GUESS PAPER 2022

کامیابی کا تعویذ

پنجاب کے تمام بورڈ کے لیے

Chemistry

For intermediate Part-I

اب فیل ہونا بھول جائیں

☆ پیپر Setter کے ذہن کو مد نظر رکھ کر تیار کیے گئے سوالات

☆ یادرکھیں ! اب وقت انتہائی کم رہ گیا ہے۔

*صرف ایک ماہ کے اندر بورڈ امتحان کی مکمل تیاری کریں۔

تالیف محمد قدیر رفیق

القدیر جناح سائنس اکیڈمی

0302474124 ملیاں کلاں مرید کے روڈ شیخوپورہ

Objective Type

No.	QUESTIONS	A	B	C	D
1	**Isotopes differ in:**	Properties which depend upon mass	Arrangement of electrons in orbitals	Chemical properties	The extent to which they may be affected in electromagnetic field
2	**Select the most suitable answer from the given ones in each question:**	Isotopes with even atomic masses are comparatively abundant	Isotopes with odd atomic masses are comparatively abundant	Isotopes with even atomic masses and even atomic numbers are comparatively abundant	Isotopes with even atomic masses and odd atomic numbers are comparatively abundant
3	**Cadmium has isotopes:**	3	4	5	9
4	**Palladium has isotopes:**	6	7	8	9
5	**Bromine has isotopes:**	8	6	4	2
6	**Nickel has isotopes:**	3	5	6	11
7	**Hemoglobin is a macro molecule and consists of approximately atoms:**	5000	10000	68000	15000
8	**Tin has isotopes:**	9	10	11	12
9	**Number of isotopes of arsenic are:**	1	2	9	11
10	**The atomicity of $C_6H_{12}O_6$ is:**	6	12	3	24
11	**The mass of two moles of electron is:**	1.10 mg	1.008 mg	0.184 mg	1.673 mg
12	**The volume occupied by 16 g of CH_4 at STP.**	224.14 dm^3	22.4 dm^3	1.12 dm^3	2.24 dm^3
13	**One mole of SO_2 contains:**	6.02×10^{23} atoms of oxygen	18.1×10^{23} molecules of SO_2	6.02×10^{23} atoms of sulphur	4 gram atoms of SO_2
14	**The number of moles of CO_2 which contain 8.0 g of oxygen:**	0.25	0.5	1	1.5
15	**The mass of one mole of electrons is:**	1.008 mg	0.55 mg	0.184 mg	1.673 mg
16	**27 g of Al will react completely with how much mass of O_2 to produce Al_2O_3?**	8 g of oxygen	16 g of oxygen	32 g of oxygen	24 g of oxygen
17	**$1^{o}A$= .... m:**	44479	44510	44540	41548
18	**In Al_2O_4, the ratio between the ions:**	0.043056	0.084028	0.085417	0.126389
19	**A limiting reactant is the one which:**	Is taken in lesser quantity in grams as compared to other reactants	Is taken in lesser quantity in volume as compared to the other reactants	Gives the maximum amount of the product which is required	Gives the minimum amount of the product under consideration
20	**Which one is not example of sublimate?**	Ammonium chloride	Iodine	NaCl	Benzoic acid
21	**Compound which undergo sublimation is:**	$KMnO_4$	$CaCO_3$	NH_4Cl	Na_2CO_3
22	**Solvent extractions is an equilibrium process and it is controlled by:**	Law of mass action	The amount of solvent used	Distribution law	The amount of solute

23	**Solvent extraction method is a particularly useful technique for separation when the product to be separated is:**	Non-volatile of thermally unstable	Volatile or thermally stable	Non-volatile of thermally stable	Volatile or thermally unstable
24	**In .... technique a solute distribute between two immiscible liquids.**	Crystallization	Solvent extraction	Filtration	Distillation
25	**The comparative rates at which the solutes move in paper chromatography, depend on:**	The size of paper	R_f values of solutes	Temperature of the experiment	Size of the chromatographic tank used
26	**During chromatography strip should be dipped into solvent mixture to a depth of:**	3-4 mm	4-5 mm	5-6 mm	6-7 mm
27	**Chromatography in which the stationary phase is a solid is classified as:**	Partition chromatography	Gas chromatography	Adsorption chromatography	Thin layer chromatography
28	**A compound having small value of K (distribution coefficient) mostly remains in the:**	Stationary phase	Mobile phase	Chromatographic tank	Solvent
29	**During paper chromatography, the stationary phase is:**	Solid	Liquid	Gas	Plasma
30	**The chromatography in which stationary phase is liquid is called:**	Thin layer chromatography	Partition chromatography	Absorption chromatography	Gel chromatography
31	**Equal masses of methane and oxygen are mixed in an empty container at 25^{0}C. the fraction of total pressure exerted by oxygen is:**	$\frac{1}{3}$	$\frac{8}{9}$	$\frac{8}{9}$	$\frac{16}{17}$
32	**How should the conditions be changed to prevent the volume of a given gas from expanding when its mass is increased?**	Temperature is lowered and pressure is increased	Temperature is increased and pressure is lowered	Temperature and pressure both are lowered	Temperature and pressure both are increased
33	**If absolute temperature of a gas is doubled and the pressure is reduced to one half, the volume of the gas will:**	Remain unchanged	Increase four times	Reduce to ¼	Be doubled
34	**Formula used for the conversion of F^o into C^o is:**	°F=9/5 (°C)+32	°C=5/9[°F-32]	°F=5/9(°C)+32	°C=9/5[°F-32]
35	**The deviation of a gas from ideal behavior is maximum at:**	-100C and 5.0 atm	-100C and 2.0 atm	1000C and 2.0 atm	00C and 2.0 atm
36	**The gases show more deviation at:**	Low temperature and low pressure	High temperature and low pressure	High temperature and high pressure	Low temperature and high pressure
37	**Number of molecules in one dm3 of water is close to:**	$\frac{6.02}{22.4}\times 10^{23}$	$\frac{12.0}{22.4}\times 10^{23}$	$\frac{18}{22.4}\times 10^{23}$	$\frac{55.6}{6.02}\times 10^{23}$
38	**Which of the following will have the same number of molecules at STP?**	280 cm^3 of CO_2 and 280 cm^3 of N_2O	11.2 dm^3 of O_2 and 32 g of O_2	44 g of CO_2 and 11.12 dm^3 of CO	28 g of N_2 and 5.6 dm^3 of oxygen
39	**Mass of 22.4 dm^3 of N_2 at STP is:**	28gm	14gm	1.4gm	2.8gm
40	**Partial pressure of oxygen in the air is:**	156torr	157torr	158torr	159torr

41	**Vapour pressure of liquid depends upon:**	Amount of liquid	Surface area	Temperature	Size of container
42	**The partial pressure of oxygen in lungs is:**	760torr	320torr	159torr	116torr
43	**Ionic solids are characterized by:**	Low melting points	Good conductivity in solid state	High vapour pressures	Solubility in polar solvents
44	**The number of water molecules in $CuSO_4.5H_2O$ attacked with Cu^+ ion:**	One	Two	Three	Four
45	**The strongest acid among Halogen acids is:**	HCl	HBr	HI	HF
46	**Dipole-induced dipole forces are also called:**	Dipole-dipole forcesDipole-dipole forces	Ion-dipole forces	Debye forces	London dispersion forces
47	**Which of the given has Hydrogen bonding:**	CH_4	CCl_4	NH_3	NaCl
48	**When water freezes, its volume increases:**	0.1	0.09	0.15	0.18
49	**The molecules of CO_2 in dry ice form the:**	Ionic crystals	Covalent crystals	Molecular crystals	Any type of crystal
50	$If\ a \neq b \neq c$ **and a=b=90° $B \neq 90^o$ then crystal system is:**	Monoclinic	Diclinic	Triclinic	Polyclinic
51	**The distillation of liquid under reduced pressure is called:**	Destructive distillation	Vacuum distillation	Fractional distillation	Simple distillation
52	**The axis (unit cell length) for Cu is:**	$a \neq b = c$	$a \neq b \neq c$	$a = b \neq c$	$a = b = c$
53	**Which impurity makes the shape of sodium chloride crystal needle like:**	$MgSO_4$	Urea	Glucose	$MgCO_3$
54	**Allotropy is the property of:**	Compound	Element	Atom	Mixture
55	**Transition temperature of KNO_3 is:**	13.2°C	95.5°C	128oC	32.2°C
56	**Amorphous solids:**	Have sharp melting points	Undergo clean cleavage when cut with knife	Have perfect arrangement of atoms	Can possesses small regions of arrangement of atoms
57	**Which of the following is a pseudo solid?**	CaF_2	Glass	NaCl	All
58	**Which of following will have H-bonding in its molecules?**	C_2H_5OH	CCl_4	I_2	NaCl
59	**The nature of the positive rays depend on:**	The nature of electrode	The nature of the discharge tube	The nature of the residual gas	All of the above
60	**The velocity of photon is:**	Independent of its wavelength	Depends on its wavelength	Equal to square of its amplitude	Depend on its source
61	**When fast neutron carries nuclear reaction with nitrogen is ejects particles:**	α	β	γ	δ
62	**Cathode rays strike alumina and produce a .... color.**	Red	Blue	Yellow	Green
63	**The e/m value for the positive rays in maximum for the gas:**	Hydrogen	Helium	Oxygen	Nitrogen

64	**Positive rays were discovered by:**	J.J Thomson	Goldstein	William Crookes	Ruther Ford
65	**The charge on proton is:**	1.6022×10^{-11}C	1.6022×10^{11}C	1.6022×10^{-19}C	1.6022×10^{19}C
66	**Name the electron is given by:**	William Crooks	Stoney	J.J Thomson	Chadwick
67	**The limiting line of Balmer series lies in:**	Visible region	U.V region	I.R region	X-rays region
68	**The element which has maximum number of unpaired electron is:**	Cr_{24}	Ca_{20}	F_{26}	CH_{29}
69	**Rutherford's model of atom failed because:**	The atom did not have a nucleus and electrons	It did not account for the attraction between protons and neutrons	It did not account for the stability of the atom	There is actually no space between the nucleus and the electrons
70	**When one beta (β) particle is emitted from the nucleus of an atom its:**	Atomic number increases by 1	Atomic number decreases by 1	Atomic mass increases by 1	Atomic mass decreases by 1
71	**The wave number of the light emitted by a certain source is 2×10^6 m^{-1}. The wavelength of this light will be:**	500nm	500m	200 nm	5x107m
72	**Bohr model of atom is contradicted by:**	Planck's quantum theory	Dual nature of matter	Heisenberg's uncertainty principle	All of the above
73	**Bohr's model atom is contradicted by:**	Planck quantum theory	Quantization of energy of electrons	Heisenberg's uncertainly principle	Quantization of angular momentum
74	**Splitting of spectral lines when atoms are subjected to strong electric field is called:**	Zeeman effect	Stark effect	Photoelectric effect	Compton effect
75	**Lyman series lies in special region:**	Infrared	Ultra violet	Visible	None of these
76	**When atoms are subjected to strong electric field, spelling of spectral lines is called:**	Zeeman effect	Stark effect	Photoelectric effect	Compton effect
77	**Number of bonds in N_2 molecules are:**	$one\ \sigma\ and\ 2\pi$	$one\ \sigma\ and\ 1\pi$	$three\ \sigma$	$two\ \sigma\ and\ 1\pi$
78	**De-Brogile equation is represented by:**	$h = \frac{\lambda}{mv}$	$m = \frac{h}{\lambda v}$	$m = \frac{\lambda}{hv}$	$\lambda = \frac{h}{mv}$
79	**In the ground state of an atom, the electron is present:**	In the nucleus	In the second shell	Nearest to the nucleus	Farthest from the nucleus
80	**Quantum number values for 2p orbitals are:**	n=2, l=1	n=1, l=2	n=1, l=0	n=2, l=0
81	**Orbitals having same energy are called:**	Hybrid orbitals	Valence orbitals	Degenerate orbitals	d-orbitals
82	**When 6d orbitals is complete, the entering electrons goes into:**	7f	7s	7p	7d
83	**Maximum number of electrons in f-subshells is:**	2	6	10	14
84	**The electron in subshell is filled according to formula:**	$2n^2$	2(2l+1)	(2n+1)	None of these

85	**When 5d orbital is completed then entering electron goes into:**	6s	6p	6d	6f
87	**An orbital which is spherical and symmetrical is:**	s-orbital	p-orbital	d-orbital	f-orbital
88	**What is the value of (n+f) for the 3rd sub-shell?**	3	4	5	6
89	**In the ground state of an atom the electron is present:**	In the nucleus	In the second shell	Nearest to the nucleus	Farthest from the nucleus
90	**Which of the following molecule obey octet rule:**	BF_3	BCl_3	NH_3	SF_6
91	**Which compound does not obey the octet rule?**	NH_3	BCl_3	H_2O	CH_4
92	**Forces of attraction between He atoms are:**	Hydrogen bonding	Debye forces	London forces	Ion dipole forces
93	**Octet rule is not followed in:**	CH_4	CF_4	CCl_3	PCl_5
94	**An ionic compound A+B- is mostly likely to be formed when:**	The ionization energy of A is high and electron affinity of B is low	The ionization energy of A is high and electron affinity of B is high	Both the ionization energy of A and electron affinity of	Both the ionization energy of A and electron affinity of
95	**Which element has highest ionization potential:**	Li	Be	B	C
96	**Percentage ionic character of HF is:**	1	0.8	0.43	0.57
97	**Which one has highest value of ionization energy:**	Be	C	O	F
98	**......... element has highest value of electron affinity:**	Fluorine	Chlorine	Bromine	Iodine
99	**The amount of energy released by absorbing an electron in the valence shell of an atom is:**	Ionization energy	Electron affinity	Electro negativity	Bond energy
100	**In nitrogen molecule (N_2) each nitrogen atom contributes in sharing for formation of bond:**	One electron	Two electrons	Three electrons	Four electrons
101	**The tendency of an atom to attract shared paired of electron towards itself is called its:**	Ionization energy	Electron affinity	Electronegativity	Dipole moment
102	**The number of bonds in nitrogen molecule is:**	$one\ \sigma\ and\ 2\pi$	$one\ \sigma\ and\ 1\pi$	$three\ \sigma$	$two\ \sigma\ and\ 1\pi$
103	**Which of the following statements is not correct regarding bonding molecular orbitals?**	Bonding molecular orbitals possess less energy than atomic orbitals from which	Bonding molecular orbitals have low electron density between the two nuclei	Every electron in the bonding molecular orbitals contributes to the attraction between atoms	Bonding molecular orbitals are formed when the electron waves undergo

		they are formed			constructive interference
104	**Which of the following species has unpaired electrons in antibonding molecular orbitals:**	O_2^{2+}	O_2^{2-}	B_2	F_2
105	**The shape of SnCl² molecule is:**	Linear	Angular	Trigonal planar	Tetrahedral
106	**The structure of water molecule is:**	Angular	Linear	Trigonal	Trigonal pyramidal
107	**Which of the following has linear structure:**	CO_2	NH_3	CH_4	H_2O
108	**In methanol, bond between carbon and oxygen:**	Ionic	Non-ionic	Polar	Co-ordinate
109	**Which of the following has coordinate covalent bond:**	NH_4Cl	NaCl	HCl	$AlCl_3$
110	**The molecular shape of SO_3 is:**	Triangular planar	Tetrahedral	Pyramidal	Linear
111	**The bond angle in NH_3 molecule is:**	109.5o	107.5 o	104.5 o	108 o
112	**The carbon atom in C_2H_4 is:**	sp_3-hybridized	sp_2-hybridized	sp_3-hybridized	Dsp_2-hybridized
113	**Carbon atom in methane in hybridized:**	sp_3	sp_2	sp	dsp_2
114	**The bond order of is O_2^{2-}**	1	2	Zero	3
115	**Molecule in which the distance between two carbon atoms in the largest is:**	C_2H_6	C_2H_4	C_2H_2	C_6H_6
116	**The bond order of N_2 molecule is:**	1	2	3	4
117	**Which of the following has bond angle of 120°:**	$BeCl_2$	BF_3	CH_4	NH_3
118	**The paramagnetic behavior of oxygen is well explained on the basis of:**	M.O theory	VESPR theory	N.B theory	CF theory
119	**The number of bonds in oxygen molecule:**	$one\ \alpha\ and\ one\ \pi$	$one\ \alpha\ and\ two\ \pi$	$three\ \alpha$	$two\ \alpha\ and\ two\ \pi$
120	**The H-H bond energy in kJ $mole^{-1}$ is:**	346	436	463	336
121	**Bond energy of hydrogen (H_2) molecule is:**	410 KJ mol^{-1}	450 KJ mol^{-1}	436 KJ mol^{-1}	415 KJ mol^{-1}
122	**In ethyne molecule the number and nature of bonds are:**	One sigma and two pi	Two sigma and one pi	Three sigma two pi	Two sigma two pi
123	**The spreading of fragrance of scent is due to:**	Osmosis	Density	Effusion	Diffusion
124	**......... is not paramagnetic:**	O_2^{2-}	O_2	N_2^{2-}	None of these
125	**The bond order of N_2 molecule is:**	Zero	1	2	3
126	**Geometry of SO_2 molecule is:**	Linear	Angular	Tetrahedral	Trigonal pyramidal
127	**Which of the following species has un-paired electrons in the anti-bonding molecular orbitals:**	O_2^{2-}	N_2^{2-}	B_2	F_2
128	**......... is not state function.**	Pressure	Volume	Temperature	Heat

129	**Which of the following statements is contrary to the first law of thermodynamics?**	Energy can neither be created nor destroyed	One form of energy can be transferred into an equivalent amount of other kinds of energy	In an adiabatic process, the work done is independent of its path	Continuous production of mechanical work without supplying an equivalent amount of heat is possible
130	**Calorie is equivalent to:**	0.4184 J	41.84 J	4.184 J	418.4 J
131	**The change in heat energy of a chemical reaction at constant temperature and pressure is called:**	Enthalpy change	Heat of sublimation	Bond energy	Internal energy change
132	**For a given process, the heat changes at constant pressure (qp) and at constant volume (qv) are related to each other as:**	$q_p=q_v$	$q_p<q_v$	$q_p>q_v$	$q_p=q_v/2$
133	**For the reaction: the change in enthalpy is called:**	Heat of reaction	Heat of formation	Heat of neutralization	Heat of combustion
134	**Enthalpy of neutralization of all the strong acids and strong bases has the same value because:**	Neutralization leads to the formation of salt and water	Strong acids and bases are ionic substances	Acids always give rise to H+ ions and bases always furnish OH- ions	The net chemical change involve the combination of H+ and OH- ions to form water
135	**The pressure of oxygen inside the bomb calorimeter is:**	100atm	50atm	25atm	20atm
136	**Enthalpies of all elements in their standard states are:**	Unity	Zero	Always positive	Always negativity
137	**The total heat content of system is called:**	Entropy	Enthalpy	Temperature	Internal energy
138	**Heat change for one mole of a solid during converting it into liquid is called:**	Molar heat of fusion	Molar heat of vaporization	Molar heat of sublimation	Enthalpy change
139	**The net heat change in a chemical reaction is same, whether it is brought about in two or more different ways in one or several steps. It is known as:**	Henry's Law	Joule's principal	Hess's Law	Law of conservation of energy
140	**The optimum temperature for the synthesis of NH3 by Haber's process is:**	200oC	300oC	400oC	500oC
141	**The born-Haber cycle is the best application of:**	Boyle's Law	Dalton's Law	Hess's Law	Graham's Law
144	**......... was derived by C.M. guldberg and P. Waage is 1864.**	Law of conservation of mass	Law of mass action	Distribution law	Law of conservation of energy
145	**The unit of Qc for reaction $N_2 + O_2 \leftrightarrow 2NO$ will be:**	mole/dm^3	Mole^{-1}/dm^3	Mole^{-2}/dm^{-6}	No unit
146	**Optimum pressure in Haber's process for synthesis of ammonia is:**	100-150atm	200-300atm	350-450atm	500-600atm
147	**The law of mass action was given by:**	D.C down and P.Waage	Gay-Lussic and C.M	C.M Guldberg and P.Waage	Hendeson and Le-Chattelier's

148	**Optimum temperature for synthesis of ammonia by Haber process is:**	370°C	390°C	400°C	410°C
149	**When Kc value of a reaction is very small, the equilibrium position lies to:**	Left	Right	May be left or right	Cannot be predicted
150	**Catalyst used in preparation of SO_3 gas from SO_2 is:**	Al2O3	CaO	V2O5	SiO2
151	**Catalyst used in preparation of NH3 from N_2 and H_2 is:**	Fe	Ni	Pt	V2O5
152	**The units of Kc for the reaction of ammonia synthesis are:**	moles^{-2}dm6	moles^{-1}dm^{6}	moles^{-2}dm^{3}	moles^{-2}dm^{2}
153	**The pH of 10^{-4} moles / dm^3 of $Ba(OH)_2$ is:**	4.5	6.4	7.5	10.3
154	**Sum of pK_a and pK_b is equal to:**	7	9	11	14
155	**The pH of human blood is:**	7.12	7.35	7.56	8
156	**The sum of pH and pOH at 25°C always equal to:**	7	Zero	14	41913
157	**The units for kw of H2O are:**	mol/dm^{3}	mol^{2}dm-^{6}	mol^{-2}dm^{6}	mol^{-2}dm^{-3}
158	**The term pH was introduced by:**	llenderson	Millikan	Le-chattlier	Sorenson
159	**The values of Kw of water at 25°C is:**	0.11x10^{-14}	0.30x10^{-14}	1.0x10^{-14}	7.5x10^{-14}
160	**The pH of milk of Magnesia is:**	10.5	3.5	8.5	11.1
161	**The pH of 10-3mol dm-3 of an aqueous solution of H2SO4 is:**	3	2.7	2	1.5
162	**The solubility of AgCl s 2.0x10^{-10} mol^2 dm^{-6}. The maximum concentration of Ag+ ions in the solution is:**	2.0x10^{-10}mol dm^{-3}	1.4x10^{-5}mol dm^{-3}	1.0x10^{-10}mol dm^{-3}	4.0x10^{-20}mol dm^{-3}
163	**An excess of aqueous silver nitrate is added to aqueous barium chloride and precipitate is removed by filtration. What are the main ions in the filtrate?**	Ag^+ and NO^{3-} only	Ag^+ and Ba^{2+} and NO^{3-}	Ba^{2+} and NO^{3-} only	Ba^{2+} and NO_3 and Cl^-
164	**pOH of water is:**	2	4	6	7
165	**The solubility product of AgCl is 2.0x10^{-10} mol^2.dm^{-6}. The maximum concentration of Ag: ions in the solution is:**	2.0x10^{-10} mol dm^{-3}	1.41x10^{-5} mol dm^{-3}	1.0x10^{-10} mol dm^{-3}	1.0x10^{-10} mol dm^{-3}
166	**By adding NH_4Cl to NH_4OH solution. The ionization of NH_4OH.**	Increases	Remains same	Decreases	Increases 100 times
167	**In the presence of common ion, the ionization of an electrolyte will.**	Increase	Decrease	No affect	Moderate change
168	**The nature of milk is:**	Acidic	Pasic	Neutral	Normal
169	**The ionization constant of pure water (Kw) at 25°C is:**	1.8x10^{-16} mole dm^{-3}	1.6x10^{-14} mole dm^{-3}	1.0x10^{-14} mole^2 dm^{-6}	1.8x10^{-14} mole^2 dm^{-6}
170	**Acid having Ka> 1 will be:**	Weak	Very weak	Moderate	Strong
171	**pH of soft drinks at 25°C is about:**	3	11	1	7
172	**pH of pure water is:**	4.4	5.4	7	8
173	**The pH of the gastric juice is:**	2	3	3.5	5.6
174	**pH of banana is:**	2.1	4.6	9.4	9.6
175	**Approximately pH of apple is:**	2.7	3.1	4.2	4.5

176	**An azeotropic mixture of two liquids boils at a lower temperature than either of them when:**	It is saturated	It shows positive deviation from Raoult's Law	It show negative deviation from Raoult's Law	It is metastable
177	**In azeotropic mixture showing positive deviation from Raoult's Law, the volume of the mixture is:**	Slightly more than the total volume of the components	Slightly less than the total volume of the components	Equal to the total volume of the components	None of these
178	**Azeotropic mixture of two liquids boils at a lower temperature than either of them, when:**	It is saturated	It shows positive deviation from Raoult's Law	It shows negative deviation from Raoult's Law	It is metastable
179	**Upper consulate temperature for water phenol system is:**	150°C	65.9°C	120°C	130°C
180	**Ideal solutions obey:**	Henry's Law	Avogadro's Law	Raoult's Law	Smith's Law
181	**Which one of the following is an ideal soltution:**	C_2H_5-OH and H_2O	C_6H_6-OH and CCl_4	$CHCl_3$ and $(CH_3)_2CO$	None of these
182	**Which of the following solutions has the highest boiling point?**	5.85% solution of sodium chloride	18.0% solution of glucose	6.0 solution of urea	All have the same boiling point
183	**Which of the following solutions has the highest boiling point?**	5.85% solution of sodium chloride	18.0% solution of glucose	6.0 solution of urea	All have the same boiling point
184	**The molal boiling point constant is the ratio of the elevation in boiling point to:**	Molarity	Molality	Mole fraction of solvent	Mole fraction of solute
185	**Colligative properties are the properties of:**	Dilute solutions which behave as nearly ideal solution	Concentrated solutions which behave as nearly non-ideal solutions	Botha a and b	Neither a nor b
186	**Melting of ice can be lowered by the use of:**	LiCl	$BeCl_2$	NaCl	AgCl
187	**One molar solutions of glucose ($C_6H_{12}O_6$) contains the amount of solute in 500cm^3 solution.**	180g	90g	45g	270g
188	**When ionic product of a solution is greater than the solubility product at a particular temperature then the solution is said to be:**	Unsaturated	Saturated	Very dilute	Super saturated
189	**The cathodic reaction in the electrolysis of dil. H_2SO_4 with Pt electrodes is:**	Reduction	Oxidation	Both oxidation and reduction	Neither oxidation or reduction
190	**Which of the following statements is not correct about galvanic cell?**	Anode is negatively charged	Reduction occurs at anode	Cathode is positively charged	Reduction occurs at cathode
191	**If the salt bridge is not used between two half cells,, then the voltage.**	Decrease rapidly	Decrease slowly	Does not change	Drops to zero

192	**If a strip of Cu metal is placed in a solution of FeSO4:**	Cu will be deposited	Fe is precipitated out	Cu and Fe both dissolve	No reaction take place
193	**Stronger the oxidizing agent, greater is the:**	Oxidation potential	Reduction potential	Redox potential	E.M.F of cell
194	**The reduction potential of Zn is:**	+0.76V	-0.34V	+0.34V	-0.75V
195	**The cathodic reaction in the electrolysis of dil H_2So_4 with Pt electrodes is:**	Reduction	Oxidation	Both a and b	Neither oxidation or reduction
196	**Standard hydrogen electrode (SHE) is made of:**	Ag foil	Au foil	Cu foil	Pt foil
197	**Stronger is the oxidizing agent, greater is the:**	Oxidation potential	Reduction potential	Redox potential	E.M.F of the cell
198	**In zero order reaction, the rate is independent of:**	Temperature of reaction	Concentration of reactants	Concentration of products	None of these
199	**If the rate equation of a reaction 2A+B→products is, rate = $k[A]^2$[B] and A is present in large excess, then order of reaction is:**	1	2	3	None of these
200	**The rate of reaction:**	Increase as the reaction proceeds	Decreases as the reaction proceeds	Remains the same as the reaction proceeds	May decrease or increase as the reaction proceeds
201	**With increase of 10oC temperature the rate of reaction doubles. This increase in rate of reaction is due to:**	Decrease in activation energy of reaction	Decrease in the number of collisions between reactant molecules	Increase in activation energy of reactants	Increase in number of effective collisions
202	**The unit of the rate constant is the same as that of the rate of reaction is:**	First order reaction	Second order reaction	Zero order reaction	Third order reaction
203	**The order of decomposition of nitrogen pentoxide $2N_2O_5 \rightarrow 2N_2O_4$**	First order	Second order	Third order	Zero order
204	**The rate of reaction determined at any given time is called:**	Average rate	Instantaneous rate	Spontaneous rate	Overall rate
205	**Unit or rate constant is the same as the rate of reaction is:**	Zero order reaction	1st order reaction	2nd order reaction	3rd order reaction
206	**All radioactive disintegration nuclear reactions are of:**	First order	Second order	Third order	Zero order
207	**Photochemical reaction are usually::**	Zero order	First order	Second order	Third order
208	**Hydrolysis of tertiary butyl bromide is:**	Zero order reaction	First order reaction	Pseudo first order reaction	Second order reaction
209	**The order reactions for the reaction $2N_2O_5 \rightarrow 2N_2O_4 + O_2$ is.**	Zero order	First order	Second order	Third order

210	The order of reaction for the reaction $NO + O_3 \rightarrow NO_2 + O_2$	Two	Three	One	Zero
211	Hydrolysis of tertiary butyl bromide has order of reaction:	First	Pseudo first	Second	Third
212	Half-life period for U_{92}^{235}	710 million years	720 million years	810 million years	820 million years
213	Half-life of a second order reaction is inversely proportional to:	Initial concentration of reactants	Final concentration of reactants	Initial concentration of products	Final concentration of products

Answers Multiple Questions

1	(A)	2	(C)	3	(D)	4	(A)	5	(D)	6	(B)	7	(B)	8	(C)	9	(B)	10	(D)
11	(A)	12	(B)	13	(C)	14	(A)	15	(B)	16	(D)	17	(A)	18	(C)	19	(D)	20	(C)
21	(C)	22	(C)	23	(D)	24	(B)	25	(B)	26	(C)	27	(C)	28	(A)	29	(B)	30	(B)
31	(A)	32	(A)	33	(A)	34	(B)	35	(A)	36	(D)	37	(A)	38	(A)	39	(A)	40	(D)
41	(C)	42	(D)	43	(A)	44	(D)	45	(C)	46	(C)	47	(C)	48	(B)	49	(A)	50	(A)
51	(B)	52	(D)	53	(B)	54	(B)	55	(C)	56	(A)	57	(A)	58	(A)	59	(A)	60	(A)
61	(A)	62	(A)	63	(A)	64	(B)	65	(C)	66	(B)	67	(A)	68	(A)	69	(A)	70	(A)
71	(A)	72	(A)	73	(C)	74	(A)	75	(B)	76	(B)	77	(A)	78	(D)	79	(A)	80	(A)
81	(A)	82	(A)	83	(D)	84	(B)	85	(B)	86	(C)	87	(A)	88	(D)	89	(C)	90	(A)
91	(B)	92	(C)	93	(D)	94	(A)	95	(D)	96	(C)	97	(D)	98	(A)	99	(B)	100	(C)
101	(C)	102	(A)	103	(A)	104	(A)	105	(C)	106	(A)	107	(A)	108	(C)	109	(A)	110	(A)
111	(B)	112	(B)	113	(A)	114	(A)	115	(A)	116	(C)	117	(B)	118	(A)	119	(A)	120	(B)
121	(C)	122	(A)	123	(D)	124	(B)	125	(D)	126	(B)	127	(B)	128	(D)	129	(A)	130	(A)
131	(A)	132	(A)	133	(A)	134	(A)	135	(D)	136	(B)	137	(B)	138	(A)	139	(A)	140	(C)
141	(C)	142	(A)	143	(A)	144	(B)	145	(D)	146	(B)	147	(C)	148	(C)	149	(A)	150	(C)
151	(A)	152	(A)	153	(D)	154	(D)	155	(B)	156	(C)	157	(A)	158	(D)	159	(C)	160	(A)
161	(A)	162	(A)	163	(A)	164	(D)	165	(B)	166	(C)	167	(B)	168	(B)	169	(C)	170	(D)
171	(A)	172	(C)	173	(A)	174	(B)	175	(B)	176	(A)	177	(A)	178	(B)	179	(B)	180	(C)
181	(D)	182	(A)	183	(A)	184	(A)	185	(A)	186	(C)	187	(B)	188	(B)	189	(A)	190	(A)
191	(A)	192	(A)	193	(A)	194	(A)	195	(A)	196	(D)	197	(A)	198	(A)	199	(A)	200	(A)
201	(A)	202	(A)	203	(A)	204	(B)	205	(A)	206	(A)	207	(A)	208	(B)	209	(B)	210	(C)
211	(A)	212	(A)	213	(A)														

Short Questions NO.2

1.	What are isotopes? Why they have same chemical but different physical properties?
2.	Explain mathematical relationship of m/e of an ion in mass spectrometry.
3.	Magnesium atom is twice heavier than carbon atom. Comment.
4.	How one mg of K_2CrO_4 has thrice the number of ions than the number of formula units when ionized.
5.	How 4.9 g of H_2SO_4 when completely ionized in water have equal number of +ve and –ve charges but the number of positively charged ions are twice the number of negativity charged ions.
6.	23 g of sodium and 39 g of potassium have equal number of atoms in them. Justify.
7.	Law of conservation of mass must be considered during stoichiometric calculations. How?
8.	Calculate the number of water molecules in 10 g of ice.
9.	Give assumptions of stoichiometry.
10.	One mole of H_2SO_4 should completely react with two moles of NaOH. How does Avogadro's number help to explain it?

11.	Why do 2 g of H_2, 16g of CH_4, 44g of CO_2 occupy separately the volume of 22.414 dm^3 although the sizes and masses of molecules of three gases are very different from each other?
12.	Define limiting reactant. Give an example.
13.	How do many chemical reactions taking place in our surrounding involve limiting reactants?
14.	Define actual yield. Write formula for the calculation of % age yield.
15.	Why theoretical yield is greater than actual yield?
16.	Why we calculate %age yield?
17.	What is Avogadro's number? Give equation to relate the Avogadro's number and mass of element.
18.	How N_2 and CO have same number of electrons, protons and neutrons.
19.	Explain Boyle's law with the help of KMT.
20.	Justify that volume of gas becomes theoretically zero at -273oC.
21.	Why lighter gases diffuse more rapidly than heavier gases?
22.	Calculate the density of methane at STP.
23.	Define Avogadro's Law.
24.	State Joule-Thomson Effect. Write its application.
25.	Hydrogen and Helium are ideal at room temperature but SO_2 and Cl_2 are non-ideal.
26.	Some of the postulates of Kinetic Molecular Theory are faulty. Justify
27.	Calculate the value of R in units' $atm.dm^3.k^{-1}mol^{-1}$.
28.	Give four fundamental postulates of KMT of gases.
29.	Derive Graham's law of diffusion in the light of KMT of gases. Rate of diffusion of ammonia is more than that of HCl. Why?
30.	Where do natural and artificial plasma exist?
31.	Write two characteristics of plasma.
32.	Define Plasma. Give its one application.
33.	Give two applications of Plasma.
34.	Explain the term reversible reaction and state of equilibrium.
35.	State Le-Chatelier's principle. And discuss the effect of change in concentration of a product on reversible reaction.
36.	How do the buffers act? Give example.
37.	How the buffer solutions are prepared?
38.	What do you mean by Buffer capacity?
39.	Write two applications of equilibrium constant?
40.	Write two uses of buffer solutions.
41.	Give two applications of solubility product.
42.	Define solubility product. Derive solubility product expression for Ag_2CrO_4?
43.	How does a catalyst affect a reversible reaction?
44.	Write the relationship of K_p with K_c.
45.	Define Chemical equilibrium. Give its any two properties
46.	Define Raoult's Law. Give one of its mathematical forms.
47.	Non-ideal solutions do not obey the Raoult's Law.
48.	Differentiate between ideal and non-ideal solution.
49.	What is solubility principle?
50.	What is solubility carve? Name its two types.
51.	Make difference between continuous and discontinuous solubility curves.
52.	Example what is meant by water of crystallization? Give two examples.
53.	Define cryoscopy constant with an example.
54.	Give two applications of colligative properties.
55.	What is discontinuous solubility curve? Give one example.
56.	Cane sugar cannot be dissolved in benzene. Give reason.
57.	Define Boiling Point Elevation.
58.	Define ebullioscopic constant with an example.
59.	What is meant by conjugate solutions?
60.	Define colligative properties, name important colligative properties.

Short Questions NO.3

1.	Define sublimate. Give two examples.

2.	Define sublimation and Chromatography.
3.	Define solvent extraction and partition law.
4.	Define chromatography. Give formula of distribution coefficient.
5.	Differentiate between stationary and mobile phase used in chromatography.
6.	What is Rf value? Why it has no units?
7.	State the distribution law with two examples.
8.	What is difference between adsorption and partition chromatography?
9.	Write down the uses of chromatography.
10.	Define distribution law. How it is helpful in solvent extraction?
11.	What is the role of Hydrogen bonding in biological compounds?
12.	What are dipole-dipole forces? How they affect thermodynamic properties of substances.
13.	Differentiate between intermolecular and intermolecular forces.
14.	Lower alcohols are soluble in water but hydrocarbons are insoluble. Give reason.
15.	What are dipole-dipole forces of attraction? Explain with an example.
16.	Why Ice floats over the surface of water?
17.	What are Debye forces?
18.	HCl is stronger acid than HF. Why?
19.	Why HF shows exceptionally low acidic strength as compared to HCl, HBr and HI?
20.	Ice has less density than liquid water. Why?
21.	How the liquid crystals help in the detection of the blockage in Veins and arteries?
22.	What are liquid crystals? Why are they so called?
23.	Give two uses of liquid crystal.
24.	Differentiate between isomorphism and polymorphism.
25.	Why graphite is a good conductor of electricity?
26.	What do you mean by cleavage and cleavage planes?
27.	Define allotropy with an example.
28.	Define Polymorphism and Anisotropy. Give one example of each.
29.	What are Pseudo-Solids (Amorphous-Solid).
30.	What is isomorphism? Give an example.
31.	The vapor pressure of diethyl ether is higher than that of water at the same temperature. Give reason.
32.	One of the unit cell angles of hexagonal crystal is 120^o. justify it.
33.	Define Unit Cell. Give an example.
34.	Write defects of Rutherford atomic model.
35.	Write down two postulates of plank's quantum theory of radiation.
36.	Differentiate between frequency and wave number.
37.	Differentiate between continuous and line spectrum.
38.	What is atomic emission spectrum?
39.	Describe Zeeman's and Stark's effect.
40.	State Heisenberg uncertainty Principle and give its mathematical form.
41.	Differentiate between orbit and orbital.
42.	Why it is necessary to decrease the pressure in the discharge tube?
43.	Differentiate between wavelength and frequency?
44.	Calculate mass of an electron from its e/m value.
45.	State Pauli's exclusion principle and Hund's rule.
46.	Why e/m of cathode rays is just equal to that of electron?

47.	Differentiate between chemical kinetics and chemical equilibrium.
48.	Define with example 2nd-order Reaction.
49.	The radioactive decay is always a first order reaction. Justify.
50.	Rate of reaction decreases with the passage of time explains.
51.	Differentiate between Average and instantaneous rate of reaction.
52.	What is effect of light on the rate of reaction?
53.	What is Pseudo First Order Reaction? Give an example.
54.	What is meant by half-life period? Give one example.
55.	Define first order reaction with example.
56.	What is Zero-order reaction? Give one example.
57.	How the rates of reaction depend upon the nature of reactants?
58.	Radioactive decay is always a first order reaction. Justify it.
59.	How does nature of reactants affect rate of reaction, give an example.
60.	What is specific rate constant or velocity constant.
61.	How can half-life be used to determine order of reaction?
62.	Define order of reaction and velocity constant.
63.	What is difference between rate constant and specific rate constant.
64.	Define half-life period of a reaction. Give one example.
65.	What is the effect of temperature on energy of activation of a reaction?

Short Questions NO.4

1.	Define octet rule. Give two examples of compounds that do not obey this rule.
2.	Why the radius of an atom cannot be determined precisely?
3.	Why cationic radii are smaller than Anionic radii?
4.	Why anionic radius is greater than parent atom?
5.	75.4pm is compromise distance between the bonded hydrogen atoms. Justify.
6.	Why ionization energy (IE) values are decreased from top to bottom in a group?
7.	Why second Electron Affinity of oxygen atom is positive but first electron affinity is negative?
8.	Define electronegativity and electron affinity of an atom.
9.	Ionization energy is index to the metallic character. Why?
10.	Why second ionization energy (I.E) of an element is always greater than first ionization energy (I.E).
11.	Helium is diamagnetic in nature Justify.
12.	Why liquids are less common then solids & gases?
13.	Bond angle in CH_4 is 109.5^o but in H_2O is 104.5^o although carbon and oxygen are sp_3 hybridized. Give reason.
14.	Define lone pair and bond pair of electron.
15.	Why CO is polar and is non-polar?
16.	Differentiate between polar and non-polar covalent bonds with examples.
17.	Write down two postulates of VSEPR theory.
18.	There is no bond with 100% ionic character. Give reason.
19.	How Sigma and pi bonds are formed?
20.	The size of chlorine atom is smaller than ion. Justify it.
21.	Differentiate between sigma and Pi bond.
22.	What is meant by Bond order? Calculate bond order for -molecule.

23.	Why sigma bond is stronger than Pi bond?
24.	How bond length is effected by change of hybridization state?
25.	Define system and surrounding. Show by diagram of any one example.
26.	What is state function? Explain with example.
27.	Define standard Enthalpy of combustion and standard enthalpy of solution
28.	Define Born-Haber cycle and Lattice energy.
29.	Differentiate between atomization energy and lattice energy.
30.	What are endothermic and exothermic reactions? Give examples.
31.	What is internal energy of a system?
32.	Differentiate between Law of conservation of energy and Hess's Law.
33.	Explain the term enthalpy of atomization.
34.	Define Enthalpy of solution and enthalpy of neutralization.
35.	Define standard enthalpy of formation and give two examples.
36.	What is state function? Give two examples.
37.	State the Hess's Law of constant heat summation.
38.	Is it true that ΔH and ΔE have the same values for the reaction taking place in solution state?
39.	Draw a labeled diagram of Bomb Calorimeter.
40.	What is spontaneous process? Give two examples.
41.	Prove that $\Delta E= q_v$
42.	Why heat energy is released in exothermic reactions?
43.	Why the work done by the system is taken as negative?
44.	Give two applications of electrochemical series.
45.	A salt bridge maintains the electrical neutrality in the cell. Explain.
46.	Write down the function of salt bridge?
47.	A porous plate or a salt bridge is not required in lead acid storage battery.
48.	Define electrochemical series.
49.	Write down two functions of salt bridge in a galvanic cell?
50.	Write down reactions taking place at the electrodes during the discharging of Nickel-Cadmium cell.
51.	What is standard electrode potential?
52.	Give chemical reactions taking place at anode and cathode in a fuel cell.
53.	Calculate the oxidation number of S in $Cr_2(SO_4)_3$ and SO_4^{2-}.
54.	Calculate the oxidation number of element underlined in the following compounds. K_2MnO_4 ,$Ca(ClO_3)_2$
55.	What is difference between electrolytic cell and voltaic cell?
56.	Voltaic cell is reversible cell. State.
57.	How fuel cells produce electricity?
58.	Give the chemistry of electrolysis of aqueous solution of sodium chloride.
59.	What is electrolysis? Give example.
60.	Lead accumulator is a chargeable battery. Justify.
61.	Write down electrode reactions of a dry cell.
62.	Differentiate between anode and cathode.

Long Questions

Q.NO.5 Ch # 1+5

A	What is difference between actual yield and theoretical yield? Why actual yield is lesser than theoretical yield?	B	Give defects of Bohr's atomic model.
A	Define limiting reactant. How is it helpful to control chemical reaction?	B	Derive the formula for calculating the energy of an electron in nth orbit using Bohr's model.
A	What is stoichiometry? Give its assumptions. Mention two important laws which help to perform the stoichiometric calculations.	B	Describe Millikan's oil drop method for the measurement of charge on electron.
A	Write down the steps to calculate empirical formula of a compound.	B	Write down the properties of cathode rays.
A	Define stoichiometry. Give its assumptions. Discuss the mass-mass relationship during stoichiometry.	B	Define Quantum numbers. Discuss briefly Azimuthal quantum number.
A	Write down four properties of Neutron.	B	How did Rutherford discover the nucleus of atom?
A	Describe J.J Thomson's experiment for determining e/m value of electron.		

Q.NO.6 Ch # 3+10

A	Calculate the density of methane at STP. What happens to…(Example#4).	B	What is electrochemical series? Explain its any three applications.
A	250 cm^3 of hydrogen effuses 4 times more ……(Example#7)	B	How can you measure electrode potential of an element using standard hydrogen electrode (SHE)?
A	1 mole of methane gas……(Example# 08)	B	What is standard hydrogen electrode (SHE)? How it is used to measure the electrode potential of Zinc.
A	Calculate the mass of $1dm^3$ of ammonia at…..(Example# 5)	B	Describe the electrolysis of molten sodium chloride and a concentrated aqueous solution of sodium chloride.
A	Calculate the mass in dm^3 of NH_3 gas at 300 and 1000 mm Hg pressure, considering that is NH_3 behaving ideally.	B	Describe Nickel-Cadmium cell (Rechargeable).
A	A sample of nitrogen gas in enclosed in a vessel of volume at and pressure of 101325 . This gas is transferred to a flask and cooled to . Calculate the pressure in exerted by the gas at .	B	Explain the construction and working of fuel cell.
A	One mole of methane gas in maintained at 300K. Its volume is calculate its pressure considering it behaves ideally.	B	Give explanation of discharging and recharging of lead accumulator along with reactions occurring at electrodes.

Q.NO.5 Ch # 6+7

A	Define electron affinity. Name the factors effecting it. How does it vary in the periodic table?	B	State and explain with an example, the Hess's law of constant heat summation.
A	Give four postulates of Valence Shell Electron Pair Repulsion Theory.	B	Define enthalpy of reaction. How is it measured by glass calorimeter?
A	Define the term electronegativity. Discuss its variation in the periodic table.	B	Define enthalpy and prove that $q_p = \Delta H$
A	Describe sp2 hybridization giving example of Ethene.	B	What is the first law of thermodynamics? Give its mathematical form.
A	Explain the paramagnetic behaviour of oxygen using molecular orbital theory.	B	Describe Bomb Calorimeter.

Q.NO.5 Ch # 4+8

A	Define hydrogen bonding. How does it explain the solubility of hydrogen bonded molecules and structure of ice?	B	Example# No. 7
A	Explain hydrogen bonding in NH_3, H_2O and HF. How is it helpful in explaining the structure of ice?	B	Example # No. 6
A	What is boiling point? What is the effect of external pressure on the boiling point? Why the temperature remains constant at boiling point although heat is continuously supplied.	B	Example # No. 4
A	What are ionic solids? Give their properties in details.	B	Example # No. 2

A	What are molecular solids? Give examples and explain their properties?	B	Example # No. 5
A	What is hydrogen bonding? Discuss hydrogen bonding in biological properties.	B	Exercise Numerical Q.23(a)
A	What is vapor pressure of a liquid? Also discuss its measurement by Monomeric method.	B	Exercise Numerical Q.25

Q.NO.5 Ch # 9+11

A	What are ideal solutions? Explain the fractional distillation of ideal mixture of two liquids.		How does Arrhenius equation help us to calculate the energy of activation of a reaction?
A	Differentiate between ideal and non-ideal solutions.		How rate of reaction depends on the following factors? a) Nature of reactants (b) Surface area c) rate of reaction
A	Give three statements of Raoult's law and also mention how Raoult's law helps us in determining the ideality of		What is order of reaction? Describe two methods (half life method and large excess method) for finding the order of reaction.
A	solution?		Define order of reaction and explain 2nd order and zero order reactions.
A	Describe freezing point depression method to determine the molecular mass of an organic compound.		What is catalytic poisoning? Give two examples.
A	Describe one method to determine the boiling elevation of a solution.		What are enzymes? Give examples in which they act as catalyst. Mention the characteristics of enzymes.