ANNUAL 2026

NEW EXAMINATION POLICY

پنجاب کے تمام بورڈز کے لیے (According to Smart Syllabus)

کامیابی کا تعویذ

صرف 2 ماہ تیاری کر کے پڑھائی میں کمزور طلباء و طالبات A+ گریڈ میں کامیابی حاصل کر سکتے ہیں۔

(Solved)

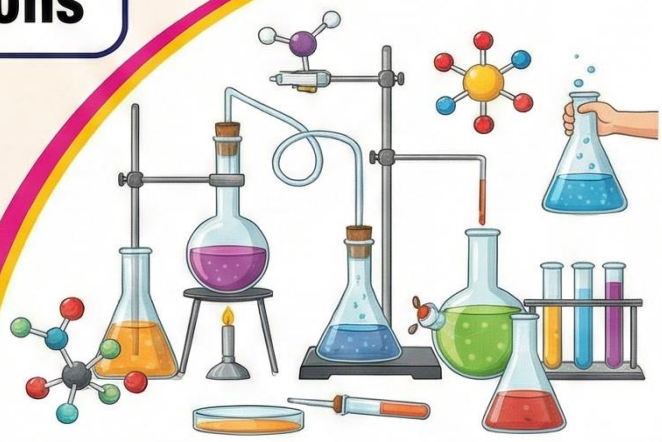
CHEMISTRY

For Inter Part-I

* Most Imp. MCQs
* Most Imp. Short Questions
* Most Imp. Long Questions

A+ CHALLENGE

اب فیل ہونا بھول جائیں



(محمد قدیر رفیق)

القدیر جناح سائنس اکیڈمی

ملیاں کلاں مرید کے روڈ شیخوپورہ

Contact: 0302-4741124

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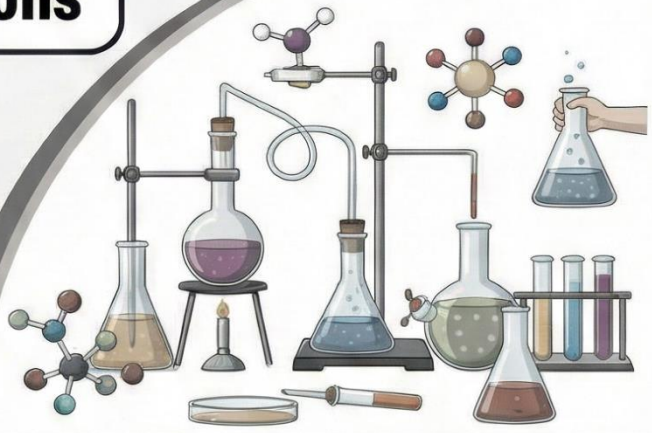

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CHEMISTRY-11

Sr.	Question	A	B	C	D
1	In modern periodic table 6^{th} period contains elements:	8	18	10	32✓
2	Select the two normal elements present in sixth period:	K, Ca	Rb, Sr	Cs, Ba✓	La, Hf
3	Out of all the elements of group VI-Athe highest melting and boiling points is shown by:	Te✓	Se	S	Po
4	Which conditions result in higher ionization energy?	**Smaller atom, higher Z_{eff}**✓	Smaller atom, lower Z	Larger atom, higher Z	Larger atom, lower Z
5	If element $x$is in period 3and group 2, which element is it?	Sodium	Magnesium✓	Aluminium	Argon
6	What does the period number (n)indicate?	**Shell number** ✓	Valence electrons	Atomic number	Mass number
7	What does the group number represent?	Meter	Picometer	Nanometer	Micrometer
8	Amount of energy released by absorbing e^-in valence shell:	I.E.	E.A.✓	Electronegativit	Atomization
9	Pick the element having least ionization energy value:	N	O✓	F	Ne
10	Why does atomic radius decrease across a period?	More shells	**Increasing Z_{eff}**✓	Less shielding	More electrons
11	Which halogen has the lowest E.A.?	Cl	Br	At✓	I
12	What happens to metallic character from left to right?	Increases	Decreases✓	Stays same	Random
13	Sodium is kept under Kerosine oil because:	Colour	**Reacts with air/H_2O** ✓	Reacts with oil	Precious
14	What type of oxides form alkalis with H_2O?	Acidic	Basic✓	Amphoteric	Neutral
15	What kind of bonds hold acidic oxides?	Ionic	Metallic	Covalent✓	Hydrogen
16	Al_2O_3reacts with HClto form:	AlO_2	$AlCl_3$✓	$NaAlO_2$	$Al(OH)_3$
17	The hydrides of Group IAare:	Ionic✓	Covalent	Metallic	Interstitial
18	Which one is an ionic hydride?	NaH✓	AlH_3	NH_3	CH_4
19	Which one is an amphoteric oxide?	SO_3	CaO	ZnO✓	Li_2O
20	Oxides of Beare:	Acidic	Basic	Amphoteric✓	Neutral
21	Oxidation number of Sin $SO_3$is:	+6✓	+3	+2	+4
22	Oxidation number of Pin $P_4O_{10}$is:	+3	+5✓	+7	+9
23	Oxidation number of Siin $SiCl_4$is:	+1	+2	+4✓	+3
24	Which atom shows variable oxidation states?	Sodium	Magnesium	Aluminum	Phosphorous ✓
25	Oxidation state of Sin SO_4^{2-}:	+4	+2	+6✓	0
26	Number of neutrons in $^{39}_{19}$Kis:	39	19	20✓	21

27	Relationship between X-ray frequency and $Z$was by:	Moseley✓	Rutherford	Bohr	Dalton
28	Nucleon number $(A)$is the sum of:	$p^+ + e^-$	$p^+ + n^0$✓	$n^0 + e^-$	$e^- + e^+$
29	Formula to find number of neutrons (N):	$N = A + Z$	$N = A - Z$✓	$N = Z - A$	$N = A \times Z$
30	How are neutrons affected by an electric field?	To positive	To negative	**No deflection** ✓	Randomly
31	Which particle has the smallest mass?	Proton	Electron✓	Neutron	All equal
32	Angle of deflection is proportional to:	mass	m/e	None	e/m✓
33	Particle deflected most in a magnetic field:	Proton	Neutron	Electron✓	All same
34	Visible range wavelength (λ):	200-400 nm	400-750 nm ✓	200-800 nm	800-1200 nm
35	Spectrum showing colored lines from hot gases:	Absorption	Emission✓	Continuous	Infra red
36	Dark lines in an absorption spectrum are due to:	Reflection	Scattering	Absorption✓	Emission
37	Wavelengths of dark lines in absorption match:	Infra red	Different	Emission lines ✓	Random
38	Number describing size and energy of orbitals:	Azimutual	Magnetic	Principal✓	Spin
39	Developed model requiring 3 quantum numbers:	Bohr	Plank	Schrodinger✓	Einstein
40	Azimutual quantum number (l)describes:	Size	Shape✓	Orientation	Spin
41	Which sub-shell has $l = 1$?	s	p✓	d	f
42	Orbital which is spherical and symmetrical:	s-orbital✓	p-orbital	d-orbital	f-orbital
43	Characteristic of d-orbitalshapes:	Spherical	Two lobes	Uniform	Diff. shapes✓
44	Differentiation of degenerate orbitals is caused by:	ΔE	Spin direction	Magnetic field ✓	Shell No.
45	Energy order after $4s$(Aufbau Principle):	$3d$✓	$4p$	$5s$	$3p$
46	Electrons being paired with opposite spins:	↑	↑↑	↑↓✓	No arrows
47	Ground state electron position:	In nucleus	2^{nd} shell	**Near nucleus** ✓	Far from nucleu
48	Max electrons in f-subshell$(2(2l + 1))$:	2	6	10	14✓
49	Orbitals having same energy:	Hybrid	Valence	Degenerate✓	d-orbitals
50	Most stable electronic configuration:	Noble gas✓	Electronegative	Alkali metal	Halogen
51	$BeCl_2$follows ____ hybridization:	sp✓	sp^3	sp^2	sp^3d^2
52	Maximum number of unpaired electrons:	O_2✓	O_2^+	O_2^-	O_2^{2-}
53	Hybridization of Cin ethane (C_2H_6):	sp	sp^2	dsp^2	sp^3✓
54	Bond angle between sp^2orbitals:	180°	109.5°	120°✓	107.5°
55	Hybridization in methane (CH_4):	sp	sp^2	sp^3✓	sp^4
56	Geometry of sphybrid orbitals:	Tetrahedral	Linear✓	Trigonal planar	Octahedral
57	Molecule explained by sphybridization:	H_2O	NH_3	CH_4	$BeCl_2$✓
58	Number of orbitals mixing in sp^3:	2	3	4✓	5
59	Bond angle in NH_3molecule:	109.5°	107.5°✓	104.5°	108°
60	Carbon in methane is hybridized:	sp^3✓	sp^2	sp	dsp^2

61	The geometry of NH_3:	Linear	Trigonal planar	Tetrahedral	**Trig. Pyramidal ✓**
62	Which molecule shows linear geometry?	CH_4	$AlCl_3$	H_2O	$BeCl_2$✓
63	Strongest electron pair repulsion:	bp-bp	bp-lp	lp-lp✓	lp-bp
64	Shape of XeF_4(excluding lp):	Square planar ✓	Tetrahedral	Octahedral	Trigonal planar
65	A receptor mostly binds to:	Ligands✓	Substrates	Antibodies	Enzymes
66	Aspirin relieves:	Fever	Infection	Allergy	Pain✓
67	Bonds in Nitrogen molecule (N_2):	$1\sigma, 1\pi$	$1\sigma, 2\pi$✓	3σ	$2\sigma, 1\pi$
68	Bond order (B.O.)of O_2^{2-}:	1✓	2	0	3
69	Bonds in Oxygen molecule (O_2):	$1\sigma, 1\pi$✓	$1\sigma, 2\pi$	3σ	$2\sigma, 2\pi$
70	Incorrect statement for BMO:	Less energy	**Low density in center ✓**	Attraction	Constructive
71	Bond order of He_2:	2	3	0✓	1
72	Symmetrical feature of BMOin H_2:	Density	**Nuclear axis ✓**	Size	Energy
73	Electron density in anti-bonding (ABMO):	Nuclear axis	Between nuclei	**Away from nuclei ✓**	Randomly
74	Why does H_2molecule not form?	B.O. $= 0$✓	No e^-	lp	High energy
75	Attracted between poles of strong magnet:	Liq. N_2	Liq. H_2	Liq. O_2✓	Solid O_2
76	9 gof H_2Ocontains H-atoms:	N_A	$2 \times N_A$✓	$0.5N_A$	$3N_A$
77	Formula for moles (n):	$m \times M$	$m + M$	m/M✓	None
78	Molar mass of 1mole of CO_2:	12 g	44 g✓	60 g	56.1 g
79	Moles in 20 gof NaOH($M = 40$):	0.2 mol	0.5 mol✓	2 mol	5 mol
80	Moles of CO_2containing 8.0 goxygen:	0.25✓	0.15	0.35	1.45
81	Same number of molecules at STP:	280 cm^3 CO_2/ N_2O✓	11.2 dm^3 /32g O_2	44g/ 11.2 dm^3	28g/ 5.6 dm^3
82	Ions in 1mole of Na^+:	6.02×10^{22}	6.02×10^{23}✓	9.61	23
83	What does Juglone kill?	Grass	Animals	Non-comp.	**Comp. plants ✓**
84	Molar volume (V_m)at STP:	54 dm^3	22.414 dm^3✓	2.24 dm^3	2.4 dm^3
85	Law of equal volumes/equal molecules:	Charles	Boyle	Dalton	Avogadro✓
86	Volume of 2.5moles of Cl_2(2.5×22.4):	44.8 dm^3	56.0 dm^3✓	89 dm^3	1.2 dm^3
87	Density (d)equals:	V/M	$M \times V$	M/V✓	$M + V$
88	Formula to find volume (V):	m/d	$m \times d$	n/V_m	$n \times V_m$✓
89	Mass of gas if $d = 1.9$ g/dm^3at STP:	44.1 g/mol✓	22.4 g/mol	1.97 g/mol	4.24 g/mol
90	Molar mass if $d = 1.43$ g/dm^3:	14.3	22.4	32.0✓	64.0
91	Relationship c, n, V:	$C = n \times V$	$C = V/n$	$C = n/V$✓	$n = C/V$
92	Formula for $n$from $V$and C:	$n = C \times V$✓	$n = V/C$	$n = C/V$	$n = V/C$
93	Molar mass of $KMnO_4$:	138.2	27.64	158✓	7.9
94	27 g Alreacts with how much O_2:	8 g	16 g	32 g	24 g✓
95	Calculation based on balanced equation:	Complex	Stoichiometric ✓	Non-stoich.	None

96	Quantitative relationship chemistry branch:	Stoichiometry ✓	Thermometry	Organic	Physical
97	Mole ratio $N_2 : H_2$ for NH_3 (1: 3):	1: 1	1: 2	1: 3✓	1: 4
98	Measured in diabetic patients:	Protein	Fat	Glucose✓	Water
99	Helps identify steroids/stimulus:	Blood test	X-ray	MRI	Urine test✓
100	Enhanced performance substances:	Vitamins	Steroids✓	Antibiotics	Glucose
1	Why does gas exert pressure on container walls?	**Particle collision ✓**	Particle volume	Temperature	Nature of container
2	Who explained that gas particles have negligible volume?	Boyle's law	Charles' law	**Ideal gas theory ✓**	Avogadro's law
3	Why are intermolecular forces considered negligible in ideal gases?	High volume	High pressure	Low temperature	**Large distance between particles ✓**
4	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 ✓**
5	What happens to a gas during Joule – Thomson expansion?	Pressure increases	**Temperature changes ✓**	Volume changes	Gas ionizes
6	Who helps to understand properties of gases through molecular theory?	**Clausius / Maxwell / Boltzmann ✓**	Lavoisier	Newlands	Moseley
7	Why is diffusion slower in liquids than gases?	**Strong inter-molecular forces ✓**	Smaller molecules	Higher energy	Volume unchanged
8	Acetone and chloroform are soluble in each other due to:	**Intermolecular hydrogen bonding ✓**	Dipole-dipole interaction	Instantaneous dipoles	All of the above
9	Liquid Hydrocarbon is:	Methane	**Pentane ✓**	Hexane	Propane
10	What usually decreases surface tension?	Pressure	**Temperature ✓**	Volume	Density
11	Which liquid has the highest surface tension at 20°C?	Ethyl alcohol	**Water ✓**	Acetone	Benzene
12	What is surface tension of acetone at 20°C?	**23.70 ✓**	72.75	22.75	22.61
13	What happens to viscosity when temperature increases?	Fluctuates	Stay same	Increases	**Decreases ✓**
14	What stronger intermolecular forces cause?	**Higher viscosity ✓**	Lower viscosity	No charge	Faster flow
15	What causes cooling during evaporation?	Volume increase	**Molecules escaping ✓**	Temperature rise	Pressure increase
16	Which vessels keep water cooler by evaporation?	Metal	Plastic	Glass	**Earthenware ✓**
17	Vapour pressure is not affected by:	Temperature	Intermolecular forces	**Surface area ✓**	Pressure

18	What is affected by intermolecular forces?	**Boiling point ✓**	Volume	Density	Colour
19	Which liquid has the highest B.P.?	CCl4	**Water ✓**	Naphthalene	Acetone
20	What happens when molecules return to the liquid?	**Condensation ✓**	Freezing	Evaporation	Heating
21	Which term describes molecules changing state in vapour pressure?	Melting	Boiling	**Phase change ✓**	Freezing
22	Which is pseudo solid?	Diamond	**Glass ✓**	NaCl	Graphite
23	Whose shape changes with impurities?	Water	Glass	Gas	**NaCl crystal ✓**
24	Liquid crystals are used in the display of:	Neon signs	Fluorescent bulbs	**T.V. displays ✓**	Lightning discharge
25	Which way do liquid crystal molecules orient?	Randomly	Perpendicular	Against each other	**Parallel ✓**
26	What does enthalpy change represent?	Mass change	Volume change	Pressure change	**Heat absorbed or evolved ✓**
27	What is denoted by "H" in Chemistry?	**Heat content ✓**	Volume	Temperature	Pressure
28	The ΔHvalue for the combustion of carbon?	+393.7 kJ/mol	**-393.7 kJ/mol ✓**	+16.2 kJ/mol	-16.2 kJ/mol
29	The change in heat contents of a chemical reaction at constant temperature and pressure is called:	**Enthalpy change ✓**	Bond Energy	Heat of Sublimation	Internal Energy Change
30	What is energy of activation denoted by?	AH	HP	**Ea ✓**	Hr
31	What happens in enthalpy change of neutralization?	**Water formed ✓**	Gas evolved	Solid formed	Heat absorbed
32	The $\Delta H_{no}$for strong acids and bases?	+57.1 kJ/mol	**-57.1 kJ/mol ✓**	+74.8 kJ/mol	-74.8 kJ/mol
33	Which of the following equations represents standard heat of formation of CH_4?	2C (diamond) + $2H_2$	2C (graphite) + $2H_2$	**C (graphite) + $2H_2$✓**	2C (diamond) + H_2
34	Which bond has energy of 615 kJ/mol?	**C=C ✓**	C-C	N-N	O=O
35	Which molecule has C-C bond energy of 376 kJ/mol?	Propane	Butane	Methane	**Ethane ✓**
36	What kind of energy does bond breaking absorb?	**Positive ✓**	Negative	Zero	Neutral
37	Which reaction is exothermic according to sign?	Positive	**Negative ✓**	Zero	Undefined
38	Which formula is used to calculate ΔH_{ro}?	Sum of reactants only	Sum of products only	Sum of reactants minus sum of products	**Sum of products minus sum of reactants ✓**
39	What does a calorimeter measure?	**Heat change ✓**	Pressure	Volume	Mass

40	How is qCalculated from heat transferred in the solution?	$q = m \times c \times \Delta T$	$q = c \times m \times \Delta T$ ✓	$q = m/c \times \Delta T$	$q = \Delta T/m \times c$
41	What unit measures calorie content in food?	Kilograms	Kilojoules	**Kilocalories** ✓	Grams
42	What is the calorie content of glucose per gram?	15.57 kcal/g	**3.72 kcal/g** ✓	4.184 kcal/g	2.8 kcal/g
43	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	**Hess's law** ✓	Joule's principle	Law of conservation of energy
44	What does the law of heat summation help to validate?	**Enthalpy values** ✓	Calorimeter Accuracy	Temperature scales	Molecular Mass
45	What is bond energy of H-H used?	242 kJ/mol	391 kJ/mol	431 kJ/mol	**436 kJ/mol** ✓
46	How is charge density defined?	**Charge per ion** ✓	Charge per mass	Charge per volume	Charge per unit area
47	What is the symbol for enthalpy change of hydration?	ΔH_{hyd} ✓	ΔH_{Latt}	ΔH_{sol}	ΔH_r
48	The born-Haber cycle is the best application of _____ law?	Boyle's	Dalton's	**Hess's** ✓	Graham's
49	What does the Born-Haber cycle calculate?	Heat capacity	Ionization energy	**Lattice energy** ✓	Enthalpy of combustion
50	Which Law is applied to Born-Haber cycle?	Boyle's Law	Charles' Law	Avogadro's Law	**Hess' Law** ✓
51	What is studied in reaction kinetics?	**Rate of chemical** ✓	Colour of chemicals	Volume of gases	Pressure of gases
52	Which reaction is very Fast?	Rusting	**NaCl with $AgNO_3$** ✓	Hydrolysis	Rusting of Iron
53	What process is slow?	Explosion	Ester hydrolysis	**Rusting of iron** ✓	NaCl with $AgNO_3$
54	How fast does an explosion happen?	Months	Years	Days	**Fraction of a second** ✓
55	Which particles must collide for a chemical reaction?	**Reactants** ✓	Products	Solvents	Solutes
56	What does collision theory explain?	Reaction Rates	**How reactions occur** ✓	Energy changes	None
57	Activation energy is related to:	Temperature	**Effective collision** ✓	Pressure	Volume
58	The rate of reaction _____ as the reaction proceeds.	Increases	**Decreases** ✓	Remains the same	May decrease or increase
59	How is rate of reaction defined?	Change in time	Change in pressure	**Change in concentration** ✓	Both A & B
60	What happens to slope as reaction proceeds?	Becomes steep	Remains same	Becomes more steep	**Becomes less steep** ✓

61	What type of graph illustrates rate of reaction?	Bar graph	**Concentration vs time ✓**	Pie chart	Bell graph
62	The electrical conductivity method measures:	Mass	Volume	Colour	**Ion concentration ✓**
63	What does activation energy represent?	**Energy needed ✓**	Energy released	Energy lost	Energy stored
64	The process in which catalyst and reactant are in different phases is called:	Homogeneous catalysis	**Heterogeneous catalysis ✓**	Auto catalysis	Negative catalysis
65	Who speeds up the reaction due to increased concentration?	**More collisions ✓**	Less particles	Decreasing temperature	Less energy
66	What molecules are involved in hydrolysis of esters catalyzed by H_2SO_4:	Gases	**Liquids ✓**	Solids	Metals
67	Who catalyzes hydrogenation of ethene?	**Ni ✓**	NO	MnO_2	$CuCl_2$
68	Increasing the temperature of a chemical reaction increases the rate of a reaction because:	**Both collision frequency and energies increase ✓**	Collision frequency increases	Activation energy increase	Activation energy decrease
69	Consider two reactions with different activation energies at the same temperature. The reaction with the lower activation energy will have:	A smaller rate constant	**A larger rate constant ✓**	The same rate constant	A rate constant that depends on enthalpy
70	Unit of rate constant is the same as that of the rate of reaction in:	**Zero order reaction ✓**	1st order reaction	2nd order reaction	3rd order reaction
71	All radioactive disintegration nuclear reactions are of:	**First-order ✓**	Second-order	Third-order	Zero order
72	Photochemical reactions are:	**Zero Order Reaction ✓**	First Order Reaction	Second Order Reaction	Third Order Reaction
73	What does a rate law relate?	Pressure & Temperature	**Reactants concentration & Rate ✓**	Time and volume	Catalyst and product
74	The overall order of reaction is equal to:	**Sum of exponents ✓**	Product of exponents	Difference of exponents	Ratio of exponents
75	What is overall order of the rate of equation rate = $k[H_2][NO]^2$?	First	second	**Third ✓**	Zero
76	Which events involve phenomenon observable with naked eye?	**Macroscopic ✓**	Microscopic	Atomic	Molecular
77	Which events cannot be observed with naked eye?	Macroscopic	**Microscopic ✓**	Atomic	Molecular
78	The reaction which proceeds in both forward and backward directions is called:	Irreversible reaction	**Reversible reaction ✓**	Spontaneous reaction	Non spontaneous reaction
79	What happens to bonds in CO & H_2O during reverse reaction?	**Broken ✓**	Formed	Ignored	Converted

80	What role does a catalyst play in equilibrium?	Changes equilibrium constant	**Attains equilibrium earlier ✓**	Shifts equilibrium position	Stops reaction
81	Which factor does not affect the equilibrium constant?	Temperature	**Catalyst ✓**	Initial concentration	Pressure
82	What type of equilibrium occurs when reactants and products in the same phase?	**Homogenous ✓**	Heterogeneous	Dynamic	Static
83	Which equilibrium involves reactants and products in different phase?	Homogenous	**Heterogeneous ✓**	Dynamic	Static
84	Among the following which equation has no unit of Kc:	$PCl_5 \rightleftharpoons PCl_3 + Cl_2$	$N_2 + O_2 \rightleftharpoons 2NO$	$N_2 + 3H_2 \rightleftharpoons 2NH_3$	$H_2 + I_2 \rightleftharpoons 2HI$ ✓
85	According to law of Mass Action what is the rate of forward reaction proportional to?	CD	AB ✓	AD	BC
86	The equilibrium constant expression (Kc) for $aA + bB \rightleftharpoons cC + dD$ is:	$[A]^a[B]^b / [C]^c[D]^d$	$[C]^c[D]^d / [A]^a[B]^b$ ✓	$[A][B] / [C][D]$	$[C][D] / [A][B]$
87	What does the position of equilibrium refer to?	Temperature	Pressure	Reaction rate	**Relative amounts ✓**
88	Which way does equilibrium shift if product concentration increases?	**Left ✓**	Right	Up	Down
89	Who proposed the principle that systems counteract disturbances?	Dalton	**Le Chatelier ✓**	Maxwell	Arrhenius
90	What happens when a system in equilibrium is disturbed?	System stops	Changes temperature	**Nullifies disturbances ✓**	Changes phases
91	Which effect is not an application of Le-Chatelier's principle?	Change in pressure	Change in concentration	Change in temperate	**Change in volume ✓**
92	Who produces more $BiOCl$ when $BiCl_3$ or HCl is added?	**Forward Reaction ✓**	Catalyst	Reverse Reaction	Water
93	What occurs if $BiCl_3$ is removed at equilibrium?	Forward reaction	**Reverse reaction ✓**	Volume increase	Kc changes
94	What condition allows pressure or volume effects?	No gases involved	Temperature constant	Pressure constant	**Reactants ≠products ✓**
95	What does the reaction favour if pressure decreases?	Forward	Reverse	Equilibrium	**Side with more moles ✓**
96	Where does the reaction move if volume is increased?	Lesser volume side	No movement	**Greater Volume side ✓**	Both sides equally
97	Consider the equilibrium $N_2 + 3H_2 \rightleftharpoons 2NH_3$. If the concentration of N_2 is increased, the concentrations of H_2 will:	Increase	**Decrease ✓**	Remain the same	Change irregularly
98	For a specific reaction, the value of the equilibrium constant, Kc:	Always remains the same	Increases if product increases	**Changes with temperature ✓**	Increases if reactant increases
99	Catalyst used in preparation of NH_3 from N_2 and H_2 is:	**Fe ✓**	Ni	Pt	V_2O_5

100	What is the effect of catalyst on equilibrium position?	Shift left	Shift right	Stops reaction	**No change ✓**
1	The ionization constant of pure water at 25°C is:	**1.8×10^{-16} moles dm^{-3} ✓**	1.6×10^{-16} moles dm^{-3}	1.0×10^{-14} $moles^2$ dm^{-6}	1.8×10^{-14} $moles^2$ dm^{-6}
2	What is the equilibrium constant symbol for water self–ionization?	Ka	**Kw ✓**	Kt	Kb
3	What is the pH of pure water at 25°C?	0	**7 ✓**	10	14
4	The pH of 10^{-3} mol dm^{-3} of an aqueous solution of H_2SO_4 is:	3.0	**2.7 ✓**	2.0	1.5
5	pOH of water is:	2.0	4.0	6.0	**7.0 ✓**
6	The pH of human blood is:	7.12	**7.35 ✓**	7.56	8.0
7	If the pH of a solution is 11, what is the $[OH^-]$ concentration?	10^{-11}	10^{-14}	**10^{-3} ✓**	10^{-7}
8	Acid having Ka > 1 will be:	Weak	Very weak	Moderate	**Strong ✓**
9	What happens to Ka as acid strength decreases?	Increases	**Decreases ✓**	Stays same	Zero
10	What assumption is usually made for weak acid dissociation?	No dissociation	**Partial, initial ✓**	Instant equilibrium	Complete dissociation
11	What is pH of 0.1 M acetic acid if Ka = 1.8×10^{-5}?	About 1	About 1.5	About 2.5	**About 2.9 ✓**
12	The pH of mixture of CH_3COONa and CH_3COOH is:	7	>7	**<7 ✓**	1
13	Which ion increases when HCl gas is passed through brine?	Na^+	**Cl^- ✓**	H^+	OH^-
14	What effect does adding a common ion have on solubility?	Increases	Doubles	**Suppresses ✓**	No change
15	What happens to H_2S in acidic solution due to common ion?	More dissociated	No dissociation	Fully ionized	**Less dissociated ✓**
16	A solution which resists to change its pH is called as:	**Buffer solution ✓**	Acid solution	Standard solution	Basic solution
17	What happens when NaOH is added to CH_3COOH?	Produce CO_2	**Neutralization ✓**	Precipitation	Ionization
18	What does the common ion effect suppress in CH_3COOH?	**Ionization ✓**	Precipitation	Evaporation	Freezing
19	Excess $AgNO_3$ added to $BaCl_2$ and filtered. Main ions in 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^-
20	What does $PbSO_4$ dissociate into water?	**Pb^{2+} & SO_4^{2-} ✓**	Pb^{2+} & SO_3^{2-}	Pb^{2+} & SO_3	Pb^{4+} & SO_2
21	Why is the solubility product temperature dependent?	Concentration change	Ion charge	Ion size	**Equilibrium constant change ✓**
22	What causes a sharp colour change in indicators?	Ionization	**pH change ✓**	Temperature	Pressure
23	What pH range suits methyl orange?	**3.1 – 4.5 ✓**	1.0 - 2.0	6.0 – 7.0	8.2 – 10.0
24	Which color determines the equivalence point (Phenolphthalein)?	**Pink ✓**	Orange	Yellow	Colourless

25	What determines the equivalence point?	pH	Volume Added	**Indicator colour ✓**	Concentration
26	What is the oxidation Number of pure elements like Na or Cl_2?	+1	-1	**Zero ✓**	+2
27	Whose oxidation state is +1 except in metal hydrides?	Oxygen	Nitrogen	Carbon	**Hydrogen ✓**
28	Substance acting as both oxidizing and reducing agent is:	Catalyst	Reactant	**Disproportionating agent ✓**	Solvent
29	What is oxidation number of oxygen in H_2O_2?	0	**-1 ✓**	-2	+1
30	What causes an element to act as an oxidizing agent?	**Gains electrons ✓**	Loses electrons	Remains neutral	Forms ions
31	What change happens to oxidation number of M in oxidation?	M^+ cation	Non-metal	M^- anion	**Increase in O.N. ✓**
32	What type of reaction is disproportionation?	Acid-base	Double displacement	**Simultaneous oxidation & reduction ✓**	Decomposition
33	Who acts as both oxidizing and reducing agent simultaneously?	Catalyst	**Single substance ✓**	Compound	Ion
34	Which electrode is negatively charged in an electrolytic cell?	Anode	**Cathode ✓**	Electrolyte	Ion
35	Which electrode is positively charged in an electrolytic cell?	Ion	Electrolyte	**Anode ✓**	Cathode
36	What happens at the cathode during electrolysis?	Oxidation	**Reduction ✓**	Ionization	Neutralization
37	Which ions lose electrons at the anode?	Cations	Neutrons	Protons	**Anions ✓**
38	What happens to cathode mass during electrolysis?	**Increases ✓**	Decreases	Unchanged	Evaporates
39	The electrode potential is measured in?	Amperes	Coulombs	Ohms	**Volts ✓**
40	Which gas is used in the standard hydrogen electrode?	Oxygen	Nitrogen	**Hydrogen ✓**	Helium
41	What covers the platinum electrode in SHE?	**Platinum black ✓**	Copper	Zinc	Silver
42	What is the standard voltage for SHE?	1.00 V	**0.00 V ✓**	0.50 V	0.25 V
43	What symbol represents standard electrode potential?	Q	V	I	**E° ✓**
44	Standard electrode potential of Zn^{2+}/Zn half-cell?	0.00 V	+0.34 V	**-0.76 V ✓**	+0.76 V
45	Standard electrode potential of Cu^{2+}/Cu half-cell?	-0.34 V	0.00 V	-0.76 V	**+0.34 V ✓**
46	Reaction responsible for electricity in voltaic cell is:	**Redox reaction ✓**	Oxidation reaction	Reduction reaction	Hydrolysis
47	What completes the electrical circuits in an electrochemical cell?	Wire only	**Salt bridge ✓**	Voltmeter	Battery
48	What does the word photovoltaic mean?	Light and current	Energy & power	Electron & hole	**Light & voltage ✓**

49	What does a P-N junction in PV cells do?	Increase light	**Separates electrons & holes ✓**	Absorbs photon	Generates heat
50	Why are photovoltaic cells sustainable?	Use fossil fuels	Work with light	Requires Batteries	**Sunlight Continuously ✓**
51	Formula of chloroform is:	CH_4	CH_3Cl	CH_2Cl_2	**$CHCl_3$ ✓**
52	Which hydrocarbons are used as fuels?	Aromatic	Aliphatic	**Both A & B ✓**	None
53	What are compounds that are not aromatic called?	Aromatic	**Aliphatic ✓**	Cyclic	Saturated
54	Which hydrocarbons may be open chain or cyclic?	Aromatic	Benzene	**Aliphatic ✓**	Phenol
55	Which hydrocarbons have high carbon to hydrogen ratio?	Aliphatic	Alkanes	Saturated	**Aromatic ✓**
56	What functional group is involved in aromatic hydrocarbons?	**Benzene ✓**	Alkane	Alkene	Alkyne
57	Which aromatic compound is a derivative of benzene?	Cyclohexane	**Toluene ✓**	Cyclopropane	Ethane
58	Where does the term aromatic come from?	Latin	Arabic	**Greek ✓**	French
59	What determines the parent name of an alkane?	Substituent	Branch length	Molecular Weight	**Longest carbon chain ✓**
60	Which alkyl group is 1-methyl propyl?	N-Butyl	**Sec-Butyl ✓**	Iso-Butyl	Ter-Butyl
61	Where do you place the substituent number in a name?	After parent	**Before parent ✓**	In the Middle	At end
62	How should substituent groups be listed in a name?	By size	By carbon number	By appearance	**Alphabetically ✓**
63	Which chain is chosen when lengths tie?	Fewer substituent	Alphabetical order	Shorter chain	**More substituents ✓**
64	What is the shape representation used for organic molecules?	Ball & Stick	Line Formula	Space Filling	**Structural Formula ✓**
65	What happens to B.P as carbon number increases?	Decrease	Stay same	**Increase ✓**	Random
66	Which compounds are insoluble in water but soluble in hexane?	Alcohols	Acids	Salts	**Alkanes ✓**
67	What must be used twice when substituents are on same carbon?	**Same number ✓**	Different numbers	No Number	Alphabetic order
68	A double bond consists of:	Two sigma bonds	**One sigma and one pi bond ✓**	One sigma and two pi	Two Pi-bonds
69	Presence of a double bond in a compound is the sign of:	Saturation	**Un-saturation ✓**	Substitution	None
70	What occurs during homolytic fission?	Ion Formation	**Free radicals ✓**	Electron pair sharing	Bond strengthening

71	What type of bond cleavage produces ions?	Homolytic	Covalent	**Heterolytic ✓**	Ionic
72	The catalytic oxidation of methane produces:	**Methanol ✓**	Ethane	CO_2	H_2
73	Why are alkanes unreactive to polar reagents?	Ionic character	High reactivity	Polar bonds	**Non-Polar bonds ✓**
74	What type of bond breaks in homolytic fission?	Ionic	**Covalent ✓**	Metallic	Hydrogen
75	What occurs during homolytic fission?	Unequal sharing	**Equal electron sharing ✓**	Radical formation	No bond breaking
76	Which theory explains basicity of ammonia?	Bronsted-Lowry	**Lewis ✓**	Arrhenius	Dalton
77	What forms Peroxy acetyl nitrate (PAN) in smog?	O_2 and NO	NO_2 and NO_2	**NO_2 and Hydrocarbons ✓**	S and NO
78	Which industry uses sulphur compounds?	Textile	Food	Paint	**Fertilizers ✓**
79	What is the Melting Point of nitrogen?	0°C	14°C	-195°C	**-210°C ✓**
80	Which property of nitrogen causes its bond to be non polar?	Size	Mass	**Symmetry ✓**	Colour
81	Nitrogen is inert in combustion because:	Denser than O_2	High specific heat	**Breaking N≡N bond energy ✓**	Noble gas
82	What shape does ammonia molecule have?	Tetrahedral	**Pyramidal ✓**	Linear	Square
83	Base used to synthesize NH_3 from ammonium chloride?	**$Ca(OH)_2$ ✓**	NH_4OH	NaOH	KOH
84	Which ion acts as an acid in the synthesis of ammonia?	OH^-	**NH_4^+ ✓**	Cl^-	Ca^{2+}
85	What is Kb value for ammonia's weak base equilibrium?	**1.8×10^{-5} ✓**	1.0×10^{-3}	1.8×10^{-4}	1.0×10^{-6}
86	Gas produced from ammonium salt and alkali (litmus blue)?	Hydrogen	Carbon dioxide	**Ammonia (NH_3) ✓**	Sulfur dioxide
87	Laughing gas is chemically?	NO	**N_2O ✓**	NO_2	N_2O_4
88	Brown gas formed when metal reduces HNO_3 to:	NO	N_2O	**NO_2 ✓**	N_2O_4
89	Which of the following is a reddish brown gas?	N_2O	NO	**NO_2 ✓**	N_2O_5
90	Brown ring test confirms the presence of:	Carbonates	Phosphates	**Nitrates ✓**	Sulphates
91	Which Nitrogen oxide is reddish brown gas?	NO	N_2O	**NO_2 ✓**	N_2O_4
92	What is the use of nitrous oxide?	Rocket fuel	**Dental Anaesthetic ✓**	Fertilizer	Explosive
93	The oxidation of NO in air produces:	N_2O	N_2O_4	**NO_2 ✓**	N_2O_5
94	Which of the elements gives acidic oxide?	**N ✓**	As	Sb	Bi
95	What are natural sources of NOx?	Vehicles	**Lightning ✓**	Chemical	Welding plants
96	Which gas dilutes oxygen to prevent fires?	CO_2	Ar	**N_2 ✓**	He
97	Which chemical initiates PAN formation?	Hydrocarbon	**NO_2 ✓**	Aldehyde	O_3

98	What is the role of NO_2 in smog?	Reducing agent	Catalyst	**Oxidizing agent ✓**	Inert gas
99	What process forms HNO_3 from nitrogen dioxide?	Haber-Bosch	Contact	Boudouard	**Ostwald ✓**
100	What does the Ostwald process produce?	NH_3	**HNO_3 ✓**	N_2O	NH_2Na
1	When methane reacts with Cl_2 in the presence of diffused light the products obtained are:	Chloroform only	Carbon tetra chloride	Methyl Chloride	**Mixture of A, B, C ✓**
2	What forms in the propagation step?	Halogen molecules	Ions	**New free radicals ✓**	Stable compounds
3	What is trichloromethane also known as?	Methane	**Chloroform ✓**	Carbon tetrachloride	Dichloromethane
4	What happens during the propagation step?	Radical termination	Radical disappearance	No reaction	**Mixture of haloalkanes formed ✓**
5	Which halogen increases anesthetic power of haloalkane?	**Chlorine ✓**	Fluorine	Bromine	Iodine
6	Which of the following reactions can an alkane undergo?	Addition	**Substitution ✓**	Polymerization	Nitration
7	Vinyl acetylene and conc. HClon reaction give:	Polyacetylene	Benzene	**Chloroprene ✓**	Divinyl acetylene
8	Preparation of vegetable ghee involves:	Halogenation	**Hydrogenation ✓**	Hydroxylation	Dehydrogenation
9	What is formed by dehydration of ethanol?	Methane	Ethanol	Water	**Ethene ✓**
10	Which acid is used in ethanol dehydration?	**Sulfuric Acid ✓**	Hydrochloric Acid	Nitric Acid	Acetic Acid
11	What results from dehydrohalogenation of ethyl bromide?	Ethane	**Ethene ✓**	Ethanol	Methane
12	What intermediate is formed during the electrophilic addition of $HBr$to an alkene?	**Carbocation ✓**	Carbanion	Radical	Epoxide
13	Why are alkenes weakly polar?	sphybridization	sp^2hybridization ✓	sp^3hybridization	No hybridization
14	Why type of bond lies perpendicular in alkenes?	Sigma	Ionic	**Pi ✓**	Metallic
15	What is bond angle around each carbon in ethene?	90°	109.5°	120°✓	180°
16	What type of carbocation is bonded to three carbons?	Methyl	Primary	Secondary	**Tertiary ✓**
17	What effect stabilizes tertiary carbocations?	Resonance	Electromeric	Mesomeric	**Inductive effect ✓**
18	What effect describes electron donating or withdrawing by adjacent groups?	**Inductive effect ✓**	Resonance effect	Magnetic effect	Ionic effect
19	Which of the following carbocations would be the least stable?	$(CH_3)_2CH^+$	CH_2CH_{2+}	CH_{3+}✓	$(CH_3)_3C^+$
20	Vinyl acetylene reach with $HCl$to form:	Polyacetylene	Benzene	**Chloroprene ✓**	Divinylacetylene

21	What causes bromine in $Br_2$to act as an electrophile?	Electron attraction	Ionization	**Electron repulsion ✓**	Heat
22	A free radical is a chemical species with an unpaired:	Proton	**Electron ✓**	Neutron	Ion
23	Which compound is formed after Ozonolysis of propene?	Ethanal	Methanal	Ethene	**Both A & B ✓**
24	What is the first step in the electrophilic addition reaction of alkenes?	Formation of a carbocation	Attack by a nucleophile	**Attack by an electrophile on the double bond ✓**	Formation of a free radical
25	The addition of unsymmetrical reagent to an unsymmetrical alkene follows:	**Markovnikov's rule ✓**	Hund's rule	Le Chatelier's principle	Aufbau principle
26	What is epoxidation producing?	Alcohol	**Epoxide ✓**	Alkane	Ketone
27	What does ozonolysis cleave?	Alkanes	Alcohols	**Alkenes ✓**	Alkynes
28	What does a catalytic converter reduce?	CO_2, N_2	H_2O	O_2	CO, NO, HC✓
29	Which catalyst is used for oxidation reactions?	Pt/Rh✓	Pt/Rd	Pt/Pd	Ni
30	Which compound is formed by PANs?	Ozone	Aldehydes	NO_2	$RCOOONO_2$ ✓
31	In a catalytic converter, $NO_x$to $N_2$and $O_2$is:	Oxidation	**Reduction ✓**	Combustion	Neutralization
32	What converts NH_3into nitrite?	Denitrifying bacteria	**Nitrifying bacteria ✓**	NH_3	O_2
33	Which bacteria function in anaerobic condition?	Nitrifying	Photosynthetic	**Denitrifying ✓**	Respiratory
34	What is released during denitrification?	NO_{3-}	N_2✓	NH_{4+}	NO_{2-}
35	Nitrification is the process by which:	$N_2$to NH_3	Nitrate to N_2gas	$NH_3$to nitrite/nitrate ✓	Organic $N$to NH_3
36	Which molecule shows cyclo-octa sulfur structure?	S_6	S_7	S_8✓	S_2
37	What is physical appearance of sulfur?	**Yellow solid ✓**	Colorless gas	Blue liquid	Red solid
38	What oxidation state does sulfur have in SO_2?	0	+2	+4✓	+6
39	Which sulfur state has six unpaired electrons?	Ground state	1st excited state	**2nd excited state ✓**	Ionized state
40	What forms when sulfur reacts with silver?	AgS	Ag_2S✓	$AgSO_4$	$AgSO_3$
41	What is sulphur hexafluoride used as?	Fuel	Explosive	**Insulator ✓**	Solvent
42	What does sulfur react with to form SF_4?	Cl_2	O_2	H_2	F_2✓
43	What is the product when sulfur reacts with cyanide?	SCN^-✓	KCN	S_2Cl_2	H_2SO_4
44	The most stable species in an acidic environment is:	SO_{42-}✓	SO_2	H_2S	S
45	What is the function of sulfur in vulcanisation?	Solvent	**Cross-linker ✓**	Catalyst	Fuel

46	Which fertilizer restores sulfur in soil?	Sodium Chloride	Calcium carbonate	Potassium nitrate	**Ammonium Sulfate ✓**
47	What are components of gun powder?	KNO_3, Charcoal, S✓	$NaNO_3$, Charcoal, S	KCl, Charcoal, S	$CaSO_4$, Charcoal, S
48	The colour of iodine in organic layer is:	Brown	Colourless	**Purple ✓**	Green
49	What colours are chlorine gases at room temperature?	Red and Blue	Dark brown	**Greenish yellow ✓**	Black
50	Who was the first halogen isolated as an element?	F	I✓	Cl	Br
51	What causes the colour change from chlorine to iodine?	Temperature	**Electron transitions ✓**	Pressure	Volume
52	Which colour is associated with Iodine?	Orange	Pale Green	**Purple ✓**	Yellow
53	Bromine emits_____ vapours.	Yellow	**Orange ✓**	Violet	Green
54	Which forces affect halogen volatility?	Ionic	Covalent	Metallic	**London dispersion ✓**
55	Why does volatility decrease down the group?	**Size increase ✓**	Pressure decrease	Temperature	Weight decrease
56	Why does bond strength decrease down the group?	**Atomic size ✓**	Electronegativity	Electron number	Pressure
57	Who has the weakest bond strength among halogens?	**Fluorine ✓**	Chlorine	Iodine	Bromine
58	What cause weak bond in Fluorine?	Large atoms	**Electron repulsion ✓**	Low electronegativity	High atomic mass
59	Which halogen molecule has the strongest bond?	F_2	Br_2	I_2	Cl_2✓
60	Which halogen has the highest bond energy?	F_2	Br_2	I_2	Cl_2✓
61	How does oxidizing power change down the group?	Increases	**Decreases ✓**	Remains same	Fluctuates
62	What is the order of oxidizing power of halogens?	$Cl_2 > F_2 > Br_2 > I_2$	$I_2 > Br_2 > Cl_2 > F_2$	$F_2 > Cl_2 > Br_2 > I_2$✓	$Br_2 > I_2 > F_2 > Cl_2$
63	Which halogen has the highest oxidizing power?	**Fluorine ✓**	Chlorine	Bromine	Iodine
64	Why is fluorine the most reactive halogen?	**Bond strength ✓**	Bond length	Electronegativity	Number of electrons
65	H_2SO_4added to solid $NaCl$initial observation:	Red fumes	Purple vapor	**White HClfumes ✓**	Black solid
66	Which halogen reacts explosively with hydrogen in dark?	**Fluorine ✓**	Chlorine	Bromine	Iodine
67	How does halogen reactivity change down the group?	Increases	**Decreases ✓**	Fluctuates	Remains same
68	Which halogen has the lowest oxidizing power?	F_2	Cl_2	Br_2	I_2✓
69	Which statement about halogen- H_2reaction is correct?	I_2reacts most	Cl_2explodes in dark	F_2combines in dark ✓	Br_2does not react

70	Bleaching powder is prepared by passing Cl_2over:	Calcium carbonate	Calcium sulphate	**Calcium hydroxide ✓**	Magnesium hydroxide
71	Which one is frequently used to disinfect water?	Sodium chloride	Hydrochloric acid	**Chlorine ✓**	Sodium hydroxide
72	Why is $HOCl$more effective than OCl^-?	More Stable	**More reactive ✓**	Less reactive	Less soluble
73	Which species must be enough for effective disinfection?	Cl_2	Cl^-	O_2	$HOCl$and OCl^-✓
74	What happens to $HOCl$in water?	Strong acid	Strong base	**Partial dissociation ✓**	Neutralization
75	Who are affected by burning fossil fuels?	Animals	Humans	Plants	**All above ✓**
76	What is the major component of the atmosphere?	**Nitrogen ✓**	Oxygen	Argon	Methane
77	What is one effect of human activities on the atmosphere?	Rain	Snow	Earthquake	**Deforestation ✓**
78	What happens to temperature in the mesosphere?	Increases	**Decreases ✓**	Fluctuates	Remains constant
79	Which layer is the coldest region of the atmosphere?	Troposphere	Stratosphere	**Mesosphere ✓**	Thermosphere
80	Which is a secondary pollutant?	**Carbonic acid ✓**	SO_2	NO_2	CO
81	Which gas absorbs UV radiation in the stratosphere?	O_2	CO_2	O_3✓	N_2
82	Which one is most toxic?	CO_2	CO✓	O_2	CH_4
83	Which natural source produces hydrogen sulphide?	**Volcanic eruptions ✓**	Thunderbolts	Oceans	Swamps
84	What human sources produce CO_2& CO?	Natural gas	**Vehicles & factories ✓**	Algae	Swamps
85	What layer extends from 85 km to 600 km?	Troposphere	Mesosphere	Stratosphere	**Thermosphere ✓**
86	What chemical reaction produces nitric oxide (NO)?	$N_2 + CO_2$	$N_2 + O_2$✓	$CO + O_2$	$SO_2 + H_2O$
87	Which compound causes ozone depletion?	CH_4	CO_2	$CFCs$✓	$VOCs$
88	What are POPs resistant to?	**Degradation ✓**	Water only	Temperature	Oxygen
89	Which POP is used in electrical equipment?	DDT	$VOCs$	$PCBs$✓	$PAHs$
90	Smog formed by sunlight, NO_x, and $VOCs$is a:	Primary pollutant	**Secondary pollutant ✓**	Natural gas	Harmless byproduct
91	What is the primary cause of climate change?	Deforestation	**Burning Fossil Fuels ✓**	Agriculture	Volcanic Eruptions
92	Which gases are included as greenhouse gases?	CO_2& CH_4✓	O_2& N_2	Ar& Ne	He& Kr
93	What form of deposition carries pollutants with rain?	Dry deposition	**Wet deposition ✓**	Gas deposition	Soil deposition

94	Who contributes to emissions of NO_2?	Power stations	Live stock	**Motor vehicles ✓**	Forests
95	What do forests absorb to help the atmosphere?	CH_4	N_2	**Greenhouse gases ✓**	CO
96	Which gas is not a pollutant?	SO_2	CO	NO_2	CO_2✓
97	Reducing smog contains high concentration of:	O_3	NO	SO_2✓	H_2O_2
98	Who affects pollutant dispersion?	Emission Sources	Natural events	Humans	**Meteorological conditions ✓**
99	What indicates 'Good' air quality in AQI table?	Yellow	Orange	**Green ✓**	Red
100	Where do microbial processes produce N_2O?	Atmosphere	Forests	**Soil & Oceans ✓**	Vehicles

SHORT Q.NO. 2

Q.NO. 1: Determine the number of protons, neutrons, and electrons in an atom with Z=17 and A=35.
Answer: For an atom with Atomic Number (Z) = 17 and Mass Number (A) = 35:

- **Number of Protons:** The atomic number (Z) represents the number of protons. Thus, Protons = 17.
- **Number of Electrons:** In a neutral atom, the number of electrons equals the number of protons. Thus, Electrons = 17.
- **Number of Neutrons:** The number of neutrons (N) is calculated by the formula $N = A - Z$. $N = 35 - 17 = 18$.

Q.NO. 2: Explain why the 4s subshell is filled before the 3d subshell in potassium.
Answer: According to the **Aufbau Principle** and the **(n + l) rule**, electrons fill orbitals with lower energy first.

- For the **4s** orbital: $n = 4, l = 0 \Rightarrow (n+l) = 4$.
- For the **3d** orbital: $n = 3, l = 2 \Rightarrow (n+l) = 5$. Since the 4s orbital has a lower $(n+l)$value than the 3d orbital, it has lower energy and is filled before the 3d subshell.

Q.NO. 3: What information about an electron can be obtained from the Principal and Azimuthal quantum numbers?
Answer:

- **Principal Quantum Number (n):** It describes the **size** and **energy** of the orbital. A larger ' n' indicates a greater distance from the nucleus and higher energy.
- **Azimuthal Quantum Number (l):** It describes the **shape** of the orbital (e.g., spherical for $l = 0$, dumbbell for $l = 1$). It also defines the subshells (s, p, d, f).

Q.NO. 4: Define the Pauli Exclusion Principle and give an example.
Answer: The **Pauli Exclusion Principle** states that no two electrons in the same atom can have the same set of four quantum numbers.

- **Example:** In a Helium atom ($1s^2$), both electrons share $n = 1, l = 0, m = 0$, but they must have opposite spins ($s = +\frac{1}{2}$and $s = -\frac{1}{2}$) to occupy the same orbital.

Q.NO. 5: What does the (n + l) rule represent in electronic configuration?
Answer: The **(n + l) rule** determines the relative energy of orbitals for electronic configuration.

1. Orbitals with a lower $(n+l)$value have lower energy and are filled first.
2. If two orbitals have the same $(n+l)$value, the one with the lower ' n' value has lower energy.

Q.NO. 6: Why are orbitals of the same subshell called degenerate?

Answer: Orbitals belonging to the same subshell that possess equal energy are called **degenerate orbitals**. For example, in the p-subshell ($l = 1$), the three orbitals p_x, p_y, and $p_z$have different spatial orientations but the same energy in the absence of a magnetic field.

Q.NO. 7: Differentiate between an orbit and an orbital.

Answer:

- **Orbit:** A definite circular path around the nucleus where an electron revolves (Bohr's concept). It is a 2-dimensional planar motion.
- **Orbital:** A 3-dimensional region in space around the nucleus where the probability of finding an electron is maximum (Wave mechanical concept).

Q.NO. 8: How many electrons can be accommodated in the M-shell? Explain using quantum numbers.

Answer: The maximum number of electrons in a shell is calculated by the formula $2n^2$. For the **M-shell**, the principal quantum number $n = 3$. Max electrons $= 2(3)^2 = 2 \times 9 = 18$. Using quantum numbers: The M-shell contains s, p, and d subshells ($2 + 6 + 10 = 18$electrons).

Q.NO. 9: Why does the f-subshell have 7 orbitals?

Answer: The number of orbitals in a subshell is given by $(2l+1)$. For the **f-subshell**, the azimuthal quantum number $l = 3$. Number of orbitals $= 2(3) + 1 = 7$. These correspond to the 7 magnetic quantum number (m) values: $-3, -2, -1, 0, +1, +2, +3$.

Q.NO. 10: Describe the significance of the magnetic quantum number.

Answer: The **Magnetic Quantum Number (m)** describes the **orientation** of an orbital in space under the influence of a magnetic field. It determines the number of degenerate orbitals in a subshell. For a given l, $m$ranges from $-l$to $+l$.

Q.NO. 11: Explain why the atomic number is more fundamental than the mass number.

Answer: Moseley (1913) discovered that the frequency of X-rays emitted by an element is directly proportional to its atomic number (Z), not its atomic mass. This proved that the properties of elements are periodic functions of their **atomic numbers** (proton numbers), making Za more fundamental property than mass number (A).

Q.NO. 12: Compare the mass of a neutron and a proton in amu.

Answer: Comparing the masses in atomic mass units (amu):

- **Proton:** 1.0073 amu.
- **Neutron:** 1.0087 amu. The neutron is slightly heavier than the proton.

Q.NO. 13: What happens to a neutron when passed through an electric field?

Answer: When a beam of **neutrons** is passed through an electric field, they are **not deflected**. They travel in a straight path because neutrons are neutral particles carrying **zero charge**.

Q.NO. 14: Why is the deflection of electrons greater than that of protons in an electric field?

Answer: The deflection of a particle in an electric field is proportional to its charge-to-mass ratio (e/m). While the magnitude of charge on an electron and proton is the same, the mass of an electron is approximately **1/1836** times the mass of a proton. Being lighter, electrons experience greater deflection.

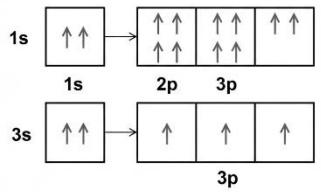
Q.NO. 15: Draw the orbital box diagram for phosphorus (Z=15) adhering to Hund's rule.

Answer:

Phosporus (Z=15)

Orbotal Box Diagram

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



Q.NO. 16: Write the electronic configuration of Sodium (Na).

Answer: Sodium (Na) has Atomic Number $Z = 11$. The electronic configuration is: $1s^2, 2s^2, 2p^6, 3s^1$. This can also be written as $[Ne]3s^1$.

Q.NO. 17: Define the azimuthal quantum number (l).

Answer: The **Azimuthal Quantum Number (l)** (also called Angular Momentum Quantum Number) describes the **shape** of the orbital and the subshell to which an electron belongs. Its values are integers ranging from 0to $(n-1)$.

Q.NO. 18: How many values of " l" are possible for n=3?

Answer: For the principal quantum number **n = 3**, the possible values of $l$are 0to $(n-1)$. Therefore, the values are **0, 1, and 2**. (These correspond to the 3s, 3p, and 3d subshells).

Q.NO. 19: What is the maximum number of orbitals in the p-subshell?

Answer: For the **p-subshell**, the azimuthal quantum number $l = 1$. The number of orbitals is determined by the magnetic quantum number m, which has values $-1, 0, +1$. Thus, the maximum number of orbitals is **3** (p_x, p_y, p_z).

Q.NO. 20: Distinguish between atomic emission and atomic absorption spectrum.

Answer:

- **Atomic Emission Spectrum:** Formed when elements release energy (e.g., when heated). It consists of bright colored lines on a dark background.
- **Atomic Absorption Spectrum:** Formed when white light passes through a cool gas. It consists of dark lines on a bright background where specific wavelengths are absorbed.

Q.NO. 21: What is 1st ionization energy? Give its standard unit.

Answer: First Ionization Energy (ΔH_{i1}) is the minimum energy required to remove the most loosely bound electron from each atom in one mole of gaseous atoms to form one mole of gaseous 1+ ions. **Standard Unit:** kJ mol^{-1} (kilojoules per mole).

Q.NO. 22: Explain why sulfur has a lower first ionization energy than phosphorus.

Answer: Phosphorus (P) ($3s^2 3p^3$) has a stable, **half-filled** p-subshell. **Sulfur (S)** ($3s^2 3p^4$) has one paired electron in the p-subshell, introducing **spin-pair repulsion**. This repulsion makes it easier to remove an electron from Sulfur than from the stable half-filled shell of Phosphorus.

Q.NO. 23: The ionization energy of Be is higher than that of B. Justify this exception.

Answer: Beryllium (Be) ($1s^2 2s^2$) has a stable, **fully-filled** 2s orbital. **Boron (B)** ($1s^2 2s^2 2p^1$) has a single electron in the 2p orbital which is higher in energy and further from the nucleus than the 2s. Therefore, it is easier to remove the 2p electron from B than the stable 2s electron from Be.

Q.NO. 24: Why does ionization energy decrease down a group?

Answer: Ionization energy **decreases** down a group because:

1. The atomic size increases (valence electrons are further from the nucleus).

2. The **shielding effect** of inner electrons increases. These factors weaken the nuclear attraction on valence electrons, making them easier to remove.

Q.NO. 25: Explain why nitrogen has a higher ionization energy than oxygen.

Answer: Nitrogen (N) ($2p^3$) has a stable **half-filled** p-subshell. **Oxygen (O)** ($2p^4$) contains a pair of electrons in one of its p-orbitals. The repulsion between these paired electrons (spin-pair repulsion) makes it easier to remove an electron from Oxygen, giving Nitrogen a higher ionization energy.

Q.NO. 26: Define electronegativity and name the scale used to measure it.

Answer: Electronegativity is the tendency or power of an atom to attract a shared pair of electrons toward itself in a covalent bond. The scale most commonly used to measure it is the **Pauling scale** (ranging from 0.7 to 4.0).

Q.NO. 27: How does atomic size affect electronegativity?

Answer: Atomic size is inversely proportional to electronegativity. As atomic size increases, the distance between the nucleus and the bonded electron pair increases, reducing the nuclear attraction. Therefore, electronegativity decreases as atomic size increases.

Q.NO. 28: Why do noble gases have positive 1st electron affinities?

Answer: Noble gases have **positive** 1st electron affinities (energy is absorbed) because they have stable, **fully-filled** valence shells (ns^2np^6). They have no tendency to accept an extra electron; any incoming electron must be forced into a higher energy level against repulsion.

Q.NO. 29: Explain why the 2nd electron affinity of oxygen is positive (+844 kJ/mol).

Answer: The 1st electron affinity of Oxygen is negative (exothermic). However, the **2nd electron affinity** is positive ($+844$kJ/mol) because the second electron is being added to an already negative ion (O^-). The electrostatic repulsion between the negative ion and the incoming electron requires energy to overcome.

Q.NO. 30: Why does fluorine have lower electron affinity than chlorine despite its smaller size?

Answer: Fluorine is very small, leading to a high electron density. The incoming electron faces significant **repulsion** from the existing electron cloud in the small 2p subshell. **Chlorine** is larger, reducing this repulsion, allowing it to accept an electron more easily (releasing more energy).

Q.NO. 31: What is the shielding effect and how does it influence atomic radius?

Answer: The **shielding effect** is the reduction in nuclear attraction on valence electrons due to the presence of inner shell electrons. As the shielding effect increases (more inner shells), the effective nuclear charge decreases, allowing the electron cloud to expand. This **increases** the atomic radius.

Q.NO. 32: Define ionic radius and compare the size of Na and Na^+.

Answer: Ionic radius is the measure of the size of an ion in a crystal lattice. The Na^+ ion is **smaller** than the neutral **Na** atom (95pm vs 186pm). This is because removing an electron reduces electron-electron repulsion and allows the nucleus to pull the remaining electrons closer.

Q.NO. 33: Why are anions always larger than their parent neutral atoms?

Answer: Anions are always **larger** than their parent neutral atoms because the addition of electron(s) increases electron-electron repulsion within the valence shell. The nuclear charge remains the same, but the electron cloud expands outward due to this increased repulsion.

Q.NO. 34: Explain why Mg^{2+} is smaller than Na^+.

Answer: Mg^{2+} and Na^+ are isoelectronic (both have 10 electrons), but Magnesium has a higher nuclear charge ($+12$) compared to Sodium ($+11$). The stronger nuclear pull in Magnesium draws the electrons closer, making Mg^{2+}(65pm) smaller than Na^+(95pm).

Q.NO. 35: What happens to atomic radius across a period from left to right?

Answer: Atomic radius **decreases** across a period from left to right. This is because the number of protons (nuclear charge) increases while electrons are added to the same shell. The increased nuclear attraction pulls the electron cloud closer to the nucleus, shrinking the atom.

Q.NO. 36: Identify semi-metals in groups 14, 15, and 16.

Answer: According to the provided text:

- **Group 14:** Silicon (Si), Germanium (Ge).

- **Group 15:** Arsenic (As), Antimony (Sb).
- **Group 16:** Tellurium (Te), Polonium (Po). These elements are situated along the "stair-step line".

Q.NO. 37: Illustrate how metallic character varies in group 14.
Answer: In **Group 14**, metallic character **increases** down the group.

- Carbon (top) is a non-metal.
- Silicon and Germanium are metalloids (semi-metals).
- Tin (Sn) and Lead (Pb) (bottom) are metals. This trend is due to increasing atomic size and shielding, which makes losing electrons easier.

Q.NO. 38: Why do non-metals have higher electronegativity than metals?
Answer: Non-metals have smaller atomic sizes and higher effective nuclear charges compared to metals in the same period. This allows their nuclei to attract shared electron pairs more strongly, resulting in **higher electronegativity** values compared to electropositive metals.
Q.NO. 39: Explain the effect of nuclear charge on ionic size. Answer: Nuclear charge is inversely related to ionic size for isoelectronic species. As nuclear charge increases, the attraction on the electron cloud becomes stronger, pulling electrons inward and **decreasing** the ionic size (e.g., $Al^{3+} < Mg^{2+} < Na^{+}$).
Q.NO. 40: Why is the 3rd ionization energy of magnesium much higher than its 2nd?
Answer: The **3rd ionization energy** of Magnesium involves removing an electron from the stable, fully-filled $2p^6$inner shell (Neon configuration) after the two valence electrons are lost. Breaking this stable noble gas configuration requires a massive amount of energy compared to the 2nd IE.

Part 3: Period 3 Elements (Oxides & Chlorides)
Q.NO. 41: Classify NaCl, $MgCl_2$, and PCl_5 as acidic, basic, or neutral.
Answer:

- **NaCl:** Neutral (pH = 7).
- **$MgCl_2$:** Neutral (or very slightly acidic due to weak hydrolysis, pH ~6.5).
- **PCl_5:** Acidic (reacts with water to form phosphoric and hydrochloric acid). The nature shifts from neutral to acidic as the bonding changes from ionic to covalent.

Q.NO. 42: Why are the oxides of Na and Mg more ionic than the oxides of N and P?
Answer: The difference in electronegativity between Oxygen and Group 1/2 metals (Na, Mg) is large (>1.7), resulting in **ionic** bonds (Na_2O, MgO). The difference between Oxygen and Non-metals (N, P) is small, resulting in **covalent** bonds.
Q.NO. 43: Describe the reaction of Na_2O with water and identify the product's nature.
Answer: Sodium oxide (Na_2O) reacts vigorously with water to form sodium hydroxide. **Reaction:** $Na_2O_{(s)} + H_2O_{(l)} \rightarrow 2NaOH_{(aq)}$. The product is a strong alkali, so the solution is **basic** (pH > 7).
Q.NO. 44: Why is $AlCl_3$ considered an acidic halide while NaCl is neutral?
Answer: NaCl dissolves to form hydrated ions without reacting chemically (Hydration), giving a neutral solution. $AlCl_3$ undergoes **hydrolysis** because the small, highly charged Al^{3+}ion polarizes water molecules, releasing H^+ions: $AlCl_{3(s)} + 3H_2O_{(l)} \rightarrow Al(OH)_{3(s)} + 3HCl_{(aq)}$.
Q.NO. 45: Predict whether PCl_5 and NCl_3 would be acidic or basic.
Answer:

- PCl_5: Acidic. It reacts with water to produce $H_3PO_4$and HCl.
- NCl_3: Acidic. Covalent chlorides of non-metals generally undergo hydrolysis to form acidic solutions.

Q.NO. 46: Why is SiO_2 considered an acidic oxide despite being a metalloid compound?
Answer: SiO_2 (Silicon dioxide) is considered an acidic oxide because, although it is insoluble in water, it reacts with hot alkalis (bases) like NaOH to form silicates. $SiO_{2(s)} + 2NaOH_{(aq)} \rightarrow Na_2SiO_{3(aq)} + H_2O_{(l)}$.
Q.NO. 47: Give the reaction of ZnO with both HCl and NaOH to prove its amphoteric nature. Answer: ZnO acts as both an acid and a base (Amphoteric):

- **With Acid:** $ZnO + 2HCl \rightarrow ZnCl_2 + H_2O$(Acts as base).

- **With Base:** $ZnO + 2NaOH \rightarrow Na_2ZnO_2 + H_2O$(Acts as acid, forming Sodium Zincate).

Q.NO. 48: What causes the difference in oxidation numbers between SO_2 and SO_3?

Answer:

- SO_2: Sulfur uses 4 valence electrons for bonding. Oxidation Number = **+4**.
- SO_3: Sulfur uses all 6 valence electrons for bonding (excitation into d-orbitals). Oxidation Number = **+6**.

Q.NO. 49: Why is sodium kept under kerosene oil while magnesium is not?

Answer: Sodium is highly reactive and reacts vigorously with oxygen and moisture in the open air to form oxides/peroxides. It is stored under kerosene to prevent this. **Magnesium** is less reactive due to a protective oxide layer and does not require such storage.

Q.NO. 50: Explain why magnesium reacts slowly with cold water but rapidly with steam. Answer: Magnesium reacts very slowly with cold water to form insoluble $Mg(OH)_2$which coats the surface. However, it reacts **vigorously with steam** to form Magnesium Oxide (MgO) and Hydrogen gas.

$$Mg_{(s)} + H_2O_{(g)} \rightarrow MgO_{(s)} + H_{2(g)}.$$

Q.NO. 51: What is hydration in the context of neutral chlorides like NaCl?

Answer: Hydration is the process where water molecules surround ions without breaking the water molecule itself. In neutral chlorides like NaCl, the ions (Na^+, Cl^-) become hydrated ($Na_{(aq)+}$) upon dissolving, leading to a neutral pH of 7.

Q.NO. 52: Compare the pH behavior of aqueous solutions of $MgCl_2$ and $AlCl_3$.

Answer:

- $MgCl_2$: Slightly acidic or near neutral (pH ≈ 6.5) as Mg^{2+}causes minimal hydrolysis.
- $AlCl_3$: Strongly acidic (pH < 7, typically around 3) due to the hydrolysis of the Al^{3+}ion releasing H^+.

Q.NO. 53: Why do sulfur and phosphorus show variable oxidation numbers?

Answer: Sulfur and Phosphorus show variable oxidation numbers (e.g., P: +3, +5; S: +4, +6) because they have empty **d-orbitals** in their valence shell. They can unpair electrons and promote them to these d-orbitals to form more bonds.

Q.NO. 54: Describe the nature of oxides formed by Period 3 non-metals.

Answer: Oxides of Period 3 non-metals (e.g., $SiO_2, P_2O_5, SO_2, SO_3, Cl_2O_7$) are **acidic** in nature. They react with water to form acids (e.g., $SO_3 \rightarrow H_2SO_4$) or react with bases to form salts.

Q.NO. 55: Write balanced equations for the reaction of Mg with oxygen.

Answer: Magnesium burns in oxygen with an intense white flame to produce white solid Magnesium Oxide.
Equation: $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$.

Q.NO. 56: Explain the term "hydrolysis" in the context of acidic chlorides.

Answer: Hydrolysis in the context of chlorides is the reaction of the cation with water to split the water molecule, typically generating H^+ions. Acidic chlorides (like $AlCl_3, PCl_5$) undergo hydrolysis to produce acidic solutions containing HClor other acids.

Q.NO. 57: Why is the oxidation number of 3rd-period elements in oxides always positive? Answer: The oxidation number of Period 3 elements in oxides is always **positive** because Oxygen is **more electronegative** than all the elements in Period 3 (Na to S, and even Cl in oxides is often treated with O as negative, though O is second only to F). The element bonded to O loses electron density.

Q.NO. 58: Compare the reactivity of Na and Mg toward chlorine.

Answer: Sodium reacts vigorously with chlorine with a golden yellow flame to form NaCl. **Magnesium** also reacts with chlorine to form $MgCl_2$. Sodium is more reactive (more electropositive) than Magnesium.

Q.NO. 59: What is the important function of the "stair-step line" on the periodic table?

Answer: The "stair-step line" on the periodic table separates **metals** (to the left) from **non-metals** (to the right). Elements adjacent to the line (Si, Ge, As, Sb, Te) are **metalloids** (semi-metals).

Q.NO. 60: How does electronic configuration help determine the block of an element?

Answer: The electronic configuration determines the block. If the last (valence) electron enters an **s-orbital**, the element is in the **s-block**. If it enters a **p-orbital**, it is in the **p-block**. For example, Na ($\ldots 3s^1$) is s-block; Al ($\ldots 3p^1$) is p-block.

Q.NO. 61: Define the rate of a reaction and give its units.

Answer: The **Rate of Reaction** is defined as the change in the concentration of a reactant or a product divided by the time taken for the change. **Formula:** Rate = $\Delta C/\Delta t$. **Units:** mol dm^{-3} s^{-1}.

Q.NO. 62: Differentiate between average rate and instantaneous rate.

Answer:

- **Average Rate:** The rate measured over a specific long time interval.
- **Instantaneous Rate:** The rate at a specific instant of time, calculated from the slope of the tangent to the concentration-time curve.

Q.NO. 63: Why does the instantaneous rate of a reaction decrease with time?

Answer: The instantaneous rate decreases with time because the **concentration of reactants decreases** as the reaction proceeds. According to the Law of Mass Action/Collision theory, a lower concentration leads to fewer effective collisions per unit time.

Q.NO. 64: Define "Activation Energy" and its role in a chemical reaction.

Answer: Activation Energy (E_a) is the minimum amount of energy required by reactant molecules to form an activated complex and undergo a chemical reaction. It acts as an energy barrier; only collisions with energy $\geq E_a$ are effective.

Q.NO. 65: Explain why wood burns more rapidly in pure oxygen than in air.

Answer: Wood burns more rapidly in pure oxygen than in air because the **concentration** of Oxygen in pure form (100%) is much higher than in air (~21%). Higher concentration leads to more frequent collisions between reactant molecules, increasing the reaction rate.

Q.NO. 66: How does temperature affect the collision frequency of molecules?

Answer: An increase in temperature increases the **kinetic energy** of the molecules. This causes them to move faster, leading to an increase in the **collision frequency** (number of collisions per second). More importantly, it increases the energy of the collisions.

Q.NO. 67: What is the Boltzmann distribution curve?

Answer: The **Boltzmann distribution curve** is a graph plotting the number of molecules (or fraction of molecules) against their kinetic energies. It shows that at a given temperature, a range of energies exists, and only a small fraction of molecules possess energy greater than the Activation Energy (E_a).

Q.NO. 68: Explain why a 10 °C rise in temperature approximately doubles the reaction rate. Answer: A 10 °C rise in temperature approximately doubles the reaction rate because it significantly increases the **fraction of molecules** possessing energy equal to or greater than the Activation Energy (E_a). This is represented by the doubling of the area under the Boltzmann curve beyond E_a.

Q.NO. 69: Define a catalyst and explain how it affects activation energy.

Answer: A **catalyst** is a substance that alters the rate of a reaction without being consumed. It works by providing an **alternative reaction pathway** with a **lower Activation Energy (E_a)**. This allows more molecules to collide successfully.

Q.NO. 70: Differentiate between homogeneous and heterogeneous catalysis.

Answer:

- **Homogeneous Catalysis:** The catalyst and reactants are in the **same phase** (e.g., both liquid or both gas). Example: Ester hydrolysis with acid.
- **Heterogeneous Catalysis:** The catalyst and reactants are in **different phases** (usually solid catalyst, gas/liquid reactants). Example: Haber process (Fe solid, N_2/H_2 gas).

Q.NO. 71: What is the "Order of Reaction"? How is it experimentally determined?

Answer: The **Order of Reaction** is the sum of the exponents (powers) of the concentration terms in the experimentally determined rate equation. Rate $= k[A]^x[B]^y$. Overall Order $= x + y$. It is determined experimentally by changing concentrations and observing the rate change.

Q.NO. 72: Define the specific rate constant (k).
Answer: The **Specific Rate Constant (k)** is the rate of the reaction when the concentration of each reactant is unity (1 mol dm^{-3}). It is a constant for a given reaction at a fixed temperature.
Q.NO. 73: Why is the sum of coefficients in a balanced equation not always equal to the order? Answer: The sum of coefficients in a balanced equation (stoichiometry) is not always equal to the order because the order depends on the **reaction mechanism** (specifically the rate-determining step), which is determined experimentally, not just by the overall equation.
Q.NO. 74: Give an example of a zero-order reaction.
Answer: A **Zero-Order Reaction** is one where the rate is independent of the concentration of reactants.
Example: Photochemical reactions or enzyme-catalyzed reactions at high substrate concentration are often cited (e.g., $Rate = k$).
Q.NO. 75: Define molecularity and distinguish it from the order of reaction.
Answer:

- **Molecularity:** The number of atoms, ions, or molecules colliding in a single elementary step. It is a theoretical whole number (1, 2, 3).
- **Order:** The sum of powers in the rate law. It is experimental and can be fractional or zero.

Q.NO. 76: What is the role of proper molecular orientation in collision theory?
Answer: For a collision to be effective, molecules must possess sufficient energy (E_a) and **Proper Orientation**. Correct orientation ensures that the atoms meant to bond come into contact directly; otherwise, the molecules may bounce off without reacting.
Q.NO. 77: How does the surface area of a solid reactant affect the rate of reaction?
Answer: For reactions involving solids, increasing the **surface area** (e.g., using powder instead of lumps) increases the rate of reaction. A larger surface area provides more sites for collisions with reactant molecules.
Q.NO. 78: Why is the rate of reaction highest at the beginning?
Answer: The rate of reaction is highest at the beginning because the **concentration of reactants** is at its maximum. As reactants are consumed, their concentration decreases, leading to fewer collisions and a slower rate.
Q.NO. 79: Give an example of a biological catalyst (enzyme).
Answer: A biological catalyst is called an **Enzyme**. Enzymes are protein molecules that catalyze biochemical reactions in living organisms (e.g., nature's catalysts) by lowering activation energy.
Q.NO. 80: What is a photochemical reaction? How does light affect its rate?
Answer: A **photochemical reaction** is a chemical reaction initiated by the absorption of light energy (photons). Light provides the necessary activation energy. The rate of such reactions often depends on the intensity of light rather than just concentration.
Q.NO. 81: Calculate the overall order for: rate = kNO^2 $NH_3{}^0$.
Answer: For the rate law: rate $= k[NO]^2[NH_3]^0$. The order with respect to NO is 2. The order with respect to $NH_3$is 0. **Overall Order** $= 2 + 0 = 2$(Second order).
Q.NO. 82: Why do chemists need to know the rate constant for a reaction?
Answer: Chemists need the **rate constant (k)** because it allows them to calculate the reaction rate for any given set of concentrations. It is also essential for comparing the speeds of different reactions and understanding the effect of temperature (via Arrhenius).
Q.NO. 83: How does the nature of reactants (ionic vs. covalent) affect the rate?
Answer: Ionic reactants typically react very fast (often instantaneously) because they involve the attraction of oppositely charged ions in solution with no bond breaking required. **Covalent reactants** react slower as bonds must be broken (requiring energy) before new ones form.
Q.NO. 84: Describe the effect of a catalyst on the "alternative path" of a reaction.
Answer: A catalyst provides an **alternative path** (mechanism) for the reaction. This new path has a lower energy barrier (Activation Energy) than the uncatalyzed path, allowing the reaction to proceed faster.
Q.NO. 85: What is meant by "effective collisions"?

Answer: Effective Collisions are collisions between reactant molecules that lead to the formation of products. They occur only when colliding molecules have:

1. Energy equal to or greater than the Activation Energy (E_a).
2. Proper molecular orientation.

Q.NO. 86: State the units of the rate constant for a first-order reaction.

Answer: For a first-order reaction, Rate $= k[Conc]^1$. Units of $k = Rate/Conc = (moldm^{-3}s^{-1})/(moldm^{-3}) = s^{-1}$(per second).

Q.NO. 87: How can color change be used to monitor the progress of a reaction?

Answer: Color change can be used if a reactant or product is colored. The intensity of the color (absorbance) is proportional to the concentration. By measuring absorbance over time using a spectrophotometer (colorimeter), the progress of the reaction can be monitored.

Q.NO. 88: Define "Fast Reactions" with an example.

Answer: Fast Reactions are those that occur almost instantaneously, often within fractions of a second. **Example:** The precipitation reaction between $AgNO_3$and $NaCl$(ionic reaction) or an explosion.

Q.NO. 89: Why is pressure important for the rate of gaseous reactions?

Answer: For gaseous reactions, **pressure** is directly related to concentration. Increasing the pressure pushes gas molecules closer together, effectively increasing their concentration. This increases the collision frequency and thus the rate of reaction.

Q.NO. 90: Explain why a catalyst is not consumed in a reaction.

Answer: A catalyst is not consumed because it is regenerated at the end of the reaction mechanism. It participates in intermediate steps (forming complexes) but is released unchanged chemically in the final step.

Q.NO. 91: Define a reversible reaction with a suitable example.

Answer: A **Reversible Reaction** is a reaction that proceeds in both the forward and reverse directions. It does not go to completion. **Example:** $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$.

Q.NO. 92: What are the macroscopic characteristics of chemical equilibrium?

Answer: Macroscopic characteristics of equilibrium:

1. The concentrations of reactants and products remain constant.
2. The system must be closed.
3. Forward and reverse rates are equal.
4. No observable change in physical properties (color, density).

Q.NO. 93: Why does a reversible reaction never go to completion?

Answer: A reversible reaction never goes to completion because as products are formed, they react to reform the reactants. Equilibrium is reached when the rate of formation of products equals the rate of reformation of reactants, leaving a mixture of both.

Q.NO. 94: Explain the term "Dynamic Equilibrium."

Answer: Dynamic Equilibrium is the state where the reaction continues to occur at the molecular (microscopic) level, but the rates of the forward and reverse reactions are equal. As a result, there is no net change in the macroscopic concentrations of substances.

Q.NO. 95: Define the Equilibrium Constant (K_c).

Answer: The **Equilibrium Constant (K_c)** is the ratio of the product of equilibrium concentrations of products to the product of equilibrium concentrations of reactants, with each term raised to the power of its coefficient in the balanced equation. It is constant at a fixed temperature.

Q.NO. 96: Write the expression for: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$.

Answer: For the reaction: $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$ the expression for $K_c$is: $K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$

Q.NO. 97: Why does a catalyst have no effect on the value of K_c?

Answer: A catalyst has **no effect** on the value of the Equilibrium Constant (K_c). It increases the rates of both the forward and reverse reactions equally, helping the system reach equilibrium faster, but it does not change the final composition of the equilibrium mixture.

Q.NO. 98: State Le-Chatelier's Principle.

Answer: Le-Chatelier's Principle states that if a system at equilibrium is subjected to a stress (change in concentration, pressure, or temperature), the system will shift its equilibrium position in a direction so as to oppose or minimize the effect of that stress.

Q.NO. 99: Explain the effect of increasing the concentration of reactants on equilibrium. Answer: Increasing the **concentration of reactants** shifts the equilibrium to the right (forward direction) to consume the added reactants and form more products. Increasing the concentration of products shifts the equilibrium to the left (reverse direction).

Q.NO. 100: How does a change in pressure affect equilibrium if the number of moles of gas is equal on both sides?

Answer: If the number of moles of gas is equal on both sides of the equation (e.g., $H_2 + I_2 \rightleftharpoons 2HI$, $\Delta n = 0$), a change in **pressure** (or volume) has **no effect** on the position of equilibrium.

Q.NO. 101: Why is temperature the only factor that changes the value of K_c?

Answer: The equilibrium constant (K_c) is a temperature-dependent constant derived from thermodynamics. While concentration, pressure, and volume changes shift the *position* of equilibrium to maintain the ratio defined by K_c, only a change in temperature changes the kinetic energy and stability of the reactants/products, altering the ratio itself, thus changing the value of K_c.

Q.NO. 102: In an exothermic reaction, in which direction does the equilibrium shift if heat is added?

Answer: In an **exothermic reaction** (where heat is a product), adding heat is equivalent to adding a product. According to **Le-Chatelier's Principle**, the system will try to consume the excess heat by shifting the equilibrium to the **left (reverse direction)**, favoring the formation of reactants.

Q.NO. 103: Describe the effect of pressure on the synthesis of Ammonia (Haber's Process). Answer: Reaction: $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$. Here, 4 moles of gaseous reactants produce 2 moles of gaseous products ($\Delta n = -2$). Increasing pressure shifts the equilibrium towards the side with fewer moles of gas to relieve the pressure. Therefore, **high pressure** favors the formation of Ammonia (yield increases).

Q.NO. 104: Why are pure solids and liquids excluded from the K_c expression?

Answer: The active mass (concentration) of a pure solid or pure liquid is constant and does not change significantly during a reaction (density remains constant). Therefore, they are incorporated into the equilibrium constant value and are not written explicitly in the K_c expression.

Q.NO. 105: Explain the soda water bottle equilibrium: $CO_2(g) + H_2O(l) \rightleftharpoons HCO_{3-}(aq) + H^+(aq)$.

Answer: In a sealed soda bottle, high pressure forces CO_2 gas to dissolve in water, establishing an equilibrium between gaseous CO_2, dissolved CO_2, and carbonic acid ions (HCO_{3-}, H^+). The high pressure shifts the equilibrium to the right, keeping the gas dissolved.

Q.NO. 106: What happens to the equilibrium in a soda bottle when the cap is removed?

Answer: When the cap is removed, the pressure of CO_2 above the liquid drops drastically. According to Le-Chatelier's Principle, the equilibrium shifts to the **left** to restore the pressure of CO_2 gas. Dissolved CO_2 escapes as bubbles, and the concentration of HCO_{3-} decreases.

Q.NO. 107: Differentiate between homogeneous and heterogeneous equilibria.

Answer:

- **Homogeneous Equilibrium:** All reactants and products are in the **same physical phase** (e.g., all gases or all liquid solutions). Example: $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$.
- **Heterogeneous Equilibrium:** Reactants and products are present in **more than one phase**. Example: $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$.

Q.NO. 108: What does a very large value of K_c indicate about the extent of a reaction?

Answer: A very large value of K_c (e.g., 10^{20}) indicates that at equilibrium, the concentration of products is much higher than that of reactants. This implies the reaction has proceeded almost to **completion**.

Q.NO. 109: How is the equilibrium constant related to the rates of forward and backward reactions?
Answer: At equilibrium, Rate(forward) = Rate(reverse). Since Rate(f) = $k_f[Reactants]$and Rate(r) = $k_r[Products]$, equating them gives: $k_f/k_r = [Products]/[Reactants]$. The ratio of the rate constants (k_f/k_r) is equal to the equilibrium constant (K_c).

Q.NO. 110: Explain why a closed system is necessary for establishing equilibrium. Answer: A closed system prevents matter from entering or escaping. If products (especially gases) escape, the reverse reaction cannot occur effectively, and equilibrium can never be established. A closed system ensures that all reactant and product molecules remain available to react in both directions.

Q.NO. 111: If $Q_c < K_c$, what does it tell you about the concentrations of reactants and products?
Answer: If the Reaction Quotient (Q_c) is less than K_c, it means the ratio of products to reactants is lower than the equilibrium ratio. To reach equilibrium, the system must form more products. Therefore, the reaction will proceed in the **forward direction** (net conversion of reactants to products).

Q.NO. 112: Define the "Position of Equilibrium."

Answer: The **Position of Equilibrium** refers to the relative concentrations of reactants and products at equilibrium. If products dominate, the position lies to the right; if reactants dominate, it lies to the left. Unlike K_c, the position can change with pressure or concentration without changing temperature.

Q.NO. 113: How does the removal of a product affect the yield in a reversible reaction?

Answer: Removing a product decreases its concentration. According to Le-Chatelier's Principle, the system opposes this change by shifting the equilibrium to the **right** (forward direction) to replace the removed product. This continuous removal drives the reaction forward, increasing the **yield**.

Q.NO. 114: Explain the effect of decreasing volume on the reaction: $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$.
Answer: Reaction: 1 mole gas ⇌2 moles gas. Decreasing volume increases pressure. The system will shift to the side with fewer moles of gas to reduce pressure. Here, it shifts to the **left** (reverse direction), favoring the formation of PCl_5.

Q.NO. 115: Why does equilibrium shift to the left in an exothermic reaction when heated?

Answer: In an exothermic reaction ($A \rightleftharpoons B + Heat$), heat is a product. Adding heat (raising temperature) increases the "concentration" of this product. The system counters this stress by consuming the heat via the reverse (endothermic) reaction, shifting equilibrium to the **left**.

Q.NO. 116: Discuss the industrial importance of Le-Chatelier's Principle.

Answer: Le-Chatelier's Principle is crucial in industry to maximize yields of economically important chemicals. For example, in the **Haber Process** (NH_3) and **Contact Process** (H_2SO_4), pressure and temperature conditions are optimized (e.g., high pressure, moderate temperature) to shift equilibrium towards the desired products.

Q.NO. 117: What is the response of an endothermic reaction to a decrease in temperature?

Answer: In an endothermic reaction ($A + Heat \rightleftharpoons B$), heat acts as a reactant. Decreasing the temperature removes heat. The system attempts to replace the lost heat by shifting in the exothermic direction (reverse direction), causing the equilibrium to shift to the **left**.

Q.NO. 118: In the reaction $BiCl_3 + H_2O \rightleftharpoons BiOCl + 2HCl$, what happens when HCl is added?

Answer: HCldissociates to provide $Cl^-$and H^+ions. Adding HClincreases the concentration of products (common ion effect). According to Le-Chatelier's Principle, the equilibrium shifts to the **left** (reverse direction), causing the white precipitate of $BiOCl$to dissolve and reform clear $BiCl_3$solution.

Q.NO. 119: Does a catalyst change the position of equilibrium? Explain.

Answer: No, a catalyst does not change the position of equilibrium or the yield. It lowers the activation energy for both forward and reverse reactions by the same amount, increasing their rates equally. It helps the system reach equilibrium **faster**, but the final ratio of concentrations (K_c) remains unchanged.

Q.NO. 120: Explain why $K_c$has no units in some reactions.

Answer: $K_c$has no units when the number of moles of gaseous reactants equals the number of moles of gaseous products ($\Delta n = 0$). In the $K_c$expression, the concentration units in the numerator cancel out exactly with those in the denominator. Example: $H_2 + I_2 \rightleftharpoons 2HI$. Units: $\frac{(moldm^{-3})^2}{(moldm^{-3})^2} = Nounits.$

Q.NO. 121: Why are halogens called "salt formers"?

Answer: The term "Halogen" is derived from Greek words *halos* (salt) and *gennan* (to form). They are called salt formers because they react directly with metals to form ionic salts (e.g., Chlorine reacts with Sodium to form NaCl, common salt).

Q.NO. 122: Explain the trend of volatility in halogens from fluorine to iodine.

Answer: Volatility decreases down the group from Fluorine to Iodine.

- Fluorine and Chlorine are gases (high volatility).
- Bromine is a volatile liquid.
- Iodine is a solid (low volatility). This is because intermolecular forces (London dispersion forces) increase with increasing atomic size and polarizability down the group.

Q.NO. 123: Why is fluorine the most powerful oxidizing agent?

Answer: Fluorine is the strongest oxidizing agent because it has the highest electronegativity, small size, and low bond dissociation energy of the F-F bond. These factors make it extremely eager to accept electrons and be reduced to F^-ions, thereby oxidizing other substances.

Q.NO. 124: Describe the role of London dispersion forces in the physical states of halogens.

Answer: Halogens are non-polar diatomic molecules held together by weak London dispersion forces. The strength of these forces depends on the number of electrons (polarizability). Down the group (F to I), the electron cloud size increases, making dispersion forces stronger. This changes the state from Gas (F_2, Cl_2) to Liquid (Br_2) to Solid (I_2).

Q.NO. 125: Why is the F-F bond weaker than the Cl-Cl bond?

Answer: The F-F bond is unexpectedly weak because the Fluorine atoms are very small. The non-bonding lone pairs of electrons on the two fluorine atoms are very close to each other, causing significant **inter-electronic repulsion**. This repulsion weakens the bond compared to the Cl-Cl bond where atoms are larger.

Q.NO. 126: Explain why HF is a weaker acid than HCl in aqueous solution.

Answer: Acid strength depends on bond strength. The **H-F bond** is very strong due to the small size of Fluorine and high polarity (hydrogen bonding). It does not dissociate easily to release H^+ions in water. The **H-Cl bond** is longer and weaker, allowing easier dissociation, making HCl a stronger acid.

Q.NO. 127: What is a disproportionation reaction? Give the reaction of chlorine with water.

Answer: A **Disproportionation Reaction** is a redox reaction where the same element is simultaneously oxidized and reduced. Reaction of Chlorine with water: $Cl_2 + H_2O \rightleftharpoons HCl + HClO$Here, Chlorine (0) is reduced to -1 in HCl and oxidized to +1 in HClO.

Q.NO. 128: Why does the reactivity of halogens with hydrogen decrease down the group?

Answer: Reactivity depends on the ability to attract electrons and bond strength. Fluorine reacts explosively (even in dark). Chlorine reacts in sunlight. Bromine requires heat. Iodine reacts slowly and reversibly. This trend mirrors the decrease in electronegativity and oxidizing power down the group.

Q.NO. 129: Name the halogen used as an antiseptic and explain how it works.

Answer: Iodine is used as an antiseptic (e.g., Tincture of Iodine). It works by oxidizing the proteins in microorganisms (bacteria/fungi), disrupting their cellular functions and killing them.

Q.NO. 130: Describe a chemical test to distinguish between Bromide and Iodide ions.

Answer: Add Chlorine water (Cl_2solution) and Carbon Tetrachloride (CCl_4) to the salt solution and shake.

- **Bromide (Br^-):** Turns the organic layer **orange/red-brown** (Br_2liberated).
- **Iodide (I^-):** Turns the organic layer **purple/violet** (I_2liberated).

Q.NO. 131: Why is HOCl a more effective disinfectant than OCl^-?

Answer: Hypochlorous acid ($HOCl$) is neutral and can easily penetrate the cell walls of bacteria (which are negatively charged) to kill them. The Hypochlorite ion (OCl^-) carries a negative charge and is repelled by the bacterial cell wall, making it a less effective disinfectant.

Q.NO. 132: Write the ionic equation for the displacement of Iodine by Bromine.

Answer: Bromine is a stronger oxidizing agent than Iodine. Equation: $Br_{2(aq)} + 2I_{(aq)-} \rightarrow 2Br_{(aq)-} + I_{2(s)}$. Bromine displaces Iodine from iodide salts.

Q.NO. 133: Which halogen is the most volatile? Explain based on intermolecular forces.

Answer: Fluorine (F_2) is the most volatile. It has the smallest molecular size and the fewest electrons, resulting in the weakest London dispersion forces between molecules. Weak forces mean very little energy is needed to separate molecules into the gas phase.

Q.NO. 134: What are the primary active species in the chlorination of water?

Answer: When Chlorine dissolves in water, it forms **Hypochlorous acid ($HOCl$)** and Hydrochloric acid (HCl). The primary active disinfectant species is **HOCl**, which kills pathogens. The OCl^- ion also contributes but is less effective.

Q.NO. 135: Why is chlorine used in the purification of drinking water?

Answer: Chlorine is a strong oxidizing agent that kills waterborne pathogens (bacteria, viruses) effectively. It is cheap, easy to apply, and leaves a "residual" effect that prevents re-contamination during distribution in pipes.

Q.NO. 136: Explain the color change when chlorine gas is passed through a KI solution.

Answer: Chlorine is a stronger oxidizing agent than Iodine. It displaces Iodine from Potassium Iodide (KI). $Cl_2 + 2KI \rightarrow 2KCl + I_2$. The solution changes from colorless to **brown** (due to dissolved Iodine) or forms a dark precipitate.

Q.NO. 137: How does the electronegativity of halogens vary down the group?

Answer: Electronegativity **decreases** down the group (F > Cl > Br > I). As atomic size increases, the valence shell moves further from the nucleus, and shielding increases. This reduces the nucleus's ability to attract the shared pair of electrons.

Q.NO. 138: Why can't iodine displace bromine from its salt?

Answer: Displacement relies on oxidizing power. A halogen can only displace a less reactive halogen below it in the group. Iodine is **less reactive** and a weaker oxidizing agent than Bromine (located above it), so it cannot take electrons from bromide ions (Br^-).

Q.NO. 139: Write the oxidation state of Chlorine in HClO and $HClO_3$.

Answer:

- **HClO:** $H(\ +1) + Cl(x) + O(\ -2) = 0 \Rightarrow x = +1.$
- **$HClO_3$:** $H(\ +1) + Cl(x) + 3O(\ -2) = 0 \Rightarrow 1 + x - 6 = 0 \Rightarrow x = +5.$

Q.NO. 140: What is the physical state of Iodine at room temperature? Why?

Answer: Iodine is a **solid** with a metallic luster. This is due to its large molecular size, which generates strong London dispersion forces between the I_2molecules, holding them in a rigid lattice structure at room temperature.

Q.NO. 141: Explain how bond length affects bond strength in halogens.

Answer: Generally, as bond length increases (due to larger atomic radii), bond strength decreases because the overlap of orbitals becomes less effective. However, the F-F bond is an exception (weaker than Cl-Cl) due to repulsion between lone pairs on the small F atoms. From Cl to I, bond strength decreases as length increases.

Q.NO. 142: What are hydrogen halides? Name the most stable one.

Answer: Hydrogen halides are binary compounds of hydrogen and halogens (HX). The most stable one is **Hydrogen Fluoride (HF)**. It has the highest bond dissociation energy due to the strong H-F bond. Thermal stability decreases from HF to HI.

Q.NO. 143: Write the conditions required for the reaction of Iodine with Hydrogen.

Answer: The reaction between Iodine and Hydrogen is slow and reversible. It requires:

1. **Heating** (high temperature).

2. Use of a **Catalyst** (typically Platinum, Pt). $H_2 + I_2 \rightleftharpoons 2HI$.

Q.NO. 144: Why is fluorine gas extremely reactive even in the dark?

Answer: Fluorine has a low bond dissociation energy (weak F-F bond) and high electronegativity. The low energy required to break the F-F bond allows it to initiate reactions spontaneously and explosively even without light energy (unlike Cl2 which needs UV light).

Q.NO. 145: Compare the oxidizing power of Cl_2 and Br_2.

Answer: Chlorine (Cl_2) is a stronger oxidizing agent than **Bromine (Br_2)**. Chlorine has a higher standard reduction potential (+1.36 V) compared to Bromine (+1.07 V), meaning Cl_2 can more easily accept electrons. Cl_2can displace $Br_2$from bromides, but not vice versa.

Q.NO. 146: How is HCl formed in the lab? Write the reaction.

Answer: HCl gas is prepared in the laboratory by heating Sodium Chloride ($NaCl$) with concentrated Sulfuric Acid (H_2SO_4). Reaction: $NaCl_{(s)} + H_2SO_{4(conc)} \rightarrow NaHSO_4 + HCl_{(g)}$.

(At higher temperatures, $2NaCl + H_2SO_4 \rightarrow Na_2SO_4 + 2HCl$).

Q.NO. 147: What is the environmental concern regarding the chlorination of water?

Answer: Chlorine can react with dissolved organic matter in water to form **Trihalomethanes (THMs)** like Chloroform ($CHCl_3$). THMs are considered carcinogenic (cancer-causing) if present in high concentrations over long periods.

Q.NO. 148: Explain the ionization of halogen acids in water.

Answer: Halogen acids (HX) ionize in water to produce Hydronium ions (H_3O^+) and Halide ions (X^-). $HX + H_2O \rightarrow H_3O^+ + X^-$. The extent of ionization determines strength: $HI > HBr > HCl \gg HF$(HF ionizes only partially).

Q.NO. 149: Why does iodine have a higher melting point than other halogens?

Answer: Melting point depends on the strength of intermolecular forces. Iodine has the largest number of electrons and largest surface area among halogens, leading to the strongest **London dispersion forces**. More energy is required to overcome these forces to melt the solid.

Q.NO. 150: Describe the reactivity of halogens with cold and hot NaOH.

Answer:

- **Cold Dilute NaOH:** $Cl_2 + 2NaOH \rightarrow NaCl + NaClO + H_2O$(Forms Chloride + Hypochlorite).
- **Hot Conc. NaOH:** $3Cl_2 + 6NaOH \rightarrow 5NaCl + NaClO_3 + 3H_2O$(Forms Chloride + Chlorate). Disproportionation occurs in both, but the oxidation state of halogen changes differently (+1 in cold vs +5 in hot).

Q.NO. 151: Why do some elements show variable oxidation numbers?

Answer: Elements (especially Transition metals and p-block non-metals) show variable oxidation numbers because the energy difference between their different subshells (e.g., ns and (n-1)d, or ns and np) is small. They can unpair electrons or use empty d-orbitals to form varying numbers of bonds.

Q.NO. 152: Define the Modern Periodic Law.

Answer: The **Modern Periodic Law** states that the physical and chemical properties of elements are a periodic function of their **atomic numbers**.

Q.NO. 153: What does the period number of an element indicate about its shells?

Answer: The **Period Number** indicates the principal quantum number (n) of the valence shell. It tells us the total number of electron shells occupied in the atom. For example, elements in Period 3 have electrons filling the 3rd shell ($n = 3$).

Q.NO. 154: Explain why lithium, sodium, and potassium have similar chemical properties.

Answer: These elements belong to **Group 1 (Alkali Metals)**. They all have the same valence electronic configuration (ns^1). Since chemical properties are determined primarily by the number of valence electrons, they exhibit similar reactivity (e.g., forming +1 ions, reacting with water).

Q.NO. 155: Why are noble gases placed at the end of each period?

Answer: Noble gases are placed at the end of each period (Group 18) because they possess a completely filled valence shell (ns^2np^6, except He). This stable configuration marks the completion of that period's electron filling before a new shell starts in the next period.

Q.NO. 156: How can ionization energy data predict the number of valence electrons?

Answer: A large "jump" or sudden increase in successive ionization energies indicates the removal of an electron from a stable inner shell. The number of ionization energies *before* this jump equals the number of valence electrons. (e.g., if IE3 >> IE2, there are 2 valence electrons).

Q.NO. 157: Differentiate between paired and unpaired electrons.

Answer:

- **Paired Electrons:** Two electrons occupying the same orbital with opposite spins (↑↓). They are stable and diamagnetic.
- **Unpaired Electrons:** A single electron in an orbital (↑). They are responsible for bonding and magnetic properties (paramagnetism).

Q.NO. 158: What is the significance of arrow direction in orbital diagrams?

Answer: The arrow direction represents the electron's **spin**. An up arrow (↑) typically represents spin $+\frac{1}{2}$, and a down arrow (↓) represents spin $-\frac{1}{2}$. Opposite directions in the same box signify that paired electrons must have opposite spins (Pauli Exclusion Principle).

Q.NO. 159: Why does electron pairing occur only after a subshell is half-filled?

Answer: According to **Hund's Rule**, electrons occupy degenerate orbitals singly first to minimize electron-electron repulsion. Pairing forces two electrons into the same region, increasing repulsion. Therefore, pairing requires more energy and only happens when no empty orbitals of the same energy are available.

Q.NO. 160: Explain the role of surface area in solid-gas reactions.

Answer: Increasing the surface area of a solid reactant (e.g., by grinding it into powder) exposes more particles to the gas phase reactant. This increases the frequency of collisions between solid and gas particles, thereby increasing the **rate of reaction**.

Q.NO. 161: What is an enzyme? Give an example of its catalytic action.

Answer: An **enzyme** is a biological catalyst, usually a protein, that speeds up specific biochemical reactions in living organisms. **Example: Urease** catalyzes the hydrolysis of urea into ammonia and carbon dioxide.

Q.NO. 162: How does pressure affect the synthesis of SO_3 in the Contact Process?

Answer: Reaction: $2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$. Since 3 moles of gas produce 2 moles ($\Delta n = -1$), increasing pressure shifts equilibrium to the right (forward). In the Contact Process, moderate pressure (1-2 atm) is sufficient to achieve high yield; extremely high pressure is not economically necessary.

Q.NO. 163: Define "Macroscopic Events" in a chemical reaction.

Answer: Macroscopic events are the large-scale, observable changes that occur during a reaction, such as changes in color, temperature, pressure, precipitate formation, or pH. These are the cumulative result of microscopic molecular collisions.

Q.NO. 164: Why is a large jump seen in the successive ionization energies of magnesium?

Answer: Magnesium ($3s^2$) has 2 valence electrons. The first two electrons are removed easily. The **3rd ionization energy** shows a massive jump because the 3rd electron must be removed from the stable, full 2p inner shell ($[Ne]$core), which requires significantly more energy.

Q.NO. 165: What is the relationship between group number and valence electrons?

Answer: For main group elements (s and p blocks), the **Group Number** (using the 1-8 system or A-groups) corresponds directly to the number of **valence electrons**. Group 1 = 1 e^-, Group 2 = 2 e^- ... Group 17 (7A) = 7 e^-.

Q.NO. 166: Explain the term "Effective Nuclear Charge" (Z_{eff}).

Answer: Effective Nuclear Charge (Z_{eff}) is the actual net positive charge experienced by a valence electron. It is less than the total nuclear charge (Z) because inner electrons shield the valence electrons. $Z_{eff} = Z - S$(where S is the shielding constant).

Q.NO. 167: Why do atoms become smaller from left to right in a period?
Answer: Across a period, the number of protons (Z) increases, increasing the nuclear pull. Electrons are added to the *same* shell, so shielding remains roughly constant. The increased Z_{eff}pulls the electron shell closer to the nucleus, reducing atomic radius.
Q.NO. 168: Define the "Spin-Pair Repulsion" effect.
Answer: Spin-Pair Repulsion is the electrostatic repulsion between two electrons sharing the same orbital. This repulsion makes the paired electrons slightly less stable and easier to remove (lower ionization energy) compared to a single electron in a similar orbital (e.g., O vs N).
Q.NO. 169: Why is the second electron affinity of oxygen positive? Answer: The first electron adds to a neutral O atom (exothermic). The second electron must be added to a negative ion (O^-). The strong electrostatic repulsion between the negative ion and the incoming negative electron requires energy input to overcome, making the process endothermic (+).
Q.NO. 170: How does the shielding effect influence the electronegativity of elements in a group?
Answer: Down a group, the number of inner electron shells increases, increasing the **shielding effect**. This screens the nucleus from the valence electrons. Reduced effective nuclear pull on the shared pair of electrons causes **electronegativity to decrease** down the group.
Q.NO. 171: Why are elements on the left side of the periodic table considered metals?
Answer: Elements on the left (Groups 1, 2) have low ionization energies, low electronegativity, and 1-2 valence electrons. They tend to **lose electrons** easily to form positive ions (cations), which is the defining chemical characteristic of metals.
Q.NO. 172: Compare the reactivity of Na and Mg with oxygen.
Answer:

- **Na:** Reacts vigorously at room temperature to form Sodium Oxide (Na_2O) and Peroxide (Na_2O_2).
- **Mg:** Reacts less readily at room temperature due to an oxide film; requires heating (burning) to react brilliantly and form Magnesium Oxide (MgO). Na is more reactive.

Q.NO. 173: Why does sodium form both an oxide and a peroxide?
Answer: Sodium is highly reactive. When burned in limited oxygen, it forms oxide (Na_2O). In excess oxygen, the vigorous reaction and high electron density allow it to form the Peroxide (Na_2O_2), where oxygen exists as the O_{22}_ion.
Q.NO. 174: Write the formula of the chloride of a group 15 element.
Answer: Group 15 elements (like Phosphorus, Nitrogen) have 5 valence electrons. They typically form trichlorides (MCl_3) using 3 electrons or pentachlorides (MCl_5, for P, As, Sb) using all 5. Example: PCl_3 or PCl_5.
Q.NO. 175: Differentiate between acidic and basic oxides with reactions.
Answer:

- **Acidic Oxide:** Non-metal oxide that reacts with bases. E.g., $SO_3 + 2NaOH \rightarrow Na_2SO_4 + H_2O$.
- **Basic Oxide:** Metal oxide that reacts with acids. E.g., $MgO + 2HCl \rightarrow MgCl_2 + H_2O$.

Q.NO. 176: Explain why $AlCl_3$ is acidic while NaCl is neutral.
Answer: $NaCl$contains distinct Na^+ions which do not hydrolyze significantly. $AlCl_3$contains small, highly charged Al^{3+}ions. In water, Al^{3+}polarizes the water molecules in its hydration shell, weakening O-H bonds and releasing H^+ions (Hydrolysis), making the solution acidic.
Q.NO. 177: How is the oxidation number of a third-period element generally related to its group?
Answer: The maximum positive oxidation number often equals the **Group Number** (valence electrons).

- Na (Group 1): +1
- Mg (Group 2): +2
- Al (Group 3): +3
- Si (Group 4): +4
- P (Group 5): +5 ... and so on.

Q.NO. 178: Why is the atomic number more fundamental than the mass number?

Answer: Atomic number (Z) determines the electronic configuration (number of electrons), which dictates chemical properties. Mass number (A) varies with isotopes (same element, different neutrons) but isotopes share the same chemical behavior. Thus, Zdefines the element's identity.

Q.NO. 179: Find the number of neutrons in $^{31}_{15}$P.

Answer: For Phosphorus-31 ($_{1531}P$): Mass Number (A) = 31 Atomic Number (Z) = 15 Neutrons = $A - Z = 31 - 15 = 16$.

Q.NO. 180: Explain why the Cl^- ion has more electrons than protons.

Answer: A neutral Chlorine atom has 17 protons and 17 electrons. The Chloride ion (Cl^-) is formed by **gaining one electron**. Total Electrons = $17 + 1 = 18$. Total Protons = 17. Since $18 > 17$, it has a net negative charge.

Q.NO. 181: Define the term "Orbit."

Answer: An **Orbit** (based on the Bohr model) is a well-defined circular path around the nucleus in which an electron revolves. It is associated with a specific fixed energy level.

Q.NO. 182: How do the values of "l" determine the type of subshell (s, p, d, f)?

Answer: The azimuthal quantum number (l) corresponds to specific subshells:

- $l = 0 \rightarrow s$-subshell
- $l = 1 \rightarrow p$-subshell
- $l = 2 \rightarrow d$-subshell
- $l = 3 \rightarrow f$-subshell

Q.NO. 183: Why are px, py, and pz orbitals called degenerate?

Answer: They are called degenerate because, in the absence of an external magnetic or electric field, they all possess the **exact same energy**. They only differ in their directional orientation along the x, y, and z axes.

Q.NO. 184: State Hund's Rule of maximum multiplicity.

Answer: Hund's Rule states that when electrons occupy degenerate orbitals (orbitals of equal energy), they fill them singly with parallel spins before pairing up. This minimizes repulsion and maximizes stability.

Q.NO. 185: Explain the importance of the rate law in chemistry.

Answer: The **Rate Law** relates the rate of reaction to the concentration of reactants experimentally. It helps determine the reaction order and provides insight into the **reaction mechanism**, specifically identifying the rate-determining step.

Q.NO. 186: Why is molecularity always a whole number?

Answer: Molecularity represents the number of distinct particles (molecules/atoms) colliding in a single step. Since you cannot have a fraction of a molecule colliding, molecularity must be a positive integer (1, 2, or rarely 3).

Q.NO. 187: How does the removal of a product affect the direction of a reversible reaction?

Answer: Removing a product lowers its concentration ($Q_c < K_c$). According to **Le-Chatelier's Principle**, the system tries to restore the product by shifting the equilibrium in the **forward direction**, favoring the conversion of reactants into products.

Q.NO. 188: What is the response of an equilibrium system to a decrease in temperature in an exothermic reaction?

Answer: In an exothermic reaction, heat is released (Product). Lowering the temperature removes heat. The system shifts to the **right** (forward direction) to produce more heat, thereby increasing the yield of products and the value of K_c.

Q.NO. 189: Why is $K_c$constant at a given temperature?

Answer: $K_c$is a ratio of rate constants (k_f/k_r). Since rate constants (k) change only with temperature (Arrhenius equation), their ratio (K_c) also remains constant unless the temperature changes. Concentration or pressure changes affect equilibrium position, not the constant itself.

Q.NO. 190: Describe the effect of volume increase on the reaction: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$.

Answer: Increasing volume decreases pressure. The system shifts to the side with **more moles of gas** to

restore pressure. Reactants: 4 moles. Products: 2 moles. Equilibrium shifts to the **left** (reverse direction), favoring decomposition of Ammonia.

Q.NO. 191: Why does a catalyst speed up both forward and reverse reactions equally?

Answer: A catalyst lowers the energy barrier (activation energy) for the transition state. Since the transition state is the peak for both forward and reverse paths, lowering it facilitates both directions equally. Thus, equilibrium is reached faster without changing the equilibrium position.

Q.NO. 192: Compare the bond strengths of Cl-Cl and Br-Br.

Answer: Cl-Cl bond is stronger (242 kJ/mol) than **Br-Br bond** (193 kJ/mol). Chlorine atoms are smaller, allowing for better orbital overlap and a shorter bond length compared to the larger Bromine atoms.

Q.NO. 193: What is observed when chlorine reacts with KI solution?

Answer: Chlorine displaces Iodine from Potassium Iodide ($2KI + Cl_2 \rightarrow 2KCl + I_2$). **Observation:** The solution turns from colorless to **brown** (due to I_2 formation), and if concentrated, black iodine crystals may precipitate.

Q.NO. 194: Explain why HF is more stable than other hydrogen halides.

Answer: HF is the most stable because the H-F bond has the highest **bond dissociation energy** (567 kJ/mol). The small size of Fluorine and high polarity create a very strong bond that requires very high temperatures to break.

Q.NO. 195: How does HClO act on microorganisms to kill them?

Answer: Hypochlorous acid ($HClO$) is a neutral molecule that easily diffuses through the cell membrane of microorganisms. Once inside, it oxidizes essential enzymes and proteins, disrupting metabolic processes and causing cell death.

Q.NO. 196: Write the electronic configuration of an element in Period 4, Group 2.

Answer: Period 4, Group 2 element is Calcium (Ca, Z=20). Configuration: $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$ or $[Ar]4s^2$.

Q.NO. 197: Why does metallic character increase down a group?

Answer: Metallic character is the tendency to lose electrons. Down a group, atomic size increases and ionization energy decreases. It becomes easier to lose valence electrons, so elements become more metallic (e.g., in Group 14, C is non-metal, Pb is metal).

Q.NO. 198: Explain why nitrogen has a higher ionization energy than oxygen.

Answer: Nitrogen ($2p^3$) has a stable half-filled p-subshell. Oxygen ($2p^4$) has one paired electron in the p-subshell. The repulsion between the paired electrons makes it easier to remove one electron from Oxygen than from the stable configuration of Nitrogen.

Q.NO. 199: What is the significance of the "rate-determining step" in a reaction?

Answer: The **rate-determining step** is the slowest step in a multi-step reaction mechanism. It acts as a bottleneck, controlling the overall speed of the reaction. The rate law for the overall reaction is derived from this step.

Q.NO. 200: Define "Amphoteric Oxides" with the example of Aluminum Oxide (Al_2O_3).

Answer: Amphoteric Oxides are oxides that can behave as both acids and bases. **Example (Al_2O_3):**

- With Acid (HCl): Forms salt $AlCl_3$(acts as Base).
- With Base ($NaOH$): Forms aluminate $NaAlO_2$(acts as Acid).

SHORT QUESTIONS: 3

Q.NO. 1: By counting electron pairs around the central atom, explain why xenon trioxide (XeO3) has a pyramidal shape.

Answer: According to the VSEPR theory, the shape of a molecule depends on the number of electron pairs (bonding and lone pairs) around the central atom. Xenon (Xe) is a noble gas with 8 valence electrons. In Xenon trioxide (XeO_3), Xenon forms double bonds with three oxygen atoms. While the detailed counting of

electrons in XeO_3involves considering the valence electrons of Xe (8) and the contribution from Oxygen, the geometry is determined by the electron pair repulsion. With three bonding domains (double bonds behave as single domains for geometry) and one lone pair of electrons on the central Xenon atom, the total electron pairs are arranged tetrahedrally. However, due to the presence of one lone pair, the molecular shape becomes trigonal pyramidal to minimize repulsion, similar to ammonia (NH_3).

Q.NO. 2: Explain the difference between the formation of $\sigma$and πbonds.

Answer: A sigma (σ) bond is formed by the **head-on** (or end-to-end) overlap of atomic orbitals. This overlap can occur between two s-orbitals (s-s overlap), an s and a p orbital (s-p overlap), or two p-orbitals lying on the same axis (p-p head-on overlap). The electron density in a σbond is concentrated symmetrically along the line joining the two nuclei. A pi (π) bond is formed by the **sideways** (or parallel) overlap of two p-orbitals ($p_y$or p_z) that are perpendicular to the nuclear axis. In a πbond, the electron density lies above and below the line joining the nuclei (the nodal plane), making it weaker than a σbond.

Q.NO. 3: Show how the central carbon atom in propanone forms $\sigma$and πbonds through hybridization.

Answer: In propanone (CH_3COCH_3), the central carbon atom is bonded to an oxygen atom by a double bond and to two methyl carbons by single bonds. This central carbon atom is sp^2hybridized. It uses its three sp^2hybrid orbitals to form three σbonds: one with the oxygen atom and one with each of the two adjacent carbon atoms. These bonds lie in a trigonal planar arrangement with bond angles of approximately 120°. The remaining unhybridized p-orbital on the central carbon overlaps sideways with the p-orbital of the oxygen atom to form a πbond, completing the carbonyl ($C = O$) double bond.

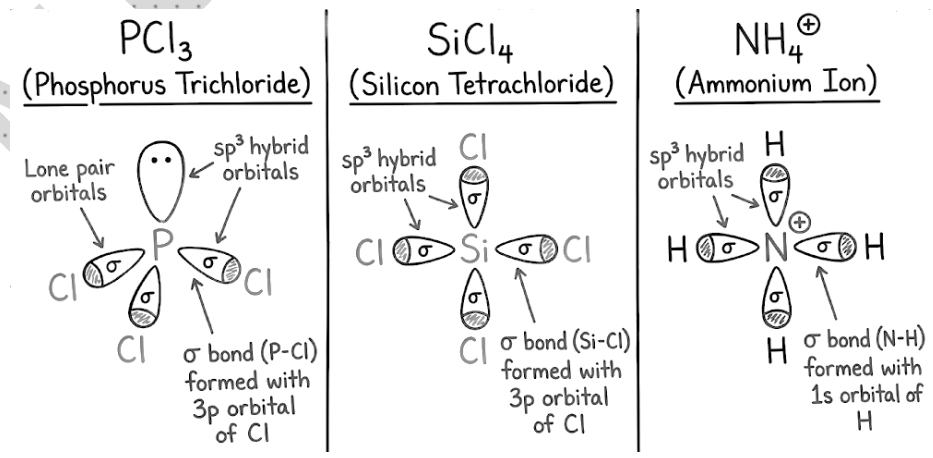
Q.NO. 4: Predict the shapes of sulfate (SO_{42-}), borate (BH_{4-}), and tri-iodide (I_{3-}) ions according to VSEPR.

Answer:

- **Sulfate (SO_{42-}):** The central sulfur atom has 6 valence electrons, and with the negative charge and bonds to 4 oxygens, it has 4 regions of electron density (4 bonding pairs, 0 lone pairs). This results in a **tetrahedral** shape.
- **Borate (BH_{4-}):** Boron (Group 13) has 3 valence electrons. Adding one electron for the negative charge gives 4 valence electrons, which form 4 single bonds with Hydrogen. There are 4 bond pairs and 0 lone pairs, resulting in a **tetrahedral** shape.
- **Tri-iodide (I_{3-}):** The central iodine atom has 7 valence electrons plus 1 from the charge, totaling 8. It forms 2 bonds with the other iodine atoms, leaving 3 lone pairs. The 5 electron pairs (2 BP, 3 LP) are arranged in a trigonal bipyramidal geometry, but the atoms form a **linear** shape to minimize lone pair repulsion.

Q.NO. 5: Sketch the hybrid orbitals and bond formation in PCl_3, $SiCl_4$, and NH_{4+}.

Answer:



Q.NO. 6: Draw the orbital structures of the CO_2molecule in terms of Valence Bond Theory (VBT).

Answer: In carbon dioxide (CO_2), the central carbon atom is bonded to two oxygen atoms via double bonds ($O = C = O$). The carbon atom undergoes **sp hybridization**. Its two sp hybrid orbitals form two σbonds with

the p-orbitals (or sp orbitals) of the two oxygen atoms in a linear geometry (180°). The two unhybridized p-orbitals on the carbon atom overlap sideways with the parallel p-orbitals of the oxygen atoms to form two separate πbonds. This results in a linear non-polar molecule.

Q.NO. 7: Can you explain why CO has a dipole moment but CO_2does not?

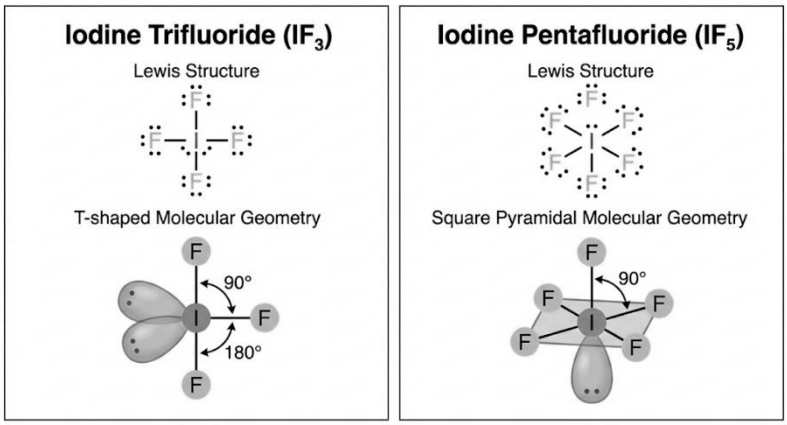
Answer: Carbon monoxide (CO) is a diatomic molecule with a polar bond due to the electronegativity difference between Carbon and Oxygen, resulting in a permanent dipole moment. Carbon dioxide (CO_2), however, is a linear molecule ($O = C = O$). Although the individual $C = O$bonds are polar due to the electronegativity difference, the linear geometry causes the two bond dipoles to point in opposite directions. These dipoles cancel each other out effectively, resulting in a net dipole moment of zero for the CO_2molecule.

Q.NO. 8: Do you think that individual bonds in $CCl_4$are polar? Explain in terms of electronegativity.

Answer: Yes, the individual bonds in carbon tetrachloride (CCl_4) are polar. Chlorine is more electronegative than Carbon, which creates an electronegativity difference between the bonded atoms. This difference causes the shared electron pair to be attracted more strongly towards the Chlorine atom, creating a partial negative charge ($\delta -$) on the Chlorine and a partial positive charge ($\delta +$) on the Carbon. Although the bonds are polar, the molecule itself is non-polar because the four bonds are arranged symmetrically in a tetrahedron, cancelling the dipole moments.

Q.NO. 9: Draw the Lewis structures for $IF_3$and $IF_5$and predict their geometry.

Answer:



Q.NO. 10: HI is a strong acid and a robust reducing agent, whereas HF is a weaker acid. Explain.

Answer: HI is a strong acid because the bond length between Hydrogen and Iodine is large and the bond energy is low, allowing it to dissociate easily into $H^+$and I^-. It is a robust reducing agent because the Iodide ion (I^-) readily loses electrons to become Iodine (I_2). HF is a weaker acid primarily due to the very strong bond between Hydrogen and Fluorine and the extensive **hydrogen bonding** between HF molecules. The hydrogen bonding entraps the H^+between electronegative Fluorine atoms, making dissociation more difficult compared to HI.

Q.NO. 11: Differentiate between s-s overlap and s-p overlap.

Answer:

- **s-s overlap:** This occurs when the s-orbital of one atom overlaps with the s-orbital of another atom. This overlap is always head-on and forms a sigma (σ) bond. An example is the bond formation in the H_2molecule.
- **s-p overlap:** This occurs when the s-orbital of one atom overlaps with the p-orbital of another atom. This overlap also takes place along the internuclear axis (head-on) and results in the formation of a sigma (σ) bond. An example is the bond in HF.

Q.NO. 12: Why are covalent bonds directional in nature?

Answer: Covalent bonds are formed by the overlap of atomic orbitals (like p, d, or hybrid orbitals) which have specific shapes and orientations in space (except for spherical s-orbitals). To form a strong bond, these orbitals must overlap effectively, which requires them to be oriented towards each other along specific axes.

This requirement for specific spatial orientation leads to fixed bond angles and definite shapes for covalent molecules, making the bonds directional.

Q.NO. 13: Explain the bonding in Cl_2 molecule using orbital overlap.

Answer: In the chlorine molecule (Cl_2), each chlorine atom has an electronic configuration ending in $3p^5$, with one half-filled 3p orbital. The bonding occurs via the **head-on overlap** of these half-filled 3p orbitals from the two chlorine atoms along the internuclear axis. This overlap results in the formation of a sigma (σ) bond ($3p - 3p\sigma$bond).

Q.NO. 14: How does the πbond form in O_2?

Answer: In the oxygen molecule (O_2), a double bond exists between the two oxygen atoms. One of these bonds is a sigma (σ) bond formed by the head-on overlap of p-orbitals. The second bond is a **pi (π) bond**, which is formed by the **sideways (parallel) overlap** of the unhybridized p-orbitals (e.g., $p_z$orbitals) that are perpendicular to the internuclear axis. This sideways overlap creates electron density above and below the nodal plane of the nuclei.

Q.NO. 15: Why does H_2S have a bent shape instead of linear?

Answer: Sulfur in H_2Shas 6 valence electrons. It forms two single bonds with Hydrogen atoms and has two lone pairs of electrons remaining. According to VSEPR theory, these four electron pairs (2 bonding, 2 lone pairs) arrange themselves tetrahedrally to minimize repulsion. However, the repulsion from the two lone pairs compresses the bond angle to be less than the ideal tetrahedral angle ($109.5°$), resulting in a **bent** or angular V-shape, similar to water (H_2O).

Q.NO. 16: Define sp^3 hybridization and give the example of methane (CH_4).

Answer: sp^3hybridization is the process of mixing one s-orbital and three p-orbitals of an atom to form four identical hybrid orbitals of equivalent energy and shape. These orbitals are directed towards the corners of a regular tetrahedron with bond angles of $109.5°$. In **methane (CH_4)**, the carbon atom undergoes sp^3hybridization. The four sp^3hybrid orbitals overlap with the 1s orbitals of four hydrogen atoms to form four sigma bonds, giving the molecule a tetrahedral geometry.

Q.NO. 17: Differentiate between sp and sp^3 hybridization in terms of bond angle and geometry.

Answer:

- **sp hybridization:** Involves the mixing of one s and one p orbital. It results in a **linear** geometry with a bond angle of $180°$. An example is $BeCl_2$.
- sp^3hybridization: Involves the mixing of one s and three p orbitals. It results in a **tetrahedral** geometry with a bond angle of $109.5°$. An example is CH_4.

Q.NO. 18: Why is the energy of hybrid orbitals lower than that of unhybridized orbitals?

Answer: Hybrid orbitals are formed to maximize bonding overlap. When hybrid orbitals form bonds (like sigma bonds), the resulting bonding molecular orbitals have lower energy than the original atomic orbitals, leading to a more stable system. Specifically, the hybridization process itself redistributes energy, and the resulting hybrid orbitals (like sp^3) are often lower in energy than the pure p-orbitals involved, though higher than the s-orbital. The primary stability comes from the stronger bonds they form due to better overlap.

Q.NO. 19: What type of hybridization is present in $BeCl_2$?

Answer: In Beryllium chloride ($BeCl_2$), the central Beryllium atom undergoes **sp hybridization**. Beryllium has two valence electrons which form bonds with two Chlorine atoms. The two sp hybrid orbitals are arranged linearly to minimize repulsion, resulting in a bond angle of $180°$and a linear shape.

Q.NO. 20: Why does BF_3 have a trigonal planar shape?

Answer: In Boron trifluoride (BF_3), the central Boron atom has 3 valence electrons and forms 3 single bonds with Fluorine atoms. There are no lone pairs on the Boron atom. According to VSEPR theory, the three bonding electron pairs arrange themselves as far apart as possible to minimize repulsion. This arrangement corresponds to a **trigonal planar** geometry with bond angles of $120°$.

Q.NO. 21: Define Bond Order.

Answer: Bond order is defined as half of the difference between the number of electrons in bonding molecular orbitals (N_b) and the number of electrons in anti-bonding molecular orbitals (N_a). The formula is given by:

$$BondOrder = \frac{N_b - N_a}{2}$$

A positive bond order indicates a stable bond, while a bond order of zero implies the molecule does not form.

Q.NO. 22: Explain why the enthalpy of hydration is always an exothermic process for gaseous ions.

Answer: The enthalpy of hydration is the energy change when one mole of gaseous ions dissolves in water to form an infinitely dilute solution. This process is always **exothermic** (negative ΔH) because it involves the formation of attractive forces between the charged ions and the polar water molecules (ion-dipole interactions). The formation of these bonds releases energy, whereas no bonds are broken in the gaseous ions themselves.

Q.NO. 23: Define standard enthalpy of atomization with an example.

Answer: The standard enthalpy of atomization ($\Delta H_{at\circ}$) is the enthalpy change required to produce one mole of gaseous atoms from the element in its standard state under standard conditions (298 K and 1 atm). For example, the atomization of hydrogen is represented as:

$$\frac{1}{2}H_2(g) \rightarrow H(g)\Delta H_{at\circ} = +218kJmol^{-1}$$

Q.NO. 24: The enthalpy of solution can be positive or negative. Explain what this indicates.

Answer: The enthalpy of solution (ΔH_{sol}) is the net result of two energy terms: the lattice energy (energy required to break the crystal lattice, endothermic) and the hydration energy (energy released when ions interact with water, exothermic).

- If ΔH_{sol} is **negative** (exothermic), it indicates that the hydration energy released is greater than the lattice energy absorbed.
- If ΔH_{sol} is **positive** (endothermic), it indicates that the lattice energy required to break the solid is greater than the hydration energy released.

Q.NO. 25: Calculate ΔH° for the formation of methane using enthalpies of combustion of C, H_2, and CH_4.

Answer: To calculate the enthalpy of formation of methane (CH_4), we use Hess's Law and the combustion data. Reaction: $C(s) + 2H_2(g) \rightarrow CH_4(g)$Given:

$$\Delta H_c(C) = -393.5kJ/mol(\text{Combustion of C})$$
$$\Delta H_c(H_2) = -285.8kJ/mol(\text{Combustion of } H_2)$$
$$\Delta H_c(CH_4) = -890.8kJ/mol(\text{Combustion of } CH_4)$$
$$\textbf{Formula: } \boldsymbol{\Delta H_f = \Sigma \Delta H_c(reactants) - \Sigma \Delta H_c(products)}$$
$$\Delta H_f = [\Delta H_c(C) + 2 \times \Delta H_c(H_2)] - [\Delta H_c(CH_4)]$$
$$\Delta H_f = [-393.5 + 2(\ \ -285.8)] - [\ \ -890.8]$$
$$\Delta H_f = [\ \ -393.5 - 571.6] + 890.8$$
$$= -965.1 + 890.8$$
$$= -74.3kJ/mol$$

Answer: $\Delta H_f = -74.3kJmol^{-1}$

Q.NO. 26: State Hess's Law and explain its advantage.

Answer: Hess's Law of Constant Heat Summation states that the total enthalpy change for a chemical reaction is the same, regardless of whether the reaction takes place in a single step or in a series of steps, provided the initial and final states are the same.

Advantage: It allows for the calculation of enthalpy changes (ΔH) for reactions that are difficult or impossible to measure directly in the laboratory, such as reactions that are too slow, have side reactions, or do not occur in isolation (e.g., formation of CO from C and O_2).

Q.NO. 27: Define Bond Dissociation Energy.

Answer: Bond Dissociation Energy (or Bond Energy) is the average amount of energy required to break one mole of a particular type of bond in gaseous molecules to produce gaseous atoms. It is a measure of the strength of a chemical bond. For example, the bond energy of the H-H bond is $436kJmol^{-1}$. Bond breaking is always an endothermic process.

Q.NO. 28: Why is bond energy always positive?

Answer: Bond energy is always positive because energy is required to overcome the attractive forces holding the atoms together in a bond. Breaking a bond is an **endothermic** process, meaning heat must be absorbed from the surroundings to separate the atoms. Therefore, the enthalpy change (ΔH) for bond breaking is always positive.

Q.NO. 29: Define Specific Heat Capacity. What is its value for water?

Answer: Specific Heat Capacity (c) is the amount of heat energy required to raise the temperature of one gram of a substance by one degree Celsius (or one Kelvin). The specific heat capacity of water is approximately $4.18Jg^{-1}K^{-1}$ (or $4.2Jg^{-1}K^{-1}$). This high value means water can absorb a lot of heat with a relatively small change in temperature.

Q.NO. 30: Explain the meaning of each symbol in the formula $q = mc\Delta T$.

Answer: In the formula $q = mc\Delta T$:

- **q** represents the amount of heat energy absorbed or released (usually in Joules, J).
- **m** represents the mass of the substance undergoing the temperature change (usually in grams, g).
- **c** represents the specific heat capacity of the substance ($Jg^{-1}K^{-1}$).
- ΔT represents the change in temperature ($\Delta T = T_{final} - T_{initial}$), usually in Kelvin (K) or Celsius ($^{\circ}C$).

Q.NO. 31: Define Lattice Energy.

Answer: Lattice energy is the standard enthalpy change that occurs when one mole of a solid ionic crystal is formed from its constituent gaseous ions. This process is always exothermic (negative ΔH) because it involves the formation of strong electrostatic attractions between oppositely charged ions. Alternatively, it can be defined as the energy required to break one mole of an ionic solid into its gaseous ions (endothermic).

Q.NO. 32: Define Hydration Energy.

Answer: Hydration energy is the enthalpy change when one mole of gaseous ions is dissolved in sufficient water to give an infinitely dilute solution. It is a measure of the attraction between ions and water molecules. Hydration energy is always negative (exothermic) as energy is released when ions are surrounded by water molecules.

Q.NO. 33: Why is lattice breaking endothermic while hydration is exothermic?

Answer: Lattice breaking involves overcoming the strong electrostatic forces of attraction between oppositely charged ions in a crystal lattice to separate them. This requires an input of energy, making it **endothermic** (positive ΔH). Hydration involves the formation of attractive forces (ion-dipole interactions) between these separated ions and polar water molecules. Forming attractive forces releases energy, making hydration **exothermic** (negative ΔH).

Q.NO. 34: What is a Born-Haber cycle?

Answer: The Born-Haber cycle is a thermochemical cycle based on Hess's Law. It relates the lattice energy of an ionic solid to other measurable thermodynamic quantities, such as ionization energy, electron affinity, enthalpy of atomization, and enthalpy of formation. It is primarily used to calculate lattice energies which cannot be measured directly.

Q.NO. 35: Is electron affinity exothermic or endothermic? Explain.

Answer: The **first electron affinity** (adding the first electron to a neutral atom) is generally **exothermic** (negative) because the incoming electron is attracted to the nucleus, releasing energy. However, the **second electron affinity** (adding a second electron to a negative ion) is always **endothermic** (positive). This is because energy is required to overcome the electrostatic repulsion between the negative ion and the incoming negative electron.

Q.NO. 36: Standard enthalpy of combustion of ethanol?

Answer: The standard enthalpy of combustion (ΔH_{c°) of ethanol (C_2H_5OH) is the enthalpy change when one mole of ethanol is completely burned in excess oxygen under standard conditions.
The reaction is: $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$
This is an exothermic process. (While the exact value is not explicitly in the snippet text, the process is described as exothermic combustion).

Q.NO. 37: Why are bond energies different in different compounds for the same bond?

Answer: Bond energies for a specific bond (e.g., C-H) can vary slightly in different compounds because the bond strength is influenced by the **chemical environment** of the bond. The presence of neighboring atoms, their electronegativity, and the overall molecular structure affect the electron distribution and, consequently, the energy required to break the bond. Therefore, average bond energies are often used.

Q.NO. 38: What happens to energy when a chemical bond is formed?

Answer: When a chemical bond is formed between two atoms, energy is **released**. This is because the bonded state is more stable (lower energy) than the state of separate atoms. Therefore, bond formation is always an **exothermic** process (negative enthalpy change).

Q.NO. 39: Explain the Joule-Thomson effect.

Answer: The Joule-Thomson effect describes the temperature change of a real gas or liquid when it is forced through a valve or porous plug while kept insulated so that no heat is exchanged with the environment. Generally, when a highly compressed gas is allowed to expand into a region of low pressure, it cools down. This cooling occurs because the gas molecules move apart against the intermolecular forces of attraction, consuming energy from the gas itself, which lowers its kinetic energy and temperature.

Q.NO. 40: Why can gases be compressed easily?

Answer: Gases can be compressed easily because there are large empty spaces between the gas molecules. According to the Kinetic Molecular Theory, the actual volume of gas molecules is negligible compared to the total volume of the gas. When pressure is applied, these molecules are pushed closer together into the available empty space, significantly reducing the volume.

Q.NO. 41: Which forces are present among the molecules of CCl_4 and SiF_4?

Answer: The molecules of Carbon tetrachloride (CCl_4) and Silicon tetrafluoride (SiF_4) are non-polar. Therefore, the only intermolecular forces present between these molecules are **London dispersion forces**, also known as instantaneous dipole-induced dipole (id-id) forces. These weak forces arise due to temporary fluctuations in electron distribution that create temporary dipoles.

Q.NO. 42: Differentiate between instantaneous dipole-induced dipole (id-id) and permanent dipole (pd-pd) forces.

Answer: Instantaneous dipole-induced dipole (id-id) forces, or London dispersion forces, are weak attractive forces present in all molecules (including non-polar ones) due to temporary fluctuations in electron clouds. Permanent dipole-permanent dipole (pd-pd) forces are stronger attractive forces that exist between polar molecules where the positive end of one molecule attracts the negative end of another, as seen in HCl.

Q.NO. 43: Can the CHF_3 molecule make a hydrogen bond? Explain.

Answer: No, the CHF_3(fluoroform) molecule cannot form hydrogen bonds with itself. Although it contains highly electronegative fluorine atoms and a hydrogen atom, the hydrogen is bonded to a carbon atom, not to an electronegative atom like Fluorine, Oxygen, or Nitrogen. For a hydrogen bond to form, the hydrogen must be directly bonded to a highly electronegative atom (F, O, N) to acquire a sufficient partial positive charge.

Q.NO. 44: Show a hydrogen bond between two molecules of ethanol.

Answer: In ethanol (C_2H_5OH), the hydrogen atom of the hydroxyl group (-OH) is bonded to an electronegative oxygen atom, creating a partial positive charge on the hydrogen ($\delta +$) and a partial negative charge on the oxygen ($\delta -$). A hydrogen bond forms between the lone pair of the oxygen atom of one ethanol molecule and the hydrogen atom of the hydroxyl group of a neighboring ethanol molecule
($C_2H_5 - O - H \cdots : O(H) - C_2H_5$).

Q.NO. 45: Explain why the boiling point difference decreases as the size of the alcohol molecules increases.

Answer: As the size of the alcohol molecules increases (e.g., from methanol to butanol), the non-polar alkyl chain becomes larger, making London dispersion forces more significant relative to hydrogen bonding. In smaller alcohols, hydrogen bonding is the dominant force causing high boiling points. In larger alcohols, the contribution of hydrogen bonding diminishes in comparison to the growing dispersion forces, leading to a smaller difference in boiling points between successive members.

Q.NO. 46: Why is the viscosity of honey higher than water?

Answer: Honey has a much higher viscosity than water due to strong intermolecular forces, particularly extensive hydrogen bonding between the sugar molecules (like glucose and fructose) and water. These strong attractive forces prevent the layers of the liquid from sliding over one another easily, creating high resistance to flow.

Q.NO. 47: Which is more viscous: glycerine or hexane? Why?

Answer: Glycerine is more viscous than hexane. Glycerine ($CH_2OH - CHOH - CH_2OH$) contains three hydroxyl (-OH) groups capable of forming extensive hydrogen bonds, resulting in very strong intermolecular attraction and high resistance to flow. Hexane (C_6H_{14}) is a non-polar hydrocarbon with only weak London dispersion forces, allowing its molecules to move past each other easily.

Q.NO. 48: Why does evaporation get faster at higher temperatures?

Answer: Evaporation gets faster at higher temperatures because the average kinetic energy of the liquid molecules increases. More molecules gain enough energy to overcome the intermolecular attractive forces holding them in the liquid phase and escape from the surface into the vapor phase. This increases the rate of evaporation.

Q.NO. 49: Why is food cooking difficult in areas with high altitudes?

Answer: At high altitudes, the external atmospheric pressure is lower than at sea level. Since a liquid boils when its vapor pressure equals the external pressure, water boils at a temperature lower than $100°C$(e.g., at Murree, water boils at $98°C$, and at Mount Everest at $69°C$). This lower boiling temperature provides less heat energy to the food, making it take longer to cook or remain undercooked.

Q.NO. 50: Explain why food cooks faster in a pressure cooker.

Answer: A pressure cooker is a closed vessel that traps steam, causing the pressure inside to rise significantly above atmospheric pressure. This increased external pressure raises the boiling point of water inside the cooker (above $100°C$). The food is cooked at this higher temperature, allowing it to absorb more heat energy and cook much faster than in an open pot.

Q.NO. 51: Why is the boiling point of water ($100°C$) higher than that of ethanol ($78°C$)? Answer: The boiling point of water is higher than that of ethanol because water has stronger and more extensive hydrogen bonding. Each water molecule has two hydrogen atoms and two lone pairs, allowing it to form up to four hydrogen bonds, creating a 3D network. Ethanol can form fewer hydrogen bonds per molecule. More energy is required to break the extensive bonding in water, leading to a higher boiling point.

Q.NO. 52: What is meant by the "habit of a crystal"?

Answer: The "habit of a crystal" refers to the characteristic external shape in which a crystal usually grows under maintained conditions. If the conditions for growth (such as solvent, temperature, or impurities) are kept constant, a crystal will always grow in the same specific shape or habit. For example, NaCl typically grows as cubic crystals.

Q.NO. 53: Name the properties of liquid crystals in which they resemble solids.

Answer: Liquid crystals resemble solids in the following properties:

1. **Optical Properties:** They show optical properties similar to crystalline solids.
2. **Order:** They possess a parallel ordered arrangement of molecules.
3. **Rigidity:** They possess some degree of rigidity compared to pure liquids.
4. **Anisotropy:** They are anisotropic, meaning their properties depend on the direction of measurement.

Q.NO. 54: Which property of liquid crystals makes them useful in temperature sensing devices?

Answer: Liquid crystals, specifically cholesteric liquid crystals, change color with changes in temperature. This property allows them to be used in temperature sensing devices like thermometers and for medical

diagnosis (e.g., locating tumors which are warmer than surrounding tissue). The arrangement of molecules changes with temperature, altering the way they reflect light.

Q.NO. 55: Why do solids not undergo translatory motion?

Answer: The constituent particles (atoms, ions, or molecules) of a solid are held together by strong cohesive forces (ionic, covalent, metallic, or van der Waals forces) in fixed positions. They are closely packed and do not have the freedom to move past one another. Therefore, they only vibrate about their mean positions and do not undergo translatory motion.

Q.NO. 56: Why does iodine exist as a solid at room temperature?

Answer: Iodine (I_2) exists as a solid at room temperature because it has a large molecular size. The large size of the iodine molecule leads to strong, polarizable electron clouds, which result in relatively strong London dispersion forces (instantaneous dipole-induced dipole forces) between the molecules. These forces are strong enough to hold the molecules in a solid lattice at room temperature.

Q.NO. 57: Why does HF have a higher boiling point than HCl?

Answer: HF has a higher boiling point ($19.9°C$) than HCl because of the presence of strong hydrogen bonding between HF molecules. Fluorine is highly electronegative, creating a strong dipole that allows for hydrogen bonds. Chlorine is less electronegative and larger, so HCl molecules interact primarily through weaker dipole-dipole forces, which require less energy to overcome.

Q.NO. 58: Why do boiling points of halogens increase down the group?

Answer: The boiling points of halogens (Group 17) increase down the group (from Fluorine to Iodine) because the atomic size and number of electrons increase. This leads to greater polarizability of the electron clouds, which results in stronger London dispersion forces (Van der Waals forces) between the molecules. Stronger forces require more energy to break, raising the boiling point.

Q.NO. 59: Why is HF less acidic than HCl despite stronger bonding?

Answer: HF is a weaker acid than HCl primarily due to the very strong bond between Hydrogen and Fluorine and the extensive hydrogen bonding in aqueous solution. The hydrogen bonding entraps the H^+ ion between highly electronegative Fluorine atoms, making its dissociation and release as a free proton more difficult compared to HCl, which dissociates completely.

Q.NO. 60: Define Surface Tension and Viscosity.

Answer:

- **Surface Tension:** A property of liquids describing the force acting along the surface, causing it to behave like a stretched elastic sheet. It arises from the inward pull of intermolecular forces on surface molecules ($Units: Nm^{-1}$).
- **Viscosity:** The internal resistance of a liquid to flow. It arises from the frictional forces between adjacent layers of the liquid moving at different speeds ($Units: kgm^{-1}s^{-1}$ or poise).

Q.NO. 61: Why can insects like mosquitoes walk on the surface of water?

Answer: Insects like mosquitoes can walk on the surface of water due to the high surface tension of water. The strong intermolecular forces (hydrogen bonds) between water molecules at the surface create a "skin" or elastic sheet-like effect that can support the light weight of the insect without breaking.

Q.NO. 62: What happens to the surface tension of water when temperature increases?

Answer: When the temperature of water increases, its surface tension decreases. This is because the increase in kinetic energy of the water molecules weakens the intermolecular forces (hydrogen bonds) holding them together. Weaker intermolecular forces result in a reduced inward pull on the surface molecules, lowering the surface tension.

Q.NO. 63: Why does vapour pressure increase with temperature?

Answer: Vapour pressure increases with temperature because heating a liquid increases the average kinetic energy of its molecules. This enables a larger fraction of molecules to overcome the intermolecular forces and escape from the liquid surface into the vapor phase. The increased rate of evaporation leads to a higher concentration of vapor molecules and thus higher pressure.

Q.NO. 64: Explain why vapour pressure is independent of container size.

Answer: Vapour pressure is an intrinsic property of a liquid at a given temperature and depends only on the nature of the liquid and temperature. It does not depend on the amount of liquid, surface area, or the volume of the container. While a larger surface area changes the *rate* of evaporation and condensation, the equilibrium pressure reached remains the same.

Q.NO. 65: Why are liquid crystals always anisotropic?

Answer: Liquid crystals are anisotropic because their molecules are rod-like and arrange themselves in an ordered, parallel fashion. This ordered arrangement causes their physical properties (such as optical, electrical, and magnetic properties) to have different values when measured in different directions, which is the definition of anisotropy.

Q.NO. 66: How are liquid crystals used in breast cancer diagnosis?

Answer: Cholesteric liquid crystals are used in breast cancer diagnosis because they are sensitive to temperature changes. Tumors often have a slightly higher temperature than the surrounding healthy tissue. When a layer of liquid crystals is applied to the skin, the warmer tumor area causes the crystals to change color, allowing for the detection of the abnormality.

Q.NO. 67: Define Crystal Lattice and Unit Cell.

Answer:

- **Crystal Lattice:** The regular arrangement of atoms, ions, or molecules in a three-dimensional space within a crystal.
- **Unit Cell:** The smallest repeating unit of the crystal lattice that shows the complete symmetry and pattern of the crystal. Stacking unit cells in three dimensions builds the entire crystal lattice.

Q.NO. 68: Why do crystalline solids have sharp melting points?

Answer: Crystalline solids have sharp melting points because their constituent particles (atoms, ions, or molecules) are arranged in a highly regular and ordered pattern. All the bonds holding the particles together are of the same strength and break at the same specific temperature. This causes the solid to melt completely at a definite temperature.

Q.NO. 69: Explain how urea impurity affects the habit of an $NaCl$ crystal.

Answer: The habit of a crystal refers to its growth shape. Pure Sodium Chloride ($NaCl$) grows as cubic crystals. However, if an impurity like 10urea is present in the solution, the urea molecules selectively interact with certain faces of the growing crystal. This alters the growth rate of the faces and causes the $NaCl$crystals to grow in a needle-like shape instead of cubes.

Q.NO. 70: Why is water more effective in temperature regulation than other liquids?

Answer: Water has a high specific heat capacity ($4.18 Jg^{-1}K^{-1}$) due to extensive hydrogen bonding. It can absorb or release large amounts of heat with only a small change in its own temperature. This property makes it highly effective at buffering temperature changes in the environment and in living organisms, aiding in temperature regulation.

Q.NO. 71: Differentiate between exothermic and endothermic reactions with examples.

Answer:

- **Exothermic reactions** release heat to the surroundings ($\Delta H$is negative). The enthalpy of products is lower than reactants. Example: Combustion of Carbon ($C + O_2 \rightarrow CO_2, \Delta H = -393.7 kJ/mol$).
- **Endothermic reactions** absorb heat from the surroundings ($\Delta H$is positive). The enthalpy of products is higher than reactants. Example: Decomposition of Calcium Carbonate ($CaCO_3 \rightarrow CaO + CO_2$, $\Delta H = +572 kJ/mol$).

Q.NO. 72: What do you understand by the enthalpy of a system?

Answer: The enthalpy of a system (H) is the total heat content possessed by a substance due to its structure (types of bonds) and physical state. It is a state function defined as the sum of the internal energy (E) and the product of pressure and volume (PV), though in chemistry it is mainly used to quantify heat changes (ΔH) at constant pressure.

Q.NO. 73: Distinguish clearly between standard enthalpy of reaction and standard enthalpy of formation.

Answer:

- **Standard Enthalpy of Reaction ($\Delta H_{r\circ}$):** The enthalpy change when stoichiometric amounts of reactants react completely under standard conditions ($298K, 1atm$).

- **Standard Enthalpy of Formation ($\Delta H_{f\circ}$):** The enthalpy change specifically when **one mole** of a compound is formed from its constituent elements in their standard states under standard conditions. Example: $C + 2H_2 \rightarrow CH_4$.

Q.NO. 74: Define and give one example of: Standard enthalpy of solution and hydration. Answer:

- **Standard Enthalpy of Solution ($\Delta H_{sol\circ}$):** The enthalpy change when one mole of a substance dissolves in sufficient solvent to form an infinitely dilute solution. Example: $NH_4Cl(s) \rightarrow NH_{4+}(aq) + Cl^-(aq), \Delta H = +16.2kJ/mol$.

- **Hydration Energy (ΔH_{hyd}):** The enthalpy change when one mole of gaseous ions dissolves in water to give an infinitely dilute solution. Example: $Na^+(g) \rightarrow Na^+(aq)$.

Q.NO. 75: Explain why the lattice enthalpy of an ionic compound is typically a large negative value.

Answer: Lattice enthalpy is the energy released when gaseous ions come together to form one mole of a solid ionic crystal lattice. This process is highly exothermic (large negative ΔH) because strong electrostatic forces of attraction are established between the oppositely charged ions ($cation$ +and $anion$ −) as they form a stable crystal structure.

Q.NO. 76: Explain why the enthalpy of hydration is always an exothermic process for gaseous ions.

Answer: The enthalpy of hydration is always exothermic (negative) because it involves the formation of attractive forces between the charged gaseous ions and the polar water molecules (ion-dipole interactions). The formation of these bonds releases energy. Since the gaseous ions are already separated, no energy is consumed to break bonds, so the net process releases heat.

Q.NO. 77: Define standard enthalpy of atomization with an example.

Answer: The standard enthalpy of atomization ($\Delta H_{at\circ}$) is the enthalpy change involved when one mole of gaseous atoms is formed from an element in its standard state under standard conditions ($298K, 1atm$). Example: $\frac{1}{2}H_2(g) \rightarrow H(g), \Delta H_{at\circ} = +218kJmol^{-1}$.

Q.NO. 78: The enthalpy of solution can be positive or negative. Explain what this indicates.

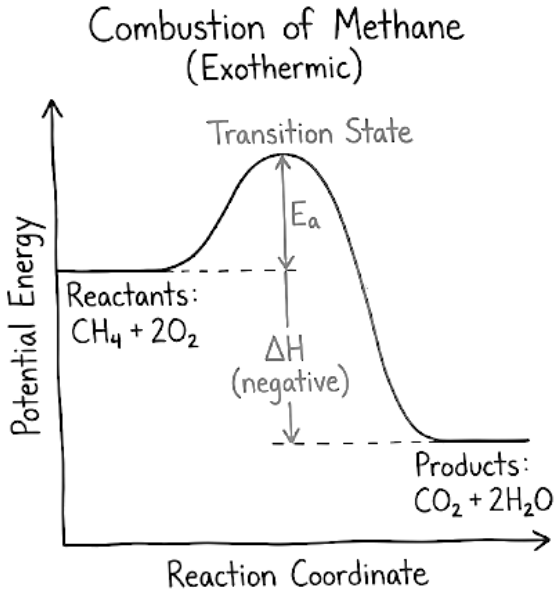
Answer: The enthalpy of solution (ΔH_{sol}) is the sum of the lattice energy (endothermic, bond breaking) and hydration energy (exothermic, bond forming).

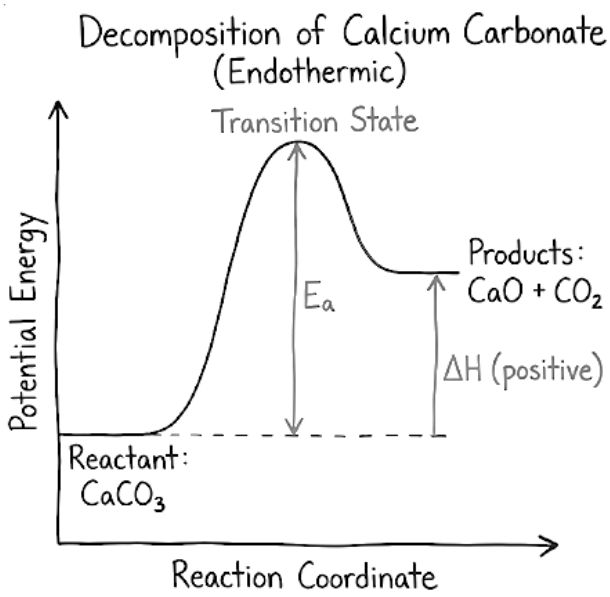
- If ΔH_{sol} is **positive** (endothermic), it indicates that the energy required to break the lattice is greater than the energy released by hydration.

- If ΔH_{sol} is **negative** (exothermic), it indicates that the hydration energy released exceeds the lattice energy.

Q.NO. 79: Draw energy profile diagrams for the combustion of methane and decomposition of calcium carbonate.

Answer:





Q.NO. 80: Calculate ΔH° for the formation of methane using enthalpies of combustion of C, H_2, and CH_4.

- **Answer:** Reaction: $C(s) + 2H_2(g) \rightarrow CH_4(g)$Formula: $\Delta H_f = \Sigma \Delta H_c(reactants) - \Sigma \Delta H_c(products)$Data: $\Delta H_c(C) = -393.5kJ/mol$, $\Delta H_c(H_2) = -285.8kJ/mol$, $\Delta H_c(CH_4) = -890.8kJ/mol$. $\Delta H_f = [-393.5 + 2(\ -285.8)] - [\ -890.8]\Delta H_f = [\ -393.5-571.6] + 890.8 = -965.1 + 890.8 = -74.3kJ/mol$. +1

Q.NO. 81: State Hess's Law and explain its advantage.

Answer: Hess's Law states that the total enthalpy change for a reaction is the same whether it occurs in a single step or a series of steps, provided the initial and final states are the same. Its advantage is that it allows us to calculate enthalpy changes for reactions that cannot be measured directly in the lab (e.g., slow reactions or those with side products) by using data from other measurable reactions.

Q.NO. 82: Calculate the enthalpy change for the Haber process using bond energies ($N \equiv N$, $H - H$, $N - H$).

Answer: Reaction: $N_2 + 3H_2 \rightarrow 2NH_3$Bonds Broken: $1 \times (N \equiv N) + 3 \times (H{-}H) = 945 + 3(436) = 945 + 1308 = 2253kJ/mol$. Bonds Formed: $2 \times 3 \times (N{-}H) = 6 \times 391 = 2346kJ/mol$. $\Delta H = BondBreaking - BondForming = 2253 - 2346 = -93kJ/mol$. (Note: Using bond energy data from Table 6.1 in the text) .

Q.NO. 83: Define Bond Dissociation Energy.

Answer: Bond Dissociation Energy is the average amount of energy required to break one mole of a particular type of bond in gaseous molecules to produce gaseous atoms. It is a measure of the strength of the bond. For example, the bond energy of the $H - H$bond is $436kJmol^{-1}$.

Q.NO. 84: Why is bond energy always positive?

Answer: Bond energy is always positive because breaking a chemical bond requires energy to overcome the attractive forces holding the atoms together. This process is endothermic (absorbs heat). Therefore, by convention, energy absorbed is represented with a positive sign.

Q.NO. 85: Define Specific Heat Capacity. What is its value for water?

Answer: Specific Heat Capacity (c) is the amount of heat energy required to raise the temperature of one gram of a substance by one degree Kelvin (or Celsius). The specific heat capacity of water is $4.18Jg^{-1}K^{-1}$

Q.NO. 86: Explain the meaning of each symbol in the formula $q = mc\Delta T$.
Answer: In the formula $q = mc\Delta T$:

- **q:** Amount of heat energy absorbed or released (Joules, J).
- **m:** Mass of the substance (grams, g).
- **c:** Specific heat capacity of the substance ($Jg^{-1}K^{-1}$).
- ΔT: Change in temperature ($T_{final} - T_{initial}$) in Kelvin or Celsius.

Q.NO. 87: Why is insulation used in a calorimeter?
Answer: Insulation (like cotton wool) is used in a calorimeter to prevent the loss or gain of heat from the surroundings. The goal of calorimetry is to measure the heat evolved or absorbed *by the reaction* alone. Insulation ensures that the heat change is confined to the calorimeter's contents (water/solution), allowing for an accurate calculation of ΔHbased on temperature change.
Q.NO. 88: Define the calorie content of food in terms of energetics.
Answer: The calorie content of food refers to the amount of energy released when the food is metabolized or "burned" in the body. It is often measured using combustion calorimetry. In energetics, it corresponds to the enthalpy of combustion (ΔH_c) of the food substances (carbohydrates, fats, proteins), which provides the energy required for bodily functions.
Q.NO. 89: Give the combustion equation of glucose with its ΔHvalue.
Answer: The complete combustion of glucose involves its reaction with oxygen to form carbon dioxide and water. The equation and enthalpy change are: $C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$
$\Delta H = -2802kJmol^{-1}$(approx, typically around -2800 kJ/mol for glucose combustion).
Q.NO. 90: On which fundamental law is Hess's Law based?
Answer: Hess's Law is based on the **First Law of Thermodynamics**, which is the Law of Conservation of Energy. It implies that energy cannot be created or destroyed, only converted. Therefore, the net energy change for a chemical process depends only on the initial and final states, not on the path taken.
Q.NO. 91: Why is Hess's Law important for reactions that are too slow or have side reactions? Answer: Hess's Law is important because it allows chemists to calculate the enthalpy change (ΔH) for reactions that are too slow to measure directly in a calorimeter or those that produce unwanted side products (mixed reactions). By using ΔHvalues from other known, fast, and clean reactions (like combustion or formation), the unknown ΔHcan be derived mathematically.
Q.NO. 92: Define Lattice Energy.
Answer: Lattice energy is the standard enthalpy change that occurs when one mole of a solid ionic crystal is formed from its widely separated gaseous ions. This process is highly exothermic (negative ΔH). Alternatively, it is the energy required to break one mole of ionic solid into gaseous ions (endothermic).
Q.NO. 93: Define Hydration Energy.
Answer: Hydration energy is the enthalpy change involved when one mole of gaseous ions dissolves in sufficient water to give an infinitely dilute solution. It reflects the attraction between the ions and water molecules and is always an exothermic process (negative ΔH).
Q.NO. 94: Why is lattice breaking endothermic while hydration is exothermic?
Answer: Lattice breaking is **endothermic** because energy must be supplied to overcome the strong electrostatic forces holding the ions together in the crystal lattice. Hydration is **exothermic** because energy is released when the separated ions form new attractive interactions (ion-dipole bonds) with the polar water molecules.
Q.NO. 95: What is a Born-Haber cycle?
Answer: A Born-Haber cycle is a thermochemical cycle based on Hess's Law. It allows for the calculation of lattice energy (which cannot be measured directly) by relating it to other measurable thermodynamic quantities such as ionization energy, electron affinity, enthalpy of atomization, and enthalpy of formation.
Q.NO. 96: Name the steps involved in the Born-Haber cycle of $NaCl$.
Answer: The steps involved in the Born-Haber cycle for $NaCl$include:

1. Atomization of Sodium ($Na(s) \rightarrow Na(g)$)

2. Ionization of Sodium ($Na(g) \rightarrow Na^+(g) + e^-$)
3. Atomization of Chlorine ($\frac{1}{2}Cl_2(g) \rightarrow Cl(g)$)
4. Electron Affinity of Chlorine ($Cl(g) + e^- \rightarrow Cl^-(g)$)
5. Lattice Energy (Formation of solid: $Na^+(g) + Cl^-(g) \rightarrow NaCl(s)$).

Q.NO. 97: Is electron affinity exothermic or endothermic? Explain.

Answer: The **first** electron affinity is generally **exothermic** (negative ΔH) because energy is released when a neutral atom attracts an incoming electron. However, the **second** electron affinity is **endothermic** (positive ΔH) because energy is required to force a negatively charged electron onto an already negative ion against electrostatic repulsion.

Q.NO. 98: What is the standard enthalpy of combustion of ethanol?

Answer: The standard enthalpy of combustion ($\Delta H_{c°}$) of ethanol is the energy released when one mole of ethanol burns completely in oxygen. Its value is $-1368 kJmol^{-1}$.

Reaction: $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$.

Q.NO. 99: Why are bond energies different in different compounds for the same bond?

Answer: Bond energies for the same bond (e.g., $C - H$) vary in different compounds because the bond strength is influenced by the chemical environment, such as the neighboring atoms and the overall structure of the molecule. The electronic distribution differs slightly in each molecule, affecting the energy required to break the bond. Therefore, "average bond energies" are used.

Q.NO. 100: What happens to energy when a chemical bond is formed?

Answer: When a chemical bond is formed between two atoms, energy is **released**. This occurs because the bonded system is more stable and has lower energy than the independent atoms. Consequently, bond formation is always an **exothermic** process (negative enthalpy change).

Q.NO. 1: Differentiate between Aliphatic and Aromatic hydrocarbons.

Answer: Aliphatic hydrocarbons are compounds that are not aromatic; they can be open-chain (acyclic) or cyclic (cycloalkanes) and may be saturated (alkanes) or unsaturated (alkenes). In contrast, aromatic hydrocarbons are a special class of cyclic hydrocarbons that have a high carbon-to-hydrogen ratio and are typically based on the benzene ring (C_6H_6) or resembling compounds, such as toluene and phenol.

Q.NO. 2: Differentiate between Homolytic and Heterolytic Fission.

Answer: Homolytic fission occurs when a covalent bond breaks evenly, with each bonded atom retaining one electron, resulting in the formation of free radicals (species with unpaired electrons). Heterolytic fission occurs when a covalent bond breaks unevenly, such that the shared pair of electrons is gained by only one atom, resulting in the formation of oppositely charged ions.

Q.NO. 3: Define Electrophile and Nucleophile with examples.

Answer: An electrophile is a species that acts as an electron pair acceptor (represented by E^+), such as the slightly positive bromine atom in a polarized Br_2molecule. A nucleophile is a species that acts as an electron pair donor (represented by Nu^-), such as the π-bond electrons in an alkene or a bromide ion (Br^-).

Q.NO. 4: Explain why alkanes do not undergo addition reactions.

Answer: Alkanes are saturated hydrocarbons containing only single sigma (σ) bonds, which are strong and stable. Because all carbon valencies are fully satisfied with hydrogen or other carbon atoms, there are no double or triple bonds to "add" reagents to. They generally undergo substitution reactions rather than addition.

Q.NO. 5: Explain why 2-bromopropane is the major product when propene reacts with HBr. Answer: The reaction follows Markovnikov's rule, which states that hydrogen adds to the carbon with more hydrogen atoms to form the more stable carbocation. In this case, the addition of $H^+$to propene forms a secondary (2°) carbocation, which is more stable than the primary (1°) carbocation due to the electron-donating inductive effect of the alkyl groups. The bromide ion then attacks the stable secondary carbocation to form 2-bromopropane.

Q.NO. 6: Explain how inductive effects from alkyl groups stabilize carbocations.

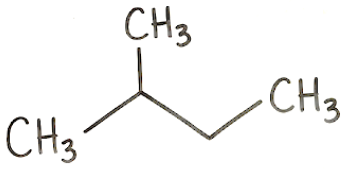
Answer: Alkyl groups possess an electron-donating positive inductive effect ($+I$). They push electron density toward the positively charged carbon atom of the carbocation. This delocalization or dispersal of the

positive charge makes the carbocation energetically more stable. Therefore, tertiary carbocations (3 alkyl groups) are more stable than secondary, which are more stable than primary ones.

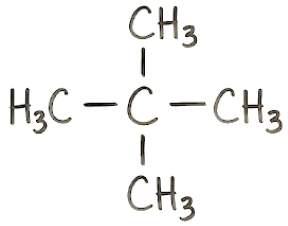
Q.NO. 7: Write down structural formulas for 2-methylbutane and 2,2-dimethylpropane.

Answer:

2-methylbutane



2,2-dimethylpropane



Q.NO. 8: Give two differences between molecules of cyclopentane and pentane.

Answer:

1. **Structure:** Pentane (C_5H_{12}) is an open-chain alkane, whereas cyclopentane (C_5H_{10}) is a cyclic alkane with carbon atoms arranged in a ring.
2. **Saturation/Formula:** While both are saturated, cyclopentane has two fewer hydrogen atoms than pentane due to the ring closure.

Q.NO. 9: Why do branched alkanes have lower boiling points than straight-chain alkanes?

Answer: Branched alkanes have a more compact structure with a smaller surface area compared to their straight-chain isomers. This reduced surface area leads to weaker intermolecular forces (Van der Waals forces) between the molecules. Consequently, less energy is required to overcome these forces, resulting in lower boiling points (e.g., n-pentane boils at $36.2°C$while neopentane boils at $9.5°C$).

Q.NO. 10: Explain why alkanes have a tetrahedral shape.

Answer: The carbon atoms in an alkane are sp^3hybridized, forming four single sigma bonds with surrounding atoms. These four electron pairs repel each other equally to maximize distance, resulting in a tetrahedral geometry with bond angles of approximately $109.5°$.

Q.NO. 11: Explain why alkanes are highly stable (paraffins).

Answer: Alkanes are called paraffins (Latin: *parum* = little, *affinis* = affinity) because they are inert towards acids, alkalis, and oxidizing agents under normal conditions. This stability is due to the non-polarity of the $C-H$and $C-C$bonds (electronegativity difference is small) and the strength of the sigma bonds, which require significant energy to break.

Q.NO. 12: Explain the mechanism of halogenation of methane.

Answer: Halogenation proceeds via a free radical substitution mechanism in three steps:

1. **Initiation:** UV light breaks the halogen bond ($Cl-Cl$) homolytically to form free radicals ($Cl^{\cdot}$).
2. **Propagation:** The radical attacks the alkane to form an alkyl radical, which then reacts with more halogen to form the product and regenerate the halogen radical.
3. **Termination:** Two free radicals combine (e.g., $Cl^{\cdot}+Cl^{\cdot}$or $CH_3.+Cl^{\cdot}$) to form a stable molecule, ending the chain.

Q.NO. 13: Give the dehydrohalogenation reaction of bromopropane.

Answer: When bromopropane is heated with alcoholic KOH, a hydrogen atom and a bromine atom are removed from adjacent carbons to form an alkene (propene).

Equation: $CH_3-CH_2-CH_2-Br+KOH(alcoholic)heat\ CH_3-CH=CH_2+KBr+H_2O$.

Q.NO. 14: How does an alkane differ from an alkene in terms of stability?

Answer: Alkenes are less stable and more reactive than alkanes due to the presence of a double bond containing a π(pi) bond. The π-electrons are held less firmly and are further from the nuclei (exposed),

making them susceptible to attack by electrophiles. Alkanes contain only strong sigma bonds, making them chemically inert.

Q.NO. 15: Why is the order of stability of carbocations 3° > 2° > 1°?

Answer: The stability of carbocations is determined by the inductive effect of alkyl groups attached to the positively charged carbon. Alkyl groups are electron-donating ($+I$effect). A tertiary (3°) carbocation has three alkyl groups helping to disperse the positive charge, whereas a primary (1°) has only one. Greater charge dispersal leads to greater stability.

Q.NO. 16: Compare the relative rates of addition to alkenes for HCl, HBr, and HI.

Answer: The rate of addition of hydrogen halides to alkenes increases from $HF$to HI. Therefore, the reaction is fastest with HI, followed by HBr, and then HCl. This trend is related to the bond strength; the $H-I$bond is weaker and breaks more easily than the $H-Cl$bond.

Q.NO. 17: Explain how Markovnikov's rule is applied in the addition of HBr to 2-pentene. Answer: Markovnikov's rule applies when adding a polar reagent like HBr to an unsymmetrical alkene. The negative part of the reagent (Br^-) attaches to the carbon atom of the double bond that has fewer hydrogen atoms. This pathway proceeds via the most stable carbocation intermediate formed by adding the proton (H^+) to the carbon with more hydrogens.

Q.NO. 18: What is a free radical?

Answer: A free radical is a chemical species that contains an unpaired electron. It is formed by homolytic fission of a covalent bond and is extremely reactive because the unpaired electron seeks to become paired.

Q.NO. 19: Why are alkenes called olefins?

Answer: (Text context implies): The term "Olefin" is often used for alkenes. Historically, it refers to "oil-forming" gas, as lower alkenes react with chlorine to form oily liquids (like 1,2-dichloroethane). Alkenes are unsaturated hydrocarbons containing a double bond.

Q.NO. 20: Describe geometrical isomerism (cis-trans) in alkenes.

Answer: Geometrical isomerism, or cis-trans isomerism, arises in alkenes due to the restricted rotation around the carbon-carbon double bond. Stereoisomers have the same structural formula but differ in the spatial arrangement of identical groups. "Cis" isomers have identical groups on the same side of the double bond, while "trans" isomers have them on opposite sides.

Q.NO. 21: How can you distinguish between an alkane and an alkene using bromine water? Answer: Bromine water (a red-brown solution) is used as a test for unsaturation. When added to an alkene (like ethene), the bromine adds across the double bond, causing the red-brown color to discharge (decolorize) immediately. Alkanes do not react with bromine water under these conditions, so the color persists.

Q.NO. 22: Which bond is weaker in a double bond: sigma or pi?

Answer: The π(pi) bond is weaker than the σ(sigma) bond. The π bond is formed by the lateral (sideways) overlap of p-orbitals, and its electron density is located above and below the plane of the nuclei, making it less firmly held and easier to break than the sigma bond formed by linear overlap.

Q.NO. 23: State the bond angle in ethene (120°).

Answer: In ethene, the carbon atoms are sp^2hybridized, leading to a trigonal planar arrangement. The bond angles between the atoms attached to the double-bonded carbons are approximately $120°$.

Q.NO. 24: Why is the boiling point of CH3OH (65°C) much higher than CH3CH3 (-89°C)? Answer: Methanol (CH_3OH) molecules are capable of forming strong hydrogen bonds due to the presence of the hydroxyl (-OH) group. Ethane (CH_3CH_3) is non-polar and interacts only through weak Van der Waals forces. The significantly stronger intermolecular forces in methanol require more energy to break, resulting in a much higher boiling point.

Q.NO. 25: Why does CCl_4 have higher viscosity than $CHCl_3$ but less than ethanol?

Answer: CCl_4has a higher viscosity than $CHCl_3$because $CCl_4$is a larger molecule with stronger dispersion forces (Van der Waals). However, Ethanol has a higher viscosity than both because it can form hydrogen bonds, which are stronger intermolecular forces that resist flow more effectively than the forces in CCl_4.

Q.NO. 26: Why does evaporation continue even at room temperature?

Answer: Evaporation is the spontaneous conversion of a liquid into vapor. It occurs at any temperature because molecules within the liquid have a distribution of kinetic energies. Some molecules possess sufficient kinetic energy to overcome the intermolecular forces and escape from the surface of the liquid into the vapor phase.

Q.NO. 27: Why do we feel cool near the bank of a river?

Answer: We feel cool due to the phenomenon of evaporation. High-energy water molecules escape the surface of the river, leaving behind lower-energy molecules, which causes the temperature of the water and the immediate surroundings to fall. This absorption of heat from the surroundings for evaporation creates a cooling effect.

Q.NO. 28: What does the peak in an energy profile diagram represent?

Answer: The peak in an energy profile diagram represents the transition state or the energy barrier that reactants must overcome to form products. The height of this peak from the reactant energy level is called the Activation Energy (E_a), which is the minimum energy required for the reaction to occur.

Q.NO. 29: What happens to the energy of a system when a bond is broken?

Answer: Bond breaking is an endothermic process. Energy must be absorbed by the system from the surroundings to overcome the attractive forces holding the atoms together. Therefore, the energy of the system increases when a bond is broken.

Q.NO. 30: What is meant by "indirect route" in Hess's Law?

Answer: Hess's Law states that the total enthalpy change of a chemical reaction is constant, regardless of whether the reaction proceeds in a single step (direct route) or through multiple steps (indirect route). The "indirect route" refers to the alternative pathway involving intermediates, the sum of whose enthalpy changes equals that of the direct path.

Q.NO. 31 Define Modern Periodic Law.

Answer: The Modern Periodic Law states that the "physical and chemical properties of elements are periodic functions of their atomic numbers." This means that when elements are arranged in order of increasing atomic number (number of protons), elements with similar properties appear at regular intervals. This law corrected the discrepancies in Mendeleev's table, which was based on atomic mass.

Q.NO. 32 Why does atomic radius decrease across a period?

Answer: As you move across a period (from left to right) in the periodic table, the atomic number increases, meaning the positive nuclear charge increases. However, the electrons are added to the same principal energy shell, so the shielding effect remains relatively constant. The stronger nuclear attraction pulls the electron cloud closer to the nucleus, resulting in a decrease in atomic radius.

Q.NO. 33 Why does atomic radius increase down a group?

Answer: Atomic radius increases down a group because, with each successive element, a new principal electron shell is added. This increases the distance between the nucleus and the valence electrons. Additionally, the number of inner electron shells increases, which creates a stronger shielding effect that outweighs the increase in nuclear charge, causing the atom to expand.

Q.NO. 34 Define Ionization Energy and give its trend in the periodic table.

Answer: The first ionization energy (H_{i1}) is the energy required to remove one electron from each atom in one mole of gaseous atoms to form one mole of gaseous 1^{+}ions. Generally, ionization energy increases across a period due to increased nuclear charge and decreases down a group due to increased atomic size and shielding effect.

Q.NO. 35 Why does shielding effect reduce ionization energy?

Answer: The shielding (or screening) effect occurs when inner shell electrons repel the valence electrons, effectively reducing the net nuclear attraction experienced by the outer electrons. Greater shielding makes the valence electrons easier to remove, thereby lowering the ionization energy. This is why ionization energy decreases down a group where shielding is higher.

Q.NO. 36 Explain why the second ionization energy of Calcium is higher than the first.

Answer: The first ionization of calcium removes an electron from a neutral atom ($Ca \rightarrow Ca^{+} + e^{-}$), while the second ionization removes an electron from a positively charged ion ($Ca^{+} \rightarrow Ca^{2+} + e^{-}$). The remaining

electrons in the Ca^{+}ion are held more tightly by the nucleus because the protons now outnumber the electrons, increasing the effective nuclear charge. Thus, removing the second electron requires significantly more energy ($1150kJmol^{-1}$) than the first ($590kJmol^{-1}$).

Q.NO. 37 Why do noble gases have the highest ionization energies?

Answer: Noble gases (Group 18) have the highest ionization energies in their respective periods because they possess a stable, fully filled valence shell configuration (ns^2np^6). This configuration is energetically very stable, and the effective nuclear charge is high. Consequently, removing an electron from such a stable system is extremely difficult and requires a large amount of energy.

Q.NO. 38 Why is the first electron affinity of Oxygen negative while the second is positive? Answer: The first electron affinity of oxygen is negative ($-141kJmol^{-1}$) because energy is released when the neutral atom attracts an electron to form O^{-}. However, the second electron affinity is positive ($+798kJmol^{-1}$) because the second electron is being added to an already negative ion (O^{-}). The electrostatic repulsion between the negative ion and the incoming electron requires energy to be absorbed to force the electron into the shell.

Q.NO. 39 Differentiate between Electron Affinity and Electronegativity.

Answer: Electron Affinity is the energy change that occurs when an electron is added to a gaseous atom to form a negative ion (a measurable thermodynamic quantity). Electronegativity, on the other hand, is a relative or dimensionless tendency of an atom to attract a shared pair of electrons towards itself in a covalent bond. While electron affinity refers to isolated atoms, electronegativity refers to bonded atoms.

Q.NO. 40 Explain the trend of Electronegativity across a period.

Answer: Electronegativity increases from left to right across a period. This is because the atomic radius decreases and the effective nuclear charge increases. A smaller atom with a higher positive charge on the nucleus exerts a stronger pull on the shared pair of electrons in a bond. For example, in Period 2, Lithium has a value of 1.0 while Fluorine has a value of 4.0.

Q.NO. 41 Why is the ionic radius of a cation smaller than its parent atom?

Answer: A cation is formed when an atom loses electrons. This loss reduces the electron-electron repulsion, allowing the remaining electrons to be pulled closer to the nucleus by the unchanged nuclear charge. In some cases, an entire valence shell is lost. Consequently, the ionic radius of a positive ion (cation) is always smaller than that of the neutral parent atom.

Q.NO. 42 Why is the ionic radius of an anion larger than its parent atom?

Answer: An anion is formed when an atom gains electrons. The addition of electrons increases the electron-electron repulsion within the valence shell. This increased repulsion causes the electron cloud to expand, pushing the electrons further apart. As a result, the nucleus holds the electrons less tightly, and the ionic radius of a negative ion (anion) becomes larger than that of the neutral atom.

Q.NO. 43 Compare the reactivity of Sodium and Magnesium with water.

Answer: Sodium (Group 1) is more reactive than Magnesium (Group 2). Sodium reacts vigorously with cold water to form sodium hydroxide and hydrogen gas ($2Na + 2H_2O \rightarrow 2NaOH + H_2$). Magnesium reacts very slowly with cold water but reacts vigorously with steam to form magnesium oxide and hydrogen

($Mg + H_2O_{(g)} \rightarrow MgO + H_2$).

Q.NO. 44 Classify Na_2O, Al_2O_3, and SO_2as acidic, basic, or amphoteric.

Answer:

- Na_2O(Sodium Oxide): It is a basic oxide because it dissolves in water to form a strong alkali (NaOH) and reacts with acids.
- Al_2O_3(Aluminum Oxide): It is an amphoteric oxide because it can react with both acids (forming $AlCl_3$) and bases (forming sodium aluminate).
- SO_2(Sulfur Dioxide): It is an acidic oxide because it reacts with water to form sulfurous acid and reacts with bases to form salts.

Q.NO. 45 Why does $AlCl_3$behave as an acidic chloride in water?

Answer: When aluminum chloride ($AlCl_3$) dissolves in water, the high charge density of the Al^{3+}ion causes strong hydration. The hydrated aluminum ion polarizes the water molecules attached to it, facilitating the release of H^+ions (hydrolysis). The reaction is: $AlCl_3 + 3H_2O \rightarrow Al(OH)_3 + 3HCl$. The production of H^+ions makes the solution acidic.

Q.NO. 46 What happens when sodium burns in excess oxygen?

Answer: When sodium burns in excess oxygen, it produces a pale yellow/golden flame and forms a mixture primarily containing sodium peroxide (Na_2O_2) along with some sodium oxide (Na_2O). The reaction for peroxide formation is: $2Na_{(s)} + O_{2(g)} \rightarrow 2Na_2O_{2(s)}$.

Q.NO. 47 Explain the variation of metallic character in the periodic table.

Answer: Metallic character, which is the tendency to lose electrons, decreases across a period from left to right due to increasing ionization energy and nuclear charge. It increases down a group because atomic size increases and ionization energy decreases, making it easier to lose valence electrons. Thus, Cesium is more metallic than Sodium.

Q.NO. 48 Define Amphoteric Oxides with a chemical equation.

Answer: Amphoteric oxides are oxides that possess both acidic and basic properties; they can neutralize both bases and acids. Aluminum oxide (Al_2O_3) is a classic example.

- Reaction with acid: $Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$
- Reaction with base: $Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$.

Q.NO. 49 What is the oxidation state of Phosphorus in $PCl_3$and PCl_5?

Answer: In PCl_3, phosphorus uses three valence electrons for bonding with chlorine, resulting in an oxidation state of +3. In PCl_5, phosphorus expands its octet using vacant d-orbitals and uses all five valence electrons for bonding, resulting in an oxidation state of +5.

Q.NO. 50 Why do chlorides of Group 1 and 2 elements form neutral solutions?

Answer: Chlorides of Group 1 (e.g., NaCl) and Group 2 (e.g., $MgCl_2$) are ionic. When dissolved in water, the metal cations (Na^+, Mg^{2+}) have relatively low charge densities compared to smaller, highly charged ions like Al^{3+}. They do not cause significant hydrolysis of water molecules to release protons. Therefore, the concentration of $H^+$and OH^-remains balanced (pH ≈7), resulting in a neutral solution.

Group 2: Atomic Structure

Q.NO. 51 Define Atomic Number and Nucleon Number.

Answer: The Atomic Number (Z), also called the proton number, is the number of protons in the nucleus of an atom. It defines the identity of the element. The Nucleon Number (A), or mass number, is the total number of protons and neutrons present in the nucleus. The number of neutrons is given by $N = A - Z$.

Q.NO. 52 Describe the behavior of protons, neutrons, and electrons in an electric field.

Answer: When passed through an electric field:

1. **Neutrons** are neutral and are not deflected, traveling in a straight line.
2. **Protons** are positively charged and deflect toward the negative plate.
3. **Electrons** are negatively charged and deflect toward the positive plate. Electrons are deflected significantly more than protons because their mass is much smaller ($1/1836$times the mass of a proton).

Q.NO. 53 What are isotopes? Give the neutron count for Cl-35 and Cl-37.

Answer: Isotopes are atoms of the same element that have the same atomic number (number of protons) but different nucleon numbers (mass numbers) due to a different number of neutrons.

- For Chlorine-35 ($Z = 17$): Neutrons = $35 - 17 = 18$.
- For Chlorine-37 ($Z = 17$): Neutrons = $37 - 17 = 20$.

Q.NO. 54 Explain the Principal Quantum Number (n).

Answer: The Principal Quantum Number (n) describes the main energy level or shell of an electron. It can have positive integer values ($n = 1,2,3, ...$). As 'n' increases, the size of the orbital and the energy of the electron increase. The maximum number of electrons in a shell is given by $2n^2$.

Q.NO. 55 What does the Azimuthal Quantum Number (l) determine?

Answer: The Azimuthal (or Angular Momentum) Quantum Number (l) determines the shape of the orbital. Its values range from 0to $n-1$.

- $l=0$: s-orbital (spherical)
- $l=1$: p-orbital (dumbbell)
- $l=2$: d-orbital (double dumbbell/cloverleaf)
- $l=3$: f-orbital (complicated)

Q.NO. 56 Describe the Magnetic Quantum Number (m).

Answer: The Magnetic Quantum Number (m) describes the spatial orientation of the orbital in a magnetic field. For a given value of l, $m$can take integer values from $-l$to $+l$, including zero. The total number of values ($2l+1$) represents the number of degenerate orbitals in a subshell (e.g., for p-subshell, $l=1$, so $m=-1,0,+1$, corresponding to p_x, p_y, p_z).

Q.NO. 57 State the Pauli Exclusion Principle.

Answer: Pauli's Exclusion Principle states that no two electrons in the same atom can have the same set of four quantum numbers. This implies that an orbital can hold a maximum of two electrons, and they must have opposite spins (represented as $+\frac{1}{2}$and $-\frac{1}{2}$or up and down arrows).

Q.NO. 58 State Hund's Rule of Maximum Multiplicity.

Answer: Hund's Rule states that when electrons occupy degenerate orbitals (orbitals of equal energy, like the three 2p orbitals), they will fill each orbital singly with parallel spins before pairing up. This distribution minimizes electron-electron repulsion and maximizes stability.

Q.NO. 59 Explain the Aufbau Principle.

Answer: The Aufbau (building up) Principle states that electrons fill atomic orbitals in order of increasing energy. Orbitals with lower energy are filled before those with higher energy. The energy order is determined by the $(n+l)$rule; orbitals with a lower $(n+l)$value have lower energy. If values are equal, the one with lower $n$is filled first (e.g., 4s fills before 3d).

Q.NO. 60 Sketch the shape of a p-orbital.

Answer: A p-orbital has a dumbbell shape consisting of two lobes located on opposite sides of the nucleus. There are three orientations aligned along the axes: p_x(along x-axis), p_y(along y-axis), and p_z(along z-axis). The node (region of zero electron probability) lies at the nucleus between the two lobes.

Q.NO. 61 Why is the electronic configuration of Chromium (Cr) anomalous?

Answer: The expected configuration of Chromium ($Z=24$) is $[Ar]3d^4 4s^2$. However, the actual configuration is $[Ar]3d^5 4s^1$. This occurs because a half-filled d-subshell ($3d^5$) offers extra stability. An electron shifts from the 4s orbital to the 3d orbital to achieve this stable half-filled state.

Q.NO. 62 Why is the electronic configuration of Copper (Cu) anomalous?

Answer: The expected configuration of Copper ($Z=29$) is $[Ar]3d^9 4s^2$. However, the actual configuration is $[Ar]3d^{10} 4s^1$. A completely filled d-subshell ($3d^{10}$) is more stable than a partially filled one. Therefore, an electron moves from the 4s orbital to complete the 3d subshell.

Q.NO. 63 What is a free radical? Give an example.

Answer: A free radical is a chemical species (atom or group of atoms) that possesses an unpaired electron. They are typically formed by the homolytic fission of bonds and are highly reactive. An example is the chlorine free radical ($Cl^{\cdot}$), formed when a Cl_2molecule is split by UV light.

Q.NO. 64 How are n-type semiconductors formed?

Answer: n-type semiconductors are formed by "doping" a pure semiconductor (Group 14, like Silicon) with a pentavalent impurity (Group 15, like Phosphorus). The phosphorus atom forms bonds with four silicon atoms, leaving one extra electron free. This free electron acts as a negative charge carrier, conducting electricity.

Q.NO. 65 How are p-type semiconductors formed?

Answer: p-type semiconductors are formed by doping a pure semiconductor (like Silicon) with a trivalent impurity (Group 13, like Aluminum). The aluminum atom can only form three bonds, creating a vacancy or "hole" in the lattice where an electron is missing. These holes act as positive charge carriers.

Q.NO. 66 Define Exothermic and Endothermic reactions with sign of ΔH.

Answer: An exothermic reaction is one that releases heat to the surroundings, resulting in a decrease in the enthalpy of the system ($\Delta H$is negative). An endothermic reaction absorbs heat from the surroundings, resulting in an increase in the enthalpy of the system ($\Delta H$is positive).

Q.NO. 67 State Hess's Law of Constant Heat Summation.

Answer: Hess's Law states that the total enthalpy change of a chemical reaction is the same, regardless of whether the reaction takes place in a single step or in several steps, provided the initial and final states are the same. It is a manifestation of the law of conservation of energy.

Q.NO. 68 Define Standard Enthalpy of Formation ($\Delta H_{f\circ}$).

Answer: The standard enthalpy of formation is the enthalpy change involved when one mole of a compound is formed from its constituent elements in their standard states under standard conditions ($25^{\circ}C$and $1atm$). For example, $C_{(s)} + 2H_{2(g)} \rightarrow CH_{4(g)}$has a specific $\Delta H_{f\circ}$.

Q.NO. 69 Define Standard Enthalpy of Combustion ($\Delta H_{c\circ}$).

Answer: The standard enthalpy of combustion is the enthalpy change occurring when one mole of a substance is completely burned in excess oxygen under standard conditions. This process is always exothermic. Example: $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$.

Q.NO. 70 Define Enthalpy of Atomization ($\Delta H_{at\circ}$).

Answer: The standard enthalpy of atomization is the energy required to form one mole of gaseous atoms from an element in its standard state. For example, forming gaseous hydrogen atoms from hydrogen gas: $\frac{1}{2}H_{2(g)} \rightarrow H_{(g)}$. This process is always endothermic.

Q.NO. 71 Define Standard Enthalpy of Neutralization.

Answer: The standard enthalpy of neutralization is the heat energy released when one mole of water is formed by the reaction of an acid and an alkali under standard conditions. For strong acids and bases, this value is approximately $-57.1kJmol^{-1}$. The net reaction is $H_{(aq)+} + OH_{(aq)-} \rightarrow H_2O_{(l)}$.

Q.NO. 72 Define Lattice Energy. Answer: Lattice energy is the enthalpy change that occurs when one mole of a solid ionic compound is formed from its constituent gaseous ions (e.g., $Na_{(g)+} + Cl_{(g)-} \rightarrow NaCl_{(s)}$). It is a measure of the strength of the ionic bonds in the lattice; a more negative lattice energy indicates a more stable solid.

Q.NO. 73 What is the Born-Haber Cycle?

Answer: The Born-Haber Cycle is a thermochemical cycle based on Hess's Law used to calculate lattice energy, which cannot be measured directly. It connects the lattice energy to other measurable quantities like enthalpy of formation, ionization energy, electron affinity, sublimation energy, and bond dissociation energy.

Q.NO. 74 Why is bond breaking an endothermic process?

Answer: Bond breaking requires energy to overcome the attractive forces holding the atoms together in a molecule. The system must absorb energy from the surroundings to separate the atoms. Since the heat content of the system increases, $\Delta H$is positive, making the process endothermic.

Q.NO. 75 Why is heat of neutralization of strong acid and strong base constant?

Answer: Strong acids and strong bases ionize completely in water. The neutralization reaction essentially involves the combination of H^+ions from the acid and OH^-ions from the base to form water ($H^+ + OH^- \rightarrow H_2O$). Since the ions involved are the same regardless of the specific strong acid or base used, the energy released per mole of water formed remains constant.

Q.NO. 76 Define Hydrogen Bonding.

Answer: Hydrogen bonding is a strong type of intermolecular force (dipole-dipole attraction) that occurs between a hydrogen atom covalently bonded to a highly electronegative atom (N, O, F) and a lone pair of electrons on a nearby electronegative atom. It is stronger than ordinary dipole-dipole forces but weaker than covalent bonds.

Q.NO. 77 Why does ice have a lower density than liquid water?

Answer: In ice, water molecules are arranged in a regular, open hexagonal lattice structure held together by hydrogen bonds. This arrangement creates significant empty spaces between the molecules. In liquid water,

the molecules are packed more closely together as the rigid structure breaks down. Therefore, ice occupies more volume for the same mass, resulting in lower density.

Q.NO. 78 What are Liquid Crystals?

Answer: Liquid crystals are a state of matter intermediate between a crystalline solid and an isotropic liquid. They possess the fluidity of liquids but retain the ordered molecular arrangement and optical properties of solids within a certain temperature range. They are used in display screens (LCDs) and temperature sensors.

Q.NO. 79 Define London Dispersion Forces.

Answer: London Dispersion Forces are weak intermolecular forces caused by temporary fluctuations in electron distribution within atoms or molecules. These fluctuations create temporary dipoles that induce dipoles in neighboring atoms. They are the only forces present in non-polar molecules and noble gases, and their strength increases with atomic size and polarizability.

Q.NO. 80 Why does evaporation cause cooling?

Answer: Evaporation is an endothermic process where high-energy molecules escape from the surface of the liquid into the vapor phase. The molecules remaining in the liquid have a lower average kinetic energy, which corresponds to a lower temperature. This reduction in the average energy of the liquid manifests as a cooling effect.

Q.NO. 81 State the Law of Mass Action.

Answer: Guldberg and Waage's Law of Mass Action states that the rate of a chemical reaction is directly proportional to the product of the active masses (molar concentrations) of the reactants, raised to the power of their stoichiometric coefficients in the balanced chemical equation.

Q.NO. 82 Define Chemical Equilibrium.

Answer: Chemical equilibrium is the state in a reversible reaction where the rate of the forward reaction equals the rate of the reverse reaction. At this stage, the concentrations of reactants and products remain constant over time, although the reaction continues at the molecular level (dynamic equilibrium).

Q.NO. 83 Write the Equilibrium Constant (K_c) expression for $aA + bB \rightleftharpoons cC + dD$.

Answer: The equilibrium constant (K_c) expression is the ratio of the product of concentrations of products to the product of concentrations of reactants, with each raised to the power of their coefficient: $K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$

Q.NO. 84 Differentiate between Homogeneous and Heterogeneous Equilibrium.

Answer: In a homogeneous equilibrium, all reactants and products are in the same physical phase (e.g., all gases or all in solution). In a heterogeneous equilibrium, the reactants and products are present in two or more different phases (e.g., a solid reacting with a gas to form a solid).

Q.NO. 85 What is the relationship between K_p and K_c?

Answer: The relationship is given by the equation $K_p = K_c(RT)^{\Delta n}$, where R is the gas constant, T is the absolute temperature, and Δn is the difference in the number of moles of gaseous products and gaseous reactants ($\Delta n = moles of gaseous products - moles of gaseous reactants$).

Q.NO. 86 How does temperature affect the equilibrium constant (K_c)?

Answer: The value of the equilibrium constant K_c is temperature-dependent. For an exothermic reaction, increasing the temperature decreases K_c (favors reactants). For an endothermic reaction, increasing the temperature increases K_c (favors products). Changes in concentration or pressure do not change the value of K_c, only the position of equilibrium.

Q.NO. 87 Why is chemical equilibrium dynamic?

Answer: Chemical equilibrium is dynamic because the reaction does not stop when equilibrium is reached. Reactants continue to convert to products and products back to reactants, but they do so at equal rates. Therefore, there is no net change in the composition of the mixture, but molecular activity persists.

Q.NO. 88 What is Le Chatelier's Principle?

Answer: Le Chatelier's Principle states that if a system at equilibrium is subjected to a change in conditions (concentration, pressure, or temperature), the system will shift its equilibrium position in a direction that opposes or minimizes the effect of that change to re-establish a new equilibrium.

Q.NO. 89 Effect of pressure on the synthesis of Ammonia ($N_2 + 3H_2 \rightleftharpoons 2NH_3$).
Answer: The synthesis of ammonia involves a decrease in the number of moles of gas (4 moles of reactants →2 moles of products). According to Le Chatelier's principle, increasing the pressure will shift the equilibrium toward the side with fewer moles of gas to reduce the pressure. Therefore, high pressure favors the formation of Ammonia (NH_3).
Q.NO. 90 Does a catalyst affect the equilibrium constant?
Answer: No, a catalyst does not affect the value of the equilibrium constant (K_c) or the position of equilibrium. It increases the rate of both the forward and reverse reactions equally, thereby allowing the system to reach the state of equilibrium faster, but it does not alter the final yield.
Q.NO. 91 What are "families" in the periodic table? Name two.
Answer: Families are groups of elements in the periodic table that share common chemical properties due to similar valence electron configurations. Examples include the **Alkali Metals** (Group 1), which are highly reactive metals, and the **Halogens** (Group 17), which are highly reactive non-metals forming salts.
Q.NO. 92 Explain the term "Blocks" in the periodic table.
Answer: The periodic table is divided into four blocks (s, p, d, f) based on the subshell being filled with the valence electrons.

- **s-block:** Groups 1 and 2.
- **p-block:** Groups 13 to 18.
- **d-block:** Transition metals (Groups 3-12).
- **f-block:** Lanthanides and Actinides (inner transition metals).

Q.NO. 93 What is the role of liquid crystals in medicine?
Answer: Liquid crystals are used in thermography for medical diagnosis. Since liquid crystals change color with temperature, they can be applied to the skin to detect areas of abnormal heat, which may indicate inflammation or tumors (e.g., detecting breast cancer or blockages in veins).
Q.NO. 94 Why are solids incompressible?
Answer: Solids are incompressible because their constituent particles (atoms, ions, or molecules) are packed closely together in a fixed, rigid arrangement with very little empty space between them. Applying pressure cannot force the particles significantly closer than they already are.
Q.NO. 95 Define Unit Cell.
Answer: A unit cell is the smallest repeating structural unit of a crystalline solid. When repeated in three dimensions, it generates the entire crystal lattice. The properties of the unit cell reflect the symmetry and structure of the whole crystal.
Q.NO. 96 What is hydration energy?
Answer: Hydration energy is the enthalpy change when one mole of gaseous ions dissolves in water to form an infinitely dilute solution. It is always exothermic because of the attraction between the ions and the polar water molecules. Smaller and highly charged ions have more negative (higher) hydration energies.
Q.NO. 97 Why is the second electron affinity of Oxygen endothermic?
Answer: (Repeated concept for clarity) The formation of the oxide ion O^{2-}involves adding an electron to the O^{-}ion. Since both the electron and the ion are negatively charged, there is significant electrostatic repulsion. Energy must be supplied to overcome this repulsion, making the process endothermic ($+798kJ/mol$).
Q.NO. 98 What is the shape of an ethene molecule?
Answer: In ethene (C_2H_4), the carbon atoms are sp^2hybridized. Each carbon forms three sigma bonds (two with hydrogen, one with the other carbon) arranged in a trigonal planar geometry with bond angles of $120°$. The remaining unhybridized p-orbitals overlap sideways to form a pi (π) bond.
Q.NO. 99 Define Isomerism.
Answer: Isomerism is the phenomenon where two or more compounds have the same molecular formula but different structural formulas (arrangements of atoms) and properties. Common types include chain isomerism, position isomerism, and functional group isomerism.
Q.NO. 100 What is the general formula of Alkanes and Alkenes? Answer:

- **Alkanes:** C_nH_{2n+2}. They are saturated hydrocarbons containing only single bonds.
- **Alkenes:** C_nH_{2n}. They are unsaturated hydrocarbons containing at least one carbon-carbon double bond.

SHORT QUESTION : 4

Question 1: How is the concept of mole derived from Avogadro's number (N_A)? **Answer:** According to the electrolytic method described, the Avogadro constant is the number of atoms, molecules, or ions in one mole of a substance. It is derived by calculating the total charge required to deposit 1 mole of a substance (1 Faraday $\approx 96500C$) divided by the charge of a single electron ($1.602 \times 10^{-19}C$).

$$N_A = \frac{96500C}{1.602 \times 10^{-19}C} \approx 6.02 \times 10^{23}$$

Thus, 1 mole is the amount of substance containing 6.02×10^{23}particles.

Question 2: Define: (a) Molar mass (b) Molar volume (c) Molar concentration.

Answer:

- **(a) Molar Mass:** The mass of one mole of a substance (element or compound), expressed in grams per mole ($g \cdot mol^{-1}$). It is numerically equal to the atomic, molecular, or formula mass.
- **(b) Molar Volume:** The volume occupied by one mole of any ideal gas at Standard Temperature and Pressure (STP), which is approximately $22.414dm^3$(or liters).
- **(c) Molar Concentration (Molarity):** The number of moles of solute dissolved per cubic decimeter (dm^3) or liter of solution. Formula: $M = \frac{moles of solute}{volume of solution(dm^3)}$.

Question 3: $39g$of Potassium (K) and $56g$of Iron (Fe) have equal number of atoms in them. Justify.
Answer: According to the Periodic Table:

- Atomic mass of Potassium (K) = $39amu$
- Atomic mass of Iron (Fe) = $56amu$Therefore, $39g$of Kequals 1 mole of Potassium, and $56g$of Feequals 1 mole of Iron. Since 1 mole of any element contains Avogadro's number of atoms (6.02×10^{23}), both samples contain the exact same number of atoms.

Question 4: $4g$of Helium (He), $17g$of Ammonia (NH_3), and $64g$of Sulfur Dioxide (SO_2) occupy $22.414dm^3$at STP despite different sizes. Explain.
Answer: These masses correspond to the molar masses of the respective gases:

- $He = 4g/mol$
- $NH_3 = 14 + 3(1) = 17g/mol$
- $SO_2 = 32 + 2(16) = 64g/mol$Since each sample represents exactly **1 mole** of gas, and according to Avogadro's Law, one mole of any ideal gas occupies the same volume ($V_m = 22.414dm^3$) at STP regardless of the size or mass of individual molecules.

Question 5: Differentiate theoretical and actual yields.
Answer:

- **Theoretical Yield:** The maximum amount of product that can be calculated from the balanced chemical equation, assuming 100% conversion of the limiting reactant.
- **Actual Yield:** The amount of product actually obtained from a chemical reaction performed in an experiment. It is almost always less than the theoretical yield.

Question 6: What are the factors mostly responsible for the low yield of products in chemical reactions?
Answer: Factors include:

1. **Side reactions:** Formation of by-products.
2. **Reversibility:** The reaction may reach equilibrium rather than going to completion.
3. **Mechanical loss:** Loss during filtration, distillation, or transferring of substances.
4. **Impurity:** Reactants may not be 100% pure.

Question 7: Calculate the molar mass of Potassium Permanganate ($KMnO_4$).
Answer: Using atomic masses from the periodic table: $K = 39$, $Mn = 55$, $O = 16$.

$$MolarMass = 39 + 55 + 4(16)$$
$$= 39 + 55 + 64$$
$$= 158g/mol$$

Question 8: How many molecules are present in $1.75g$ of Hydrogen Peroxide (H_2O_2)?
Answer:

1. Molar mass of $H_2O_2 = 2(1) + 2(16) = 34g/mol$.
2. Moles = $\frac{1.75}{34} \approx 0.0515mol$.
3. Molecules = $Moles \times N_A = 0.0515 \times 6.02 \times 10^{23}$.
$$\approx 3.10 \times 10^{22} molecules$$

Question 9: How many atoms are present in a $15g$ Gold (Au) ring?
Answer:

1. Atomic mass of Gold (Au) $\approx 197g/mol$.
2. Moles = $\frac{15}{197} \approx 0.076mol$.
3. Atoms = $0.076 \times 6.02 \times 10^{23}$.
$$\approx 4.58 \times 10^{22} atoms$$

Question 10: Determine the volume of $2.5moles$ of Chlorine molecules (Cl_2) at STP.
Answer: Using molar volume at STP ($22.414dm^3$):

$$V = n \times V_m$$
$$V = 2.5mol \times 22.414dm^3/mol$$
$$V = 56.035dm^3$$

Question 11: Calculate the molar mass of a gas which has a density of $1.97g/dm^3$ at STP.
Answer: Density (d) = $\frac{Mass}{Volume}$. For 1 mole, Volume = $22.414dm^3$.

$$MolarMass = Density \times MolarVolume$$
$$MolarMass = 1.97g/dm^3 \times 22.414dm^3/mol$$
$$\approx 44.16g/mol (likely CO_2)$$

Question 12: Calculate the molar concentration of a solution containing $7.9g$ of $KMnO_4$ dissolved in $1dm^3$ (Molar mass = 158).
Answer:

1. Moles of $KMnO_4 = \frac{Mass}{MolarMass} = \frac{7.9g}{158g/mol} = 0.05mol$.
2. Molarity (M) = $\frac{Moles}{Volume(dm^3)} = \frac{0.05mol}{1dm^3}$.
$$= 0.05M$$

Question 13: When $3.3moles$ of Nitrogen react with Hydrogen, how many moles of Hydrogen are consumed to form Ammonia?
Answer: Reaction: $N_2 + 3H_2 \rightarrow 2NH_3$. Ratio $N_2 : H_2$ is $1:3$.

$$Moles of H_2 = 3 \times Moles of N_2$$
$$= 3 \times 3.3 = 9.9 moles of H_2$$

Question 14: Define Stoichiometry. Mention its two basic assumptions.
Answer: Stoichiometry is the branch of chemistry that deals with the quantitative relationships between reactants and products in a chemical reaction. **Assumptions:**

1. All reactants are completely converted into products.
2. No side reactions occur.

Question 15: Calculate the mass of Aluminum (Al) needed to react with $32.0g$ of Iron(III) Oxide (Fe_2O_3) to produce Iron.

Answer: Reaction (Thermite): $2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$.

$$\text{Molar mass } Fe_2O_3 = 2(56) + 3(16) = 160g/mol.$$
$$\text{Moles } Fe_2O_3 = \frac{32.0}{160} = 0.2mol.$$
$$\text{Ratio } Al: Fe_2O_3 = 2:1. \text{ Moles } Al = 2 \times 0.2 = 0.4mol.$$
$$\text{Mass } Al = 0.4mol \times 27g/mol.$$
$$= 10.8g$$

Question 16: What is the mass of $10^{-3}mol$of Magnesium Sulfate ($MgSO_4$)?
Answer:

1. Molar mass $MgSO_4 = 24 + 32 + 4(16) = 120g/mol.$
2. Mass = $Moles \times MolarMass.$

$$Mass = 10^{-3} \times 120 = 0.12g$$

Question 17: How many molecules are in $1 \times 10^{-6}g$of isopentyl acetate ($C_7H_{14}O_2$)?
Answer:

1. Molar mass $C_7H_{14}O_2 = 7(12) + 14(1) + 2(16) = 84 + 14 + 32 = 130g/mol.$
2. Moles = $\frac{1\times10^{-6}}{130} \approx 7.69 \times 10^{-9}mol.$
3. Molecules = $7.69 \times 10^{-9} \times 6.02 \times 10^{23}$.

$$\approx 4.63 \times 10^{15}molecules$$

Question 18: Calculate the volume of CO_2produced at STP when $4.5dm^3$of Methane (CH_4) is burnt.
Answer: Reaction: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$. According to Avogadro's Law, volume ratios match mole ratios for gases. Ratio $CH_4: CO_2$is 1: 1. Therefore, $4.5dm^3$of CH_4produces $4.5dm^3$ of CO_2.
Question 19: Why does 1 mole of Uranium weigh more than 1 mole of Hydrogen?
Answer: Although 1 mole of both elements contains the same number of atoms (6.02×10^{23}), the mass of an individual Uranium atom (^{238}U) is approximately 238 times greater than that of a Hydrogen atom (^{1}H) due to the large number of protons and neutrons in the Uranium nucleus. Therefore, the total mass of the mole is proportionally higher.
Question 20: Define oxidation and reduction in terms of electron transfer with examples.
Answer:

- **Oxidation:** Loss of electrons. Example: $Zn \rightarrow Zn^{2+} + 2e^-$.
- **Reduction:** Gain of electrons. Example: $Cl_2 + 2e^- \rightarrow 2Cl^-$.

Question 21: Define oxidation and reduction in terms of change in oxidation number.
Answer:

- **Oxidation:** Increase in oxidation number (e.g., $Fe^{2+} \rightarrow Fe^{3+}$).
- **Reduction:** Decrease in oxidation number (e.g., $Mn^{+7} \rightarrow Mn^{+2}$).

Question 22: Determine the Oxidation Number of: (i) Crin $K_2Cr_2O_7$(ii) Nin NO_{3-}.
Answer: (i) Crin $K_2Cr_2O_7$:

$$2(\ \ +1) + 2(Cr) + 7(\ \ -2) = 0$$
$$2 + 2Cr - 14 = 0 \Rightarrow 2Cr = 12 \Rightarrow Cr = +6$$

(ii) Nin NO_{3-}:

$$N + 3(\ \ -2) = -1$$
$$N - 6 = -1 \Rightarrow N = +5$$

Question 23: Identify the oxidized and reduced species in: $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$.
Answer:

- **Oxidized:** Carbon (C) in CO. Oxidation state changes from $+2$to $+4$.
- **Reduced:** Iron (Fe) in Fe_2O_3. Oxidation state changes from $+3$to 0.

Question 24: Define a Disproportionation reaction with an example involving Chlorine.
Answer: A reaction where the same element is simultaneously oxidized and reduced.
Example (Chlorine in cold alkali):

$$Cl_2 + 2NaOH \rightarrow NaCl + NaClO + H_2O$$

Here, Cl_2(0) becomes Cl^-(-1, reduced) and OCl^-($+1$, oxidized).

Question 25: Balance the equation by oxidation number method: $MnO_2 + HCl \rightarrow MnCl_2 + H_2O + Cl_2$.

Answer:

1. Oxidation: $2Cl^- \rightarrow Cl_2 + 2e^-$(Change of 1 per Cl, 2 total).
2. Reduction: $Mn^{+4} + 2e^- \rightarrow Mn^{+2}$(Change of 2).
3. Electrons are balanced. Combine:

$$MnO_2 + 4HCl \rightarrow MnCl_2 + Cl_2 + 2H_2O$$

Question 26: Balance the equation: $K_2Cr_2O_7 + FeSO_4 + H_2SO_4 \rightarrow Cr_2(SO_4)_3 + Fe_2(SO_4)_3 + K_2SO_4 + H_2O$. **Answer:**

1. Reduction half: $Cr_2O_{72-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$.
2. Oxidation half: $Fe^{2+} \rightarrow Fe^{3+} + 1e^-$.
3. Multiply oxidation by 6 to balance electrons: $6Fe^{2+} \rightarrow 6Fe^{3+} + 6e^-$.
4. Combine:

$$K_2Cr_2O_7 + 6FeSO_4 + 7H_2SO_4 \rightarrow Cr_2(SO_4)_3 + 3Fe_2(SO_4)_3 + K_2SO_4 + 7H_2O$$

Question 27: Explain the construction and function of a Salt Bridge in a galvanic cell.

Answer:

A salt bridge is a U-shaped tube containing an inert electrolyte (like KClor KNO_3) in a gel (agar). Its functions are:

1. To connect the two half-cells and complete the electrical circuit.
2. To maintain electrical neutrality by allowing ions to migrate into the half-cells to neutralize the accumulating charges (cations to the cathode side, anions to the anode side).

Question 28: During electrolysis of aqueous $NaCl$, why is Nanot liberated at the cathode?

Answer: In aqueous solution, both $Na^+$and H^+ions migrate to the cathode. The reduction potential of hydrogen ($0.00V$) is much higher than that of sodium ($-2.71V$). Therefore, $H^+$is preferentially reduced to H_2gas, and Na^+remains in the solution.

Question 29: Define Standard Hydrogen Electrode (SHE). What is its potential?

Answer: The SHE is the reference electrode consisting of a platinum foil coated with platinum black, immersed in a $1.0MH^+$solution, with H_2gas at 1 atm bubbling over it. Its electrode potential is arbitrarily defined as $0.00V$ at all temperatures.

Question 30: How and why is an electrical double layer formed?

Answer: When a metal is dipped into a solution of its own ions, a potential difference develops. Metal atoms may oxidize into ions leaving electrons on the metal (making it negative), or ions from solution may reduce onto the metal (making it positive). A layer of oppositely charged ions forms in the solution adjacent to the electrode surface, creating an "electrical double layer" (Helmholtz layer).

Question 31: Why is the electrode potential of Cucalled reduction potential?

Answer: According to IUPAC convention, standard electrode potentials are reported as reduction potentials (tendency to gain electrons). For Copper:

$$Cu_{(aq)2+} + 2e^- \rightarrow Cu_{(s)}(E^\circ = +0.34V)$$

The positive value indicates $Cu^{2+}$is easier to reduce than H^+.

Question 32: Calculate the standard cell potential ($E_{cell\circ}$) for a Ni-Co cell.

Answer: Given standard potentials (approximate values):

- $Ni^{2+} + 2e^- \rightarrow Ni(E^\circ = -0.25V)$
- $Co^{2+} + 2e^- \rightarrow Co(E^\circ = -0.28V)$Nickel has the higher (less negative) potential, so it acts as the Cathode. $E_{cell\circ} = E_{cathode\circ} - E_{anode\circ}E_{cell\circ} = (\quad -0.25) - (\quad -0.28) = +0.03V$

Question 33: What is the effect of variation in ion concentration on the standard electrode potential?

Answer: According to the Nernst Equation, electrode potential depends on concentration. For a reduction $M^{n+} + ne^- \rightarrow M$:

- If $[M^{n+}]$increases ($> 1M$), the potential (E) becomes more positive.
- If $[M^{n+}]$decreases ($<1M$), the potential (E) becomes more negative.

Question 34: Arrange Ag, Cr, and Fein increasing order of their reducing powers based on $E°$values.
Answer: Reducing power is the ability to lose electrons (oxidize). The more negative the reduction potential, the stronger the reducing agent.

- $Cr^{3+}/Cr = -0.74V$(Most negative →Strongest Reducer)
- $Fe^{2+}/Fe = -0.44V$
- $Ag^{+}/Ag = +0.80V$(Most positive →Weakest Reducer) **Order (Increasing Reducing Power):** $Ag < Fe < Cr$

Question 35: Describe the redox changes during the electrolysis of molten $CuBr_2$.
Answer:

- **Anode (Oxidation):** $2Br_{(l)-} \rightarrow Br_{2(g)} + 2e^{-}$(Bromine gas evolved).
- **Cathode (Reduction):** $Cu_{(l)2+} + 2e^{-} \rightarrow Cu_{(s)}$(Copper metal deposited).

Question 36: What happens if the salt bridge is removed from a galvanic cell?
Answer: The circuit becomes incomplete. Charge accumulation (polarization) occurs immediately in both half-cells, creating an opposing potential that stops the flow of electrons. The voltage drops to zero and the reaction stops.
Question 37: Why is DC used instead of AC in electrolysis?
Answer: Direct Current (DC) ensures constant polarity at the electrodes. The cathode always attracts cations, and the anode always attracts anions. Alternating Current (AC) reverses polarity 50-60 times a second, causing ions to oscillate without being permanently discharged or deposited, thus preventing effective electrolysis.
Question 38: Define Electrolysis. How is it used to refine metals?
Answer: Electrolysis: The process of using electricity to drive a non-spontaneous chemical reaction.
Refining:

- **Anode:** Impure metal (e.g., Impure Copper). Oxidation occurs: $Cu \rightarrow Cu^{2+} + 2e^{-}$.
- **Cathode:** Thin strip of pure metal. Reduction occurs: $Cu^{2+} + 2e^{-} \rightarrow Cu$.
- Result: Pure metal builds up on the cathode, while impurities settle as "anode mud."

Question 39: Calculate the voltage of a cell having iron and copper electrodes ($E°Fe = -0.44V$, $E°Cu = +0.34V$).
Answer: The cell potential ($E_{cell°}$) is calculated by subtracting the anode potential from the cathode potential. Since copper has a more positive potential ($+0.34V$), it acts as the cathode (reduction), and iron acts as the anode (oxidation).

$$E_{cell°} = E_{cathode°} - E_{anode°}$$
$$E_{cell°} = (\quad +0.34V) - (\quad -0.44V)$$
$$E_{cell°} = +0.78V$$

Question 40: Explain why the electrode potential is measured relatively using SHE.
Answer: The potential of a single electrode cannot be measured in isolation because oxidation or reduction cannot occur without a complementary half-reaction. To measure the potential, the half-cell must be connected to a reference electrode to form a complete circuit. The Standard Hydrogen Electrode (SHE) is used as the universal reference because its potential is arbitrarily fixed at **0.00 V** at all temperatures, providing a standard baseline for comparison.
Question 41: Describe the movement of ions through the salt bridge.
Answer: The salt bridge contains an inert electrolyte (like KNO_3or KCl). Its function is to maintain electrical neutrality in the half-cells. Negative ions (anions) from the salt bridge move toward the anode compartment to neutralize the excess positive charge produced by oxidation. Positive ions (cations) move toward the cathode compartment to neutralize the excess negative charge left behind by the reduction of metal ions.
Question 43: Calculate the mass of 1.5 moles of $Ca(OH)_2$.
Answer:

1. Calculate the molar mass of Calcium Hydroxide ($Ca(OH)_2$):

$$Ca(40) + 2 \times [O(16) + H(1)] = 40 + 2(17) = 40 + 34 = 74g/mol$$

2. Multiply moles by molar mass:

$$Mass = 1.5mol \times 74g/mol = 111g$$

Question 44: Determine the number of molecules in 1.75 g of water.
Answer:

1. Molar mass of water (H_2O) = $2(1) + 16 = 18g/mol$.
2. Moles of water = $\frac{1.75}{18} \approx 0.0972mol$.
3. Number of molecules = $Moles \times N_A$

$$= 0.0972 \times 6.02 \times 10^{23} \approx 5.85 \times 10^{22} molecules$$

.

Question 45: How many atoms of Carbon are in 1 mole of isopentyl acetate?
Answer: The formula for isopentyl acetate is $C_7H_{14}O_2$. One mole of the molecule contains **7 moles** of Carbon atoms. Number of Carbon atoms = $7 \times 6.02 \times 10^{23} = 4.214 \times 10^{24} atoms$.
Question 46: Define Molar Volume. What is its value at STP?
Answer: Molar volume is the volume occupied by one mole of any ideal gas at Standard Temperature and Pressure (STP). Its value is approximately $22.414dm^3$ (or liters) per mole.
Question 47: State Avogadro's Law in terms of gas volumes.
Answer: Avogadro's Law states that equal volumes of all ideal gases, at the same temperature and pressure, contain the same number of molecules (or moles). This implies that the molar volume of all ideal gases is constant at STP ($22.414dm^3$).
Question 48: Calculate the mass of sodium hypochlorite produced from 2.25 moles of chlorine.
Answer: Reaction (Cold dilute alkali): $Cl_2 + 2NaOH \rightarrow NaCl + NaClO + H_2O$. From the stoichiometry, 1 mole of Cl_2 produces 1 mole of $NaClO$. So, 2.25 moles of Cl_2produce 2.25 moles of $NaClO$. Molar mass of $NaClO = 23(Na) + 35.5(Cl) + 16(O) = 74.5g/mol$.

$$Mass = 2.25mol \times 74.5g/mol = 167.625g$$

Question 49: Balance: $MnO_2 + HCl \rightarrow MnCl_2 + Cl_2 + H_2O$.
Answer: Using the oxidation number method or inspection:

$$MnO_2 + 4HCl \rightarrow MnCl_2 + Cl_2 + 2H_2O$$

(Mn is reduced from +4 to +2, and Cl is oxidized from -1 to 0) .
Question 50: Calculate the Ox. No. of Chromium in Cr_2O_{72-}.
Answer: Let the oxidation number of Cr be x. Oxygen is -2.

$$2x + 7(\quad -2) = -2$$
$$2x - 14 = -2$$
$$2x = +12 \Rightarrow x = +6$$

The oxidation number of Chromium is **+6**.
Question 51: List two reasons for the chemical inertness of N_2gas.
Answer:

1. **Triple Bond Strength:** The nitrogen molecule has a triple bond ($N \equiv N$) with a very high bond dissociation energy ($941kJ/mol$), making it extremely difficult to break.
2. **Non-polar Nature:** The molecule is non-polar, meaning it is not easily attacked by electrophiles or nucleophiles under normal conditions.

Question 52: Why is ammonia (NH_3) considered a weak base?
Answer: Ammonia is a weak base because it has a lone pair of electrons on the nitrogen atom, which can accept a proton (H^+) to form the ammonium ion (NH_{4+}). However, in aqueous solution, this ionization is incomplete, and the equilibrium lies far to the left, producing a low concentration of hydroxide ions compared to strong bases.
Question 53: Describe the laboratory preparation of ammonia gas.
Answer: In the laboratory, ammonia is prepared by heating an ammonium salt (like ammonium chloride) with a strong alkali (like calcium hydroxide).

$$2NH_4Cl + Ca(OH)_2 \rightarrow CaCl_2 + 2H_2O + 2NH_3(g)$$

The gas is dried by passing it over quicklime (CaO).

Question 54: Why is N_2gas used in food packaging?

Answer: Nitrogen is chemically inert and does not react with food components (fats, oils, proteins). It is used to displace oxygen in packaging (modified atmosphere packaging), which prevents oxidation (rancidity) and inhibits the growth of aerobic bacteria, thereby extending shelf life.

Question 55: Both CO and N_2have triple bonds. Why is CO more reactive?

Answer: Although both are isoelectronic with triple bonds, Carbon Monoxide (CO) is polar due to the electronegativity difference between C and O. This polarity makes the carbon atom susceptible to attack, allowing CO to act as a ligand in coordination compounds and as a reducing agent. $N_2$is non-polar and therefore inert.

Question 56: Why shouldn't a farmer treat a field with ammonium fertilizer and lime at the same time?

Answer: Lime (calcium hydroxide/oxide) is basic. If mixed with ammonium fertilizers (like ammonium nitrate or sulfate), a chemical reaction occurs that releases ammonia gas (NH_3) into the atmosphere. This results in the loss of nitrogen from the soil, reducing the effectiveness of the fertilizer.

$$2NH_{4+} + Ca(OH)_2 \rightarrow Ca^{2+} + 2NH_3(g) + 2H_2O$$

Question 57: Draw the structures of N_2O, NO, and $NO_2$and explain their bonding.

Answer:

- N_2O(Nitrous Oxide): Linear structure ($N \equiv N - O$). It is a resonance hybrid.
- NO(Nitric Oxide): Contains an odd number of electrons (paramagnetic). The bond order is 2.5 ($N = O$).
- NO_2(Nitrogen Dioxide): Bent molecule with an unpaired electron on Nitrogen. It often dimerizes to $N_2O_4$to pair the electron. It contains coordinate covalent bonding.

Question 58: Explain why sulfur is unreactive at room temperature.

Answer: Sulfur exists as stable S_8ring molecules at room temperature. The strong S-S covalent bonds within the ring structure require significant energy to break (high activation energy) before reaction can occur. Therefore, heating is usually required to initiate reactions.

Question 59: Determine the oxidation state of S in: $H_2SO_4$and S_2O_{32-}.

Answer:

- H_2SO_4: $2(\ \ +1) + S + 4(\ \ -2) = 0 \Rightarrow 2 + S - 8 = 0 \Rightarrow S = +6$.
- S_2O_{32-}(Thiosulfate): Average oxidation state is calculated as $2S + 3(\ \ -2) = -2 \Rightarrow 2S = 4 \Rightarrow S = +2$. (Note: structurally, the two sulfurs have different environments, but +2 is the average calculation) .

Question 60: Explain the involvement of d-orbitals in the variable oxidation states of sulfur.

Answer: Sulfur is in Period 3 and has empty 3d orbitals available. It can unpair electrons from its 3p and 3s orbitals and promote them to the 3d orbitals. This allows sulfur to expand its valency from 2 (ground state) to 4 (e.g., SO_2) and 6 (e.g., SO_3), exhibiting variable oxidation states.

Question 61: What is the role of sulfur in the vulcanization of rubber?

Answer: In vulcanization, sulfur is heated with natural rubber. The sulfur atoms form cross-links (disulfide bridges, $-S-S-$) between the long polymer chains of rubber. This cross-linking prevents the chains from sliding past each other, making the rubber harder, more elastic, and resistant to heat and wear.

Question 62: Give the reaction of Sulfur with NaOH.

Answer: Sulfur undergoes a disproportionation reaction when heated with concentrated sodium hydroxide solution:

$$3S + 6NaOH \rightarrow 2Na_2S + Na_2SO_3 + 3H_2O$$

It forms sodium sulfide (Na_2S) and sodium sulfite (Na_2SO_3).

Question 63: What is the function of $SO_2$and sulfite salts in food preservation?

Answer: Sulfur dioxide (SO_2) and sulfites act as antioxidants and antimicrobial agents. They inhibit the growth of bacteria and fungi in dried fruits, juices, and wine. They also prevent enzymatic browning by reducing the enzymes responsible for oxidation.

Question 64: Explain the oxidation of $SO_2$to SO_3.
Answer: The oxidation of sulfur dioxide (SO_2) to sulfur trioxide (SO_3) is a reversible, exothermic reaction. It requires a catalyst (Vanadium Pentoxide, V_2O_5) and an optimum temperature ($400 - 450°C$) to proceed at a practical rate (Contact Process).

$$2SO_2 + O_2 \rightleftharpoons 2SO_3(\Delta H = -197kJ/mol)$$

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Question 65: Why does sulfur show more oxidation states than oxygen?
Answer: Oxygen is in Period 2 and lacks d-orbitals in its valence shell, limiting it primarily to -2 (and rarely -1 or +2 with Fluorine). Sulfur is in Period 3 and possesses vacant 3d orbitals. This allows Sulfur to expand its octet and exhibit higher oxidation states like +4 and +6.
Question 66: Compare the reactions of sulfur with metals and non-metals.
Answer:

- **With Metals:** Sulfur acts as an oxidizing agent, reacting with heated metals to form sulfides (e.g., $Fe + S \rightarrow FeS$, $Cu + S \rightarrow CuS$).
- **With Non-metals:** Sulfur reacts with non-metals like oxygen and carbon. With oxygen, it forms oxides ($S + O_2 \rightarrow SO_2$). With carbon, it forms carbon disulfide ($C + 2S \rightarrow CS_2$) at high temperatures.

Group: Environmental Chemistry
Question 67: What does PAN stand for? Give its general formula.
Answer: PAN stands for **Peroxyacetyl Nitrate**. It is a component of photochemical smog. General Formula: $R - CO - O - O - NO_2$(where R is usually a methyl group, CH_3, for Peroxyacetyl Nitrate).
Question 68: Describe the role of a catalyst in a catalytic converter.
Answer: Catalytic converters use Platinum (Pt), Palladium (Pd), and Rhodium (Rh) catalysts to convert harmful exhaust gases into less harmful ones.

- **Oxidation:** Converts carbon monoxide (CO) and unburnt hydrocarbons (HC) into carbon dioxide (CO_2) and water (H_2O).
- **Reduction:** Converts nitrogen oxides (NO_x) into nitrogen gas (N_2) and oxygen (O_2).

Question 69: Write the conversion reaction carried out by Nitrobacter.
Answer: Nitrobacter bacteria play a crucial role in the nitrogen cycle (nitrification) by oxidizing nitrite ions (NO_{2-}) into nitrate ions (NO_{3-}).

$$2NO_{2-} + O_2 \rightarrow 2NO_{3-}$$

Question 70: Explain why denitrification occurs in anaerobic conditions.
Answer: Denitrification is performed by bacteria (like *Pseudomonas*) in waterlogged soils where oxygen is scarce (anaerobic). These bacteria use nitrate (NO_{3-}) as an electron acceptor instead of oxygen for respiration, converting it back into nitrogen gas (N_2) which is released into the atmosphere.
Question 71: Name the four major layers of the atmosphere and their approximate heights.
Answer:

1. **Troposphere:** 0 to 12 km
2. **Stratosphere:** 12 to 50 km
3. **Mesosphere:** 50 to 85 km
4. **Thermosphere:** 85 to 500+ km

Question 72: Why is the Troposphere the most important layer for life?
Answer: The troposphere contains 75% of the atmosphere's mass and almost all water vapor and dust. It is the layer where all weather phenomena occur and where the air we breathe resides. It maintains the temperature range suitable for life.
Question 73: Explain why the Stratosphere is warmer than the Mesosphere.
Answer: The stratosphere contains the Ozone Layer, which absorbs high-energy ultraviolet (UV) radiation from the sun. This absorption releases heat, causing the temperature to rise with altitude in this layer. The mesosphere lacks ozone and other heat-absorbing gases, so temperature decreases with altitude.
Question 74: Give the equations for the formation and depletion of ozone in the stratosphere.

Answer:

- **Formation:** $O_2 + h\nu(UV) \rightarrow 2O O + O_2 \rightarrow O_3$
- **Depletion (e.g., by CFCs):** $Cl + O_3 \rightarrow ClO + O_2 ClO + O \rightarrow Cl + O_2$(The Chlorine radical Cl is regenerated).

Question 75: Identify three major natural sources of air pollutants.

Answer:

1. **Volcanoes:** Release sulfur dioxide (SO_2), ash, and dust.
2. **Forest Fires:** Release carbon monoxide (CO), smoke, and particulate matter.
3. **Decaying Vegetation:** Releases methane (CH_4) and hydrogen sulfide (H_2S).

Question 76: Differentiate between Classical and Photochemical smog.

Answer:

- **Classical Smog (London Smog):** Reducing in nature. Formed by smoke + fog + sulfur dioxide (SO_2). Occurs in cool, humid climates.
- **Photochemical Smog (Los Angeles Smog):** Oxidizing in nature. Formed by the reaction of sunlight with nitrogen oxides (NO_x) and volatile organic compounds (VOCs). Occurs in warm, dry, sunny climates.

Question 77: Name the major Greenhouse Gases (GHGs). How do they cause global warming?

Answer: Major GHGs: Carbon Dioxide (CO_2), Methane (CH_4), Water Vapor (H_2O), Nitrous Oxide (N_2O), and Ozone (O_3). **Mechanism:** These gases allow short-wave solar radiation to pass through to the Earth's surface but absorb and re-emit the long-wave infrared radiation (heat) reflected back from the Earth. This traps heat in the lower atmosphere, raising the global temperature (Greenhouse Effect).

Question 78: What are the main chemical processes involved in the formation of Acid Rain?

Answer: Acid rain is formed when $SO_2$and NO_xreact with water and oxidants in the air.

- **Sulfuric Acid:** $2SO_2 + O_2 \rightarrow 2SO_3$then $SO_3 + H_2O \rightarrow H_2SO_4$.
- **Nitric Acid:** $4NO_2 + 2H_2O + O_2 \rightarrow 4HNO_3$.

Question 79: How do volatile organic compounds (VOCs) affect air quality?

Answer: VOCs (like hydrocarbons from fuel evaporation) react with nitrogen oxides (NO_x) in the presence of sunlight to form ground-level ozone (O_3) and Peroxyacetyl Nitrate (PAN), which are key components of photochemical smog. This causes respiratory issues and eye irritation.

Question 80: What is the impact of PAN on human health and plants?

Answer:

- **Humans:** PAN is a powerful lachrymator (causes tearing of eyes) and causes respiratory irritation and breathing difficulties.
- **Plants:** It damages plant tissues by inhibiting photosynthesis ("bronzing" of leaves) and attacks the enzyme systems, retarding growth.

Question 81: How does deforestation impact air quality and CO_2levels?

Answer: Trees act as "carbon sinks" by absorbing $CO_2$for photosynthesis. Deforestation reduces this capacity, leading to higher atmospheric CO_2levels (global warming). Burning cleared forests also releases stored carbon back as $CO_2$and smoke (particulates).

Question 82: Why is CO called a "silent killer"?

Answer: Carbon Monoxide (CO) is colorless, odorless, and tasteless. It binds to hemoglobin in blood 200 times more strongly than oxygen, forming carboxyhemoglobin. This prevents oxygen delivery to vital organs, causing suffocation and death without warning.

Question 83: Explain the principle of Winkler's method to determine Dissolved Oxygen (DO). **Answer:** The Winkler method uses the oxidizing property of dissolved oxygen. The DO in a water sample oxidizes Manganese(II) to Manganese(IV) in alkaline solution. Upon acidification in the presence of Iodide ions (I^-), the Mn(IV) liberates Iodine (I_2) quantitatively. The amount of Iodine released is titrated with Sodium Thiosulfate to calculate the original DO level.

Question 84: Define BOD. How is it carried out?

Answer: BOD (Biochemical Oxygen Demand): The amount of dissolved oxygen consumed by aerobic bacteria to decompose organic matter in a water sample over 5 days at $20°C$. **Procedure:** A water sample's initial DO is measured. It is then incubated in the dark at $20°C$ for 5 days. The final DO is measured. The difference ($DO_{initial} - DO_{final}$) is the BOD.

Question 85: Why are BOD and DO inversely related in polluted water?

Answer: High pollution (organic waste) stimulates the growth of bacteria that decompose the waste. These bacteria consume large amounts of oxygen during respiration. Therefore, as the waste load (BOD) increases, the available Dissolved Oxygen (DO) decreases rapidly.

Question 86: What are the anthropogenic sources of NO_x?

Answer: The primary man-made sources are high-temperature combustion processes in:

1. **Vehicles:** Internal combustion engines.
2. **Power Plants:** Burning of coal and oil.
3. **Industrial Boilers:** Chemical production (like nitric acid plants).

Question 87: Explain how NO_2 contributes to the brown color of smog.

Answer: Nitrogen Dioxide (NO_2) is a reddish-brown gas. When present in high concentrations in photochemical smog, it absorbs blue light from the solar spectrum and transmits/reflects yellow-brown light, giving the smog its characteristic brownish haze.

Question 88: What is the role of sunlight in photochemical smog formation?

Answer: Sunlight (UV radiation) provides the energy to initiate the breakdown of NO_2 into NO and atomic oxygen (O). This atomic oxygen reacts with O_2 to form ozone (O_3). Sunlight also drives the reactions between VOCs, NO_x, and O_3 to form secondary pollutants like PAN.

Question 89: How do particulates (Particulate Matter) cause lung damage?

Answer: Fine particulates ($PM_{2.5}$) are small enough to penetrate deep into the lungs (alveoli). They can irritate the respiratory tract, trigger asthma, cause inflammation, and may carry absorbed toxic/carcinogenic chemicals directly into the bloodstream.

Question 90: Write a balanced equation for the formation of sulfuric acid from atmospheric SO_3.

Answer:

$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

This reaction occurs in cloud droplets, contributing to acid rain.

Question 91: Explain the "Greenhouse Effect" in terms of heat absorption.

Answer: The Earth's surface absorbs solar energy and radiates it back as infrared (heat) energy. Greenhouse gases (like CO_2, CH_4) have molecular structures that vibrate when struck by this specific infrared frequency. They absorb the energy and re-emit it in all directions, including back toward Earth, preventing the heat from escaping into space and keeping the planet warm.

Question 92: What are the effects of SO_2 on plants and aquatic life?

Answer:

- **Plants:** Causes chlorosis (yellowing of leaves) by interfering with chlorophyll synthesis. It damages cell membranes leading to tissue death (necrosis).
- **Aquatic Life:** SO_2 causes acid rain, which lowers the pH of lakes and rivers. This acidity can kill fish eggs, leach toxic aluminum from soil into water (which clogs fish gills), and destroy the food chain (plankton).

Question 1: Calculate the mass of 0.5 moles of HCl.

Answer:

1. Molar mass of $HCl = 1(H) + 35.5(Cl) = 36.5g/mol.$
2. Mass = $Moles \times MolarMass$
   $= 0.5 \times 36.5$
   $= 18.25g$

Question 2: Find the moles in 20 g of NaOH.

Answer:

1. Molar mass of $NaOH = 23(Na) + 16(O) + 1(H) = 40g/mol.$
2. Moles = $\frac{Mass}{MolarMass}$

$$= \frac{20}{40}$$
$$= 0.5mol$$

Question 3: How many atoms are in 2.3 g of Sodium?
Answer:

1. Atomic mass of Sodium (Na) = $23g/mol.$
2. Moles = $\frac{2.3}{23} = 0.1mol.$
3. Number of atoms = $Moles \times Avogadro'sNumber$

$$= 0.1 \times 6.02 \times 10^{23}$$
$$= 6.02 \times 10^{22} atoms$$

Question 4: Calculate the molar mass of Juglone ($C_{10}H_6O_3$).
Answer:

$$MolarMass = 10(12) + 6(1) + 3(16)$$
$$= 120 + 6 + 48$$
$$= 174g/mol$$

Question 5: Calculate the moles in 0.87 g of Juglone ($C_{10}H_6O_3$).
Answer:

$$Moles = \frac{Mass}{MolarMass}$$
$$= \frac{0.87}{174}$$
$$= 0.005mol$$

Question 6: Determine the volume of 4.75 mol of methane at STP.
Answer: According to Avogadro's Law, 1 mole of gas at STP occupies $22.414dm^3$.

$$Volume = Moles \times 22.414$$
$$= 4.75 \times 22.414$$
$$= 106.47dm^3$$

Question 7: Calculate the molar mass of a gas with density $1.34g/dm^3$at STP.
Answer:

$$MolarMass = Density \times MolarVolume(STP)$$
$$= 1.34g/dm^3 \times 22.414dm^3/mol$$
$$= 30.03g/mol$$

Question 8: Calculate the molar concentration of $27.64gK_2CO_3$in $1dm^3$.
Answer:

1. Molar mass of $K_2CO_3 = 2(39) + 12 + 3(16) = 78 + 12 + 48 = 138g/mol.$
2. Moles = $\frac{27.64}{138} \approx 0.2mol.$
3. Molarity = $\frac{Moles}{Volume} = \frac{0.2}{1} = 0.2M.$

Question 9: How much iron can be produced from $160g$of Fe_2O_3?
Answer: Reaction (Thermite or Reduction): $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2.$

1. Molar mass of $Fe_2O_3 = 2(56) + 3(16) = 112 + 48 = 160g/mol.$
2. Moles of $Fe_2O_3 = \frac{160}{160} = 1mol.$
3. Mole ratio $Fe_2O_3: Fe = 1:2$. So, $1mol$of Fe_2O_3produces $2moles$of Fe.
4. Mass of $Fe = 2mol \times 56g/mol = 112g.$

Question 10: Define the Law of Conservation of Mass in stoichiometry.

Answer: The Law of Conservation of Mass states that matter is neither created nor destroyed during a chemical reaction. In stoichiometry, this means the total mass of the reactants must equal the total mass of the products. Consequently, chemical equations must be balanced to reflect the same number of atoms of each element on both sides.
+1

Question 11: What is a Mole-Mole relationship?

Answer: A Mole-Mole relationship is a quantitative relationship between the reactants and products in a balanced chemical equation. It allows chemists to calculate the number of moles of a product formed from a given number of moles of reactant, or vice versa, using stoichiometric coefficients.

Question 12: Write the mole ratio for the synthesis of ammonia.

Answer: Reaction: $N_2 + 3H_2 \rightarrow 2NH_3$. The mole ratios are:

- $1molN_2 : 3molH_2$
- $1molN_2 : 2molNH_3$
- $3molH_2 : 2molNH_3$

Question 13: Define the Law of Definite Proportions.

Answer: The Law of Definite Proportions states that a given chemical compound always contains its component elements in a fixed ratio by mass, regardless of its source or method of preparation. For example, water (H_2O) always contains Hydrogen and Oxygen in a 1:8 mass ratio.
+1

Question 14: Calculate the percentage yield if the actual yield is $2.5kg$and theoretical is $4.5kg$.

Answer:

$$= \left(\frac{2.5}{4.5}\right) \times 100$$
$$= 55.56$$

Question 15: Mention two physical processes that reduce actual yield.

Answer:

1. **Mechanical Loss:** Loss of substance during filtration, transferring between containers, or washing of precipitates.
2. **Incomplete Separation:** Difficulty in completely separating the desired product from the reaction mixture (e.g., during distillation or crystallization).

Question 16: Why is the sum of oxidation numbers zero in a neutral molecule?

Answer: In a neutral molecule, the total positive charge from the protons in all nuclei is exactly balanced by the total negative charge from all electrons. Since oxidation numbers represent the distribution of these electrons, the sum of the oxidation states of all atoms must equal the net charge of the molecule, which is zero.

Question 17: How does electronegativity affect the assignment of oxidation numbers?

Answer: In a covalent bond between two different atoms, the shared electrons are assigned to the more electronegative atom. The more electronegative atom gets a negative oxidation number (e.g., -1, -2), while the less electronegative atom gets a positive oxidation number. For example, in HCl, Chlorine is more electronegative than Hydrogen, so Cl is -1 and H is +1.

Question 18: Define Reducing Agent with an example.

Answer: A reducing agent (or reductant) is a substance that donates electrons to another species, thereby causing the reduction of that species. In the process, the reducing agent itself is oxidized.

Example: In $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$, Zinc (Zn) loses electrons and acts as the reducing agent.

Question 19: Explain why an oxidizing agent is itself reduced.

Answer: An oxidizing agent accepts electrons from another substance to enable oxidation. By accepting these electrons, its own oxidation state decreases (becomes more negative). Therefore, the process of accepting electrons is, by definition, reduction.

Question 20: Identify oxidizing/reducing agents in: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$.

Answer:

- **Carbon in CH_4:** Oxidation state changes from -4 to +4 (Loses electrons →Oxidized). Therefore, **Methane (CH_4) is the Reducing Agent**.
- **Oxygen in O_2:** Oxidation state changes from 0 to -2 (Gains electrons →Reduced). Therefore, **Oxygen (O_2) is the Oxidizing Agent**.

Question 21: Define an Electrolytic Cell.

Answer: An electrolytic cell is an electrochemical device that uses electrical energy from an external source (like a battery) to drive a non-spontaneous chemical reaction (electrolysis).

Question 22: Where does oxidation occur in an electrolytic cell?

Answer: Oxidation always occurs at the **Anode**. In an electrolytic cell, the anode is the positive electrode where anions (negative ions) lose electrons.

Question 23: Where does reduction occur in a galvanic cell?

Answer: Reduction always occurs at the **Cathode**. In a galvanic cell, the cathode is the positive electrode where cations (positive ions) gain electrons.

Question 24: Write the half-reaction of Cu^{2+}at the cathode.

Answer:

$$Cu_{(aq)2+} + 2e^- \rightarrow Cu_{(s)}$$

Question 25: Compare electrolytic and galvanic cells (3 points).

Answer:

1. **Energy Conversion:** Electrolytic converts electrical energy to chemical energy; Galvanic converts chemical energy to electrical energy.
2. **Spontaneity:** Electrolytic drives non-spontaneous reactions; Galvanic is driven by spontaneous redox reactions.
3. **Electrode Charge:** In Electrolytic, Anode is Positive (+); in Galvanic, Anode is Negative (-). (Cathode is reversed).

Question 26: Why do metals like K and Na have highly negative electrode potentials?

Answer: Metals like Potassium (K) and Sodium (Na) are highly electropositive, meaning they have a strong tendency to lose electrons to form positive ions ($M \rightarrow M^+ + e^-$). A highly negative standard reduction potential indicates a very low tendency to *gain* electrons (be reduced), which corresponds to a high tendency to *lose* them (be oxidized).

Question 27: What does a positive electrode potential indicate?

Answer: A positive standard electrode potential ($E^\circ > 0$) indicates that the species is easier to reduce than Hydrogen ions (H^+). It signifies a strong tendency to gain electrons. Elements with high positive potentials (like Fluorine, Gold) are strong oxidizing agents.

Question 28: Define Standard Conditions for measuring E°.

Answer: Standard electrode potentials are measured under the following conditions:

1. **Concentration:** $1.0 moldm^{-3}$for all aqueous ions.
2. **Pressure:** $1atm$($101.3kPa$) for any gases involved.
3. **Temperature:** Typically $298K$($25^\circ C$).

Question 29: Describe the structure of a Standard Hydrogen Electrode.

Answer:

The SHE consists of a platinum wire connected to a platinum foil covered with finely divided platinum black. This electrode is immersed in a solution containing hydrogen ions at $1M$concentration ($1MHCl$). Pure hydrogen gas at $1atm$pressure is continuously bubbled over the foil.

Question 30: Why is Cu^{2+}/Cumore "noble" than Zn^{2+}/Zn?

Answer: Copper has a positive reduction potential ($+0.34V$), indicating it prefers to remain in the metallic state and resist oxidation. Zinc has a negative reduction potential ($-0.76V$), indicating it easily loses electrons to form ions. Therefore, Copper is less reactive (more noble) than Zinc.

Question 31: How can E°values predict if a redox reaction is spontaneous?

Answer: For a redox reaction to be spontaneous, the calculated standard cell potential ($E_{cell\circ}$) must be positive.

$$E_{cell\circ} = E_{reduction\circ}(cathode) - E_{reduction\circ}(anode)$$

If $E_{cell\circ} > 0$, the reaction is feasible. If $E_{cell\circ} < 0$, the reaction is non-spontaneous.

Question 32: Calculate E_{cell}using $Ag^+/Ag(+0.80V)$ and $Zn^{2+}/Zn(-0.76V)$.

Answer:

- Cathode (Reduction): Ag^+(Higher potential, $+0.80V$)
- Anode (Oxidation): Zn(Lower potential, $-0.76V$)

$$E_{cell} = E_{cathode} - E_{anode}$$
$$= (+0.80V) - (-0.76V)$$
$$= +1.56V$$

Question 33: What is the role of a voltmeter in measuring cell potential?

Answer: A high-resistance voltmeter is connected in parallel between the two electrodes of the galvanic cell. It measures the potential difference (Voltage or EMF) generated by the electron flow from the anode to the cathode without drawing significant current, thus preventing polarization of the electrodes.

Question 34: Explain the feasibility of a reaction using $E°$values.

Answer: $E°$values allow us to determine if a specific oxidizing agent can react with a specific reducing agent. If the oxidizing agent has a more positive electrode potential than the reducing agent, the reaction is feasible. Essentially, electrons flow from the species with the more negative potential to the species with the more positive potential.

Question 35: Why does Zn react with $CuSO_4$but not with $NaCl$?

Answer:

- **With $CuSO_4$:** Zinc ($-0.76V$) is a stronger reducing agent than Copper ($+0.34V$), so it can displace Cu from solution ($Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$).
- **With $NaCl$:** Zinc ($-0.76V$) is a weaker reducing agent than Sodium ($-2.71V$). Zinc cannot force electrons onto Sodium ions (Na^+) to reduce them. Thus, no reaction occurs.

Question 36: Define Displacement Reaction.

Answer: A displacement reaction is a chemical reaction where a more reactive element displaces a less reactive element from its compound (usually in aqueous solution). General Form: $A + BC \rightarrow AC + B$(where A is more reactive than B).

Question 37: How is HNO_3formed from NOin the atmosphere?

Answer: Nitric oxide (NO) released into the atmosphere reacts with oxygen and water vapor:

1. $2NO + O_2 \rightarrow 2NO_2$(Nitrogen Dioxide)
2. $4NO_2 + 2H_2O + O_2 \rightarrow 4HNO_3$(Nitric Acid) This nitric acid contributes to acid rain.

Question 38: Name the compound added first in the Winkler method.

Answer: Manganese(II) Sulfate ($MnSO_4$) is added first, followed by an alkaline iodide solution (NaOH + KI).

Question 39: Why is starch used as an indicator in DO determination?

Answer: In the final step of Winkler's method, Iodine (I_2) is titrated against Sodium Thiosulfate. Starch is used as an indicator because it forms a deep blue-black complex with trace amounts of Iodine. The disappearance of this blue color marks the precise endpoint of the titration.

+1

Question 40: Explain the inverse relationship between DO and water pollution.

Answer: Water pollution often involves organic waste. Bacteria and other microorganisms decompose this organic matter using dissolved oxygen (DO) for respiration. As pollution increases (higher organic load), the microbial population explodes and consumes more oxygen, leading to a rapid decrease in DO levels.

Question 41: What is the significance of DO for aquatic life?

Answer: Dissolved Oxygen is essential for the respiration of aquatic organisms like fish, invertebrates, and plants. If DO levels fall below critical limits (typically $< 4-5mg/L$), aquatic life becomes stressed, leading to suffocation ("fish kills") and loss of biodiversity.

Question 42: What is the bond enthalpy of the $N \equiv N$bond?

Answer: The bond enthalpy (bond dissociation energy) of the nitrogen-nitrogen triple bond ($N \equiv N$) is very high, approximately $941kJ/mol$ ($225kcal/mol$). This high energy requirement for bond breaking is responsible for nitrogen's inertness.

Question 43: Why is nitrogen gas used in preserving edible oils?

Answer: Nitrogen is chemically inert and does not react with lipids. Flushing food containers with nitrogen displaces oxygen, preventing oxidative rancidity (oxidation of fats/oils) which causes bad odor and taste. It creates a protective atmosphere to maintain freshness.

Question 44: Explain the non-polarity of the nitrogen molecule.

Answer: The nitrogen molecule (N_2) consists of two identical nitrogen atoms bonded together. Since the electronegativity difference between the two atoms is zero, the electron pair is shared equally. This results in a non-polar covalent bond and a non-polar molecule with no dipole moment.

Question 45: Write the equation for the Haber process.

Answer: The Haber process for the industrial synthesis of ammonia is:

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)} (\Delta H = -92.4kJ/mol)$$

It requires an iron catalyst, high pressure ($200atm$), and optimal temperature ($450°C$).

Question 46: What is the $K_b$value of ammonia?

Answer: The base dissociation constant (K_b) for ammonia at $25°C$is approximately $1.8 \times 10^{-5}mol/dm^3$. This low value confirms that ammonia is a weak base.

Question 47: Why does red litmus paper turn blue in ammonia gas?

Answer: Ammonia gas is extremely soluble in water. When it comes into contact with the moisture on the damp red litmus paper, it dissolves to form ammonium hydroxide (NH_4OH), which dissociates to release hydroxide ions (OH^-).

$$NH_3 + H_2O \rightarrow NH_{4+} + OH^-$$

The alkaline OH^-ions turn the red litmus paper blue.

Question 48: How is ammonia displaced from ammonium salts?

Answer: Ammonia is displaced from ammonium salts by heating them with a strong base (alkali) like Calcium Hydroxide ($Ca(OH)_2$) or Sodium Hydroxide ($NaOH$). Reaction:

$$2NH_4Cl + Ca(OH)_2 \rightarrow CaCl_2 + 2H_2O + 2NH_3 \uparrow$$

Question 49: Name two neutral oxides of nitrogen.

Answer:

1. **Nitrous Oxide** (N_2O) - also known as laughing gas.
2. **Nitric Oxide** (NO). (Note: Nitrogen dioxide, NO_2, and Dinitrogen pentoxide, N_2O_5, are acidic).

Question 93 What is the common name of N_2O?

Answer: The common name of Dinitrogen Oxide (N_2O) is **Laughing Gas**. It is a colorless gas with a faint, sweet smell. It is used as a mild anesthetic in dentistry and minor surgeries because, when inhaled in small quantities, it produces hysterical laughter.

Question 94 Which nitrogen oxide is paramagnetic?

Answer: Nitric Oxide (NO) and **Nitrogen Dioxide (** NO_2**)** are paramagnetic. This property is due to the presence of an odd number of valence electrons (an unpaired electron) in their molecular structures. For example, NO has 11 valence electrons, leaving one unpaired, which causes it to be attracted by a magnetic field.

Question 95 Which nitrogen oxide is used in rocket propellants?

Answer: Dinitrogen Tetroxide (N_2O_4) is used as an oxidizer in rocket propellants, often in combination with hydrazine fuel. It is preferred because it can be stored as a liquid at room temperature and reacts hypergolically (spontaneously) with hydrazine.

Question 96 List two natural sources of NOx.

Answer:

1. **Lightning:** The high energy of lightning strikes provides sufficient heat for atmospheric nitrogen (N_2) and oxygen (O_2) to combine and form nitric oxide (NO).
2. **Bacterial Action:** Denitrifying bacteria in the soil decompose nitrogenous compounds (nitrates) under anaerobic conditions, releasing nitrogen oxides into the atmosphere.

Question 97 Define Anthropogenic sources of pollution.

Answer: Anthropogenic sources refer to pollution sources originating from **human activities**. These include emissions from industrial factories, combustion of fossil fuels in vehicles and power plants, agricultural activities (fertilizers), and deforestation. These are distinct from natural sources like volcanoes.

Question 98 What gives the brown color to urban smog?

Answer: The characteristic brown haze of photochemical smog is primarily due to the presence of **Nitrogen Dioxide (NO_2)**. $NO_2$is a reddish-brown gas that absorbs blue light from the solar spectrum and transmits yellow-brown light, giving the polluted air its distinct color.

Question 99 Explain the role of sunlight in smog formation.

Answer: Sunlight (specifically UV radiation) acts as the initiator for photochemical smog. It dissociates Nitrogen Dioxide (NO_2) into Nitric Oxide (NO) and atomic Oxygen (O). This atomic oxygen is highly reactive and combines with molecular oxygen (O_2) to form Ozone (O_3), a key component of smog.

$$NO_2 + h\nu \rightarrow NO + O$$

Question 100 Name the metals used in a catalytic converter.

Answer: Catalytic converters use noble metals such as **Platinum (Pt)**, **Palladium (Pd)**, and **Rhodium (Rh)** coated on a ceramic honeycomb structure. These metals catalyze the oxidation of toxic CO and hydrocarbons to $CO_2$and H_2O, and the reduction of $NO_x$to N_2.

Question 101 Define Nitrification and name the bacteria involved.

Answer: Nitrification is the biological oxidation of ammonia or ammonium ions (NH_{4+}) into nitrite (NO_{2-}) and then into nitrate (NO_{3-}).

- Bacteria converting Ammonia to Nitrite: **Nitrosomonas**.
- Bacteria converting Nitrite to Nitrate: **Nitrobacter**.

Question 102 Write the conversion reaction of NH4+ to NO_2-.

Answer: This reaction is the first step of nitrification carried out by *Nitrosomonas* bacteria:

$$2NH_{4+} + 3O_2 \rightarrow 2NO_{2-} + 4H^+ + 2H_2O + Energy$$

Question 103 What is the atomic number and physical state of sulfur?

Answer:

- **Atomic Number:** 16.
- **Physical State:** At room temperature, sulfur is a **yellow solid**. It exists as brittle, crystalline molecular solids, predominantly in the rhombic form.

Question 104 Define Catenation in sulfur.

Answer: Catenation is the ability of atoms of the same element to link together to form long chains or rings. Sulfur shows a strong tendency for catenation, forming stable homo-atomic ring structures, most notably the puckered S_8ring (crown shape), due to the strength of the S-S single bond.

+1

Question 105 Which is the most stable molecular form of sulfur?

Answer: The most stable molecular form of sulfur at room temperature is the **Rhombic Sulfur** (α-Sulfur). It consists of S_8molecules packed in a crystal lattice. It is stable below $95.5°C$.

Question 106 Compare oxidation states of sulfur in SO_2 and H_2SO_4.

Answer:

- **In SO_2(Sulfur Dioxide):** Oxygen is -2. $S + 2(\ \ -2) = 0 \Rightarrow S = +4$.
- **In H_2SO_4(Sulfuric Acid):** Hydrogen is +1, Oxygen is -2. $2(\ \ +1) + S + 4(\ \ -2) = 0 \Rightarrow S = +6$. Sulfur is more oxidized in sulfuric acid.

Question 107 Explain the stability of the +6 oxidation state in SO_3.

Answer: In Sulfur Trioxide (SO_3), sulfur is in the +6 oxidation state, using all six valence electrons for bonding (involving d-orbital hybridization). This state is stable because the small, highly electronegative oxygen atoms effectively shield the central sulfur atom and satisfy its bonding requirements, forming a stable resonance hybrid structure.

Question 108 What happens when SO2 is dissolved in water?

Answer:

Sulfur dioxide is highly soluble in water and reacts to form **Sulfurous Acid** (H_2SO_3), which is a weak, unstable acid.

$$SO_2(g) + H_2O(l) \rightleftharpoons H_2SO_3(aq)$$

Question 109 Why is SCl2 unstable in moist air?

Answer: Sulfur dichloride (SCl_2) is unstable in moist air because it undergoes hydrolysis. It reacts with the moisture (water vapor) to produce hydrogen chloride fumes (HCl) and sulfur dioxide or thiosulfuric acid products, decomposing the molecule.

Question 110 How is sulfur used in the bleaching of paper?

Answer: Sulfur dioxide (SO_2) is used as a bleaching agent for delicate materials like paper pulp, wool, and silk. It bleaches by **reduction** (removing oxygen from the color pigment or adding hydrogen). However, the bleaching is temporary because atmospheric oxygen eventually re-oxidizes the reduced pigment, restoring the color.

Question 111 Describe the role of sulfur in making dyes and explosives.

Answer: Sulfur, primarily in the form of Sulfuric Acid (H_2SO_4), is essential in these industries.

- **Dyes:** It is used for the sulfonation of organic compounds to make them water-soluble dyes.
- **Explosives:** It acts as a dehydrating agent in the nitration reactions required to produce explosives like TNT (Trinitrotoluene) and Nitroglycerin.

Question 112 How is sulfuric acid used in the battery industry?

Answer: Sulfuric acid acts as the **electrolyte** in lead-acid storage batteries (car batteries). It provides the sulfate ions (SO_{42-}) and hydrogen ions necessary for the electrochemical reactions at the lead anode and lead dioxide cathode that store and release electrical energy.

Question 113 What is the physical appearance of sulfur?

Answer: Pure sulfur is a **pale yellow, brittle solid** at room temperature. It is tasteless and odorless (though often associated with the smell of its compounds like H_2S). It is insoluble in water but soluble in organic solvents like carbon disulfide (CS_2).

Question 114 Explain the role of sulfur in the vulcanization of rubber.

Answer: (Repeated concept) Vulcanization involves heating raw rubber with sulfur. The sulfur atoms form **cross-links** (disulfide bridges) between the polymer chains of the rubber. This process transforms soft, sticky raw rubber into a harder, more durable material that retains its elasticity over a wider range of temperatures.

Question 115 Which atmospheric layer contains the ozone layer?

Answer: The **Stratosphere** contains the ozone layer. The highest concentration of ozone (O_3) is found at an altitude of approximately 25 to 30 km. This layer shields the Earth from harmful ultraviolet (UV) radiation.

Question 116 Why is the Mesosphere the coldest layer?

Answer: The Mesosphere (50–85 km) is the coldest layer (temperatures drop to $-93°C$) because it lacks significant concentrations of ozone or water vapor to absorb solar radiation. Unlike the stratosphere below it (which is warmed by ozone absorbing UV), there is no internal heat source in the mesosphere.

Question 117 Describe how UV radiation contributes to ozone formation.

Answer: High-energy UV radiation (wavelength < 242 nm) strikes oxygen molecules (O_2) in the stratosphere, splitting them into two atomic oxygen atoms (O). These highly reactive oxygen atoms then collide with other molecular oxygen (O_2) molecules to form ozone (O_3).

$$O_2 + h\nu \rightarrow 2O$$
$$O + O_2 \rightarrow O_3$$

Question 118 Name two primary air pollutants.

Answer: Primary pollutants are emitted directly from sources. Examples include:

1. **Carbon Monoxide (CO):** From incomplete combustion.
2. **Sulfur Dioxide (SO_2):** From burning sulfur-containing coal or oil.

Question 119 Give an example of a secondary pollutant.

Answer: Ozone (O_3) (specifically ground-level ozone) is a classic secondary pollutant. It is not emitted directly but is formed by the chemical reaction of primary pollutants like $NO_x$and VOCs in the presence of sunlight. **Peroxyacetyl Nitrate (PAN)** is another example.

Question 120 How are secondary pollutants formed?

Answer: Secondary pollutants are formed in the atmosphere through chemical reactions between primary pollutants and normal atmospheric constituents, usually powered by energy from sunlight (photochemical reactions). For example, Sulfuric acid (H_2SO_4) is formed when SO_2reacts with water and oxygen.

Question 121 Why are aerosols harmful?

Answer: Aerosols (minute liquid or solid particles suspended in air) are harmful because:

1. They scatter sunlight, reducing visibility and altering the Earth's heat balance (cooling effect).
2. When inhaled, they penetrate deep into the respiratory system, causing inflammation, asthma, and other pulmonary diseases.

Question 122 Why is fossil fuel combustion a major cause of air pollution?

Answer: Fossil fuels (coal, oil, natural gas) contain carbon, hydrogen, sulfur, and nitrogen impurities. When burned, they release vast quantities of **Carbon Dioxide (CO_2)** (greenhouse gas), **Sulfur Dioxide (SO_2)**, **Nitrogen Oxides (NO_x)**, and **Particulate Matter** into the atmosphere, which are the primary drivers of smog, acid rain, and global warming.

Question 123 How does agriculture release ammonia into the air?

Answer: Agriculture releases Ammonia (NH_3) primarily through the decomposition of livestock waste (manure and urine) and the application of nitrogenous fertilizers (like urea) to soils. The volatile ammonia gas escapes from the soil and waste lagoons into the atmosphere.

Question 124 Compare industrial and domestic sources of air pollutants.

Answer:

- **Industrial Sources:** Include power plants, factories, and refineries. They emit large-scale pollutants like SO_2, metallic dust, and chemical solvents.
- **Domestic Sources:** Include household heating, cooking, and waste burning. They emit pollutants on a smaller, localized scale, such as $CO$from stoves and smoke from burning trash.

Question 125 What is the impact of urbanization on air quality?

Answer: Urbanization concentrates population, vehicles, and industries in small areas. This leads to a high density of emission sources, resulting in elevated levels of NO_x, CO, and particulates. The "Urban Heat Island" effect also accelerates the formation of photochemical smog.

Answer: Global warming is the long-term rise in the average temperature of the Earth's climate system. It is primarily caused by the **Greenhouse Effect**, where increased concentrations of greenhouse gases (like CO_2) from human activities trap heat in the atmosphere that would otherwise escape into space.

Question 127 Which gas has the highest heat-absorbing potential?

Answer: Per molecule, **Chlorofluorocarbons (CFCs)** and **Methane (CH_4)** have a much higher heat-absorbing potential (Global Warming Potential) than Carbon Dioxide. However, $CO_2$is the most significant contributor overall due to its sheer abundance.

Question 128 How do particulates cause cardiovascular issues?

Answer: Fine particulates ($PM_{2.5}$) are small enough to pass from the lungs into the bloodstream. Once in the blood, they cause systemic inflammation, oxidative stress, and can lead to plaque buildup in arteries (atherosclerosis), increasing the risk of heart attacks and strokes.

Question 129 What is the impact of PAN on plant life?

Answer: Peroxyacetyl Nitrate (PAN) is highly toxic to plants. It damages the waxy coating of leaves and attacks the cellular structure, leading to **"bronzing"** or "glazing" of the leaves. It also inhibits photosynthesis, reducing plant growth and crop yields.

Question 130 Explain the role of NO2 in acid rain.

Answer: Nitrogen Dioxide (NO_2) reacts with water vapor and oxygen in the atmosphere to form **Nitric Acid (HNO_3)**.

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

This strong acid dissolves in rain droplets, lowering their pH and resulting in acid rain.

Question 131 How can smog formation be reduced in cities?

Answer:

1. **Catalytic Converters:** Installing them in vehicles to reduce $NO_x$and VOC emissions.
2. **Public Transport:** Reducing the number of private cars on the road.
3. **Fuel Standards:** Using cleaner fuels (unleaded, low-sulfur).
4. **Regulations:** Controlling industrial emissions and open burning.

Question 132 Define the Greenhouse Effect.

Answer: The Greenhouse Effect is the natural process by which the Earth's atmosphere traps energy from the Sun. Greenhouse gases allow short-wave solar radiation to reach the surface but absorb the long-wave infrared radiation reflected back by the Earth. This trapped heat maintains the planet's temperature at a level suitable for life.

Question 133 Name one heavy metal pollutant affecting human health.

Answer: Lead (Pb) is a significant heavy metal pollutant, primarily from old leaded fuels and industrial paints. It is a neurotoxin that accumulates in the body, causing brain damage, developmental delays in children, and kidney failure.

Question 134 Why are motor vehicles considered mobile pollution sources? Answer: They are classified as mobile sources because they move from place to place while emitting pollutants. This mobility makes their pollution difficult to contain in one specific area, as they distribute exhaust gases ($CO, NO_x, VOCs$) across vast urban and rural networks.

Question 135 Calculate moles of HCl neutralized by 2.1 g of baking soda ($NaHCO_3$).

Answer: Reaction: $NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$(1:1 mole ratio).

1. Molar mass $NaHCO_3 = 23 + 1 + 12 + 3(16) = 84g/mol$.
2. Moles $NaHCO_3 = \frac{2.1g}{84g/mol} = 0.025mol$.
3. Since ratio is 1:1, Moles HClneutralized = **0.025 mol**.

Question 136 What volume of H2 at STP is produced from 7.0 g of Iron and H2SO4?

Answer: Reaction: $Fe + H_2SO_4 \rightarrow FeSO_4 + H_2$.

Molar mass $Fe = 56g/mol$.

Moles $Fe = \frac{7.0}{56} = 0.125mol$.

Ratio $Fe: H_2$is 1:1, so Moles $H_2 = 0.125mol$.

Volume at STP = $0.125 \times 22.414dm^3$.

$= 2.80dm^3$

Question 137 Differentiate between Actual Yield and Percentage Yield.

Answer:

- **Actual Yield:** The amount of product experimentally obtained from a reaction. It is a mass quantity (e.g., grams).

- **Percentage Yield:** The efficiency of the reaction, calculated as the ratio of actual yield to theoretical yield multiplied by 100. It is a dimensionless percentage.

Question 138 Why must electrons lost equal electrons gained in balancing redox reactions?
Answer: This requirement stems from the **Law of Conservation of Charge**. In any chemical process, electric charge cannot be created or destroyed. Therefore, the total number of electrons released by the reducing agent (oxidation) must be exactly accepted by the oxidizing agent (reduction) to maintain electrical neutrality.
Question 139 Define Oxidation Number.
Answer: Oxidation number (or state) is the apparent charge assigned to an atom in a compound, assuming that the bond is completely ionic. It represents the number of electrons lost, gained, or shared by the atom to form the bond.
Question 140 Find the Oxidation Number of Mn in KMnO4.
Answer:
$K$is Group 1 (+1), $O$is usually (-2).

$$1(\ \ +1) + Mn + 4(\ \ -2) = 0$$
$$1 + Mn - 8 = 0$$
$$Mn = +7$$

Question 141 Why is the mole used by chemists as a unit?
Answer: Atoms and molecules are incredibly small and react in vast numbers. The mole (6.02×10^{23}particles) provides a bridge between the microscopic world of atoms and the macroscopic world of grams. It allows chemists to weigh substances in the lab (grams) and know they are working with specific ratios of atoms.
+1
Question 142 How is molar mass related to the periodic table?
Answer: The molar mass of an element (in grams per mole) is numerically equivalent to its **relative atomic mass** (atomic weight) listed in the periodic table. For example, Carbon's atomic mass is 12.01 amu, so its molar mass is 12.01 g/mol.
Question 143 Do 1 mole of glucose and 1 mole of water contain equal molecules?
Answer: Yes. By definition, 1 mole of *any* substance contains Avogadro's number of particles (6.02×10^{23}). Therefore, 1 mole of glucose ($C_6H_{12}O_6$) and 1 mole of water (H_2O) both contain exactly 6.02×10^{23}molecules, despite having different masses and sizes.+1
Question 144 What happens to a cation during electrolysis?
Answer: During electrolysis, a cation (positively charged ion) is attracted to the **Cathode** (negative electrode). Upon reaching the cathode, the cation **gains electrons** (is reduced) to become a neutral atom or molecule. Example: $Cu^{2+} + 2e^- \rightarrow Cu$.

LONG QUESTION :5

1. Explain the four quantum numbers with their names, symbols, possible values, and significance. Also write their formulae where applicable.
2. Draw the shapes of s, p, and d-orbitals. Justify these by keeping in view the azimuthal and magnetic quantum numbers.
3. Describe Aufbau Principle, Pauli Exclusion Principle, and Hund's Rule with diagrams.
4. Discuss the variation of first ionization energies across a period and down a group with reasons.
5. Explain the Bohr and Schrodinger models of the atom. How did the concept of energy levels change with quantum mechanics?
6. Explain the method to determine the number of protons, neutrons, and electrons in atoms and ions, with the help of suitable examples.
7. Hydrogen bonding is present in H_2O, NH_3, HF, $(CH_3)_2CO$ and $CHCl_3$ molecules. Sketch structures and discuss briefly.

8. Compare the hydrogen bonding in water, ammonia, and hydrogen fluoride. How does this bonding affect their physical states and boiling points?
9. What are London dispersion forces? Give examples, and discuss the factors affecting these forces.
10. Discuss the structural changes when water turns into ice. Justify the empty spaces in its crystals as compared to H_2O at 4 °C and lower density of ice.
11. Discuss the factors that affect the boiling point of liquids with examples. Explain how intermolecular forces influence the boiling points of water and ethanol.
12. What are liquid crystals? How are they different from solids and liquids? Explain with their properties and uses.
13. Compare crystalline and amorphous solids.
14. Describe the following properties of crystalline solids: (i) Geometrical shape (ii) Melting point (iii) Cleavage plane (iv) Habit of a crystal.
15. Explain the Ideal Gas Equation ($PV = nRT$).

LONG QUESTION :6

1. What are the postulates of VSEPR model? Discuss the structures of the following species with reference to this theory: (i) CH_4 (ii) NH_3 (iii) H_3O^+ (iv) PCl_5 (v) SO_2 (vi) SF_6
2. Explain the orbital hybridization for CH_4, NH_3, BF_3, and $BeCl_2$.
3. Describe the molecular orbital diagram of oxygen (O_2). Explain its bond order and magnetic behavior.
4. Draw and explain the molecular orbital diagram for nitrogen (N_2). What type of bonding exists?
5. Differentiate between Sigma and Pi Bonds in detail.
6. Differentiate between sp, sp^2, and sp^3 hybridization.
7. Explain the postulates of collision theory. Also describe the terms "effective collision" and "activation energy" with the help of diagrams or examples.
8. Explain how a catalyst works. Support your answer with an energy profile diagram.
9. Discuss in detail the five major factors that affect the rate of a chemical reaction with examples.
10. Explain the difference between order of reaction and molecularity with examples.
11. Define rate law, rate constant, and order of reaction. Derive units of k for first and second-order reactions.
12. Relate the order of a reaction to the rate law for the reaction. How do you distinguish between zero order, first order and second order reaction?
13. Calculate the reaction rate if the concentration of A is 0.5 M, the concentration of B is 0.2 and the rate constant k is $4.0 M^{-2}s^{-1}$. Given the rate law for a reaction: Rate = $k[A][B]^2$.

LONG QUESTION :7

1. Define and explain the law of mass action and derive the expression for the equilibrium constant (K_c) for a general reaction: $aA + bB \rightleftharpoons cC + dD$.
2. State Le-Chatelier's Principle. Explain its application with respect to changes in concentration, temperature, pressure, and use of a catalyst.
3. Synthesis of ammonia by Haber's process is an exothermic reaction. What should be the possible effect of change of temperature, pressure, and concentration at the equilibrium stage?
4. The change of volume or pressure for the following reactions only changes the equilibrium position. Discuss how the direction changes for: (i) $2H_2 + O_2 \rightleftharpoons 2H_2O$ (ii) $N_2O_4 \rightleftharpoons 2NO_2$.
5. Ethanol reacts with ethanoic acid to form ethyl ethanoate and water.
6. Propanone reacts with hydrogen cyanide. At equilibrium, the concentration of the product is 0.0233 mol dm^{-3}. Calculate K_c given initial concentrations were 0.0500 mol dm^{-3}.
7. Define pH and pOH. Derive their mathematical expressions and prove that $pH + pOH = 14$ at $25°C$.

8. Define ionization constant (K_a) of acids. Derive the expression for K_a of a weak acid and explain its significance in distinguishing strong vs. weak acids.
9. What is a buffer solution? Explain the working mechanism of acidic (CH_3COOH/CH_3COONa) and basic buffers with examples.
10. What is the common ion effect? Explain with examples how the addition of a common ion affects the ionization of weak acids using Le Chatelier's Principle.
11. The ionic product of water (K_w) is 1.0×10^{-14} at $25°C$. If $[H^+] = 1.0 \times 10^{-7} M$, calculate $[OH^-]$, pH, and pOH.
12. Explain the use of Solubility Product (K_{sp}) in predicting precipitation vs. solubility.
13. Calculate the solubility of a sparingly soluble salt lead (II) iodide (PbI_2) in water. It has $K_{sp} = 1.4 \times 10^{-8}$.
14. $Ca(OH)_2$ is a sparingly soluble compound. Its solubility product is 6.5×10^{-6}. Calculate the solubility of $Ca(OH)_2$.
15. A buffer solution has a pH of 5.0. It is made from a weak acid HA with a pK_a of 4.8. What is the ratio of $[A^-]$ to $[HA]$?

LONG QUESTION :8

1. Describe the halogenation of methane using the free radical chain mechanism. Include initiation, propagation, and termination steps.
2. Describe the mechanism of electrophilic addition of hydrogen halides to alkenes. Discuss Markovnikov's Rule in the context of hydrogen halide addition.
3. Describe the important reactions of alkenes including hydrogenation, halogenation, hydration, and ozonolysis with equations.
4. Describe the following methods for the preparation of alkenes: (i) Dehydrohalogenation of alkyl halides (ii) Dehydration of alcohols.
5. Explain the structure and bonding in ethene. How does it affect the reactivity of alkenes?
6. What is geometrical isomerism in alkenes? Illustrate your answer with appropriate structural examples.
Unit: Environmental Chemistry
7. Explain the characteristics, height ranges, and temperature variations of the four main layers of the atmosphere.
8. Write short notes on the following: (i) CFCs and ozone layer depletion (ii) Greenhouse effect and global warming.
9. How does fossil fuel burning cause acid rain? Discuss in detail with chemical reactions.
10. Discuss the effects of major air pollutants (CO, SO_2, NO_2, O_3, PAN, particulates) on human health and the environment.
11. Define primary and secondary pollutants. Explain their sources and harmful effects with examples.
12. Explain the major natural and anthropogenic (man-made) sources of air pollution with examples.

LONG QUESTION :9

1. Describe the industrial and laboratory preparation of ammonia. Also give chemical equations and tests for its identification.
2. How oxides of nitrogen (NOx) cause the formation of photochemical smog and PAN? Give its mechanism.
3. Describe the processes of nitrification and denitrification with the help of relevant chemical equations and bacterial names.

4. Explain the Bronsted-Lowry basic character of ammonia. How is it converted into NH_4^+ ion and what is the geometry of this ion?

5. Discuss the structures, oxidation states, physical properties, and uses of at least three oxides of nitrogen.

Unit: Sulfur and its Compounds

6. Describe the different oxidation states shown by sulfur and explain their stability with reference to compounds like H_2S, SO_2, SO_3, and S_8.

7. Discuss the important chemical reactions of sulfur with oxygen, hydrogen, chlorine, water, and metals. Include balanced equations.

8. Describe the uses of sulfur and its compounds in industry, agriculture, and medicine.

9. Explain the trend in oxidizing power of halogens down group 17. Support your answer with reactions and ionic equations.

10. Describe the reactions of halogens with hydrogen. Compare their reactivity trend and write balanced chemical equations.

11. Describe how chlorine purifies drinking water. Include chemical reactions and the role of HClO and ClO^-.

12. Describe the displacement reactions of halogens with halide ions and relate them to oxidizing strength.