Some Basic Concepts of Chemistry

EXERCISE-I (MHT CET LEVEL)

Q.1 **(3)**

Each has three significant figure. When zero is used to locate the decimal point it is not considered as significant figure.

Q.2 (3)

Vapour density = $\frac{48}{2} = 24$

Q.3 (1)

100 amu of He = $\frac{100}{4}$ atoms of He = 25 atoms.

[1 a.m.u. = mass of one proton (approx.)]

Q.4 (3)

Exa = 10^{18}

Q.5 (4)

The concerned chemical reactions are

(i) $\underset{64kg}{CaC_2} + 2H_2O \rightarrow Ca(\mathrm{OH})_2 + \underset{Ethyne\,26kg}{\mathrm{C_2H_2}}$

(ii) $C_2H_2 + H_2 \rightarrow \underset{Ethylene,28kg}{C_2\,\mathrm{H}_4}$

(iii) $\underset{\substack{n\times 28kg \\ or\ 28kg}}{nC_2H_2} \rightarrow \underset{\substack{n\times 28kg\ \text{polythrne} \\ or\ 28\ \text{kg}}}{[-\mathrm{CH_2}-CH_2-]_n}$

Thus 64kg of CaC_2 gives 26 kg of acetylene which in turn gives 28kg of ethylene whose 28kg gives 28kg of the polymer, polythene.

Q.6 (1)

Given, mass ratio is C : H : O (6 : 1 : 24) so, molar ratio will be 6/12 : 1/1 : 24/16 = 1 : 2 : 3
therefore, HO-(C=O)-OH has molar ratio 1 : 2 : 3.

Q.7 (3)

This is Avogadro's hypothesis.
According to this, equal volume of all gases contain equal no. of molecules under similar condition of temperature and pressure.

Q.8 **(3)**

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Q.9 (3)

Q.10 (4)

Q.11 (3)

Q.12 (3)

Q.13 (3)

Q.14 (4)

Q.15 (2)

Q.16 (4)

Q.17 (1)

Q.18 (2)

Q.19 (1)

Q.20 (3)

Q.21 (4)

According to Dulong and Pettit's law
Atomic weigth × Specific heat =6.4(approx)
This law is applicable only to solid elements but it fails to explain very high specific heat of diamond.

Q.22 (1)

Q.23 (1)

Q.24 (3)

Q.25 (3)

Q.26 (1)

$$\mathrm{Zn} + \mathrm{H_2SO_4} \rightarrow \mathrm{ZnSO_4} + \mathrm{H_2}$$
$$\mathrm{Zn} + 2\mathrm{NaOH} \rightarrow \mathrm{Na_2ZnO_2} + \mathrm{H_2}$$

$\therefore$ Ratio of volumes of $\mathrm{H_2}$ evolved is 1:1

Q.27 (1)

Q.28 (1)

Q.29 (3)

Q.30 (1)

Q.31 (2)

Q.32 (3)

Q.33 (4)

Let E be the equivalent weight of metal

So, $\frac{E+17}{E+8} = \frac{1.52}{0.995}$

[17 is equivalent weight of OH and 8 is equivalent weigth of oxygen]

$$\Rightarrow 0.995E + 17 \times 0.995$$
$$= E \times 1.52 + 8 \times 1.52$$

$$\Rightarrow 0.525E = 16.915 - 12.16 = 4.755$$

$$\therefore E = \frac{4.755}{0.525} = 9$$

Q.34 (2)

Eq. of acid = Eq of base,

$$\therefore \frac{0.45}{\text{E.wt}} = \frac{20 \times 0.5}{1000} = \text{E.wt} = 45$$

$$\text{Basicity} = \frac{\text{M.wt}}{\text{E.wt}} = \frac{90}{45} = 2$$

Q.35 (1)

Eq. of A = Eq. of B

$$\frac{m_A}{E_A} = \frac{m_B}{E_B}; E_A = \frac{m_A}{E_B} \times E_B$$

Q.36 (3)

Q.37 (4)

Q.38 (1)

Q.39 (1)

$$\text{Eq of } KMnO_4 \text{ used } = \frac{50 \times 1}{1000 \times 10} = 0.005$$

$\therefore$ Eq of Ferrous ammonium sulphate reacted =0.005

$\therefore$ weight of Ferrous ammonium sulphate needed

$= 0.005 \times 392 = 1.96\,g$

Thus percentage purity of Ferrous ammonium sulphate is 50%

Q.40 (3)

Q.41 (4)

Q.42 (2)

Q.43 (3)

The element A is ns^2p^1 and B is ns^2p^4. They can form compound of the type A_2B_3.

Q.44 (4)

Xe = 53.5% $\therefore$ F = 4635%

Relative number of atoms Xe

$$= \frac{53.5}{131.2} = 0.4 \text{ and } F = \frac{46.5}{19} = 204$$

Simple ratio Xe = 1 and F = 6: Molecular formula is XeF_6

O.N. of Xe is + 6

EXERCISE-II (NEET LEVEL)

Q.1 (1)

Q.2 (1)

Q.3 (3)

Q.4 (4)

$\because$ 17*gm* NH_3 contains 6×10^{23} molecules of NH_3

$\therefore$ 4.25*gm* NH_3 contains = $\frac{6 \times 10^{23}}{17} \times 4.25$

$\therefore$ No. of atoms $= \frac{6 \times 10^{23} \times 4.25}{17} \times 4 = 6 \times 10^{23}$.

Q.5 (2)

$\because$ 22400*cc* of gas at STP has 6×10^{23} molecules

$\therefore$ 1.12 $\times 10^{-7}$ of gas at STP has $\frac{6 \times 10^{23} \times 1.12 \times 10^{-7}}{22400}$

$= .03 \times 10^{14} = 3 \times 10^{12}$.

Q.6 (4)

$d = \frac{M}{V}$ (*d* = density, *M*= mass, *V* =volume)

Since *d* = 1

So, $M = V$

18*gm* = 18*ml*

18*ml* = N_A molecules (N_A = avogadro's no.)

1000*ml* $= \frac{N_A}{18} \times 1000 = 55.555\,N_A$.

Q.7 (2)

2*gm* of oxygen contains atom = $\frac{2}{16} = \frac{1}{8}$ mole

also 4*g* of sulphur $= \frac{4}{32} = \frac{1}{8}$ mole.

Q.8 (1)

100*gm* caffeine has 28.9*gm* nitrogen

194*gm* caffeine has = $\frac{28.9}{100} \times 194 = 56.06\,gm$

$\therefore$ No. of atoms in caffeine $= \frac{56.06}{14} \approx 4$.

Q.9 (1)

$\because$ 40*gm NaOH* contains 16*gm* of oxygen

$\therefore$ 100*gm* of *NaOH* contains $\frac{16}{40} \times 100$ =40% oxygen.

Q.10 (1)

Urea- $NH_2 - CO - NH_2$

$\because$ 60*gm* of urea contains 28*gm* of nitrogen

$\therefore$ 100*gm* of urea contains $\frac{28}{60} \times 100 = 46.66$.

Q.11 (4)

$C = 24\,gm$, $H = 4\,gm$, $O = 32\,gm$

So, Molecular formula $= C_2H_4O_2$

So, Empirical formula $= CH_2O$

(Simplest formula).

Q.12 (3)

In $Fe(CNS)_3 . 3H_2O$

% of $H_2O = \frac{3\times18}{284}\times100 = 19\%$.

Q.13 (2)

Element	At.wt.	Mole	Ratio
Empiricalformula			
C=86%	12	7.1	1
CH_2			
H=14%	1	14	2

Belongs to alkene

Q.14 (2)

Element	%(1)	At.wt.(2)
a/b Ratio		
X	50	10
5 2		
Y	50	20
2.5 1		

Simplest formula $= X_2Y$

Q.15 (3)

$$N = \frac{\text{W}(gm)\times1000}{V\times\text{Eq.wt.}}$$

1500ml of 0.1N HCl = 150ml (N)

$1 = \frac{\text{W}(gm)\times1000}{150\times40}$, W $(gm) = \frac{150\times40}{1000} = 6gm$ ..

Q.16 (3)

$$
\begin{array}{lcccc}
 & 4NH_{3(g)} + & 5O_{2(g)} \rightarrow & 4NO_{(g)} + & 6H_2O_{(g)} \\
t=0 & 1 & 1 & 0 & 0 \\
t=t & 1-4x & 1-5x & 4x & 6x
\end{array}
$$

Oxygen is limiting reagent

So, $X = \frac{1}{5} = 0.2$ all oxygen consumed

Left $NH_3 = 1 - 4\times0.2 = 0.2$.

Q.17 (3)

0.1 M $AgNO_3$ will react with 0.1 M $NaCl$ to form 0.1 M $NaNO_3$. But as the volume is doubled, conc. of

$NO_3^- = \frac{0.1}{2} = 0.05\,M$

Q.18 (2)

$Mg^{+2} \equiv H_2$

$n = \frac{12\,gm}{24\,gm} = \frac{1}{2}$ mole of H_2

Q.19 (3)

$$\underset{100\,g}{CaCO_3} + \underset{2N}{2HCl} \rightarrow CaCl_2 + \underset{44\,g}{CO_2} + H_2O$$

100 g $CaCO_3$ with 2 N HCl gives 44 g CO_2

100 g $CaCO_3$ with 1 N HCl gives 22 g CO_2

Q.20 (4)

H_3PO_4 is tribasic so $N = 3M = 3\times1 = 3$.

EXERCISE-III (JEE MAIN LEVEL)

Q.1 (3)

In $Ca_3(PO_4)_2$

$$\frac{\text{mole of Ca atom}}{\text{mole of O atom}} = \frac{3}{8}$$

Mole of 'O' atom = $\frac{8}{3}$ (mole of Ca atom)

Mole of 'Ca' atom = 3

Q.2 (3)

NaI mass = $\frac{3\times0.5}{100}$ = 0.015 gm

No. of moles of NaI = $\frac{0.015}{150} = 1\times10^{-4}$

No. of I^- ions = $10^{-4}\times6.023\times10^{23} = 6.023\times10^{19}$

Q.3 (3)

(1) n = $\frac{10\times1}{18}$ = 0.55

(2) n = 0.1 × 5 = 0.5

(3) n = $\frac{12}{48}\times3$ = 0.75

(4) n = $\frac{N}{NA}$ = 0.2 × 2 = 0.4

Q.4 (3)

(1) n = $\frac{12}{12}$ = 1 (2) n = $\frac{8}{16}$ = 05

(3) n = $\frac{32}{32}$ = 1 (4) n = $\frac{24}{24}$ = 1

Q.5 (2)

Moles of $Mg_3(PO_4)_2 = \frac{1}{8}\times0.25 = 3.125\times10^{-2}$

Q.6 (1)

Sol.

C	H	O
Mass	24	8
32		
Moles	$\frac{24}{12}$	$\frac{8}{1}$
$\frac{32}{16}$		
Ratio	2	8
2		
Simple integer ratio	1	4
1		

Hence empirical formula is CH_4O

Q.7 (4)

X	Y
$\frac{75.8}{75}$	$\frac{24.2}{16}$
1.01	1.5×2
2	3

Q.8 (1)

$$\underset{(g)}{C_xH_y} + \left(x + \frac{y}{4}\right) O_2 \longrightarrow \underset{(g)}{XCO_2} + \frac{y}{2} H_2O \text{ (g)}$$

$$\frac{\left(1 + x + \frac{y}{4}\right)}{\left(x + \frac{y}{2}\right)} = \frac{600}{700}$$

$$x + 7 = \frac{5y}{4}$$

by option (1)

Q.9 (3)

$A : B : C \Rightarrow 1 : 3 : 5$

$b \Rightarrow x : y = 32 : 84$ by mass

$= 1 : 3$ by mole

$C \Rightarrow x : y = 16 : 5 \Rightarrow 16 : 70$

Q.10 (1)

$$\text{amount of butter} = \frac{2 \times 10^{-3}}{5.5 \times 10^{-6}} = 3.63.6 \text{ gm}$$

Q.11 (3)

$$\frac{\Delta x}{x} = \frac{\Delta y}{y}$$

$$\Rightarrow y' = y + \Delta y = \frac{16.006}{16} \times 107.868$$

Q.12 (2)

M_2O_3 $0.30 \times (2M + 48) = 48$

$0.6\,M = 0.7 \times 48$

$M = 7 \times 8 = 56$

Q.13 (1)

On balacing Na atoms on both sides of reaction, we get :

$y = 6x$.

$\therefore\ x : y = 1 : 6$ (only A option matches).

Q.14 (2)

$C = 84/12 = 7$ mole

$H_2 = 12\text{ g} = 6$ mole

$O_2 = 56/22.4 = 5/2$ mole

$12C + 11H_2 + 11/2\, O_2 \longrightarrow C_{12}H_{22}O_{11}$

$L.R. = O_2$

11/2 mole O_2 produce 1 mole sucrose

5/2 mole O_2 will for 5/11 mole sucrose

mass of sucrose = 5/11 × (mol. mass)

$= 5/11 \times 342$

$= 155.45$ g

Q.15 (2)

$C_6H_5\text{–}OH + 7\,O_2 \rightarrow 6CO_2 + 3H_2O$

30 $\quad 30 \times 7 = 210 \quad 30 \times 6 = 180$

$\Delta V = 210 + 30 - 180 = 60$

Q.16 (1)

Explanation : $2\,Ag + S \rightarrow Ag_2\,S$

2 × 108 g of Ag reacts with 32 g of sulphur

10 g of Ag reacts with $\frac{32}{216} \times 10 = \frac{320}{216} > 1$ g

It means 'S' is limiting reagent

32 g of S reacts to form 216 + 32 = 248 g of Ag_2S

1 g of S reacts to form $= \frac{248}{32} = 7.75$ g

Alternately

n_{eq} of $Ag = \frac{10}{108} = 0.0925$ n_{eq} of $S = \frac{1}{16} = 0.0625$

(n_{eq} = number of equivalents)

Since n_{eq} of S is less than n_{eq} of Ag

$\Rightarrow$ 0.0625 eq of Ag will react with 0.0625 eq of S to form 0.0625 eq of Ag_2S

Hence , amount of $Ag_2S = n_{eq} \times$ Eq. wt. of $Ag_2S = 0.0626 \times 124 = 7.75$ g

Q.17 (4)

$$M = \frac{\%\text{ by weight} \times 10 \times d}{Mw_2} = \frac{36.5 \times 10 \times 1.2}{36.5} = 12 \text{ M}$$

$$m = \frac{36.5 \times 1000}{36.5 \times (100 - 36.5)} = \frac{1000}{63.5} = 15.7 \text{ m}$$

Q.18 (2)

1000 mL solution contain 2 mole of ethanol or 1000 × 1.025 g solution contain 2 mole of ethanol

wt. of solvent = 1000 × 1.025 – 2 × 46

$$m = \frac{2}{1000 \times 1.025 - 2 \times 46} \times 1000$$

$$m = \frac{2}{933} \times 1000 = 2.143$$

Q.19 (2)

Q.20 (1)

$$\text{Mole fraction of A i.e. } X_A = \frac{n_A}{\text{Total moles}}$$

$$\text{So } X_{H2O} = \frac{n_{H_2O}}{\text{Total moles}}$$

$$\text{Now } \frac{X_A}{X_{H_2O}} = \frac{n_A}{n_{H_2O}}$$

and molality = $\frac{n_A \times 1000}{n_{H_2O} \times 18} = \frac{X_A \times 1000}{X_{H_2O} \times 18}$ =

$\frac{0.2 \times 1000}{0.8 \times 18} = 13.9$ Ans.

Q.21 (4)

Moles of Cl^- in 100 ml of solution = $\frac{2}{58.5} + \frac{4}{111} \times 2 + \frac{6}{53.5} = 0.2184$

Molarity of Cl^- = $\frac{0.2184}{100} \times 1000 = 2.184.$

Q.22 (4)

Conc. of cation = $\frac{400 + 300 + 200}{400}$

Conc. of anion = $\frac{200 + 300 + 400}{400}$

$\therefore$ Ratio of the conc. = 1

EXERCISE-IV

NUMERICAL VALUE BASED

Q.1 8

$x + 4(-2) = 0$

$x = +8$

Q.2 – 3

$x + 4(+1) = +1$

$x = -3$

Q.3 500

$M = \frac{18 \text{ moles of solute}}{(1 \text{ lit. solution})}$

Mass of solution = 1000 × 1.8 = 1800 g.

Mass of solute = 18 × 98 = 1764.

Mass of solvent = (1800 – 1764) g = 36 g.

Hence molality = $\frac{\text{moles of solute}}{\text{mass of solvent}} \times 1000 = \frac{18}{36} \times 1000$

= 500.

Q.4 40

$\because$ 40gm NaOH contains 16gm of oxygen

$\therefore$ 100gm of NaOH contains $\frac{16}{40} \times 100$ =40% oxygen.

Q.5 +3

$[Co(NO_2)_6]^{3-}$

$x + 6(-1) = -3$

$x = +3$

Q.6 6g

$N = \frac{W(gm) \times 1000}{V \times \text{Eq.wt.}}$

1500ml of 0.1N HCl = 150ml (N)

$1 = \frac{W(gm) \times 1000}{150 \times 40}$, W (gm) = $\frac{150 \times 40}{1000} = 6gm$.

Q.7 1

$2Al + KClO_3 \longrightarrow Al_2O_3 + KCl$

(excess) 1 mole ($KClO_3$ is L.R.)

1 mole ($KClO_3$ L.R.)

From mole-mole analysis

$\frac{n_{KClO_3}}{1} = \frac{n_{Al_2O_3}}{1} \Rightarrow n_{Al_2O_3} = 1$ mole

Q.8 22g

$\underset{100\,g}{CaCO_3} + \underset{2N}{2HCl} \rightarrow CaCl_2 + \underset{44\,g}{CO_2} + H_2O$

100 g $CaCO_3$ with 2 N HCl gives 44 g CO_2

100 g $CaCO_3$ with 1 N HCl gives 22 g CO_2

Q.9 1.

Conc. of cation = $\frac{400 + 300 + 200}{400}$

conc. of anion = $\frac{200 + 300 + 400}{400}$

$\therefore$ ratio of the conc. = 1

Q.10 27g

Let wg water in added to 16 g CH_3OH

molality = $\frac{16 \times 1000}{W \times 32} = \frac{500}{W}$

$\frac{500}{W} = \frac{x_A \times 1000}{(1 - x_A)m_B} = \frac{0.25 \times 1000}{0.75 \times 18}$

W = 27 gm.

Q.11 (1)
Q.12 (3)
Q.13 (1)
Q.14 (1)
Q.15 (3)
Q.16 (1)

PREVIOUS YEAR'S

Q.1 (1)
Q.2 (1)
Q.3 (2)
Q.4 (2)
Q.5 (4)
Q.6 (2)
Q.7 (3)

Q.8 (1)
Q.9 (2)
Q.10 (1)
Q.11 (4)
Q.12 (2)
Q.13 (3)

$\because$ 25.4 g of I_2 combines with 8g of oxygen
$\because$ 254 g of I_2 will combine with 80 g of oxygen

Element	Mass %	Atomic mass	Moles of element Molar ratio	Simplest whole number ratio
I	25.4	127	$\frac{25.4}{127} = 0.2 \frac{0.1}{0.1} =$	$\frac{2}{1}$
O	8	16	$\frac{8}{16} = 0.5, \frac{0.5}{0.2} =$	$\frac{5}{25}$

$\therefore$ Formula of oxide iodine will be I_2O_5

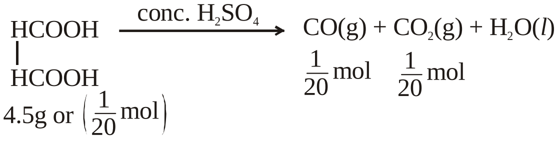
Q.14 (2)
The correct scientific notation of 0.0034 will be 3.4×10^{-3}
Thus , $0.0034 = 3.4\times10^{3}$ is incorrect

Q.15 (1)
Q.16 (1)

NEET/AIPMT

Q.1 (1)

$$\underset{2.3\text{g or }\left(\frac{1}{20}\text{mol}\right)}{HCOOH} \xrightarrow{\text{conc. } H_2SO_4} \underset{\frac{1}{20}\text{mol}}{CO(g)} + H_2O(l)$$



$$\underset{4.5\text{g or }\left(\frac{1}{20}\text{mol}\right)}{\begin{array}{l}HCOOH\\ |\\ HCOOH\end{array}} \xrightarrow{\text{conc. } H_2SO_4} \underset{\frac{1}{20}\text{mol}}{CO(g)} + \underset{\frac{1}{20}\text{mol}}{CO_2(g)} + H_2O(l)$$

Gaseous mixture formed is CO and CO_2 when it is passed through KOH, only CO is absorbed. So the remaining gas is CO.
So, weight of remaining gaseous product CO is

$$\frac{2}{20}\times 28 = 2.8\text{g}$$

So, the correct option is (1)

Q.2 (3)

(1) Moles of water = $\frac{0.00224}{22.4} = 10^{-4}$
Molecules of water = mole × $N_A = 10^{-4}\,N_A$

(2) Molecules of water = mole × $N_A = \frac{0.18}{18} N_A = 10^{-2}$ N_A
(3) Mass of water = 18 × 1 = 18 g
Molecules of water = mole × $N_A = \frac{18}{18} N_A = N_A$
(4) Molecules of water = mole × $N_A = 10^{-3}\,N_A$

Q.3 (3)
Formation of ammoina
$N_2 + 3H_2 \rightarrow 2NH_3$
2 mole of NH_3 is formed by 3 mole of H_2
20 mole of NH_3 is formed by 30 mole of H_2

Q.4 (3)

Q.5 (1)

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(q)} + H_2O_{(\ell)}$$

no of moles of $CaCO_3$ (pure) $= \frac{1}{2}\times$ mole of HCl
[Mole = molarity × volume (in ltr.)]

$$= \frac{1}{2}\times 0.5\times\frac{50}{1000} = 0.0125$$

weight of $CaCo_3$(pure) = mole × mol. wt
= 0.0125 × 100 = 1.25 g

$$\% = \frac{1.25}{\text{wt. of impure sample}}\times 100$$

$$\text{wt of impure sample} = \frac{1.25\times100}{95} = 1.32\text{g}$$

JEE MAIN

Q.1 (3)
Mass of $C_{15}H_{30}$=V×d=1000×0.756 g/ml =756 gm

Moles of $C_{15}H_{30} = \frac{75.6}{180+30} = \frac{75.6\text{g}}{210\text{g}} = 3.6$ moles of

$$\begin{array}{lllll} & C_{15}H_{30} & + \frac{45}{2}O_2 & \rightarrow 15CO_2(g) & + 15H_2O \\ \text{Moles} & 3.6 & \frac{45}{2}\times3.6 & 15\times3.6 & 15\times3.6 \end{array}$$

Mass $\quad O_2 \text{ mass} = \frac{45}{2}\times3.6\times32 = 2592\text{g of } O_2$
(mole ×mol. cot.) CO2 mass=15×44×3.6=2376 g of CO_2

Q. 2 [5418]
Molar massof $C_7H_5N_3O_6$ is 84 + 5 + 42 + 96 = 227

$$n_{C_7H_5N_3O_6} = \frac{681}{227} = 3 \text{ moles}$$

$$n_N = \frac{681}{227}\times 3 = 9 \text{ moles of N}$$

No. of N atoms = 9 × 6.02 × 10^{23}
= 5418 × 10^{21}
= 5418

Q.3 (25)

$$\% \text{ of glycine} = \left[\frac{75\times1}{GMM}\right]\times100 = 0.3$$

$$GMM = \left[\frac{75\times100}{0.3}\right] = 25\times10^3$$

Q.4 (4)

$$C_xH_y + (x+\frac{y}{4})O_2 \to xCO_2 + \frac{y}{2}H_2O$$

120 g   mass = 330 g   270 g

$$\text{moles} = \frac{330}{44} = 7.5 \quad \text{moles} = \frac{270}{18} = 15$$

Moles of CO_2 = moles of 'C'
⇒ so mass of carbon = n × Atomic mass
= 7.5 × 12
= 90 gram

$$\% \text{ of carbon in organic compound} = \frac{90}{120}\times100 = 75\%$$

$$\% \text{ of Hydrogen in organic compound} = \frac{30}{120}\times100$$

= 25 %

Q.5 (3)
Given HCl by mass = 35% = 0.35

$$M = \frac{10\times d}{\text{molecular weight}}\times100$$

$$M = \frac{10\times0.35\times1.46}{36.5}\times100$$

= 14 M

Q. 6 (225)
$\Rightarrow CH_4 + 2O_2 \to CO_3 + 2H_2O$
moles = 2.25   Mass = 81

$$\text{Moles} = \frac{81}{18}$$

⇒ 4.5

So, 225 × 10^{-2} moles of CH^4 are required to produce. 81 g fo water after complete combustion.

Q. 7 (143)

$CH_4 + 20_2(g) \to CO_2(g) + H_2O$

Mass 100   208   Moles ⇒ 3.25

$$\text{Moles} \quad \frac{100}{16} \quad \frac{208}{32} \quad \text{Mass} = 143 \text{ gram}$$

$$\Rightarrow 6.25 \quad \frac{6.5}{2} = 3.25$$

(L.R.)

$$n_{O_2} = \frac{n_{CO_2}}{1}$$

$$\frac{6.5}{2} = n_{CO_2}$$

$$\text{Mass of } CO_2 = \frac{6.5}{2}\times 44 = 143 \text{ gm}$$

Q.8 (2)
Q.9 (1)
Q.10 (2)
For

$$\% \text{ Hydrogen} = \frac{7.5}{116}\times100 = 6.5 : \text{ Relation atomicities}$$

$$\% \text{ Oxygen} = \frac{60}{116}\times100 = 51.7 \quad : H = 6.5$$

$$\% \text{ Carbon} = \frac{48.5}{116}\times100 = 41.8 \quad : O = \frac{51.7}{16} = 3.25$$

Empirical formula CH_2O; Weight = 30 g; $C = \frac{41.8}{12} = 3.5$

(A) $C_2H_4O_2$ (B) CH_2O relate to this formula.

Q. 11 (4)

C	74%	$\frac{74}{12} = 6.16$	$\frac{6.16}{1.23} = 5$
N	17.3%	$\frac{17.3}{14} = 1.23$	$\frac{1.23}{1.23} = 1$
H	8.7%	$\frac{8.7}{1} = 8.7$	$\frac{8.7}{1.23} = 7$

Emperical formula = C_5NH_7
Emperical weight = 81

$$\text{Multiplying factor} = \frac{162}{81} = 2$$

Molecular formula = $C_{10}N_2H_{14}$

Q.12 (3)
$4CO + Fe_3O_4 \to 3Fe + 4CO_2$
According to equation
1 mole (232 g) of Fe_3O_4 is reacting with
4 mole (112 g) of CO and
3 mole (168 g) of Fe and 4 mole (176 g) of CO_2 with produce.
According to question
4.640 × 10^3 g of Fe_3O_4 and 2.52 × 10^3 g of CO is reacting.
Since

232 g Fe_3O_4 needs 112 g CO to react

1 g Fe_3O_4 needs $\frac{112}{232\,g}$ CO to react

So 4.64×10^3 g Fe_3O_4 needs $\frac{112}{232} \times 4.640 \times 10^3$ CO to react

$= 2.24 \times 10^3$ g CO to react

$\Rightarrow$ That means Fe_3O_4 is limiting reagent (L.R)

$\Rightarrow$ 232 g Fe_3O_4 produce 168 g Fe

1 g Fe_3O_4 produce $\frac{168}{232}$ g Fe

4.640×10^3 Fe_3O_4 produce $\frac{168}{232} \times 4.640 \times 10^3$ Fe

= 3360 g

$\Rightarrow$ Hence option (3) is correct.

Q.13 (2)

$\Rightarrow \frac{0.02585 \times 0.112}{0.5702}$ Result should be of three significant figures.

$\Rightarrow \frac{0.00320096}{0.5702}$

$\Rightarrow$ 0.005613749....

Hint:- Least number of significant figure

Q.14 [46]

$$\underset{56\,L}{N_2(g)} + \underset{excess}{3H_2(g)} \rightarrow \underset{20\,lit}{2NH_3(g)}$$

Used N_2 = 10 lit, unused N_2 = 46 L

Q.15 (3)

$So_2Cl_2 + 2H_2O \rightarrow$ $H_2SO_4 + 2HCl$

a mole a mole 2a mole

(1) $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$

a mole 2a mole

Total mole of NaOH required = 4a = 16

a = 16

Q.16 [24]

$\because d_{solution} \approx 1$ gm/ml

Volume of solution = 2L = 2000 ml

Mass of solution = 2000 gm

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

$$48 = \frac{\text{mass of Mg}}{2000} \times 10^6$$

$48 \times 2 \times 10^{-3}$ = mass of Mg

Mole of Mg = $\frac{48 \times 2 \times 10^{-3}}{24} = 4 \times 10^{-3}$ mol

No. of Mg-atoms

= mole × N_A

$= 4 \times 10^{-3} \times 6.02 \times 10^{23}$

$= 24.08 \times 10^{20}$

$= x \times 10^{20}$

x = 24

Q.17 (2)

$C_6H_{12}O_6 \rightarrow$ Glucose

We know : $\frac{\text{mass of C}}{\text{mass of glucose}} = \frac{72}{180}$

Given : %C = 10.8 = $\frac{\text{mass of C}}{\text{mass of solution}} \times 100$

$\frac{10.8 \times 250}{100}$ = mass of C $\Rightarrow$ Mass of C = 27 g

$\therefore$ mass of glucose = 67.5 gm

$\therefore$ moles of glucose = 0.375 moles

Mass of solvent = 250 – 67.5 gm = 182.5 gm

$\therefore$ Molality = $\frac{0.375}{0.1825} = 2.055 \approx 2.06$

Q.18 [3]

$$CH_3CH_2MgBr + CH_3OH \longrightarrow \underset{2.24\,ml}{CH_3 - CH_3(g)}$$

Mole of ethane = $\frac{2.24}{22400}$

$= 10^{-4}$ mole

Mass of ethane = $10^{-4} \times 30$

$= 3 \times 10^{-3}$ gm

= 3 mg

Q.19 [8]

Mass of mixture = 10 g = H_2 + He

Volume of container = 0.0125m^3

= 12.5 *l*

P = 6 bar = 5.922 atm

T = 27°C = 300 K

PV = nRT

Total mole (n) =

$$\frac{PV}{RT} = \frac{5.922 \times 12.5 \times 101.33}{8.314 \times 300}$$

$\because$ 1 *l*-atm = 101.33 J

$= 3.01 \approx 3$ mol

Let mass of H_2 in mixture = x gm

Mass of He = (10 - x) gm

Mole of H_2 = $\frac{x}{2}$

Mole of He = $\frac{10 - x}{4}$

Total mole $= \frac{x}{2} + \frac{10-x}{4} = 3$

$2x + 10 - x = 12$

$X = 2\text{ gm}$

Mass of $H_2 = 2g$

Mass of He $= 10 - x$

$= 8\text{ gm}$

Q.20 (3)

% of Iron in hemoglobin $= 0.34\%$

Given mass of hemoglobin = 3.3g

Mass of Iron in hemoglobin $= 3.3 \times \frac{0.34}{100}$

$= 0.01122\text{ gm}$

No. of Iron atoms $= \frac{\text{mass}}{\text{GAM}} \times N_A$

$= \frac{0.01122}{56} \times 6.023 \times 10^{23}$

$= 1.206 \times 10^{20}$

$\approx 1.21 \times 10^{20}$

Q.21 (2)

	x +	y + 3z	$\rightleftharpoons$	xyz_3
n		1 1		0.05
% mass		10 20		30

Limiting reagent = Z

Moles of product formed $= \frac{1}{3} \times 0.05$

$= \frac{0.05}{3}$ Moles

Amount = Moles × Molecular mass

$= \frac{0.05}{3} \times (10 + 20 + 3 \times 30)$

$\Rightarrow \frac{0.05}{3} \times 120 = 2\text{ g}$

Q.22 (3)

$N_2(g) \quad + \quad 3H_2(g) \rightleftharpoons 2NH_3(g)$

$W_2 = 20g \quad 5$

$n = \frac{20}{28} \qquad 5/2$

Stoichiometric Amount

$N_2 \rightarrow \frac{20/28}{1} = \frac{20}{28} H_2 \rightarrow \frac{5/2}{3} = \frac{5}{6}$

$\therefore$ N_2 is the limiting reagent

$\therefore n(NH_3) = 2 \times n(N_2) = 2 \times \frac{20}{28} = 1.42$

Q.23 (25)

$H_2SO_4 + 2NaOH \rightarrow \quad Na_2SO_4 \quad + \quad 2H_2O$

$M = 0.2 \qquad M = 0.1$

Moles = 0.1

$V = 2\,L \qquad V = 2\,L$

$V = 4\,L$

$n = M \times V \quad n = M \times V$

$m = \frac{0.1}{4} \times 1000$

$= 0.2 \times 2 \qquad = 0.1 \times 2$ **m = 25 Ans.**

$= 0.4 \qquad = 0.2$

L.R. $\frac{0.4}{1} \qquad \frac{0.2}{2} = 0.1$

L.R. × NaOH

Q.24 (6)

Mole of H_2 gas $= \frac{1.344}{22400} = 6 \times 10^{-5}$

No. of H.atoms per molecule of $H_2 = 2$

Moles of alcoholic hydrogens $= 6 \times 10^{-5} \times 2$

$n \times \frac{1.84 \times 10^{-3}}{92} = 2 \times 6 \times 10^{-5}$

$n = \frac{12 \times 92}{184} \Rightarrow 6$

Q.25 (3)

$4HNO_3(l) + 3KCl(s) \rightarrow Cl_2(g) + NOCl(g)$
$+ 2H_2O(g) + 3KNO_3$

Moles of $HNO_3 = \frac{4}{3} \times \frac{110}{101}$

mass = 110

Mass of $HNO_3 = \frac{4}{3} \times \frac{110}{101} \times 63$

Molecular mass =

$39 \times 1 + 14 \times 1 + 16 \times 3$

$= 91.485\text{ g}$

$= 39 + 14 + 48$

$= 39 + 62$

$= 101$

Moles $= \frac{110}{101}$

Structure of Atom

EXERCISE-I (MHT CET LEVEL)

Q.1 (3)

Q.2 (3)

$$\lambda = \frac{h}{p} = \frac{6.6\times10^{-34}}{3.3\times10^{-24}} = 2\times10^{-10}\,m = 2\overset{0}{A}$$

Q.3 (1)

Radius of orbit = $\frac{n^2 a_o}{z}$ (a_o = 0.529 Å)

Radius of H = $\frac{(1)^2\times0.529Å}{1} = 0.53Å$

Thus, the radius of $_3Li^{2+}$ will be :

$= \frac{(1)^2\times0.529}{3} = 0.17Å$

Q.4 (4)

Radius of hydrogen atom $= 0.530Å$, Number of excited state (n)=2 and atomic number of hydrogen atom (Z)=1. We know that the Bohr radius.

$$(r) = \frac{n^2}{Z}\times \text{ Radius of atom } = \frac{(2)^2}{1}\times0.530$$

$$= 4\times0.530 = 2.12Å$$

Q.5 (4)

For H atom, $E_n = -\frac{13.6Z^2}{n^2}eV$

For second orbit, = 2

Z = At. no. = 1(for hydrogen)

$$\therefore\ E_2 = -\frac{13.6\times(1)^2}{(2)^2} = \frac{-13.6}{4}eV$$

$$= -\frac{13.6\times1.6\times10^{-19}}{4}J$$

$= -5.44\times10^{-19}$ J

Q.6 (1)

Q.7 (1)

Q.8 (4)

For Balmer $n_1 = 2$ and $n_2 = 3$;

$$v = R\left(\frac{1}{2^2}-\frac{1}{3_2}\right) = \frac{5R}{36}cm^{-1}$$

Q.9 (1)

Q.10 (3)

Q.11 (1)

Q.12 (4)

Q.13 (4)

When = –3, ℓ = 3, $\therefore$ n = 4.

Q.14 (1)

For a given orbital with principal quantum number (n) and azimuthal quantum number (ℓ) number of radial nodes = (n – ℓ – 1)

for 3s orbital: n = 3 and ℓ = 0

therefore, number of radial nodes = 3 – 0 – 1 = 2

for 2p orbital: n = 2 and ℓ = 1

therefore, number of radial nodes = 2 – 1 – 1 = 0

Q.15 (3)

For 4*p* electron n = 4, 1 = 1, m =–1, 0, + 1 and s = $+\frac{1}{2}$or$-\frac{1}{2}$

Q.16 (3)

Q.17 (2)

Q.18 (3)

Q.19 (1)

Q.20 (3)

Q.21 (3)

Q.22 (3)

Q.23 (3)

Q.24 (1)

Q.25 (4)

Q.26 (3)

Q.27 (2)

Q.28 (3)

Q.29 (3)

Q.30 (3)

Q.31 (2)

Q.32 (3)

Q.33 (4)

EXERCISE-II (NEET LEVEL)

Q.1 (4)

Neutrons and protons in the nucleus and electrons in the extranuclear region.

Q.2 (3)

Proton is the nucleus of $H-$ atom ($H-$ atom devoid of its electron).

Q.3 (2)
Cathode rays are made up of negatively charged particles (electrons, e^-)

Q.4 (1)
Discoverer of Anode Rays – E. Goldstein

Q.5 (1)
Charge on electron = 1.6×10^{-19}C

Q.6 (3)
When $c = v \times \lambda$ than $\lambda = \frac{c}{v} = \frac{3\times10^8}{2\times10^6} = 1.5\times10^2 m$

Q.7 (2)
According to quantum theory of radiation, a hot body emits radiant energy not continuously but discontinuously in the form of small packets of energy called quanta or photons.

Q.8 (2)
$\lambda = \frac{h}{p}$ or $\frac{h}{mv}$ or $\frac{h}{mc}$ de-Broglie equation.

Q.9 (1)

T.E. = – I. E. =

Q.10 (1)
Energy is released when electron jumps from a higher energy level to a low energy level.

Q.11 (4)
It represents Heisenberg's uncertainty principle.

Q.12 (1)
According to Hydrogen spectrum series.

Q.13 (3)
$\lambda = \frac{h}{mv}$

Q.14 (3)
$\lambda = h/mv$

Q.15 (3)
According to de-Broglie equation $\lambda = \frac{h}{mv}$ or $\frac{h}{p}$ or $\frac{h}{mc}$.

Q.16 (1)
Debroglie predicted the wave nature.

Q.17 (2)
Hunds rule of maximum multiplicity

Q.18 (4)
Acording to paulis no two electron in an orbit will have all quantum number to be same.

Q.19 (2)
Hund's rule states that pairing of electrons in the orbitals of a subshell (orbitals of equal energy) starts when each of them is singly filled.

Q.20 (2)
Orbital angular momentum = $\frac{h}{2\pi}\sqrt{\ell(\ell+1)}$

Q.21 (2)
3d subshell filled with 5 electrons (half-filled) is more stable than that filled with 4 electrons. $1,4s$ electrons jumps into $3d$ subshell for more sability.

Q.22 (4)
When $n = 3$ shell, the orbitals are $n^2 = 3^2 = 9$.
No. of electrons $= 2n^2$
Hence no. of orbital $= \frac{2n^2}{2} = n^2$.

Q.23 (2)
$N_7^{14} = 1s^2\,2s^2\,2p_x^1\,2p_y^1\,2p_z^1$.

Q.24 (2)

$$O_8 = \overset{1s^2}{\boxed{\uparrow\downarrow}}\ \ \overset{2s^2}{\boxed{\uparrow\downarrow}}\ \ \overset{2p^4}{\boxed{\uparrow\downarrow}\underbrace{\boxed{\uparrow}\boxed{\uparrow}}_{\text{Unpaired electron}}}$$

Q.25 (1)
Aufbau state that electron are first filled in those subshell of lower energy.

EXERCISE-III (JEE MAIN LEVEL)

Q.1 (1)
Hydrogen atom contains 1 proton, 1 electron and no neutrons.

Q.2 (4)
$\frac{(e/m)_e}{(e/m)_\alpha} = \frac{e/m_e}{2e/4\times1836\,m_e} = \frac{3672}{1}$

Q.3 (1)
Volume fraction = $\frac{\text{Volume of nucleus}}{\text{Total vol. of atom}}$ = $\frac{(4/3)\pi\,(10^{-13})^3}{(4/3)\pi\,(10^{-8})^3} = 10^{-15}$

Q.4 (1)
Ne contains 10 electrons
O^{2-} and F^- contain 10 electrons

Q.5 (1)
$r\,\alpha\left(\frac{n^2}{Z}\right)$ As Z increases, radius of I orbit decreases.

Q.6 (2)

Given :

$P = 1\ kW$

$P = 1 \times 10^3$ watt

$E = 10^3$ J/S in one sec

$v = 880$ Hz

$\therefore E = nhv$

$\Rightarrow 10^3\ x \times 6.626 \times 10^{-34} \times 880 \Rightarrow x = 1.71 \times 10^{33}$ in one sec

Q.7 (3)

$$E = \frac{hc}{\lambda}$$

$$= \frac{1240}{31} = 40$$

K.E. max = 40 – 12.8

= 27.2 eV

$$\frac{1}{2}mv^2 = 27.2 \times 1.6 \times 10^{-19}$$

$$V^2 = \frac{54.4 \times 1.6 \times 10^{-19}}{9.1 \times 10^{-31}}$$

$V^2 = 9.56 \times 10^{12}$

$V = 3.09 \times 10^6$ m/sec.

Q.8 (4)

Photoelectric effect is a random phenomena. So, electron It may come out with a kinetic energy less than (hν – w) as some energy is lost while escaping out.

Q.9 (2)

We know that, for wave no.

$$\bar{v}_3 = \bar{v}_1 + \bar{v}_2$$

$$\frac{1}{\lambda_3} = \frac{1}{\lambda_2} + \frac{1}{\lambda_1}$$

Q.10 (2)

$$r = 0.529 \times \frac{n^2}{Z}\ \text{Å} = 0.529 \times \frac{1^2}{1}\ \text{Å} = 0.529 \times 10^{-10}\ m = 0.529 \times 10^{-8}\ cm$$

Q.11 (4)

$$\frac{r_3 = 0.529 \times 3^2/Z}{r_1 = 0.529 \times 1^2/Z}$$

$\therefore r_3 = 9r_1$

Q.12 (4)

$r = \infty$

Q.13 (4)

$$\lambda = \frac{hc}{\Delta E}\ \therefore \lambda \alpha \frac{1}{\Delta E}$$

Q.14 (2)

Balmer means transition

to n = 2

1. line → 3 to 2
2. line → 4 to 2
3. line → 5 to 2

Q.15 (2)

$$\frac{1}{\lambda} = 109677 \times 9$$

$\lambda = 1.01 \times 10^{-6}$ cm

$$E = \frac{6.62 \times 10^{-34} \times 3 \times 10^8}{1.01 \times 10^{-8}}$$

$= 19.66 \times 10^{-18}$ J

$$\Rightarrow \frac{1}{2}mv^2 = 19.66 \times 10^{-18}$$

$$V^2 = \frac{39.32 \times 10^{-18}}{9.1 \times 10^{-31}}$$

$V^2 = 4.32 \times 10^{13}$

$V^2 = 43.2 \times 10^{12}$

$V = 6.57 \times 10^6$

$$\lambda_{debroglie} = \frac{6.62 \times 10^{-34}}{9.1 \times 10^{-31} \times 6.57 \times 10^6}$$

$= 1.17 \times 10^{-10}$ m

= 1.17 Å

Q.16 (3)

Visible lines ⇒ Balmer series (5 → 2, 4 → 2, 3 → 2). So, 3 lines.

Q.17 (4)

$$\frac{1}{\lambda} = R_H \times 4 \left\{\frac{1}{9} - \frac{1}{16}\right\}$$

$$\frac{1}{m} = R_H \times \frac{7}{36}$$

$$\frac{1}{\lambda_{required}} = \frac{36}{7m} \times 16 \left\{\frac{1}{9}\right\}$$

$$\frac{1}{\lambda_{required}} = \frac{36}{7m} \times 16 \left\{\frac{1}{9}\right\}$$

$$\lambda_{req.} = \frac{7m}{64}$$

Q.18 (3)

For an α particle, $\lambda = \dfrac{0.101}{\sqrt{V}}$ Å.

Q.19 (3)

$$\lambda = \frac{h}{mv} = \frac{6.625\times10^{-34}}{0.2\times5}\times3600 \approx 10^{-30}\,\text{m}.$$

Q.20 (1)

Mass of α particle = 4 (mass of proton)

Mass of proton = 1840 (mass of e^-)

Let

Mass of e^- = m

$\therefore$ Mass of p^+ = 1840 m

and mass of α particle = 7360 m

$$\frac{1}{2}\,meV^2 = 16\,E$$

$$V_e^2 = \frac{32E}{m}$$

$$V_p^2 = \frac{8E}{1840m}$$

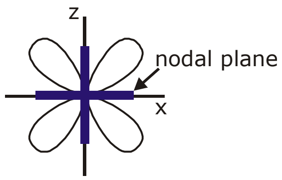
$$V_\alpha^2 = \frac{E}{7360m}$$

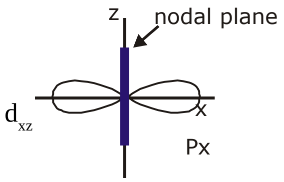
Q.21 (1)

5-fold degenerate

All d-orbitals are of same energy.

Q.22 (3)





Q.23 (2)

Na^+, Co^{+2}, Cr^{2+}, Fe^{+3}

M. M. (μ) = $\sqrt{n(n+2)}$

we get Na^+, Co^{+2}, Cr^{2+}, Fe^{+3}

Q.24 (3)

$n\ell$ no. of e^-

3s $\ell = 0 \rightarrow s$

Q.25 (1)

${}_{30}Zn^{2+} \rightarrow 3d^{10}\,4s^0$

no. of unpaired e^- = 0

EXERCISE-IV

Q.1 6

Q.2 22

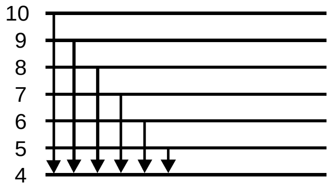
One molecule of CO_2 have 22 electrons.

Q.3 0014

Radius $= 0.529\,\dfrac{n^2}{Z}$ Å $= 10 \times 10^{-9}$ m

So, $n^2 = 189$ or, $n \approx 14$ **Ans.**

Q.4 0006



First Excited level = 2
∴ ninth level = 10

Total line = 6

Q.5 0003

Visible lines ⇒ Balmer series (5 → 2, 4 → 2, 3 → 2). So, 3 lines.

Q.6 0006

infrared lines = total lines – visible lines – UV lines

$$= \frac{6(6-1)}{2} - 4 - 5 = 15 - 9 = 6.$$

(visible lines = 4 6→2, 5→2, 4→2, 3→2)

(UV lines = 5 6→1, 5→1, 4→1, 3→1, 2→1)

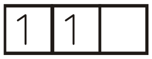
Q.7 0005

Q.8 0014

Maximum no. of electrons in a subshell $= 2(2l+1)$ for f-subshell, l = 3 so 14 electrons accommodated in f -subshell.

Q.9 0002

${}_6C \rightarrow 1s^2\,2s^2\,2p^2$



Q.10 0026

$\sqrt{n(n+2)} = \sqrt{35}$

$\therefore n = 5$

$x^{3+} \rightarrow 4s^o\,3d^5$

$x \rightarrow 4s^2\,3d^6$

i.e., ${}_{26}Fe$

Q.11 (3)

Q.12 (1)

Q.13 (1)
Q.14 (2)
Q.15 (3)
Q.16 (1)

PREVIOUS YEAR'S

MHT

Q.1 (1)

Q.2 (3)

Q.3 (3)

Q.4 (4)

$^{14}_{6}C$ and $^{16}_{8}O$ have same number of neutrons i.e.8. Therefore , they are the examples of isotones .

Q.5 (2)

Balmer series of transitions in the specturm of hydrogen atom fall in visible region Lyman series falls in ultaviolet region while paschen and Brackett series fall in infra red region.

Q.6 (2)

The vaule of Rydberg constant is 2.18×10^{-18} J.

Q.7 (1)

The 3d - orbitals of Cr and Cu have half - filled and completely filled electronic configuration respectively which are the most stable elemctronic configurations Thus, these both elements do not follow Aufbau priciple according to which 4s orbial should be filled first with 2 electrons as it have lower energy than 3d-orbital . So their electrons configuration are

Cr (Z = 24) = $1s^2 2s^2 2p^3 3s^2 3p^6 3d^5 4s^1$
Cr (Z = 29) = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$

Q.8 (2)

Hund' s rule represent the pairing of electrons in a subshell after each orbital is filled with one electron Thus this rule states that during filling of electrone in a subshell pairing of electrons cannot take palce until there is no empty orbital avaliable This rule is also called Hund ,S rule of maxium multiplicity

Q.9 (2)

Energy of an electron is given by

$$E_n = \frac{-13.6}{n^2} \text{ev}$$

For n = 2

$$E = \frac{-13.6}{2 \times 2} \text{eV} = \frac{-13.6}{4} \text{eV}$$

$$[1\text{ev} = 1.602 \times 10^{-19} \text{J}]$$

$$= -3.4 \times 1.602 \times 10^{-19} \text{ J}$$
$$= -5.45 \times 10^{-19} \text{ J}$$
$$= -0.545 \times 10^{-18} \text{J}$$

Q.10 (4)

Q.11 (4)

NEET/AIPMT

Q.1 (1)

According to Hund's Rule of maximum multiplicity, the correct electronic configuration of N-atom is

↑↓ ↑↓ ↑ ↑ ↑
$1s^2$ $2s^2$ $2p^3$

OR

↑↓ ↑↓ ↓ ↓ ↓
$1s^2$ $2s^2$ $2p^3$

∴ Option (1) violates Hund's Rule.

Q.2 (1)

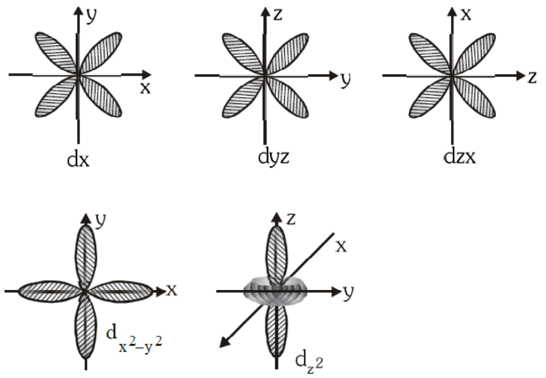
Energy α value of (n+l)
5f 5 + 3 = 8
6p 6 + 1 = 7
4d 4 + 2 = 6
5p 5 + 1 = 6

Q.3 (2)

1st four line of Balmer series of spectrum of hydrogen atom falls in visible region.

Q.4 (4)

Q.5 (3)



JEE MAIN

Q.1 (2)

(A) (B)–3p (a)–3d
(B) (a) –3d (a)–3d
(C) (a)–4p (a)–3d

Q. 2 [4]

Ele. conf. of Li = $1s^2 2s^1$

$$\left(E_{Li}\right)_{n=2} = \left(E_H\right)\frac{z^2}{n^2} = -2.2 \times 10^{-18} \times \frac{9}{4} \text{J}$$

$$E = \left(\frac{hc}{\lambda}\right) = 2.2\times10^{-18}\times\frac{9}{4}J$$

$$\frac{6.63\times10^{-34}\times3\times10^{8}}{\lambda} = 2.2\times10^{-18}\times\frac{9}{4}J$$

$\lambda = 4\times10^{-18}$m
= 4

Q.3 (3)
$E = E_0 + KE$
For minimum energy, $E = E_0$

$$E = \left(\frac{hc}{\lambda}\right) = hv$$

$= 6.6\times10^{-34}\times1.3\times10^{15}$
$= 8.58\times10^{-19}$ J

Q.4 (3)
We know:

$$E = \frac{hc}{\lambda}$$

$$\Rightarrow \frac{6.63\times10^{-34}\times3\times10^{8}}{300\times10^{-9}}$$

$\Rightarrow 6.63\times10^{-19}$ J
Energy of 1 mole of photons
$\Rightarrow 6.63\times10^{-19}\times6.02\times10^{23}\ mol^{-1}$
$\Rightarrow$ 399 kJ / mol

Q.5 (1)
4d orbital
Angular nodes $\Rightarrow \ell = 2$
Radial nodes = $(n - \ell - 1) \Rightarrow 4 - 2 - 1 = 1$

Q.6 (22)

Q.7 (4)

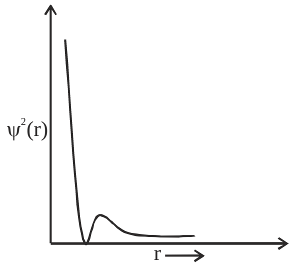
Q.8 (300)

Q.9 (2)
For possible set $n > \ell \leftarrow$ This condition should be exist.
So, here is 2.

Q.10 (3)
(a) Number of values of n = 1, 2, 3, ...∞
(b) Number of values of ℓ = 0 to (n – 1)
(c) Number of values of m = – ℓ to + ℓ
Total values = $2\ell + 1$
(d) Values of spin = $\pm\frac{1}{2}$
(e) For ℓ = 5 number of orbitals = $2\ell + 1 = 11$

Q.11 (4)
Option (1), (2) and (3) are correct.
Hence option (4) is correct

Q.12 (2)
ONE radial node is there



Radial Node $\Rightarrow n - l - 1$
$\Rightarrow 2 - 0 - 1$
$\Rightarrow 1$ radial Node
1 = 0 zero angular Node

Q.13 (2)

$$r = \frac{n^2}{z}$$

for H-atom, we know =1
So, $r \propto n^2$
$r_3 \propto 9$
$r_4 \propto 16$

$$\text{Ratio} = \frac{r_3}{r_4} = \frac{9}{16}$$

$$r_4 = \frac{16}{9} r_3$$

Q.14 (4)
O^{2-}, Mg^{2+} and Al^{3+} have 10 electronic so they are isoelectronic

Q.15 [1758]

$$\lambda_{e^-} = \frac{h}{m_e v_n} \lambda_n = \frac{h}{m_n v_n}$$

If $V_e = xV_n$
then $\lambda_e = \lambda_n$

$$\frac{h}{m_e v_e} = \frac{h}{m_n v_n}$$

$$\frac{h}{m_e \times x v_n} = \frac{h}{m_n v_n}$$

$$X = \frac{m_n}{m_e} = \frac{1.6\times10^{-27}}{9.1\times10^{-31}}$$

$= 175.82$

≈ 176

Q.16 [4]

Number of emission line

$$= \frac{n(n-1)}{2} = \frac{5\times(5-1)}{2} = 10$$

Q.17 (3)

For n = 3 $\Rightarrow$ only 3s, 3p & 3d

$\downarrow$

$\ell = 0 \;\; \ell = 1 \;\; \ell = 2$

$\ell = 3$ not possible

Q.18 (1)

(A) Energy of 2s orbital H atom > Li atom

(R) Energy of orbital in semi sub shell dec. with line in atomic number

Q.19 [4]

$${}^{48}_{22}x^{3-} \Rightarrow \begin{array}{l} \text{no. of } e^{-} = 22+3 = 25 \\ \text{no. of } n^{0} = 48-22 = 26 \end{array}$$

No. of neutron = 9% of no of e^- + no. of e^-

$$26 = \left(\frac{a}{100}\times25\right)+25$$

$$1 = \frac{a}{4}$$

$a = 4$

Q.20 (1)

Energy of orbital α value of (n + I)

When same value of (n + I) then

Energy of orbital α value of n

	n		ℓ		$(n+\ell)$
[A]	3	+	0	=	3
[B]	4	+	0	=	4
[C]	3	+	1	=	4
[D]	3	+	2	=	5

Q.21 (4)

Orbital is represented by n, l, m.

Q.22 (548)

$$\Delta x \cdot \Delta P \ge \frac{h}{4\pi}$$

$\Delta x = 2a_0$ $\qquad$ $\Delta P = m\Delta v$

$a_0 = 52.9 \text{ pm} = 52.9\times10^{-12}$ m

$$\Delta x \cdot m\Delta v = \frac{h}{4\pi}$$

$$\Delta x = \frac{h}{4\pi\times m\times\Delta x} = \frac{6.63\times10^{-34}}{4\times3.14\times9.1\times10^{-31}\times2\times52.9\times10^{-12}}$$

= 548273 m/sec

= 548.273 km/sec

$\approx$ 548 km/sec

Q.23 (2)

Greater the value of (n + l) grater is energy.

Classification of Elements and Periodicity in Properties

EXERCISE-I (MHT CET LEVEL)

Q.1 (3)

Q.2 (3)

Q.3 (3)
It is factual.

Q.4 (3)
According to Aufbau '4s' is lower in energy then 3d.

Q.5 (3)
It is fact.

Q.6 **(3)**

Q.7 **(3)**

Q.8 (2)

$$_{11}\mathrm{Na} \Rightarrow 2,8,1;\ _{17}\mathrm{CI} \Rightarrow 2,8,7$$

These have same number of shells. Hence, they are the elements of the same period.

Q.9 (2)

Q.10 (1)
F is most electronegative.

Q.11 (4)
Except Li–Na, all have diagonal relationship.

Q.12 **(2)**

Q.13 **(1)**

Q.14 (1)
The ionic radii follows the order

$$O^2 > F^- > Li^+ > B^{3+}$$

Q.15 (4)

Q.16 (2)
Lower is ionisation energy more easily cation is formed.

Q.17 (1)
$BI_{HCl} = r_H + r_{Cl} - .09\,|X_H - X_{Cl}|$
$= 5 + 1 - 0.9\,(0.9) = 5.919$ Å

Q.18 (2)

Q.19 (2)
Effective nuclear charge (i.e. Z/e ratio) decreases from F^- to N^3 hence the radii follows the order : $F^- < O^2 < N3^-$. Z/c for $F^- = \frac{9}{10} = 0.9$ for

$$O^2 = \frac{8}{10} = .8 \text{ for } N^{3-} = 0.7$$

Q.20 (4)
The screening effect of inner electron of the nucleus causes the decrease in ionization potential, therefore the order of the screening effect is

$$f < d < p < s$$

Hence, the screening effect of d -electron is less than p -electron.

Q.21 (1)
Isoelectronic species have same number of electrons.

Ions	O^{2-}	F^-	Na^+	Mg^{2+}
8+29 + 1	11 - 1	12 - 2		
No. of e^- = 10		10	10	10

$\therefore$ O^{2-}, F^-, Na^+, Mg^{+2} are isoelectronic

Q.22 (4)

Q.23 (3)
1E increases from left to right in the period.

Q.24 (2)
From data it can loose 4 electron easily.

Q.25 (1)

Q.26 (3)
IE_2(Na) > IE_2(Mg
because second electron is Na is removed from noble gas configuration.

Q.27 (2)
First ionisation energy of Pb is higher than Sn. It is due to enhanced nuclear charge in Pb because of poor shielding of d- and f-electrons. Hence IE_1 of Sn is less than Pb.

Q.28 (2)

Q.29 (2)

Q.30 (3)

Q.31 (2)
Closed shell (Ne), half filled (P) and completely filled comfiguration (Mg) are the cause of higher value of I.E.

Q.32 (4)
The smaller the atomic size, larger is the value of ionisation potential. Further the atoms having half filled or fully filled orbitals are comparatively more stable, hence more energy is required to remove the electron from such atoms.

Q.33 (3)
Element with atomic number 9 has the smallest size and highest ionization enthalpy
Fluorine– Atomic Number 9

Q.34 (2)

Q.35 (4)

Increase in mass is more withy respect to volume : $\left(\frac{\text{Mass}}{\text{Volume}}\right)$ ratio increases.

Q.36 (1)

F is most electronegative.

Q.37 (3)

Electron affinity is the measure of the ease with which an atom receives the additional electron in its valence shell in gaseous phase.

Generally down the group, the electron affinity decreases due to increase in atomic size.

Q.38 (2)

Q.39 (1)

As size of atom decreases across the period, the attraction between the nucleus and shared pair of electrons increases. So electronegativity increases across the period.

Q.40 (3)

In half filled orbital it is difficult to give electron.

Q.41 (1)

Q.42 (4)

Due to strong repulsion second EA is endothermic.

Q.43 (3)

Q.44 (2)

Q.45 (2)

More the oxidation state of the central atom (metal) more is its acidity. Hence SeO2 (O.S. of Se = +4) is acidic. Further for a given O.S., the basic character of the oxides increases with the increasing size of the central with the increasing size of the central atom. Thus Al_2O_3 and Sb_2O_3 are amphoteric and Bi_2O_3 is basic.

Q.46 (3)

EXERCISE-II (NEET LEVEL)

Q.1 (1)

Q.2 (3)

Q.3 (1)

$[Kr]^{36}\ 5s^1$

Last electron enters in s-subshell, so it belongs to s-block.

For s-block group number = number of valence electron(s).

Q.4 (4)

Q.5 (1)

Al(13) and Ga(31) both have same valence shell electron configuration i.e. $ns^2\ np^1$. Hence belongs to 13th group.

Q.6 (3)

Q.7 (2)

If n = 7 then it will be 5f or actinide series.

Q.8 (4)

The d-sub shells are not filled with electrons monotonically with increase in atomic number. There are some exceptions like Cr, Cu etc.

Q.9 (4)

Q.10 (1)

Q.11 (3)

$$\text{Elements in 5}^{\text{th}}\text{ period} = \left(\frac{5+1}{2}\right)^2 = 18$$

$$\text{Elements in 6}^{\text{th}}\text{ period} = \left(\frac{6+2}{2}\right)^2 = 32$$

Q.12 (2)

$15 = [Ne]\ 3s^2\ ,\ 3p^3$

$33 = [Ar]\ 4s^2,\ 3d^{10},\ 4p^3$

$51 = [Kr]\ 5s^2,\ 4d^{10},\ 5p^3$

So all belongs to nitrogen family as outermost configuration is $ns^2\ np^3$ or belongs to 15th group.

Q.13 (3)

n represent period & total electrons in outermost shell represent group.

Q.14 (2)

Alkali metals are most electropositive & among alkali metals Cs is most.

Q.15 (1)

Density increases across the period as volume decreases and atomic weight increases, and generally increase down the group due to increase in atomic weight as compared to atomic volume.

Q.16 (3)

Ne is bigger than oxygen due to interelectronic repulsion on account of completely filled 2p subshell.

Q.17 (3)

$$\text{Radius of ion} \propto \frac{1}{\text{oxidation state}}$$

Q.18 (2)

O^+ is smaller than parent atom while anion is bigger than parent atom. O^{2-} and N^{3-} are isoelectronic species.

So ionic size $\propto \frac{1}{\text{nuclear charge}}$. Hence the correct order is $O^+ < O^{2-} < N^{3-}$.

Q.19 (c)

Amongst isoelectronic species, ionic radii of anion is more than that of cations. Further size of anion increase with increase in -ve charge and size of cation decrease with in crease in + ve charge. Hence ionic radii decreases from O^- to Al^{3+}.

Q.20 (1)

Due to lanthamide contraction size of 4d $\approx$ 5d elements.

Q.21 (2)

For isoelectronic species ionic radius $\propto \dfrac{1}{\text{nuclear charge}}$ so incorrect order is $N^{3-} < N^{5+}$.

Q.22 (1)

In noble gas vander walls radii is calculated which is normally then double of other type of radii.

Q.23 (1)

As we go from left to right in a period radius decreases.

Q.24 (4)

Down the group radius increases.

Q.25 (1)

Q.26 (2)

Among isoelectronic ions lesser is $\frac{z}{e}$, more is radius.

	O^{2-}	Mg^{2+}	Al^{+3}
$\frac{z}{e}$ =	$\frac{8}{10}$	$\frac{12}{10}$	$\frac{13}{10}$

Q.27 (3)

Element	no. of unpaired electron
Mg	O
Al	1
O	2
N	3

Q.28 (1)

Q.29 (2)

Fe, Co, Ni have nearly same atomic radii on account of cumulate effect of increased nuclear charge and shielding effect across the period. (Electrons are filled in d-subshell which has poor shielding effect).

Q.30 (1)

In isoelectronic anion is bigger in size.

Q.31 (2)

Q.32 (4)

Q.33 (3)

Second 1E of O > F

Q.34 (1)

Depends upon the different energies of the various sub-shells for the same energy level, e.g., electrons in s-subshell will have lower energy and thus will be closer to the nucleus and will have high penetration power while p-subshell electrons will penetrate the electron cloud to lesser extent and so on.

Q.35 (2)

Rare gases; as the e^- is to be removed from stable configuration.

Q.36 (2)

The right sequence of $I.E_1$ of Li < B < Be < C.

Q.37 (3)

Al^+ is more stable as second electron is removed from fully filled orbital.

Q.38 (1)

As elements are ionized of the proton to electron ratio increases, so the attraction between valence shell electron and nucleus increases and as a result the size decreases. Therefore, the removal of electron from smaller cation requires higher energy. Hence the second ionisation energy is higher than its first ionisation energy.

Q.39 (3)

(1) Down the group size increases and therefore, ionization energies decrease. Hence the order is correct.

(2) Cation is smaller but anion is bigger than its parent atom. As charge on cation increases the ionic radius decreases. Similarly as charge on anion increases the ionic radius increases. IE $\propto \dfrac{1}{\text{atomic / ionic radius}}$ and $IE_1 < IE_2 < IE_3$. Hence the order is correct.

(3) N has stable half filled configuration thus has higher ionization energy than O. Hence the order is correct.

(4) The correct order is Mg > Al < Si. Mg (ns^2) has higher ionization energy than Al ($ns^2\,np^1$) because s-sub shell electrons have higher penetration power than that of p-sub shell electrons. Further across the period size decreases and nuclear charge increases and therefore, ionization energy increases.

Q.40 (2)

$Mg \longrightarrow Mg^+ + e^- \qquad 1E_1 = 178$

$Mg^+ \longrightarrow Mg^{2+} + e^- \qquad 1E_2 = 348$

$Mg \longrightarrow Mg^{2+} + e^- \qquad$ so $\quad 1E = 1E_1 + 1E_2$

$= 178 + 348 = 526$

Q.41 (2)

1E increases from left to right along the period.

Q.42 (4)

From data it seems that it can loose two electrons easily.

Q.43 (1)

It can loose 2 electron easily.

Q.44 (3)

1E increases along the period.

Q.45 (1)

It is difficult to remove e^- from positive ion.

Q.46 (4)

1E of anion < 1E of atom.

Q.47 (3)

Removal of Ist electron is easier because of bigger size but 2nd electron is to be removed from $ns^2\ np^6$ configuration i.e. stable noble gas configuration. So IE_2 is quite higher.

Q.48 (4)

Second ionisation energy of sodium is greater than that of Mg. In case of sodium ion (i.e. Na^+) the electron removal from the stable inert gas configuration ($1s^2\ 2s^2\ 2p^6$) requires much higher energy.

Q.49 (1)

Q.50 (4)

(1) As electronegativity increases the non-metallic character increases as tendency to form anion increases.

(2) It is based on their SRP values. (Oxidising power may be cumulative effect of hydration energies, electronegativities, bond dissociation energies and electron gain enthalpies).

(3) C = – 121 ; Si = – 135 ; P = – 60 ; N = + 31 (all values are in KJ /mole) . It depends on various factors like size of atom, nuclear charge, partially filled, half filled and completely filled electronic configurations.

Q.51 (2)

Cl have more electron gain enthalpy than F

{due to smaller size of F and 'O' it will show repulsion}

S have more electron gain enthalpy than O

$\therefore\ \boxed{\text{Order will be O} < \text{S} < \text{F} < \text{Cl}}$

Q.52 (2)

Electronic configuration of element with atomic number 118 will be $[Rn]5f^{14}6d^{10}7s^27p^6$. Since its elctronic configuration in the outer most orbit (ns^2np^6) resemble with that of inert or noble gases, theregore it will be noble gas elemennt.

Q.53 (1)

On descending a group, the atoms and ions increase in size. On moving from left to right the size decreases. Thus on moving diagonally the size remains nearly the same.

Q.54 (1)

EA of third period elements is greater than second period in the group.

Q.55 (2)

EA increases from left to right in period & nitrogen family is having endothermic EA.

Q.56 (3)

Since N is having half filled stability of orbital.

Q.57 (4)

IP = –EA

Q.58 (4)

In chlorine, the addition of additional electron to larger 3p-subshell experiences less electron-electron repulsion than smaller 2p-subshell of fluorine. Phosphorus has very low electron affinity because there is high electron repulsion when the incoming electron enters an orbital that is already half filled.

Q.59 (2)

EA increases from left to right in period & nitrogen family is having endothermic EA.

Q.60 (3)

According to Mulliken's, the electronegativity =

$$\frac{\text{Ionisation energy} + \text{Electron affinity}}{2}$$

Q.61 (1)

Negative electron gain enthalpy of F is less than Cl. This is due to the fact that when an electron is added to F, the added electron goes to the smaller n = 2 energy level and experiences significant repulsion from the other electrons present in this level. In Cl, the electron goes to the larger n = 3 energy level and consequently occupies a larger region of space leading to much less electron-electron repulsion.

Q.62 (2)

Q.63 (2)

Q.64 (4)

IE (I) of N is greater than IE (I) of O because of stable half filled electronic configuration.

Q.65 (1)

EA of F is less than Cl.

Q.66 (2)

Q.67 (3)

In isoelectronic if $\frac{z}{e}$ is high, size is less.

	Cl^-	K^+
$\frac{z}{e} =$	$\frac{17}{18}$	$\frac{19}{18}$

Q.68 (3)

Q.69 (3)

Q.70 (1)

Cl has highest E.A.

Q.71 (3)

EN decreases down the group.

Q.72 (4)

EN increases along the period.

Q.73 (2)

Acidic nature increases along the period from left to right.

Q.74 (4)

Acidic nature of hydroxide decreases down the group.

Q.75 (1)

Alkali metals are basic.

Q.76 (1)

Basic nature of oxide decreases along the period.

Q.77 (2)

$1E_2$ of Be > B because of fully filled stability of orbital.

Q.78 (3)

(1) Across a period, with increase in atomic number, electron gain enthalpy becomes more negative because left to right across a period effective nuclear charge increases and consequently it will be easier to add an electron to a small atom.

(2) Nitrogen has very low electron affinity because there is high electron repulsion when the incoming electron enters an orbital that is already half filled.

(3) As we move in a group from top to bottom, electron gain enthalpy becomes less negative because the size of the atom increases and the added electron would be at larger distance from the nucleus.

Negative electron gain enthalpy of F is less than Cl. This is due to the fact that when an electron is added to F, the added electron goes to the smaller n = 2 energy level and experiences significant repulsion from the other electrons present in this level. In Cl, the electron goes to the larger n = 3 energy level and consequently occupies a larger region of space leading to much less electron-electron repulsion. So the correct order is Cl > F > Br > I.

(4) Refer option (3).

Q.79 (d)

$$SO_2 > \underset{Acidic}{P_2O_3} > \underset{Weak\ acidic}{SiO_2} > \underset{Amphoteric}{Al_2O_3}$$

Q.80 (c)

Q.81 (4)

EXERCISE-III (JEE MAIN LEVEL)

Q.1 (2)

Mg Ca
Sr
24 40
96

$$\frac{96+24}{2} = \frac{120}{2} = 60$$

In this atomic mass of central element was not arithmatic mean of atomic mass of ofter two element so, Mg, Ca, Sr is not a Dobereiner triad.

Q.2 (4)

Q.3 (2)

Q.4 (3)

Q.5 (2)

The d-block has 10 columns, because a maximum of 10 electrons can occupy all the (5) orbitals in a d-subshell.

Q.7 (2)

Down the group the size of atomic increases as new atomic orbits are added with increasing atomic number.

Q.8 (4)

Yes, hydrogen can be placed in 1st group on the basis of its valency +1 (H^+).

Q.9 (3)

Silver belongs to Vth period. So the atomic number of elements placed above and below will be $47 - 18 = 29$ and $47 + 32 = 79$ respectively.

Q.10 (4)

Z = 118 $[Rn]^{86}\,5f^{14}\,6d^{10}\,7s^2\,7p^6$; as last electron enters in p-subshell, it belongs to p-block. Thus its group number will be 10 + 2 + 6 = 18. Hence the element is a noble gas.

Q.11 (2)

Moseley work on X-ray spectrum.

Q.12 (2)

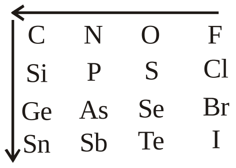
Atomic no. is the base of modern periodic table/long form of periodic table.

Q.13 (1)

Q.14 (4)

Q.15 (1)

Q.16 (1)



C	N	O	F
Si	P	S	Cl
Ge	As	Se	Br
Sn	Sb	Te	I

Q.17 (3)

Q.18 (3)

Penetration of p-subshell electron is less than s-subshell electrons. Thus 3s electron of magnesium is closer to the nucleus as compared to the 3p electron of aluminium.

Q.19 (1)

Across the period (i.e. 2nd period) the size of atom decreases and nuclear charge increases. So generally the ionisation energy increases. However the ionisation energy of Be is greater than B because of more penetration power of $2s^2$ sub-shell electrons of Be as compared to that of the $2p^1$ sub-shell electron of B.

Q.20 (4)

Q.21 (2)

Half filled electron configuration has extra stability. Hence the removal of electron from outer most orbit requires higher energy.

Q.22 (2)

Completely filled electron configurations and half filled electron configurations are expected to have higher ionisation energies. $ns^2\ np^5$ will have higher first ionisation energy than $ns^2\ np^4$ on account of smaller size of atom and higher nuclear charge.

Q.23 (3)

It can be seen by electronic configuration.

Q.24 (4)

Second ionisation energy of potassium is greater than that of Ca. In case of potassium ion (i.e. K^+) the electron removal from the stable inert gas configuration ($1s^2\ 2s^2\ 2p^6 3s^2 3p^6$) requires much higher energy.

Q.25 (2)

N has higher first ionisation energy than that of O because of extra stable half-filled electron configuration.

Q.26 (2)

Q.27 (2)

As fourth electron is to be removed from inert gas configuration i.e. $1s^2$, $2s^2$, $2p^6$. Thus I.E. (IV) is very high.

Q.28 (4)

Q.29 (3*)

(i) Due to more effective nuclear charge in Be^+ than Be.

(iii) Across the period, atomic size decreases and nuclear charge increases and thus valence shell electron(s) is/are tightly held by nucleus. So, ionisation energy of C is greater than Be.

Q.30 (4)

Q.31 (3)

Q.32 (1)

Second electron affinity of an element is alway exothermic.

Q.33 (3)

Q.34 (1)

Q.35 (4)

There is more interelectronic repulsion in 2p-subshell of fluorine than chlorine (3p). So extra electron will be added easily in 3p-subshell of chlorine as compared to 2p-subshell of fluorine.

Q.36 (4)

Q.37 (3)

The addition of extra electron is difficult to the atom having stable configuration and so electron gain enthalpy will be positive. Similarly the removal of electron is quite difficult from stable configuration and so ionisation enthalpy is higher. However EN remains unaffected because it neither involves gain nor loss of electron.

Q.38 (3)

Q.39 (1)

Higher the metallic character, greater will be the basic character of its oxide as ΔE_n (difference in electronegativities between element and oxygen) increases.

EXERCISE-IV

Q.1 0015

$Ni^+ = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^7$

Maximum number of electrons having
same spin = 15

Q.2 350

Lattice enthalpy = 772 KJ/Mol

$$\frac{(\Delta H.E)_{Na^+}}{(\Delta H.E)_{Cl^-}} = \frac{30}{25} = \frac{6}{5};$$

$$\underset{(x)}{(\Delta HE)_{Na^+}} = \frac{6}{5}\ \underset{(y)}{(\Delta HE)_{Cl^-}}$$

$$x = \frac{6}{5}y$$

$(\Delta H)_{solution}$ = Lattice enthalpy + H.E. of Na^+ + H.E. of Cl^-

$$2 = 272 + (y) + \frac{6}{5}(y)$$

$$-770 = \frac{11}{5}y\ ;\ y = \frac{770\times 5}{11} = 350\text{ KJ}$$

Q.3 7

N, O, F, P, C, S, Cl

Q.4 8

$r_A + r_A = 10Å$ ...(i)
$r_B + r_B = 6Å$ ...(ii)
(i) + (ii)
$2(r_A + r_B) = 16Å$
$r_A + r_B = 8Å$

Q.5 30

Zn

Q.6 1

$A(g) + e^- \rightarrow A^-(g)$ $\Delta H = +E.G.E$ ....(1)
$A(g) \rightarrow A^{2+}(g) + 2e^-$ $\Delta H = 1200$ KJ/mol.........(2)
$A^-(g) \rightarrow A^{2+}(g) + 3e^-$ $\Delta H = 1100$ KJ/mol.........(3)
Eq. (3) = (2) – (1)
– E.G.E + 1200 = 1100
– E.G.E = –100 kJ/mol
P = 1.

Q.7 **23**

$A + e^- \longrightarrow A^- + 3$ eV

Number of mole of A = 10/30.

As 1 mole of A releases the amount of energy = 3 × 23 kcal.

∴ Energy released for conversion of 10/30 mole of gaseous A into A^- ions = $\frac{3 \times 23}{30} \times 10$ = **23 kCal**

Q.8 42

According to triad rule at wt. of $Y = \frac{X+Z}{2}$

$$26 = \frac{10+Z}{2};$$

Z = 42

Q.9 116

Un – 1

Un – 1

Hex – 6

So at no. = 116

Q.10 33

$1E_2$ will be $> 1E_1 < 1E_3$

Q.11 (2)

Q.12 (2)

Q.13 (1)

Q.14 (1)

Q.15 (2)

Q.16 (2)

PREVIOUS YEAR'S

Q.1 (3)

Q.2 (4)

Q.3 (4)

Q.4 (2)

Q.5 (2)

Q.6 (4)

Q.7 (1)

Q.8 (2)

NEET/AIPMT

Q.1 (2)

$$\underrightarrow{\text{BeO < MgO < CaO < BaO}}$$

Basic character increases.

So, the most acidic should be BeO. In fact, BeO is amphoteric oxide while other given oxides are basic

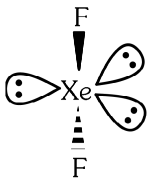
Q.2 (3)

Q.3 (1)

Q.4 (2)

Q.5 (3)

XeF_2



XeF_2 has maximum 3 lone-pair – lone-pair repulsions

Q.6 (4)

.

IUPAC nomenclature

119 → Ununennium → Uue

JEE MAIN

Q.1 **(2)**

Acidic → Cl_2O_7

Neutral → N_2O, NO

Amphoteric → As_2O_3

Basic → Na_2O

Q.2 **(b)**

Electron gain enthalpy of F is less than Cl due to compact size

Q. 3 (4)

Sc, Pb, Bi are metals

Te is metalloid

Q.4 **(A)**

Element "E" = Selenium (Se) valence electronic configuration =[Ar] $3d^{10}4s^24p^4$

Q.5 (A)

$Mg^{+2} < Na^+ \; F^- < O^{2-} < N^{3-}$

$10e^-s \; 10e^-s \; 10e^-s \; 10e^-s \; 10e^-s$

so they are isoelectronic species.

or Number of protons × $\frac{1}{\text{size}}$

Size × $\frac{1}{\text{Atomic number}}$

Q.6 **(C)**

Polarizing power (pp) $\propto \frac{1}{\text{size of cation}} \propto$ Covalent nature

Salt are –

(A) LiCl	$\rightarrow Li^+ + Cl^-$
(B) NaCl	$\rightarrow Na^+ + Cl^-$
(C) KCl	$\rightarrow K^+ + Cl^-$
(D) CsCl	$\rightarrow Cs^+ + Cl^-$

Order of size of cations

$Li^+ < Na^+ < K^+ < Cs^+$

So decreasing order of covalent nature

$LiCl > NaCl > KCl > CsCl$

According to option.

$A > B > C > D$

Option (C) is correct.

Q.7 **(A)**

I^{st} I.E of O < I^{st} I.E. of N

${}_8O = 1s^2\ 2s^2\ 2p^4$ ${}_7N = 1s^2\ 2s^2\ 2p^3$

Partially filled Half filled

Both, A and R are correct and R is the correct explanation of A.

Q.8 (4)

(1) IE of Be is more than B due to stable full filled 2s-orbital

(2) IE of N is more than O due to more stable half filled 2p-orbital

Q.9 (4)

Atomic no = 103(Unt)

Q.10 (4)

Radius of Anionic species is always greater than their parent atomic Radius but it can not say that covalent radius is double of the atomic radius.

Q.11 (4)

(1) Li
Na
K (i.E↓)

So, K < Na < Li

(2) In the period, noble gas has maximum 1.E

(3) atomic number 37 ⇒ 2, 8, 8 18 + 1 ⇒ s Block S^1
atomic number 38 ⇒ 2, 8, 8 18 + 2 ⇒ s Block S^2

So, Z = 37 has lesser i.E than

Z = 38

(4) actually the first i.E of Zn > $i.E_1$ of Ga

Due to completely filled orbitals

Q.12 (2)

$3s^2 3p^1 < 3s^2 < 3s^2 3p^4 < 3s^2 3p^3$

Q.13 (3)

(A) Rb and Cs

(C) Ar and Kr are nearly same electron gain enthalpy.

Q.14 (3)

Order of 1^{st} I.E.

I.E. ↑ → along the period
But Mg has more ionization energy because of stable electronic configuration

$$\underset{496}{Na} < Al < \underset{731}{Mg} < \underset{786}{Si}$$

So, the value of I.E. of Al must lie in between 496 and 737 i.e. ⇒ (3) 577

Q.15 (1)

Na > Mg > Be > Si > P

Chemical Bonding and Molecular Structure

EXERCISE-I (MHT CET LEVEL)

Q.1 (4)

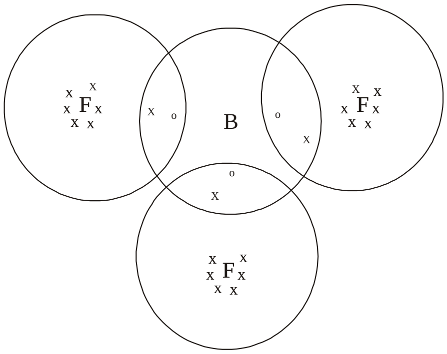
Q.2 (3)
In N_2 molecule each Nitrogen atom contribute $3e^-$ so total no. of electron's are 6.

Q.3 (3)
In $CaCl_2$ calcium loose 2 electrons and transfer to Cl atom thus both acquire outermost 8 electron in valence shell.

Q.4 (3)
Covalent bond is directional.

Q.5 (3)
In co-ordinate bond acceptor contain vacant orbital.

Q.6 (2)
BF_3 does not follow octate rule because central atom, boron lacks an electron pair, Thus, it also acts as Lewis acid,



Q.7 (3)
Sulphuric acid contain, covalent and co-ordinate bond.

Q.8 (1)
In graphite all carbon atoms are sp^2-hybridised and have covalent bond.

Q.9 (3)
In fluorine molecule formation p-p orbitals take part in bond formation.

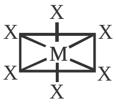
Q.10 (2)
1σ and 2π

Q.11 (1)
σ bond is formed on internuclear axis overlap.

Q.12 (1)

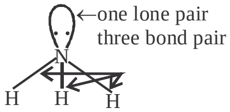
Q.13 (1)

Q.14 **(4)**
d^2sp^3 hybridisation has octahedral structure as follow:-



Q.15 (2)
For T-shape geometry the molecule must have 3 bonded pair and 2 lone pair of electrions.

Q.16 **(3)**
Pyramidal strueture has three bond pair & one lone pair & hybridisation is sp^3 like – NH_3



Q.17 (4)

Q.18 (1)

Q.19 **(3)**
The correct electron dot structure of N_2O is


$:\ddot{N}=\!=N=\!=\ddot{\underset{..}{O}} \longleftrightarrow N\equiv\overset{\oplus}{N}-\overset{\ominus}{O}$

Q.20 (4)
Choice (1), (2), (3) are the resonance structures of CO_2

Q.21 (2)

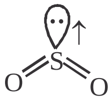
Q.22 (4)
lp – lp repulsion is more in H_2O

Q.23 (1)
Bond angle : $sp > sp^2 > sp^3 > sp^3d^2$

Q.24 (3)
Compound is carbontetrachloride because CCl_4 has sp^3 - hybridization 4 orbitals giving regular tetrahedron geometry. In others the geometry is little distorted inspite of sp^3 - hybridization due to different atoms on the vertices of tetrahedron.

Q.25 (3)
SO_4^{2-} has 42 electrons; CO_3^{2-} has 32 electrons; NO_3^- has 32 electrons.

Q.26 **(1)**
The SO_2 molecule has not dipole moment while other molecule has net dipole moment zero.

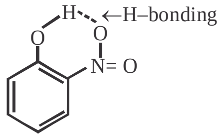


So, SO_2 is a palar molecule

Q.27 (4)
BF_3 is planar while NF_3 is pyramidal due to the presence of lone pair of electron on nitrogen in NF_3

Q.28 (1)

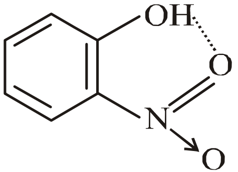
Q.29 **(2)**
Intramolecular H – bonding is present in O-nitrophenol as follow:-



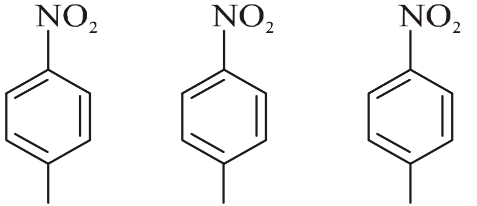
Q.30 (3)

Ortho-nitrophenol has intramolecular

H-bonding



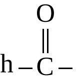
and paranitrophenol has intermolecular H-bonding.



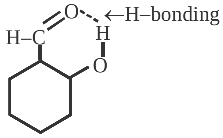
Hence former is more volatile than latter.

Q.31 (2)

Intramolecular hydrogen bonding is formed in salicylaldehyde it is due to adjacent phenolic hydrogen



interact with $-\overset{\overset{\large O}{\|}}{C}-$ group of aldehyde.



Q.32 (2)

Ice has hydrogen bonding.

Q.33 (2)

H – F has highest boiling point because it has hydrogen bonding.

Q.34 (3)

o-Nitrophenol has intramolecular hydrogen bonding but p-Nitrophenol has intermolecular hydrogen bonding so boiling point of p-Nitrophenol is more than o-Nitrophenol.

Q.35 (4)

Mercury has very weak interatomic forces so it remains in liquid state.

Q.36 (4)

Q.37 **(3)**

Nitrogen being smaller in size can effectively form $P\pi - P\pi$ bonds with other atoms like C & O as well as itself.

$$:N\overset{\pi}{\underset{\pi}{\equiv}}N:$$

Q.38 (1)

Q.39 (3)

Q.40 (4)

Boron does not have d-orbital.

Q.41 (3)

Q.42 (1)

Q.43 (2)

Ionic bonding is non directional, whereas covalent bonding is directional. So, CO_2 is directional.

Q.44 **(4)**

HCl gas is covalent and NaCl is ionic as the electronegativeity difference between H & Cl is less than 2.1

Q.45 (1)

Bond formation is always exothermic. Compounds of sodium are ionic.

Q.46 (2)

Because CsF is electrovalent compound.

Q.47 (2)

Cation will be formed easily when IE is low.
Anion will be formed easily when EA is high.

Q.48 (1)

Ionic compounds are polar.

Q.49 (2)

When electronegativity difference is more between two joined atoms then covalent bond becomes polar and electron pair forming a bond don't remain in the centre.

Q.50 (4)

AgI is most covalent.

Q.51 (3)

Q.52 (4)

Q.53 (2)

Q.54 (4)

Q.55 (2)

Q.56 (2)

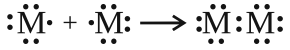
Sol. Bond length $\propto \dfrac{1}{\text{Bond order}}$

EXERCISE-II (NEET LEVEL)

Q.1 (3)

Q.2 (2)

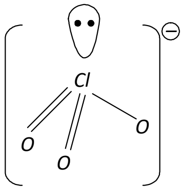
Non-metals readily form diatomic molecules by sharing of electrons. Element M($1s^2\ 2s^2\ 2p^5$) has seven electrons in its valence shell and thus needs one more electron to complete its octet. Therefore, two atoms share one electron each to form a diatomic molecule (M_2)

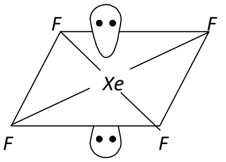


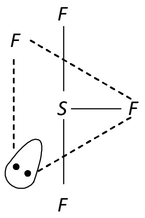
Q.3 (3)
BCl_3 is electron deficient compound because it has only '6' electrons after forming bond.

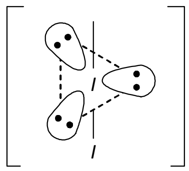
Q.4 (3)

Q.5 (4)









Q.6 (4)
HNO_2 does not have co-ordinate bond. Structure is $H-O-N=O$.

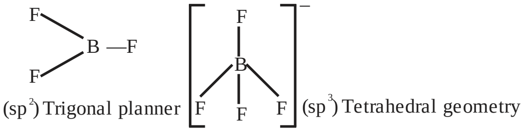
Q.7 (2)

Q.8 (3)

Q.9 (2)

Q.10 (4)
$(C_2H_5)_3B \leftarrow :N(CH_3)_3$
$BF_3 \leftarrow NH_3$

Q.11 **(3)**
In BF_3 hybridisation is sp^2 where as in BF_4^- hybridisation is sp^3 as follows:-



Q.12 (2)
The hybridisation of diamond is sp^3 in graphite sp^2 & in ethyne sp hybridised

Q.13 (4)
In graphite every carbon is linked with three carbon & form sp^2 hybridisation while in diamond every carbon is linked or bonded with four carbon atom with 3D – Network by sp^3 hybridisation.

Q.14 (3)
The pair of compound having identical shapes is XeF_2 & $ZnCl_2$ both has linear shape.

Q.15 (2)

Q.16 (3)

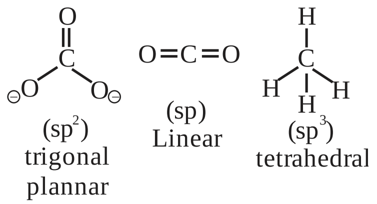
Q.17 (3)

Q.18 (4)
BF_4^- hybridisation sp^3, tetrahedral structure.
NH_4^+ hybridisation sp^3, tetrahedral structure.

Q.19 (3)

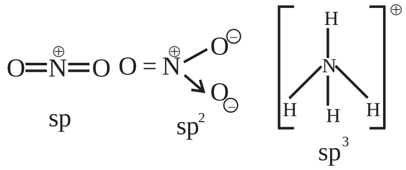
Q.20 (4)

Q.21 (3)



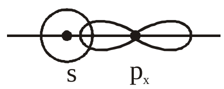
The hybrid orbital of carbon in CO_2 is sp in CH_4 sp^3 and in CO_3^{2-} is sp^2 hybrid.

Q.22 (3)
In NO_2^+, NO_3^- and NH_4^+ the percentage s-character in hydrid orbital of nitrogen is 50%, 33.3% and 25% it is due to NO_2^+ is sp hybrid NO_3^- is sp^2 hybrid & NH_4^+ as sp^3 hybrid & the structure is as follow:



Q.23 (4)
In BF_3 hybridisation charge from sp^2 to sp^3 & NH_3 donate lone pair acting as Lewis base.

Q.24 (4)



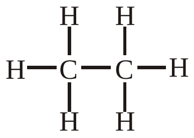
Q.25 (3)
Overlap order is $p-p > s-p > s-s$.

Q.26 CO_2 is sp-hybridised

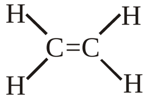
Q.27 (2)
NH_3 has sp^3 hybridised central atom so it is non planar.

Q.28 (3)
In a double bond connecting two atom sharing of 4 electrons take place as in $H_2C=CH_2$.

Q.29 (4)



(SP^3 Carbon)



(SP^2 Carbon)

ethane

ethene

H—C ≡ C—H (SP Carbon)
Acetylene

EN of carbon SP > SP^2 > SP^3 ,C – H bond energy $\propto$ EN of carbon

$\therefore$ bond energy Acetylene > ethene > ethane

Q.30 (2)

π bond are formed by lateral overlap of orbitals.

Q.31 (2)

$$\text{bond length} \propto \frac{\mathbf{1}}{\mathbf{bond\,order}}$$

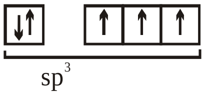
Q.32 (1)

$N \equiv C - C \equiv N$ Linear

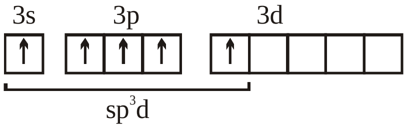
Q.33 (2)

Q.34 (3)

$PCl_3 \rightarrow sp^3$



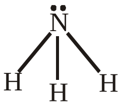
$PCl_5 \rightarrow sp^3d$



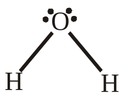
Q.35 (3)

$$\%\text{s character in } sp^2 = \frac{1}{3} \times 100$$

Q.36 (3)



Pyramidal;

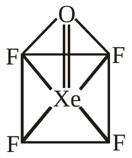


bent

Q.37 (4)

:SF_4
sp^3d with one l.p.
CF_4
sp^3 zero l.p.
XeF_4 sp^3d^2 two l.p.

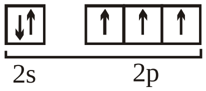
Q.38 (4)



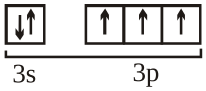
Q.39 (4)

Q.40 (3)

$\ddot{N}H_3 - sp^3$



$\ddot{P}Cl_3 - sp^3$

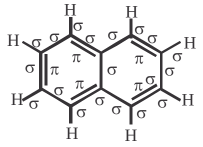


Q.41 (3)

Q.42 (4)

Q.43 (2)

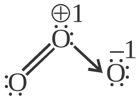
The no. of π bond & σ bond in naphtalene is 5 & 19.



$\pi = 5\,,\ \sigma = 19$

Q.44 (4)

The formal charge on central oxygen atom in ozone is +1 & its structure as follow:-



Q.45 (3)

CO has triple bond so bond will be shortest, CO_2 has double bond, CO_3^{2-} and HCO_2^- will be in conjugation in which conjugation of carbontate ion will be more so bond length will be higher.

Q.46 (1)

The number of lone pairs of electrons on central atom in various given species are
Species Number of lone pairs on central atom
IF_7
0
IF5
1
ClF_3
2
XeF_2
3
Thus the correct increasing order is

$$IF_7 < IF_5 < ClF_3 < XeF_2$$

0 1 2 3

Q.47 (1)

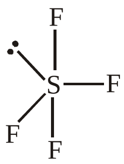
If bond order is high stronger bond is formed & bond length decreases.

Q.48 (2)

In NH_3 nitrogen has one lone pair of electron.

Q.49 (2)

Q.50 (3)



In TBP axial and equatorial bond length are unequal.

Q.51 (1)
The bond angle in PH_3 would be expected to be close to 90°. (The bond angle $H-P-H$ in PH_3 is 93°)

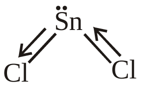
Q.52 (2)
In BF_3 molecule Boron is sp^2 hybridised so its all atoms are co-planar.

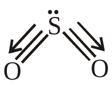
Q.53 (3)
Due to *l*p–*l*p repulsions, bond angle in H_2O is lower (104°.5°) than that in NH_3(107°) and CH_4(109°28'). BeF_2 on the other hand, has sp-hybridization and hence has a bond angle of 180°.

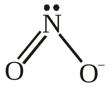
Q.54 (1)
As the s-character of hybridized orbitals decreases the bond angle also decreases
In sp^3 hybridisation: s-character 1/4, bond angle 109°
In sp^2 hybridisation: s-character 1/3, bond angle 120°
In sp hybridisation: s-character 1/2, bond angle 180°

Q.55 (1)

Q.56 (3)



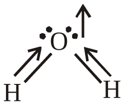




$\frac{\text{Polar}}{\mu \neq 0}$ $\mu \neq 0$ $\mu \neq 0$

Q.57 (1)

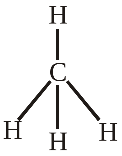
Q.58 (2)

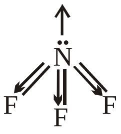


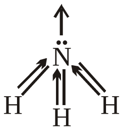
$\mu \neq 0$ $\mu = 0$

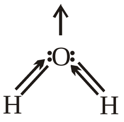
Q.59 (1)

Q.60 (1)





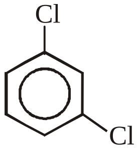




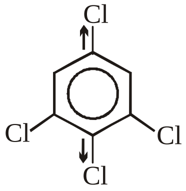
$\mu = 0$ Dipole of lp Dipole of lp Dipole of lp
& N–F cancel each other & N–F in opposite direction
& bp add $\mu \neq 0$

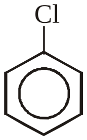
Q.61 (4)

$$\mu = \sqrt{\mu_1^2 + \mu_2^2 + 2\mu_1\mu_2 \cos\theta}$$



$\cos\theta = 120° \mu_1 = \mu_2$





$\cos\theta = 120°$ so same.

Q.62 (1)
All dipole cancel each other.

Q.63 (3)

Q.64 (2)
PH_3 has the lowest boiling point because it does not form Hydrogen bond.

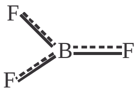
Q.65 (2)
Benzene is non-polar.

Q.66 (1)

Q.67 (4)

Q.68 (4)
Glycerol has a three OH group hence it is viscous in nature.

Q.69 (3)
There is maximum $P\pi - P\pi$ back bonding present in BF_3 molecule it is because BF_3 is a weaker lewis acid than other boron halides & its fluorine atom can easily donate lone - pair to vacant orbital of boron . as follow:-



Q.70 (4)
Main axis of diatomic molecule is Z. If the orbitals of different axis overlap there will be no bond formed b/w different orbital . So P_x & P_y orbital do not overlap with each other .

Q.71 (1)

Q.72 (3)

Q.73 (3)
Nitrogen does not contain d-orbital so can not expent octet.

Q.74 (4)
The correct order of solublity of sulphides is
$Na_2S > ZnS > CuS$

Q.75 (3)

Q.76 (2)
NaCl is ionic crystal so it is formed by Na^+ and Cl^- ions.

Q.77 (4)

Covalent character depend on the size of cation and anion.

Q.78 (2)

Due to its small size (fajan's rule)

Q.79 (3)

NaCl is formed by electrovalent bonding.

Q.80 (4)

Valency of metal is + 2 by formula MO so its phosphate would be $M_3(PO_4)_2$ because valency of $[PO_4]$ is – 3.

Q.81 (4)

The value of lattice energy depends on the charges present on the two ions and distance between them. Lattice energy is high if charges are high and ionic radii are small.

Q.82 (3)

$I^{\delta+} - Cl^{\delta-}$ polar attraction is more.

Q.83 (1)

According to Fajan's rule, polarisation of anion is influenced by charge and size of cation more is the charge on cation, more is polarisation of anion.

Q.84 (1)

Lattice energy $\propto \dfrac{z + z^-}{r^+ + r^-}$

Q.85 (3)

Q.86 (1)

We know that in O_2 bond, the order is 2 and in O_2^- bond, the order is 1.5. Therefore the wrong statements is (1).

Q.87 (2)

Bond length decreases with an increase in bond order. Therefore, the order of bond length in these species is $O_2^+ < O_2^- > O_2 < O_2^{2-}$ (bond order – $O_2^+ = 2.5$, $O_2 = 2$, $O_2^- = 1.5$, $O_2^{2-} = 1$).

Q.88 (2)

Q.89 (1)

Q.90 (3)

Q.91 (4)

Q.92 (1)

Q.93 (1)

Q.94 (4)

Q.95 (3)

Q.96 (4)

N_2 molecule has diamagnetic structure because its all bonding molecular orbitals are paired and its molecular orbital configuration

$\Rightarrow \sigma 1s^2 < \sigma^* 1s^2, < \sigma 2s^2 < \sigma^* s^2,$

$< \pi 2P_x^2 = \pi 2P_y^2 < \sigma 2P_Z^2$

Q.97 (3)

The no. of anti-bonding electron present in O_2^+ molecular ion is 5 & it has 10 bonding electron because it contain total 15 electron.

Q.98 (4)

$O_2^{2-} < O_2^- < O_2 < O_2^+$

O_2^{2-} has bond order $1 \Rightarrow \dfrac{6-4}{2} = 1$

O_2^- has bond order $3/2 \Rightarrow \dfrac{6-3}{2} = 3/2$ or 1.5

O_2 has bond order $2 \Rightarrow \dfrac{6-2}{2} = 2$

O_2^+ has bond order $2.5 \Rightarrow \dfrac{6-1}{2} = 5/2$ or 2.5

Q.99 (3)

O_2^+ and O_2^- both are paramagnetic as O_2^- has 1 unpaired electrons & O_2^+ also has 1 unpaired electrons but the bond order O_2^+ has 5/2 which is greater from O_2^- which is 3/2 and as we know higher the bond order higher will be stability of molecule.

$O_2^- \rightarrow \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2P_z^2 \pi 2p_x^2 \pi 2P_y^2 \pi^* 2P_x^2$

$\pi^* 2P_y^1$

$O_2^- \rightarrow B.O. = 3/2$

$O_2^+ \rightarrow \sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2 \sigma 2P_z^2 \pi 2p_x^2 \pi 2P_y^2 \pi^* 2P_x^1$

$\pi^* 2P_y^0$

$O_2^+ \rightarrow B.O. = 5/2$

Q.100 (1)

O_2^+ B.O. $= \dfrac{1}{2}(10 - 5) = 2.5$

NO B.O. $= \dfrac{1}{2}(10 - 5) = 2.5$

NO^{+2} B.O. $= \dfrac{1}{2}(9 - 4) = 2.5$

N_2^+ B.O. $= \dfrac{1}{2}(10 - 5) = 2.5$

Q.101 (2)

Bond order of O_2^+ is highest so its bond length is smallest.

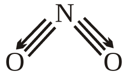
Q.102 (3)

B.O. of $N_2 = \dfrac{1}{2}(10 - 4) = 3$

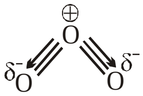
$N_2^+ = \dfrac{1}{2}(9 - 4) = 2.5$

Q.103 (1)

Q.104 (2)



$\mu \neq 0$



$\mu \neq 0$

Q.105 (1)

Q.106 (3)

Q.107 (1)

Bond order in NO = $\frac{1}{2}(10-5) = 2.5$

Bond order in NO⁺ = $\frac{1}{2}(10-4) = 3$

NO is paramagnetic and NO^+ is diamagnetic.

Q.108 (1)

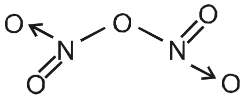
$$\left.\begin{aligned}&\text{In } CN^- \text{ no. of electron} = 14\\&\text{In } NO^+ \text{ no. of electron} = 14\end{aligned}\right\}\text{ so bond order is same.}$$

EXERCISE-III (JEE MAIN LEVEL)

OBJECTIVE QUESTIONS

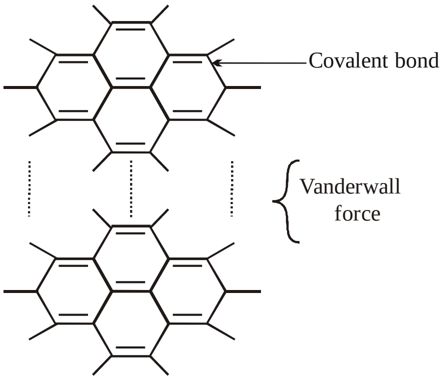
Q.1 (4)
The maximum covalency of representative element is equal to the number of s & p electrons in valence shell.

Q.2 (4)



Q.3 (3)

Sol.

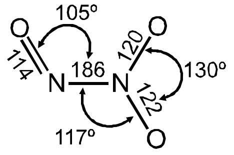


Q.4 (4)
In SF_6, PCl_5 and IF_7 the valence shell has 12, 10 and 14 electrons. As all contain more than 8 electrons in their valence shell they are example of super octet molecules.

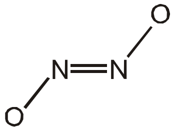
Q.5 (4)
In all species the valence shell contain more than 8 electrons. Thus the octet rule is not applicable to BrF_5, SF_6 and IF_7.

Q.6 (3)

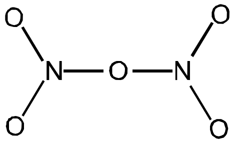
(1) N_2O_3 Dinitrogen trioxide



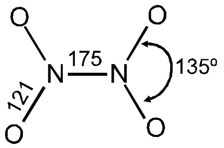
(2) $N_2O_2^{2-}$ Hyponitrite ion



(3) N_2O_5 Dinitrogen pentoxide

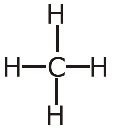


(4) N_2O_4 Dinitrogen tetroxide



Q.7 (3)

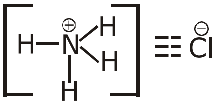
(1)



(only covalent)

(2) H—O—O—H (only covalent)

(3)



(covalent and ionic)

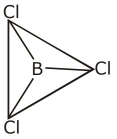
(4) H—C$\equiv$N (only covalent)

Q.8 (4)
The strength of metallic bonds depends upon the number of mobile electron(s) per atom. Sodium has only one mobile electron per atom where as iron has 8 mobile electrons per atom.

Q.9 (4)

(1) F–Be–F Linear

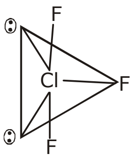
(2)



Trigonal planar

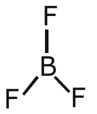
(3) $\ddot{N}H_3$ Trigonalpyramidal

(4)



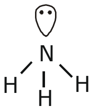
T-shapal

Q.10 (2)



sp^2, trigonal planar.

Q.11 (2)



sp^3

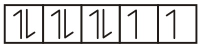
$[PtCl_4]^{2-}$

⇒ Pt is of 3rd transition series

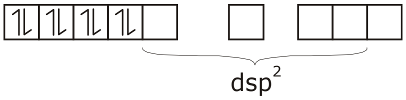
⇒ all ligand will be strong field ligand

⇒ Hybridization ⇒ dsp^2

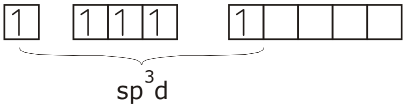
Originally

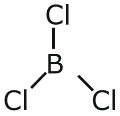


In this compound due to SFL.



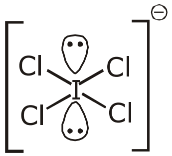
$PCl_5 \Rightarrow sp^3d$





⇒ sp^2

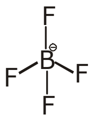
Q.12 (4)



⇒ No of bond pair-lone pair repulsion = 4

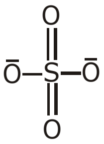
Q.13 (3)

(1)



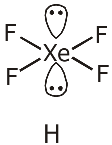
$SP^3 \Rightarrow$ Terahedral

(2)



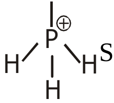
$SP^3 \Rightarrow$ Terahedral

(3)



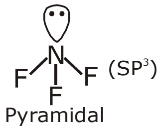
$SP^3d^2 \Rightarrow$ Square Planer

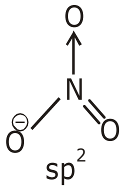
(4)



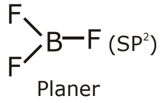
$SP^3 \Rightarrow$ Terahedral

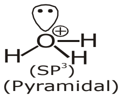
Q.14 (3)

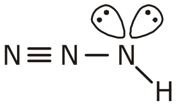




(planar)





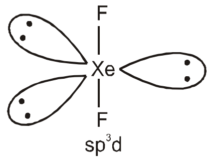


Q.15 (4)

N_3^-, $(CNO)^-$ and $(NCN)^{2-}$ all have same number of electrons i.e., 22 ; so all are isoelectronic with CO_2 which also has 22 electrons.

Q.16 (4)

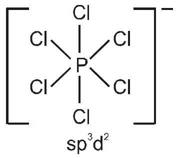
(i)



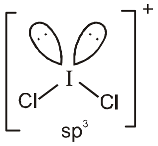
(ii) $[N = \overset{+}{N} = N]^{2-}$

sp

(iii) PCl_5 (s) exists as $[PCl_4]^+$ and $[PCl_6]^-$,

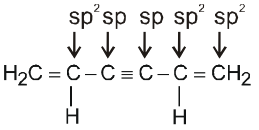


(iv) $I_2Cl_6(\ell) \rightleftharpoons [ICl_2]^+ + [ICl_4]^-$ self ionisation



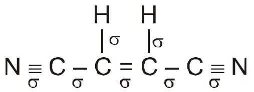
Q.17 (4)

(1)



$$H_2C = \overset{sp^2}{\underset{H}{C}} - \overset{sp}{C} \equiv \overset{sp}{C} - \overset{sp^2}{\underset{H}{C}} = \overset{sp^2}{C}H_2$$

(2)

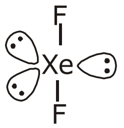


$$N \underset{\sigma}{\equiv} C \underset{\sigma}{-} \overset{H}{C} \underset{\sigma}{=} \overset{H}{C} \underset{\sigma}{-} C \underset{\sigma}{\equiv} N$$

(3) In diamond each carbon atom is in sp^3 hybridisation.

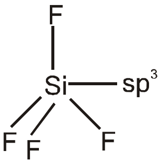
(4) $O = \underset{sp}{C} = \underset{sp}{C} = \underset{sp}{C} = O$

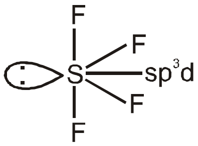
Q.18 (3)



= SP^3d

Q.19 (1)





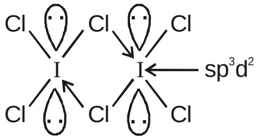
(1) Tetrahedral and see-saw shaped.
(2) Both are sp^3 hybridised and trigonal pyramid.
(3) Both are sp^3 hybridised and tetrahedral.
(4) Both are sp^3d^2 hybridised and octahedral.

Q.20 (2)

(i) SF_4	Steric no. = 5
	Lone pair = 1
(ii) $[PCl_4]^+$	Steric no. = 4
	Lone pair = 0
(iii) XeO_2F_2	Steric no. = 5
	Lone pair = 1
(iv) $ClOF_3$	Steric no. = 5
	Lone pair = 1

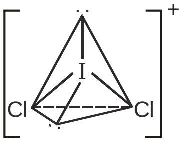
Q.21 (4)

ICl_3 does not exist, but the dimer I_2Cl_6 is a bright yellow solid. Its structure is planar.



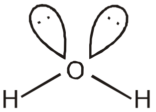
Note : $I_2Cl_4Br_2$ will have the same hybridisation as that of iodine in I_2Cl_6. But it exists in 3 different forms.

Q.22 (4)

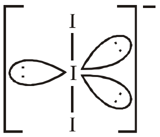


Q.23 (2)

(1)



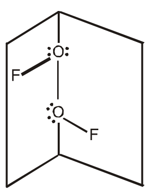
(2)



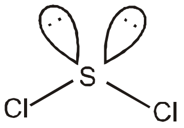
Number of bond pairs around I = 2.

Number of lone pairs around I = 3.

(3)



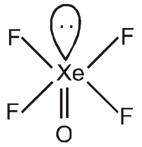
(4)

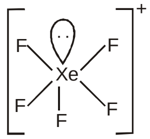


Q.24 (4)

(1) XeO_3 is trigonal pyramid. (2) IOF_4^+ is see-saw.
(3) PCl_5 is trigonal bipyramidal.

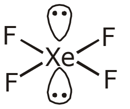
(4)





square pyramidal.

Q.25 (4)



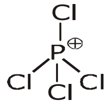
$= sp^3d^2$

Q.26 (1)

Hybridization is process of mixing of atomic orbital of nearly equal energy

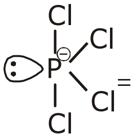
Q.27 (2)

PCl_4^+



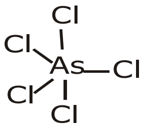
(Tetrahedral)

PCl_4^-



= see saw

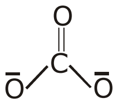
sp^3d



sp^3d Trigonal bi pyramidal

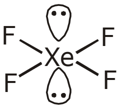
Q.28 (1)

(I)



$SP^2 = 33\%$ S

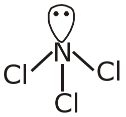
(II)



SP^3d^2 $\% S = \frac{1}{6} \times 100 = 16\%$

(III) $I - \overset{\ominus}{I} - I$ SP^3d $\frac{1}{5} \times 100 = 20\%S$

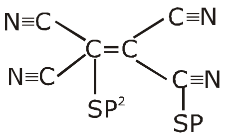
(IV)



$SP^3 = 25\%S$

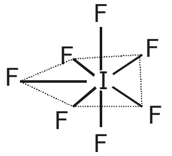
(V) Cl—Be—Cl 50%S
SP

Q.29 (3)



⇒ SP and SP^2 hybridized

Q.30 (2)

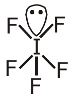


no. of B.P. =7

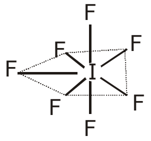
Q.31 (4)

$\overset{\ominus}{C}\equiv N$

Q.32 (1)

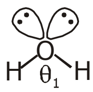


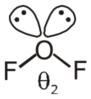
SP^3d^2 sq. Pyramidal ;

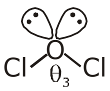


SP^3d^3 Pentagonal bipyramidal

Q.33 (2)

(1)


(2)


(3)


$= Cl_2O > H_2O > F_2O$
(exception)
As EN of surrounding atom ↑ BA ↓
(VSEPR theory)

Q.34 (1)

Atomic size arguments can be used for these species. Larger outer atoms result in larger angles due to steric repulsion.

Q.35 (1)

⇒ more s, more EN of c

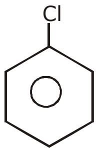
Q.36 (1)

$PH_3 \approx 91^o$
$NH_3 \approx 107^o$
(Drago's rule in PH_3)

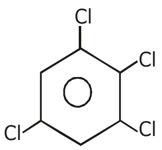
Q.37 (4)

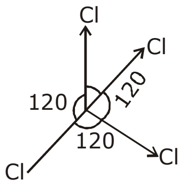
⇒ H_2O is polar molecule
⇒ Hence H_2O has higher critical temp

Q.38 (2)



$\mu = 1.5$ D





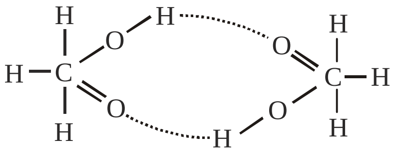
= d.M. of Cl at 120° will cancel each other
⇒ $\mu = 1.5$ D

Q.39 (1)

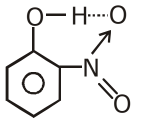
Q.40 (1)

H_2O due to H-bond has highest b.p.

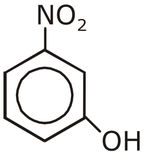
Q.41 (3)

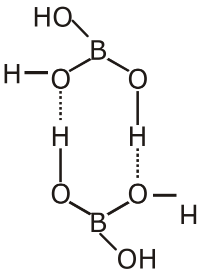
(I)


(Inter H-bond)

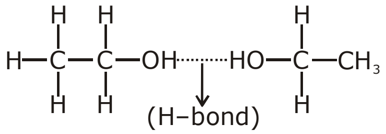
(II)


(Intra H-bond)

(III)


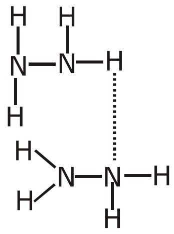
(IV)


Q.42 (3)



⇒ due to intermolecular H-bond, it has higher b.p. than CH_3OCH_3

Q.43 (3)



Q.44 (4)

due to H–bond, HF has Low volatility

Q.45 (2)

Hydrogen bond acts as a bridge between atoms which holds one atom by covalent bond and the other by hydrogen bond and this increase in the attraction between molecules.

Q.46 (4)

London forces are extremely short range in action and the weakest of all attractive forces.

The order of strength of bonds/ forces is ionic bond > covalent bond > hydrogen bond > London force.

Q.47 (4)

All statements are true.

Q.48 (4)

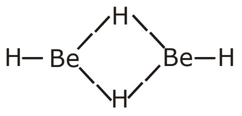
Boiling point of SbH_3 is greater than NH_3. The higher boiling point of SbH_3 is attributed to higher van der Waal forces because of its higher molecular weight.

Q.49 (2)

The boiling points of ICl, H_2S and CO can be explained on the basis of dipole-dipole attraction between their respective molecules. London dispersion force exists among the non-polar molecules like H_2, O_2, Br_2 etc. in solid or liquid states. Even in atoms in molecules which have no permanent dipole, instantaneous dipoles will arise as a result of momentary unbalances in electron distribution.

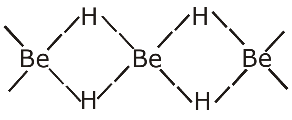
Q.50 (3)

(1)



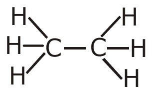
3C—2e⁻

(2)



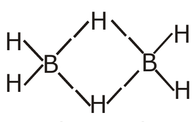
(3C—2e⁻)

(3)



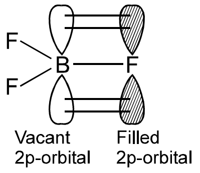
(does not contain 3C—2e⁻)

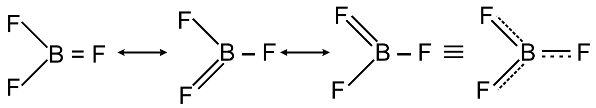
(4)



(3C—2e⁻)

Q.51 (4)





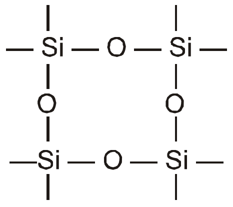
Decrease in B – F bond length is due to delocalised pπ –pπ bonding between filled p-orbital of F atom and vacant p-orbital of B atom.

Q.52 (1)

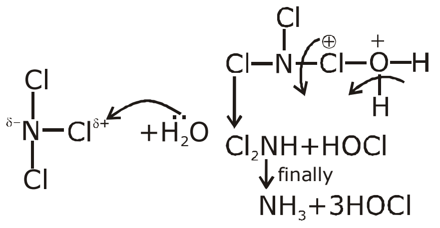
The central atom of PCl_5 have vacant d orbital therefore represent extended covalent bonding while in NCl_5 N have no vacant d orbital

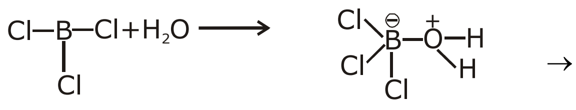
Q.53 (2)

Sol.



Q.54 (3)





$HCl+Cl_2BOH \xrightarrow{finally} B(OH)_3+3HCl$

Q.55 (4)

In CH_3Cl carbon does not carry vacant orbital so it can not be hydrolysed.

Q.56 (2)

Cs has lowest IE_1 amongst the metals and F has higher electron affinity. So Cs and F form most ionic compound.

Q.57 (2)

Sn^{4+} has highest polarising power amongst Na^+, Pb^{2+}, Sn^{4+} and Al^{3+} because of smaller size and higher charge. So $SnCl_4$ is most covalent and thus have least melting point.

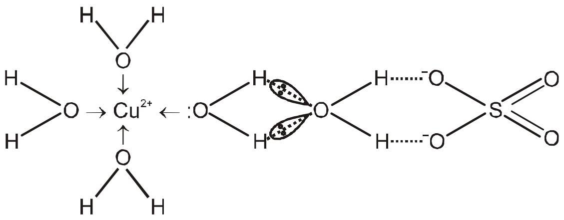
Q.58 (2)

Because of high charge density on Sn^{4+} it has high polarising power and thus leads to greater polarisation of anion i.e., greater distortion of electron clouds of the Cl^- ions. So $SnCl_4$ is most covalent.

Q.59 (2)

Increase in oxidation state (Ni^{4+}) increases the polarising power of cation and thus increases the polarisation of Br^- ion.

Q.60 (4)



Q.61 (1)

$C_3^{4-} \Rightarrow {}^{3-}C-C\equiv C^-$

$= \sigma-\text{bond} = 2$

$\pi-\text{bond} = 2$

Q.62 (1)

Polarizibility size of anions

$l^->Br^->Cl^->F^-$

Q.63 (2)

$Cl_2 > Br_2 > F_2 > I_2$

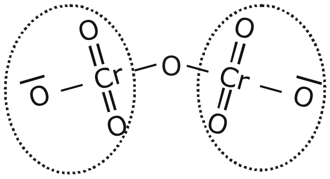
very less due to interelectronic repulsion due to small 2p orbital.

Q.64 (2)

According fajan's rule

Q.65 (4)

Q.66 (2)



$\Rightarrow$ 6   Cr — O

Bond equal due to resonance effect.

Q.67 (2)

(OF)

$$B.0 = \frac{1}{2}\left|\begin{array}{cc} 2 & 2 \\ 2 & 2 \\ 6 & 3 \end{array}\right| = \frac{1}{2}\left|\begin{array}{c} 6-3 \\ =1.5 \end{array}\right| = \frac{3}{2}$$

Q.68 (4)

(1) $NO \rightarrow N\overset{+}{O}$

3.0=2.5   B.O.=3

(Para)   (dia)

(2) $O_2^+ \longrightarrow O_2$ (unpaired=2)

unpaired e =1

B.O = 2.5   B.O = 2

(Para)   (Para)

(3) $N_2 \longrightarrow N_2^-$

(dia)   (para)

Q.69 (2)

B_2 bond order = 1 ; C_2 bond order = 2 ; F_2 bond order = 1 ; O_2^- bond order = 1.5

bond order $\propto$ 1/bond length.

Q.70 (3)

(3) $O_2 : (\sigma 1s)^2\ (\sigma^*1s)^2\ (\sigma 2s)^2\ (\sigma^*2s)^2\ (\sigma 2p_z)^2\ (\pi 2p^2{}_x = \pi 2p^2{}_y)\ (\pi^*2p_x{}^1 = \pi^*2p^1{}_y)$

$O_2^- : (\sigma 1s)^2\ (\sigma^*1s)^2\ (\sigma 2s)^2\ (\sigma^*2s)^2\ (\sigma 2p_z)^2\ (\pi 2p^2{}_x = \pi 2p^2{}_y)\ (\pi^*2p_x{}^2 = \pi^*2p^1{}_y)$

EXERCISE-IV

Q.1 4

$N_2^+,\ N_2^-,\ O_2^+, C_2^+$ have fractional bond order.

Q.2 4

(1) PCl_6^- (Coordinate Bond = 1)

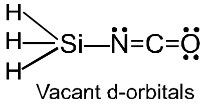
(2) $NH_3 \cdot BF_3$ (Coordinate Bond = 1)

(3) HNO_3 (Coordinate Bond = 1)

(4) CO   (Coordinate Bond = 1)

Q.3 5

Q.4 20



LP of N can be delocalised back bonding ($p\pi-d\pi$). SN of N = 2 sp

Q.5 4

Q.6 6

Q.7 4 (ii, iv, vi, vii)

Q.8 9

$(NO_3^-, CO_3^{2-}, F_2, Cl_2, Br_2, O_2^{2-}, O_2^-, Li_2^+, He_2^+)$

Q.9 4

Q.10 3

Q.11 (2)

Q.12 (4)

Q.13 (1)

Q.14 (4)

Q.15 (2)

Q.16 (3)

PREVIOUS YEAR'S

MHT

Q.1 (3)

Q.2 (4)

Q.3 (1)

Q.4 (1)

Q.5 (4)

Q.6 (3)

Q.7 (4)

Q.8 (3)

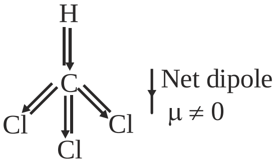
Q.9 (2)

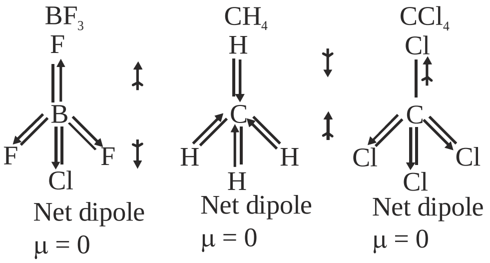
Q.10 (4)

Q.11 (2)

Q.12 (3)
Q.13 (4)
Q.14 (2)
Q.15 (3)

$CHCl_3$

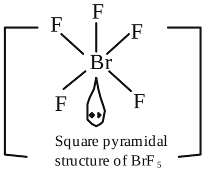




Among the given compound $CHCl_3$ is having dipole moment.

Q.16 (2)

According to VSEPR theory the shap of BrF_5 is square pyramidal and its electron geometry is octahedral because bromine being the central atom has five bonds connected with surrounding flurine atoms. Each F – Br — F bond making an angle od 90° in the same plane.

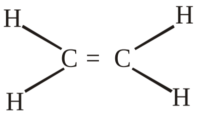


Q.17 (3)
Q.18 (4)
Q.19 (1)
Q.20 (2)

NEET/AIPMT

Q.1 (1)
Q.2 (1)
Q.3 (1)
Q.4 (1)
Q.5 (2)
Q.6 (3)

Sol.



5σ bond 1σ bond

Q.7 (2)

NF_3 and H_2O have sp^3 hybridisation

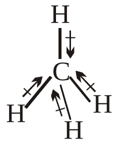
Q.8 (1)

O_2^-(17) Superoxide has one unpaired electron

$\sigma 1s^2\ \sigma^*1s^2\ \sigma 2s^2\ \sigma^*2s^2\ \sigma 2p_z^2\ \ \pi 2p_x^2 = \pi 2p_y^2\ \ \pi^*2p_x^2 = \pi^*2p_y^1$

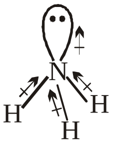
Q.9 (3)

(μ=0, symmetrical);



;

(μ=0, symmetrical)



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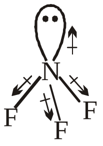
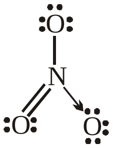


(μ≠0, Pyramidal) (μ≠0, Pyramidal)

In NH_3, H is less electronegative than N and hence dipole moment of each N – H bond is towards N and create high net dipole moment whereas in NF_3, F is more electronegative than N, the dipole moment of each N – F bond is opposite to that of lone pair, hence reducing the net dipole moment.

Q.10 (2)



sp^2 hybridised,
trigonal planes

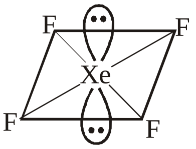
Q.11 (3)

In diamond and silicon carbide, central atom is sp^3 hybridised and hence, both are isostructural.

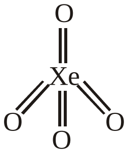
NH_3 and PH_3 both are pyramidal and central atom in both cases is sp^3 hybridised.

$SiCl_4$ and PCl_4^+, both are tetrahedral and central atom in both cases is sp^3 hybridised.

In XeF_4, Xe is sp^3d^2 hybridised and structure is square planar while in XeO_4, Xe is sp^3 hybridised and structure is tetrahedral.



XeF_4



XeO_4

Q.12 (4)

$O_2^-(16) \rightarrow KK\ \sigma 2s^2 \sigma^* 2s^2\ \sigma 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2\ \pi^* 2p_x^1 = \pi^* 2p_y^1$

$$\text{B.O.} = \frac{1}{2}[8-4] = 2$$

$O_2^-(17) \rightarrow KK\ \sigma 2s^2 \sigma^* 2s^2\ \sigma 2p_z^2 \pi 2p_x^2 = \pi 2p_y^2\ \pi^* 2p_x^1 = \pi^* 2p_y^1$

$$\text{B.O.} = \frac{1}{2}[8-5] = 1.5$$

$O_2^-(15) \rightarrow KK\ \sigma 2s^2 \sigma^* 2s^2\ \sigma 2p_z^2\ \pi 2p_x^2 = \pi 2p_y^2\ \pi^* 2p_x^1$

$$\text{B.O.} = \frac{1}{2}[8-3] = 2.5$$

As, Bond order $\propto$ stability

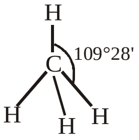
The decreasing order of stability is

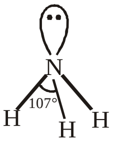
$O_2^+ > O_2 > O_2^- > O_2^{2-}$

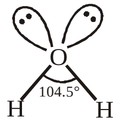
Q.13 (3)

According to VSEPR theory, the repulsive forces between lone pair and lone pair are greater than between lone pair and bond pair which are further greater than bond pair and bond pair.

Q.14 (4)







Q.15 (3)

H_2O_2, HCN and conc. CH_2COOH form intermolecular hydrogen bonding while cellulose has intramolecular hydrogen bonding.

Q.16 (1, 4)

(1) CO_3^{2-} : 6 + 24 + 2 = 32; sp^2 ; trigonal planar ; NO_3^- : 7 + 24 + 1 = 32; sp^2 ; trigonal planar

Hence, these are isoelectronic as well as isostructural.

(2) ClO_3^- : 17 + 24 + 1 = 42; sp^3 ; trigonal pyramidal ; CO_3^{2-} : 6 + 24 + 2 = 32; sp^2 ; trigonal planar

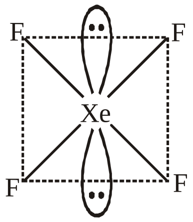
Hence, these are neither isoelectronic nor isostructural.

(3) SO_3^{2-} : 16 + 24 + 2 = 42; sp^3 ; trigonal pyramidal ; NO_3^- : 7 + 24 + 1 = 32; sp^2 ; trigonal planar

These are neither isoelectronic nor isostructural.

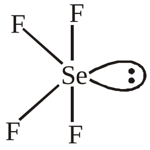
(4) ClO_3 : 17 + 24 + 1 = 42; sp^3 ; trigonal pyramidal ; SO_3^{2-} : 16 + 24 + 2 = 42; sp^3 ; trigonal pyramidal

Q.17 (4)

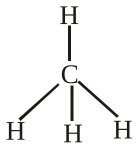


sp^3d^2 hybridisation (octahedral geometry, square planar shape)

Q.18 (3)



sp^3d hybridisation (see-saw shape)



sp^3 hybridisation (tetrahedral)

Q.19 (2)

Note : In this question, in place of isoelectronic there should be same nubmer of valence electrons.

Q.20 (3)

BCl_3 - Trigonal planar, sp^2 - hybridised, 120° angle.

Q.21 (2)

Molecular orbital electronic configurations and bond order values are :

$O_2 = \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_z^2, \pi 2p_x^2 = \pi 2p_y^2, \pi^* 2s_x^1 = \pi^* 2s_y^1$

$$\text{B.O.} = \frac{1}{2}(N_b - N_a) = \frac{1}{2}(10-6) = 2$$

$NO^+ = \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_z^2, \pi 2p_x^2 = \pi 2p_y^2$

$$\text{B.O.} = \frac{1}{2}(10-4) = 3$$

$CN^- = \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \pi 2p_x^2 = \pi 2p_y^2, \pi^* 2p_y^1$

$$\text{B.O.} = \frac{1}{2}(10-4) = 3$$

$CO = \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \pi 2p_x^2 = \pi 2p_y^2, \sigma 2p_z^2$

$$\text{B.O.} = \frac{1}{2}(10-4) = 3$$

$N_2 = \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \pi 2p_x^2 = \pi 2p_y^2, \sigma 2p_z^2$

$$\text{B.O.} = \frac{1}{2}(10-4) = 3$$

$O_2^- = \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_z^2, \pi 2p_x^2 = \pi 2p_y^2, \pi^* 2p_x^2 = \pi^* 2p_y^2$

$$\text{B.O.} = \frac{1}{2}(10-7) = 1.5$$

$NO = \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_z^2, \pi 2p_x^2 = \pi 2p_y^2, \pi^* 2p_x^1$

$$\text{B.O.} = \frac{1}{2}(10-5) = 2.5$$

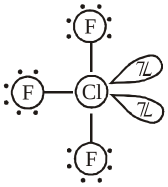
Q.22 (1)

$\therefore$ 'B' has no vacant d-orbitals in its valence shell, so it can't extend its covalency beyond 4. i.e. 'B' cannot form the ion like $MF_6^{3(-)}$ i.e. $BF_6^{3(-)}$.

Hence, the correct option is (1).

Q.23 (2)

The structure of ClF_3 is



The number of lone pair of electrons on central Cl is 2 .

Q.24 (3)

For 2nd group hydrides, on moving down the group metallic character of metals increases so ionic character of metal hydride increases. Hence the option (3) should be correct option.

Q.25 (2)

NO : $(\sigma 1s)^2, (\sigma^*1s)^2, (\sigma 2s)^2, (\sigma^*2ps)^2, (\sigma 2p_z)^2, (\pi 2p_x)^2 = (\pi^*2p_x)^1 = (\pi^*2p_y)^0$

$$BO = \frac{10-5}{2} = 2.5$$

$CN^- = (\sigma 1s)^2, (\sigma^*1s)^2, (\sigma 2s)^2, (\sigma^*2ps)^2, (\pi 2p_x)^2 = (\pi 2p_x)^2, (\sigma 2p_y)^2$

$$BO = \frac{10-4}{2} = 3$$

$CN = (\sigma 1s)^2, (\sigma^*1s)^2, (\sigma 2s)^2, (\sigma^*2s)^2, (\pi 2p_x)^2 = (\pi 2p_y)^2, (\sigma 2p_z)^1$

$$BO = \frac{9-4}{2} = 2.5$$

$CN^+ = (\sigma 1s)^2, (\sigma^*1s)^2, (\sigma 2s)^2, (\sigma^*2s)^2, (\pi 2p_x)^2 = (\pi 2p_y)^2$

$$BO = \frac{8-4}{2} = 2$$

Hence, option(2) should be the right answer.

Q.26 (4)

Element (X) electronic configuration

$1s^2\ 2s^2\ 2p^3$

So, valency of X will be 3.

Valency of Mg is 2.

Formula of compound formed by Mg and X will be Mg_3X_2.

Q.27 (1)

NCERT Page No. 327

Pent-2-ene-4-yne

$CH_3-CH=CH-C\equiv CH$

$\sigma C-C = 4$

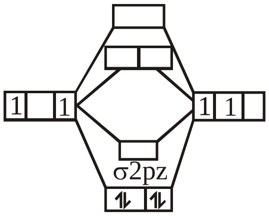
$\sigma C-H = 6$

$\pi C-C = 1$

$\pi \equiv C \equiv C = 2$

Q.28 (3)

Ace^4 MOT :-



C_2 contains 2π bond as it have 4π electron in molecular orbiritals

Q.29 (1)

PbF_4 and SnF_4 are exceptions they are ionic in nature

Q.30 (3)

Q.31 (4)

Q.32 (4)

Q.33 (2)

Q.34 (2)

Q.35 (3)

O_2 ion is having 15 electrons, so it contain one unpaired electron. Hence it is paramagnetic in nature.

JEE Main Previous Year's

Q.1 (3)

B_2H_6, LiH and BCl_3 are electron deficient

Q.2 (A)

$O_2 \rightarrow BO = 2$

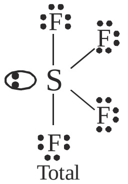
$O_2^+ \rightarrow BO = 2.5$

$O_2^- \rightarrow BO = 1.5$

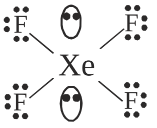
$O_2^{-2} \rightarrow BO = 1.0$

Q.3 (3)

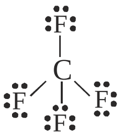
Allen Ans. (1)



Total lone pairs = 13

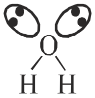


Total lone pairs = 14



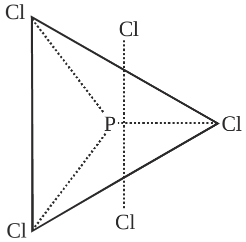
Total lone pairs = 12

Total Ione pairs = 2



Q.4 (2)

Q.5 **(C)**



PCl_5 $z = \frac{1}{2}[5+5] = 5 \leftarrow sp^3d$

trigonalbipyramidal

lone pair = Z - number of surrounding atoms

= 5 – 5 = 0

As axial bond is longer than equatorial bond

Hence 2 axial bonds are weaker than equatorial bond

Q.6 (C)

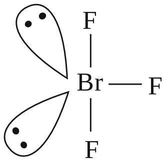
BrF_3

$Hyb = \frac{7+3}{1} = 5 sp^3d$

B.P + L.P = 5

3+2=5

Number of lone pair = 2



Shape : Bent T -shape

Q.7 (B)

NH_4

Number of electron = 6 + 4 = 10

$Hyb \Rightarrow \frac{4+4}{2} = 4 \rightarrow sp^3$

GEO $\Rightarrow$ Tetrahedral

NH_4^+

Number of electron = 7 + 4 - 1=10

$Hyb = \frac{5+4-1}{2} = 4 \rightarrow sp^3$

GEO $\Rightarrow$ Tetrahedral

BH_4^-

Number of electron = 5 + 4 + 1 = 10

$Hyb = \frac{3+4+1}{2} = 4 \rightarrow sp^3$

GEO $\Rightarrow$ Tetrahedral

$\Rightarrow$ They are isoelectronic and all have tetrahedral structures.

Q.8 (1)

The hybridization of P in PF_5 is sp^3d

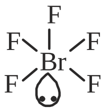
$\boxed{y = 1}$

Q.9 (B)

(A) $[PtCl_4]^{2-} \rightarrow dsp^2$
(B) $BrF_5 \rightarrow sp^3d^2$
(C) $PCl_5 \rightarrow sp^3d$
(D) $[Co(NH_3)_6]^{3+} \rightarrow d^2sp^3$

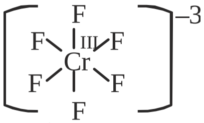
Q.10 [3]

(A)



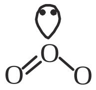
sp^3d^2 (square pyramidal)

(B)



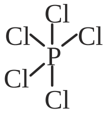
d^2sp^3 (octahedral)

(C)



sp^2(Bent)

(D)



sp^3d (T.B.P)

Q.11 [3]

Number of electrons in species		**Bond Order**
10	→	1
11	→	1.5
12	→	2
13	→	2.5
14	→	3
15	→	2.5
16	→	2
17	→	1.5
18	→	1

short trick

$CN^- = 14e^- \Rightarrow B.O. = 3$
$NO^+ = 14e^- \Rightarrow B.O. = 3$
$O_2 = 16e^- \Rightarrow B.O. = 2$
$O_2^+ = 15e^- \Rightarrow B.O. = 2.5$
$O_2^{2+} = 14e^- \Rightarrow B.O. = 3$

So, here CN^-, NO^+, O_2^{2+} have identical bond order

Q.12 (2)

Size of anion increase. Covalent character increase

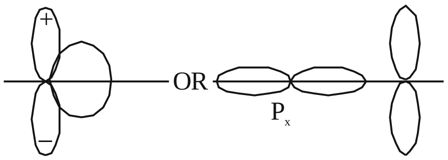
covalent character $\rightarrow CaF_2 < CaCl_2 < CaBr_2 < CaI_2$

Q.13 **[6]**

Molecules	**Plamar/ non planar**
NO_3^-	Planar
H_2O_2	non planar
BF_3	Planar
PCL_3	non planar
XeF_4	Planar
SF_4	non planar
XeO_3	non planar
PH_4^+	non planar

SO_3	Planar
$[Al(OH)_4]^-$	non planar

Q.14 (1)

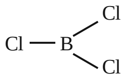


Zero overlapping

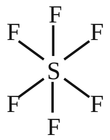
P_y
Zero overlapping
due to improper
orientation of
orbitals

Q.15 (2)

Expanded octet – Central atom has more than $8e^-$
Odd electron species ⇒ Number of total electrons in valence shell is odd

(i) BCl_3 ⇒


⇒ total electrons (valence shell) for boron ⇒ $6e^-$
for each chloride ⇒ $8e^-$

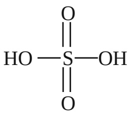
(ii) $5F_6$ ⇒


→ total electron (valence shell)

⇒ $12e^-$ For sulphur
6 × 2 ⇒ 12
for 'F' each ⇒ 8

(iii) No. ⇒ N = O total valence electron in nitrogen is 5e
Total valence electron in oxygen is 8e
So total electron is 13ē
⇓
∵ It is odd electronic species Odd number

$$Pbs \rightleftharpoons Pb^{2+} + s^{2-}$$

(iv) H_2SO_4 ⇒


$$HO-\overset{\overset{O}{||}}{\underset{\underset{O}{||}}{S}}-OH$$

→ Here also it is not odd electronic species

Now, if we want to identify expanded octet

(i) BCl_3	(ii) SF_6	(iii) NO	(iv) H_2SO_4
⇓	⇓	⇓	⇓
$6e^-$	$12e^-$	$5e^-$	$2e^-$

STATES OF MATTER

EXERCISE-I

Q.1 (4)

$\frac{P}{T}$ = constant (Gay Lussac's law)

$$\Rightarrow \frac{P_1}{T_1} = \frac{P_2}{T_2} \Rightarrow P_1T_2 = P_2T_1$$

PV = constant

$P_1V_1 = P_2V_2$ [Boyle's law]

Q.2 (3)

Q.3 (3)

Q.4 (4)

Q.5 (2)

Q.6 (1)

Q.7 (2)

RMS velocity of molecules depends on mass. If mol. wt. increases, rms velocity of melocules decreases.

$$\text{rms} \propto \frac{1}{\sqrt{\text{mol.wt}}}$$

The order of increasing m. wt. is

$$H_2 < N_2 < O_2 < HBr$$

Order of V_{rms} of molecules.

$$HBr < O_2 < N_2 < H_2$$

Q.8 (3)

$$r \propto U \text{and } U = \sqrt{\frac{3RT}{M}}$$

$$\therefore \frac{r_1}{r_2} = \sqrt{\frac{T_1M_2}{T_2M_2}} \; or \; \frac{r_{N_2}}{r_{SO_2}} = \sqrt{\frac{T_1 \times 64}{323 \times 28}} = 1.625$$

or $T_2 = 373\,K$

Q.9 (2)

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{16}{1}} = 4:1$$

Q.10 (2)

$$\overline{KE} = \frac{3}{2}kT = \frac{3}{2} \times \frac{8.313}{6.023 \times^{23}} \times 298 = 6.17 \times 10^{-21} J$$

(Average Kinetic energy $\overline{KE} = \frac{3}{2}kT = \frac{3}{2}\frac{R}{N}T$)

Q.11 (2)

Q.12 (1)

Q.13 (4)

Q.14 (4)

Q.15 (3)

Q.16 (1)

Q.17 (3)

Q.18 (4)

van der Waals's constant b = 4 times the actual volume of 1mole molecules = $4VN_0$

Q.19 (2)

Q.20 (4)

In an ideal gas, the intermolecular forces of attraction are negligible and hence it cannot be liquefied.

Q.21 (1)

Q.22 (3)

Q.23 (3)

EXERCISE-II (NEET LEVEL)

Q.1 (3)

Boyle's law is $V \propto \frac{1}{P}$ at constant T

Q.2 (4)

Boyle's Law

$V \propto \frac{1}{P}$ (Mass & Temperature constant)

Q.3 (1)

At constant $T, P_1V_1 = P_2V_2$

$1 \times 20 = P_2 \times 50$ **;**

Q.4 (1)

$P_1 = 5$ atm $V_1 = 10\,L$

$P_2 = 1$ atm $V_2 = ?$

$P_1V_1 = P_2V_2$ or $\frac{P_1}{P_2} = \frac{V_1}{V_2}$

$$\frac{5}{1} = \frac{V_2}{10}$$

$V\text{-}_2 = 50\,L$

Q.5 (4)
According to Boyle's Law *PV* = constant, at constant temperature either *P* increases or *V* increases both (*2*) & (*3*) may be correct.

Q.6 (3)

Graham's law of diffusion rate $\propto \frac{P}{\sqrt{\text{mol.wt}}}$

Q.7 (2)
Molecular weight = *V.d.* × 2 = 11.2 × 2 = 22.4
Volume of 22.4 *gm* Substance of NTP = 22.4 *litre*

1 *gm* substance at NTP = $\frac{22.4}{22.4}$ *litre*

11.2 *gm* substance of NTP = 11.2 *litre*

Q.8 (2)

$$\frac{M.wt.of\,O_2}{M.wt.of\,SO_2} \Rightarrow \frac{M_1}{M_2} \Rightarrow \frac{32}{64} = \frac{1}{2}$$

The weight of oxygen will be $\frac{1}{2}$ that of SO_2

Q.9 (4)
$P_{total} = P_1 + P_2 + P_3 + \ldots\ldots\ldots$
There should be non reacting gases, behave ideally and temperature of the component gases remain constant

Q.10 (3)
From Daltons partial pressure law

$$P_1 = \frac{1000}{1500} \times 600 + \frac{500}{1500} \times 800$$

$$= 400 + \frac{800}{3} = \frac{2000}{3} \text{ torr.}$$

$V_1 = 1500\,\text{mL}$
$P_2 = ?$
$V_2 = 2\text{L} = 2000\,\text{mL}$

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

$$\frac{P_1}{P_2} = \frac{V_2}{V_1} =$$

$$\Rightarrow \frac{2000/3}{P_2} = \frac{2000}{1500} \Rightarrow \qquad P_2 = \frac{1500}{3} = 500 \text{ torr}$$

Q.11 (3)

Initial $N_2 + 3H_2 \longrightarrow 2NH_3$
1 3
final – – 2

ratio = $\frac{4}{2} = \frac{2}{1}$.

Q.12 (1)
Rate of effusion ∝ Area of cross section

$$\therefore \quad \frac{r_A}{r_B} = \frac{A_A}{A_B} = \frac{\pi r^2}{r^2} = \frac{\pi}{1}$$

Q.13 (1)
$P = P_A + P_B + P_C$

Q.14 (4)

Rate of diffusion $\propto \frac{1}{\sqrt{\text{molar mass of gas}}}$

Q.15 (1)

$$r_g = \frac{1}{5}.r_{H_2}$$

$$\frac{M_g}{M_{H_2}} = \left[\frac{r_{H_2}}{r_g}\right]^2 = (5)^2 = 25 \;;\; M_g = 2 \times 25 = 50$$

Q.16 (2)

$$\frac{d_1}{d_2} = \frac{1}{16} \;;\; \frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{16} = \frac{4}{1}$$

Q.17 (2)

Most probable velocity = $\sqrt{\frac{2RT}{M}}$

Q.18 (1)

$$KE = \frac{3}{2}RT = \frac{3}{2}PV$$

$\therefore P = \frac{2}{3}\frac{E}{V}$ for unit volume (*V* = 1) ⇒ $P = \frac{2}{3}E$

Q.19 (1)
When average speed of molecule is increased due to increase in temperature then the change in momentum during collision between wall of container and molecules of gas also increases

Q.20 (1)

$$V_{av} : V_{rms} : V_{\text{most probable}} = V : U : \alpha$$

$$\sqrt{\frac{8RT}{\pi M}} : \sqrt{\frac{3RT}{M}} : \sqrt{\frac{2RT}{M}}$$

$$\alpha : V : U = \sqrt{2} : \sqrt{\frac{8}{\pi}} : \sqrt{3} \qquad = 1 : 1.128 : 1.224$$

Q.21 **(2)**

$$C_xH_y + O_2 \longrightarrow CO_2 + H_2O$$

$x \times n_{C_xH_y} = n_{CO_2}$ (POAC on C)

$x \times 500 = 2500$ $(x = 5)$

$y \times nC_xH_y = 2 \times n_{H_2O}$ (POAC on H)

$y \times 500 = 2 \times 3000$ $y = 12$

Formula $= C_5H_{12}$

Q.22 (1)

Z < 1 shows that the gas has greate compressibility at intermediate pressure and all reactive forces are dominant, higher value of 'a' and lower value of 'b' above Boyle's temperature it is not possible because Z > 1

Q.23 (2)

$Z = \frac{PV}{RT}$; for ideal gas $PV = RT$; so $Z = 1$

Q.24 (3)

Value of constant a is greater in NH_3 from other that's why NH_3 can be most easily liquefied

Q.25 (2)

If $Z < 1$ then molar volume is less than 22.4 L

$$Z = \frac{V_{real}}{V_{ideal}}$$

EXERCISE-III (JEE MAIN LEVEL)

Q.1 (3)

$$\frac{V_1}{V_2} = \frac{T_1}{T_2}$$

Q.2 (2)

$n_1 + n_2 = n_f$ $\frac{PV}{RT} = n$

$$\frac{1000 \times 500}{RT} + \frac{800 \times 1000}{RT} = \frac{P_f \times 2000}{RT}$$

$P_f = 650$ torr.

Q.3 (2)

n = const.

no of molecules = const

same number of molecules

Q.4 (3)

$$\frac{10}{V_2} = \frac{273}{373}$$

Q.5 (3)

$P = 1$

$\frac{n}{V} = 1$ mol/lt

$T = ?$

$PV = nRT$

$\Rightarrow 1 = 1 \times 0.0892 \times T \Rightarrow T = \frac{1}{0.0892}$

$T = \frac{1}{12}$; $T = 12$ K

Q.6 (4)

$$\frac{p_{H_2}}{p_{C_2H_6}} = \frac{n_{H_2}}{n_{C_2H_6}} = \frac{30}{2} = \frac{15}{1}.$$

Q.7 (4)

Weight of H_2 = 20 g in 100 g mixture ; Weight of O_2 = 80 g

$\therefore$ Moles of $H_2 = \frac{20}{2} = 10$; $\therefore$ Moles of $O_2 = \frac{80}{32} = \frac{5}{2}$

$\therefore$ Total moles $= 10 + \frac{5}{2} = \frac{25}{2}$

$\therefore P'_{H_2} = P_T$ x mole fraction of $H_2 = 1$ x $\frac{10}{25/2} = 0.8$ bar

Q.8 (4)

$$r \propto \frac{1}{\sqrt{M}}$$

Q.9 (2)

$$\frac{20}{60} \times \frac{30}{V} = \sqrt{\frac{32}{64}}.$$

Q.10 (3)

$$r \propto \frac{1}{\sqrt{M}}$$

So NH_3 diffuses with faster rate.

Q.11 (2)

$$\frac{r_1}{r_2} = \frac{t_2}{5} = \sqrt{\frac{M_2}{2}}$$

Q.12 (4)

Given $\frac{r_A}{r_B} = \frac{16}{3}$; $\frac{w_A}{w_B} = \frac{2}{3}$

we have $\frac{r_A}{r_B} = \frac{n_A}{n_B}\sqrt{\frac{M_B}{M_A}}$

$$\frac{16}{3} = \frac{w_A}{M_A}\frac{M_B}{w_B}\sqrt{\frac{M_B}{M_A}}$$

$$\frac{16}{3} = \frac{2}{3}\left(\frac{M_B}{M_A}\right)^{3/2} \Rightarrow \left(\frac{M_B}{M_A}\right)^{3/2} = 8 \Rightarrow \frac{M_B}{M_A} = 4$$

$\therefore$ mole ratio $= \frac{8}{3}$

Q.13 (4)

$C_3H_6O \longrightarrow CO_2$

$3 \times nC_3H_6O = nCO_2$

Q.14 (2)

$$\frac{u_1}{u_2} = \sqrt{\frac{T_1 \times M_2}{T_2 M_1}}$$

Q.15 (1)

$m_A = 2\,m_B$

$u_A = 2\,u_B$

$n_A = n_B$

$v_A = v_B$

$$\frac{P_A V_A}{P_B V_B} = \frac{\frac{1}{3} m_A n_A u_A^2}{\frac{1}{3} m_B n_B u_B^2}$$

Q.16 (2)

$$U_{rms} = \sqrt{\frac{3RT}{M}}$$

$$\frac{T_1}{M_1} = \frac{T_2}{M_2}$$

$$\frac{T_1}{32} = \frac{300}{20}$$

$T_1 = 480\text{ K}$

Q.17 (3)

$$\frac{u_1}{u_2} = \sqrt{\frac{\frac{3R \times 50}{2}}{\frac{2R \times 800}{32}}} = \frac{u_1}{u_2} = 1$$

Q.18 (2)

$$C_xH_y + \left(x + \frac{y}{4}\right)O_2 \longrightarrow xCO_2 + \frac{y}{2}H_2O$$

$$\frac{x + \frac{y}{4}}{x} = \frac{7}{4}$$

$$\frac{y}{4x} = \frac{3}{4} \quad \frac{y}{x} = \frac{3}{1}$$

Q.19 $C_4H_{10} + \frac{13}{2}O_2 \longrightarrow 4\,CO_2 + 5\,H_2O$

xml n-butane

y ml isobutane

Volume of $O_2 = x \times \frac{13}{2} + y \times \frac{13}{2}$

Q.20 (1)

$T_C = \frac{8a}{27Rb}$

$T_B = \frac{a}{Rb}$

$T_i = 2T_B$

$T_C < T_B < T_i$

Q.21 (3)

Factual question

Q.22 (1)

$(P)\,(V - nb) = nRT$

$$P = \frac{nRT}{V - nb}$$

Q.23 (3)

$PV = Pb + RT$

$$\frac{PV}{RT} = 1 + \frac{Pb}{RT}$$

Q.24 (1)

$$\left(P + \frac{a}{V^2}\right)(V) = RT$$

$$PV + \frac{a}{V} = RT$$

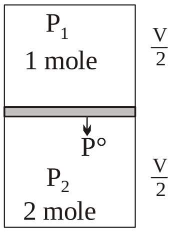
$$\frac{PV}{RT} = 1 - \frac{a}{VRT}$$

Q.25 (1)

$$4 \times \frac{4}{3}\pi r^3 \times N_A = 24$$

EXERCISE-IV

Q.1 0090



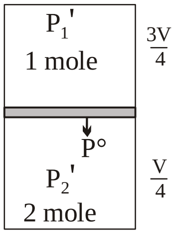
at 300K

P° is the pressure
due to weight of piston

$$P_1 + P° = P_2$$

$$P° = \frac{2 \times R \times 300}{\frac{V}{2}} - \frac{1 \times R \times 300}{\frac{V}{2}}$$

at temperature T



$$P_1' + P° = P_2' \qquad P° = \frac{2 \times R \times T}{\frac{V}{4}} - \frac{1 \times R \times T}{\frac{3V}{4}}$$

$$\frac{2 \times R \times 300}{\frac{V}{2}} - \frac{1 \times R \times 300}{\frac{V}{2}} = \frac{2RT}{\frac{V}{4}} - \frac{1 \times R \times T}{\frac{3V}{4}}$$

Solving T = 90 K

Q.2 0001

$V_1 = 500$ ml
$V_2 = 1800$ ml
$P_{total} = 3.2$ atm
$T_2 = 360$ K
$T_1 = 320$ K
V.P. = 0.6 atm
V.P. = 0.2 atm

$P_{gas} = 3.2 - 0.2 = 3$ atm
For H_2O(vap.)
For gas :

$$\frac{0.2 \times 500}{320} = \frac{P_2 \times 1800}{360}$$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} = \frac{3 \times 500}{320} = \frac{P_2 \times 1800}{360}$$

$$P_2 = \frac{20}{320} = \frac{2}{32} \text{atm } (< \text{V.P.})$$

$$P_2 = \frac{6 \times 500}{320 \times 10} = \frac{30}{320} \text{ atm}$$

$$P_{total} = \frac{30}{32} + \frac{2}{32} =$$

1 atm

Q.3 (4) $\left(\sqrt{\frac{3RT}{m}}\right)_X = \left(\sqrt{\frac{2RT}{m}}\right)_Y$

$$\sqrt{\frac{3R400}{40}} = \sqrt{\frac{2R60}{m}}$$

$$\sqrt{30} = \sqrt{\frac{4 \times 30}{m}}$$

m = 4

Q.4 72

$$\frac{1}{6} = \sqrt{\frac{2}{x}}$$ (Where X is molecular weight of gas)

$$\frac{1}{36} = \frac{2}{x}$$

x = 72

Q.5 (4)
Q.6 (*)
Q.7 (2)
Q.8 (3)
Q.9 (1)
Q.10 (4)

PREVIOUS YEAR'S

MHT CET
Q.1 (4)
Q.2 (1)
Q.3 (3)

Q.4 (2)
$T_1 = 273$K
$T_2 = 373$K
$V_1 = 10$ L
$V_2 = ?$
From charles law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$V_2 = 373 \times \frac{10}{273} = 13.66\text{L} = 13.66\,\text{dm}^3$$

Q.5 (3)

$T_1 = 50K$ for H_2

$T_2 = 800$ K for O_2

$$\mu_{rms} = \sqrt{\frac{3RT}{M}}$$

$$\Rightarrow \frac{\mu_{H_2}}{\mu_{O_2}} = \sqrt{\frac{T_1}{M_1} \times \frac{M_2}{T_2}} = \sqrt{\frac{50}{2} \times \frac{32}{800}} = 1$$

Q.6 (3)

Ideal gas equation

$$PV = \eta RT$$

is obeyed by an idela gas in both adiabatic and isothemal processes.

Q.7 (1)

According to Graham's law, $\frac{r_A}{r_B} = \sqrt{\frac{MB}{MA}}$

Here , $M_A = \left(\frac{100}{2}\right)$ Kg/ molecule

Q.8 (3)

PREVIOUS YEAR'S

NEET/AIPMT

Q.1 (3)

* van der waal constant 'a', signifies intermolecular forces of attraction.

* Higher is the value of 'a', easier will be the liquefaction of gas.

Q.2 (4)

In real gas equation, $\left(P + \frac{an^2}{V}\right)(V - nb) = nRT$ van der Waal's constant, 'a' signifies intermolecular forces of attraction.

Q.3 **(3)**

$V_i = V,$

$V_r = V - 0.2\,V \qquad = 0.8\,V$

if value of Z < 1 then attractive forces are dominant.

$$\left(Z = \frac{V_r}{V_i} = 0.8\right)$$

Q.4 (2)

Q.5 (3)

Q.6 (3)

$V = 10\,L \qquad W_{O_2} = 64g$

$T = 27^o \qquad n_{O_2} = 2$

$R = 0.086.$ L bar K^{-1} mol^{-1}

Ideal gas equation PV = nRT

$$P = \frac{2 \times 0.0831 \times 300}{10}$$

P = 4.9 bar

Q.7 (3)

Dalton's law of partial pressure :

Partial pressure of gas = mole fraction of gas in gaseous mixture × Total pressure of gaseous mixture.

$p_1 = X_1 p$

$p_2 = X_2 p$

$p_3 = X_3 p$

Total pressure,

$p = p_1 + p_2 + p_3$

Therefore, statement-3 is incorrect.

JEE MAIN

Q.1 [45]

$$(PV)_{gas} = \left(\frac{W_{gas}}{MM_{gas}}\right)RT_{gas}$$

$$(PV)_{H_2} = \left(\frac{W_{H_2}}{MM_{H_2}}\right)RT_{H_2}$$

According to question $(PV)_{gas} = (PV)_{H_2}$

$$\left(\frac{3}{M_{gas}}\right)300 = \left(\frac{0.2}{2}\right)200$$

$$M_{gas} = \frac{3 \times 300 \times 2}{0.2 \times 200}$$

$M_{gas} = 45$

Q.2 (9960)

$$V_{H_2(g)} = \frac{nRT}{P} \Rightarrow \frac{\left(\frac{2}{2}\right) \times 0.0831 \times 300}{1} \Rightarrow 24.9L$$

$\therefore$ Volume of H_2(g) adsorbed on 2.5 g Pt=24.9 L

$\therefore$ Volume of H_2(g)adsorbed on 1 g Pt $= \frac{24.9}{2.5} = 9.96L$

We know 1L=1000 ml

So 9.96 L $\Rightarrow$ 9960 ml.

Q.3 (4)

Weight of liquid = 135 – 40 = 95g

Volume of liquid $= \frac{mass}{density} = \frac{95}{0.95} = 100ml$

Volume of vessel = 100 ml

PV = nRT

$$0.82\times\frac{100}{1000}=\text{nRT}$$

$$\text{n}=\frac{0.82\times\frac{100}{1000}}{0.0821\times 250}=\frac{1}{250}$$

mass of ideal gas 40.5–40 ⇒ 0.5 g

We know, $\text{n}=\frac{\text{Given mass}}{\text{M.mass}}$

$$\Rightarrow\frac{1}{250}=\frac{0.5}{\text{M Mass}}$$

M Mass = 0.5×250
= 125 g

Q.4 (29)

$$\text{V}=\frac{\text{nRT}}{\text{P}}=\frac{0.90\times 0.82\times 300\times 760}{18\times 32}29.21$$

Q. 5 (4)

$$1.5\times 4.16\frac{100}{\text{M}}\times 0.83\times 300$$

M = 3.99

Q.6 (2)

PM = dRT

Where P = Pressure of an ideal gas; d = density of ideal gas

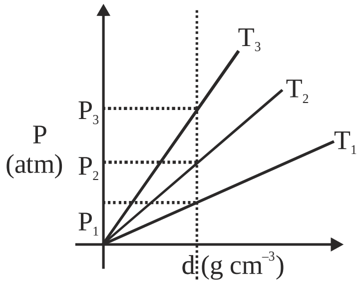
R = Universal gas constant; T = Temperature

P ∝ density

on plotting a vertical line, where to put density is constant

P ∝ Temperature $P_3 > P_2 > P_1$

from there $T_3 > T_2 > T_1$



Q.7 [1655]

Pv = nRT

$$2\times 10^{6}\times 2\times 10^{-3}=\frac{11}{44}\times 8.3\times \text{T}$$

T = 1927.7 K
T = 1654.7°C

Q.8 [2]

$H_2 + O_2$
↓
40% by mass

Let mass of mixture = 100 gm

Mass of H_2 = 40 gm

Mass of O_2 = 60 gm

$$\text{Mole of } H_2=\frac{40}{2}=20\text{ mol}$$

$$\text{Mole of } O_2=\frac{60}{32}\text{ mol}=\frac{15}{8}\text{ mol}$$

Partial pressure of H_2

$$=\frac{{}^{n}H_2}{{}^{n}H_2+{}^{n}O_2}\times P_{\text{total}}\quad=\frac{20}{20+\frac{15}{8}}\times 2.2$$

$$=\left(\frac{20\times 8}{160+15}\right)\times 2.2\quad=\frac{160\times 2.2}{175}$$

$$=\frac{352}{175}=2.011\approx 2\text{ bar}$$

Q.9 [25]

p = 99 bar z = 2 b = ?

= 99 × 0.987 atm

Vander wall equation –

$$(P+\frac{a}{v^2})(V-b)=RT\text{ for 1 mole}$$

At high p $\rightarrow\frac{\alpha}{v^2}\rightarrow$ Can be neglect

But b → can not be neglected

$$P(V-b)=RT$$
$$PV-Pb=RT$$
$$\frac{PV}{RT}=\frac{pb}{RT}+\frac{RT}{RT}$$
$$\downarrow$$
$$Z=\frac{Pb}{RT}+1$$
$$2=\frac{Pb}{RT}+1$$
$$\frac{pb}{RT}=2-1=1$$

$99\times 0.987\times b = 8.3\times 298$

$b = 0.253 = 25.3\times 10^{-2}$

Ans.: 25 (nearest integer)

Q.10 [1]

Normality of Acid - Base mixture solutions

$$N_{MIX} = \left|\frac{N_A V_A - N_B V_B}{V_A + V_B}\right|$$

$$\left|\frac{M_A n_A v_A - M_B n_B v_B}{V_A + V_B}\right|$$

$$= \left|\frac{0.1 \times 2 \times 100 - 0.1 \times 1 \times 50}{100 + 50}\right|$$

$$= \frac{20-5}{150} = \frac{15}{150} = 0.1N$$

Ans: 1x 10^{-1}, Ans (1)

Q.11 (2)

$$P_A = P_A^0 \times X_A = P_{total} \times Y_A$$

$$\Rightarrow P_A^0 \times 0.2 = 0.8 \times 0.5$$

$$P_A^0 = 2atm$$

Q.12 (2)

$$\lambda = \frac{h}{\sqrt{2mK}}$$

$$K = \frac{h^2}{2m\lambda^2}$$

$$K = \frac{h^2}{2m\lambda^2} = \frac{43.9 \times 10^{-68}}{2 \times 9.1 \times 10^{-31} \times 10.89 \times 10^{-20}}$$

$$K = 2.215 \times 10^{-18}$$

$$E_{abs} = E_{req} + K$$

$$\frac{E_{abe}}{E_{req}} = 1 + \frac{K}{E_{req}} = 1 + \frac{2.215 \times 10^{-18}}{13.6 \times 1.602 \times 10^{-19}} = 2.0166$$

Q.13 (80)

$$P_{Ne} = P_{total} \times X_{Ne}$$

$$\Rightarrow 20 = 25 \times X_{Ne}$$

$$[X_{Ne}] = \frac{20}{25} = \frac{4}{5}$$

$$\Rightarrow \left[\frac{\frac{200}{20}}{\frac{200}{20} + \frac{x}{32}}\right] = \frac{4}{5}$$

$$\Rightarrow \frac{10}{10 + \frac{x}{32}} = \frac{4}{5}$$

$$\Rightarrow 400 = 320 + x$$

$$\Rightarrow x = 80$$

Q.14 (32)

At constant volume and mole

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$\frac{30}{300} = \frac{P_2}{318}$$

$$P_2 = \left(\frac{30}{300}\right) \times 318 = 31.8 \approx 32$$

Q.15 [22]

Press . of moist = 4 atm

Press. of gas = 4-v.p of H_2O

$= 4 - 0.4$

$= 3.6$

When volume is doubled ⇒ P = P/2

$$= \frac{3.6}{2} = 1.8$$

Total Press. = 1.8 + v.p of H_2O

$= 1.8 + 0.4 = 2.2$ atm

Thermodynamics

EXERCISE-I (MHT CET LEVEL)

Q.1 (4)

Q.2 (1)

Q.3 (1)

$$q = \frac{c}{m}$$

$$\Rightarrow C_v = q \times m$$

$= 0.075 \times 40 = 3.0$ cal

$C_P - C_V = R$

$C_P = C_V + R = 3 + 2 = 5$

$$\frac{C_P}{C_V} = \frac{5}{3} = 1.66$$

Monoatomic gas

Q.4 (4)

Q.5 (2)

$$N_2 + 3H_2 \leftrightarrow 2NH_3$$

According to thermodynamics's Ist law

$$\Delta H = \Delta E + nRT$$

Where $\Delta H =$ enthalpy of reaction at constant pressure

$\Delta E =$ heat of reaction at constant volume

R= molar gas constant

T= temperature of the reaction

N= (no. of moles of product)- (no. of moles of reactant)

From reaction, $n = n_p - n_R = 2 - 4 = -2$

Hence $\Delta H = \Delta E - 2RT$

Q.6 (2)

For isothermal process, $\Delta E = 0$

Q.7 (2)

$$\Delta n = \frac{1}{2}; \Delta H = \Delta E - \frac{1}{2}RT$$

$\Rightarrow \Delta E > \Delta H$

Q.8 (3)

Q.9 (3)

Q.10 (4)

Condition of equilibrium, hence $\Delta G = 0$.

Q.11 (1)

Since, in the first reaction gaseous products are forming from solid carbon hence entropy will increases i.e. $\Delta =$ + ve.

$$C(gr.) + \frac{1}{2}O_2(g) \rightarrow CO(g); \Delta S° = + ve$$

Since, $\Delta G° = \Delta H° - T\Delta S$ hence the value of ΔG decrease on increasing temperature.

Q.12 (1)

Given enthalpy of vaporization,

$\Delta H = 186.5\ kJmol^{-1}$

Boiling point of water = $100^0C = 100+273 = 373$ k

Entropy change,

$$\Delta S = \frac{\Delta H}{T} = \frac{186.5 KJmol^{-1}}{373K}$$

Q.13 (4)

$\Delta G = \Delta H - T\Delta S$; ΔG is positive for a reaction to be non-spontaneous when ΔH is positive and ΔS is negative.

Q.14 (2)

At equilibrium $\Delta G = 0$

Hence, $\Delta G = \Delta H - T_e \Delta S = 0$

$$\therefore\ \Delta H = T_e \Delta S \text{ or } T_e = \frac{\Delta H}{\Delta S}$$

For a spontaneous reaction Δ G must be negative which is possible only if Δ H–T Δ S<0

$\therefore\ \Delta H = < T \Delta S$

$$or\ T > \frac{\Delta H}{\Delta S} \qquad \text{So } T > T_e$$

Q.15 (1)

Given ΔH 35.5 kJ mol^{-1}

$\Delta S = 83.6\ JK^{-1}\ mol^{-1}$

$\because \Delta G = \Delta H - T\Delta S$

For a reaction to be spontaneous, $\Delta G = -ve$

i.e., $\Delta H < T\Delta S$

$$\therefore\ T > \frac{\Delta H}{\Delta S} = \frac{35.5 \times 10^3 Jmol^{-1}}{83.6JK^{-1}mol^{-1}}$$

So, the given reaction will be spontaneous at T > 425 K.

Q.16 (4)

$$H_2O(\ell)\xrightarrow{1atm}H_2O(g)$$

$$\Delta H = 40630 Jmol^{-1}$$

$$\Delta S = 108.8 JK^{-1}mol^{-1}$$

$\Delta G = \Delta H - T\Delta S$ when $\Delta G = 0$,

$$\Delta H - T\Delta S = 0$$

$$T = \frac{\Delta H}{\Delta S} = \frac{40630 Jmol^{-1}}{108.8 Jmol^{-1}} = 373.4K$$

Q.17 (2)

$$\Delta G^\circ = -RT\ln K_P; K_P = (2X)^2\ X = 4X^3$$

$$\Delta G^\circ = -RT\ell n 4X^3$$

$$\Delta G^\circ = -RT\ln 4 - 3RT\ln X$$

Q.18 (4)

ΔG = ΔH – TΔS = +ve

for spontaneous change, ΔH < 0, ΔS > 0

for non-spontaneous change, ΔH < 0, ΔS > 0

Q.19 (1)

Q.20 (3)

Q.21 (2)

Q.22 (3)

Q.23 (4)

Q.24 (2)

Q.25 (1)

Q.26 (1)

Q.27 (3)

The third law helps to calculate the absolute entropies of pure substances at different temperature.

Q.28 (1)

Q.29 (3)

Ozone is allotropic from of oxygen and is less stable than oxygen.

Q.30 (3)

Formation of CO and CO_2 illustrates the law of multiple proportion that is constant mass of C reacts with different masses of oxygen. These masses here bears simple ratio of 1:2.

Q.31 (3)

Q.32 (3)

By bomb calorimeter we get ΔE .

$$2C_6H_6(l) + 15O_2(g) \rightarrow 12CO_2(g) + 6H_2O(l)$$

$$\Delta H - \Delta E = \Delta nRT$$

$$= (12-15)\times 8.314\times 300 = -7.483\text{kJ}$$

Q.33 (1)

Q.34 (1)

Q.35 (1)

$$\frac{1}{2}H_2 + \frac{1}{2}X_2 \longrightarrow HX$$

Let the bond enthalpy of X – X bond be x.

ΔH_f (HX) = – 50

$$= \frac{1}{2}\Delta H_{H-H} + \frac{1}{2}\Delta H_{X-X} - \Delta H_{H-X}$$

$$\frac{1}{2}2x + \frac{1}{2}x - 2x = \frac{-x}{2}$$

$\therefore$ x = 50 × 2 = 100 kJ mol^{-1}

Q.36 (1)

$$\Delta H = \Delta H_{(\text{product })} - \Delta H_{(\text{reactant})}$$

$$162 = 2\times\Delta H_H - \Delta H_{H_2}$$

$$\Delta H_H = \frac{162}{2}\quad (\because \Delta H_{H_2} = 0)$$

$$\Delta H_H = 81\text{Kcal}$$

Q.37 (4)

Q.38 (2)

Heat needed to be supplied per mol

= 330+ 580 + 1820 + 2740 = 5470 kJ

No. of mols of Al taken $= \frac{13.5}{27} = 0.5\text{mol}$

Q.39 (3)

Appliying Hess's Law

$$\Delta_f H^o = \Delta_{sub} H + \frac{1}{2}\Delta_{diss} H + I.E + E.A + \Delta_{lattice} H$$

–617 = 161 + 520 + 77 + E.A. + (–1047)

E.A. = –328 KJ mol^{-1}

$\therefore$ electron affinity of fluorine = –328 KJ mol^{-1}

Q.40 (2)

Q.41 (Bonus)

Q.42 (4)

EXERCISE-II (NEET LEVEL)

Q.1 (3)

In isolated system neither exchange of matter nor exchange of energy is possible with surroundings.

Q.2 (4)

Q.3 (3)

Work is not a state function as during a process its value depends on the path followed. The value of enthalpy, internal energy and entropy depends on the state and not on the path followed to get that state, hence these are state functions.

Q.4 (3)

Surface tension is an intensive property which do not depend upon the quantity of matter present in the system.

Q.5 (4)

During adiabatic process, no heat transfer occurs.

Q.6 (3)

Q.7 (2)

$$W = -2.303\ nRT \log \frac{V_2}{V_1}$$

$$= -2.303 \times 1 \times 8.314 \times 10^7 \times 298 \log \frac{20}{10}$$

$$= -298 \times 10^7 \times 8.314 \times 2.303 \log 2 .$$

Q.8 (3)

$$W = -P\Delta V = -1 \times 10^5 (1 \times 10^{-2} - 1 \times 10^{-3})$$

$$= -1 \times 10^5 \times 9 \times 10^{-3} = -900\ J$$

Q.9 A

We have,

$$W = -\,2.303\ nRT \log \frac{V_2}{V_1}$$

$$= -\,2.303 \times 2 \times 2 \times 298 \times \log \frac{50}{15}$$

$= -1426.87$ calories.

Q.10 D

If $n_g = 0$ $\Delta H = \Delta E$

Q.11 (2)

Q.12 (2)

$\Delta H = \Delta E + P\Delta V \cdot$

Q.13 (3)

$\Delta n_g = 1 - \frac{3}{2} = \frac{-1}{2}$, As Δn_g is negative, thus

$\Delta H < \Delta E \cdot$

Q.14 **(2)**

According to FLOT

$\Delta H = \Delta E + P\Delta V$

So $\Delta H = \Delta E - P\Delta V$ should be incorrect.

Q.15 (3)

At constant P or T

$\Delta H = \Delta U + \Delta nRT \Rightarrow \Delta n = n_p - n_R = 2 - 4 = -2$

$\therefore \Delta H < \Delta U \cdot$

Q.16 (3)

According to FLOT; $\Delta U = q + w$

Q.17 (3)

Q.18 (1)

We know that $\Delta E = Q + W = 600 + (-300) = 300\ J$

Q.19 (1)

We know that work done, $W = C_v(T_1 - T_2)$ (for reversible adiabatic process)

$3 \times 1000 = 20\,(300 - T_2)\,;\ \therefore 3000 = 6000 - 20\,T_2$

$\therefore T_2 = \frac{3000}{20} = 150\ K$.

Q.20 (3)

For reversible spontaneous change;

$$\Delta S = \frac{q}{T}$$

Q.21 (4)

When $\Delta S = +ve$ the change is spontaneous.

Q.22 (4)

The role of calcination in metallurgical operations is to remove moisfure, to decompose carbonate & to drive off organic matter

Q.23 (3)

The plot of ΔG° verrus temperature for the formation fo oxide of elements is called ellingation diagram

Q.24 (2)

Solid $\longrightarrow$ Gas, ΔS is maximum.

Q.25 (1)

the most electropositive metals are obtained from their ores by electrolysis of their fused ionic salts like : - NaCl

Q.26 (2)

Q.27 (1)

When $\Delta H = -ve$, $\Delta S = +ve$ and $\Delta G = -ve$ than reaction is spontaneous .

Q.28 (4)

At equilibrium; $\Delta G = 0$

Q.29 (2)

Aluminium is produced on a large scale by electrolysis of alumina dissolved in fused cryolite and little fluorspar & the formula of cryolite is Na_3AlF_6 & the formula of fluorspar is CaF_2

Q.30 (4)

When $\Delta H = +ve$ and $\Delta S = -ve$ than the reaction is non-spontaneous.

Q.31 (2)
Hess law's statement.

Q.32 (4)
Heat of neutralisation of a strong acid and strong base is equal to $-13.7\, kcal$.

Q.33 (4)
Heat of neutralisation will be maximum for strong acid and a strong base is about $-13.7\, Kcal$.

Q.34 (2)
Effect of temperature in heat of reaction is given by Kirchoff's equation.

EXERCISE-III (JEE MAIN LEVEL)

Q.1 (3)
Boiling point, pH & density are intensive properties. Entropy is an extensive property.

Q.2 (2)

Q.3 (1)
PV = nRT [V in litre and T in Kelvin]
PV = 1 × 0.0821 × 373
PV = 30 litre
V = 30 litre
$W = -P\Delta V$
$= -1\,(30 - V_{gas})$ [18 ml is negligible as compared to 30 litre]
= - 1 × 30 = – 30 litre atm.

Q.4 (1)
$W = -10^5 \times 2 = -2 \times 10^5$ J

Q.5 (3)
$w = -n\,C_v (T_2 - T_1)$
$T_2 = 150$ K

Q.6 (1)
$$C_P = \frac{q}{n\Delta T}$$
$\Delta T = 0$
$C_P = \infty$

Q.7 (2)

Q.8 (1)

Q.9 (2)

Q.10 (2)

Q.11 (4)
$C_2H_4 + HCl \rightarrow C_2H_5Cl$
$\Delta n_g = -1$
HCl is limiting reagent
$\Delta U = \Delta H - \Delta n_g RT$
$= -185 + (1 \times 8.314 \times 10^{-3} \times 300)$
$= -69.8$ kJ
For two mole of HCl $= -69.8 \times 2 = -139.6$ kJ

Q.12 (2)
W = –100J , q = 150 J
$\because q = \Delta E - W$
$\therefore 150 = \Delta E - (-100)$
$= 150 - 100$
$\Delta E = 50$ J

Q.13 (1)
$W = -P\Delta V$
$= -1.5\,(1 - 0.5)$
$= -0.75$ atm-litre $= -0.75 \times 101$ Joule
q = 200 J
$\because q = \Delta E - W$
$200 = \Delta E - (-0.75 \times 101)$
$\Delta E = 124.25$ Joule

Q.14 (4)
$$U = 0,\ W = -nRT \ln\left(\frac{V_2}{V_1}\right) = P_1V_1 \ln\left(\frac{V_2}{V_1}\right) = -20 \ln 2 = -14 \text{ lit-atm}$$

Q.15 (2)
$$\Delta S_f = \frac{\Delta H_f}{T} = \frac{6025}{273}$$
$= 22.1$ JK^{-1} mol^{-1}

Q.16 (1)
$$\Delta S_{system} = nR\ln\frac{V_2}{V_1}$$
$\Rightarrow 2 \times R \times \ln 2$
$\Rightarrow$ 11.52 J/K
$$\Delta S_{surrounding} = -\frac{3.3 \times 1000}{300} \Rightarrow -11 \text{ J/K}$$

Q.17 (3)
$$\Delta S_{vap} = +\frac{300 \times 30}{300} = -30 \text{ J/mol.K}$$
$\Delta S_{cond} = -30$ J/mol-K

Q.18 (1)
Sol. $\Delta H = -3600$ cal
$\Delta G = \Delta H - T\Delta S$
$\Delta G = -600$ cal

Q.19 (1)
$\Delta G^o = -RT \ln k$

Q.20 (4)

Q.21 (2)

Q.22 (1)

Q.23 (2)

Q.24 (1)

Q.25 (1)

EXERCISE-IV

Q.1 1900 cal

$\Delta U = \Delta U_1 + \Delta U_2$

$= n_1 C_{V_1 m} \Delta T + n_2 C_{V_1 m} \Delta T$

$= [3 \times \frac{3}{2} R \times 100] + [2 \times \frac{5R}{2} \times 100]$

$= 450\,R + 500\,R = 950\,R = 1900$ cal

Q.2 $\Delta E = 0.993$ kcal, $\Delta H = 1$ kcal

$w = -P_{ext}(V_2 - V_1) = -1\,(1.5 - 1.2) = -0.3$ L atm $= -7.235$ cal

$\Delta U = q + w$

$= 1 - 0.007235 = 0.993$ kcal

$q = \Delta H = 1$ kcal

Q.3 200 J/K mole

$$\Delta S = \frac{\Delta H}{T} = \frac{-392 - 75 - (-484)}{85} = 200 \text{ J/K mole}$$

Q.4 – 88 kJ/mol

$C_2H_6 + \frac{7}{2} O_2 \rightarrow 2CO_2 + 3H_2O$ $\Delta H = -1560$

..................(1)

$C + O_2 \rightarrow CO_2$ $\Delta H = -395$

...................(2)

$H_2 + \frac{1}{2} O_2 \rightarrow H_2O$

$\Delta H = -286$ ....(3)

Target reaction $2C + 3H_2 \rightarrow C_2H_6$ $\Delta H = ?$

ΔH can be obtained by (2) × 2 + (3) × 3 – (1)

$\Delta H = -88$ kJ mol^{-1}

Q.5 (4)
Q.6 (4)
Q.7 (3)
Q.8 (3)
Q.9 (1)
Q.10 (2)

PREVIOUS YEAR'S

MHT CET

Q.1 (3)
Q.2 (1)
Q.3 (3)
Q.4 (2)
Q.5 (3)
Q.6 (2)
Q.7 (2)
Q.8 (2)
Q.9 (3)
Q.10 (1)
Q.11 (2)
Q.12 (3)
Q.13 (4)
Q.14 (3)
Q.15 (3)
Q.16 (4)
Q.17 (2)
Q.18 (2)
Q.19 (1)
Q.20 (2)
Q.21 (2)
Q.22 (1)
Q.23 (1)
Q.24 (4)
Q.25 (2)
Q.26 (4)
Q.27 (2)
Q.28 (1)
Q.29 (3)
Q.30 (1)
Q.31 (2)
Q.32 (3)
Q.33 (2)
Q.34 (1)
Q.35 (1)
Q.36 (2)
Q.37 (4)
Q.38 (1)
Q.39 (3)
Q.40 (3)
Q.41 (2)
Q.42 (3)

Q.43 (1)
Q.44 (1)
Q.45 (3)
Q.46 (Bonus)
Q.47 (3)
Q.48 (2)
Q.49 (1)
Q.50 (4)
Q.51 (3)
Q.52 (4)

$N_2 + 3H_2 \rightarrow 2NH_3$

$\Delta_f H\,(NH_3)$ = Bond energy of reactant - Bond energy of product

$$= \left(\frac{1}{2} \times \text{BE of } N \equiv N \text{ bond} + \frac{3}{2} \text{BE of } H - H \text{ bond}\right)$$

$-$ (3 BE of $N - H$ bond)

Q.53 (2)

Standard enthalpy of formation is the change in enthapy that accompanies the formation of one mole of a compound from its elements with all substances in their standard standard standard states it is also called as' standard heat of formation '

Q.54 (1)

$S + O_2 \rightarrow SO_2;\ \Delta H_f = -4.6$ KJ

0.5 g of sulphur on burning produce 1 g of SO_2

$\therefore$ 32 g of suplphur on buring will product 64 g of SO_2

$\therefore \quad \Delta H_f = (-4.6\text{KJ}) \times 64$

$= -294.4\ \text{kJ mol}^{-1}$

Q.55 (3)

In the given reaction

$C_2H_4\,(g) + 3O_2\,(g) \rightarrow 2CO_2\,(g) + 2H_2O\,(1)$

$\Delta H = -1410.0$ KJ

For ideal gas $\Delta H = \Delta U + RT(\Delta n_g)$

where, R = 8.314 J/ mol K

T = 298 K

Δn_g = number of gaseous products – number of gaseous reactants

$= 2 - 4 = -2$

$\therefore -1410 = \Delta U + [8.314 \times 298(-2)]$

$\Delta U = -1405.045$ KJ

$\simeq -1405.05$ KJ

Now $\Delta H - \Delta U = -1410 - (-1405.05) = -4.95$ KJ

Q.56 (2)
Q.57 (2)
Q.58 (2)
Q.59 (3)
Q.60 (3)

PREVIOUS YEAR'S

NEET/AIPMT

Q.1 (1)

The reaction for $\Delta_f H°(XY)$

$$\frac{1}{2}X_2(g) + \frac{1}{2}Y_2(g) \longrightarrow XY(g)$$

Bond energies of X_2, Y_2 and XY are X, $\frac{X}{2}$, X respectively.

$$\therefore \Delta H = \left(\frac{X}{2} + \frac{X}{4}\right) - X = -200$$

On solving, we get

$$\Rightarrow -\frac{X}{2} + \frac{X}{4} = -200$$

X = 800 kJ / mole

Q.2 (1)

A gas expands against a constant external pressure is irreversible process.

The work done in irreversible process $= -P_{ext}\Delta V$

$= -P_{ext}(V_2 - V_1)$

$= -2(0.25 - 0.1)$

$= -2 \times 0.15$ bar – L/L-bar

$= -30 \times 100$ J

$= -30$ J

and work done by the gas is + 30 J

Q.3 (4)

Entropy = measurement of disorderness if $\Delta n_g < 0$ then $\Delta S < 0$

Q.4 (3)
Q.5 (4)
Q.6 (2)
Q.7 (1)
Q.8 (1)

In P-V graph area under the curve represent magnitude of work.

As it is maximum in graph-1

So correct answer is (1)

JEE MAIN

Q.1 [300]

$\Delta G° = \Delta H°\ T\Delta S° = 0$ at equilibrium

$\Rightarrow -165 \times 10^3 - T \times (-505) = 0$

$\Rightarrow T = 300$ K

$= 300$

Q.2 (3)

$\Delta H_r = (\Delta H_C)_{reactant} - (\Delta H_C)_{product}$

$= 3\,[\Delta H_C(C_2H_2,g)] - [\Delta H_C\,(C_6H_6,\ell)]$

$= 3 \times (-1300) - (-3268)$

$= -3900 + 3268 = -632$ kJ/mole

Q.3 (727)

$$CH_3OH(\ell)\frac{+3}{2}O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

We know, $\Delta H = \Delta U + \Delta n_g RT$

$$= -726 + \left(\frac{-1}{2}\right) \times 8.314 \times 10^{-3} \times 300$$

$= -726 - 1247.1 \times 10^{-3}$

$= -726 - 1.2471$

$= -727$

So, x=727

Q.4 (38)

T = 373 K

Given weight of water = 36 gm

Moles of water = $\frac{36}{18}$ = 2 moles

Given – $\Delta_{vap}H° = 41.1$ kJ/mol

$\Delta U = \Delta H - nR\Delta T$

$= 41.1 - 8.31 \times 373 \times 10^{-3}$

$= 38$ kJ mol^{-1}

Q.5 (2)

Q.6 (0)

In case of free expansion, $w = -(P_{ext} \times \Delta V)$

$P_{ext} = 0 \quad \therefore w = 0$

First law of thermodynamics; $\Delta \bigcup = q = w$

$\Delta \bigcup = 0$

$q = -w = 0$

Amount of heat absorbed = 0

Q.7 (51)

51 kJ/mol.

$Pt(s)|H_2(g)|H^+(aq)||Ag^+(aq)|Ag(s)$

$E^0_{cell} = +0.5332$ V. $\quad Ag^+ + e^- \rightarrow Ag$

$\Delta_f G^0 = -nF E^0_{cell}$

$= -1 \times 96500 \times 0.5332$

$= -51.453$ kJ/mol.

$= 51$ kJ/mol.

Q.8 (8630)

n = 5; T = 300; V_1 = 10L & V_2 = 20 L.

$$W = -nRT\ln\frac{v_2}{v_1} = -5 \times 8.314 \times 300 \times 2.303 \text{ by}\left(\frac{20}{10}\right)$$

$= -5 \times 8.314 \times 300 \times 2.303 \times 0.3010$

$= 8630$ J

Q.9 (195)

$$N_2O \text{ moles} = \frac{2.2}{4.4} = \frac{1}{20}$$

$\Delta H = nC_p\Delta T = 1/2 \times 100 (-40) = -200$ J

$\Delta H = q_p + w$

$w = -P_{ext}.DV$

$$W = -1\frac{(167.75 - 217.1)}{1000} \times 101.3J$$

$w = +5J$

$\Delta U = -200 + 5 = -195J$

Q.10 (117)

A/q → For 17 g (1 mole NH_3) enthalpy change for vapourisation = 23.4 kJ mol^{-1}

So for 85 (g) (5 mole NH_3)

Value will be = 23.4 × 5 = 117 kJ

Q.11 (600)

$$Mg(s) + \frac{1}{2}O_2(g) \rightarrow MgO(s)$$

$\Delta H = \Delta U + \Delta n_g RT$

$$-601.70 \times 10^3 = \Delta U - \frac{1}{2} \times 8.3 \times 300$$

$-601.70kJ = \Delta U - 1.245$ kJ

$\Delta U = -600.455$ kJ

Q.12 (3)

(3)

Given

$\Delta H_{comb}(C_2H_6, g) = -1560$ kJ / mole

$\Delta H_{comb}(C, S) = -394$ kJ / mole

$\Delta H_{comb}(H_2, g) = -286.0$ kJ / mole

(i) $C_2H_6(g) + \frac{7}{2}O_2(g) \rightarrow 2CO_2(g) +$

$3H2O(\ell)\ (\ell), \Delta H^o_{comb} = -1560$ kJ / mole

(ii) $C(s) + O_2(g) \rightarrow CO_2(g),\ \Delta H^o_{comb} = -394$ kJ / mole

(iii) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(g),\ \Delta H^o_{comb} = -286$ kJ / mole

$2C(s) + 3H_2(g) \rightarrow C_2H_6(g)\ \Delta H^o_{rxn} = \Delta H^o_f(C_2H_6, g)$

$\Delta H^o_f = \Delta H^o_C(\text{reactant}) - \Delta H^o_C(\text{Product})$

$= (-394) \times 2 + (-286) \times 3 - (-1560)$

$= -788 - 858 + 1560$

$= -1646 + 1560$

$= -86.0$ kJ / mol

Q.13 (747)

$$2O_3(g) \rightleftharpoons 3O_2(g)$$

t=0 (mole) 1 0

t_{eq}(mole) 1–0.5

$$\left(\frac{3}{2}\right)\times 0.5 = 0.75\,\text{Mole}$$

(α=0.5)
Total final moles = 1.25, P_{total}=1atm
Partial pressure =mole fraction ×total pressure

$$k_p = \frac{(Po_2)^3}{(Po_3)^2} \frac{\left(\frac{0.75}{1.25}\times 1\right)^3}{\left(\frac{0.5}{1.25}\times 1\right)^2} = 1.35$$

We know →
$\Delta G = -RT \ln Kp$
$\Delta G = -8.3\times 300 \ln 1.35$
$= -747$ J mol = 747

Q.14 (3)
State variable → Volume, Enthalpy, Internal Energy
Path Variable → Heat

Q.15 [2]
$C_{p,m} = C_{v,m} + R$
$\Rightarrow C_{v,m} = 20.785 - 8.314 = 12.471\ Jk^{-1}\ ml^{-1}$
$\Delta U = nC_{v,m}\Delta T$

$$\Rightarrow\ n = \frac{5000}{12.471\times 200} = \frac{25}{12.471} \approx 2$$

Q.16 [104]

$$C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O$$

$$\Delta H^o{}_c = 3\Delta H^o{}_f\ CO_2 + 4\Delta H^o{}_f\ H_2O - \Delta H_f\, C_3H_8$$

$$-2220 = 3(-393.5) + 4(-285.8) - x$$

$$-2220 = -1180.5 - 1143.2 - x$$

$x = -2323.7 + 2220$
$= 103.7$
$= 104$ kJ

Q.17 [2]
Ionisation energy of $CH_3COOH = 57.3 - 55.3 = 2$ kJ

Q.18 [200]
$\Delta T = 2K$
$\Delta H = -x$ kJ/mole

$$\text{Coal} \to 2.4\ \text{gm} \Rightarrow \frac{2.4}{12}$$

= 0.2 mole
ΔH for 0.2 mole
$= 0.2 \times (-x)$ kJ
$= -0.2\,x$ kJ
$q_{cal} = C\Delta T$
$0.2x = 20 \times 2 = 40$ kJ

$$X = \frac{40}{0.2} = 200 \qquad \because\ q_{cal} = -\Delta H$$

Q.19 [57]

$$H_2F_2(g) \longrightarrow H_2(g) + F_2(g)$$

$\Delta U = -59.6$ kJ/mol, $\Delta n_g = 2 - 1 = 1$
T = 27°C = 300 K
$\Delta H = \Delta U + \Delta ngRT$
$= -59.6 + 1 \times 8.314 \times 300 \times 10^{-3}$
$= -59.6 + (8.314 \times 0.3)$
$= -57.10$ kJ/mol
≈ -57KJ/mol

Q.20 [35]
At constant volume → ΔU

$$\Delta u\ (\text{kJ/Mol.}) = \frac{\text{Heat capacity} \times \text{change in temp.}}{\text{No. of moles of gas}}$$

$\Delta T = 298.5 - 298 = 0.45$ K
Heat capacity (C_V) = 2.5 kJ/K
$\Delta u = 9$ KJ

$$9kJ = \frac{2.5KJ/k \times 0.45k}{\text{No. of moles of gase}}$$

No. of moles of gas = 0.125 mol
Mass of gas = 280 x 0.125 = 33 gram
Ans: 35

Q.21 (1)
$\Delta H = \Delta U + P\Delta V$

Q.22 (54)
$HNO_3 + NaOH \to NaNO_3 + H_2O$
600 ml   400 ml
0.2M   0.1 M
mili moles of $HNO_3 = 600 \times 0.2 = 120$ m mol
mili moles of NaOH = $400 \times 0.1 = 40$ m mol
$HNO_3 + NaOH \to NaNO_3 + H_2O$
120 40
80 0   40
After the reaction
$\because$ 40 m moles = 40×10^{-3} moles are reacting
Enthalpy change for reaction
$\Delta_r H = 40 \times 10^{-3} \times 57 \times 10^3$ J
= 2280 J
C = Specific heat of water
M = mass of solution
= density × volume
= 1 gm/ml × 1000 ml
= 1000 gm
$q = mc\Delta T \quad \because\ q = \Delta_r H$
$2280 = 1000 \times 4.2 \times \Delta T$
$\Delta T = 54.286 \times 10^{-2}k = 54.286 \times 10^{-2}\,°C \approx 54 \times 10^{-2}\,°C$

Q.23 (4)

$$C_{(s)} + \frac{1}{2}O_2(g) \to CO_{(g)};\quad \Delta H = -100\text{KJ/mole}$$

$C(s) + O_2(g) \rightarrow CO_2(g);\quad \Delta H = -100 KJ/mole$

$\text{Mass of carbon} = (0.6 \times 10^3)\ \frac{60}{100} = \frac{600 \times 60}{100} = 360 \text{ gram}$

$60\% \text{ of carbon} \Rightarrow \frac{360 \times 60}{100} = 216 \text{ gram}$

(1) $C_{(s)} + \frac{1}{2} O_2 \rightarrow CO(g);\quad \Delta H = -100 KJ/mole$

$\left(\frac{216}{12}\right)\quad \Delta H = -100 \times \frac{216}{12} = -1800 KJ/$
mole

(2) $C_{(s)} + O_2 \rightarrow CO_2(g);\quad \Delta H = -400 KJ/mole$

$\left(\frac{144}{12}\right)\quad \Delta H = -400 \times \frac{144}{12} = -4800 KJ$

Total heat released = 1800 + 4800 = 6600 KJ