺ will cause backward shift in the equilibrium reaction.
 Statement-I : When concentration of a strong acid is less than 10⁻⁶ M
 10⁻⁷, 10⁻⁸, 10⁻⁹.....than $[\text{H}^+]_{\text{mix}} = [\text{H}^+]_{\text{strong acid}} + [\text{H}^+]_{\text{H}_2\text{O}}$ is considered.
 \therefore For 10⁻⁸ HCl solution
 $[\text{H}^+]_{\text{solution}} = [\text{H}^+]_{\text{HCl}} + [\text{H}^+]_{\text{H}_2\text{O}} = 10^{-8} + 10^{-7}$
 $= 11 \times 10^{-8} \text{ M}$
 $\therefore \text{pH} = -\log [\text{H}^+]_{\text{solution}} = -\log 11 \times 10^{-8} = 6.958$
 Statement-II :
 $\text{CH}_3\text{COOH(aq)} \rightleftharpoons \text{CH}_3\text{COO}^-\text{(aq)} + \text{H}^+\text{(aq)}$
 weak acid
 Adding of H⁺ or CH₃COO⁻ in the above reaction will cause a shift in the equilibrium reaction in backward direction.

31. (4) P. Ionic product of water

$$K_w = [H^+][OH^-] = 10^{-14}$$

P → (ii)



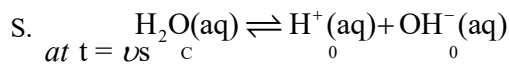
$$K_c = \text{dissociation constant} = \frac{[H^+][OH^-]}{[H_2O]} = \frac{10^{-7} \times 10^{-7}}{1000/18}$$

$$= \frac{10^{-14}}{55.55}$$

Q → (iii)

R. Molar concentration of $H_2O = \frac{1000}{18} = 55.55 \text{ M}$

R → (iv)



at t = teq $C - c\alpha$ $c\alpha$ $c\alpha$

$$[H^+] = c\alpha = 10^{-7}$$

$$\Rightarrow \alpha = \frac{10^{-7}}{C} = \frac{10^{-7}}{55.55}$$

Where C = molar concentration of water

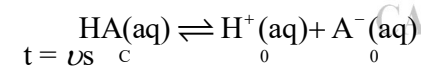
S → (v)

T. $pK_w = -\log K_w = -\log 10^{-14}$

$$pK_w = 14$$

T → (i)

32. (4) Monobasic acid = HA



t = teq $C - c\alpha$ $c\alpha$ $c\alpha$

$$pH = -\log(H^+) = -\log c\alpha = -\log \sqrt{KaC}$$

$$\Rightarrow pH = -\frac{1}{2} \log k_a - \frac{1}{2} \log C$$

$$\Rightarrow pH = \frac{pka}{2} - \frac{\log C}{2}$$

$$\Rightarrow 4.5 = \frac{pka}{2} - \frac{\log 0.1}{2}$$

$$\Rightarrow 9 = pka + 1$$

$$\Rightarrow pka = 8$$

33. (2) We know that

$$K_a \times K_b = K_w = 10^{-14}$$

$$\Rightarrow K_b = \frac{10^{-14}}{K_a} = \frac{10^{-14}}{5.6 \times 10^{-10}}$$

$$= 0.178 \times 10^{-4}$$

$$= 1.78 \times 10^{-5}$$

$$pkb = -\log k_b$$

$$= -\log (1.78 \times 10^{-5})$$

$$= -\log 10^{-5} - \log 1.78$$

$$= 5 - \log 1.78$$

$$\therefore pkb = 5 - \log 1.8$$

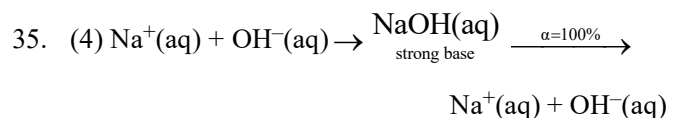
34. (4) For polyprotic acids

$$ka_1 > ka_2 > ka_3 > \dots$$

\therefore statement 1 is true

Removing a positively charged proton charged ion will be difficult

\therefore statement 2 is false



$\therefore Na^+(aq)$ ion will not get hydrolysed but will get hydrated

From the above we may conclude ion from weak electrolyte will get hydrolyzed but that from strong electrolyte will not get hydrolyzed.

$ClO_4^- \rightarrow$ anion from $HClO_4$ a strong acid so will not get hydrolyzed and will get hydrated

$CH_3COO^- \rightarrow$ anion from CH_3COOH a weak acid so will get hydrolyzed

$NO_3^- \rightarrow$ anion from $HClO_4$ a strong acid so will not get hydrolyzed and will get hydrated.

36. (3) (P) $NaNO_3$ salt of $NaOH$ (strong base) + HNO_3 (strong acid)

$$\therefore pH = 7$$

P → (ii)

(Q) CH_3COONa salt of CH_3COOH (weak acid) + $NaOH$ (strong base)

$\Rightarrow CH_3COO^-$ ion will undergo hydrolysis anionic hydrolysis \Rightarrow solution will be alkaline

$$\therefore pH > 7$$

(Q) → (i)

(R) NH_4Cl salt of NH_4OH (weak base) + HCl (strong acid)

$\Rightarrow NH_4^+$ ion will undergo hydrolysis cationic hydrolysis \Rightarrow solution will be acidic

$$\therefore pH < 7$$

(R) \rightarrow (iii)

(S) $\text{CH}_3\text{COONH}_4$ salt of CH_3COOH (weak acid) + NH_4OH (weak base)

$\Rightarrow \text{CH}_3\text{COO}^-$ ion and NH_4^+ ion both will undergo hydrolysis

\therefore pH can have any value

(S) \rightarrow (iv)

37. (4) Acidic buffer can be prepared by mixing a weak acid and salt of this weak acid with strong base.

38. (1) pH = 9.25 \Rightarrow it's a basic buffer

for basic buffer

$$\text{pH} = 14 - \text{pOH}$$

$$\text{where } \text{pOH} = \text{pK}_b + \log \frac{[\text{conjugate acid}]}{[\text{Base}]}$$

\therefore conjugate acid is from salt a strong electrolyte ($\alpha = 100\%$)

\therefore dilution will not have any effect on [conjugate acid] but [base] will change

after 10 time dilution

$$[\text{Base}]^1 = \frac{[\text{Base}]}{10}$$

$$\Rightarrow (\text{pOH})^1 = \text{pK}_b + \log \frac{[\text{Conjugate acid}]}{[\text{Base}]}$$

$$= \text{pK}_b + \log \frac{[\text{Conjugate acid}]}{[\text{Base}]}$$

$$= \text{pK}_b + \log \frac{[\text{Conjugate acid}]}{[\text{Base}]} \times 10$$

$$(\text{pOH}) = \text{pOH} + 1$$

$$\Rightarrow (\text{pH}) = 14 - (\text{pOH})^1 = 14 - (\text{pOH} + 1)$$

$$= 14 - \text{pOH} - 1 = \text{pH} - 1$$

$$= 9.25 - 1$$

$$= 8.25$$

39. (3) For a salt to dissolve in a given solvent (liquid)

$$\Delta_{\text{Lattice energy}}^{\text{H}} < \Delta_{\text{solution}}^{\text{H}}$$

40. (2) $\text{Zr}_3(\text{PO}_4)_4 \longrightarrow 3\text{Zr}^{4+} + 4\text{PO}_4^{3-}$

$$K_{\text{sp}} = [\text{Zr}^{4+}]^3 [\text{PO}_4^{3-}]^4$$

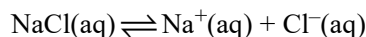
$$= (3S)^3 (4S)^4$$

$$= 27 S^3 \times 256 S^4$$

$$= 6912 S^7$$

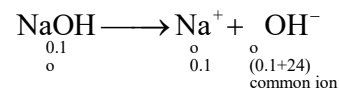
$$\therefore S = \left(\frac{K_{\text{sp}}}{6912} \right)^{\frac{1}{7}}$$

41. (1) For a saturated solution of NaCl the following equilibrium will exist



by passing HCl gas into this solution we are adding Cl^- a common ion which will cause NaCl to precipitate out of the solution.

42. (2) $\text{Ni}(\text{OH})_2 \longrightarrow \text{Ni}^{2+} + 2\text{OH}^-$



$$(K_{\text{ip}}) \text{Ni}(\text{OH})_2 = [\text{Ni}^{2+}] [\text{OH}^-]^2$$

$$\Rightarrow 2 \times 10^{-15} = S (25 + 0.1)^2$$

$$\therefore S \ll \ll 1$$

$$\Rightarrow 25 + 0.1 \approx 0$$

$$\Rightarrow S = - \frac{2 \times 10^{-15}}{(0.1)^2} = 2 \times 10^{-3} \text{ M}$$

43. (1) $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$

$$(1) (\text{ng}) \text{ product} = 2 (\text{ng})_{\text{R}} = 1$$

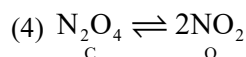
\therefore Decreasing pressure will shift the reaction in forward direction and increasing pressure will shift reaction backward direction.

$$(2) K_p = \frac{P_{\text{NO}_2}^2}{P_{\text{N}_2\text{O}_4}} \Rightarrow \text{Unit of } k_p = \text{atm}$$

$$(3) \text{At } 500 \text{ K } K_p = 107 \times 10^3$$

$$\text{and at } 600 \text{ K } K_p = 107 \times 10^4$$

\Rightarrow Increasing the temperature increases the value of k_p means the reaction is endothermic i.e. $\Delta H_r = +ve$.



$$n \quad C-C \alpha \quad 2C \alpha \quad n_{\text{total}} = C + C \alpha$$

$$P_i \frac{C(1-\alpha)}{C(1+\alpha)} P \quad \frac{2C\alpha}{C(H\alpha)} P$$

$$= \frac{1-\alpha}{1+\alpha} P \quad = \frac{2\alpha}{1+\alpha} P$$

where P = equilibrium pressure

$$K_p = \frac{P_{\text{NO}_2}^2}{P_{\text{N}_2\text{O}_4}} = \frac{\left(\frac{2\alpha}{1+\alpha} P \right)^2}{\frac{(1-\alpha)}{1+\alpha} P}$$

$$\Rightarrow K_p = \frac{4\alpha^2 P^2 (1+\alpha)}{(1+\alpha)^2 (1-\alpha)P} = \frac{4\alpha^2 P}{(1+\alpha)(1-\alpha)} = \frac{4\alpha^2 P}{1-\alpha^2}$$

$$\text{If } P^1 = 2P \quad \alpha^1 = ??$$

$$K_{p^1} = \frac{4(\alpha^1)^2 P^1}{1-(\alpha^1)^2} = \frac{4(\alpha^1)^2 2P}{1-(\alpha^1)^2} = \frac{8(\alpha^1)^2 P}{1-(\alpha^1)^2}$$

$\therefore K_p$ only depends upon temperature

$$\Rightarrow K_p = K_{p^1}$$

$$\Rightarrow \frac{4\alpha^2 P}{1-\alpha^2} = \frac{8(\alpha^1)^2 P}{1-(\alpha^1)^2}$$

assuming α and α^1 to be small as compared to 1

$$\Rightarrow 1-\alpha^2 \approx 1 \text{ and } 1-(\alpha^1)^2 \approx 1$$

$$\Rightarrow 1\alpha^2 = 8(\alpha^1)^2$$

$$\Rightarrow (\alpha^1)^2 = \frac{\alpha^2}{8} \Rightarrow \alpha^1 = \frac{\alpha}{\sqrt{2}}$$

i.e. increasing pressure by 100% cause degree of

dissociation to change by a factor of $\frac{1}{\sqrt{2}}$

44. (3) $\text{CH}_3\text{COONH}_4$ is a salt of CH_3COOH (weak acid) and NH_4OH (weak base)

$$\Rightarrow \text{pH} = \frac{\text{pk}_w}{2} + \frac{\text{pk}_a}{2} - \frac{\text{pk}_b}{2}$$

$$\text{Given } (K_a)\text{CH}_3\text{COOH} = 1.8 \times 10^{-5}$$

$$\Rightarrow \text{pk}_a = -\log K_a = -\log 1.8 \times 10^{-5} = 4.74$$

$$\text{Given } (k_b)\text{NH}_4\text{OH} = 1.8 \times 10^{-5}$$

$$\Rightarrow \text{pk}_b = -\log k_b = -\log 1.8 \times 10^{-5} = 4.74$$

$$\therefore \text{pk}_a = \text{pk}_b$$

$$\Rightarrow \text{pH} = \frac{\text{pk}_w}{2} = \frac{14}{2} = 7$$

45. (4) $\Delta G^\circ = -2.303 RT \log K_c$

$$\Delta G^\circ > 0$$

$$\Rightarrow -2.303 RT \log K_c > 0$$

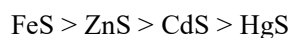
$$\Rightarrow K_c < 1$$

46. (2) For AB type salt

$$s = \sqrt{K_{sp}}$$

\therefore Higher the K_{sp} greater will be the solubility.

K_{sp} order (given)



47. (4) For basic buffer

$$\text{pOH} = \text{p}k_b + \log \frac{[\text{conjugate acid}]}{[\text{Base}]}$$

$$= 5 + \log \frac{0.01}{0.1}$$

$$= 5 + \log 0.1 = 5 - 1 = 4$$

$$\text{pH} = 14 - \text{pOH} = 14 - 4 = 10$$

48. (1) HNO_2 is weak acid and NaNO_2 is salt of HNO_2 with NaOH a strong base

$\therefore \text{HNO}_2 + \text{NaNO}_2$ is a buffer.

49. (3) $\text{NaCO}_3 \longrightarrow$ salt of NaOH (strong base) and H_2CO_3 (weak acid)

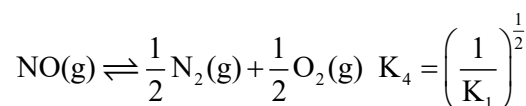
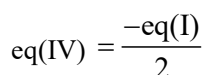
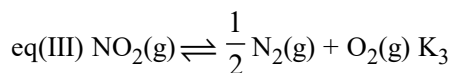
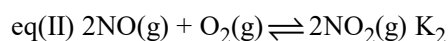
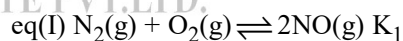
$\therefore \text{CO}_3^{2-}$ will undergo hydrolysis and result in alkaline solution

$$\Rightarrow \text{pH} > 7$$

KCl , NaCl and CuSO_4 all are strong electrolytes and hence will not undergo hydrolysis

$$\Rightarrow \text{pH} = 7$$

50. (4) Correction in Question



REDOX REACTION (SOLUTION)

01. (3) Oxidation means addition of E.N element (like 'O') or increase of '+ve' charge.
02. (1) Addition of E.N element is oxidation.
03. (3) $2\text{Na}_{(s)} + \text{Cl}_2 \longrightarrow 2\text{NaCl}$
Na losses $2 e^-$ (one each) & Cl gain $2 e^-$.
04. (4) Oxidation involve loss of e^- & act as reducing agent.
05. (3) Na is oxidised & H_2 is reduced.
06. (2) Zn replace Cu from CuSO_4 .
$$\text{Zn} + \text{CuSO}_4 \longrightarrow \text{ZnSO}_4 + \text{Cu}$$

(Blue) (Colourless)
07. (3) Cu donot replace Zn from ZnSO_4 i.e. no Rex^n occurs.
08. (2) Electron is always transferred from low E.N to high E.N. atom.
09. (4) $\text{O}_3 \rightarrow$ in elemental state o.s is always zero.
10. (2) Oxygen shows $-2, -1, -1/2, 0$ & even '+2' o.s. with 'F'.
11. (1) Oxide '-2' peroxide = -1 , superoxide = $-1/2$
12. (3) LiAlH_4
$$\begin{matrix} +1 & +3 & & 4x \\ \text{Li} & \text{Al} & \text{H}_4 & \\ & & & \end{matrix}$$

 $1 + 3 + 4x = 0$
 $4x = -4$
 $x = -1$.
13. (4) $\text{Si} \rightarrow +4$
 $\text{P} \rightarrow +5$
 $\text{S} \rightarrow +6$
 $\text{Cl} \rightarrow +7$.
14. (3) $\text{Hg}_2(\text{I}) \text{Cl}_2$ is correct notation.
15. (2) $\text{Ca} \overset{+2}{\text{O}} \overset{-2}{\text{C}} \overset{+4}{\text{O}} \overset{-2}{\text{O}} \longrightarrow \text{Ca} \overset{+2}{\text{C}} \overset{+4}{\text{O}} \overset{-2}{\text{O}}_3$. no oxidation no. is changed.
16. (1) Above question.
17. (3) As per definition.
18. (4) Sodium is more reactive while Fe is less.
Alkali metals are good oxidising agent.
19. (4) Ag is less reactive than hydrogen & so donot react to replace H from HCl.
20. (3) Br is more reactive than I_2 & So. I_2 donot replace Br.
 $\text{F}_2 > \text{Cl}_2 > \text{Br}_2 > \text{I}_2$
21. (1) Due to more reactive nature of F.
22. (3) More E.N atoms act as oxidising agent i.e F_2 is strongest O.A.
23. (4) Intermediate O.S. can both oxidise or reduce.
24. (4) $\text{ClO}_4^- \rightarrow \overset{+7}{\text{Cl}}$ is in highest O.S. & So act as only oxidising agent.
25. (2)
$$\begin{matrix} \text{O} & & \text{O} & & \text{O} \\ || & & || & & || \\ \text{O}=\text{Br} & \text{---} & \text{Br} & \text{---} & \text{Br}=\text{O} \\ || & & || & & || \\ \text{O} & & \text{O} & & \text{O} \end{matrix}$$

Average O.S. $\Rightarrow \frac{6+4+6}{3}$
 $\Rightarrow \frac{16}{3}$
26. (2)
$$\begin{matrix} \text{O} & & \text{O} & & \text{O} \\ || & & | & & || \\ \text{O}-\text{S} & \text{---} & \text{S} & \text{---} & \text{S} \\ || & & | & & || \\ \text{O} & & \text{O} & & \text{O} \end{matrix}$$

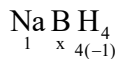
 $n = 2 \times 3 = 6$
27. (2) $\overset{+6}{\text{Cr}_2\text{O}_7^{2-}} + \text{SO}_3^{2-} \rightarrow 2\text{Cr}^{3+} + \text{SO}_4^{2-}$
 $n = \cancel{6}$
28. (3) $\overset{+7}{\text{MnO}_4^-} + \text{Br}^- \rightarrow \overset{+4}{\text{MnO}_2} + \overset{+5}{\text{BrO}_3^-}$
 $n = \cancel{2}$
$$2\text{MnO}_4^- + \text{Br}^- + 2\text{H}^+ \rightarrow 2\text{MnO}_2 + \text{BrO}_3^- + \text{H}_2\text{O} \text{ (In acidic)}$$

$$2\text{MnO}_4^- + \text{Br}^- + \text{H}_2\text{O} \rightarrow 2\text{MnO}_2 + \text{BrO}_3^- + 2\text{OH}^- \text{ (In basic)}$$
29. (2) $\text{Zn} \rightarrow$ anode (oxidation)
 $\text{Cu} \rightarrow$ cathode (reduction)
30. (3) $E^\circ = -ve' \rightarrow$ strong reducing agent.
 $E^\circ = '+ve' \rightarrow$ strong oxidising agent.
31. (4) Theory
32. (2) More E.N. i.e. more oxidising power.
 $\text{F}_2 > \text{Cl}_2 > \text{Br}_2 > \text{I}_2$
33. (3) $\text{Tl} \text{SO}_4 \longrightarrow \text{Tl}_2\text{SO}_4$
$$\begin{matrix} & \text{SO}_4 & \\ & \nearrow & \searrow \\ 1 & & 2 \end{matrix}$$

34. (1) More '-ve' electrode potential i.e. more reducing power. i.e. $p > s > t > r > q$.

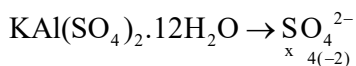
35. (3) $\text{NaH}_2\text{PO}_4 \rightarrow 1 + 2 + x - 8 = 5$
 $\begin{matrix} 1 & 2(1) & x & 4(-2) \end{matrix}$

$x = +5$

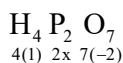


$1 + x - 4 = 0$

$x = 3$



$x + (-8) = -2$ $x = 6$



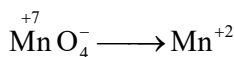
$4 + 2x - 14 = 0$

$x = 5$

36. (4) Reduction potential

$\propto \frac{1}{\text{reduced}} \propto \text{Oxidising power.}$

37. (2) KMnO_4 efficiently work in acidic medium.



MnO_4^- in basic medium with I^\ominus



38. (3) $\text{O} = \overset{+2}{\text{C}} = \overset{0}{\text{C}} = \overset{+2}{\text{C}} = \text{O}$

39. (3) S O_2
 $\begin{matrix} x & 2(-2) \end{matrix}$ $x + (-4) = 0$ $x = 4$

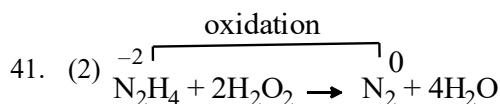
Intermediate O.S. [S \rightarrow -2 to +6]

So, SO_2 act as both oxidising as well as reducing agent.

40. (1) $\text{AgF}_2 \rightarrow \text{AgF} + \text{F}^\ominus$
 (unstable)

(so it is strong oxidising agent).

Ag^{2+} have more tendency to gain $1e^-$.



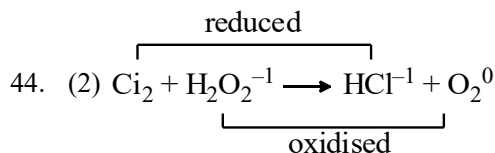
So; N_2H_4 act as reducing agent.

[Slight mistake in question]

42. (2) P_4 molecule is both oxidation & reduced in same rex^n & so this rex^n is called disproportionation rex^n .

43. (3) $\text{f} \rightarrow -1$; $\text{Ne} \rightarrow 0$

$\text{Cs} \rightarrow +1$; $\text{I} \rightarrow -1$ to $+7$



$\text{H}_2\text{O}_2 \rightarrow$ oxidising agent.

45. (4) $\text{N}_2\text{O}_2 \rightarrow 2\text{NO}$ (Combination).

$2\text{Pb(NO}_3)_2 \rightarrow 2\text{PbO} + 4\text{NO}_2 + \text{O}_2$ (decomposition).

$\text{NaH} + \text{H}_2\text{O} \rightarrow \text{NaOH} + \text{H}_2$ (displacement).

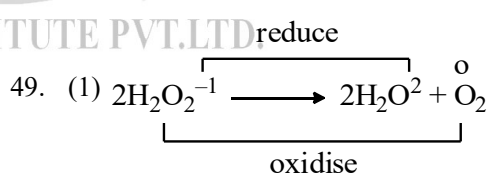
$2\text{NO}_2 + 2\text{OH}^- \rightarrow \text{NO}_2^- + \text{NO}_3^- + \text{H}_2\text{O}$ (disproportionation).

i) s ii) r iii) q iv) p

46. (1) $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$
 (Blue) (Colourless)

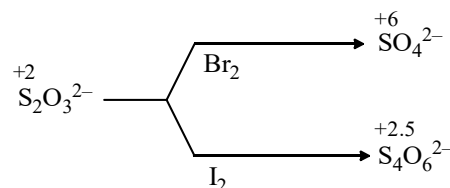
47. (2) O.N of oxygen may be -1, -2, -1/2 or even '+ve' in same compounds.

48. (4) $E^\ominus = \text{'-ve'}$ $\left\{ \begin{array}{l} \text{strong reducing agent} \\ \text{weak oxidising agent} \end{array} \right.$



(disproportionation reaction)

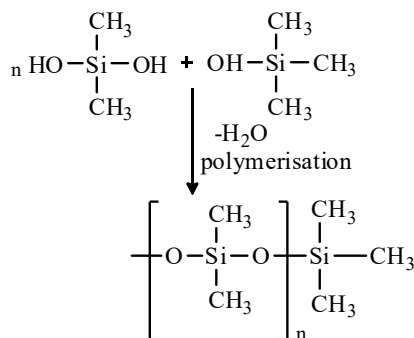
50. (1) Br_2 is stronger oxidising agent than I_2 & So Br_2 oxidise mole i.e. from +2 to +6 while I_2 cannot.



p-BLOCK ELEMENTS [GROUP 13 & 14] (SOLUTION)

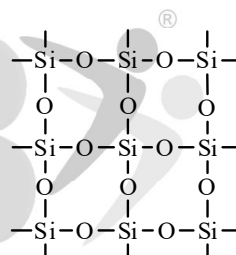
01. (3) Increasing order of atomic radii of group-13 elements is Ga (135 pm) < Al (143 pm) < In (167 pm) < Tl (170 pm)
02. (2) Successive ionisation enthalpy values increases in the order $\Delta_1 H_1 < \Delta_1 H_2 < \Delta_1 H_3$
03. (1) Aluminium in air is ordinarily protected by a molecule-thin layer of its own oxide.
04. (2) COO reacts with B_2O_3 to give blue bead of $\text{CO}(\text{BO}_2)_2$
05. (1) $3\text{B}_2\text{H}_6 + 2\text{NH}_3 \longrightarrow \text{B}_2\text{H}_6 \cdot 2\text{NH}_3$
 $\text{B}_2\text{H}_6 \cdot 2\text{NH}_3 \xrightarrow{450\text{K}} 2\text{B}_3\text{N}_3\text{H}_6 + 3\text{H}_2$
06. (1) On heating orthoboric acid (H_3BO_3) at 370K or above, it changes to metaboric acid (HBO_2) on further heating this yields boric oxide B_2O_3
07. (2) $\text{Na}_2\text{B}_4\text{O}_7 + 2\text{HCl} + 5\text{H}_2\text{O} \longrightarrow 2\text{NaCl} + 4\text{H}_3\text{BO}_3$
08. (3) Four $2\text{C}-2\text{e}^-$ bonds and two $3\text{C}-2\text{e}^-$ bonds.
09. (2) Copper is a better conductor as compared to aluminium, aluminium is about 40% less conductive than copper
10. (3) G-14 General electronic configuration is ns^2np^2
11. (3) Due to ineffective shielding of the d and f orbital, Pb has a higher ionization enthalpy than Si.
12. (3) Boron, Silicon, germanium, arsenic, antimony, tellurium and polonium all are radioactive.
13. (4) Lead does not show catenation
14. (4) 1, 2 and 3 are the resonance structures of CO_2 . Carbondioxide, has three resonance structures, out of which one is a major contributor
15. (3) $\text{Fe}_2\text{O}_3 + \text{CO} \rightarrow \text{Fe} + \text{CO}_2$
16. (2) The tendency to show (+2) oxidation state increases as we move down the group this is due to inert pair effect. Thus, the stability of elements in +2 O.S increases as we move down the group in the order $\text{C} < \text{Si} < \text{Ge} < \text{Sn} < \text{Pb}$
17. (2) $\text{AlCl}_3 + 3\text{H}_2\text{O} \rightarrow \text{Al}(\text{OH})_3 + 3\text{HCl}$.
18. (3) Limestone is heated to produce carbondioxide
 $\text{CaCO}_3 \xrightarrow{\Delta} \text{CaO} + \text{CO}_2$.
19. (1) Buckminster fullerene is the first discovered fullerene which contains 60 carbon atoms. The molecule has shape of soccer ball in which 12 five membered and 20 Six membered rings are present.

20. (4) Orthoboric acid consists of 6 hydrogen bonds.
21. (3) Mixture of CO and H_2 is known as water gas
22. (4) The chain length of silicon polymer can be controlled by adding $(\text{CH}_3)_3\text{SiCl}$ which block the end [PTO]



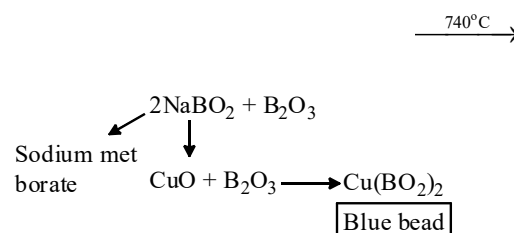
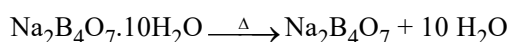
23. (4) Silicon dioxide is a covalent, three-dimensional network solid in which silicon atom is bonded to four oxygen atoms which are arranged tetrahedrally around it and each oxygen atom is attached to two silicon atoms by covalent bonds.

Each corner is further shared by another tetrahedron.



The entire crystal may thus be considered as a giant molecule in which eight-membered rings are formed with alternate silicon and oxygen atoms.

24. (2) $\text{SiO}_2 + 4\text{HF} \rightarrow \text{SiF}_4 + 2\text{H}_2\text{O}$.
25. (3) $\text{SiO}_2 + 2\text{Na}_2\text{O} \rightarrow \text{Na}_2\text{SiO}_3 + \text{H}_2\text{O}$
 $\text{SiO}_2 + 4\text{HF} \rightarrow \text{SiF}_4 + 2\text{H}_2\text{O}$.
26. (4) Boron is considered as typical non-metal.
27. (2) Meta borate is formed in borax bead test

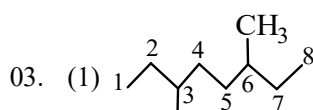


28. (1) A combustible mixture of nitrogen, carbon monoxide, and hydrogen are called producer gas
29. (2) ZSM-5 converts alcohol into gasoline
30. (4) Weak acid and its salt with strong base acts as acidic buffer $[H_2CO_3/NaHCO_3]$
This buffer solution does not alter the pH of blood hence it helps to maintain pH of blood between 7.26 to 7.42.
31. (4) Quartz, cristobalite and tridymite all are crystalline form of silica.
32. (1) $HCOOH \xrightarrow{Conc.H_2SO_4} CO + H_2O$
33. (1) Fullerenes have closed cage like structures unlike diamond and graphite. They do not have dangling bonds.
34. (1) 1st member of each group in p-block has 2p orbitals able to form strongest p π -p π bonds. Due to its size.
35. (2) Graphite is the most thermodynamically stable form of carbon.
36. (2) Fact based
37. (3) $Na_2B_4O_7 \cdot 10H_2O \xrightarrow{\Delta} Na_2B_4O_7 \xrightarrow{\Delta} 2NaBO_2 + B_2O_3$
 $Na_2B_4O_7 + 7H_2O \rightarrow 2NaOH + 4H_3BO_3$
38. (2) Fact based
39. (3) Due to smaller size of oxygen atom, Si-O bond is stronger than Si-Si bond.
40. (1) Fact based
41. (4) Gallium has smaller atomic radius than aluminium because the d-orbital which is highly diffused offers poor screening effect which results in increased nuclear charge. So, radius decreases.
42. (1) PbI_4 does not exist because the iodine reduces the lead to Pb(II) and the Pb oxidises the iodide to iodine (I₂) since the iodine is not a reducing agent to reduce Pb(II) to Pb, the compound PbI_2 is formed.
43. (1) The presence of 4th shell increases the size of gallium atom hence, an increase in ionization enthalpy due to high Z_{eff} overcomes the decrease in ionization due to increased size as a result the ionization of gallium is close to that of aluminium.
44. (4) $NaBH_4 + I_2 \rightarrow B_2H_6 + 2NaI + H_2$
 $3NaBH_4 + 4BF_3 \rightarrow 3NaBF_4 + 2B_2H_6$
 $2BF_3 + 6NaH \xrightarrow{450K} B_2H_6 + 6NaF$
45. (2) Boron atom in trihalides has only six electrons in valence shell and can accept e^- . So, acts as Lewis acid. Due to back bonding the relative Lewis acid strength is $BI_3 > Br_3 > BCl_3 > BF_3$
46. (4) Fact based
47. (3) B_2H_6 consists of two electrons each, three centered bonds. Each boron atom is in a link with four hydrogen atoms. This makes tetrahedral geometry hence boron is sp^3 hybridised
48. (3) Carbon shows maximum catenation and ability to form p π -p π multiple bonds.
49. (4) $AlCl_3 \xrightarrow[\text{aqueous solution}]{\text{Acidified}} [Al(H_2O)_6]^{3+}$
Hybridisation = sp^3d^2
50. (3) When air is passed over red-hot coke producer gas is obtained
 $2C + O_2 + N_2 \rightarrow 2CO + N_2$

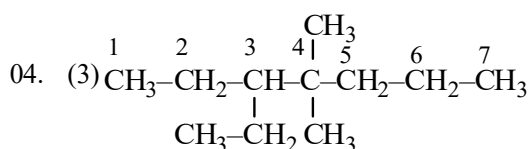
ORGANIC CHEMISTRY SOME BASIC PRINCIPLE & TECHNIQUES (SOLUTION)

01. (4) Propanone and propanal are functional group isomer to each other.

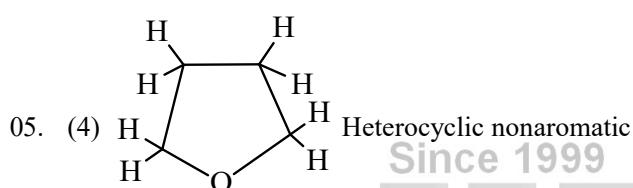
02. (4)



6-methyl (secondary prefix)

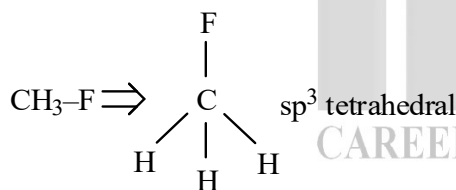


3-ethyl-4,4-dimethylheptane



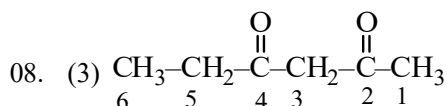
Tetrahydro furan (THF)

06. (4) Methyl fluoride



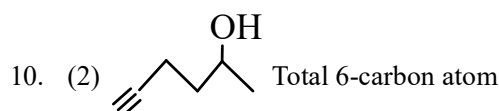
sp^3 tetrahedral

07. (3) $-\text{NO}_2$ (-I, -R)



Hexane-2,4-dione

09. (3) Cyclohexene \Rightarrow



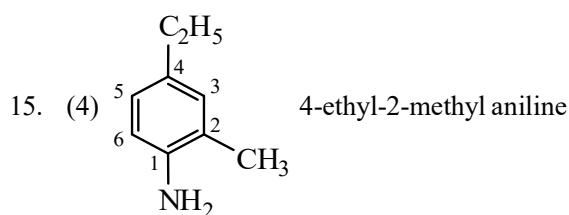
Total 6-carbon atom

11. (2) $\text{N} \equiv \text{C}-\text{CH}(\text{OH})-\text{C} \equiv \text{N}$ condensed formula $\text{HOCH}(\text{CN})_2$

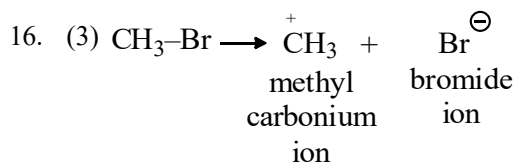
12. (3) C_{60} cluster

13. (1) Homologous series compounds differ molecular formula

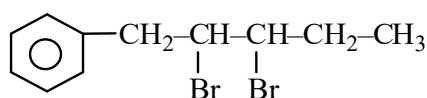
14. (3) Carbamoyl \Rightarrow 2^o prefix \Rightarrow $-\text{CONH}_2$ functional group in IUPAC



4-ethyl-2-methyl aniline



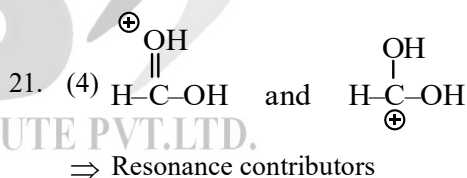
17. (4) 2,3-dibromo-1-phenyl pentane



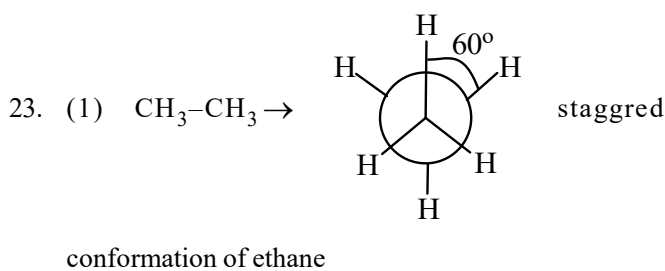
18. (2) 2,2-Dimethyl propane and 2-methyl butane are chain isomer to each other.

19. (3) $\text{CH}_3\text{-C}^{\oplus}(\text{O})$ electrophile

20. (3) Phosphoric acid (H_3PO_4)



22. (3) $\text{CH}_2=\text{CH}-\text{CH}_2-\text{CH}_2-\text{C}\equiv\text{CH}$
 $\text{C}_1 - \text{C}_3 \Rightarrow \text{sp}^2 - \text{sp}^3$



staggered

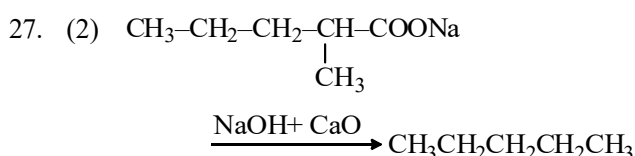
conformation of ethane

24. (2)

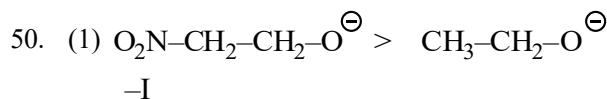
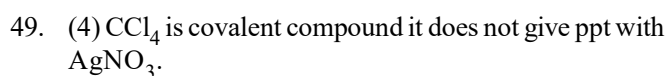
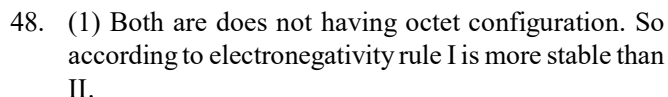
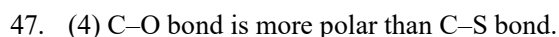
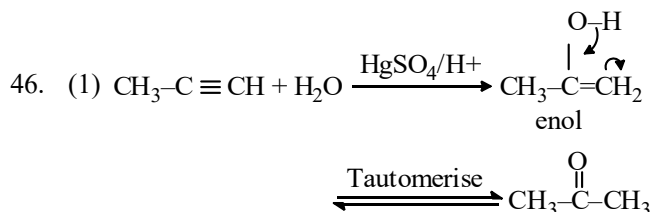
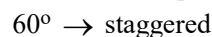
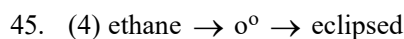
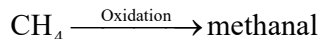
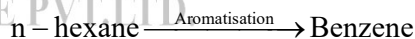
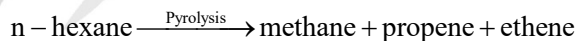
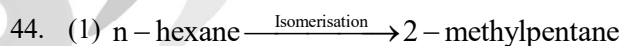
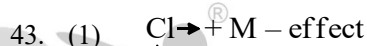
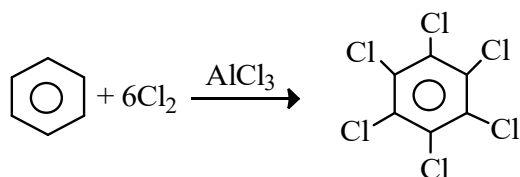
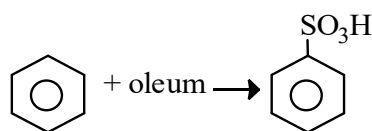
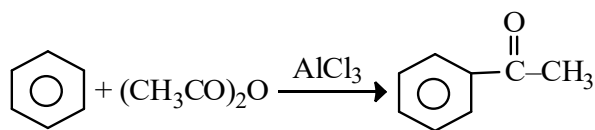
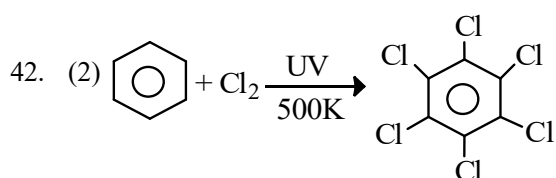
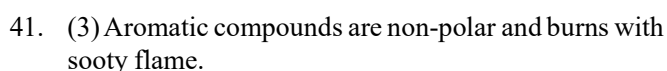
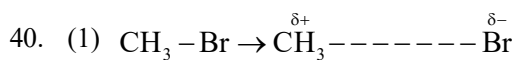
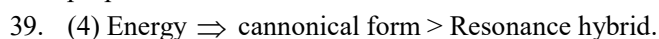
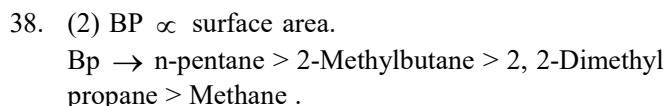
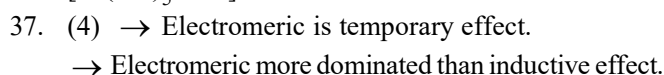
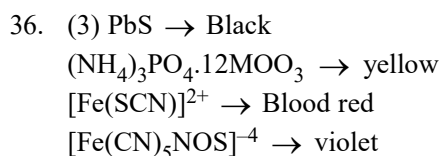
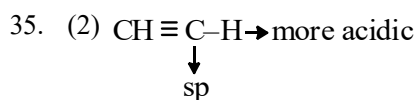
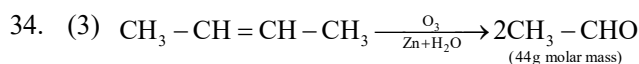
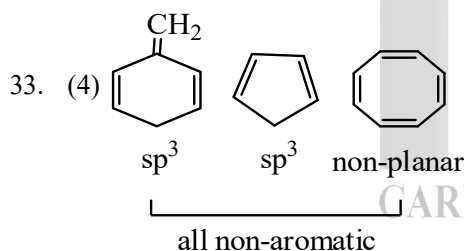
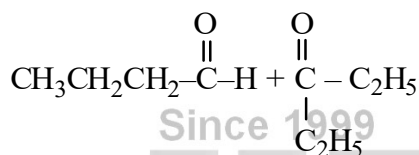
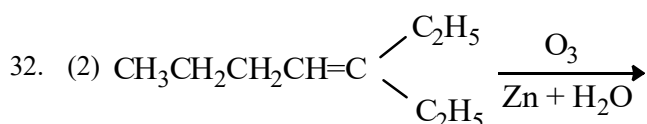
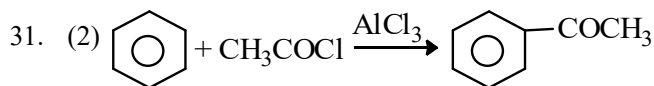
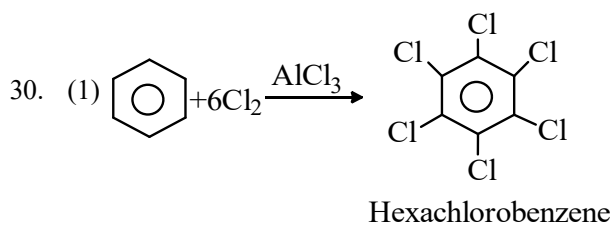
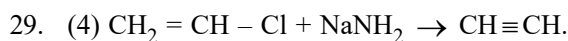
Only one structure exhibits geometrical isomers

25. (4)

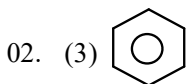
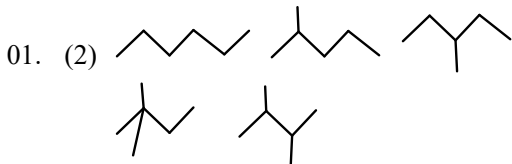
26. (3) Correct IUPAC is 5-sec-butyl-4-isopropyldecane



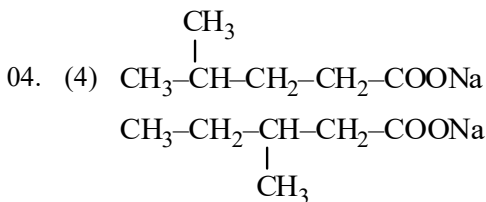
28. (3) $-\text{COCl} > -\text{CONH}_2 > -\text{CN} > -\text{CHO}$



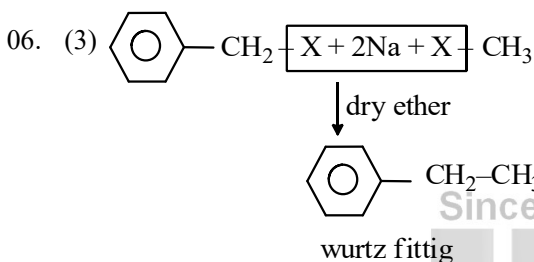
HYDROCARBONS (SOLUTION)



03. (2) Diagram from NCERT.



05. (3) CH₄ can't be prepared.



07. (1) $BP \propto \frac{1}{\text{Branching}}$ (Isomers).

08. (1) 1°H = 6 × 1 = 6

1°H = 3 × 1 = 3

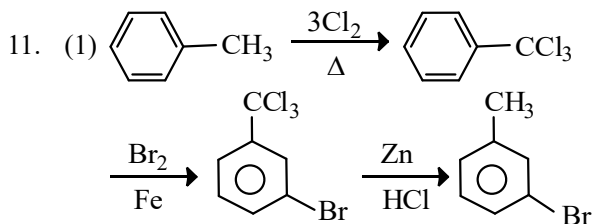
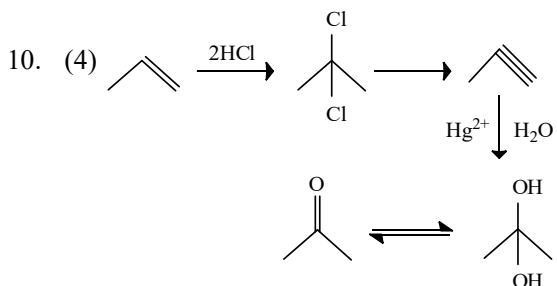
2°H = 2 × 3.8 = 7.6

3°H = 1 × 5 = 5

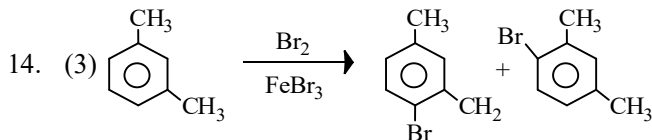
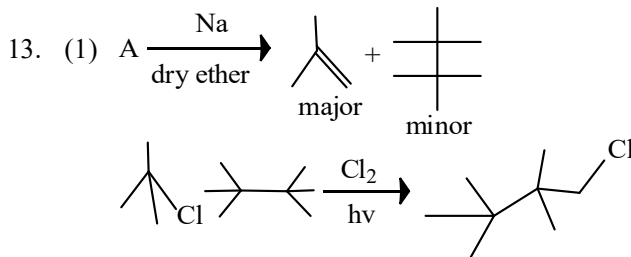
21.6

2°H % yield = $\frac{7.6}{21.6} \times 100 = 35.18\%$

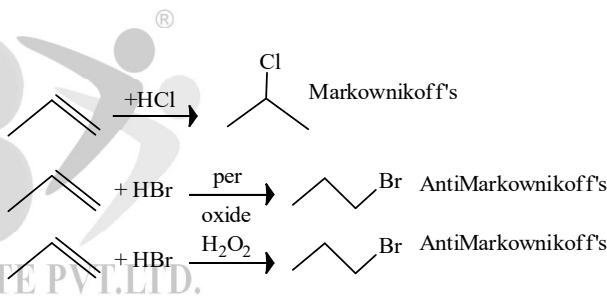
09. (3)



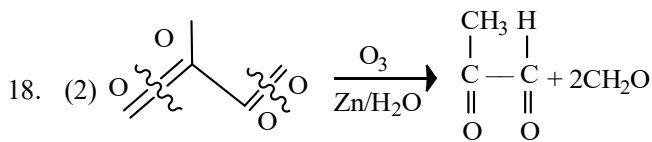
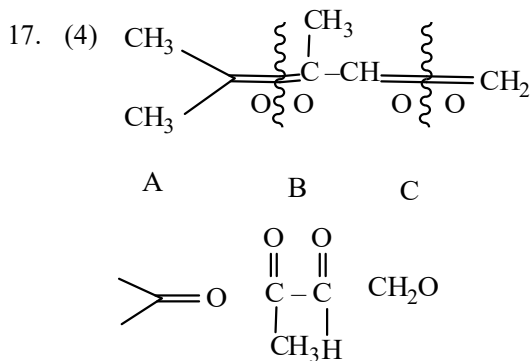
12. (2) Pyrolysis

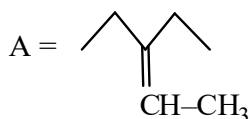
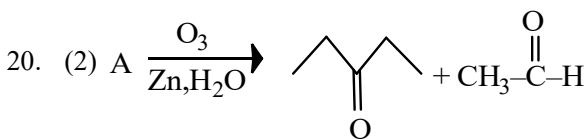
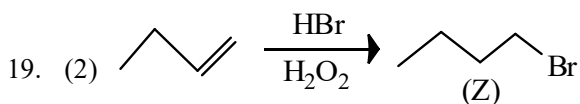


15. (4)

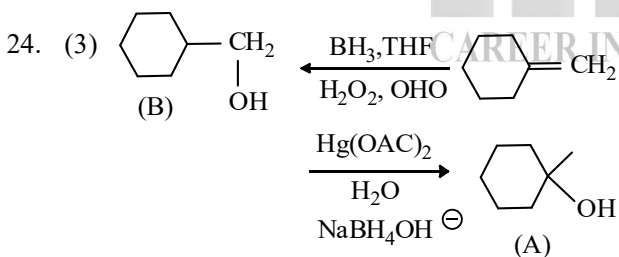
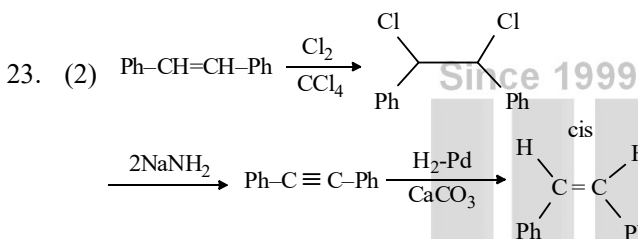
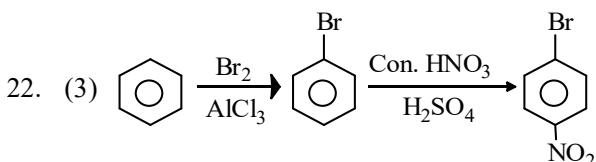
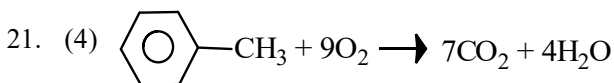


16. (3) In Kharasch effect reaction follows free radical addition

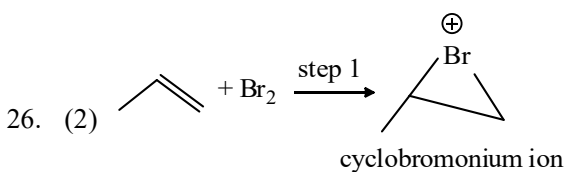
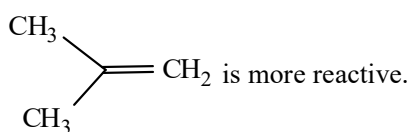




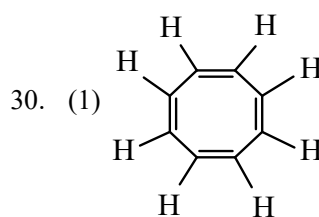
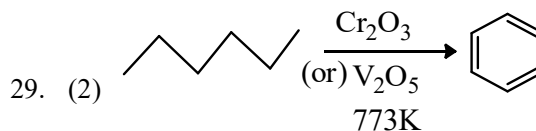
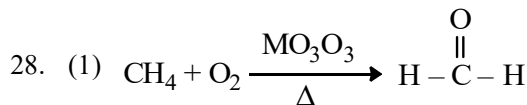
3-ethyl pent-2-ene



25. (3) $r_{\text{EAR}} \propto$ nucleophilicity of alkene (stability of C^\oplus).

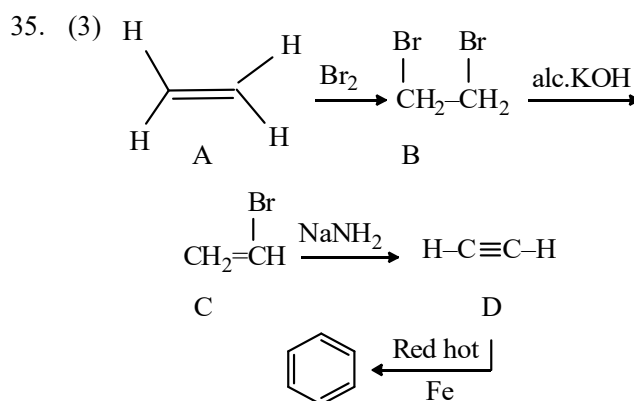
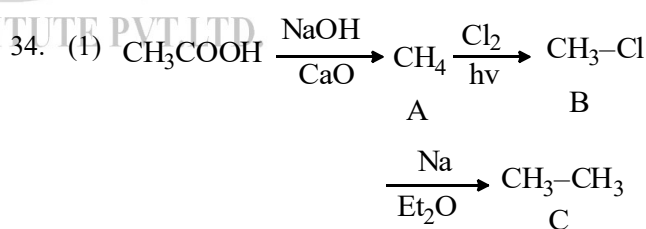
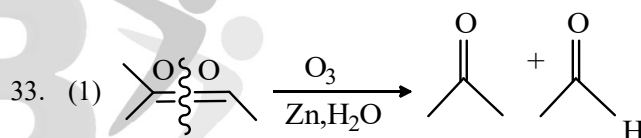
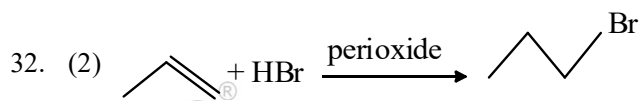


27. (3) $r_{\text{EAR}} \propto$ nucleophilicity of alkane
 $r_{\text{EAR}} \propto$ stability of C^\oplus .
 $\text{R} > \text{Q} > \text{P}$.



16 σ and 4 π (option incorrect)

31. (4) No change in B. Angle and B.length in conformers of ethane.



36. (2) Hydrocarbons are also used for the manufacture of polymers like polythene, polystyrene due to presence of unsaturated bonds.
37. (3) Wurtz reaction is preferred for the preparation of higher alkanes containing even no. of carbon atoms
38. (4) Rate of nitration of benzene and hexadeutero benzene are not different since removal of H is not slow step.
39. (2) Both are correct answers.
40. (2) Both are correct answers.
41. (1) Alkene that can form stable carbocation will react faster.
42. (3) Alkene with even no. of carbon atom have higher melting point than adjacent alkane of odd no. of carbon atoms. (Incorrect statement)

Melting point \propto molar mass


propane has least melting point. (Exception)

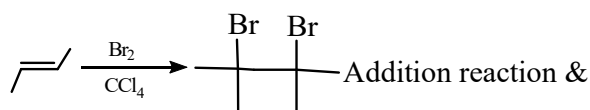
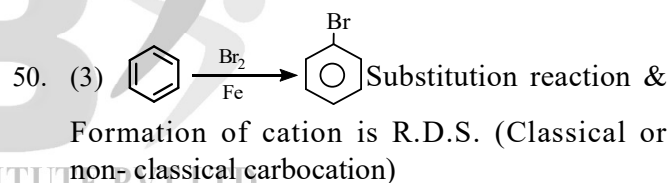
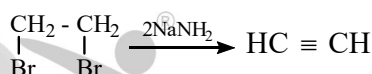
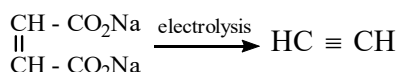
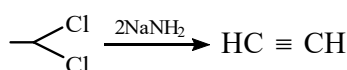
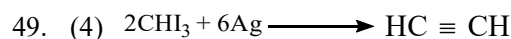
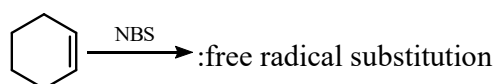
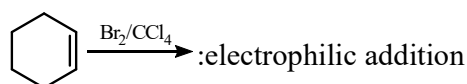
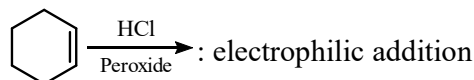
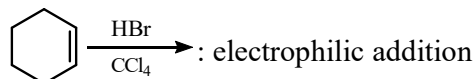
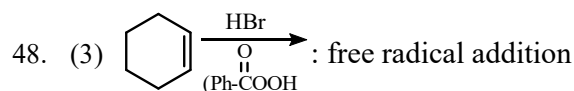
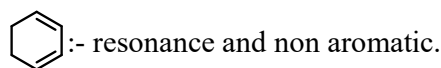
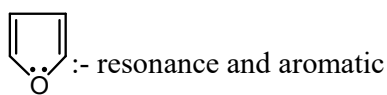
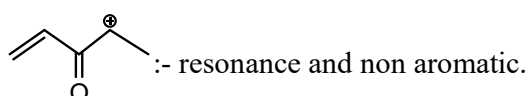
43. (2) In Kolbe's electrolytic method generally carbon dioxide liberates at anode while H_2 (Hydrogen gas) liberates at cathode.

Methane can not be prepared by Kolbe's electrolytic method.

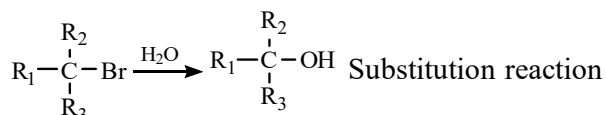
44. (4) HCl does not show peroxide effect therefore 1, 4 addition
45. (2) Propane gives 2-bromo propane when reacted with HBr & it is free radical substitution.
46. (1) (A)-(p,r); (B)-(p,r), (C)-(p,r), (D)-(p,r)

π bonded functional group is -M & lone pair functional group is +M.

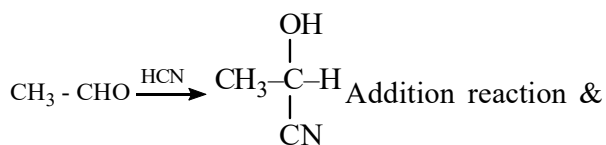
47. (4)  :- resonance and non aromatic.



Formation of cation is R.D.S. (Classical or non-classical carbocation)



, Formation of cation is R.D.S. (Classical or non-classical carbocation) & Racemic mixture



Racemic mixture