Questio Bank

Chapter: Structure of Atom

Question 1

How does Bohr's theory account for stability of an atom?

Ans.

According to Bohr, as long as an electron remains in a particular permitted circular orbit or stationary state, it neither emits nor absorbs energy. As a result, an electron can not spiral down towards the nucleus loosing energy continuously (as per Maxwell's theory of electromagnetic radiation). This explains why atoms are stable and do not collapse due to electrostatic attraction between the nucleus and the electrons.

Question 2

How atomic spectra of Hydrogen and Hydrogen like atom are explained by Bohr's theory?

Ans.

According to Bohr, when an atom is subjected to high temperature or electric discharge, an electron in an atom may jump from its normal energy levels i.e. from ground state to some higher energy level or excited state by absorbing photon of radiation of definite energy. The lifetime of such excited electrons is short and hence they return to some lower energy level or even o their stable ground state level in one or more steps. During each such jumping back energy is emitted I form of a photon of light or electromagnetic radiation of definite wavelength and frequency depending upon the energy difference between two energy levels. In Hydrogen spectra, corresponding to each wavelength of such photon emitted, there appears a bright line.

Question 3

What do you mean by Shell or energy level according to Bohr's model of atom?

Ans.

Out of infinite number of circular orbits possible around the nucleus of atom, electrons revolve only in certain definite or fixed circular orbits called stationary states. The term stationary state implies that the motion of electron in this orbit does not cause any change in the energy of the electron and the effect is as if they are stationery. According to Bohr each orbit is associated with a certain definite amount of energy. Stationary states are also known as or Shell or Bohr's energy levels. They are denoted by K, L, M, N and so on or by integers n =1, 2, 3, 4 and so on.

Question 4

What is the energy of electron in Bohr's nth orbit of hydrogen like atom?

Ans.

Energy of electron in any orbit for hydrogen like atom like He^{+}, Li^{2+} etc is given by the following expression

$E_n = -13.6z^2/n^2$ eV where E_n is the energy of nth orbit of H-like atom and z is At. No. of the element.

Question 5

What is the energy of electron in Bohr’s nth orbit of hydrogen atom?

Ans.

Energy of electron in any orbit for hydrogen atom is given by the following expression
$E_n = -13.6/n^2$ eV where E_n is the energy of nth orbit of H-atom

Question 6

What are limitations of Bohr’ model?

Ans.

According to Bohr’s theory, atom is planer, but now it has been established that atom has 3D existence. Bohr’s model is applicable for single electron system. It fails to explain the line spectra of multi-electron atoms. Each spectral line splits into finer lines when subjected to influence of magnetic field (Known as Zeeman effect) and electric field (known as Stark effect) Bohr’s atomic theory can’t explain this. It was also unable to explain de Broglie’ dual character of matter and Heisenberg’s uncertainty principle.

Question 7

What were the findings of Rutherford’s scattering experiment?

Ans.

Rutherford (1911) discovered the nucleus and provided the basis for the modern atomic structure through his alpha particle scattering experiment. According to Rutherford, the atom is made of two parts: the nucleus and the extra-nuclear part. His experiments proved that the atom is largely empty and has a heavy positively-charged body at the center called the nucleus. The central nucleus is positively-charged and the negatively-charged electrons revolve around the nucleus.

Question 8

What are drawbacks of Rutherford's model of atom?

Ans.

Rutherford's model could not explain the stability of atom as a moving electron in an atomic orbit will continuously loose energy and will fall into the nucleus.
It also fails to explain discontinuous linear atomic spectrum.

Question 9

Describe Rutherford's gold foil experiment.

Ans.

Rutherford shot alpha particles through thin gold foil. He expected all of the particles to go straight through. Most did, but some deflected. This led him to two conclusions: first, that atom is mostly empty space, and secondly, that there is a positively charged, dense nucleus present.

Question 10

State the differences between the following atomic models: J.J. Thomson's, Rutherford's, Bohr's, Quantum mechanical model.

Ans.

The model presented by J.J. Thomson is known as the "plum pudding model" - a positively charged blob with negative electrons stuck throughout it. Rutherford stated that there was a small, dense, positively charged nucleus with electrons outside the nucleus. Bohr also had a positive nucleus, but stated that the electrons were in definite orbits, or energy levels, around the nucleus. The quantum mechanical model also gives a positive nucleus, but states that the electrons are in not in a definite orbit, but rather in a probability region, or electron cloud.

Question 11

Two isotopes of calcium (atomic number 20) are ^{40}Ca and ^{44}Ca. For both ^{40}Ca and $^{44}Ca^{2+}$ give the following information: atomic number & mass number.

Ans.

The atomic number is unique for each element, and is the same for all isotopes of an element and it does not depend on the number of neutrons present. Therefore, the atomic number for both isotopes is 20.

Atomic number depends only on the number of protons present, so the charge on the $^{44}Ca^{2+}$ ion (which arises from the loss of two electrons) does not alter this.

The mass number is the number of (protons + neutrons) present For ^{40}Ca this is 40. Note that the charge present on the $^{44}Ca^{2+}$ ion does not affect the composition of the nucleus, and the mass number is 44.

Question 12

For one atom of the isotope ^{34}X (atomic number 16) give the number of neutrons.

Ans.

In a neutral atom the mass number equals the sum of the number of protons and neutrons. Since the number of protons in the atom is 16, hence the number of neutrons in the atom is 34-16 = 18. The radius of the various nuclei (r) can be calculated from; r = (1.3 × 10-13) m 1/3 cm. Then find the radius of the nucleus of the atom ^{34}X.

r = 1.3 × 10-13 × 341/3

Question 13

For one atom of the isotope ^{34}X (atomic number 16) give the number of protons the number of electrons.

Ans.

The number of protons equals atomic number, which is 16.

We assume here that the atom in question is neutral, since no charges have been explicitly mentioned. In a neutral atom the number of electrons equals the number of protons, which is 16.

Question 14

The mass to charge ratio for A+ ion is 1.97 x 10-7 kg C-1. Calculate the mass of A atom.

Ans.

Given m/e = 1.97 x 10-7

Hence, m = 1.97 x 10-7 x 1.6 x 10-19 = 3.16 x 10-26 kg

Question 15

The element boron consists of two isotopes, $^{10}_{5}B$ and $^{11}_{5}B$. Their masses, based on the carbon scale, are 10.01 and 11.01, respectively. The abundance of $^{10}_{5}B$ is 20.0% and the abundance of $^{11}_{5}B$ is 80.0%. What is the atomic mass of boron?

Ans.

The percentages of multiple isotopes must add up to 100%.

Apply the following equation to the problem:

atomic mass = (atomic mass X_1) $\times$ (% of X_1)/100 + (atomic mass X_2) $\times$ (% of X_2)/100

Substituting for boron in this equation:

atomic mass of B = 10.01 $\times$ 20.0/100 + 11.01 $\times$ 80.0/100

atomic mass of B = 2.00 + 8.81

atomic mass of B = 10.81

Question 16

What are isotope, isobars and isotones?

Ans.

Isotopes are atoms of the same element having same atomic number but different mass numbers. They have similar chemical properties but different physical properties. Isobars are atoms of different elements having same mass numbers (i.e. the sum of their Protons and Neutrons are same). Isotones are atoms of different elements having same number of Neutrons.

Question 17

What is atomic Number and who introduced the term?

Ans.

Moseley in 1913 introduced the term as fundamental property of an element. According to him “It is the number of unit +ve charges carried by the nucleus of an atom. It is generally denoted by Z.”

Question 18

How were Neutrons discovered and by whom?

Ans.

Neutron as a fundamental particle was first predicted by Rutherford but discovered first by Chadwick in 1932 when he bombarded Beryllium foil with high speed alpha particle

Question 19

How were protons discovered and by whom?

Ans.

protons were discovered in the Cathode ray experiment and the name proton was proposed by Rutherford

Question 20

What are important characteristics of Anode rays?

Ans.

Important characteristics of Anode rays:

(i) They consists of heavy positively charged particles.

(ii) The charge to mass ratio of these particles was found to be dependent upon the type of gas used in the discharge tube.

(iii) The charge to mass ratio for hydrogen gas is found to be equal to that of protons. Anode rays get deflected into a direction opposite to that of cathode rays in electrical and magnetic fields.

Question 21

How were electrons discovered and by whom?

Ans.

Electrons were discovered by J.J. Thomson using the cathode ray tube.

Question 22

What are important characteristics of Cathode rays?

Ans.

Important characteristics of Cathode rays:

1. They consist of negatively charged material particles; electrons.
2. They produce X rays. When a beam of cathode rays is made to fall upon a hard metallic target like Tungsten, X-rays are produced.
3. They cause ionization of gas through which they pass. They are deflected by electric and magnetic fields towards +ve field.

Question 23

How were Cathode rays discovered and by whom?

Ans.

Cathode rays were first discovered by Julius Plucker .The Ideal condition for producing cathode rays in discharge tube are 0.01 mm Hg and 10,000 V potential.

Question 24

What is the photoelectric effect? Give an example.

Ans.

The photoelectric effect is where a photon (packet of light) can eject an electron from a metal. The photon must be at least at the threshold energy for ejecting a electron from metal surface. For example sodium metal.

Dual nature

Question 25

Explain de Broglie's wave length?

Ans.

De Broglie hypothesized that all matter can exhibit wave like properties. He derived an equation (λ=h/ mv) to describe the wavelength of a particle where h is Planck's constant, m is the mass of the particle, and v is its velocity.

Question 26

A microscopic particle of mass 10^{-26} kg is moving with a kinetic energy of 5×10^{-25} joule. Calculate its de Broglie wavelength.

Ans.

We know, $K E = \frac{1}{2} m v^2$

Or, $v = \sqrt{2} \times K E / m$

Substituting the values of KE and m

We get , v = 10 m/s

We also know, from de Broglie relation ; $\lambda = h/mv$

$= 6.63\times10^{-34}/10^{-26}\times10$

$= 6.63 \times 10^{-8}$ m

= 66.3 nm.

Question 27

Calculate the momentum of a moving object which has a de Broglie wavelength of 210 pm.

Ans.

$210pm = 2.1 \times 10^{-10}$

We know , from de Broglie relation ;

$\lambda = h/mv$

$mv = h/ \lambda$

$= 6.63\times10^{-34}/2.1 \times 10^{10}$

$= 3.15 \times 10^{-24}$ $kgms^{-1}$

Question 28

Calculate the frequency and wavelength of a body of mass 1 mg moving with a velocity of 10 ms-1.

Ans.

Mass = 10 mg = 10 x10^{-6} kg

We know , from de Broglie relation ; λ = h/mv

= 6.63x10^{-34}/10^{-6}x10

= 6.63 x 10^{-29} m

We also know that, v = c/ λ

Or , v = 3.0 X 10^{8} / 6.63 x 10^{-29} = 4.52 x 10^{20} s-1

Question 29

Calculate the wavelength of a particle of mass 8.6 × 10^{-27} kg moving with a kinetic energy of 3.2 × 10^{-12} joules.

Ans.

We know, K E = ½ m v^2

Or, v = $\sqrt{2}$ x K E/ m

Substituting the values of KE and m

We get , v = 3.11 x 10^{6} m/s

Now using de Broglie relation ; λ = h/mv

λ = 6.6 x 10^{-34}/ 8.6 x 10^{-27} x 3.11 x 10^{6}

= 2.4 x 10^{-14} m

Question 30

A dust particle of mass 10^{-11} gm has velocity of 10^{-4} cm/s. The error in measurement of velocity is 0.1 %. Calculate uncertainty in its position.

Ans.

ΔV = 0.1 x 10^{-4}/100 = 10^{-7} cm/s = 10^{-9} m/s

We know, ΔX. mΔV ≥ h/4π

Hence, ΔV = 6.626 x 10^{-34}/ 4 x 3.14 x 10^{-14} x 10^{-9} = 5.27 x 10^{-12}m

Question 31

Calculate the uncertainty in velocity of an electron if uncertainty in its position is of the order of 1 AÂ°.

Ans.

We know, ΔX. mΔV ≥ h/4π

Hence, ΔV = 6.626 x 10^{-34}/ 4 x 3.14 x 9.1 x 10^{-31} x 10^{-10} = 5.8 x 10^{5} m/s

Question 32

Calculate the uncertainty in velocity of a cricket ball of mass 0.15 kg if its uncertainty in position is of the order of 1A^0

Ans.

We know, $\Delta X.\ m\Delta V \geq h/4\pi$

Hence, $\Delta V = 6.626 \times 10^{-34} / 4 \times 3.14 \times 0.15 \times 10^{-10} = 3.51 \times 10^{-24}$ m/s

Question 33

On the basis of Heisenberg's uncertainty principle, show that the electron can not exist within the nucleus.

Ans.

Radius of the nucleus is of the order of 10^{-15} m and thus maximum uncertainty in the position of electron i.e Δx if it is within the nucleus will be 10^{-15} m.

Now, $\Delta X.\ m\Delta V \geq h/4\pi$

Or, $\Delta V = 6.626 \times 10^{-34} / 4 \times 3.14 \times 9.1 \times 10^{-31} \times 10^{-15} = 5.97 \times 10^{10}$ m/s

As this level of uncertainty in velocity is impossible to be achieved hence electron can not exist in the nucleus.

Question 34

What is Heisenberg's uncertainty principle?

Ans.

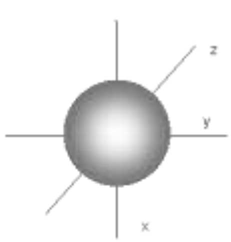
Heisenberg's uncertainty principle states that both the position and momentum of an electron cannot be known precisely at the same time.

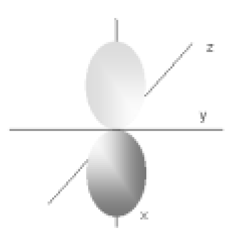
Electronic configuration of atoms

Question 35

Draw the boundary surface diagram for 2s orbital and 3p$_x$ orbital.

Ans. (a) A 2s orbital





(b) A 3p orbital

Question 36

Write the ground state electron configuration of an atom of Tungsten.
How many of the electrons in this atom have a principal quantum number of 4?

How many electrons are in the valence shell of this atom?

Ans.

Electronic configuration of Tungsten having at. no. 74

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^4$

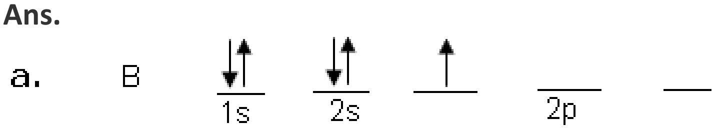
Total no. of possible electrons in 4th Sub shell = 32

No. of electron in the valance shell= 2

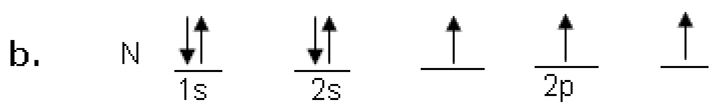
Question 37

Give orbital diagram (using lines) for ^5B and ^7N and indicate the number of unpaired electrons for each:

Ans.



1 unpaired electron



3 unpaired electrons

Question 38

Write the electron configuration for each of the following: Aluminum, Vanadium and Chlorine ion(Cl-)

Ans.

Al

$1s^2 2s^2 2p^6 3s^1$ or $[Ne]3s^1$

V

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$ or $[Ar]4s^2 3d^3$

Cl^-

$1s^2 2s^2 2p^6 3s^2 3p^6$ or $[Ar]$

Question 39

Fill in the missing values or names of the following sets of quantum numbers.

a. $n = 2, l = ?, ml = +1$; name: $2p$

b. $n = ?, l = 0, ml = ?$; name: $4s$

c. $n = 3, l = 1, ml = +1$; name: ?

d. $n = 3, l = ?, ml = 0$; name: $3d$

Ans.

a. $n = 2, l = 1, ml = +1$; name: $2p$

b. $n = 4, l = 0, ml = 0$; name: $4s$

c. $n = 3, l = 1, ml = +1$; name: $3p$

d. $n = 3, l = 2, ml = 0$; name: $3d$

Question 40

Indicate which of the following sets of quantum numbers in an atom are unacceptable and explain why: (a)1, 0, 1/2, 1/2 (b)3, 0, 0, 1/2 and (c)2, 2, 1, 1/2

Ans.

(a) (1, 0, ½, ½)

Unacceptable: Cannot have m_ℓ with non-integer value

Since $\ell = 0$, m_ℓ can only be 0

(b) (3, 0, 0, ½)

Acceptable--this describes an electron in a 3s orbital

(c) (2, 2, 1, ½)

Unacceptable: The values of ℓ can only be 0, ..., n-1. Since $n = 2$, ℓ can only be 0 or 1, not 2.

Question 41

Use the Aufbau principle to write the ground state electron configuration of the following ions(a) Co^{3+}(b) In^{3+}(c) Bi^{3-}(d) Hg^{+}

Ans.

(a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^4$

(b) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^8$

(c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^6$

(d) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^9$

Question 42

Determine the maximum number of electrons that can be found in each of the following sub shells: 3s,3d and 4p

Ans.

3s

Each orbital can accommodate 2 electrons. The number of orbitals in a sub shell is given by $2\ell + 1$.

An s sub shell (1 orbital) can hold 2 electrons.

3d

A d sub shell ($\ell = 2$) can hold 10 electrons.

4p

A p sub shell ($\ell = 1$) can hold 6 electrons.

Question 43

What is the difference between a $2p_x$ and a $2p_y$ orbital?

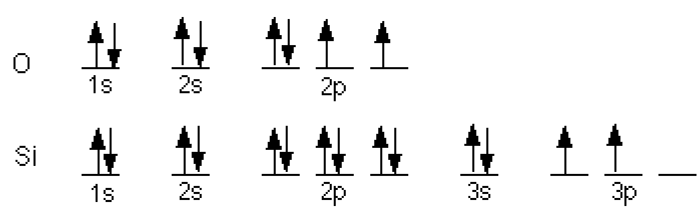
Ans.

The orbitals have the same size (energy) and shape, but are oriented differently in space (different magnetic quantum numbers). The $2p_x$ orbital lies on the x-axis while the $2p_y$ orbital lies along the y-axis.

Question 44

Draw the orbital diagrams for O and Si. How many unpaired electrons are in each of these?

Ans.



There are 2 unpaired electrons in each.

Question 45

What is the number of unpaired electron in Cr^{3+}? (At No. Cr = 24)

Ans.

The Electronic configuration of $_{24}Cr$:

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^3$

Thus, there are 3 unpaired electrons.

Question 46

Write electronic configuration of a metal ion which is isoelectronic with Cr^{3+}. What is the number of occupied energy shells, sub shells and Orbitals in it?

Ans.

V^{2+} is isoelectronic with Cr^{3+}

V^{2+}: $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $3d^3$

Hence, Occupied energy shells = 3, Sub shells=6 and Orbitals = 12

Question 47

One unpaired electron in an atom contributes a magnetic moment of 1.1 BM Calculate the magnetic moment for Chromium.

Ans.

Electronic Configuration of Chromium atom is

24Cr: $1s^2,2s^2,2p^6,3s^2,3p^6,3d^5,4s^1$

There are six unpaired in Cr atom

Hence, the Magnetic moment for Chromium= 1.1 x 6 = 6.6 BM

Question 48

Which of the following atoms/ions are diamagnetic Zn, Cu^{2+}, Cd, Ti^{2+}, Cu^{+}

Ans.

Zn, Cd and Cu^{+} are diamagnetic as they do not contain unpaired electrons.

Question 49

Calculate the ionization energy of the Li^{2+} ion with the electron in its ground state.

Ans.

It is important to note that this is Li^{2+}, has an electron configuration of $1s^1$

Hence, $E = -(Z^2 / n^2) \times 13^{10}$ kJ/mol = -(32/12) x13^{10} kJ/mol = 1.18×10^2 kJ

Question 50

Give the electron configurations of potassium and calcium.

Which of the following will have the greatest first ionization energy?

Which of the following will have the greatest second ionization energy?

Ans.

Electron configurations are as follows;

K (alkali metal); [Ar] $4s^1$

Ca (alkali earth metal); [Ar] $4s^2$

The effective nuclear charge experienced by the outermost valence electrons on Ca is greater than that on K. This explains why Ca is actually smaller than K and why the first ionization energy of Ca would be greater than that for K.

The second ionization energy however is much greater for K than for Ca, since in removing the second electron one must remove one of the core electrons from K.

Question 51

Of the following electron configurations, state whether each represents an atom in the ground state, a possible excited state, or is incorrect.

a) $1s^2 2s^2 2p^1$ b) $1s^2 2s^1 3s^1$ c) $1s^2 2s^2 2p^6 3s^2 2d^2$ d) $1s^2 2s^4 2p^2$ e) $1s^1 2s^1$

Ans.

a) $1s^2 2s^2 2p^1$ - correct

b) $1s^2 2s^1 3s^1$ -excited

c) $1s^2 2s^2 2p^6 3s^2 2d^2$ -incorrect

d) $1s^2 2s^4 2p^2$ -incorrect

e) $1s^1 2s^1$-excited

Question 52

Account for the fact that manganese has a greater number of oxidation states than zinc even when they have similar electronic configurations.

Ans.

Mn is a transition element with an incompletely filled 3d shell. The completely filled 3d shell in zinc leads to stability and removal of the 4s electrons with only a relatively low ionization energy to form the 2+ ion is the only common oxidation state of zinc (+II). The filled 3d orbital cannot participate in bonding. For manganese, both the 3d and 4s sub shells are available for occupation by bonding electrons, so a wide range of oxidation states is observed.

Question 53

Write the detailed electronic configurations for the atoms of the following elements: Mn & Zn

Ans.

Electronic configurations of Mn : $1s^2\ 2s^2\ 2p^6\ 3s2\ 3p^6\ 3d^5\ 4s^2$
Electronic configurations of Zn : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2$

Question 54

How many electrons can be accommodated in the l=1 orbitals? What is the value of l for the f-orbitals?

Ans.

For l=1,

ml = -1, 0, +1

Each orbital can accommodate 2 electrons hence total no. of electrons = 6

The value for l for f-orbital is 3.

Question 55

For the principle quantum no. n = 4; How many types of orbitals are there? How many electrons can be accommodated in the complete principle shell.

Ans.

For n = 4, there are four possible values for l.

They are;

0 .... s orbitals

1 .... p orbitals

2 .... d orbitals

3 .... f orbitals

For each of these there are values for ml

l=0, ml = 0 = 2 electrons

l = 1, ml = -1, 0, +1 = 6 electrons

l = 2, ml = -2, -1, 0, 1, 2 = 10 electros

l = 3, ml = -3,-2, -1, 0, 1, 2,3 = 14 electrons

Each orbital can accommodate 2 electrons hence total no. of electrons = 32

Question 56

What is the maximum number of electrons in a given atom that can have n = 3, ml =1.

Ans.

For n = 3, there are three possible values for l.

They are;

0 s orbitals

1 p orbitals

2 d orbitals

For each of these there are values for ml

l=0, ml = 0

l=1, ml = -1, 0, +1

l = 2, ml = -2, -1, 0, 1, 2

Question 57

Give the quantum numbers l, ml, and the number of orbitals for n = 4

Ans.

l = (n-1) = 4

ml = -3, -2, -1, 0, 1, 2, 3

7 orbitals (count the ml)

Question 58

What are the three types of the Electromagnetic Spectrum?

Ans.

The Electromagnetic Spectrum is divided into three portions for convenience of study:
Infra red zone consisting of waves greater than visible zone
Visible zone consists of seven distinct colored light waves which are sensed by our eyes
Ultra-Violet Zone consisting of low wavelengths but high frequency.

Question 59

What do you mean by Quantum Numbers?

Ans.

Quantum Numbers specify the properties of atomic orbitals and their electrons.
Quantum numbers can be used to write electron configurations which show how electrons are most likely distributed around the nucleus
There are four quantum numbers:
principal quantum number
orbital quantum number
magnetic quantum number
spin quantum number
The principal quantum number (n) specifies the main energy levels around the nucleus. As n increases, the distance from the nucleus increases. Currently the values for n are 1 to 7 Orbital Quantum Number (l) indicates the shape of the orbital where the electron can be found These orbitals are called sub shells or sublevels The four most common orbital quantum numbers are given letter abbreviations l" values range from 0 to (n-1)
Orbital Quantum Numbers:
l= 0, s orbital
l = 1, p orbital
l = 2, d orbital
l = 3, f orbital
Magnetic Quantum Number (m) indicates the orientation of an orbital about the nucleus it tells which axis that sublevel is located on (x, y, or z axis)
ml ranges from - 1 to 1
Spin Quantum Number (ms) indicates the two possible states of an electron within an orbital their Values are + 1/2 or -1/2

Question 60

In the Quantum Mechanical Model, what are electrons considered to be?

Ans.

In the Quantum Mechanical Model electrons are considered to be clouds of electromagnetic waves which occupy the space around the nucleus in different energy levels. The space around the nucleus where probability of finding an electron is maximum is called an orbital.

Question 61
Describe the Quantum Mechanical Model of the atom.
Ans.
The Quantum Mechanical Model is a complex mathematical model. The Quantum Model is based on understanding the behavior of light which is considered to be composed of electromagnetic (EM) waves.

Based on knowing some information about electrons, the location of the electron is predicted. Electrons exit around the nucleus in form of electron clouds called orbitals in different energy levels. Low energy electrons are found near the nucleus; high energy electrons are found further away from the nucleus. Electrons emit energy when fall back in lower energy levels. Energy is released in the form of electromagnetic radiation. When this energy is emitted, it can be observed using special instruments called spectrophotometers. EM spectrum shows all forms of radiation

Quantum mechanical model of atom

Question 62
What are the values of n, ℓ and m_ℓ for 3p-orbitals?
Ans.
For 3p-orbital
n = 4, ℓ = 1 and m_ℓ can have any of three values -1, 0, +1.

Question 63
How many quantum numbers are required to specify an orbital?
Ans.
Three (n, ℓ and m_ℓ)

Question 64
What is the maximum number of electrons that may be present in all the atomic orbitals with principal quantum number 3 and azimuthal quantum number 2?
Ans.
Ten

Question 65
Write down all the four quantum numbers for outermost electron of sodium atom (Z=11).
Ans.
N = 3, ℓ =0 m= 0, s = +1/2

Question 66

(i) An atom orbital has n= 3, What are the possible values of ℓ and m_ℓ.

(ii) List the quantum number (m_ℓ and ℓ) of electrons for 3d-orbital.

(iii) Which of the following orbital are possible?
1p, 2s, 2p and 3f.

Ans.

(i) ℓ = 0,1,2 and m_ℓ = -2, -1, 0, +1, +2

(ii) ℓ = 2, m_ℓ = -2, -1, 0, +1, +2

(iii) 2s, 2p are possible.

Question 67

If the quantum number 'n' has a value of 3, what are the permitted values of the quantum number 'ℓ'?

Ans.

0, 1, 2.

Question 68

How many electrons in a given atom can have the following quantum numbers?

(i) $n = 3,\ \ell = 1$ (ii) $n = 3,\ \ell = 2,\ m_\ell = 0$

(iii) $n = 3,\ \ell = 2,\ m_\ell = +2,\ m_s = +\frac{1}{2}$ (iv) $n = 3$

Ans.

(i) 6 (ii) 2 (iii) 1 (iv) 18

Question 69

Give all possible values of ℓ, m_ℓ and ms for electrons when n = 3.

Ans.

$\ell = 0,\ m_\ell = 0;\ \ell = 1,\ m_\ell = -1, 0, +1;$

$\ell = 2,\ m_\ell = -2, -1, 0, +1, +2$ and

$m_\ell = +\frac{1}{2}$ and $-\frac{1}{2}$ for each value of m_ℓ.

Question 70

(i) What subshells are possible in n = 3 energy level?

(ii) How many orbitals (of all kinds) are possible in this level?

Ans.

(i)Subshells in n = 3 energy level

We know that the subshells are given by different values of ℓ .

For n = 3, the possible values of ℓ are 0, 1 and 2.

The corresponding subshells are:

ℓ = 0, s-subshell ℓ = 1, p-subshell ℓ = 2, d-subshell.

(ii) Number of orbitals

For n = 3, there are one s, three p and five d-orbitals.

This makes total of nine orbitals in n = 3 level.

Question 71

Using the s, p, d notations describe the orbital with the following quantum numbers.

(i)	$n = 1, \ell = 0$	(ii)	$n = 3, \ell = 2$
(iii)	$n = 3, \ell = 1$	(iv)	$n = 2, \ell = 1$
(v)	$n = 4, \ell = 3$	(vi)	$n = 4, \ell = 2$

Ans.

(i)	$n = 1, \ell = 0$	:	1s-orbital
(ii)	$n = 3, \ell = 2$	:	3d-orbital
(iii)	$n = 3, \ell = 1$	:	3p-orbital
(iv)	$n = 2, \ell = 1$	:	2p-orbital
(v)	$n = 4, \ell = 3$	:	4f-orbital
(vi)	$n = 4, \ell = 2$	:	4d-orbital

Question 72

List all the values of ℓ and m for n = 2.

Ans.

When $n = 2$, ℓ can have values 0 and 1

For $\ell = 0$, $m_\ell = 0$

For $\ell = 1$, $m_\ell = -1, 0, +1$.

Question 73

(i) If the quantum number 'ℓ' has value of 2, what are the permitted values of the quantum number ?

(ii) An atomic orbital has n = 3, what are the possible values of ℓ ?

(iii) An atomic orbital has ℓ = 3, what are the possible values of m_ℓ ?

Ans.

(i) If ℓ = 2, the permitted values of m are:

me = -2, -1, 0, +1, +2

(ii) For n = 3, ℓ may have the value ℓ = 0, 1, 2

(iii) For ℓ = 3, me may have the values

m = -3, -2, -1, 0, +1, +2, +3

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