2 STRUCTURE OF ATOM

INTRODUCTION

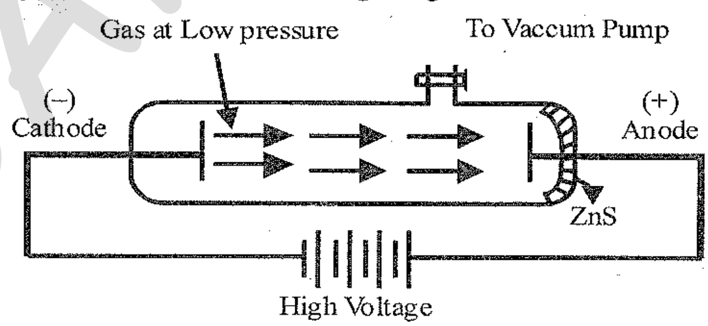
The existence of atom has been proposed since the time of early Indian and Greek philosophers. The atom are the fundamental building blocks of matter. According to them the continued subdivisions of matter would ultimately gives atom which would not be further divisible. The word 'atom' has been derived from Greek word 'a-tomio' which means uncuttable or non-divisible.

DALTON'S ATOMIC THEORY

The atomic theory of matter was first proposed on a firm scientific basis by John Dalton in 1808. His theory called Dalton's atomic theory.
Dalton's atomic theory was able to explain the law of conservation of mass, law of constant composition and law of multi proportion very successfully

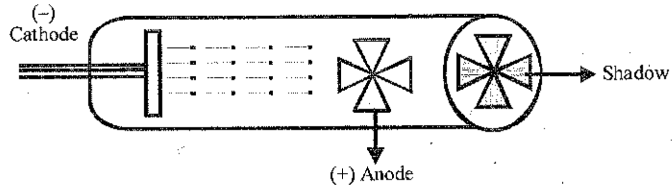
DISCOVERY OF THE ELECTRON

The cathode rays were discovered by **Julius Plucker** in 1859 during the experiments with gas discharge tubes.
A discharge tube is long glass tube as shown in fig. This tube is fitted with metal electrodes on either end across which high voltage can be applied. The tube is also connected to vacuum pump for controlling the pressure of gas inside the discharge tube.
Note : A pressure of 0.001 mm of Hg is applied because at normal pressure gases are non-conductors. At this low pressure of 0.001 mm of Hg the atoms and molecules of gases get ionised.

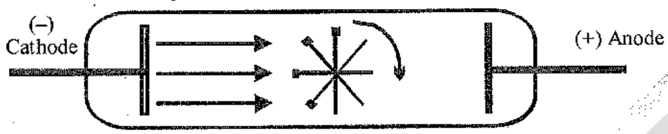


PROPERTIES OF CATHODE RAYS : The cathode rays possess the following properties :

1. **Cathode rays travel in straight lines :** An object placed in the path of cathode rays casts a sharp shadow. It shows that cathode rays travel in straight lines.



2. **Heating Effect :** When cathode rays are focused on a thin metal foil, it gets heated up to incandescence.
3. **Cathode rays consist of material particles :** This was indicated by the fact that a light paddle wheel placed in the path of cathode rays starts rotating.



Light Paddle wheel

4. **Effect of electric field :** When electric field is applied to a stream of cathode rays, they get deflected towards positive plate. It shows that cathode rays are negatively charged.
5. **Effect of Magnetic field :** When magnetic field is applied, perpendicular to the path of the cathode rays, they get deflected in the direction expected for negative particles.
6. On striking against walls of the discharge tube, cathode rays produce faint greenish fluorescence.

Result

1. The cathode ray start from cathode and moves towards anode.
2. There rays themselvesare not visible but their behaviour can be observed within the help of certain kind of material (zns) which glow when hit by them.
3. In the absence of electrical or magnetic field these rays travel in straight line.
4. The cathode rays consist of negatively charged particles called electron.
5. The characterstics of cathode rays don't defend upon the material.

"Thus we can conclude that electron are basic constituent of all the atoms."

ANODE RAYS

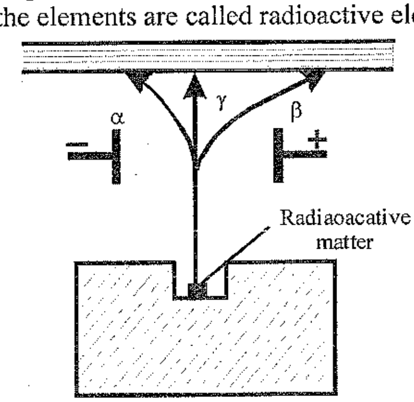
Canal Rays

Goldstein, in 1886, discovered the existence of a new type of rays in the discharged tube. He used a perforated cathode in the discharge tube. On passing the electric discharge at low pressure he observed a new type of rays streaming behind the cathode. These rays were **anode rays.**

Note : Charge to mass ratio of the particle in the anode rays depends upon nature of the gas taken in the discharge tube.

- **A proton is a fundamental particle of atom carrying one unit positive charge and having mass nearly equal to the mass of atom of hydrogen.**

  RADIOACTIVITY : The spontaneous emission of active radiations by certain elements like Uranium is called radioactivity and the elements are called radioactive elements.



1. Unline cathode rays, mass of positively charged particles depends upon the nature of gas present in cathode ray tube.
2. The charge to mass ratio of particles depends on the gas from which there originates.
3. Some of the positively charged particles carry a multiple of the fundamental unit of electric charge.
4. The behaviour of these particle in the magnetic or electrical field in opposite to that observed for electron or cathode rays.

The smallest and lightest positive ion obtained from hydrogen was called Proton.

Compt. Point → β-rays more deflect than α-rays due to less momentum.

- β-particle is a high velocity electron.
- β-rays emits from nucleus.
  $^1_0n \rightarrow {}^1_1p + {}^0_{-1}e$
- Protons remain in nucleus and electron goes out.
- γ-rays are just an electromagnetic radiation. There are no particles present.

Property	α	β	γ
Velocity	1/10 of c	Almost equal to c	Equal to c
Penetration power	Small	Large (100 times of a rays)	Very large (10,000 times of a-rays)
Ionisation power	Very Large	Large	Very Small

DISCOVERY OF NEUTRON

Neutron was discovered by James **Chadwick.**

He bombarded a thin foil of beryllium with fast moving α-particle and observed that highly penetrating rays consisting of neutral particles were produced.

$$_4Be^9 + \frac{4}{2}He \rightarrow \frac{12}{6} + {}^1_0n$$

A neutron is a subatomic particle carrying no charge and having mass 1.675 × 10^{-27} kg which is almost equal to that of a hydrogen atom.

Note : Neutron has 0.18% more mass than proton. When electrically neutral particles having a mass slightly greater than that protons were emitted these particle as neutrons.

	Symbol	Charge (Coulomb)	Charge (e.s.u.)	Mass (gram)	Mass (a.m.u.)	Inventor
Electron	$_{-1}e^0$	-1.602×10^{-19} C	-4.8×10^{-10}	9.109×10^{-28}	0.000548	J.J. ThomSon
Proton	$_1H^1$	$+1.602 \times 10^{-19}$ C	$+4.8 \times 10^{-10}$	1.673×10^{-24}	1.00727	Goldstein
Neutron	$_0n^1$	Zero	—	1.6748×10^{-24}	1.00867	ChadWick

ATOMIC MODELS

Atom is composed of sub-atomic particles carrying positive and negative charges. Different atomic models were proposed to explain the distributions of these charged particles in an atom.

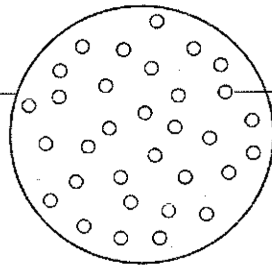
Two models are briefly explained below-

1. **THOMSON'S MODEL OF ATOM**
2. **RUTHERFORD'S NUCLEAR MODEL**
3. **BOHR'S MODEL OF ATOM**

THOMSON'S MODEL OF ATOM

Thomson was the first to propose a detailed model of the atom. He proposed that an atom consists of a uniform sphere of positive electricity in which the electrons are distributed more or less uniformly. An atom possesses a spherical shape (radius approximately 10^{-10} m)

This model of atom is known as the **'plum-pudding model'** or **'raisin pudding model'** or **watermelon model. This model assumed that mass of the atom is evenly spread over the entire atom.**



Thomson model of atom

An important feature of this model is that the mass of the atom is assumed to be uniformly distributed over the atom.

RUTHERFORD'S NUCLEAR MODEL

Rutherford and his students (Hans Geiger and Ernest Marsden) bombarded very thin gold foil with α–particles.

Rutherford's famous α–particle scattering experiment - A stream of high energy α–particles from a radioactive source was directed at a thin foil (thickness ~ 100 nm) of gold metal. The thin gold foil had a circular fluorescent zinc sulphide screen around it. Whenever α–particles struck the screen, a tiny flash of light was produced at that point.

It was observed that :

(i) most of the α–particles passed through the gold foil undeflected.

(ii) a small fraction of the α–particles was deflected by small angles.

(iii) a very few α-particle bounced back that is deflected by 180°.

CONCLUSION:

(i) Since most of the α-particles pass through the foil undeflected, it indicates that the most of the space in an atom is empty.

(ii) α-particles being positively charged and having considerable mass could be deflected only by some heavy, positively charged centre. The small angle of deflection of α-particles indicated the presence of a heavy positive centre in the atom. Rutherford named this positive centre as **nucleus.**

(iii) α-particles which make head-on collision with heavy positive centre are deflected through large angles. Since the number of such α-particles is very small, the space occupied by the heavy positive centre must be very small.

Result

1. Most of the mass and all the positive charge of an atom is concentrated in a very small region called nucleus.
2. The positive charge on the nucleus is due to protons. The magnitude of the charge on the nucleus is different for atoms of different elements.
3. The nucleus is surrounded by electrons which are revolving around it at very high speeds in circular path called **orbits**. The electrostatic force of attraction between electrons and the nucleus holds them together.
4. **Nuclear model of atom can be compared with the solar system.** In an atom electron revolve around the nucleus in just the same way as the planets revolve around the sun. Due to this comparison revolving electrons are sometimes called **planetary electrons** and Rutherford's nuclear model of atom is known **planetary model of atom.**

NOTE :

1. Size of the nucleus is extremely small as compared with the size of the atom.
2. Radius of the nucleus is of the order of $\mathbf{10^{-15}}$ **m**, whereas radius of atoms is of the order of $\mathbf{10^{-10}}$ **m.**

Drawbacks of Rutherford's model

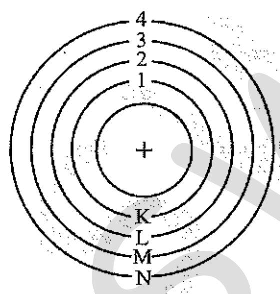
1. **Rutherford model failed in view of electromagnetic theory** given by Maxwell.

According to this theory a charged particle when accelerated emits energy in the form of electromagnetic radiation.
According to Rutherford's model, electrons are revolving around the nucleus. This means, electrons would be in a state of acceleration all the time. Since electrons are charged particles, therefore, electron revolving in an orbit should continuously emit radiations. As a result of this, it would slow down and would no longer be able to withstand the attractive force of the nucleus. Hence, it would more closer and closer to the nucleus and would finally fall in the nucleus by following a spiral path. This means atom should collapse. But actually we know atom is stable. Thus, **Rutherford's model failed to explain stability of atom.**

2. Rutherford's model also failed to explain the existence of certain definite lines in the hydrogen spectrum. **According to theory of Rutherford atomic spectrum should be continuous but atomic spectrum is a line spectrum**
3. Rutherford's model does not explain the distribution of electrons around the nucleus and their energies.

BOHR'S MODEL OF ATOM

Bohr model of atom was proposed by Neil Bhor in 1915. It came into existence with the modification of Rutherford's model of atom.
Bohr modified this atomic structure model by explaining that e^- move in fixed orbital and also explain each shell.



Bohr's Model of an Atom

NATURE OF ELECTRO MAGNETIC RADIATIONS

Wave Theory of Radiations

Light is the form of radiation known from early days and speculation about its nature dates back to remote ancient times. In earlier days (Newton) light was supposed to be made of particles (corpuscles). It was only in the 19th century when wave nature of light was established.

Maxwell was again the first to reveal that light waves are associated with oscillating electric and magnetic character.

Properties of Electromagnetic Waves

(i) The oscillating electric and magnetic fields produced by oscillating charged particles are perpendicular to each other and both are perpendicular to the direction of propagation of the wave. Simplified picture of electromagnetic wave is shown in Fig.

(ii) Unlike sound waves or water waves, electromagnetic waves do not require medium and can move in vacuum.

(iii) There are many types of electromagnetic radiations, which differ from one another in wavelength (or frequency). These constitute what is called electromagnetic spectrum .
Different regions of the spectrum are identified by different names.
Some examples are: **radio frequency region** around 10^6 Hz, used for broadcasting.
Microwave region around 10^{10} Hz used for radar;
Infrared region around 10^{13}Hz used for heating; **ultraviolet region** around 10^{16}Hz a component of sun's radiation.
The small portion around 10^{15} Hz, is what is ordinarily called **visible light**. It is only this part which

our eyes can see (or detect). Special instruments are required to detect non-visible radiation.

(iv) Different kinds of units are used to represent electromagnetic radiation. These radiations are characterised by the properties, namely, frequency (ν) and wavelength (λ).

Electromagnetic Spectrum

When electromagnetic radiations are arranged in order of their increasing wavelengths or decreasing frequencies ,the complete spectrum obtained is called electromagnetic spectrum.

Cosmic rays < γ-rays <X-rays < ultraviolet-rays <visible-rays <infrared –rays <Micro waves <radio waves (CGXUVIMR).

WAVE RELATED TERM

Wave

The wave is short of disturbance which originates from vibrating source and travel onwards perpendicular to their axis as alternate crest and through known as wave.

- Wave is a form of energy.
- No medium required to propagate the wave.

Each wave has six character

Wavelength (λ):

The distance between two adjacent crests and troughs know as wavelength.

Frequency (ν):

Number of wave passes in one second known as frequency. It is reciprocal of wavelength and it's unit is Hz, m^{-1}, cm^{-1}

$$\nu = \frac{1}{\lambda}$$

Time Period:

The time taken by the wave to complete one cycle.

$$\nu = \frac{1}{T}$$

Wave Number:

Number of wavelength per unit length is known as wave number. It is reciprocal of wavelength

$$\bar{\nu} = \frac{1}{\lambda}$$

Relation Between Speed of light (c), frequency (ν) and wavelength (λ):

$$c = \nu\lambda, \quad \lambda = \frac{c}{\nu}, \quad \nu = \frac{c}{\lambda}$$

PRACTICLE NATURE OF ELECTROMAGNETIC RADIATION

Planck's Quantum Theory

The failure of the classical electromagnetic theory of radiation to explain the phenomenon of photoelectric effect led Max Planck (1901) to propose a new theory known as

Quantum Theory of Radiation.

The main points of this theory are :

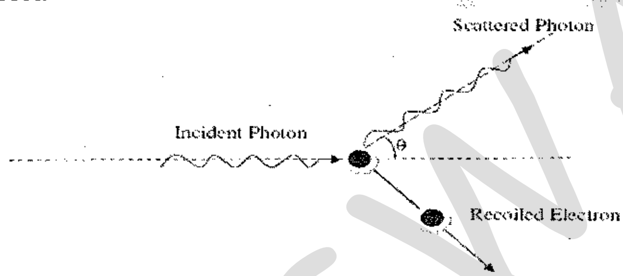
1. Radiant energy is emitted or absorbed not continuously but discontinuously in the form of small packets of energy called quanta.
   Note : In case of light, quanta is normally called Photons.
2. The amount of energy associated with a quantum of radiation is proportional to the frequency of radiation.

$$E \propto \nu$$
$$E = h\nu$$
$$E = \frac{hc}{\lambda}$$
$$\boxed{\nu = \frac{c}{\lambda}}$$

3. A body can emit or absorb energy only in terms of the integral multiples of quantum ,i-e:
$$E = nh\nu$$
Where n= 1,2,3,4,5........., h = Plank's constant 6.626×10^{-34} Js
4. Ejection of electron from metal surface when radiation strikes it.
5. Variation of heat capacity of solid as a function of temperature.
6. Line spectra of atoms with special reference to hydrogen.
7. The nature of emission of radiation from hot bodies.

Compton Effect

"The phenomenon of scattering of X-rays from suitable material and hence increase in its wavelength is called Compton Effect."



PHOTOELECTRIC EFFECT

Photoelectric effect may be defined as the phenomenon of ejection of electrons from the surface of a metal when light of suitable frequency strikes on it.

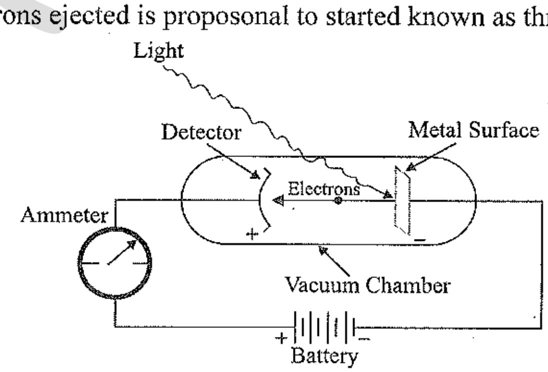
The electrons, thus, ejected are called **photoelectrons**

NOTE : Photoelectric effect is shown by any metal in a photoelectric cell only.

It comprises an evacuated tube with metal having low ionisation energy and constituting the negative electrode. Now when light of sufficiently high energy strikes the metal, electrons are knocked off from its surface and move towards the positive electrode. These electrons constitute a current flowing through the circuit.

Condition and Features:

1. The electrons are ejected from the metal surface as soon as the bean of light strikes the surface
2. The number of electrons ejected is proposonal to started known as thresold frequency (ν_0)



Equipment for studying the photoelectric effect. Light of a particular Frequency strikes a clean metal surface inside a vacuum chamber. Electrons are ejected from the metal and are counted by a detector that measures their kinetic energy.

Only a few metals show this effect under the action of visible light but many more can show it under the action of more energetic ultraviolet light.
Caesium with lowest I.E. emits electrons very easily.

$$h\nu = h\nu_0 + \frac{1}{2}mev^2$$

DUAL NATURE OF ELECTROMAGNETIC RADIATIONS

The photoelectric effect could be explained considering that electromagnetic radiations consist of small packets of energy called quanta or quantum. These packets of energy can be treated as particles.
On the other hand, radiations exhibit phenomena of interference and diffraction which indicate that they possess wave nature. So it may be concluded that electromagnetic radiations possess dual nature, i.e particle nature as well as wave nature.
Einstein even calculated the mass of the photon associated with a radiation of frequency ν as given below :
The energy E of the photon is given as

$$E = h\nu \qquad ...(i)$$

Also according to Einstein's equation,

$$E = mc^2 \qquad ...(ii)$$

where *m* is the mass of photon.
From equations (i) and (ii), we get

$$h\nu = mc^2$$

$$m = \frac{h\nu}{c^2} \qquad ...(iii)$$

$$m = \frac{h}{c^2} \cdot \frac{c}{\lambda} \qquad \left(\therefore \nu = \frac{c}{\lambda}\right)$$

$$m = \frac{h}{c\lambda} \qquad ...(iv)$$

Equations (iii) and (iv) can be used to calculate the mass of the photon.

Electronic Energy Levels

Atomic spectra

1. When an element is excited by some method such as by heating, by passing electric current or by passing electric discharge, the atoms of the element emit electromagnetic radiations of definite frequencies.
2. **This phenomenon of splitting of a beam of light into radiations of different frequencies after passing through the prism is called dispersion and the pattern of radiations obtained after dispersion of beam is called spectrum.**
3. **The radiations in the spectrum are emitted due to energy changes taking place in the atoms, this spectrum is also known as atomic spectrum.**
   Atomic spectrum of an element can be used to identify the element and is sometimes called fingerprint of its atoms.
4. The branch of science dealing with the study of spectra is called **spectroscopy.**
5. SPECTRUM IS BROADLY CLASSIFIED INTO -
   (A) Emission spectra
   (B) Absorption spectra

Differences

Emission spectra	Absorption spectra
1. It is obtained when the radiation from the source are directly analysed in the spectroscope.	1. It is obtained when the white light is passed through the substance and the transmitted light is analysed in the spectroscope.
2. It consists of bright coloured lines separated by dark spaces.	2. It consists of dark lines in otherwise continuous spectrum
3. It may give a continuous or discontinuous spectrum.	3. It is always discontinuous spectrum

Hydrogen Spectrum

1. When an electric discharge is passed through hydrogen gas its molecules dissociate into hydrogen atoms. The excited hydrogen atoms, thus produced, emit electromagnetic radiations of discrete frequencies.
2. The spectrum obtained consists of a large number of sharp lines. Each line corresponds to a particular frequency of light emitted by hydrogen atoms.
3. The lines in the **emission spectrum** of hydrogen are classified of hydrogen are classified into five series as follows :
   (i) **Lyman Series** ...Ultraviolet region
   (ii) **Balmer Series** ...Visible region
   (iii) **Paschen Series** ...Infra-red region
   (iv) **Brackett Series** ...Infra-red region
   (v) **Pfund Series** ...Infra-red region
   The complete spectrum of hydrogen is shown in Fig.

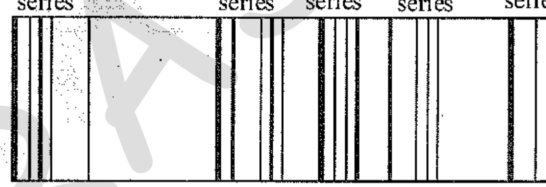


Fig. Atomic spectrum of hydrogen

4. The lines are grouped into different series, named after the discoverers.
5. **RYDBERG FORMULA**
   Although a large number of lines are present in the hydrogen spectrum , Rydberg gave a very simple theoretical equation for the calculation wavelengths and wave numbers of the lines by the formula

$$\bar{\nu} = \frac{1}{\lambda} = 109,677\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) cm^{-1}$$

when n_1 and n_2 are integers, such that $n_2 > n_1$. For a particular series n_1 is constant.

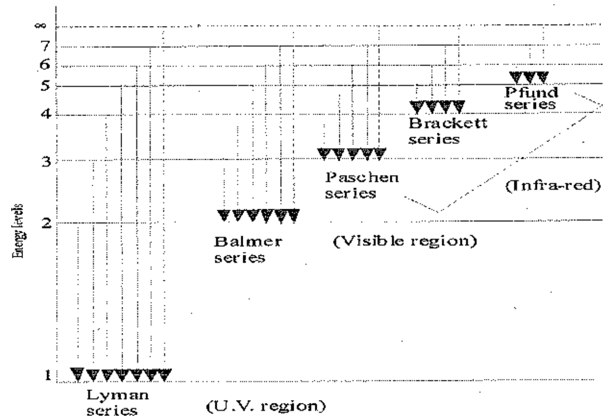


Fig. Generation of various spectral series in hydrogen spectrum.

Lyman Series	$n_1 = 1$	$n_2 = 2, 3, 4.....$
Balmer Series	$n_1 = 2$	$n_2 = 3, 4, 5.....$
Paschen Series	$n_1 = 3$	$n_2 = 4, 5, 6.....$
Brackett Series	$n_1 = 4$	$n_2 = 5, 6, 7.....$
Pfund Series	$n_1 = 5$	$n_2 = 6, 7, 8.....$

The value 109,677 cm^{-1} is called **Rydberg constant** for hydrogen.

For H – like particles

$$\bar{v}(cm^{-1}) = \frac{1}{\lambda} = RZ^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

where Z is atomic number of the H –like particles

LIMITING LINE : The limiting line of any spectral series in the hydrogen spectrum is the line when n_2 in the Rydberg's formula is infinity , i-e , $\mathbf{n_2 = \infty}$

BOHR'S MODEL FOR HYDROGEN ATOM

1. The electron in the hydrogen atom revolves around the nucleus only in certain elected circular orbits. These orbits are associated with definite energies and are also called **energy shells or energy levels.** These are numbered as 1, 2, 3, 4...etc. or designed as K, L, M, N....etc. shells (Fig).

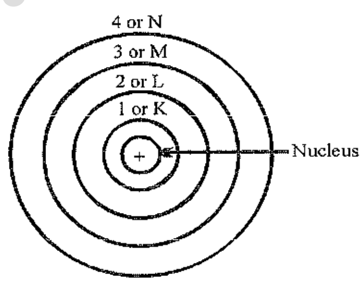


Fig. 2.23. Bhor's orbits.

2. Only those orbits are permitted in which the angular moment of the electron is a whole number multiple of h/2π where h is Planck's constant. That is, angular moment of the electron,

$$mvr = n\frac{h}{2\pi}, \text{where } n = 1, 2, 3....n.$$

where, *m* is the mass of the electron
v is the velocity of the electron
r is the radius of the orbit.
In other words, **angular moment of electrons in an atom is quantized.**

3. As long as the electron remains in particular orbit, it does not lose or gain energy. This means that energy of an electron in a particular orbit remains constant. That is why, these orbits are also called **stationary states.**
4. When energy from some external source is supplied to the electron, it may jump to some higher energy level by absorbing a definite amount of energy (equal to the difference in energy between the two energy levels). When the electron jumps back to the lower energy level it radiates same amount of energy in the form of a photon of radiation such that

$$\Delta E = E_2 - E_1 = h\nu$$

$$\nu = \frac{E_2 - E_1}{h}$$

where ν is the frequency of the radiation emitted when electron jump from the energy level having energy E_2 to the energy level having energy E_1.

$$E_n = -2.18\times10^{-18}\left(\frac{z^2}{n^2}\right)\text{J}$$

$$R_n = \frac{52.9n^2}{z}\text{Pm}$$

LIMITATIONS OF BOH'S MODEL

1. Bohr's model could not explain the spectra of atoms containing more than one electron. Bohr's model could not explain even hydrogen spectrum obtained using high resolution spectroscopes. Each spectral line, on high resolution was found to consist of two closely spaced lines.
2. It was observed that in the presence of a magnetic field, each spectral line gets split up into closely spaced lines. This phenomenon, known as **Zeeman effect,** could not be explained by Bohr's model. Similarly, the splitting of spectral line under the effect of applied electric field (**Stark effect**), could not be explained by Bohr's model.
3. The main objection to Bohr's theory was raised by **Heisenberg's uncertainty principle.** According to this principle, it is impossible to determine simultaneously the exact position and the momentum of a small moving particle like an electron. The postulate of Bohr that electrons move in well-defined orbits around the nucleus, therefore, is not valid. Thus, Bohr's model contradicts Heisenberg's uncertainty principle.
4. Bohr's model could not explain ability of atoms to form molecules and the geometry and shapes of molecules

Sommer's Field Atomic Model

Note : Sommer field atomic model explain the fine spectrum of H atom.

1. Sommer-field suggested that orbit may be elliptical or circular.
2. Sommer-field suggested that orbits are made up of sub-orbits or sub-shells. (s-p-d-f)

   s → sharp p → principle

   d → diffused f → fundamental

Defects : 1. This model doesn't explain Zeeman's and stark's effect.

2. This model explain the behaviour of those atoms and ions which contain only one electron.

The Dual Nature of Matter

Einstein had suggested in 1905 that light has dual nature, that is, wave nature as well as particle nature. Louis de Broglie, the French physicist proposed that like light, matter also has dual character. It exhibits wave as well as particle nature.

Louis de Broglie even derived a relationship for the calculation of wavelength (λ) of the wave associated with a particle of mass m, moving with velocity v as given below :

$$\lambda = \frac{h}{mv} \text{ or } \lambda = \frac{h}{p}$$

where h is Planck's constant and p is momentum of the particle.

Heisenberg's Uncertainity Principle

It is not possible to determine simultaneously the position and momentum of a small moving particle, such as electron, with entire certainty.

$$\Delta x \times \Delta p \geq \frac{h}{4\pi}$$

$$or \quad \Delta x \times \Delta v \geq \frac{h}{4\pi m}$$

where Δx = uncertainity of position
Δp = uncertainity in momentum
m = mass of the particle
Δv = uncertainity of velocity
h = Planck's constant.

Significance of Uncertainty Principle

1. It rules out existence of definite paths or trajectories of electrons and other similar particles. The trajectory of an object is determined by its location and velocity at various moments.
2. The effect of Heisenberg Uncertainty Principle is significant only for motion of microscopic objects and is negligible for that of macroscopic objects.

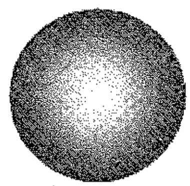
CONCEPT OF ORBITALS

An orbital may be defined as that region of space around the nucleus where the probability of finding an electron is maximum (90-95%).
Difference between an Orbit and an Orbital

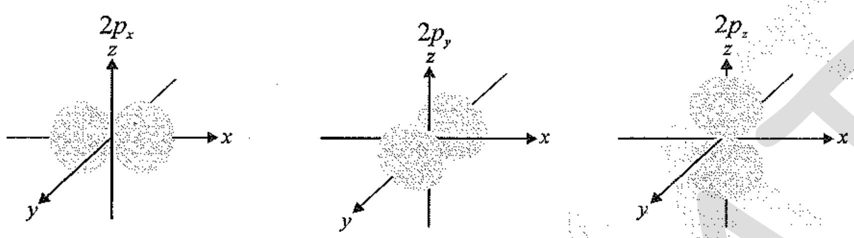
Orbit	Orbital
1. It is a well-defined circular path followed by revolving electron around the nucleus.	1. It is region of space around the nucleus of the atom where the electron is most likely to be found.
2. It is represents planar motion of an electron	2. It represents three-dimensional motion of an electron around the nucleus.
3. The maximum number of electrons in an orbit is $2n^2$, where n stands for number of the orbit.	3. An orbital cannot accommodate more than two electrons.
4. Orbits are circular in shape.	4. Orbitals have different shapes, e.g., s-orbitals are spherically symmetrical whereas p-orbitals are dumb-bell shaped.
5. Orbits are non-directional in character hence they cannot explain shapes of molecules.	5. Orbitals (except s-orbitals) have directional character and hence they can account for shapes of molecules.
6. Concept of well-defined orbit is against Heisenberg's principle.	6. Concept of orbitals is in accordance with Heisenberg's principal.

Shapes of Orbitals

- s subshell has spherical shape

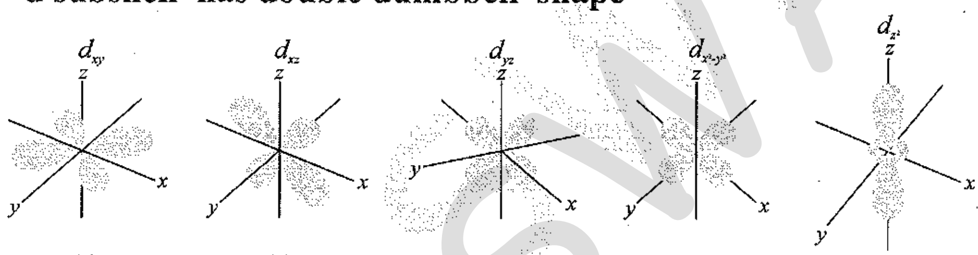


- **p subshell has dumbell shape**



(a) (b) (c)

- **d subshell has double dumbbell shape**



(a) (b) (c) (d) (e)

Pauli's Exclusion Principle

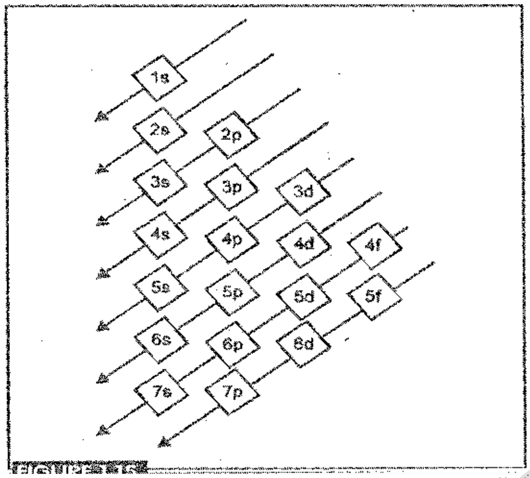
No two electrons in an atom can have same values for all the four quantum numbers

Or

Acc. to this principle an orbital can accommodate maximum of two electrons and these two electrons must have opposite spin

AUFBAU RULE

According to this rule,
the electrons are added progressively to the various orbitals in their order of increasing energies, starting with the orbital of lowest energy.
Increasing order of energies of various orbitals is :
1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s…



Aufbau is not a name of scientist; it is a german word meaning building up. It states that in the ground state of an atom an electron enters the orbital of lowest energy first & then electron enters in the order of increasing energy.

(n + *l*)'s rule

1. In neutral atom the orbital with lower value of (n + *l*) has lower energy than the orbital with higher (n + *l*) value. The e^- entering in a vacant subshell with lowest value of (n + *l*).
2. If two orbitals have some value of n + *l* than the orbital with lower value of n as has lower energy. If the value of (n + *l*) is same, then e^- enters the subshell having lowest value of n.

Exceptional Cases

At. No.	Element	Expected configuration	Actual configuration
24	Cr	$[Ar]^{18} 3d^4 4s^2$	$[Ar]^{18} 3d^5 4s^1$
29	Cu	$[Ar]^{18} 3d^9 4s^2$	$[Ar]^{18} 3d^{10} 4s^1$
41	Nb	$[Kr]^{36} 4d^3 5s^2$	$[Kr]^{36} 4d^4 5s^1$
42	Mo	$[Kr]^{36} 4d^4 5s^2$	$[Kr]^{36} 4d^5 5s^1$
44	Ru	$[Kr]^{36} 4d^6 5s^2$	$[Kr]^{36} 4d^7 5s^1$
45	Rh	$[Kr]^{36} 4d^7 5s^2$	$[Kr]^{36} 4d^8 5s^1$
46	Pd	$[Kr]^{36} 4d^8 5s^2$	$[Kr]^{36} 4d^{10} 5s^0$
47	Ag	$[Kr]^{36} 4d^9 5s^2$	$[Kr]^{36} 4d^{10} 5s^1$
64	Gd	$[Xe]^{54} 6s^2 4f^8 5d^0$	$[Xe]^{54} 6s^2 4f^7 5d^1$
78	Pt	$[Xe]^{54} 4f^{14} 5d^8 6s^2$	$[Xe]^{54} 4f^{14} 5d^9 6s^1$
79	Au	$[Xe]^{54} 4f^{14} 5d^9 6s^2$	$[Xe]^{54} 4f^{14} 5d^{10} 6s^1$

Other examples are La_{57}, Ce_{58}, R_{g111}.

It is because of half filled orbitals or fully filled orbitals have extra stability.

Quantum Mechanical Model of Atom

Quantum mechanics is a theoretical science that takes into account the dual nature of matter.

Quantum mechanics was developed independently in 1926 by Werner Heisenberg and Erwin Schrödinger. In the present context we shall deal with the quantum mechanics of Schrödinger.

Based on quantum mechanics a new model of atom was developed .This model is known as **quantum mechanical model.** In this model behaviour of the electron in an atom is described by an equation known as the **Schrödinger wave equation**

$$\frac{\partial^2\psi}{\partial x^2}+\frac{\partial^2\psi}{\partial y^2}+\frac{\partial^2\psi}{\partial^2 z}+\frac{8\pi^2 m}{h^2}(E-V)\psi=0$$

where, x, y and z are the three space co-ordinates,

m is mass of the electron,
h is the Planck's constant,
E is the total energy and V is the potential energy of the electron,

ψ (Greek letter psi) is amplitude of the electron wave and is called **wave function,** and $\frac{\partial^2 y}{\partial x^2}$ refers to the second derivative of ψ with equation is written in the form

$$H\psi = E\psi$$

where H is a mathematical operator known as **Hamiltonian operator.**

Significance of ψ AND ψ^2

Knowledge of $|\Psi|^2$ is helpful in assessing the probability of electron in a particular region. Thus, $|\Psi|^2$ is called **probability density** and $|\Psi|$ is referred to as **probability amplitude.**

ORBITALS AND QUANTUM NEMBER:

Atomic orbitals precisely distinguished by what are known as quantum numbers. Quantum nu ber is the complete address of electron in an atom just like postal address of any person.
It has four types

1. Principal Q.No. (n)	2. Azimuthal Q. No. (ℓ)
It is given by Bohr. It explain name of shell, energy and radius of shell The value start from K L M N (1 2 3 4)	It is given by Sommer field. It explain the sub shell S, P, D, F $\ell = 0$ 1 2 3 It explain shape and name of sub-shell. $\mu_m = \sqrt{\ell(\ell+1)}$ B.M
3. Marnetic Q.No. (m_ℓ)	**4. Spin Q.No. (s)**
It is given by Linde. It explaination of orbital. It explain the number of sub-shell $\mu_m = \sqrt{\ell(\ell+1)}\cos\theta$	It is given by Gold Smidth. It explain that the moving electron around the nucleus must has spin just like Earth. It explain the filling of electron. $\mu_s = \sqrt{s(s+1)}$ B.M

SUB-SHELL NOTATION

n	ℓ	Sub-Shell Notation
1	0	1s
2	0	2s
2	1	2p
3	0	3s
3	1	4p
3	2	3d

n	ℓ	Sub-Shell Notation
4	0	4s
4	1	4p
4	2	4d
4	3	4f

Value of 'ℓ'	0	1	2	3	4	5
Subshell Notation	s	p	d	f	g	h
Number of orbital	1	3	5	7	9	11

The Four Quantum number provide the following information

1. n-defines the shell, determines the size of the orbital and also to a large extent the energy of the orbital.
2. There are 'n' subshell in the n^{th} shell. ℓ quantify the subshell and determine the shape of orbital.
3. m_ℓ designates the orientation of orbital for a given value of 'ℓ' m_ℓ has $(2\ell + 1)$ values.
4. m_s refers orientation of the spin of the electron ⇅.

NODES AND NODAL PLANES

NODES – The region where probability of finding the electron is zero are called nodes.

NODAL PLANES

The plane where the probability of finding the electron is zero is known as nodal plane.

The no of nodal planes (angular nodes) in a orbital =*l*

The no. of nodal planes for s,p,d,f orbital are 0,1,2.3 respectively .

RADIAL NODES –The region where the probability of finding the electron is zero, radial nodes (i.e., probability density function is zero) **is known as radial nodes.**

Number of radial nodes in a given orbital are given as (n-*l*-1)

Thus for 1s no. of radial nodes = 1-0-1 = 0

2p no. of radial nodes = 2-1-1 = 0

Total number of nodes = (n-*l*-1)+*l*= n-1

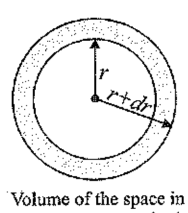
RADIAL WAVE FUNCTION AND RADIAL PROBABILITY FUNCTIONS

RADIAL WAVE FUNCTION (R) : Radial wave function (R) depends upon the distance (r) of the electron from the nucleus. The plots between the different values of r and corresponding values of R for 1s, 2s and 2p orbitals.

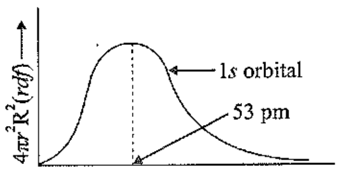
In all the cases as r approaches infinity, the radial wave function tends to become zero. There is a node in case of 2s radial function. At the node, the value of the radial wave function changes from positive (+) to negative (–). These signs simply indicate the sign of the wave functions at different distances from the nucleus in the same way as we notice the positive and negative signs of the amplitude in case of plane waves. They have no connection with normal positive and negative charges. These plots simply give information as to how the radial wave function changes with distance r from the nucleus and any node if present. In general, it has been found that *ns*-orbitals have (n – 1) nodes, *np*-orbitals have (n – 2) nodes etc.

RADIAL PROBABILITY FUNCTIONS ($4\pi r^2R^2$) : The plots between R^2 and r give the probable electron density around the nucleus. In case, the total probability of finding the electron in an infinitesimally small region is to be determined, then the volume of that region (dV) has also to be considered.

The product of the probable electron density and the corresponding volume of the space is known as **radial probability function.**



Volume of the space in between two concentric shells



Radial probability distribution curve for 1*s* electron orbital

Comparison of Radial Probability distribution curves for 1s, 2s and 2p electron orbitals :

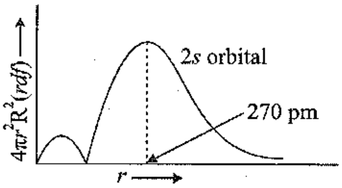
Following points are noteworthy when we compare the radial probability curves for the different electron orbitals.

(i) The radial probability distribution curve for 2s electron orbital shows two maxima, a smaller one

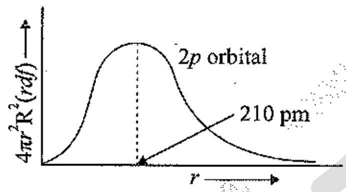
near the nucleus and a bigger one away from it at a larger distance. In between two maxima there is one minima where the probability of finding the electron is almost nil. This represents the **nodal point.**

(ii) The radius of maximum probability for 2s and 2p electron orbitals are more as compared to 1s electron orbital. This is quite expected also since these orbitals are bigger in size.

(iii) Although the radius of maximum probability for 2p electron (210 pm) is slightly less than for 2s electron (270 pm), but because of the presence a small additional maxima, 2s electron spends some more time near the nucleus than 2p electron. In other words, 2s electron is held more tightly by the nucleus than 2p electron. This is what we actually observe.



Radial probability distribution curve for 2*s* electron orbital



Radial probability distribution curve for 2*p* electron orbital

Stability of Completely Filled and Half Filled Subshell:

The ground state electronic configuration of atom more stable in half filled or full filled state.
Eg: The valence electronic configuration of Cr and Cu therefore $3d^5 4s^1$ and $3d^{10}4s^1$ not $3d^4 4s^2$ and $3d^9 4s^2$.

QUESTIONS

NCERT EXERCISE & EXAMPLES

RYDBERG FORMULA

1. What is the wavelength of light emitted when the electron in a hydrogen atom undergoes transition from an energy level with n=4 to an energy level with **n = 2** ? What is the colour corresponding to this wavelength ? (R_H = 109677 cm^{-1})

2. Calculate the wavelength of the first and the last line in the Balmer series of hydrogen spectrum.

DE BROGILE EQUATION

3. What will be the wavelength of a ball of mass 0.1 kg moving with a velocity of 10 m s^{-1} ?

4. The mass of an electron is 9.1×10^{-31} kg. If its K.E. is 3.0×10^{-25} J, calculate its wavelength.

5. Calculate the mass of a photon with wavelength 3.6 Å.

6. Calculate the wavelength of an electron moving with a velocity of 2.05×10^{7} m sec^{-1}.

7. The mass of an electron is 9.1×10^{-31} kg. If its K.E. is 3.0×10^{-25} J. Calculate its wavelength.

8. Show that the circumference of the Bohr orbit for the hydrogen atom is an integral multiple of the de Broglie wavelength associated with the electron revolving around the orbit.

9. Dual behavior of matter, proposed by de Broglie, led the discovery of electron microscope often used for the highly magnified images of biological molecules and other type of material. If the velocity of the electron in this microscope is 1.6×10^{6} m sec^{-1} , calculate de Broglie wavelength associated with this electron.

10. Similar to electron diffraction, neutron diffraction microscope is also used for the determination of the structure of molecules. If the wavelength used here is 800 pm, calculate the characteristic velocity associated with the neutron.

11. If the velocity of the electron in Bohr's first orbit 2.19×10^{6} m sec^{-1}, calculate the de Broglie wavelength associated with it.

12. The velocity associated with a proton moving in a potential difference of 10000 V is 4.37×10^{5} m sec^{-1}. If the hockey ball of mass 0.1 kg is moving with this velocity, calculate the wavelength associated with this velocity ?

ELECTROMAGNETIC SPECTRUM

13. The Vividh Bharati station of All India Radio, Delhi, broadcasts on a frequency of 1,368 kHz (kilo hertz). Calculate the wavelength of the electromagnetic radiation emitted by transmitter. Which part of the electromagnetic spectrum does it belong to?

14. The wavelength range of the visible spectrum extends from violet (400 nm) to red (750 nm). Express these wavelengths in frequencies (Hz). (1 nm = 10^{-9} m).

15. Calculate (a) wavenumber and (b) frequency of yellow radiation having wavelength 5800 Å.

16. Yellow light from a sodium lamp has a wavelength (λ) of 580 nm. Calculate frequency (v) and wave number($\bar{\nu}$) of the yellow light.

17. Calculate the wavelength, frequency and wave number of a light wave whose period is 2.0×10^{-10}S

18. Arrange the following type of radiations in increasing order of frequency :
(a) radiation from microwave oven
(b) amber light from traffic single
(c) radiation from radio
(d) cosmic rays from outer
(e) X-rays.

HEISENBERG'S UNCERTAINTY PRINCIPLE

19. A microscope using suitable photons is employed to locate an electron in an atom within a distance of 0.1 Å. What is the uncertainty involved in the measurement of its velocity?

20. A golf ball has a mass of 40g, and a speed of 45 m/s. If the speed can be measured within accuracy of 2%, calculate the uncertainty in the position.

HUND'S RULE OF MAXIMUM MULTIPLICITY

21. What is the total number of orbitals associated with the principal quantum number n = 3 ?

22. Using s, p, d, f notations, describe the orbital with the following quantum numbers (a) n = 2*l*

= 1, (b) n = 4l = 0, (c) n = 5l = 3, (d) n = 3l = 2

23. (i) Write the electronic configurations of the following ions :
(a) H^- (b) Na^+
(c) O^{2-} (d) F^-
(ii) What are the atomic number of elements whose outermost electrons are represented by :
(a) $3s^1$ (b) $2p^3$ and
(c) $3p^5$?
(iii) Which atoms are indicated by the following configurations ?
(a) [He] $2s^1$ (b) [Ne] $3s^2 3p^3$
(c) [Ar] $4s^2 3d^1$

24. What is the lowest value of n that allows g orbitals to exist ?

25. An electron is in one of the 3d orbitals. Give the possible values of n, l, and m_l for this electron.

26. (i) An atomic orbital has n = 3. What are possible values of l and m_l ?
(ii) List the quantum numbers (m_l and l) of electrons for 3d orbital.
(iii) Which of the following orbitals are possible? 1p, 2s, 2p and 3f

27. Using s, p, d notations, describe the orbital with the following quantum numbers.
(a) n = 1, l = 0; (b) n = 3, l = 1;
(c) n = 4, l = 2; (d) n = 4, l = 3.

28. Explain, giving reasons, which of the following sets of quantum numbers are not possible.
(a) n = 0, l = 0, m_l = 0, m_s = +1/2
(b) n = 1, l = 0, m_l = 0, m_s = −1/2
(c) n = 1, l = 1, m_l = 0, m_s = +1/2
(d) n = 2, l = 1, m_l = 0, m_s = −1/2
(e) n = 3, l = 3, m_l = -3, m_s = +1/2
(f) n = 3, l = 1, m_l = 0, m_s = +1/2

29. How many electrons in an atom may have the following quantum numbers ?
(a) n = 4 m_s = −1/2
(b) n = 3 l = 0

30. The quantum numbers of six electrons are given below. Arrange them in order of increasing energies. If any of these combination (s) has/have the same energy lists :
1. n = 4, l = 2, m_l = −2, m_s = −1/2
2. n = 3, l = 2, m_l = 1, m_s = +1/2
3. n = 4, l = 1, m_l = 0, m_s = +1/2
4. n = 3, l = 2, m_l = −2, m_s = −1/2
5. n = 3, l = 1, m_l = −1, m_s = +1/2
6. n = 4, l = 1, m_l = 0, m_s = +1/2

31. Among the following pairs of orbitals which orbital will experience the larger effective nuclear charge ?
(i) 2s and 3s, (ii) 4d and 4f,
(iii) 3d and 3p.

32. The unpaired electrons in Al and Si are present in 3p orbital. Which electrons will experience more effective nuclear charge from the nucleus?

33. Indicate the number of unpaired electrons in :
(a) P (b) Si
(c) Cr (d) Fe and
(e) Kr.

34. (a) How many sub-shells are associated with n = 4 ?
(b) How many electrons will be present in the sub-shells having m_s value of −1/2 for n=4?

HYDROGEN SPECTRUM

35. Calculate the energy associated with the first orbit of He^+. What is the radius of this orbit?

36. Calculate the wave number for the longest wavelength transition to Balmer series of atomic hydrogen.

37. How much energy is required to ionize a H atom if the electron occupies n = 5 orbit ? Compare your answer with the ionization energy of H atom (energy required to remove the electron n = 1 orbit).

38. What is the energy in joules, required to shift the electron of the hydrogen atom from the first Bohr orbit to the fifth Bohr orbit and what is the wavelength of the light emitted when the electron returns to the ground state ? The ground state electron energy is -2.18×10^{-11} ergs.

39. The electron energy in hydrogen atom is given by

$$E_n = -\frac{2.18 \times 10^{-18}}{(n)^2} J.$$

Calculate the energy required to remove an electron completely from n = 2 orbit. What is the longest wavelength of light in cm that can be used to cause this transition ?

40. What is the maximum number of emission lines when the excited electron of a hydrogen atom in n=6 drops to the ground state (in different hydrogen atoms).

41. (i) The energy associated with the first orbit in hydrogen atom is -2.18×10^{-18} J $atom^{-1}$. What is the energy associated with the fifth orbit ?

(ii) Calculate the radius of the Bohr's fifth orbit for H atom.

42. What transition is the hydrogen spectrum would have the same wavelength as the Balmer transition $n = 4$ to $n = 2$ of He^+ spectrum ?

43. Calculate the energy required for the process: $He^+(g) \longrightarrow He^{2+}(g) + e^-$. The ionization energy for the H atom in ground state is 2.18×10^{-18} J/atom.

PHOTO-ELECTRIC EFFECT

44. When electromagnetic radiation of wavelength 300 nm falls on the surface of sodium, electrons are emitted with a kinetic energy of 1.68×10^5 J/mol. What is the minimum energy needed to remove an electron from sodium? What is the maximum wavelength that will cause a photoelectron to be emitted ?

45. A photon of wavelength 4×10^{-7} m strikes on metal surface, the work function of the metal being 2.23 eV. Calculate :
(i) the energy of the photon (in eV)
(ii) the kinetic energy of emission
(iii) the velocity of the photoelectron.
$(1\,eV = 1.602 \times 10^{-19}\,J)$

46. Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength 6800 Å. Calculate threshold frequency (v_o) and work function (W_o) of the metal.

47. Electromagnetic radiation of wavelength 242 nm is just sufficient to ionize the sodium atom. Calculate the ionization energy of sodium in kJ mol^{-1}.

48. The work function for Caesium atom is 1.9 eV. Calculate
(a) the threshold wavelength and
(b) the threshold frequency of the radiation.
If the Caesium element is irradiated with a wavelength 500 nm, calculate the kinetic energy and the velocity of the ejected photoelectron.

49. The ejection of the photoelectron from the silver metal in the photoelectric effect experiment can be stopped by applying the voltage of 0.35 V when the radiation 256.7 nm is used.
Calculate the work function for silver metal.

50. If the photon of the wavelength 150 pm strikes an atom and one of its inner bound electrons is ejected out with a velocity of $1.5 \times 10^7\,m\,sec^{-1}$, calculate the energy with which it is bound to the nucleus.

PLANCK'S QUANTUM THEORY

51. Calculate energy of one mole of photons of radiation whose frequency is 5×10^{14} Hz.

52. A 100 watt bulb emits monochromatic light of wavelength 400 nm. Calculate the number of photons emitted per second by the bulb.

53. Find energy of each of the photons which
(i) correspond to light of frequency 3×10^{15} Hz
(ii) have wavelength of 0.50 Å

54. What is the number of photons of light with wavelength 4000 pm that provide 1J of energy?

55. Neon gas is generally used in the sign boards. If it emits strongly at 616 nm, calculate:
(a) the frequency of emission,
(b) distance traveled by this radiation in 30 s,
(c) energy of quantum and
(d) number of quanta present if it produces 2 J of energy.

56. In astronomical observations, signals observed from the distant stars are generally weak. If the photon detector receives a total of 3.15×10^{-18} J from the radiations of 600 nm, calculate the number of photons received by the detector.

MISCELLANEOUS

57. (i) Calculate the number of electrons which will together weigh one gram.
(ii) Calculate the mass and charge of one mole of electrons.

58. (i) Calculate the total number of electrons present in one mole of methane.
(ii) Find :
(a) the total number and
(b) the total mass of neutrons in 7 mg of ^{14}C. (Assume that mass of a neutron = 1.675×10^{-27} kg)
(iii) Find (a) the total number and (b) the total mass of protons in 34 mg of NH_3 at STP. Will the answer change if the temperature and pressure are changed ?

59. How many neutrons and protons are there in the following nuclei ?
$^{13}_{6}C,\ ^{16}_{8}O,\ ^{24}_{12}Mg,\ ^{56}_{26}Fe,\ ^{88}_{38}Sr$

60. Write the complete symbol for the atom with the given atomic number (Z) and atomic mass (A)
(i) $Z = 17, A = 35$ (ii) $Z = 92, A = 233$

(iii) Z = 4, A = 9.

61. A 25 watt bulb emits monochromatic yellow light of wave length of 0.57 μ m. Calculate the rate of emission of quanta per second.

62. Which of the following are isoelectronic species i.e., those having the same number of electrons? Na^+, K^+, Mg^{2+}, Ca^{2+}, S^{2-}, Ar

63. An atom of an element contains 29 electrons and 35 neutrons. Deduce
(i) the number of protons and
(ii) the electronic configuration of the element.

64. Give the number of electrons in the species H_2^+, H_2 and O_2^+.

65. If the diameter of a carbon atom is 0.15 nm, calculate the number of carbon atoms which can be placed side by side in a straight line across length of scale of length 20 cm long.

66. 2×10^8 atoms of carbon are arranged side by side. Calculate the radius of carbon atom if the length of this arrangement is 2.4 cm.

67. The diameter of zinc atom is 2.6 Å. Calculate
(a) radius of zinc atom in pm and
(b) number of atoms present in a length of 1.6 cm if the zinc atoms are arranged side by side lengthwise.

68. A certain particle carries 2.5×10^{-16} C of static electric charge. Calculate the number of electrons present in it.

69. In Milikan's experiment, static electric charge on the oil drops has been obtained by shining X-rays. If the static electric charge on the oil drip is -1.282×10^{-18} C, calculate the number of electrons present on it.

70. In Rutherford's experiment, generally the thin foil of heavy atoms, like gold, platinum etc. have been used to be bombarded by the α-particles. If the thin foil of light atoms like aluminium etc. is used, what difference would be observed from the above results ?

71. Symbols $^{79}_{35}Br$ and ^{79}Br can be written, whereas symbols $^{35}_{79}Br$ and $_{35}Br$ are not acceptable. Answer briefly.

72. An element with mass number 81 contains 31.7% more neutrons as compared to protons. Assign the atomic symbol.

73. An ion with mass number 37 possesses one unit of negative charge. If the ion contains 11.1% more neutrons than the electrons, find the symbol of the ion.

74. An ion with mass number 56 contains 3 units of positive charge and 30.4% more neutrons than electrons. Assign the symbol to this ion.

75. Nitrogen laser produces a radiation at a wavelength of 337.1 nm. If the number of photons emitted is 5.6×10^{-24}, calculate the power of this laser.

76. Lifetimes of the molecules in the excited states are often measured by using pulsed radiation source of duration nearly in the nano second range. If the radiation source has the duration of 2 ns and the number of photons emitted during the pulse source is 2.5×10^{-15}, calculate the energy of the source.

77. The longest wavelength doublet absorption transition is observed at 589 and 589.6 nm. Calculate the frequency of each transition and energy difference between two excited states.

78. Following results are observed when sodium metal is irradiated with different wavelengths. Calculate (a) threshold wavelength and (b) Planck's constant.

S.No.→	(i)	(ii)	(iii)
λ(nm)	400	500	450
$v \times 10^5 (ms^{-1})$	5.35	2.55	4.35

79. Emission transitions in the Paschen series end at orbit n = 3 and start from orbit n and can be represented as $v = 3.29 \times 10^{15}$ (Hz) $[1/3^2 - 1/n^2]$ Calculate the value of n if transition is observed at 1285 nm. Find the region of the spectrum.

80. Calculate the wavelength for the emission transition if it starts from the orbit having radius 1.3225 nm and ends at 211.6 pm. Name the series to which this transition belongs and the region of the spectrum.

81. If the position of the electron is measured within an accuracy of +0.002 nm, calculate the uncertainity in the momentum of the electron. Suppose the momentum of the electron is $h/4\pi_m \times 0.05$ nm, is there any problem in defining this value.

82. The bromine atom possesses 35 electrons. It contains 6 electrons in 2p orbital, 6 electrons in 3p orbital and 5 electron in 4p orbital. Which of these electrons experiences the lowest effective nuclear charge ?

SOLUTIONS

NCERT EXERCISE & EXAMPLES

Sol.1 $\bar{v} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] = 109677\left[\frac{1}{2^2} - \frac{1}{4^2}\right] cm^{-1}$

$= 20564.4\ cm^{-1}$ $[n_1 = 2; n_2 = 4]$

Now, $\lambda = \frac{1}{v} = \frac{1}{20564.4} = 486 \times 10^{-7}$ cm

$= 486 \times 10^{-9}$ m = 486 nm

The colour corresponding to this wavelength is blue.

Sol.2 Here we will use the following relation

$$\left(\frac{1}{\lambda}\right) = 1.09677\times10^7\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

For balmer series, $n_1 = 2$ and $n_2 = 3,4,5....$

The first line will be obtained for $n_2 = 3$. So, we will have

$$\left(\frac{1}{\lambda}\right) = 1.09677\times10^7\left[\frac{1}{4} - \frac{1}{9}\right]$$

Solving for λ, we get $\lambda = 6.563\times10^{-7}\,m$

Similarly, last line of Balmer series will be obtained for $n_2 = \infty$. So, we will have

$$\left(\frac{1}{\lambda}\right) = 1.09677\times10^7\left[\frac{1}{4} - \frac{1}{\infty}\right] = 1.09677\times10^7\left[\left(\frac{1}{4}\right) - 0\right]$$

Solving for λ, we get: $\lambda = 3.647\times10^{-7}\,m$

Sol.3 According to de Brogile equation

$$\lambda = \frac{h}{mv} = \frac{(6.266\times10^{-34}\,J\,s)}{(0.1\,kg)(10\,m\,s^{-1})}$$

$= 6.626\times10^{-34}\,m\,(J = kg\,m^2\,s^{-2})$

Sol.4 $\lambda = 8967\times10^{-10}\,m = 896.7\,nm$.

Sol.5 $\lambda = 3.6$ Å $= 3.6\times10^{-10}\,m$

Velocity of photon = velocity of light

$$m = \frac{h}{\lambda v} = \frac{6.626\times10^{-34}\,Js}{(3.6\times10^{-10}\,m)(3\times10^8\,ms^{-1})}$$

$= 6.135\times10^{-29}\,kg$

Sol.6 $\lambda = \frac{h}{mv}$

$$= \frac{6.63\times10^{-34}\,kg\,m^2\,s^{-1}}{9.1\times10^{-31}\,kg\times2.05\times10^7\,ms^{-1}} = 3.55\times10^{-11}\,m$$

Sol.7 K.E. $= \frac{1}{2}mv^2$

Velocity of the electron $= \left(\frac{2\times K.E.}{m}\right)^{1/2}$

$= \left[\frac{2\times3.0\times10^{-25}}{9.1\times10^{-31}}\right]^{1/2} = 0.812 \times 10^3\ ms^{-1}$

Now $\lambda = \frac{h}{m\times v} = \frac{6.625\times10^{-34}}{9.1\times10^{-31}\times0.812\times10^3}$

$= 0.8965 \times 10^{6}$ m

Wave length = l = $0.8965 \times 10^{-6} \times 10^{10}$ Å

= 8965 Å.

Sol.8 According to de-Broglie equation

$\lambda = \frac{h}{mv}$ ...(i)

de-Broglie pointed out such closed orbits can exist only when the circumference of orbit equals to an integral multiple of the wavelength of electron

i.e. $2\pi r = n\lambda$

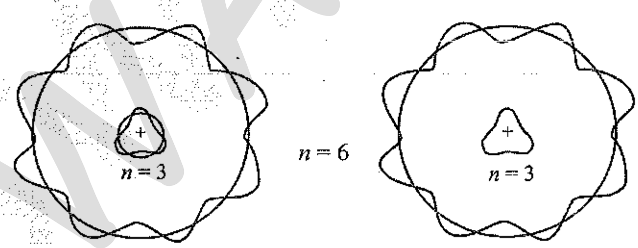


Fig. De, Broglie's waves

Substituting (i) from de-Broglie equation

$2\pi r = n\frac{h}{mv}$ or $mvr = n\frac{h}{2\pi}$

i.e. angular momentum is quantized for closed orbits. There is a stationary wave system for the de-Broglie waves for an electron in that orbit and the orbit must contain a whole number of electron waves. Fig. above clearly indicates 3 waves in third orbit, 6 waves in sixth orbits and so on.

Sol.9 Velocity, $v = 1.6\times10^6\,m\,sec^{-1}$

Mass of electron, $m = 9.11\times10^{-31}\,kg$

$\lambda = \frac{h}{mv}$

$$= \frac{6.626\times10^{-34}\,kg\,m^2\,s^{-1}}{9.11\times10^{-31}\,kg\times1.6\times10^6\,ms^{-1}} = 4.55\times10^{-10} = 455\,\text{pm}.$$

Sol.10 $\lambda = 800\ pm = 8.00\times10^{-10}\,m$

Mass of neutron, $m = 1.675\times10^{-27}\,kg$

$\lambda = \frac{h}{mv}$

$$v = \frac{h}{m\lambda} = \frac{6.626\times10^{-34}\,kg\,m^2\,s^{-1}}{1.675\times10^{-27}\,kg\times8.00\times10^{-10}\,m} = 4.95\times10^2\,m\,s^{-1}.$$

Sol.11 $v = 2.19\times10^6\,m\,sec^{-1}$, $m = 9.11\times10^{-31}\,kg$

$$\lambda = \frac{h}{mv} = \frac{6.626\times10^{-34}\,kg\,m^2\,s^{-1}}{9.11\times10^{-31}\,kg\times2.19\times10^6\,ms^{-1}} = 3.32\times10^2\,ms^{-1} = 332\,pm$$

Sol.12 Velocity of the hockey ball = $4.37\times10^5\,m\,s^{-1}$
Mass of the ball = 0.1 kg

$$\lambda = \frac{h}{mv} = \frac{6.626\times10^{-34}\,kg\,m^2\,s^{-1}}{0.1\,kg\times4.37\times10^5\,ms^{-1}} = 1.52\times10^{-38}\,m$$

This wavelength is too small to be detected by any means.

Sol.13 The wavelength, λ, is equal to c/v, where c is the speed of electromagnetic radiation in vacuum and v is the frequency. Substituting the given values, in

$$\lambda = \frac{c}{v} = \frac{3.00\times10^8\,m\,s^{-1}}{1368\,kHz} = \frac{3.00\times10^8\,m\,s^{-1}}{1368\times10^3\,s^{-1}}$$

= 219.3 m.
This is a characteristic radiowave wavelength.

Sol.14 Using equation 2.5, frequency of violet light

$$v = \frac{c}{\lambda} = \frac{3.00\times10^8\,m\,s^{-1}}{400\times10^{-9}\,m} = 7.50 \times 1014\text{ Hz}$$

Frequency of red light

$$v = \frac{c}{\lambda} = \frac{3.00\times10^8\,m\,s^{-1}}{750\times10^{-9}\,m} = 4.00\times10^{14}\text{Hz}$$

The range of visible spectrum is from 4.0×10^{14} to 7.5×10^{14} Hz in terms of frequency units.

Sol.15 (a) Calculation of wave number ($\bar{v}$)
$\lambda = 5800Å = 5800\times10^{-8}$ cm $= 5800\times10^{-8}$ m

$$\bar{v} = \frac{1}{\lambda} = \frac{1}{5800\times10^{-10}\,m} = 1.724\times10^6\text{ m}$$

$= 1.724\times10^4$ cm

(b) Calculation of the frequency (v)

$$v = \frac{c}{\lambda} = \frac{3\times10^8\,m\,s^{-1}}{5800\times10^{-10}\,m} = 5.172\times10^{14}\,s^{-1}$$

Sol.16 Wavelength (l) of yellow light = 580 nm = 580 $\times 10^{-9}$ m
Frequency of yellow light (v)

$$= \frac{c}{\lambda} = \frac{3\times10^8\,ms^{-1}}{580\times10^{-9}\,m} = 5.17\times10^{14}\,s^{-1}$$

Wavenumber $\{\bar{v}\}$ of the yellow light

$$\frac{1}{\lambda} = \frac{1}{580\times10^{-9}\,m} = 0.0172\times10^8\,m^{-1}$$

$= 1.72\times10^6\,m^{-1}$

Sol.17 (T) Period of the wave = 2.0×10^{-10} s
∴ Frequency =

$$v = \frac{1}{T} = \frac{1}{2.0\times10^{-10}} = 5\times10^9\,sec^{-1}$$

$c = 3\times10^8\,ms^{-1}$; Now $c = v \times l$;
$\therefore 3\times10^8 = 5\times10^9\times l$

$$\therefore \lambda = \frac{3\times10^8}{5\times10^9} = 6.0\times10^{-2}\text{ m}$$

Wave number

$$\bar{v} = \frac{1}{\lambda} = \frac{1}{6.0\times10^{-2}} = \frac{100}{6} = 16.66\,m^{-1}$$

Sol.18 The given radiations in increasing order of frequency – fmradio waves < Micro waves < amber colour < X-rays < Cosmic rays

Sol.19 $\Delta x\times\Delta p = \frac{h}{4\pi}$ or $\Delta x\times m\Delta v = \frac{h}{4\pi}$

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

$$\Delta v = \frac{6.626\times10^{-34}\,Js}{4\times3.14\times0.1\times10^{-10}\,m\times9.11\times10^{-31}\,kg}$$

$= 0.579\times10^7\,m\,s^{-1}$ ($1J = 1\,kg\,m^2\,s^{-2}$)
$= 5.79\times10^6\,m\,s^{-1}$

Sol.20 The uncertainty in the speed is 2%, i.e.,

$$\Delta v = 45\times\frac{2}{100} = 0.9\,m\,s^{-1}$$

Using the equation $\Delta x = \frac{h}{4\pi m\Delta v}$

$$= \frac{6.626\times10^{-34}\,Js}{4\times3.14\times40\,g\times10^{-3}\,kg\,g^{-1}(0.9\,m\,s^{-1})}$$

$= 1.46\times10^{-33}\,m$

This is nearly ~ 10^{18} times smaller than the diameter of a typical atomic nucleus. As mentioned earlier for large particles, the uncertainty principle sets no meaningful limit to the precision of measurements.

Sol.21 For n = 3, the possible values of l are 0, 1 and 2. Thus there is one 3s orbital (n = 3, l = 0 and ml = 0); there are three 3p orbitals (n = 3, l = 1 and m_l = –1, 0, + 1); there are five 3d orbitals (n = 3, l = 2 and m_l = –2, –1, 0, +1 +, + 2). Therefore, the total number of orbitals is 1 + 3 + 5 = 9
The same value can also be obtained by using the relation; number of orbitals = n^2, i.e. $3^2 = 9$.

Sol.22

	n	l	orbital
(a)	2	1	2p
(b)	4	0	4s
(c)	5	3	5f

(d) 3 2 3d

Sol.23 (i) (a) Electronic configuration of H^- = 1 + 1 = 2 = $1s^2$
(b) Na^+ = 11 - 1 = 10 = $1s^2, 2s^22p^6$
(c) O^{2-} = 8 + 2 = 10 = $1s^2, 2s^22p^6$
(d) F^- = 9 + 1 = 10 = $1s^2, 2s^22p^6$
(ii) (a) $3s^1$; It is sodium whose configurations is $1s^2, 2s^2 2p^6, 3s^1$. Its Atomic number is 11
(b) $2p^3$; It is nitrogen whose config. is $1s^2, 2s^2 2p^3$. Its Atomic number is 7
(c) $3p^5$; It is chlorine whose config. is $1s^2, 2s^2 2p^6, 3s^2 3p^5$. Its Atomic number is 17
(iii) (a) [He] $2s^1$, i.e., $1s^2\ 2s^1$. It is Lithium(Li) whose Z = 3.
(b) [Ne] $3s^2\ 3p^3$, i.e., $1s^2, 2s^2 2p^6, 3s^2 3p^3$. It is phosphorus (P) whose Z = 15 .
(c) [Ar] $4s^2\ 3d^1$, i.e., $1s^2, 2s^2 2p^6, 3s^2 3p^6, 4s^2 3d^1$. It is Scandium (Sc) whose Z = 21.

Sol.24 For n; the value of l are 0, 1, …(n - 1)
∴ For n = 5; the values of l are 0, 1, 2, 3, 4
For l = 4; g sub shell can exist.
Hence lowest value of n = 5.

Sol.25 When n = 3 l = 2 [Since it is in 3d subshell]
$m_l = -2, -1, 0, +1, +2$ [any one of the values]

Sol.26 (i)

l	ml
0	0
1	-1, 0, +1
2	-2 -1, 0, +1, +2

(ii) For 3d orbital l = 2 , The values of ml are -2 -1, 0, +1, +2.
(iii) Out of 1p, 2s, 2p and 3f only 2s, and 2p are possible.

Sol.27 (a) When n = 1, l = 0 1s orbital
(b) n = 3, l = 1 3p orbital
(c) n = 4, l = 2 4d orbital
(d) n = 4, l = 3 4f orbital.

Sol.28 (a) Not possible because n cannot be = 0
(b) When n = 1, l = 0, m_l = 0, m_s = +1/2 is possible
(c) When n = 1, l cannot be = 1. It is not possible
(d) Possible
(e) When n = 3; possible value of l can be = 0, 1, 2. ∴ $l \neq 3$ ∴ It is not possible
(f) n = 3, l = 1, m_e = 0, m_s = +1/2 is possible.

Sol.29 (a) n = 4; Total electrons possible are $2n^2 = 32$ Out of these 32 electrons, 16 electrons will have $m_s = -1/2$
(b) When n = 3, l = 0; Only 2 electrons are possible in 3s.

Sol.30 5 < 2 = 4 < 6 = 3 < 1.

Sol.31 (i) 2s, as it is nearer the nucleus
(ii) 4d
(iii) 3p

Sol.32 The unpaired electrons present in Aluminium and Silicon are in $3p^1$ and $3p^2$ orbitasls. Due to more nuclear charge in the case of Si [14p] than in Al [13p], electrons in Si will experience more nuclear charge than Al.

Sol.33 (a) P has 3 unpaired electrons [P = [Ne] $3s^2\ 3p^3$]
(b) Si has 2 unpaired electrons [Si = [Ne] $3s^2\ 3p^2$]
(c) Cr = 24 = [Ar] $3d^5\ 4s^1$ has 6 unpaired electrons.
(d) Fe = 26 = [Ar] $3d^6\ 4s^2$ has 4 unpaired electrons.
(d) Kr = 36 = [Ar] $3d^{10}\ 4s^2\ 4p^6$ has no unpaired electrons.

Sol.34 (a) When n = 4, l = 0, 1, 2, 3.

$$\left.\begin{array}{l} l = 0;\ 4s - \text{subshell} \\ l = 1;\ 4p - \text{subshell} \\ l = 2;\ 4d - \text{subshell} \\ l = 3;\ 4f - \text{subshell} \end{array}\right\} \text{These are 4 subshells}$$

(b) No. of electrons present in the subshells having m_s value of $-\frac{1}{2}$ for n = 4 are 16.

Sol.35 $E_n = -\frac{(2.18\times10^{-18} J)Z^2}{n^2} atoms$

For He+, n = 1, Z = 2

$$E_1 = -\frac{(2.18\times10^{-18} J)(2^2)}{1^2} = -8.72\times10^{-18} J$$

The radius of the orbit is given by equation

$$r_n = \frac{(0.0529\,nm)n^2}{Z}$$

Since n = 1, and Z = 2

$$r_n = \frac{(0.0529\,nm)1^2}{2} = 0.02645\,nm$$

Sol.36 For Balmer series n_1 = 2, Hence

$$\bar{\nu} = R\left(\frac{1}{2^2} - \frac{1}{n_2^2}\right)$$

Now $\bar{\nu} = \frac{1}{\lambda}$ Hence l will be longest when $\bar{\nu}$ is shortest, which in turn will be shortest if n_2 is shortest,
i.e., when n_2 = 3.

$$\therefore \bar{\nu} = 109677\left(\frac{1}{2^2} - \frac{1}{3^2}\right) = 109677 \times \frac{5}{36} cm^{-1} =$$

15232.9 cm^{-1} = 1.523 × 10^6 m^{-1}

Sol.37 Energy required to ionize H atom when the electron is in the 5th Orbit can be thus calculated :

$n_1 = 5; \quad n_2 = \infty$

$\therefore \; DE = E_2 - E_1 = -21.8\times10^{-19}\left(\frac{1}{n_2^2}-\frac{1}{n_1^2}\right)$

$= +21.8\times10^{-19}\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)$

$= +21.8\times10^{-19}\left(\frac{1}{5^2}-\frac{1}{\infty}\right)$

= +8.72 × 10^{-20} J atom^{-1}

∴ Energy required to ionize H atom from 5th orbit = +8.72 × 10^{-20} J

Ionization energy [i.e. energy required to remove the electron from n = 1 orbit] of H atom

$= \frac{21.79\times10^{-19}}{(1)^2}$ J atom^{-1}

= 2.179 × 10^{-18} J atom^{-1}

Thus, far higher energy is required to remove an electron from 1st orbit.

Sol.38 The ground state electron energy is the energy of the energy of the electron in the first Bohr orbit.

= −2.18 × 10^{-11} erg (given)

E_1 = −2.18 × 10^{-18} J [1 J = 10^7 ergs)

E_n = Energy in the nth orbit

$= \frac{-2.18\times10^{-18}}{n^2}$ Where n = no. of the orbit

$E_5 = \frac{-2.18\times10^{-18}}{5^2}$ = 8.72 × 10^{-20} J

Energy has to be absorbed to shift electron from Bohr's first orbit to fifth orbit.

$\therefore DE = E_5 - E_1 = -2.18\times10^{-18}\left[\frac{1}{5^2}-\frac{1}{1^2}\right]$

$= -2.18\times10^{-18}\left[\frac{24}{25}\right]$

= 2.08 × 10^{-18} J = 2.08 × 10^{-11} ergs.

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

$\lambda = \frac{hc}{\Delta E} = \frac{6.6\times10^{-34}\times3\times10^8}{2.08\times10^{-18}} = 956\ \text{Å}$

Wavelength of the light emitted = 956 Å.

Sol.39 For n = 2

$E_2 = -\frac{2.18\times10^{-18}}{2^2}J$ = −5.45 × 10^{-19} J

Energy required to remove this electron completely in the case of H atom

= +5.45 × 10^{-19} J

[∵ $E_\infty = 0$; So DE = $E_\infty - E_2$

= 0 − (− 5.45 × 10^{-19}) J = 5.45 × 10^{-19} J]

Now DE = hv = hc/l.

$\lambda = \frac{hc}{\Delta E} = \frac{6.66\times3\times10^8}{5.45\times10^{-19}}m$ = 3.647 × 10^{-7} m

= 3647 Å.

Sol.40 Maximum number of emission lines

$= \frac{n(n-1)}{2} = \frac{6(6-1)}{2} = 15$

Sol.41 (i) Energy of the nth orbit of H atom

$= \frac{-2.18\times10^{-18}}{n^2}$ J atom^{-1}

For first orbit $(E_1) = \frac{-2.18\times10^{-18}}{1^2}$

= −2.18 × 10^{-18} J atom^{-1} (given)

For the fifth orbit : E_5

$= \frac{-2.18\times10^{-18}}{5^2} = \frac{-2.18\times10^{-18}}{25}$

= −8.72 × 10^{-20} J

(ii) Calculate of the radius of the Bohr's fifth orbit for Hydrogen atom

$r_\circ = \frac{n^2h}{4\pi^2 mZe^2}$

For Hydrogen n = 1; Z = 1; putting usual values of pm, h, e radius ($r_\circ$) = 0.529 Å. This is known as Bohr's radius. For H−atom, the radius of any orbit (r_n)

= n^2 × $r_\circ$ where n = number of the orbit

= 5^2 × (0.529) Å = 1.3225 nm

Sol.42 For an atom $\bar{v} = \frac{1}{\lambda} = R_H Z^2\left[\frac{1}{n_1^2}-\frac{1}{n_2^2}\right]$

For He$^+$ spectrum, Z = 2, n_2 = 4, n_1 = 2

$\bar{v} = \frac{1}{\lambda} = R_H \times\left(\frac{1}{(2)^2}-\frac{1}{(4)^2}\right) = \frac{3R_H}{4}$

For hydrogen spectrum $\bar{v} = \frac{3R_H}{4}$ and Z = 1

$\therefore \bar{v} = \frac{1}{\lambda} = R_H\times1\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)$ or,

$R_H\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right) = \frac{3R_H}{4}$

or, $\frac{1}{n_1^2}-\frac{1}{n_2^2} = \frac{3}{4}$ which can be so for n_1 = 1 and n_2 = 2.

This means that n_1 = 1 and n_2 = 2.

Therefore, the transition is from to n = 1 in case of hydrogen spectrum.

Sol.43 For H-like particles. $E_n = -\frac{2\pi^2 mZ^2e^2}{n^2h_2}$
For H-atom :

$$I.E. = \frac{2\pi^2 me^2}{h^2} = 2.18 \times 10^{-18} \text{ J atom}^{-1} \text{ (given)}$$

For the given process,
Energy required =

$$E_\infty - E_1 = 0 - \left(-\frac{2\pi^2 m \times 2^2 \times e^4}{1^2 \times h^2} \right)$$

$$= 4 \times \frac{2\pi^2 me^4}{h^2} = 4 \times 2.18 \times 10^{-18} J$$

$$= 8.72 \times 10^{-18} \text{ J}$$

Sol.44 The energy (E) of a 300 nm photon is given by ; $E = hv = hc/\lambda$

$$= \frac{6.626 \times 10^{-34} Js \times 3.0 \times 10^8 ms^{-1}}{300 \times 10^{-9} m}$$

$$= 6.626 \times 10^{-19} J$$

The energy of one mole of photons

$$= 6.626 \times 10^{-19} J \times 6.022 \times 10^{23} mol^{-1} = 3.99 \times 10^5 Jmol^{-1}$$

The minimum energy needed to remove a mole of electrons from sodium

$= (3.99 - 1.68)\ 10^5$ J mol^{-1} $= 2.31 \times 10^5\ Jmol^{-1}$

The minimum energy for one electron

$$= \frac{2.31 \times 10^5 J mol^{-1}}{6.022 \times 10^{23} electron mol^{-1}} = 3.84 \times 10^{-19} J$$

This corresponds to the wavelength

$$\therefore \lambda = \frac{hc}{E} = \frac{6.626 \times 10^{-34} Js \times 3.0 \times 10^8 ms^{-1}}{3.84 \times 10^{-19} J}$$

= 517 nm(This corresponds to green light)

Sol.45 Energy (E) = hc/λ
h = 6.6 × 10^{-34} Js; c = 3 × 10^8 ms^{-1};
λ = 4 × 10^{-3}m

$$= \frac{6.6 \times 10^{-34} Js \times 3 \times 10^8 ms^{-1}}{4 \times 10^{-7} m} = 4.97 \times 10^{-19} \text{ J}$$

$$= \frac{4.97 \times 10^{-19}}{1.602 \times 10^{-19}} eV \quad [1 \text{ eV} = 1.602 \times 10^{-19} \text{ J}]$$

= 3.10 eV

(i) ∴ Energy of the photon in eV = 3.10

(ii) Kinetic energy of the emission of photoelectron (K.E.)
= hv − W_0 where W_0 = work function
= 3.10 eV − 2.13 eV = 0.97 eV

(iii) Velocity of the photoelectron

Since K.E. $= \frac{1}{2}mv^2$ or $v^2 = \frac{2 \times K.E.}{m}$
or

$$v = \sqrt{\frac{2 \times K.E.}{m}} = \sqrt{\frac{2 \times 0.97 \times 1.602 \times 10^{-19}}{9.1 \times 10^{-31}}}$$

= 5.84 × 10^5 ms^{-1}

Sol.46 K.E. = hv − hv_0
Since electrons are emitted with zero velocity
∴ K.E. = 0
∴ h (v − v_0) = 0 [v_0 = v]

$$v_0 = \frac{c}{\lambda} = \frac{3 \times 10^8}{6800 \times 10^{-10}} = 4.41 \times 10^{14} \text{ s}^{-1}$$

∴ Threshold frequency (v_0) = 4.41 × 10^{14} s^{-1}
Work function W_0 = hv_0
= 6.6 × 10^{-34} × 4.41 × 10^{14} J
= 2.91 × 10^{-19} J

Sol.47 Wave length λ = 242 nm = 242 × 10^{-9} m

Energy per photon (E) $= \frac{hc}{\lambda}$

$$= \frac{6.6 \times 10^{-34} Js \times 3 \times 10^8 ms^{-1}}{242 \times 10^{-9}} = 0.0821 \times 10^{-17} \text{J}$$

Since the energy is sufficient to cause ionization of one atom of Na
∴ Ionisation energy of 1 mole

$$= \frac{0.0821 \times 10^{-12} \times 6.02 \times 10^{23}}{1000} kJ = 494 \text{ kJ mol}^{-1}$$

Sol.48 W_0 = work function for Caesium = 1.9 eV
= 1.9 × 1.602 × 10^{-19} J
[1 eV = 1.602 × 10^{-19} J]
W_0 = hn_0 where n_0 = thereshold frequency

$$\therefore v_0 = \frac{W_0}{h} = \frac{1.9 \times 1.602 \times 10^{-19}}{6.625 \times 10^{-34}} s^{-1}$$

(b) ∴ Thereshold frequency
= 0.4594 × 10^{15} s^{-1} = 4.59 × 10^{14} s^{-1}

(a) Thereshold wavelength

$$\lambda_0 = \frac{c}{v_0} = \frac{3 \times 10^8}{4.59 \times 10^{14}} = 0.6536 \times 10^{-6} \text{ m}$$

= 654 nm

Kinetic energy = hn − W_0 $= \frac{hc}{\lambda} - W_0$

$$= \frac{6.625 \times 10^{-34} \times 3.0 \times 10^8}{500 \times 10^{-9}} - 3.04 \times 10^{-19}$$

3.975 × 10^{-19} − 3.04 × 10^{-19}
= 0.935 × 10^{-19} J

K.E. $= \frac{1}{2}mv^2$

$$\therefore v = \sqrt{\frac{2 \times K.E.}{m}} = \sqrt{\frac{2 \times 0.935 \times 10^{-19}}{9.1 \times 10^{-31}}}$$

$= \sqrt{0.20549 \times 10^{12}} = 4.53 \times 10^5\ ms^{-1}$

Sol.49 Energy of the incident radiation
= Work function + Kinetic energy of photoelectron
Energy of the incident radiation (E)
$= h\nu = hc/\lambda$

$$= \frac{(6.626 \times 10^{-34} Js)(3 \times 10^8 ms^{-1})}{(256.7 \times 10^{-9} m)}$$

$$= 7.74 \times 10^{-19} J = 4.83 eV \quad (1 eV = 1.602 \times 10^{-19} J)$$

Work function = 4.83 eV − 0.35 eV = 4.48 eV.

Sol.50 $l = 150$ pm $= 150 \times 10^{-12}$ m, v = velocity = $1.5 \times 10^7\ ms^{-1}$
∴ Kinetic energy of the ejected electron

$$= \frac{1}{2} mv^2$$

$$= \frac{1}{2} \times 9.1 \times 10^{-31} \times (1.5 \times 10^7) J = 0.102\ 10^{-19}\ J$$

Energy of the striking photon $E = h\nu = \frac{hc}{\lambda}$

$$= \frac{6.625 \times 10^{-34} \times 3.0 \times 10^8 J}{150 \times 10^{-12}} = 1.325 \times 10^{-15}\ J$$

∴ Minimum energy required to eject the electron
W_o = E − K.E. of the ejected the electron
$= [1.325 \times 10^{-15} - 0.102 \times 10^{-15}]$ J
$= 1.223 \times 10^{-15}$ J

$$= \frac{1.223 \times 10^{-15}}{1.602 \times 10^{-19}} eV \quad [\because 1 J = 1.602 \times 10^{-19} eV]$$

$= 7.6 \times 10^3$ eV
∴ Energy with which the electron is bound to the nucleus $= 7.6 \times 10^3$ eV.

Sol.51 Energy (E) of one photon is given by the expression $E = h\nu$
$h = 6.626 \times 10^{-34}$ J s, $\nu = 5 \times 10^{14}\ s^{-1}$ (given)
$E = (6.626 \times 10^{-34}\ J\ s) \times (5 \times 10^{14}\ s^{-1})$
$= 3.313 \times 10^{-19}$ J
Energy of one mole of photons
$= (3.313 \times 10^{-19}\ J) \times (6.022 \times 10^{23}\ mol^{-1})$
= 199.51 kJ /mol

Sol.52 100 watt means the bulb is emitting 100 Joules per second.
Power of the bulb = 100 watt = 100 J s^{-1}
Energy of one photon $E = h\nu = hc/\lambda$

$$= \frac{6.26 \times 10^{-34} Js \times 3 \times 10^8 ms^{-1}}{400 \times 10^{-9} m} = 4.969 \times 10^{-19} J$$

Number of photons emitted $= \frac{100 J s^{-1}}{4.969 \times 10^{-19} J}$

$= 2.012 \times 10^{20}\ sec^{-1}$

Sol.53 (i) To calculate the energy of a photon whose frequency (ν) $= 3 \times 10^{15}$ Hz $= 3 \times 10^{15}$ cps
(1 Hz = 1 cycle sec^{-1})
$E = h\nu$ where h = Planck's constant $= 6.6 \times 10^{-34}$ Js
$E = 6.6 \times 10^{-34} \times 3 \times 10^{15} = 1.98 \times 10^{-18}$ J

(ii) $E = h\nu = h\frac{c}{\lambda}\left[\because \nu = \frac{c}{\lambda_1}\right]$

$$= \frac{6.6 \times 10^{-34} Js \times 3 \times 10^8 ms^{-1}}{0.50 \times 10^{-10} m}$$

$= 3.98 \times 10^{-15}$ J.

Sol.54 Energy of photon $E_{photon} = h\nu = \frac{hc}{\lambda}$

$$= \frac{6.6 \times 10^{-34} Js \times 3.0 \times 10^8 ms^{-1}}{4000 \times 10^{-12} m}$$

$= 4.9687 \times 10^{-17}$ J

∴ No. of photons $= \frac{1.00}{4.9687 \times 10^{-17}}$

$= 2.012 \times 10^{16}$.

Sol.55 (a) Frequency (n) of emission can be calculated as : $c = n \times \lambda$
[where λ = wavelength = 616 nm $= 616 \times 10^{-9}$ m, c = velocity of light $= 3.0 \times 10^8\ ms^{-1}$]

$$\therefore \nu = \frac{c}{\lambda} = \frac{3.0 \times 10^8}{616 \times 10^{-9}} = 4.87 \times 10^{14}\ s^{-1}$$

(b) Distance travelled in 30 seconds
Distance travelled in 1 second
$= 3.0 \times 10^8$ m
Distance travelled in 30 sec $= 30 \times 3.0 \times 10^8$ m $= 9 \times 10^9$ m

(c) Energy of the photon (quantum)
$E = h\nu = 6.625 \times 10^{-34} \times 4.87 \times 10^{14}$ J
$= 32.26 \times 10^{-20}$ J

(d) Number of quanta $= \frac{\text{Total energy produced}}{\text{Energy of one quantum}}$

$$= \frac{2}{32.26 \times 10^{-20}} = 6.2 \times 10^{18}.$$

Sol.56 Total Energy = Energy received by photon detector $= 3.15 \times 10^{-18}$ J
Wavelength of radiations = 600 nm
$= 600 \times 10^{-9}$ m
Energy of one Quanta(E)

$$= \frac{hc}{\lambda} = \frac{6.6 \times 10^{-34} \times 3 \times 10^8}{600 \times 10^{-9}}$$

No. of photons = $\frac{Total\ Energy}{Energy\ of\ one\ Quanta}$

$= \frac{3.15\times10^{-18}\times600\times10^{-9}}{6.6\times10^{-34}\times3\times10^{8}} = 10$

Sol.57 (i) 9.1×10^{-28} g is the weight of 1 electron

$\therefore$ 1g is the wt. of $\frac{1}{9.1\times10^{-28}}$ electrons

$= 1.099 \times 10^{27}$ electrons.

No. of electrons weighing 1 gm = 1.099×10^{27} electrons.

(ii) Mass of one electron = 9.1×10^{-31} kg

$\therefore$ Mass of 1 mole (6.022×10^{23}) electrons

$= 9.1 \times 10^{-31} \times 6.022 \times 10^{23}$ kg

$= 5.48 \times 10^{-7}$ kg

Charge on one electron

$= 1.602 \times 10^{-19}$ coulomb

$\therefore$ Charge on 1 mole ($= 6.022 \times 10^{23}$) electrons

$= 1.602 \times 10^{-19} \times 6.022 \times 10^{23}$ coulomb

$= 9.65 \times 10^{4}$ C.

Sol.58 (i) 1 mole of methane (CH_4) contains 1 mole of C atoms and 4 moles of H atoms.

Each mole of C atoms contains 6 moles of electrons and each mole of H atoms contains 1

mole of electrons

$\therefore$ Total no. of electrons present in one mole of CH_4

$= (6 \times 6.022 \times 10^{23} \times 10^{23} + 4 \times 6.022 \times 10^{23})$ electrons

$= 3.614 \times 10^{24} + 2.409 \times 10^{24}$ electrons

$= 6.022 \times 10^{24}$ electrons

(ii) (a) Mass of a neutron = 1.675×10^{-27} kg

14.0 g of ^{14}C contains 1 mole of atoms of C–14. [each atom of C–14 contains 8 neutrons]

$\therefore$ 14.0 g of ^{14}C contain $8 \times 6.022 \times 10^{23}$ neutrons

7 mg ($= 7 \times 10^{-3}$ g) of ^{14}C contains

$\frac{8\times6.023\times10^{23}\times7\times10^{-3}}{14} = 2.409 \times 10^{21}$ neutrons

(ii) (b) Total mass of neutrons :

As in (a) above 7 mg of ^{14}C contains 2.409×10^{21} neutrons

The mass of 1 neutrons = 1.675 ´ 10^{-27} kg (given)

$\therefore$ Mass of 2.409×10^{21} neutrons (contained in 7 mg of ^{14}C)

$= 2.409 \times 10^{21} \times 1.675 \times 10^{-27}$

$= 4.0347 \times 10^{-6}$ kg

(iii) (a) The total number of protons in 34 mg of NH_3 17.0 g of NH_3 contains 1 mole of N atoms and 3 moles of H atoms.

i.e., 17.0 g of NH_3 contains 7 mole of protons of N and 3 moles of protons of H

$= 10$ moles of protons $= 10 \times 6.022 \times 10^{23}$ protons

34×10^{-3} g of NH_3 contains

$= \frac{6.022\times10^{24}}{17} \times 34 \times 10^{-3}$ protons

$= 1.2044 \times 10^{22}$ protons.

(b) No. of protons in 34×10^{-3} g of NH_3

$= 1.2044 \times 10^{22}$ protons.

Mass of 1 proton = 1.675 ´ 10^{-27}

$\therefore$ Mass of 1.2044×10^{22} protons

$= 1.675 \times 10^{-27} \times 1.2044 \times 10^{22}$ kg

$= 2.015 \times 10^{-5}$ kg

There is no effect of temperature and pressure.

Sol.59 No. of neutrons present in $^{13}_{6}C$ =

Mass no. - Atomic no. = 13 – 6 = 7

No. of protons present in $^{13}_{6}C$ = Atomic no. = 6

No. of neutrons present in $^{16}_{8}O$ = 16 – 8 = 8

No. of protons = 8

No. of neutrons present in $^{24}_{12}Mg$ = 24 – 12 = 12

No. of protons = 12

No. of neutrons present in $^{56}_{26}Fe$ = 56 – 26 = 30

No. of protons = 26

No. of neutrons present in $^{88}_{38}Sr$ = 88 – 38 = 50

No. of protons = 38

Sol.60 (i) Z = 17, A = 35

Since the no. of protons = 17 = no. of electrons

$\therefore$ The atom is chlorine $= Cl : ^{35}_{17}Cl$

(ii) Z = 92, A = 233 ; No. of protons = 92

$\therefore$ The atom is Uranium = U : $^{233}_{92}U$

(iii) Z = 4, A = 9; No. of protons = 4

$\therefore$ The atom is Beryllium = Be : $^{9}_{4}Be$

Sol.61 Energy of one photon (E) = hν $= \frac{hc}{\lambda}$

$= \frac{6.62\times10^{-34}\times3\times10^{8}}{0.57\times10^{-6}} = 3.48 \times 10^{-19}$ J

Now 25 watt = 25 Js^{-1}

$\therefore$ Rate of emission of quanta per second

$= \frac{25}{3.48\times10^{-19}} = 7.18 \times 10^{19}\ s^{-1}$

Sol.62 No. of electrons in $Na^+ = 10$ [11–1]
No. of electrons in $K^+ = 18$ [19–1]
No. of electrons in $Mg^{2+} = 10$ [12–2]
No. of electrons in $Ca^{2+} = 18$ [20–2]
No. of electrons in $S^{2-} = 18$ [16+2]
No. of electrons in Ar = 18
$\therefore$ Na^+, Mg^{2+} are isoelectronic [$10e^-$ each]
$\therefore$ Ca^{2+}, K^+, S^{2-}, Ar are isoelectronic [$18e^-$ each]

Sol.63 (i) No. of protons = No. of electrons = 35 (Since atom is electrically neutral.)
(ii) Atomic No. = 29
$\therefore$ Electronic configuration = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$

Sol.64 $H_2^+ = 2 - 1 = 1$; $H_2 = 2$; $O_2^+ = 16 - 1 = 15$.

Sol.65 Diameter of a C-atom = 0.15 nm = 15×10^{-9} m
Length of the straight line = 20 cm = 20×10^{-2} m
$\therefore$ No. of atoms of carbon which can be placed on 20.0 cm length $= \frac{20\times10^{-2}}{0.15\times10^{-9}} = 1.33 \times 10^9$

Sol.66 No. of atoms of C = 2×10^8
Length of this arrangement = 2.4 cm = 2×10^{-2}m
$\therefore$ Diameter of one carbon atom $= \frac{2.4\times10^{-2}}{2\times10^8}m = 1.2 \times 10^{-10}$ m

$\therefore$ Radius of C atom $= \frac{1.2\times10^{-10}}{2} = 0.6 \times 10^{-10}$m

= 0.06 nm.

Sol.67 (a) Diameter of zinc atom = 2.6 Å = 2.6×10^2 pm
Radius $= \frac{2.6\times10^2}{2} = 1.3 \times 10^2$ pm = 130 pm
(b) No. of atoms of zinc on 1.6 cm length
$= \frac{1.6\times10^{-2}}{2.6\times10^{-10}} = 0.615 \times 10^8 = 6.15 \times 10^7$.

Sol.68 Static electric charge carried by particle = 2.5×10^{-16} C
Charge carried by a single electron = 1.602×10^{-19} C
$\therefore$ No. of electrons in the given particle
$= \frac{2.5\times10^{-16}}{1.602\times10^{-19}} = 1.5605 \times 10^3$.

Sol.69 The static electric charge on oil drop = -1.282×10^{-18} C
Charge on one electron = -1.602×10^{-19} C
$\therefore$ No. of electrons present in oil drop
$= \frac{-1.282\times10^{-18}}{-1.602\times10^{-19}} = 8$

Sol.70 More number of α-particles will pass as the nucleus of the lighter atoms like aluminium is small, smaller number of α-particles will be deflected as the number of positive charges is less than on the lighter nuclei.

Sol.71 For a given element the number of protons is the same for its isotopes whereas the mass number can be different for the given atomic number.

Sol.72 Mass no. of the element = 81
Let the no. of protons in it = x
$\therefore$ no. of neutrons in it $= x + \frac{31.7}{100}x$
But Mass number is the sum of protons and neutrons
$\therefore x + x + \frac{31.7}{100}x = 81 \Rightarrow x\left(2 + \frac{31.7}{100}\right) = 81$
$231.7x = 8100$ or $x = 35$
no. of protons = 35; no. of neutrons = 81 – 35 = 46
$\therefore$ The symbol of the element is $^{81}_{35}Br$.

Sol.73 Mass number of the ion = 37
i.e. no. of protons + no. of neutrons = 37
This ion carries one unit of negative charge.
Let the no. of protons in it = x; no. of electrons = x + 1; no. of neutrons = 37 – x
Acc. to the question

$x + 1 + (x+1)\times\frac{11.1}{100} = 37 - x$ or

$(x+1)\left[1 + \frac{11.1}{100}\right] = 37 - x$

$(x+1)\frac{111.1}{100} = 37 - x$

$111.1\ (x + 1) = 3700 - 100\ x$

$211.1\ x = 3700 - 111.1 = 3588.9$ or

$x = \frac{3588.9}{211.1} = 17$

$\therefore$ No. of protons = 17
No. of electrons = 17 + 1 = 18
No. of neutrons = 37 – 17 = 20
This ion is : $^{37}_{17}Cl^{-1}$ (Chloride ion) and its symbol is $^{37}_{17}Cl^{-1}$.

Sol.74 Mass no. of ion = 56 *i.e.* n + p = 56
Let the no. of protons = x
No. of neutrons = $56 - x$
Since it carries 3 units of positive charge. no. of electrons = $x - 3$
$\therefore$ According to the question ;

$$x-3+(x-3)\times\frac{30.4}{100}=56-x$$

$$(x-3)\left[1+\frac{30.4}{100}\right]=56-x$$

$$230.4\,x-391.2=5600 \text{ or } x=\frac{5991.2}{230.4}=26$$

At. No. = p = 26

Hence the symbol of the ion will be ${}^{56}_{26}Fe^{3+}$.

Sol.75 Wavelength of Nitrogen laser radiation = 337.1 nm

(n) No. of photons emitted = 5.6×10^{-24}

Power of this laser E = nhν $=\frac{nhc}{\lambda}$

$$=\frac{5.6\times10^{24}\times6.625\times10^{-34}\times3.0\times10^{8}}{337.1\times10^{-9}}J$$

= 3.3×10^{6} J.

Sol.76 Duration of the radiation source = 2 ns (Period) = 2×10^{-9} s

Frequency $(\nu)=\frac{1}{Period}=\frac{1}{2\times10^{-9}}s^{-1}=0.5\times10^{9}\ s^{-1}$

No. of photons emitted = 2.5×10^{15}

∴ Energy of the source nhν

= $2.5 \times 10^{15} \times 6.626 \times 10^{-34} \times .5 \times 10^{9}$ J

= 8.28×10^{-10} J.

Sol.77 To find frequency (n_1); l_1 = 589 nm = 589×10^{-9} m. (given)

To find frequency (n_2); l_2 = 589.6 nm = 589.6×10^{-9} m (given)

$$\nu_1=\frac{c}{\lambda_1}=\frac{2\times10^{8}}{589\times10^{-9}}=5.0934\times10^{14}\ s^{-1}$$

$$\nu_2=\frac{c}{\lambda_2}=\frac{3\times10^{8}}{589.6\times10^{-9}}=5.0881\times10^{14}\ s^{-1}$$

Energy = hn : $E_1 = hn_1$, $E_2 = hn_2$

∴ $DE = E_1 - E_2 = h(n_1 - n_2)$

= $6.625 \times 10^{-34} \times .0053 \times 10^{14}$

= $0.035 \times 10^{-20} = 3.46 \times 10^{-22}$ J

Sol.78 Suppose threshold wavelength

$=\lambda_0 m=\lambda_0\times10^{-9}$ m

Then $h(\nu-\nu_0)=\frac{1}{2}m\nu^2$ or

$$hc\left(\frac{1}{\lambda}=\frac{1}{\lambda_0}\right)=\frac{1}{2}m\nu^2$$

Substituting the given results of the three experiments, we get

$$\frac{hc}{10^{-9}}\left(\frac{1}{500}-\frac{1}{\lambda_0}\right)=\frac{1}{2}m\left(2.55\times10^{6}\right)^2 \quad \text{...(i)}$$

$$\frac{hc}{10^{-9}}\left(\frac{1}{450}-\frac{1}{\lambda_0}\right)=\frac{1}{2}m\left(4.35\times10^{6}\right)^2 \quad \text{...(ii)}$$

$$\frac{hc}{10^{-9}}\left(\frac{1}{400}-\frac{1}{\lambda_0}\right)=\frac{1}{2}m\left(5.20\times10^{6}\right)^2 \quad \text{...(iii)}$$

Dividing equation (ii) by equation (i), wet

$$\frac{\lambda_0-450}{450\lambda_0}\times\frac{500\lambda_0}{\lambda_0-500}=\left(\frac{4.35}{2.55}\right)^2 \text{ or}$$

$$\frac{\lambda_0-450}{\lambda_0-500}=\frac{450}{500}\left(\frac{4.35}{2.55}\right)^2=2.619$$

or $\lambda_0-450=2.16\lambda-1309.5$ or

$1.619\lambda_0=859.5$ ∴ = 531 nm

Substituting this value in equation (iii), we get

$$\frac{h\times\left(3\times10^{8}\right)}{10^{-9}}\left(\frac{1}{400}-\frac{1}{531}\right)=\frac{1}{2}\left(9.11\times10^{-31}\right)\left(5.20\times10^{6}\right)^2$$

or $h=\mathbf{6.66\times10^{-34}\ J\,s}$

Sol.79 According to the question

$\nu = 3.29\times10^{15}\text{ (Hz)}\left[\frac{1}{3^2}-\frac{1}{n^2}\right]$ But $\nu=\frac{c}{\lambda}$

Hence $\frac{c}{\lambda}=3.29\times10^{15}\left[\frac{1}{3^2}-\frac{1}{n^2}\right]$

$$\frac{3.0\times10^{8}}{1285\times10^{-9}}=3.29\times10^{15}\left[\frac{1}{9}-\frac{1}{n^2}\right]$$

$$\therefore\frac{1}{9}-\frac{1}{n^2}=\frac{3.0\times10^{8}}{1285\times10^{-9}\times3.29\times10^{15}}=0.07096$$

or $\frac{1}{n^2}=\frac{1}{9}-0.07096=0.401511$

$\therefore n^2=\frac{1}{0.0401511}=24.9059$ ∴ n = 5

∴ The electron jumps from n = 5 to n = 3

Since this is a transition occurring in the Paschan series and Paschen series lie in the **infra-red portion of light.**

Sol.80 Radius of nth orbit of H-like particles

$$=\frac{0.529\,n^2}{Z}\mathring{A}=\frac{52.9\,n^2}{Z}pm$$

$r_1=1.3225\,nm=1322.5\,pm=52.9\,n_1^2$

$$r_2=211.6\,pm=\frac{52.9\,n_2^2}{Z}$$

$$\therefore\frac{r_1}{r_2}=\frac{1322.5}{211.6}=\frac{n_1^2}{n_2^2} \text{ or } \frac{n_1^2}{n_2^2}-6.25 \text{ or}$$

$\frac{n_1}{n_2} = 2.5$

$\therefore$ If $n_2 = 2, n_1 = 5$. Thus, the transition is from 5^{th} orbit to 2^{nd} orbit. It belongs to Balmer series.

$\bar{v} = 1.097 \times 10^7 m^{-1}\left(\frac{1}{2^2} - \frac{1}{5^2}\right)$

$= 1.097 \times \frac{21}{100} \times 10^7 m^{-1}$

or $\lambda = \frac{1}{v} = \frac{100}{1.097 \times 21 \times 10^7} m = 4.34 \times 10^{-7} m$

= 434 nm

It lies in the visible region.

Since this is the transition from $n_2 = 5$ to $n_1 = 2$, this transition belongs to the BALMER SERIES. It lies in visible region.

Sol.81 $\Delta x \times \Delta p = \frac{h}{4\pi}$

$0.002 \times 10^{-9} \times Dp = \frac{6.625 \times 10^{-34} \times 7}{4 \times 22}$

$\Delta p = \frac{6.625 \times 10^{-34} \times 7}{4 \times 22 \times 0.002 \times 10^{-9}} = 2.63 \times 10^{-23}$ kg ms^{-1}

Actual momentum

$\frac{h}{4\pi \times 0.05} = \frac{h}{4\pi \times 5 \times 10^{-11} m}$

$= \frac{6.626 \times 10^{-34} kgm^2 s^{-1}}{4 \times 3.14 \times 5 \times 10^{-11} m}$

$= 1.055 \times 10^{-24} kgms^{-1}$

This value cannot be defined as the actual magnitude is smaller than the uncertainty.

Sol.82 As we go away from the nucleus, effective nuclear pull goes on decreasing. Hence electrons present in 4p orbital experience the lowest effective nuclear charge.

QUESTION ALIKE

NUCLEAR STRUCTURE

1. Calculate the number of protons, neutrons and electrons in $^{88}_{35}Br$.

2. Find out the atomic number, mass number, number of protons, electrons and neutrons present in the element with the notation $^{238}_{92}U$.

3. The nuclear radius is of the order of 10^{-13} cm while atomic radius is of the order 10^{-8} cm. Assuming the nucleus and the atom to be spherical, what fraction of the atomic volume is occupied by the nucleus?

4. Calculate the percentage of higher isotope of neon which has atomic mass 202 and the isotopes have the mass numbers 20 and 22.

THE RELATION $c=\nu\lambda$ AND $\bar{\nu}=\frac{1}{\lambda}$

5. Calculate (a) wave number and (b) frequency of yellow radiations having wavelength of 5800 Å.

6. A particular radio station broadcasts at a frequency of 1120 kHz (kilohertz). Another radio station broadcasts at a frequency of 98.7 MHz (Megahertz). What are the wavelengths of the radiations from each station ?

PLANCK'S QUANTUM THEORY AND PHOTOELECTRIC EFFECT

7. Calculate the frequency and energy of a photon of radiation having wavelength 6000 Å.

8. A 100 watt bulb emits monochromatic light of wavelength 400 nm. Calculate the number of photons emitted per second by the bulb.

9. Calculate the kinetic energy of the electron ejected when yellow light of frequency $5.2\times10^{14}\,sec^{-1}$ falls on the surface of potassium metal. Threshold frequency of potassium of $5\times10^{14}\,sec^{-1}$.

10. When electromagnetic radiation of wavelength 300 nm falls on the surface of sodium, electrons are emitted with a kinetic energy of $1.68\times10^{5}\,J\,mol^{-1}$. What is the minimum energy needed to remove an electron from sodium ? What is the maximum wavelength that will cause a photoelectron to be emitted ?

RYDBERG FORMULA/BALMER FORMULA PHOTOELECTRIC EFFECT

11. Calculate the frequency and the wavelength of the radiation in nanometers emitted when an electron in the hydrogen atom jumps from third orbit to the ground state. In which region of the electromagnetic spectrum will this line lie ? (Rydberg constant $=109,677\,cm^{-1}$)

12. The wavelength of the first line in the Balmer series is 656 nm. Calculate the wavelength of the second line and the limiting line in Balmer series.

13. Calculate the wavelength of the spectral line obtained in the spectrum of Li^{2+} ion when the transition takes place between two levels whose sum is 4 and the difference is 2.

BOHR'S MODEL PHOTOELECTRIC EFFECT

14. Calculate the wavelength of the radiation emitted when an electron in a hydrogen atom undergoes a transition from 4th energy level to the 2nd energy level. In which part of the electromagnetic spectrum does this line lie ?

15. Calculate the energy associated with the first orbit of He^{+}. What is the radius of this orbit ?

16. Calculate the velocity of electron in the first Bohr orbit of hydrogen atom. Given that Bohr radius = 0.529 Å.
Planck's constant, $h=6.626\times10^{-34}\,J\,s$,
mass of electron
$=9.11\times10^{-31}\,kg$ and $1J=1\,kg\,m^2s^{-2}$

17. The ionization energy of He^{+} is 8.72×10^{-18} J atom^{-1}. Calculate the energy of the first stationary state of Li^{2+}.

DE-BROGLIE EQUATION PHOTOELECTRIC EFFECT

18. Calculate the wavelength associated with an electron (mass 9.1×10^{-31}kg) moving with a velocity of $10^{3}\,m\,sec^{-1}$
($h=6.6\times10^{-34}\,kg\,m^2\,sec^{-1}$).

19. A moving electron has 4.55×10^{-25} joules of kinetic energy. Calculate its wavelength (mass $=9.1\times10^{-31}$ kg and $h=6.6\times10^{-34}\,kg\,m^2\,s^{-1}$).

20. Calculate the mass of a photon with wavelength

3.6 Å.

21. Two particles A and B are in motion. If the wavelength associated with particle A is 5×10^{-8} m, calculate the wavelength associated with particle B if its momentum is half of A.

UNCERTAINTY PRINCIPLE PHOTOELECTRIC EFFECT

22. Calculate the uncertainty in the velocity of an electron if the uncertainty in its position is 1 Å or 100 pm $\left(100^{-1}\,\text{m}\right)$.

23. Calculate the uncertainty in the velocity of a wagon of mass 3000 kg whose position is known to an accuracy of ± 10 pm (Planck's constant $=6.63\times10^{-34}$ J s).

24. Calculate the uncertainty in the position of an electron if the uncertainty in its velocity is 5.7×10^{5} m / sec ($h=6.6\times10^{-34}\,\text{kg}\,\text{m}^2\text{s}^{-1}$, mass of the electron $=9.1\times10^{-31}$ kg).

25. A golf ball has a mass of 40 g and a speed of 45 m/s. If the speed can be measured within accuracy of 2%, calculate the uncertainty in position.

THE CALCULATION OF QUANTUM NUMBERS, DESGINATION OF ORBITALS PHOTOELECTRIC EFFECT

26. An electron is in a 4f orbitals. What possible values for the quantum numbers n, *l*, m and s can it have?

27. Write down the quantum number n, *l* and m for the following orbitals :

(i) $3d_{x^2-y^2}$ (ii) $4d_{z^2}$
(iii) $3d_{xy}$ (iv) $4d_{xz}$
(v) $2p_z$ (vi) $3p_z$

28. What designation is given to an orbital having
(i) n = 2, *l* = 1, (ii) n = 3, *l* = 0
(iii) n = 5, *l* = 3 and (iv) n = 4, *l* = 2 ?

29. Which of the following sets of quantum numbers are not permitted ?

(i) n = 2, *l* = 2, m = –1, s = + ½
(ii) n = 2, *l* = 1, m = –1, s = – ½
(iii) n = 2, *l* = 0, m = 0, s = 0
(iv) n = 2, *l* = 1, m = 2, s = + ½

30. Which of the following orbitals are not possible? 1p, 2s, 3f and 4d.

QUESTION ALIKE SOLUTIONS

Sol.1 Here, $Z = 35, A = 80$
$\therefore$ Number of protons = Atomic number = 35
Number of neutrons $= A - Z = 80 - 35 = 45$
As the atom is neutral, Number of electrons
= Number of protons = 35.

Sol.2 Atomic number (Z) = 92
Mass number (A) = 238
But we know that
Number of protons = Number of electrons
= Atomic number (Z)
$\therefore$ Number of proton = 92
and Number of electrons = 92
further, Number of neutrons = Mass number – Atomic number
$= A - Z = 238 - 92 = 146$.

Sol.3 The volume of a sphere
$= 4\pi r^3 / 3$ where r is the radius of the sphere.
$\therefore$ Volume of the nucleus $= 4\pi r^3 / 3$
$= 4\pi\left(10^{-13}\right)^3 / 3cm^3$

Similarly, volume of the atom
$= 4\pi r^3 / 3 = 4\pi\left(10^{-8}\right)^3 / 3cm^3$

$\therefore$ Fraction of the volume of atom occupied by the nucleus

$$= \frac{4\pi\left(10^{-13}\right)^3 / 3cm^3}{4\pi\left(10^{-8}\right)^3 / 3cm^3} = 10^{-15}$$

Sol.4 Suppose $^{22}Ne = x\%$.
Then $^{20}Ne = (100 - x)\%$
Average atomic mass

$$= \frac{x \times 22 + (100 - x)20}{1000} = 20.2$$

(Given)
or $22x + 2000 - 20x - 2020$
or $2x = 20$ or $x = 10$
i.e., $^{22}Ne = 10\%$

Sol.5 (a) Calculation of the wave number
Wave number $(\bar{v}) = 1/\lambda$
But $\lambda = 5800\,Å$ (Given)
$= 5800 \times 10^{-10}\,m$

$$\therefore\ \bar{v} = \frac{1}{5800 \times 10^{-10}\,m}$$

$= 1.72 \times 10^6\,m^{-1}$

(b) Calculation of the Frequency

Frequency $v = \frac{c}{\lambda}$

Substituting $c = 3 \times 10^8\,m/\sec$
and $\lambda = 5800 \times 10^{-10}\,m$
we get

$$v = \frac{3 \times 10^8\,m/\sec}{5800 \times 10^{-10}\,m}$$

$= 5.172 \times 10^{14}\,\sec^{-1}$ or cycles/sec or Hz.

Sol.6 (a) Calculation of wavelength corresponding to a frequency of 1120 kHz.

Wavelength, $\lambda = \frac{c}{v}$
Substituting $c = 3 \times 10^8\,m\,s^{-1}$
and $v = 1120\,kHz$ (Given)
$= 1120 \times 10^3\,cycles\,s^{-1}$
(1 kHz $= 10^3\,cycles\,s^{-1}$)
$= 1120 \times 10^3\,s^{-1}$

we have, $\lambda = \frac{3 \times 10^8\,m\,s^{-1}}{1120 \times 10^3\,s^{-1}}$
$= 267.85\,m$.

(b) Calculation of wavelength corresponding to a frequency of 98.7 MHz.

Wavelength, $\lambda = \frac{c}{v}$
Substituting $c = 3 \times 10^8\,m\,s^{-1}$
and $v = 98.7\,MHz$
$= 98.7 \times 10^6\,cycles\,s^{-1}$
(1 MHz $= 10^6\,cycles\,s^{-1}$)
we get, $\lambda = \frac{3 \times 10^8\,m\,s^{-1}}{98.7 \times 10^6\,s^{-1}}$
$= 30.095\,m$.

Sol.7 (i) Frequency $v = \frac{c}{\lambda}$
Substituting $c = 3 \times 10^8\,m\,s^{-1}$
$\lambda = 6000\,Å = 6000 \times 10^{-10}\,m$

we get $v = \frac{3 \times 10^8}{6000 \times 10^{-10}} = 5 \times 10^{14}\,s^{-1}$

(ii) Energy of the photon $E = hv$
Substituting $h = 6.625 \times 10^{-34}\,Js$
$v = 5 \times 10^{14}\,s^{-1}$
we get $E = 6.625 \times 10^{-34} \times 5 \times 10^{14}$
$= 3.3125 \times 10^{-19}\,J$

Sol.8 Power of the bulb = 100 watt
= 100 Js^{-1}

Energy of one proton, $E = hv = \frac{hc}{\lambda}$

$$= \frac{(6.626\times10^{-34} J s)\times(3\times10^{8} m s^{-1})}{400\times10^{-9} m}$$

$= 4.969\times10^{-19} J$

$\therefore$ Number of photons emitted

$$= \frac{100 J s^{-1}}{4.969\times10^{-19} J} = 2.012\times10^{20} s^{-1}$$

Sol.9 K.E. of the ejected electrons is given by

$$\frac{1}{2}mv^2 = hv - hv_0 = h(v - v_0)$$

$= 6.625\times10^{-34}(5.2\times10^{14} - 5.0\times10^{14})$

$= 6.625\times10^{-34}\times0.2\times10^{14}$ joules

$= 1.325\times10^{-20}$ joules.

Sol.10 Energy of a proton of radiation of wavelength 300 nm will be

$$E = hv = h\frac{c}{\lambda}$$

$$= \frac{(6.626\times10^{-34} J s)\times(3.0\times10^{8} m s^{-1})}{(300\times10^{-9} m)}$$

$= 6.626\times10^{-19} J$

$\therefore$ Energy of 1 mole of photons

$= (6.626\times10^{-19} J)\times(6.022\times10^{23} mol^{-1})$

$= 3.99\times10^{5} J mol^{-1}$

As $E = E_0 +$ K.E. of photoelectrons emitted.

$\therefore$ Minimum energy (E_0) required to remove 1 mole of electrons from sodium $= E - K.E.$

$= (3.99 - 1.68)10^{5} J mol^{-1}$

$= 2.31\times10^{5} J mol^{-1}$

$\therefore$ Minimum energy required to remove one electron

$$= \frac{2.31\times10^{5} J mol^{-1}}{6.022\times10^{23} mol^{-1}} = 3.84\times10^{-19} J$$

The wavelength corresponding to this energy can be calculated using the formula :

$$E = hv = h\frac{c}{\lambda}$$

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

$$= \frac{(6.626\times10^{-34} J s)\times(3.0\times10^{8} m s^{-1})}{(3.84\times10^{-19} J)}$$

$= 5.17\times10^{-7} m = 517\times10^{-9} m$

$= 517$ nm

which corresponds to the green light.

Sol.11 According to Rydberg formula,

$$\bar{v} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

Here, $R = 109,677 cm^{-1}$

$n_2 = 3$

$n_1 = 1$ (for ground state)

$$\therefore \bar{v} = 109,677\left(\frac{1}{1^2} - \frac{1}{3^2}\right) cm^{-1}$$

$$= 109,677\times\frac{8}{9} cm^{-1} = 97490.7 cm^{-1}$$

$$\lambda = \frac{1}{\bar{v}} = \frac{1}{97490.7 cm^{-1}}$$

$= 103\times10^{-7} cm$

$= 103\times10^{-7} m = 103$ nm

$$v = \frac{c}{\lambda} = \frac{3\times10^{8} m s^{-1}}{103\times10^{-9} m}$$

$= 2.91\times10^{15} s^{-1}$

The wavelength, as calculated above, lies in the ultraviolet region. Otherwise too, as the jump is on the

1st orbit, the line will belong to Lyman series and hence lie in the ultraviolet region.

Sol.12 According to Rydberg's formula,

$$\bar{v} = \frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

For the Balmer series, $n_1 = 2$ and for the 1st line, $n_2 = 3$

$$\therefore \frac{1}{656} = R\left(\frac{1}{2^2} - \frac{1}{3^2}\right) = R\times\left(\frac{1}{4} - \frac{1}{9}\right)$$

$$= R\times\frac{5}{36} = \frac{5R}{36} \quad \ldots(i)$$

For the second line, $n_1 = 2, n_2 = 4$

$$\therefore \frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{4^2}\right) = R\left(\frac{1}{4} - \frac{1}{16}\right)$$

$$= R\times\frac{3}{16} = \frac{3R}{16} \quad \ldots(ii)$$

Dividing (i) by (ii), we get

$$\frac{\lambda}{656} = \frac{5}{36}\times\frac{16}{3}$$

or $\lambda = 485.9 nm$

for the limiting line,

$n_1 = 2$, $n_2 = \infty$

$\therefore \ \frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{\infty^2}\right) = \frac{R}{4}$ ...(iii)

Dividing (i) by (iii), we get

$\frac{\lambda}{656} = \frac{5}{36} \times 4$

or $\lambda = 364.4\,nm$

Alternatively first calculate R from (i) and substitute (ii) and (iii).

Sol.13 Suppose the transition takes place between levels n_1 and n_2.

Then $n_1 + n_2 = 4$

and $n_2 - n_1 = 2$

solving these equations, we get

$n_1 = 1$, $n_2 = 3$

$\therefore \ \frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) Z^2$

For $Li^{2+}, Z = 3$

$\therefore \ \frac{1}{\lambda} = 109{,}667\,cm^{-1}\left(\frac{1}{1^2} - \frac{1}{3^2}\right) \times 3^2$

$= 109{,}677 \times \left(\frac{1}{1} - \frac{1}{9}\right) \times 9\,cm^{-1}$

$= 109677 \times 8\,cm^{-1}$

or $\lambda = \frac{1}{109677 \times 8\,cm^{-1}}$

$= 1.14 \times 10^{-6}$ cm

Sol.14 For hydrogen atom

$E_n = -\frac{21.8 \times 10^{-19}}{n^2} J\,atom^{-1}$

Energy emitted when the electron jumps from $n = 4$ to $n = 2$ will be given by

$\Delta E = E_4 - E_2 = 21.8 \times 10^{-19}\left(\frac{1}{2^2} - \frac{1}{4^2}\right)$

$= 21.8 \times 10^{-19} \times \frac{3}{16} = 4.0875 \times 10^{-19} J$

The wavelength corresponding to this energy can be calculated using the expression

$E = h\nu = h.\frac{c}{\lambda}$ $(\because c = \nu\lambda)$

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

$= \frac{(6.626 \times 10^{-34} J\,s)(3 \times 10^8\,m\,s^{-1})}{(4.0875 \times 10^{-19}\,J)}$

$= 4.86 \times 10^{-7}\,m$

$= 4863\,\mathring{A}$ (or 486.3 nm)

It lies in the visible region.

Sol.15 $E_n = -\frac{2.18 \times 10^{-18}\,Z^2}{n^2} J\,atom^{-1}$

For $He^+, n = 1, Z = 2$

$\therefore \ E_1 = -\frac{(2.18 \times 10^{-18} J)(2^2)}{12}$

$= -8.72 \times 10^{-18} J$

Radius of H-like particles is given by

$r_n = \frac{(0.0529)n^2}{Z} nm$

For $He^+, n = 1, Z = 2$

$\therefore \ r_1 = \frac{0.0529 \times 1^2}{2^2} = 0.02645\,nm$

Sol.16 $mvr = \frac{nh}{2\pi}$ or $v = \frac{nh}{2\pi mr}$

$= \frac{(1)(6.626 \times 10^{-34}\,Js)}{2 \times 3.14 \times (9.11 \times 10^{-31}\,kg) \times (0.529 \times 10^{-10}\,m)}$

$= 2.189 \times 10^6\,J\,s\,kg^{-1}\,m^{-1}$

But $1\,J = 1 kg\,m^2 s^{-2}$ *.

* $1\,J$ = work done = Force × Distance

= (Mass × Acceleration) × Distance

$= (kg\,m\,s^{-1}) \times m = kg\,m^2\,s^{-2}$

Hence, $v = 2.189 \times 10^6$ m s^{-1}

Sol.17 $E_n = -\frac{2\pi^2 m Z^2 e^4}{n^2 h^2}$

$= -K\frac{Z^2}{n^2}$ (K = constant)

I.E. of $He^+ = E_\infty - E_1$

$= 0 - \left(-K\frac{2^2}{1^2}\right) = 4K$

Hence, $4K = 8.72 \times 10^{-18} J\,atom^{-1}$ (Given)

or $K = 2.18 \times 10^{-10}\,J\,atom^{-1}$

For $Li^{2+}, Z = 3$ and for 1st stationary state, $n = 1$

$\therefore \ E_1 = -K\frac{Z^2}{n^2} = -2.18 \times 10^{-18} \times \frac{3^2}{1^2}$

$= -19.62 \times 10^{-18}$ J atom^{-1}

Sol.18 Here, we are given

$m = 9.1 \times 10^{-31}\,kg$

$v = 10^3\ m\ \sec^{-1}$

$h = 6.6 \times 10^{-34}\ kg\ m^2\ s^{-1}$

$\lambda = \frac{h}{mv} = \frac{6.6\times10^{-34}\ kg\ m^2\ s^{-1}}{(9.1\times10^{-31}\ kg)\times(10^3\ m\ s^{-1})}$

$= 7.25\times10^{-7}$ m

Sol.19 Here, we are given
Kinetic energy

i.e., $\frac{1}{2}mv^2 = 4.55\times10^{-25} J$

$m = 9.1\times10^{-31}\ kg$

$h = 6.6\times10^{-34}\ kg\ m^2\ s^{-1}$

$\therefore\ \frac{1}{2}\times(9.1\times10^{-31})v^2 = 4.55\times10^{-25}$

or $v^2 = \frac{4.55\times10^{-25}\times2}{9.1\times10^{-31}} = 10^6$

or $v = 10^3\ m\ \sec$

$\therefore\ \lambda = \frac{h}{mv}$

$= \frac{6.6\times10^{-34}\ kg\ m^2\ s^{-1}}{(9.1\times10^{-31}\ kg)\times10^3\ m\ s^{-1}}$

$= 7.5\times10^{-7}$ m

Alternatively,

$K.E. = \frac{1}{2}mv^2$ or $v = \sqrt{\frac{2K.E.}{m}}$

$\lambda = \frac{h}{mv} = \frac{h}{m}\times\sqrt{\frac{m}{2K.E.}}$

$= \frac{h}{\sqrt{2m\times K.E.}}$

$= \frac{6.6\times10^{-34}\ kg\ m^2 s^{-1}}{\sqrt{2\times9.1\times10^{-31} kg\times4.55\times10^{-25} J}}$

$= 7.25\times10^{-7}$ m $(J = kg\ m^2\ s^{-2})$

Sol.20 Here, $\lambda = 3.6\,\mathring{A} = 3.6\times10^{-10} m$. As photon travels with the velocity of light,
$v = 3.0\times10^8 m\ s^{-1}$
By de Broglie equation,

$\lambda = \frac{h}{mv}$

or $m = \frac{h}{\lambda v}$

$= \frac{6.26\times10^{-34}\ J\ s}{(3.6\times10^{-10} m)(3.0\times10^8 m\ s^{-1})}$

$= 6.135\times10^{-29}$ kg.

Sol.21 By de Broglie equation,

$\lambda_A = \frac{h}{p_A}$ and $\lambda_B = \frac{h}{p_B}$

$\therefore\ \frac{\lambda_A}{\lambda_B} = \frac{p_B}{p_A}$

But $p_B = \frac{1}{2}p_A$ (Given)

$\frac{\lambda_A}{\lambda_B} = \frac{1/2 p_A}{p_A} = \frac{1}{2}$

or $\lambda_B = 2\times\lambda_A = 2\times5\times10^{-8} m$

$= 10^{-7}$m

Sol.22 Here, we are given

$\Delta x = 10^{-10} m$

$h = 6.6\times10^{-34} kg\ m^2\ s^{-1}$

$m = 9.1\times10^{-31} kg$

Applying uncertainty principle,

$\Delta x.(m\Delta v) = \frac{h}{4\pi}$

i.e., $\Delta v = \frac{h}{4\pi\times m\times\Delta x}$

$= \frac{6.6\times10^{-34} kg\ m^2 s^{-1}}{4\times\frac{22}{7}\times9.1\times10^{-31} kg\times10^{-10} m}$

$= 5.77\times10^5$ m s^{-1}

Sol.23 Here, we are given $m = 3000 kg$
$\Delta x = 10\ pm = 10\times10^{-12} m = 10^{-11} m$

$\therefore$ By uncertainty principle,

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

$= \frac{6.63\times10^{-34} kg\ m^2 s^{-1}}{4\times\frac{22}{7}\times3000 kg\times10^{-11} m}$

$= 1.76\times10^{-27}$ m s^{-1}

Sol.24 Here, we are given

$\Delta v = 5.7\times10^5\ m\ s^{-1}$

$m = 9.1\times10^{-31} kg$

$h = 6.6\times10^{-34} kg\ m^2\ s^{-1}$

Substituting these values in the equation for uncertainty principle, i.e.,

$\Delta x\times(m\times\Delta v) = \frac{h}{4\pi}$,

we have, $\Delta x = \frac{h}{4\pi\times m\times\Delta v}$

$$= \frac{6.6\times10^{-34}\,kg\,m^2 s^{-1}}{4\times\frac{22}{7}\times9.1\times10^{-31}kg\times5.7\times10^5 m\,s^{-1}}$$

$$=1.0\times10^{-10}\,m$$

i.e., uncertainty in position $=\pm10^{-10}$ m

Sol.25 Uncertainty in speed $=2\%$ of $40\,m\,s^{-1}$

i.e., $\Delta v=\frac{2}{100}\times45=0.9\,m\,s^{-1}$

Applying uncertainty principle

$$\Delta x(m\times\Delta v)=\frac{h}{4\pi} \text{ or } \Delta x=\frac{h}{4\pi m\,\Delta v}$$

$$=\frac{6.626\times10^{-34}\,kg\,m^2 s^{-1}}{4\times3.14\times(40\times10^{-3}\,kg)(0.9\,m\,s^{-1})}$$

$$=1.46\times10^{-33}\,\text{m}$$

Sol.26 Since the electron is in a 4f orbital, the value of the principal quantum number, $n=4$.
For the f-orbital, the secondary quantum number, $l=3$.
The values of the magnetic quantum (m) are $-l$ to $+l$ including zero. Therefore, when $l=3, m$ has seven
values, i.e., $-3,-2,-1,0+1,+2$ and 3.
For each value of m, the spin quantum number, s has two values, i.e., $s=+1/2$ and $s=-1/2$.

Sol.27 (i) $n=3, l=2, =+2$
(ii) $n=4, l=2, m=0$
(iii) $n=3, l=2, m=-2$
(iv) $n=4, l=2, m=+1$
(v) $n=2, l=1, m=0$
(vi) $n=3, l=1, m=+1$

Sol.28 (i) $n=2, l=1$ means 2p-orbital
(ii) $n=3, l=0$ means 3s-orbital
(iii) $n=5, l=3$ means 5f-orbital
(iv) $n=4, l=2$ means 4d orbital

Sol.29 (i) This set of quantum numbers is not permitted since the value of l cannot be equal to n.
(ii) This set of quantum numbers is permitted.
(iii) This set of quantum numbers is not permitted because the values of spin quantum number cannot be zero.
(iv) The set of quantum numbers is also not permitted since the value of 'm' cannot be greater than 1.

Sol.30 (i) The first shell has only one sub-shell, i.e., 1s, which has only one orbital, i.e., 1s orbital. Therefore, 1p orbital is not possible.
(ii) The second sub-shell has two subshells, i.e., 2s and 2p. Therefore, 2s orbitals are possible.
(iii) The third subshell has three subshells, i.e., 3s, 3p and 3d. Therefore, 3f-orbitals are not possible.
(iv) The fourth shell has four subshells, i.e., 4s, 4p, 4d and 4f. Therefore, 4d-orbitals are possible.

Objective Questions

1. Compared to the mass of lightest nuclei, the mass of an electron is only (approximately)
(a) 1/80 (b) 1/800
(c) 1/1800 (d) 1/2800

2. Atoms consists of protons, neutrons and electrons. If the mass of neutrons and electrons were made half and two times respectively to their actual masses, then the atomic mass of ${}_6C^{12}$
(a) Will remain approximately the same
(b) Will become approximately two times
(c) Will remain approximately half
(d) Will be reduced by 25%

3. The total number of valence electrons in $4.2\,gm$ of N_3^- ion is (N_A is the Avogadro's number)
(a) $1.6N_A$ (b) $3.2N_A$
(c) $2.1N_A$ (d) $4.2N_A$

4. The argument that favours the particle nature of cathode rays is that they
(a) produce fluorescence.
(b) travel through vacuum.
(c) get deflected by electric and magnetic fields.
(d) cast shadows of objects present in their way.

5. Which of the following reaction led to the discovery of neutrons?
(a) ${}^{12}_{6}C + {}^{1}_{1}H \rightarrow {}^{14}_{7}N + {}^{1}_{0}n$
(b) ${}^{9}_{4}Be + {}^{4}_{2}He \rightarrow {}^{12}_{6}C + {}^{1}_{0}n$
(c) ${}^{11}_{5}B + {}^{2}_{1}D \rightarrow {}^{11}_{5}C + {}^{1}_{0}n$
(d) ${}^{8}_{4}Be + {}^{4}_{2}He \rightarrow {}^{11}_{6}C + {}^{1}_{0}n$

6. When a gold sheet is bombarded by a beam of ∝ - particles, only a few of them get deflected, whereas most go straight, undeflected. This is because
(a) force of attraction exerted on ∝ - particle by electrons is insufficient.
(b) volume of nucleus is smaller than atom.
(c) force of repulsion acting on a fast – moving ∝- particle is very small.
(d) neutrons have no effect on ∝ - particle.

7. The nucleus of an atom can be assumed to be spherical. The radius of the nucleus of mass number A is given by $1.25\times10^{-13}\times A^{1/3}cm$ Radius of atom is one $\AA$. If the mass number is 64, then the fraction of the atomic volume that is occupied by the nucleus is
(a) 1.0×10^{-3} (b) 5.0×10^{-5}
(c) 2.5×10^{-2} (d) 1.25×10^{-13}

8. Which of the following is true for Thomson's model of atom?
(a) The radius of an electron can be calculated by using this model.
(b) In an undisturbed atom, the electrons will be at their equilibrium position, where the attraction between the cloud of positive charge and the electrons balances their mutual repulsion.
(c) It can explain the existence of protons.
(d) None of these.

9. The energy of an electron in the first Bohr orbit of H atom is $-13.6eV$. The possible energy value(s) of the excited state(s) for electrons in Bohr orbits to hydrogen is(are)
(a) $-3.4eV$ (b) $-4.2eV$
(c) $-6.8eV$ (d) $+6.8eV$

10. Rutherford's scattering formula fails for very small scattering angles because
(a) the kinetic energy of ∝- particles is larger.
(b) the gold foil is very thin.
(c) the full nuclear charge of the target atom is partially screened by its electron.
(d) All of these.

11. The line spectrum of two elements is not identical because
(a) they do not have the same number of neutrons.
(b) they have dissimilar mass number.
(c) they have different energy level schemes.
(d) they have different number of valence electrons.

12. The series limit for Balmer series of H-spectra is
(a) 3800 (b) 4200
(c) 3646 (d) 4000

13. The first line in the Balmer series in the hydrogen atom will have the frequency
(a) $4.57\times10^{14}s^{-1}$
(b) $3.29\times10^{15}s^{-1}$
(c) $8.22\times10^{15}s^{-1}$
(d) $8.05\times10^{13}s^{-1}$

14. Einstein's theory of photoelectric effect is based on
(a) Maxwell's electromagnetic theory of light.
(b) Planck's quantum theory of light.
(c) Both a and b.

(d) None of these.

15. In photoelectric effect, the number of photoelectrons emitted is proportional to
(a) intensity of incident beam.
(b) frequency of incident beam.
(c) wavelength of incident beam.
(d) All of these.

16. Heisenberg's uncertainty principle rules out the exact simultaneous measurement of
(a) probability and intensity.
(b) energy and velocity.
(c) charge density and radius.
(d) position and velocity.

17. Uncertainty in the position of an electron (mass $= 9.1 \times 10^{-31}$ kg) moving with a velocity of 300 m s^{-1} accurate upon 0.001% will be (h = 6.63×10^{-34} Js)
(a) 19.2×10^{-2} m (b) 5.76×10^{-2} m
(c) 1.92×10^{-2} m (d) 3.84×10^{-2} m

18. Which phenomenon best supports the theory that matter has a wave nature?
(a) Electron momentum
(b) Electron diffraction
(c) Photon momentum
(d) Photon diffraction

19. If wavelength of photon is 2.2×10^{-11} m, $h = 6.6 \times 10^{-34}$ *J-sec,* then momentum of photon is
(a) $3 \times 10^{-23}\, kg\, ms^{-1}$
(b) $3.33 \times 10^{22}\, kg\, ms^{-1}$
(c) $1.452 \times 10^{-44}\, kg\, ms^{-1}$
(d) $6.89 \times 10^{43}\, kg\, ms^{-1}$

20. Wavelength associated with electron motion
(a) Increases with increase in speed of electron
(b) Remains same irrespective of speed of electron
(c) Decreases with increase in speed of e^-
(d) Is zero

21. The volume of space where probability of finding an electron is zero is known as
(a) orbit (b) orbital
(c) node (d) None of these

22. An ion that has 18 electrons in the outermost shell is
(a) Cu^+ (b) Th^{4+}
(c) Cs^+ (d) K^+

23. Which of the following combination of quantum numbers is not allowed?
(a) n = 3, $l = 0$, $m_1 = 0$
(b) n = 4, $l = 4$, $m_1 = 0$
(c) n = 3, $l = 1$, $m_1 = 1$
(d) None of these.

24. A region in space around the nucleus of an atom where the probability of finding the electron is maximum is called
(a) sublevel. (b) orbit.
(c) orbital. (d) electron shell.

25. Which orbital notation does not have spherical node?
(a) n = 2; $l = 0$ (b) n = 3; $l = 1$
(c) n = 3; $l = 0$ (d) n = 1; $l = 0$

Answer Key

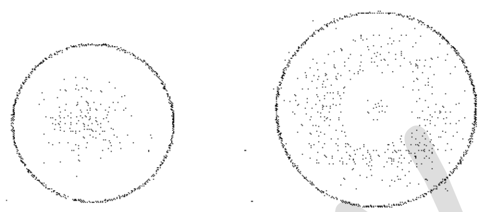
1. (c)	2. (d)	3. (a)
4. (c)	5. (b)	6. (b)
7. (d)	8. (a)	9. (b)
10. (d)	11. (c)	12. (c)
13. (a)	14. (b)	15. (a)
16. (d)	17. (c)	18. (b)
19. (a)	20. (c)	21. (c)
22. (a)	23. (d)	24. (c)
25. (b)		

Examplar Questions

1. Which of the following conclusions could not be derived from Rutherford's α -particle scattering experiement?
   (i) Most of the space in the atom is empty.
   (ii) The radius of the atom is about 10^{-10} m while that of nucleus is 10^{-15} m.
   (iii) Electrons move in a circular path of fixed energy called orbits.
   (iv) Electrons and the nucleus are held together by electrostatic forces of attraction.

2. Which of the following options does not represent ground state electronic configuration of an atom?
   (i) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^8\ 4s^2$ (ii) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^9\ 4s^2$
   (iii) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$ (iv) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$

3. The probability density plots of 1s and 2s orbitals are given in Fig. :



1*s* 2*s*

The density of dots in a region represents the probability density of finding electrons in the region.
On the basis of above diagram which of the following statements is incorrect?
   (i) 1*s* and 2*s* orbitals are spherical in shape.
   (ii) The probability of finding the electron is maximum near the nucleus.
   (iii) The probability of finding the electron at a given distance is equal in all directions.
   (iv) The probability density of electrons for 2*s* orbital decreases uniformly as distance from the nucleus increases.

4. Which of the following statement is **not** correct about the characteristics of cathode rays?
   (i) They start from the cathode and move towards the anode.
   (ii) They travel in straight line in the absence of an external electrical or magnetic field.
   (iii) Characteristics of cathode rays do not depend upon the material of electrodes in cathode ray tube.
   (iv) Characteristics of cathode rays depend upon the nature of gas present in the cathode ray tube.

5. Which of the following statements about the electron is incorrect?
   (i)I t is a negatively charged particle.
   (ii) The mass of electron is equal to the mass of neutron.
   (iii) It is a basic constituent of all atoms.
   (iv) It is a constituent of cathode rays.

6. Which of the following properties of atom could be explained correctly by Thomson Model of atom?
   (i) Overall neutrality of atom.
   (ii) Spectra of hydrogen atom.
   (iii) Position of electrons, protons and neutrons in atom.
   (iv) Stability of atom.

7. Two atoms are said to be isobars if.
   (i) they have same atomic number but different mass number.
   (ii) they have same number of electrons but different number of neutrons.
   (iii) they have same number of neutrons but different number of electrons.
   (iv) sum of the number of protons and neutrons is same but the number of protons is different.

8. The number of radial nodes for $3p$ orbital is __________.
(i) 3 (ii) 4
(iii) 2 (iv) 1

9. Number of angular nodes for $4d$ orbital is __________.
(i) 4 (ii) 3
(iii) 2 (iv) 1

10. Which of the following is responsible to rule out the existence of definite paths or trajectories of electrons?
(i) Pauli's exclusion principle.
(ii) Heisenberg's uncertainty principle.
(iii) Hund's rule of maximum multiplicity.
(iv) Aufbau principle.

11. Total number of orbitals associated with third shell will be __________.
(i) 2 (ii) 4
(iii) 9 (iv) 3

12. Orbital angular momentum depends on __________.
(i) l (ii) n and l
(iii) n and m (iv) m and s

13. Chlorine exists in two isotopic forms, Cl-37 and Cl-35 but its atomic mass is 35.5. This indicates the ratio of Cl-37 and Cl-35 is approximately
(i) 1:2 (ii) 1:1
(iii) 1:3 (iv) 3:1

14. The pair of ions having same electronic configuration is __________.
(i) Cr^{3+}, Fe^{3+} (ii) Fe^{3+}, Mn^{2+}
(iii) Fe^{3+}, Co^{3+} (iv) Sc^{3+}, Cr^{3+}

15. For the electrons of oxygen atom, which of the following statements is correct?
(i) Z_{eff} for an electron in a $2s$ orbital is the same as Z_{eff} for an electron in a $2p$ orbital.
(ii) An electron in the $2s$ orbital has the same energy as an electron in the $2p$ orbital.
(iii) Z_{eff} for an electron in $1s$ orbital is the same as Z_{eff} for an electron in a $2s$ orbital.
(iv) The two electrons present in the $2s$ orbital have spin quantum numbers m_s but of opposite sign.

16. If travelling at same speeds, which of the following matter waves have the shortest wavelength?
(i) Electron (ii) Alpha particle (He^{2+})
(iii) Neutron (iv) Proton

Examplar Solutions

1. (iii)	2. (ii)	3. (iv)	4. (iv)	5. (ii)	6. (i)
7. (iv)	8. (iv)	9. (iii)	10. (ii)	11. (iii)	12. (i)
13. (iii)	14. (ii)	15. (iv)	16. (ii)		

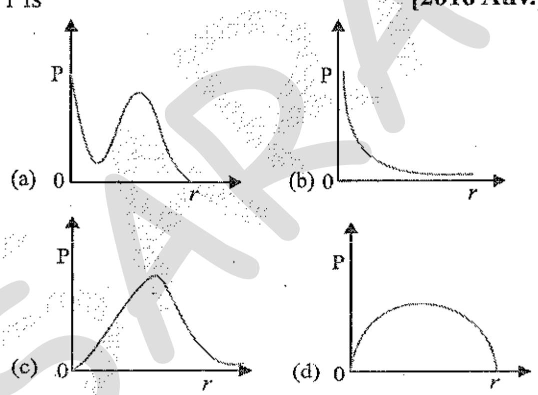
JEE and NEET Previous Year

Jee Previous Year

Q.1. For emission line of atomic hydrogen from $n_i = 8$ to $n_f = n$, the polt of wave number (v) against $\left(\frac{1}{n^2}\right)$ will be (The Rydberg constant, R_H is in wave number unit) **[2019 Mains]**
(a) non linear
(b) linear with slope-R_H
(c) linear with slope R_H
(d) linear with intercept-R_H

Q.2. The radius of the second Bohr orbit for hydrogen atom is (Planck's constant $(h) = 6.6262 \times 10^{-34}$ Js; mass of electron $= 9.1091 \times 10^{-31}$ kg; charge of electron $(e) = 1.60210 \times 10^{-19}$ C; permitivity of vacuum $(\in_0) = 8.854185 \times 10^{-12}$ kg^{-1} m^{-3} A^2) **[2017 Mains]**
(a) 1.65 Å (b) 4.76 Å
(c) 0.529 Å (d) 2.12 Å

Q.3. P is the probability of finding the 1s electron of hydrogen atom in a spherical shell of infinitesimal thickness, dr, at a distance r from the nucleus. The volume of this shell is $4\pi r^2 dr$. the qualitative sketch of the dependence of P on r is **[2016 Adv.]**



Q.4. A stream of electrons form a heated filament was passed between two charged plates kept at a potential difference V esu. If e and m are charge and mass of an electron, respectively, then the value of h/λ (where, λ is wavelength associated with electron wave) is given by **[2016 Mains]**
(a) $2\,meV$ (b) $\sqrt{meV}$
(c) $\sqrt{2meV}$ (d) meV

Q.5. Not considering the electronic spin, the degeneracy of the second excited state (n = 3) of H-atom is 9, while the degeneracy of the second excited state of H-atom is **[2015 Adv.]**

Q.6. In an atom, the total number of electrons having quantum numbers **[2014 Adv.]**
$n = 4$, $|m_l| = 1$ and $m_s = -\frac{1}{2}$ is

Q.7. Energy of an electron is given by $E = -2.178 \times 10^{-18}\,J\left(\frac{Z^2}{n^2}\right)$ Wavelength of light required to excite an electron in an hydrogen atom from level $n = 1$ to $n = 2$ will be ($h = 6.62 \times 10^{-34}$ Js and $c = 3.0 \times 10^8$ ms^{-1}) **[2013 Mains]**
(a) 1.214×10^{-7} m (b) 2.816×10^{-7} m
(c) 6.500×10^{-7} m (d) 8.500×10^{-7} m

Q.8. The atomic masses of He and Ne are 4 and 20 amu, respectively. The value of the de-Broglie wavelength of He gas at –73°C is 'M' times that of the de-Broglie wavelength of Ne at 727°C. M is **[2013 Adv.]**

Q.9. The kinetic energy of an electron in the second Bohr orbit of a hydrogen atom is [a_0 is Bohr radius] **[2012 Mains]**
(a) $\frac{h^2}{4\pi^2 ma_0^2}$ (b) $\frac{h^2}{16\pi^2 ma_0^2}$
(c) $\frac{h^2}{32\pi^2 ma_0^2}$ (d) $\frac{h^2}{64\pi^2 ma_0^2}$

Q.10. The work function (ϕ) of some metals is listed below. The number of metals which will show photoelectric effect when light of 300 nm wavelength falls on the metal is **[2011]**

Metal	Li	Na	K	Mg	Cu	Ag	Fe	Pt	W
ϕ (eV)	2.4	2.3	2.2	3.7	4.8	4.3	4.7	6.3	4.75

Q.11. The maximum number of electrons that can have principal quantum number, $n = 3$ and spin quantum number, $m_s = -1/2$, is **[2011]**

The hydrogen-like species Li^{2+} is in a spherically symmetric state S_1 with one radial node. Upon absorbing light the ion undergoes transition to a state S_2. The state S_2 has one radial node and its energy is equal to the ground state energy of the hydrogen atom. (12-13)

Q.12. The state S_1 is **[2010]**

(a) 1*s* (b) 2*s*
(c) 2*p* (d) 3*s*

Q.13. Energy of the state S_1 in units of the hydrogen atom ground state energy is **[2010]**

(a) 0.75 (b) 1.50
(c) 2.25 (d) 4.50

Q.14. The wavelength associated with a golf ball weighing 200 g and moving at a speed of 5 m/h is of the order **[2001]**

(a) 10^{-10} m (b) 10^{-20} m
(c) 10^{-30} m (d) 10^{-40} m

Q.15. The number of nodal planes in a p_x orbital is **[2001]**

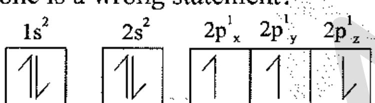
(a) one (b) two
(c) three (d) zero

NEET Previous Year

Q.16. Which of the following series of transitions in the spectrum of hydrogen atom fall in visible region? **[2019]**

(a) Paschen series (b) Brackett series
(c) Lyman series (d) Balmer series

Q.17. Which one is a wrong statement? **[2018]**



(a) Total orbital angular momentum of electron in 's' orbital is equal to zero.
(b) An orbital is designated by three quantum numbers while an electron in an atom is designated by four quantum numbers
(c) The value of m for $d_z{}^2$ is zero.
(d) The electronic configuration of N atom is

Q.18. Which one is the wrong statement? **[2017]**

(a) De–brogile's wavelength is given by $\lambda = \frac{h}{mv}$, where m= mass of the particle,v = group velocity of the particle.
(b) The uncertainity principle is $\Delta E \Delta t \geq \frac{h}{4\pi}$
(c) Half filled and fully filled orbitals have greater stability due to greater exchange energy, greater symmetry and more balanced arrangement.
(d) The energy of 2s orbital is less than the energy of 2p orbital in case of Hydrogen like atoms.

Q.19. Two electrons occupying the same orbital are distinguished by **[2016]**

(a) Principal quantum number
(b) Magnetic quantum number
(c) Azimuthal quantum number
(d) Spin quantum number

Q.20. Magnetic moments 2.84 B.M is given by (At nos, Ni = 28, Ti = 22, Cr = 24, Co = 27) **[2015]**

(a) Ni^{2+} (b) Ti^{3+}
(c) Cr^{2+} (d) CO^{2+}

Q.21. The number of d-electrons in Fe^{2+} (Z = 26) is not equal to the number of electrons in which one of the following? **[2015]**

(a) s-electrons in Mg (Z = 12)
(b) p – electrons in Cl (Z = 17)
(c) d – electrons in Fe (Z = 26)
(d) p – electrons in Ne (Z = 10)

Q.22. The angular momentum of electron in 'd' orbital is equal to **[2015]**

(a) $\sqrt{6}\,\hbar$ (b) $\sqrt{2}\,\hbar$
(c) $\sqrt{3}\,\hbar$ (d) $0\,\hbar$

Q.23. What is the maximum number of orbitals that can be identified with the following quantum numbers? n=3, l = 1, and m = 1 **[2014]**

(a) 1 (b) 2
(c) 3 (d) 4

Q.24. Calculate the energy in joule corresponding to light of wavelength 45 nm (Planck's constant, h = 6.63×10^{-34}Js speed of light , c = 3×10^{8} ms^{-1}) **[2014]**

(a) 6.67×10^{15} (b) 6.67×10^{11}
(c) 4.42×10^{-15} (d) 4.42×10^{-18}

Q.25. The value of Planck's constant is 6.63×10^{-34} Js. The speed of light is 3×10^{17}nm s^{-1}. Which value is closest to the wavelength in nanometer of a quantum of light with frequency 6×10^{15} s^{-1}? **[2013]**

(a) 10 (b) 25
(c) 50 (d) 75

Q.26. What is the maximum numbers of electrons that can be associated with the following set of quantum numbers? n = 3, l= 1 and m = –1 **[2013]**

(a) 10 (b) 6
(c) 4 (d) 2

Q.27. Maximum number of electrons ins subshell with l = 3 and n = 4 is **[2013]**

(a) 14 (b) 16

(c) 10 (d) 12

Q.28. The correct set of four quantum numbers for the valence electron of rubidium atom (at. no . = 37) is **[2012]**

(a) $5,1,1,+\frac{1}{2}$ (b) $6,0,0,+\frac{1}{2}$

(c) $5,0,0,+\frac{1}{2}$ (d) $5,1,0,+\frac{1}{2}$

Q.29. The energies E_1 and E_2 of two radiations are 25 eV and 50 eV respectiely. The relation between their wavelengths, i.e. λ_1 and λ_2 will be **[2011]**

(a) $\lambda_1 = 2\lambda_2$ (b) $\lambda_1 = 4\lambda_2$

(c) $\lambda_1 = \frac{1}{2}\lambda_2$ (d) $\lambda_1 = \lambda_2$

Q.30. If n = 6, the correct sequence of filling of electrons will be **[2011]**

(a) ns→(n–1)d→(n–2)f→np
(b) ns→(n–2)f→np→(n–1)d
(c) ns→np→(n–1)d→(n–2)f
(d) ns→(n–2)f→(n–1)d→np

Q.31. Which of the following is not permissible arrangement of electrons in an atom? **[2009]**

(a) n = 4, l = 0, m = 0, s = –1/2
(b) n = 5, l = 3, m = 0, s = +1/2
(c) n = 3, l = 2, m = –3, s = –1/2
(d) n = 3, l = 2, m = –2, s = –1/2

Q.32. The energy absorbed by each molecule [A_2] of a substance is 4.4×10^{-19} J and bond energy per molecule is 4×10^{-19} J. The kinetic energy of the molecule per atom will be **[2009]**

(a) 2.0×10^{-20} J (b) 2.2×10^{-19} J
(c) 2.0×10^{-19} J (d) 2.0×10^{-19} J

Q.33. Maximum number of electrons in a subshell is determined by the following **[2009]**

(a) 4l+2 (b) 2l+1
(c) 4l–2 (d) $2n^2$

Q.34. If uncertainly is position and momentum are equal, then uncertainty in velocity is

(a) $\frac{1}{2m}\sqrt{\frac{h}{\pi}}$ (b) $\sqrt{\frac{h}{2\pi}}$

(c) $\frac{1}{m}\sqrt{\frac{h}{\pi}}$ (d) $\sqrt{\frac{h}{\pi}}$

Q.35. The frequency of radiation emitted when the electron falls from n = 4 to n =1 in a hydrogen atom will be (given ionization energy of H = 2.18×10^{-18}Jatom^{-1} and h=6.625×10^{-34}J s **[2004]**

(a) $1.54\times10^{15}\,s^{-1}$ (b) $1.03\times10^{15}\,s^{-1}$
(c) $3.08\times10^{15}\,s^{-1}$ (d) $2.00\times10^{15}\,s^{-1}$

Q.36. The following quantum numbers and possible for how many orbital (s) n = 3, l = 2 and m = +2? **[2001]**

(a) 1 (b) 2
(c) 3 (d) 4

Answer Key

1. (c)	2. (c)	3. (c)	4. (c)	5. ()	6. (6)
7. (a)	8. (5)	9. (c)	10. ()	11. ()	12. (6)
13. (c)	14. (c)	15. (a)	16. (d)	17. (d)	18. (d)
19. (d)	20. (a)	21. (b)	22. (a)	23. (a)	24. (d)
25. (c)	26. (d)	27. (a)	28. (c)	29. (a)	30. (d)
31. (c)	32. (a)	33. (a)	34. (a)	35. (c)	36. (a)