Structure of Atom

- Atoms are composed of fundamental particles, electrons, protons and neutrons.

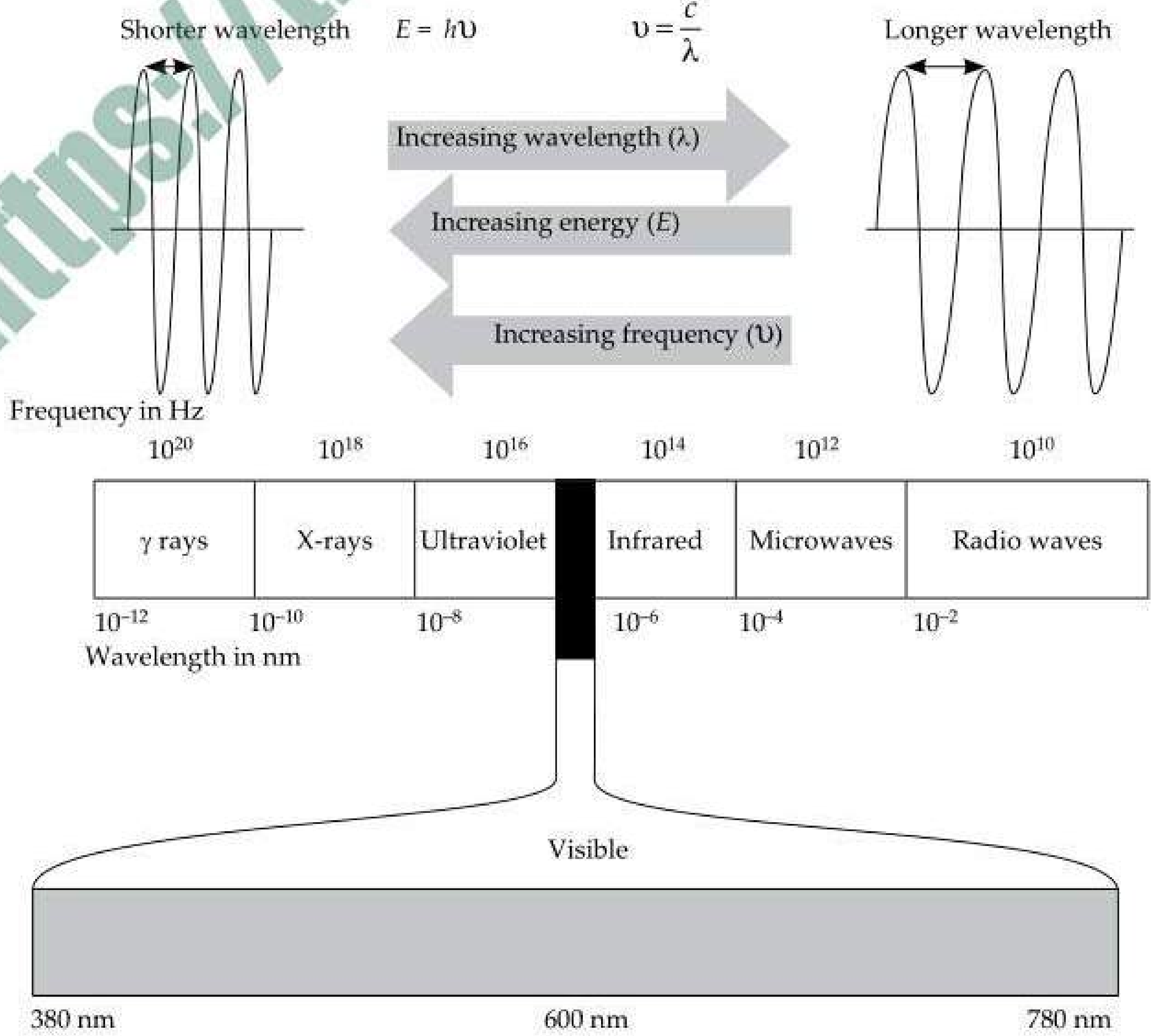
Properties of electron, proton and neutron

Properties	Electron, *e*	Proton, *p*	Neutron, *n*
Mass	9.101×10^{-31} kg	1.67262×10^{-27} kg	1.67495×10^{-27} kg
Charge	1.6022×10^{-19} C	1.6022×10^{-19} C	0
Mass relative to the electron	1	1836	1839
Spin	1/2	1/2	1/2
Charge relative to the proton	–1	+1	0
Discovery	Thomson	Goldstein	Chadwick

- **Nature of electromagnetic radiations :** Wave theory considers light to be a form of wave motion of wavelength λ, related to frequency υ and velocity of light c as $\upsilon = \frac{c}{\lambda}$

  Also, wavenumber $(\bar{\upsilon}) = 1/\lambda$

- Light waves were also considered electromagnetic in nature (i.e., they are oscillations of electric and magnetic fields in space) by James Maxwell in 1873.
- Various types of electromagnetic radiations in order of increasing wavelengths or decreasing frequencies is known as electromagnetic spectrum.



Electromagnetic spectrum

- **Quantum theory put forward by Planck :**
$$E = h\upsilon = \frac{hc}{\lambda}$$
- The energy is radiated or absorbed by a body not continuously but discontinuously in the form of small packets called quantum. In case of light, the quantum is called photon.
- **Moseley's equation :** Relates frequency of the X-rays produced to the charge present on the nucleus of an atom of the element used as anti-cathode.
$$\sqrt{\upsilon} = a(Z-b)$$
where, υ = frequency, Z = nuclear charge (atomic number) and a and b are constants.

Photoelectric Effect

- It is the phenomenon in which electrons are ejected when certain metals are exposed to a beam of light. For each metal, there is a characteristic minimum frequency, υ_0 (also known as threshold frequency) below which photoelectric effect is not observed.
- At a frequency $\upsilon > \upsilon_0$, the ejected electrons come out with certain kinetic energy.
$$h\upsilon = h\upsilon_0 + \frac{1}{2}mv^2$$
where, υ = Frequency of incident radiation
υ_0 = Threshold frequency
$h\upsilon_0$ = Energy of photon used in knocking out the electron from the atom (*i.e.* work function)
$h\upsilon$ = Energy of incident photon
$h\upsilon > h\upsilon_0$ and
$h(\upsilon - \upsilon_0)$ = Energy of photon converted into K.E.
- The energy of an individual photon depends only on its frequency and not on the intensity of the light beam.

Dual Nature of Matter

- $\lambda = \dfrac{h}{mc}$
According to de Broglie's equation,
$$\lambda = \frac{h}{p} = \frac{h}{mv} = \frac{h}{\sqrt{2mE_k}}$$
where E_k = Kinetic energy of the particle.
- According to Heisenberg's uncertainty principle,
$$\Delta x \cdot \Delta p \approx \frac{h}{4\pi} \text{ or } \Delta x \cdot \Delta v \geq \frac{h}{4\pi m}$$

Bohr's theory of Hydrogen Atom

- The main postulates are :
  - Atom consists of a small, heavy and positively charged nucleus in centre, and electrons revolve around the nucleus in fixed paths called orbits.
  - Energy of an electron in the orbit does not change with time.
  - The electron can revolve only in those orbits whose angular momentum is an integral multiple of $h/2\pi$
  i.e., $mvr = \dfrac{nh}{2\pi}$, $n = 1, 2, 3, \ldots$
  - When electron jumps from one level to another, energy is either emitted or absorbed.
  The energy difference between two states is given by $\Delta E = E_2 - E_1$
  As the distance of the orbits increases from the nucleus, the energy gap goes on decreasing, *i.e.*,
  $$E_2 - E_1 > E_2 - E_3 > E_3 - E_4 > \ldots\ldots$$
- **Energy of n^{th} orbit :**
$$E_n = -R_H\left(\frac{Z^2}{n^2}\right), \quad n = 1, 2, 3, \ldots$$
$$E_n = \frac{-2\pi^2 me^4 Z^2}{n^2h^2} = \frac{-2.18\times10^{-18}\times Z^2}{n^2} \text{ J/atom}$$
$$= \frac{-1312Z^2}{n^2} \text{ kJ mol}^{-1}$$
Here Z = nuclear charge for H-atom and single electron species like He^+, Li^{2+}, Be^{3+}, ..., etc.
- **Bohr's radius :** $r = \dfrac{n^2h^2}{4\pi^2 mkZe^2}$
For H-atom, n = 1, Z = 1
$\therefore$ r = 0.529 Å
In general, $r_n = \dfrac{0.529n^2}{Z}$ Å
- Number of waves in n^{th} orbit = $\dfrac{2\pi r}{\lambda}$
- **Rydberg's equation :** $\bar{v} = \dfrac{1}{\lambda} = R\left(\dfrac{1}{n_1^2} - \dfrac{1}{n_2^2}\right)$
where, R = Rydberg constant = 109677 cm^{-1}
when electron drops from a higher orbit (n_2) to lower orbit (n_1).

The spectral lines for atomic hydrogen

Series	Region	n_1	n_2
Lyman	Ultraviolet	1	2, 3, 4, 5 ...
Balmer	Visible	2	3, 4, 5, 6 ...
Paschen	Infra-red	3	4, 5, 6, 7 ...
Brackett	Infra-red	4	5, 6, 7, 8 ...
Pfund	Infra-red	5	6, 7, 8, 9 ...

Quantum Mechanical Model of Atom

- Schrodinger derived an equation for calculating energy of an electron possessing wave-particle dualism. The Schrodinger wave equation in three-dimension :
$$\frac{\partial^2\psi}{\partial x^2} + \frac{\partial^2\psi}{\partial y^2} + \frac{\partial^2\psi}{\partial z^2} + \frac{8\pi^2 m}{h^2}(E - V)\psi = 0$$
- The wave function (ψ) is the amplitude function of the particle wave and ψ^2 is the measure of the probability of finding an electron in a volume element $\partial x\partial y\partial z$.

Quantum Numbers

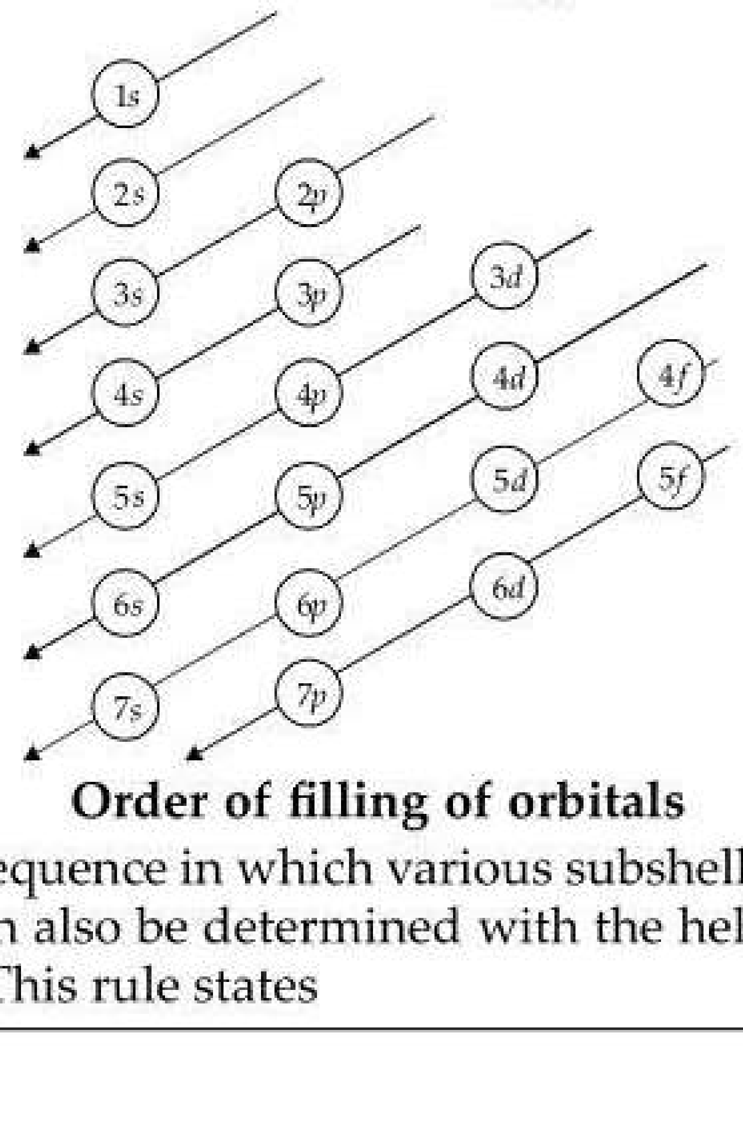
- Each electron in an atom can be identified by a set of four quantum numbers.
  - Principal quantum number (n) corresponds to the main energy level or shell in which the electron is present.

The value of n : 1 2 3 4
Corresponding shell : K L M N

- Azimuthal quantum number (l) gives the orbital angular momentum and corresponds to the subshell in a given principal energy shell.
  $l = 0, 1, 2, 3, ..., (n-1)$
  The various subshells are designated by the letters s, p, d, f.
  The value of l : 0 1 2 3
  Designation : s p d f
- Magnetic quantum number (m_l) describes the behaviour of an electron in magnetic field and it corresponds to the number of orbitals in a subshell.
  $m_l = -l$ to 0 to $+l = (2l + 1)$ values.
  Subshell : s p d f
  No. of orbitals : 1 3 5 7
- Spin quantum number (m_s) corresponds to the direction of electron spin in each orbital.
  m_s can have only two values, *i.e.*, ±1/2, represented as $\uparrow$ and $\downarrow$.

- The point where there is zero probability of finding an electron is called nodal point or node.
- For an atom with principal quantum number n,
  - There are $(n - l - 1)$ radial nodes.
  - l angular nodes.
  - Total $(n - 1)$ nodes.

  Electrons are distributed in different orbitals in an atom and this arrangement is referred to as electronic configuration.
- The filling up of electrons in different orbitals is based on the following rules:
  - **Aufbau's principle :** Electrons are progressively added to the different orbitals in their increasing order of energy.
    $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < ....$
  - **Pauli's exclusion principle :** No two electrons in an atom can have the same set of all the four quantum numbers.
  - **Hund's rule of maximum multiplicity :** Electron pairing in an orbital of same energy takes place only when each orbital is singly filled.



Order of filling of orbitals

  - The sequence in which various subshells are filled up can also be determined with the help of $(n + l)$ rule. This rule states
    "The subshell with lowest $(n + l)$ value is filled up first. When two or more subshells have same $(n + l)$ value, the subshell with lowest value of 'n' is filled up first."

	n	l	$(n+l)$	
$1s$	1	0	1	
$2s$	2	0	2	
$2p$	2	1	3	} (Lowest value of n)
$3s$	3	0	3	
$3p$	3	1	4	} (Lowest value of n)
$4s$	4	0	4	

Isotopes, Isobars, Isotones, Isoelectronics and Isosters

- Atoms of same element having same atomic number but different mass number are related as isotopes, which differ from each other in their physical properties and number of neutrons. *e.g.* Protium ($_1H^1$), Deuterium ($_1H^2$), Tritium ($_1H^3$).
- Atoms of different elements having same mass number but different atomic number are related as isobars, which differ in their physical as well as chemical properties.
  e.g. $_{18}Ar^{40}$ and $_{20}Ca^{40}$, $_7N^{14}$ and $_6C^{14}$, etc.
- Atoms of different elements having same number of neutrons are called as isotones.
  e.g. $_6C^{13}$ and $_7N^{14}$, having 7 neutrons each.
  $_{32}Ge^{76}$ and $_{33}As^{77}$, having 44 neutrons each.
- Species having same number of electrons are called as isoelectronics. *e.g.* Na^+, Mg^{2+}, Al^{3+} each contains 10-electrons. CH_4, NH_3, H_2O are also isoelectronic containing 10 electrons each.
- Species having same number of atoms as well as same number of electrons in the valence shell are termed as isosters, *e.g.* CO and N_2; CO_2 and N_2O, etc.

Greater Stability of Exactly Half - Filled and Completely Filled Configuration

- Elements with atomic numbers 24 (Cr), 42 (Mo) and 74 (W) have $ns^1(n-1)d^5$ configuration and not $ns^2(n-1)d^4$, due to extra stability of half-filled orbitals.
- Elements with atomic numbers 29 (Cu), 47 (Ag) and 79 (Au) have $ns^1(n-1)d^{10}$ configuration and not $ns^2(n-1)d^9$ due to extra stability of fully-filled orbitals.
- **Causes of extra stability :**
  - **Symmetry :** The half-filled and completely filled configurations are more symmetrical and symmetry leads to greater stability.
  - **Exchange energy :** The electrons present in the different orbitals of the same subshell can exchange their positions. Each such exchange results in release of energy (called exchange energy), which leads to a greater stability. As the number of exchanges that can take place is maximum in the exactly half-filled and completely filled arrangements (*i.e.* more in d^5 than in d^4 and more in d^{10} than in d^9), therefore exchange energy is maximum in these two cases and hence the stability is maximum.

EXAM DRILL

1. The nucleus of an atom is located at $x = y = z = 0$. If the probability of finding an s-orbital electron in a tiny volume around $x = a$, $y = z = 0$ is 1×10^{-5}, what is the probability of finding the electron in the same sized volume around $x = z = 0$, $y = a$?
   (a) 1×10^{-5} (b) $1 \times 10^{-5} \times a$
   (c) $1 \times 10^{-5} \times a^2$ (d) $1 \times 10^{-5} \times a^{-1}$

2. The limiting line in Balmer series will have a frequency of
   (a) $6.22 \times 10^{15}\ s^{-1}$ (b) $7.22 \times 10^{14}\ s^{-1}$
   (c) $8.22 \times 10^{14}\ s^{-1}$ (d) $9.22 \times 10^{14}\ s^{-1}$

3. If the electron falls from $n = 5$ to $n = 4$ in the H-atom, then emitted energy is
   (a) 0.306 eV (b) 12.09 eV
   (c) 1.89 eV (d) 0.65 eV

4. Photoelectric effect is the phenomenon in which
   (a) photons come out of a metal when it is hit by a beam of electrons
   (b) photons come out of the nucleus of an atom under the action of an electric field
   (c) electrons come out of a metal with a constant velocity which depends on the frequency and intensity of incident light wave
   (d) electrons come out of a metal with different velocities not greater than a certain value which depends only on the frequency of the incident light wave and not on its intensity.

5. Which of the folllowing statements is not correct?
   (a) The shape of an atomic orbital depends on the azimuthal quantum number.
   (b) The orientation of an atomic orbital depends on the magnetic quantum number.
   (c) The energy of an electron in an atomic orbital of multi-electron atom depends on principal quantum number.
   (d) The number of degenerate atomic orbitals of one type depends on the values of azimuthal and magnetic quantum numbers.

6. Rydberg is
   (a) also called Rydberg constant and is a universal constant
   (b) unit of wavelength and one Rydberg equals $1.097 \times 10^7\ m^{-1}$
   (c) the unit of wave number and one Rydberg equals $1.097 \times 10^7\ m^{-1}$
   (d) unit of energy and one Rydberg equals 13.6 eV.

7. The total number of orbitals in a shell with principal quantum number n is
   (a) $2n$ (b) $2n^2$ (c) n^2 (d) $n + 1$

8. The spectrum of white light ranging from red to violet is called a continuous spectrum because
   (a) different colours are seen as different bands in the spectrum
   (b) the colours continuously absorb energy to form a spectrum
   (c) the violet colour merges into blue, blue into green, green into yellow and so on
   (d) it is a continuous band of coloured and white light separating them.

9. Effective nuclear charge ($Z_{eff}\, e$) for a nucleus of an atom is defined as
   (a) shielding of the outermost shell electrons from the nucleus by the innermost shell electrons
   (b) the net positive charge experienced by electron from the nucleus
   (c) the attractive force experienced by the nucleus from electron
   (d) screening of positive charge on nucleus by innermost shell electrons.

10. The number of radial nodes and angular nodes for d-orbital can be represented as
    (a) $(n - 2)$ radial nodes + 1 angular node $= (n - 1)$ total nodes
    (b) $(n - 1)$ radial nodes + 1 angular node $= (n - 1)$ total nodes
    (c) $(n - 3)$ radial nodes + 2 angular nodes $= (n - l - 1)$ total nodes
    (d) $(n - 3)$ radial nodes + 2 angular nodes $= (n - 1)$ total nodes

11. If the radius of first Bohr orbit is x pm, then the radius of the third orbit would be
    (a) $(3 \times x)$ pm (b) $(6 \times x)$ pm
    (c) $\left(\frac{1}{2} \times x\right)$ pm (d) $(9 \times x)$ pm

12. Millikan's oil drop method is used to find
    (a) e/m ratio of electron (b) mass of electron
    (c) velocity of electron (d) charge of electron.

13. The correct set of quantum numbers for the outermost electron of Rubidium (37) is
    (a) $5, 0, 0, +\frac{1}{2}$ (b) $4, 3, 2, -\frac{1}{2}$
    (c) $5, 1, 0, -\frac{1}{2}$ (d) $5, 1, 1, +\frac{1}{2}$

14. In any subshell, the maximum number of electrons having same value of spin quantum number is
    (a) $\sqrt{l(l+1)}$ (b) $l + 2$ (c) $2l + 1$ (d) $4l + 2$

15. An ion with mass number 56 contains 3 units of positive charge and 30.4% more neutrons than electrons. The ion is
    (a) $^{56}_{28}Ni^{3+}$ (b) $^{56}_{26}Fe^{3+}$
    (c) $^{56}_{27}Co^{3+}$ (d) $^{56}_{24}Cr^{3+}$

16. Energy of H-atom in the ground state is –13.6 eV, hence energy in the second excited state is
    (a) – 6.8 eV (b) – 3.4 eV
    (c) –1.51 eV (d) – 4.53 eV

17. The total spin resulting from a d^7 configuration is
(a) $\pm 1/2$ (b) ± 2 (c) ± 1 (d) $\pm 3/2$

18. Nuclides
(a) have same number of protons
(b) have specific atomic numbers
(c) have specific atomic and mass numbers
(d) are isotopes.

19. ^{18}O isotope of oxygen will have
(a) 18 protons
(b) 9 protons and 9 neutrons
(c) 8 neutrons and 10 protons
(d) 10 neutrons and 8 protons.

20. The concept that atoms combine in small whole number ratio was proposed by
(a) Dalton (b) Avogadro
(c) Gay-Lussac (d) Berzelius.

21. For a shell of principal quantum number $n = 4$, there are
(a) 16 orbitals (b) 8 subshells
(c) 18 electrons (maximum)
(d) 4 electrons with $l = 3$.

22. Which among the following statements is/not correct?
(a) ψ^2 represents the molecular orbitals.
(b) The number of peaks in radial distribution is $(n - l)$.
(c) Radial probability density $\rho_{nl}(r) = 4\pi r^2 R^2_{nl}(r)$.
(d) A node is a point in space where the wave function (ψ) has zero amplitude.

23. 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?
(a) 5.79×10^6 cm s^{-1} (b) 5.79×10^6 m s^{-1}
(c) 5.79×10^{10} cm s^{-1} (d) 5.79×10^{10} m s^{-1}

24. 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 the position.
(a) 1.46×10^{-33} m (b) 1.46×10^{-33} cm
(c) 1.59×10^{-33} m (d) 1.39×10^{33} km

25. Calculate the uncertainty in the momentum of an electron if it is confined to a linear region of length 1×10^{-10} metre.
(a) 5.37×10^{-27} kg m s^{-1} (b) 5.27×10^{-25} g m s^{-1}
(c) 5.37×10^{-25} g m s^{-1} (d) 5.27×10^{-25} kg m s^{-1}

26. What is the maximum numbers of electrons that can be associated with the following set of quantum numbers?
$$n = 3, l = 1 \text{ and } m = -1$$
(a) 4 (b) 2 (c) 10 (d) 6

27. Based on equation $E = -2.178 \times 10^{-18}\ \text{J}\left(\dfrac{Z^2}{n^2}\right)$, certain conclusions are written. Which of them is not correct?
(a) Equation can be used to calculate the change in energy when the electron changes orbit.
(b) For $n = 1$, the electron has a more negative energy than it does for $n = 6$ which means that the electron is more loosely bound in the smallest allowed orbit.
(c) The negative sign in equation simply means that the energy of electron bound to the nucleus is lower than it would be if the electrons were at the infinite distance from the nucleus.
(d) Larger the value of n, the larger is the orbit radius.

28. The ionization enthalpy of He^+ ion is 19.60×10^{-18} J atom^{-1}. The ionization enthalpy of Li^{2+} ion will be
(a) 84.2×10^{-18} J atom^{-1} (b) 44.10×10^{-18} J atom^{-1}
(c) 63.20×10^{-18} J atom^{-1} (d) 21.20×10^{-18} J atom^{-1}

29. Which transition in the hydrogen atomic spectrum will have the same wavelength as the transition, $n = 4$ to $n = 2$ of He^+ spectrum?
(a) $n = 4$ to $n = 3$ (b) $n = 3$ to $n = 2$
(c) $n = 4$ to $n = 2$ (d) $n = 2$ to $n = 1$

30. The energy of an electron in first Bohr orbit of H-atoms is 13.6 eV. The possible energy value of electron in the excited state of Li^{2+} is
(a) –122.4 eV (b) 30.6 eV
(c) –30.6 eV (d) 13.6 eV

31. Which one of the following has the lowest ionisation energy?
(a) $1s^2 2s^2 2p^6$ (b) $1s^2 2s^2 2p^6 3s^1$
(c) $1s^2 2s^2 2p^5$ (d) $1s^2 2s^2 2p^3$

32. The representation of the ground state electronic configuration of He by box-diagram as $\boxed{\uparrow\uparrow}$ is wrong because it violates
(a) Heisenberg's uncertainty principle
(b) Bohr's quantization theory of angular momenta
(c) Pauli's exclusion principle
(d) Hund's rule.

33. The electronic transitions from $n = 2$ to $n = 1$ will produce shortest wavelength in (where n = principal quantum state)
(a) Li^{2+} (b) He^+ (c) H (d) H^+

34. 13.6 eV is needed for ionization of a hydrogen atom. An electron in a hydrogen atom in its ground state absorbs 1.50 times as much energy as the minimum required for it to escape from the atom. What is the wavelength of the emitted electron?
($m_e = 9.109 \times 10^{-31}$ kg, $e = 1.602 \times 10^{-19}$ Coloumb, $h = 6.63 \times 10^{-34}$ J s)
(a) 1.5×10^{-6} m (b) 2.3×10^{-11} m
(c) 4.7×10^{-10} m (d) 1.4×10^{-9} m

35. Number of waves made by an electron in one complete revolution in 3rd Bohr orbit is
(a) 2 (b) 3 (c) 4 (d) 1

36. The degeneracy of the level of hydrogen atom that has energy $-\dfrac{R_H}{16}$ is
(a) 16 (b) 4 (c) 2 (d) 1

37. What is the minimum error in position of an electron moving with a speed of 500 m s^{-1} measured to an accuracy of 0.006% (mass of the electron $= 9.1 \times 10^{-31}$ kg)?
(a) 2.73×10^{-3} m (b) 1.932×10^{-3} m
(c) 3.112×10^{-5} m (d) 0.119×10^{-5} m

38. 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 radiation of 600 nm, calculate the number of photons received by the detector.

(a) 4 (b) 9 (c) 6 (d) 7

39. Time taken for an electron to complete one revolution in the Bohr orbit of hydrogen atom is

(a) $\frac{4\pi^2 mr^2}{nh}$ (b) $\frac{nh}{4\pi^2 mr}$

(c) $\frac{2\pi mr}{n^2h^2}$ (d) $\frac{nh}{4\pi^2 mr^2}$

40. What is the wavelength of the light emitted when the electron in a hydrogen atom undergoes transition from an energy level with $n = 4$ to an energy level with $n = 2$?

(a) 205 nm (b) 123 nm

(c) 486 nm (d) 329 nm

41. The energy of the first electron in helium will be

(a) –13.6 eV (b) –54.4 eV

(c) –5.44 eV (d) zero

42. Emission transitions in the Paschen series end at orbit $n = 3$ and start from orbit n and can be represented as

$$\nu = 3.29 \times 10^{15}\,\text{Hz}\left(\frac{1}{3^2} - \frac{1}{n^2}\right)$$

Calculate the value of n if the transition is observed as 1285 nm.

(a) 3 (b) 4 (c) 5 (d) 6

43. If the uncertainty in the position of an electron is zero, the uncertainty in its momentum would be

(a) zero (b) greater than $h/4\pi$

(c) less than $h/4\pi$ (d) infinite.

44. If the position of an electron is known to be within 10^{-12} m, then uncertainty in its momentum is

(a) 5.27×10^{-23} kg m s^{-1} (b) 2.130×10^{-23} kg m s^{-1}

(c) 1.69×10^{-24} kg m s^{-1} (d) 1.69×10^{-23} kg m s^{-1}.

45. The electrons, identified by quantum numbers n and l (i) $n = 4, l = 1$ (ii) $n = 4, l = 0$ (iii) $n = 3, l = 2$ (iv) $n = 3, l = 1$ can be placed in order of increasing energy from the lowest to highest as

(a) (iv) < (ii) < (iii) < (i) (b) (ii) < (iv) < (i) < (iii)

(c) (i) < (iii) < (ii) < (iv) (d) (iii) < (i) < (iv) < (ii)

DAY 02 OMR SHEET

Time : 45 min

INSTRUCTIONS

- Use HB pencil only and darken each circle completely.
- If you wish to change your answer, erase the already darkened circle completely and then darken the appropriate circle.
- Mark only one choice for each question as indicated.

Correct marking ● ⓑ ⓒ ⓓ

Wrong marking ⊗ ⓑ ◐ ⊙

1. ⓐ ⓑ ⓒ ⓓ	10. ⓐ ⓑ ⓒ ⓓ	19. ⓐ ⓑ ⓒ ⓓ	28. ⓐ ⓑ ⓒ ⓓ	37. ⓐ ⓑ ⓒ ⓓ
2. ⓐ ⓑ ⓒ ⓓ	11. ⓐ ⓑ ⓒ ⓓ	20. ⓐ ⓑ ⓒ ⓓ	29. ⓐ ⓑ ⓒ ⓓ	38. ⓐ ⓑ ⓒ ⓓ
3. ⓐ ⓑ ⓒ ⓓ	12. ⓐ ⓑ ⓒ ⓓ	21. ⓐ ⓑ ⓒ ⓓ	30. ⓐ ⓑ ⓒ ⓓ	39. ⓐ ⓑ ⓒ ⓓ
4. ⓐ ⓑ ⓒ ⓓ	13. ⓐ ⓑ ⓒ ⓓ	22. ⓐ ⓑ ⓒ ⓓ	31. ⓐ ⓑ ⓒ ⓓ	40. ⓐ ⓑ ⓒ ⓓ
5. ⓐ ⓑ ⓒ ⓓ	14. ⓐ ⓑ ⓒ ⓓ	23. ⓐ ⓑ ⓒ ⓓ	32. ⓐ ⓑ ⓒ ⓓ	41. ⓐ ⓑ ⓒ ⓓ
6. ⓐ ⓑ ⓒ ⓓ	15. ⓐ ⓑ ⓒ ⓓ	24. ⓐ ⓑ ⓒ ⓓ	33. ⓐ ⓑ ⓒ ⓓ	42. ⓐ ⓑ ⓒ ⓓ
7. ⓐ ⓑ ⓒ ⓓ	16. ⓐ ⓑ ⓒ ⓓ	25. ⓐ ⓑ ⓒ ⓓ	34. ⓐ ⓑ ⓒ ⓓ	43. ⓐ ⓑ ⓒ ⓓ
8. ⓐ ⓑ ⓒ ⓓ	17. ⓐ ⓑ ⓒ ⓓ	26. ⓐ ⓑ ⓒ ⓓ	35. ⓐ ⓑ ⓒ ⓓ	44. ⓐ ⓑ ⓒ ⓓ
9. ⓐ ⓑ ⓒ ⓓ	18. ⓐ ⓑ ⓒ ⓓ	27. ⓐ ⓑ ⓒ ⓓ	36. ⓐ ⓑ ⓒ ⓓ	45. ⓐ ⓑ ⓒ ⓓ

(1) Number of questions attempted : ____________ (3) Marks scored : ___________

(2) Number of questions correct : ____________

For every correct answer award yourself 4 marks. For every incorrect answer deduct 1 mark.

HINTS & SOLUTIONS

1. (a) : Since the distance from the nucleus is same, and the s-orbital is spherically symmetrical, hence probability is identical.

2. (c) : The limiting line of Balmer series refers to the transition of electron from ∞ to 2^{nd} orbit $\upsilon = c.\bar{\upsilon}$

$$= 3\times10^{10}\times109677\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right) = 3.29\times10^{15}\left(\frac{1}{n_1^2}-\frac{1}{n_2^2}\right)\text{ sec}^{-1}$$

$= 8.22 \times 10^{14}\text{ sec}^{-1}$ ($n_1 = 2$, $n_2 = \infty$)

3. (a) : $E_n = -\frac{13.6}{n^2}$ eV

$$E_5 - E_4 = 13.6\left(\frac{1}{4^2}-\frac{1}{5^2}\right)\text{ eV} = 13.6\left(\frac{1}{16}-\frac{1}{25}\right) = 0.306\text{ eV}$$

4. (d): $\frac{1}{2}mu_{\text{max}}^2 = h\upsilon - W$.

5. (c) : The energy level of an orbital in multi-electron system depends upon l.

6. (a)

7. (c) : Number of orbitals in a shell = n^2.

8. (c) : In a continuous spectrum, the colours merge into each other in a continuous pattern.

9. (b): Effective nuclear charge is the net positive charge experienced by the electrons from the nucleus. Due to presence of electrons in the innermost shells the electrons in the outermost shell do not experience the full positive charge from the nucleus. This is known as shielding of the outermost shell electrons from the nucleus by the innermost shell electrons.

10. (d) : Total number of nodes = $n - 1$
For d-orbital, radial nodes = $n - 3$ and there are 2 angular nodes.
The number of angular nodes is given by l. *i.e.,* for p, 1 angular node, for d, 2 angular nodes and so on.

11. (d) : $r = \frac{5.29n^2}{Z}$ or $r = x\times n^2$ or $(9\times x)$ pm

12. (d)

13. (a) : Outermost electron of Rb (At. no. 37) is 5s. Hence, its quantum numbers are 5, 0, 0, $\pm 1/2$.

14. (c) : Maximum number of electrons with same spin is equal to the maximum number of orbitals *i.e.*, $(2l + 1)$.

15. (b): Let no. of electrons in the ion $M^{3+} = x$

$\therefore$ No. of neutrons = $x + \frac{30.4x}{100} = 1.304x$

No. of electrons in the neutral atom = $x + 3$
$\therefore$ No. of protons = $x + 3$
Mass no. = No. of protons + No. of neutrons
$56 = x + 3 + 1.304x$ or $2.304x = 53$ or $x = 23$
No. of protons = $x + 3 = 23 + 3 = 26$
Hence the ion is ${}^{56}_{26}Fe^{3+}$.

16. (c) : $E_n = -\frac{13.6}{n^2}$ eV
Second excited state means $n = 3$

Thus, $E_3 = -\frac{13.6}{3^2} = -\frac{13.6}{9} = -1.51$ eV

17. (d): d^7 configuration has three unpaired electrons.
Thus, total spin = $\pm1/2 \times$ no. of unpaired electrons

$= \pm\frac{1}{2}\times3 = \pm\frac{3}{2}$

18. (c) : A nuclide has a definite number of protons and neutrons, *e.g.*, ${}_8O^{16}$.

19. (d): Atomic number of oxygen is 8.

20. (a)

21. (a) : Maximum number of orbitals = $n^2 = 16$
Maximum number of electrons = $2n^2 = 32$
Number of subshells when $n = 4$
i.e., n = 1, 2, 3, 4 (K, L, M, N)

22. (a) : ψ^2 represents atomic orbital.

23. (b): $\Delta x\,\Delta p = \frac{h}{4\pi}$ or $\Delta x\, m\Delta v = \frac{h}{4\pi} \Rightarrow \Delta v = \frac{h}{4\pi\,\Delta x\, m}$

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

$= 0.579 \times 10^7\text{ m s}^{-1}$ ($1\text{ J} = 1\text{ kg m}^2\text{ s}^{-2}$) $= 5.79 \times 10^6\text{ m s}^{-1}$.

24. (a) : The uncertainty in the speed is 2%, *i.e.*,

$45\times\frac{2}{100} = 0.9\text{ m s}^{-1}$. Using the equation, $\Delta x = \frac{h}{4\pi m\Delta v}$

$$= \frac{6.626\times10^{-34}\text{ J s}}{4\times3.14\times40\times10^{-3}\text{ kg }(0.9\text{ m s}^{-1})} = 1.46\times10^{-33}\text{ m}$$

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.

25. (d): According to uncertainty principle,

$\Delta x\cdot\Delta p = \frac{h}{4\pi}$ or $\Delta p = \frac{h}{4\pi\Delta x}$

$$\Delta p = \frac{(6.626\times10^{-34}\text{ kg m}^2\text{ s}^{-1})}{4\times3.143\times(10^{-10}\text{ m})} = 5.27\times10^{-25}\text{ kg m s}^{-1}.$$

26. (b): The orbital associated with $n = 3$, $l = 1$ is 3p. One orbital (with $m = -1$) of 3p-subshell can accomodate maximum 2 electrons.

27. (b): The electron is more tightly bound in the smallest allowed orbit.

28. (b) : Ionisation enthalpy $\propto Z^2$ (Z = Atomic number)

Ionisation enthalpy of Li^{2+} = *I.E.* of $He^+ \times \left(\frac{3^2}{2^2}\right)$

I.E. of $Li^{2+} = 19.6\times10^{-18}\times\frac{9}{4} = 44.1\times10^{-18}$ J atom^{-1}

29. (d): For He^+ ion,

$$\frac{1}{\lambda} = Z^2 R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \Rightarrow (2)^2 R\left[\frac{1}{2^2} - \frac{1}{4^2}\right] = \frac{3R}{4}$$

For hydrogen atom, $\frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

$$\frac{3R}{4} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right] \quad \text{or} \quad \frac{1}{n_1^2} - \frac{1}{n_2^2} = \frac{3}{4}$$

$n_1 = 1$ and $n_2 = 2$.

30. (c) : $E_n = E_1 \times \frac{Z^2}{n^2}$

E_1 = energy of hydrogen in first orbit
$n = 2$, Z = atomic number

$$E = -13.6 \times \frac{(3)^2}{(2)^2} = -30.6 \text{ eV}$$

31. (b): The element which has only one electron in the 3*s*-subshell has the lowest ionisation energy.

32. (c) : Pauli exclusion principle states that an orbital can have maximum two electrons and these must have opposite spins.

33. (a) : $\frac{1}{\lambda} = Z^2 R_H\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$

$n_1 = 1$, $n_2 = 2$

$$\frac{1}{\lambda} = Z^2 R_H\left[\frac{1}{1} - \frac{1}{4}\right] \Rightarrow \frac{1}{\lambda} = \frac{3}{4} R_H Z^2 \therefore \lambda \propto \frac{1}{Z^2}$$

Hence elements having high atomic number will produce shortest wavelength.

34. (c) : 1.5 times of 13.6 eV *i.e.*, 20.4 eV is absorbed by the hydrogen atom out of which 6.8 eV (20.4 – 13.6) is converted to kinetic energy.

KE = 6.8 eV = 6.8 (1.602×10^{-19} coulomb)(1 volt) = 1.09×10^{-18} J

Now, $KE = \frac{1}{2}mv^2$

$$\text{or, } v = \sqrt{\frac{2KE}{m}} = \sqrt{\frac{2(1.09 \times 10^{-18} \text{ J})}{(9.109 \times 10^{-31} \text{ kg})}} = 1.55 \times 10^6 \text{ m/s}$$

$$\therefore \quad \lambda = \frac{h}{mv} = \frac{(6.63 \times 10^{-34} \text{ J s})}{(9.109 \times 10^{-31} \text{ kg})(1.55 \times 10^6 \text{ m/s})} = 4.7 \times 10^{-10} \text{ m}$$

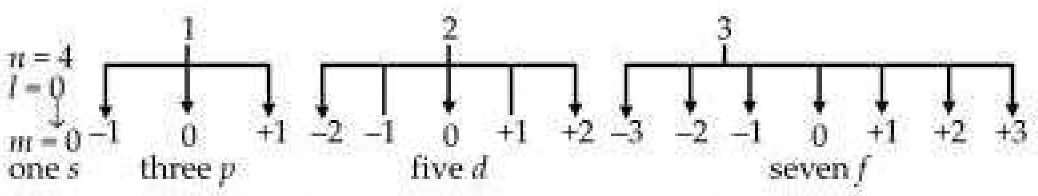
35. (b): Velocity of the electron in 3rd orbit

By de-Broglie equation, $= \frac{3h}{2\pi mr}$

$$\lambda = \frac{h}{mv} = \frac{h}{m} \times \frac{2\pi mr}{3h} = \frac{2\pi r}{3}$$

No. of waves $= \frac{2\pi r}{\lambda} = \frac{2\pi r}{2\pi r} \times 3 = 3$

36. (a) : $E_n = -\frac{R_H}{n^2} \quad \therefore \quad -\frac{R_H}{n^2} = -\frac{R_H}{16}$ *i.e.*, for 4th sub-shell



i.e., $1 + 3 + 5 + 7 = 16 \therefore$ Degeneracy is 16.

37. (b): $\Delta p = m\Delta v = 9.1 \times 10^{-31} \times 500 \times \frac{0.006}{100}$

$= 2.73 \times 10^{-32}$ kg m s^{-1}

$$\Delta x = \frac{h}{4\pi m \Delta v} = \frac{6.626 \times 10^{-34}}{4 \times \pi \times 2.73 \times 10^{-32}} = 1.932 \times 10^{-3} \text{ m}$$

38. (b): Total energy = 3.15×10^{-18} J
Wavelength, λ = 600 nm = 600×10^{-9} m
No. of photons received by detector = ?

Energy per photon $= \frac{hc}{\lambda}$

$$= \frac{6.626 \times 10^{-34} \text{J s} \times (3 \times 10^8 \text{ m s}^{-1})}{600 \times 10^{-9} \text{ m}} = 3.313 \times 10^{-19} \text{ J}$$

So, no. of photons received by the detector

$$= \frac{3.15 \times 10^{-18} \text{ J}}{3.313 \times 10^{-19} \text{ J}} = 9.5 = 9$$

39. (a) : By Bohr postulate, $mvr = n\frac{h}{2\pi}$ or $v = \frac{nh}{2\pi mr}$

No. of revolutions per sec

$$= \frac{\text{Velocity}}{\text{Circumference of the orbit}} = \frac{v}{2\pi r}$$

Substituting value of v, we get

$$= \frac{nh}{2\pi mr} \times \frac{1}{2\pi r} = \frac{nh}{4\pi^2 mr^2}$$

$\frac{nh}{4\pi^2 mr^2}$ revolutions completed in 1s

$$\therefore \quad \text{1 revolution will take } \frac{1}{\frac{nh}{4\pi^2 mr^2}} = \frac{4\pi^2 mr^2}{nh}$$

40. (c) : Given that $n_1 = 2$ and $n_2 = 4$
Using the formula,

$$\bar{\nu} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) = 109677\left(\frac{1}{2^2} - \frac{1}{4^2}\right) = 20564.4 \text{ cm}^{-1}$$

$$\lambda = \frac{1}{\bar{\nu}} = \frac{1}{20564.4} = 4.86 \times 10^{-5} \text{ cm} = 4.86 \times 10^{-7} \text{ m} = 486 \text{ nm}$$

41. (b)

42. (c) : $\nu = 3.29 \times 10^{15}\left(\frac{1}{3^2} - \frac{1}{n^2}\right)$

$$\nu = \frac{c}{\lambda} = \frac{3 \times 10^8}{1285 \times 10^{-9}} = 3.29 \times 10^{15}\left(\frac{1}{3^2} - \frac{1}{n^2}\right)$$

$$\frac{1}{n^2} = \frac{1}{9} - \frac{3 \times 10^8}{1285 \times 10^{-9}} \times \frac{1}{3.29 \times 10^{15}} = \frac{1}{25} \Rightarrow n^2 = 25 \Rightarrow n = 5$$

The radiation corresponding to 1285 nm lies in the infrared region.

43. (d): $\Delta x \cdot \Delta p = \frac{h}{4\pi}$ or $\Delta p = \frac{h}{4\pi} \cdot \frac{1}{\Delta x} = \frac{h}{0} = \infty$.

44. (a) : $\Delta p = \frac{6.626 \times 10^{-34}}{4\pi \times 10^{-12} \text{ m}} = 5.27 \times 10^{-23}$ kg ms^{-1}

45. (a) : (i) 4*p* (ii) 4*s* (iii) 3*d* (iv) 3*p*
These are arranged in increasing order of energy as

$3p < 4s < 3d < 4p$
(iv) (ii) (iii) (i)