ATOMIC STRUCTURE

TOPIC-WISE STUDY MATERIAL

BY

EXAMS ROAD

"Bringing Excellence to Students"

The concept of atomic structure refers to the organization and composition of atoms, which are the building blocks of matter. Here's a complete overview of the atomic structure:

Atoms: Atoms are the smallest units of an element that retain the chemical properties of that element. They consist of three main subatomic particles:

a. Protons: Protons carry a positive charge and are located in the nucleus (central core) of an atom. Each proton has a relative mass of 1 and a charge of +1.

b. Neutrons: Neutrons are electrically neutral particles found in the nucleus alongside protons. They have a relative mass of 1 but carry no charge (neutral).

c. Electrons: Electrons are negatively charged particles that orbit the nucleus in specific energy levels called electron shells or orbitals. They have a negligible mass compared to protons and neutrons.

Nucleus: The nucleus is the central part of an atom that contains protons and neutrons. It carries almost all of the atom's mass but occupies a tiny fraction of its volume.

Electron Shells: Electrons occupy specific energy levels around the nucleus, known as electron shells or orbitals. The shells are labeled with the letters K, L, M, N, and so on, starting from the one closest to the nucleus (K shell). Each shell can hold a maximum number of electrons: 2 for the first shell (K), 8 for the second shell (L), 18 for the third shell (M), and so on. The outermost shell is called the valence shell.

Atomic Number: The atomic number (Z) of an atom represents the number of protons in its nucleus. It determines the identity of the element. In a neutral atom, the atomic number also corresponds to the number of electrons.

Mass Number: The mass number (A) of an atom is the total number of protons and neutrons in its nucleus. It represents the mass of the atom.

Isotopes: Isotopes are atoms of the same element that have the same atomic number (number of protons) but different mass numbers (different number of neutrons). Isotopes exhibit similar chemical properties but may have different physical properties, such as varying atomic masses and radioactive behavior.

Atomic Mass: The atomic mass of an atom is the average mass of all its naturally occurring isotopes, taking into account their relative abundance. It is typically expressed in atomic mass units (amu).

Electron Configuration: The electron configuration of an atom describes the distribution of electrons among its various electron shells and subshells. It follows a specific notation, such as the Aufbau principle, Pauli exclusion principle, and Hund's rule.

Valence Electrons: Valence electrons are the electrons in the outermost shell (valence shell) of an atom. They play a crucial role in chemical bonding and determining the atom's reactivity. The number of valence electrons largely determines an atom's chemical properties.

Atomic Structure Models: Over time, different models have been proposed to explain atomic structure, including the Thomson model, Rutherford model, and Bohr model. The modern understanding is based on quantum mechanics, which describes electrons as existing in probability clouds around the nucleus, represented by atomic orbitals.

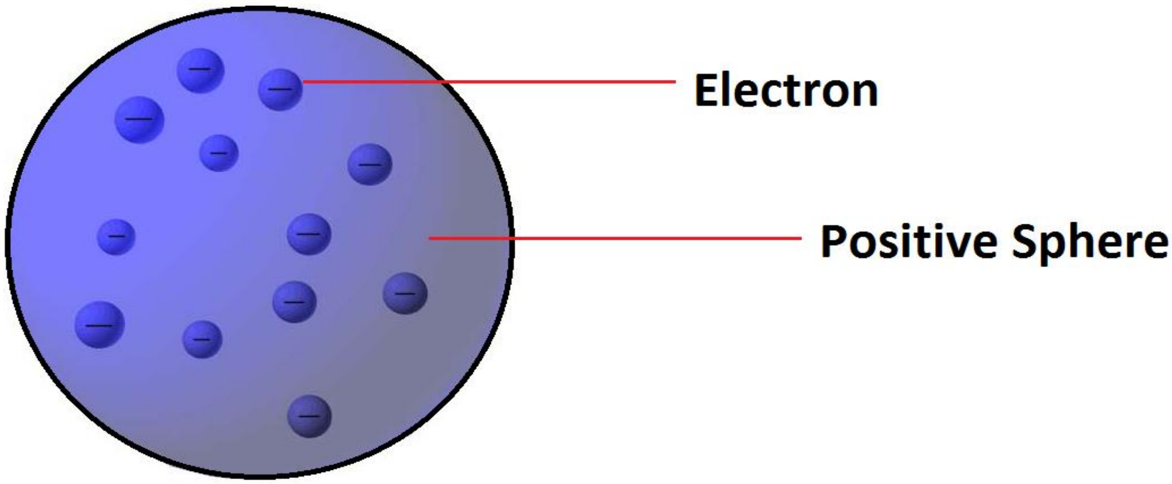
Understanding atomic structure is essential for comprehending chemical reactions, the periodic table, bonding, and various other aspects of chemistry and physics.

Subatomic Particles:

Name	Electron	Proton	Neutron
Symbol	e-	p	n
Approximate relative mass	1/1836	1	1
Mass in kg	9.109×10^{-31}	1.673×10^{-27}	1.675×10^{-27}
Mass in amu	5.485×10^{-4}	1.007	1.008
Charge (coulomb)	1.602×10^{-19}	1.602×10^{-19}	0
Actual Charge (e.s.u)	4.8×10^{-10}	4.8×10^{-10}	0

Atomic Models:

Thomson’s Atomic Model (Plum – pudding model):-



Postulate: -Atom is a sphere of positive charge in which number of electrons are embedded.

Limitations: - Could not satisfactorily explain the results of scattering experiment carried out by Rutherford.

Rutherford’s Model:

Postulates:-

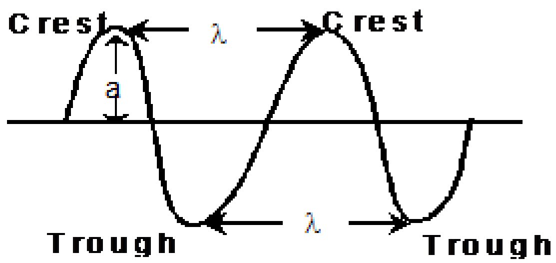
- Almost all the positive charge and mass of atom is present in its nucleus.
- Electrons revolve around the nucleus in circular orbits.
- There is strong electrostatic attraction between nucleus and electrons

Limitations: - Could not explain stability and electronic structure of atom.

Atomic Terms

Terms	Definition / Explanation
Atomic Number (Z)	Number of protons or electrons of neutral atom.
Mass Number (A)	Total number of protons and neutrons in an atom
Nucleons	Protons and neutrons are present in a nucleus. So, these fundamental particles are collectively known as nucleons
Isotopes	Atoms of the element with same atomic number but different mass number e.g. $_1H^1$, $_1H^2$, $_1H^3$.
Isobars	Atoms having the same mass number but different atomic numbers, e.g. $_{15}P^{32}$ and $_{16}S^{32}$
Isotones	Atoms having the same number of neutrons but different number of protons or mass number, e.g. $_6C^{14}$, $_8O^{16}$, $_7N^{15}$
Isoelectronic	Atoms, molecules or ions having same number of electrons e.g. N_2,CO, CN^-
Nuclear isomers	atoms with the same atomic number and same mass number but with different radioactive properties. Example of nuclear isomers is Uranium –X (half life 1.4 min) and Uranium –Z (half life 6.7 hours
Isosters	Molecules having same number of atoms and also same number of electrons are called isosters. E.g., N_2 and CO

Wave



Terms	Explanation
Wave length (λ)	Distance between two neighbouring troughs or crests.
Frequency (v)	Number of times a wave passes through a given point in a medium in one second. $v = c/\lambda$
Velocity (c)	The distance travelled by the wave in one second. $c = v\lambda$

Wave number	Number of wavelengths per cm.
Amplitude (a)	Height of the crest or depth of the trough. Determines the intensity of the beam of light.

Electromagnetic Waves

Radiations	Wave length (Å)
Radio waves	3×10^{14} to 3×10^{7}
Micro waves	3×10^{9} to 3×10^{6}
Infrared (IR)	6×10^{6} to 7600
Visible	7600 to 3800
Ultra violet (UV)	3800 to 150
X–rays	150 to 0.1
Gamma rays	0.1 to 0.01

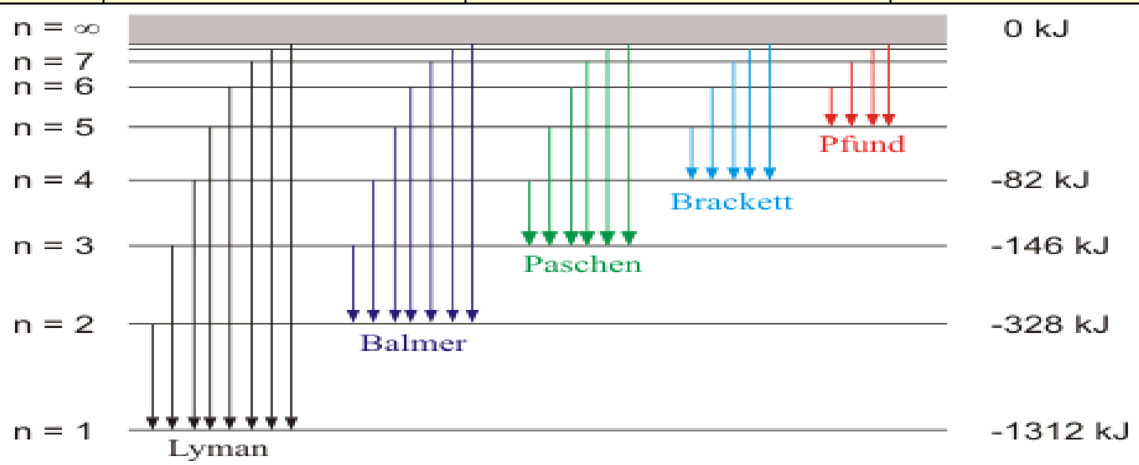
Atomic spectrum of hydrogen atom:

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

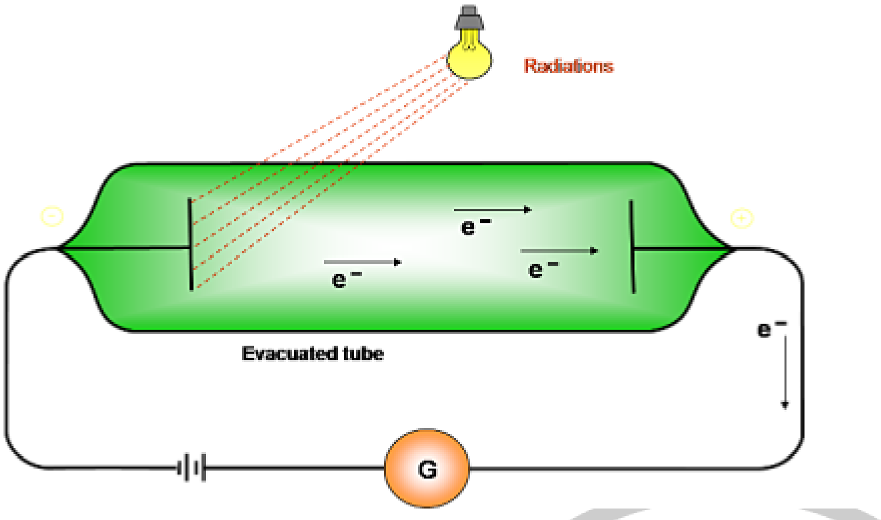
Where, R_H = Rydberg constant (108978 cm^{-1})

n_1 and n_2 have integral values as follows

n_1	n_2	Spectral Series	Spectral region
1	2,3,4…	Lyman	UV
2	3,4,5…	Balmer	Visible
3	4,5,6…	Pascher	IR
4	5,6,7…	Brackett	IR
5	6,7,8…	Pfund	IR



Photoelectric effect:



- Ejection of electrons takes place from the surface of metal when light of suitable frequency fall on it.
- Minimum frequency required for ejection of electron is called threshold frequency (v_o).
- Energy of the ejected electrons is directly proportional to the frequency of radiation.
- Number of electrons ejected per second depends on the intensity of radiation.
- $hv - hv_o = 1/2 m_e v^2$

Planck's quantum theory:

Substances radiate or absorb energy discontinuously in the form of energy packets

The smallest packet of energy is called quantum. In case of light the quantum is known as photon.

The energy of a quantum is directly proportional to the frequency of the radiation.

E = hv were v is the frequency of radiation and h is Planck's constant having the value 6.626×10^{-27} erg sec or 6.626×10^{-34} J sec.

A body can radiate or absorb energy in whole number multiples of quantum hn, 2hv,3hv..........nhv, where n is the positive integer.

Bohr's atomic model:

Electrons revolve around the nucleus in circular orbits of fixed energy.

Electron revolve only in those orbits whose angular momentum (mvr) is an integral multiple of h/2Π.

Electron absorbs energy in the form of EMR, when it jumps from lower energy level (ground state) to higher energy level (excited state) and vice-versa.

Energy absorbed or released in an electron jump, (dE) is given by $dE = E_2 - E_1 = hv$

Energy of stationary state oh hydrogen atom $(E_n) = -R_H (1/n^2)$

For an hydrogen like species i.e. He^+, Li^{2+} with atomic number Z

Radius of n^{th} orbit (r_n) = 52.9 × n^2/z pm

Energy of n^{th} orbit (E_n) = $-2.18 \times 10^{-18}(Z^2/n^2) = -13.6 \times (Z^2/n^2)$ eV = $313.6 \times (Z^2/n^2)$ kcal /mole

Velocity of electron (v) = (2.18×10^8) z/n cms^{-1}

Where n = 1,2,3,4…

Limitations of Bohr's theory:

- Failed to explain the spectra of atoms having more than one electron.
- Failed to account for the splitting of spectral line source of a spectrum is placed in a strong magnetic or electric field.
- Dual nature of particle and the uncertainty principle was ignored in the Bohr's atomic model.

de- Broglie equation:

$\lambda = h/mv = h/p$

Heisenberg's uncertainty principle:

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

It is impossible to determine simultaneously, the exact position and exact momentum of an electron.

Quantum Numbers:

Principal quantum number (n):

- It tells the main shell in which the electron resides and the approximate distance of the electron from the nucleus.
- Maximum number of electrons a shell can accommodate is $2n^2$.

Azimuthal or angular momentum quantum number (l):

- It represents the number of subshells present in the main shell.
- These subsidiary orbits within a shell will be denoted as s,p,d,f…
- This tells the shape of the subshells.
- For a given value of n, there are n possible values of l starting from 0 to (n-1)

Value of l	0	1	2	3	4
Notation of sub shell	s	p	d	f	g

The magnetic quantum number (m):

- It determines the number of preferred orientations of the electron present in a subshell.
- For a given value of l, there are (2l+1) possible values of m starting from –l to +l.

Value of l	0	1	2	3
Notation of sub shell	s	p	d	f

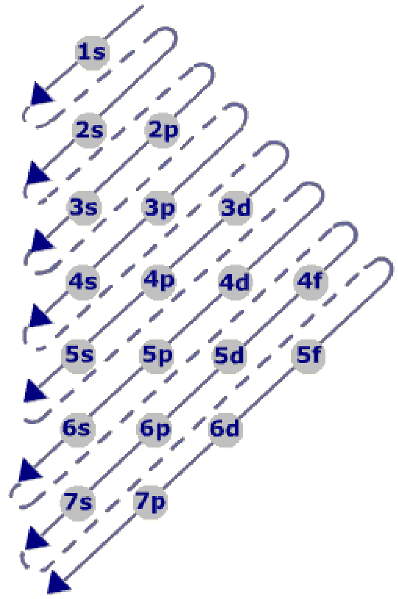
Values of m	0	-1,0,1	-2,-2,0,1,2	-3,-2,-1,0,1,2,3

The spin quantum number (s)

- It determines the direction of spin of electron in an orbit.
- There are only two possible values for spin quantum number i.e. -1/2 ,+1/2.

Rules for filling of electrons in various orbitals :

Aufbau Principle:



- Electrons are added one by one to the various orbitals in order of their increasing energy starting with the orbital of lowest energy.
- In neutral isolated atom, the lower the value of (n+ l) for an orbital, lower is its energy
- The increasing order of energy of various orbital is : 1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d ..

Pauli's Exclusion principle :-

An orbital can contain a maximum number of two electrons and these two electrons must be of opposite spin.

Hund's rule of maximum multiplicity :-

Electron pairing in p, d and f orbital cannot occur until each orbital of a given subshell contains one electron each or is singly occupied".

Exceptional Configurations

Stability of half filled and completely filled orbitals

Cu has 29 electrons. Its expected electronic configuration is

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^9$

But a shift of one electron from lower energy 4s orbital to higher energy 3d orbital will make the distribution of electron symmetrical and hence will impart more stability.

Thus the electronic configuration of Cu is

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1, 3d^{10}$

Fully filled and half filled orbitals are more stable

Question 1:

Choose isosteres from the following.

(a) NO_2^- and PO_4^{3-}
(b) NO_2^- and O_3
(c) CO_2, N_2O, NO_3^-
(d) ClO_4^- and OCN^-

Question 2:

Which of the following is a pair of isoelectronic species?

(a) Cl_2O_3, ICl_2^-
(b) ICl_2^-, ClO_2
(c) IF_2^+, I_3^-
(d) ClO_2^-, ClF_2^+

Question 3:

Which of the following statements about the electron is incorrect?
(a) It is negatively charged particle
(b) The mass of electron is equal to the mass of neutron
(c) It is a basic constituent of all atoms
(d) It is a constituent of cathode rays

Question 4:

The value of Planck's constant is 6.63×10^{-34} Js. The velocity of light is 3×10^8 ms^{-1}. Which value is closest to the wavelength in nanometres of a quantum of light with frequency of 8×10^{15} s^{-1}?

(a) 2×10^{-25}
(b) 3×10^7
(c) 4×10^1
(d) 5×10^{-18}

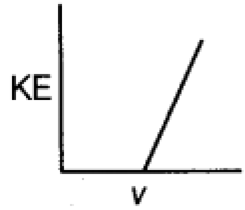
Question 5:

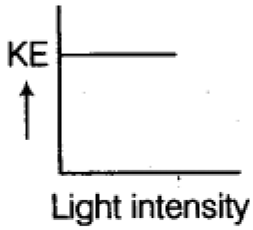
The kinetic energy of an electron emitted from a surface of a metal by light of wavelength 5.5×10^{-8} cm (Threshold frequency for the metal is $3.62 \times 10^{12} s^{-1}$), is **[NCERT Exemplar]**

(a) 6.6×10^{-27} erg $(5.5 \times 10^{17} s^{-1} - 3.62 \times 10^{12} s^{-1})$
(b) 6.6×10^{-27} erg $(-5.5 \times 10^{17} s^{-1} - 3.62 \times 10^{12} s^{-1})$
(c) 6.6×10^{-27} erg $(-5.5 \times 10^{17} s^{-1} + 3.62 \times 10^{12} s^{-1})$
(d) 6.6×10^{-27} erg $(5.5 \times 10^{17} s^{-1} + 3.62 \times 10^{12} s^{-1})$

Question 6:

Which of the following options represents the correct graph for photoelectrons?

(a) 

(b) 

(c) Both (a) and (b)

(d) None of these

Question 7:

If E_e, E_α and E_p represents the kinetic energies of an electron, alpha particle and a proton respectively, each moving with same de-Broglie wavelength then

(a) $E_e = E_\alpha = E_p$
(b) $E_e > E_\alpha > E_p$
(c) $E_\alpha > E_p > E_e$
(d) $E_e > E_p > E_\alpha$

Question 8:

The wavelength corresponding to maximum energy for hydrogen is 91.2 nm. The corresponding wavelength for He^+ ion is

(a) 2.28 nm (b) 22.8 nm (c) 182.4 nm (d) 364.8 nm

Question 9:

The wavelength of the radiation emitted when in a hydrogen atom electron falls from infinity to first stationary state would be ($R_H = 1.097 \times 10^7 \ m^{-1}$)

(a) 91 nm (b) 191 nm (c) 209 nm (d) 314 nm

Question 10:

The ratio of area covered by second orbital to the first orbital is

(a) 1 : 2
(b) 1 : 16
(c) 8 : 1
(d) 16: 1

Question 11:

Bohr's theory is applicable to

(a) He (b) Li^{2+}

(c) He^{2+} (d) None of these

Question 12:

According to Bohr's theory, the angular momentum of electron in 5th orbit is

(a) $25 \frac{h}{\pi}$ (b) $1.0 \frac{h}{\pi}$ (c) $10 \frac{h}{\pi}$ (d) $2.5 \frac{h}{\pi}$

Question 13:

The energy of second Bohr orbit of the hydrogen atom is -328 kJ mol^{-1}, hence the energy of fourth Bohr orbit would be

(a) -41 kJ mol^{-1} (b) -1312 kJ mol^{-1}

(c) -164 kJ mol^{-1} (d) -82 kJ mol^{-1}

Question 14:

Atomic emission spectra of an element cannot be used to

(a) identify the atom

(b) determine the mass number of the nucleus of atom

(c) measure the difference in energy between pairs of stationary state of atom

(d) All of the above

Question 15:

In hydrogen spectrum, the different line of Lyman series are present in

(a) UV field

(b) IR field

(c) visible field

(d) far IR field

Question 16:

The energy of the electron in the second and the third Bohr orbits of the hydrogen atom are -5.42×10^{-12} erg and -2.41×10^{-12} erg respectively. The wavelength of emitted light when the electrons drops from the third to the second orbit is

(a) 0.660 nm (b) 330 nm (c) 660 nm (d) 590 nm

Question 17:

The frequency of radiation emitted when electron falls from $n=4$ to $n=1$ in a hydrogen atom will be (Given ionisation energy of H$= 2.18 \times 10^{-18}$ J atom^{-1} and $h = 6.625 \times 10^{-34}$ Js)

(a) 1.54×10^{15} s^{-1}
(b) 1.03×10^{15} s^{-1}
(c) 3.08×10^{15} s^{-1}
(d) 2.00×10^{15} s^{-1}

Question 18:

The de-Broglie wavelength associated with a ball of mass 1 kg having kinetic energy 0.5 J is

(a) 6.626×10^{-34} m
(b) 13.20×10^{-34} m
(c) 10.38×10^{-21} m
(d) 6.626×10^{-34} Å

Question 19:

A particle *A* moving with a certain velocity has the de-Broglie wavelength of 1 Å. For particle *B* with mass 25% of *A* and velocity 75% of *A*, calculate the de-Broglie wavelength.

(a) 3 Å (b) 5.33 Å (c) 6.88 Å (d) 0.48 Å

Question 20:

Uncertainty in the position of an electron (mass $= 9.1 \times 10^{-31}$kg) moving with a velocity 300 ms^{-1}, accurate upon 0.001% will be ($h = 6.63 \times 10^{-34}$ Js)

(a) 19.3×10^{-2} m
(b) 5.76×10^{-2} m
(c) 1.93×10^{-2} m
(d) 3.84×10^{-2} m

Question 21:

Which of the following options does not represent ground state electronic configuration of an atom?

[NCERT Exemplar]

(a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^8\ 4s^2$
(b) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^9\ 4s^2$
(c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$
(d) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5 4s^1$

Question 22:

Choose the correct statement among the following.

(a) ψ^2 represents the atomic orbital
(b) The number of peaks in radial distribution is $(n - l)$
(c) A node is a point in space around nucleus where the wave function ψ has zero value
(d) All of the above

Question 23:

The number of nodal planes in p_x is
(a) 1
(b) 2
(c) 3
(d) 0

Question 24:

Four different sets of quantum numbers for 4 electrons are given below:

$$e_1 = 4, 0, 0, -\frac{1}{2}; \quad e_2 = 3, 1, 1, -\frac{1}{2};$$

$$e_3 = 3, 2, 2, +\frac{1}{2}; \; e_4 = 3, 0, 0, +\frac{1}{2}.$$

The order of energy of e_1, e_2, e_3 and e_4 is

(a) $e_1 > e_2 > e_3 > e_4$
(b) $e_4 > e_3 > e_2 > e_1$
(c) $e_3 > e_1 > e_2 > e_4$
(d) $e_2 > e_3 > e_4 > e_1$

Question 25:

The set of quantum number for 19th electrons of chromium ($Z = 24$) is

(a) $4, 0, 0, +\frac{1}{2}$
(b) $4, 1, -1, +\frac{1}{2}$
(c) $3, 2, 2, +\frac{1}{2}$
(d) $3, 2, -2, +\frac{1}{2}$

Question 26:

The two electrons in s-subshell will differ in
(a) principal quantum number
(b) azimuthal quantum number
(c) magnetic quantum number
(d) spin quantum number

Question 27:

If an electron has spin quantum number of + 1/2 and a magnetic quantum number of -1, it cannot be represented in a
(a) s-orbital
(b) p-orbital
(c) d-orbital
(d) f-orbital

Question 28:

Consider the following sets of quantum numbers.

	n	*l*	*m*	*s*
(i)	3	0	0	$+\frac{1}{2}$
(ii)	2	2	1	$+\frac{1}{2}$
(iii)	4	3	-2	$-\frac{1}{2}$
(iv)	1	0	-1	$-\frac{1}{2}$
(v)	3	2	3	$+\frac{1}{2}$

Which of the following sets of quantum numbers is not possible?
(a) (ii), (iii) and (iv)
(b) (i), (ii), (iii) and (iv)
(c) (ii), (iv) and (v)
(d) (i) and (iii)

Question 29:

If principal quantum number n = 4 , the total number of orbitals having l = 3 is

(a) 3
(b) 7
(c) 5
(d) 9

Question 30:

Direction (Q. NO. 31) In the following question, more than one of the answers given may be correct. Select the correct answers and mark it according to the codes.

Codes

(a) 1,2 and 3 are correct
(b) 1 and 2 are correct
(c) 2 and 4 are correct
(d) 1 and 3 are correct

Question 31:

Which of the following statement(s) is (are) correct?

1. The electronic configuration of Cr is [Ar] $3d^5 4s^1$ (Atomic number of Cr = 24)
2. The magnetic quantum number may have a negative value.
3. In silver atom, 23 electrons have a spin of one type and 24 of the opposite type. (Atomic number of Ag = 47)
4. For valency electron of potassium, value of n is 2.

Question 32:

Direction (Q. Nos. 32-33) *In the following questions, match Column* I *with Column* II *and choose the correct codes.*

32.

	Column I		Column II
A.	$mvr = \frac{nh}{2\pi}$	1.	de-Broglie equation
B.	$\lambda = \frac{h}{P}$	2.	Electrons total energy
C.	$\frac{-e^2}{2r}$	3.	Paschen series
D.	Infrared	4.	Bohr's equation

Codes

	A	B	C	D		A	B	C	D
(a)	4	1	2	3	(b)	4	1	3	2
(c)	1	3	4	2	(d)	2	1	3	4

Question 33:

	Column I		Column II
A.	V_n/K_n	1.	0
B.	If radius of nth orbit $\propto E_n^x$, $x = ?$	2.	+ 1
C.	Angular momentum in lowest orbital	3.	–1
D.	If $\frac{1}{r^n} \propto Z^y$, $y = ?$	4.	–2

Codes

	A	B	C	D
(a)	1	2	3	4
(b)	3	4	2	1
(c)	4	3	1	2
(d)	2	1	4	3

Question 34:

Which of the following pairs of ions are isoelectronic and isostructural? **[CBSE-AIPMT 2015]**

(a) CO_3^{2-}, SO_3^{2-}
(b) ClO_3^-, CO_3^{2-}
(c) SO_3^{2-}, NO_3^-
(d) ClO_3^-, SO_3^{2-}

Question 35:

The number of d-electrons in Fe2+ (Z = 26) is not equal to the number of electrons in which one of the following?

(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)

Question 36:

Be^{2+} is isoelectronic with which of the following ions? **[CBSE-AIPMT 2014]**

(a) H^+ (b) Li^+ (c) Na^+ (d) Mg^{2+}

Question 37:

What is the maximum numbers of electrons that can be associated with the following set of quantum numbers? n = 3, l =1, m = -1

(a) 6
(b) 4
(c) 2
(d) 10

Question 38:
Maximum number of electrons in a subshell with l= 3 and n= 4 is
(a) 14
(b) 16
(c)10
(d) 12

Question 39:
The correct set of four quantum numbers for the valence electron of rubidium atom (Z = 37) is
(a)5,1, 1, + 1/2
(b)6,0, 0, + 1/2
(c)5,0, 0, + 1/2
(d) 5,1, 0, + 1

Question 40:
Which of the following is non-permissible?
(a) n = 4, l = 3, m = 0
(b) n = 4, l = 2, m = 1
(c) n = 4, l= 4, m = 1
(d) n = 4, l= 0, m = 0

Question 41:
The ratio of the difference in energy between the first and second Bohr orbit to that between the second and the third Bohr orbit is
(a)1/2
(b)1/3
(c)4/9
(d)27/5

Question 42:
Calculate the wavelength of light required to break the bond between two chlorine atoms in a chlorine molecule. The Cl—Cl bond energy is 243 kJ mol^{-1} ($h = 6.6 \times 10^{-34}$ Js; $c = 3 \times 10^{8}$ ms^{-1}, Avogadro's number $= 6.02 \times 10^{-23}$ mol^{-1}) **[AIIMS 2010]**
(a) 4.91×10^{-7} m
(b) 4.11×10^{-6} m
(c) 8.81×10^{-31} m
(d) 6.26×10^{-21} m

Question 43:

The energies E_1 and E_2 of two radiations are 25 eV and 50 eV respectively. The relation between their wavelengths i.e. λ_1 and λ_2 will be **[CBSE-AIPMT 2011]**

(a) $\lambda_1 = 2\lambda_2$ (b) $\lambda_1 = 4\lambda_2$

(c) $\lambda_1 = \frac{1}{2}\lambda_2$ (d) $\lambda_1 = \lambda_2$

Question 44:

The total number of atomic orbitals in fourth energy level of an atom is

(a) 16
(b) 32
(c) 4
(d) 8

Question 45:

If $n = 6$, the correct sequence for filling of electrons will be **[CBSE-AIPMT 2011]**

(a) $ns \longrightarrow (n-1)d \longrightarrow (n-2)f \longrightarrow np$
(b) $ns \longrightarrow (n-2)f \longrightarrow np \longrightarrow (n-1)d$
(c) $ns \longrightarrow np \longrightarrow (n-1)d \longrightarrow (n-2)f$
(d) $ns \longrightarrow (n-2)f \longrightarrow (n-1)d \longrightarrow np$

Question 46:

Which of the following statement in relation to the hydrogen atom is correct?

(a) 3s, 3p and 3 d-orbitals all have the same energy
(b) 3s and 3p-orbitals are of lower energy than 3d-orbital
(c) 3p-orbital is lower in energy than 3d-orbital
(d) 3s-orbital is lower in energy than 3p-orbital

Question 47:

The wavelength of a spectral line emitted by hydrogen atom in the Lyman series is cm. What is the value ofn2? (where, R= Rydberg constant)

(a) 2
(b) 3
(c) 4
(d) 1

Question 48:

Which one of the following ions has electronic configuration $[Ar]\,3d^6$? (At. no : Mn = 25, Fe = 26, Co = 27, Ni = 28) **[CBSE-AIPMT 2010]**

(a) Ni^{3+} (b) Mn^{3+} (c) Fe^{3+} (d) Co^{3+}

Question 49:

Deuterium nucleus contains

(a) 1p + 1n

(b) 2p + On

(c) 1p + 1e

(d) 2p+ 2n

Question 50:

If the photon of the wavelength 150 pm strikes an atom and one of its inner bound elec velocity of $1.5\times 10^7\ ms^{-1}$, what is the energy with which it is bound to the nucleus?

(a) 1.2×10^2 eV (b) 2.15×10^3 eV (c) 7.6×10^3 eV (d) 8.12×10^3 eV

Answers:

1. (b)	**2.** (d)	**3.** (b)	**4.** (c)	**5.** (a)	**6.** (c)	**7.** (d)	**8.** (b)	**9.** (a)	**10.** (d)
11. (b)	**12.** (d)	**13.** (d)	**14.** (b)	**15.** (a)	**16.** (c)	**17.** (c)	**18.** (a)	**19.** (b)	**20.** (c)
21. (b)	**22.** (d)	**23.** (a)	**24.** (c)	**25.** (a)	**26.** (d)	**27.** (a)	**28.** (c)	**29.** (b)	**30.** (a)
31. (b)	**32.** (a)	**33.** (c)	**34.** (d)	**35.** (b)	**36.** (b)	**37.** (c)	**38.** (a)	**39.** (c)	**40.** (c)
41. (d)	**42.** (a)	**43.** (a)	**44.** (a)	**45.** (d)	**46.** (a)	**47.** (c)	**48.** (d)	**49.** (a)	**50.** (c)

Hints and Solution

1. NO_2^- and O_3 are isosteres because both species have same atoms (= 3) and number of electrons (24).

2. Number of electrons in $ClO_2^- = 17 + 16 + 1 = 34$ electrons

Number of electrons in $ClF_2^+ = 17 + 18 - 1 = 34$ electrons

Hence, ClO_2^- and ClF_2^+ are isoelectronic species.

3. $m_e = 9.109 \times 10^{-31}$ kg

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

4. $\because$ $$\lambda = \frac{c}{\nu} = \frac{3 \times 10^8}{8 \times 10^{15}} = 3.75 \times 10^{-8} \text{ m}$$
$$= 3.75 \times 10^{-8} \times 10^9 \text{ nm} = 4 \times 10^1 \text{ nm}$$

5. $$v = \frac{c}{\lambda} = \frac{3 \times 10^{10} \text{ cm s}^{-1}}{5.5 \times 10^{-8} \text{ cm}} = 5.5 \times 10^{17} \text{s}^{-1}$$

Further, $\frac{1}{2} mu^2 = h(\nu - \nu_o)$
$$= 6.6 \times 10^{-27} \text{ erg} (5.5 \times 10^{17} \text{s}^{-1} - 3.62 \times 10^{12} \text{s}^{-1})$$

6. $KE = \frac{1}{2} mv^2$ and KE does not depend on light intensity.

7. $$\lambda = \frac{h}{mv} \quad \ldots\text{(i)}$$
$$KE = \frac{1}{2} mv^2 \quad \ldots\text{(ii)}$$

Now, putting the value of v from Eq. (i), we get
$$KE = \frac{1}{2} m \times \left(\frac{h}{m \times \lambda}\right)^2 = \frac{1}{2} \times \frac{h^2}{m \times \lambda^2}$$
$$KE \propto \frac{1}{m} \text{ (if } \lambda \text{ same)}$$
$$\therefore \quad E_e > E_p > E_\alpha \quad (\because m_e < m_p < m_\alpha)$$

8. From Rydberg's equation,

$$\bar{\nu} = \frac{1}{\lambda} = RZ^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \Rightarrow \frac{1}{\lambda} \propto Z^2$$

$$\therefore \quad \frac{\lambda(He^+)}{\lambda(H)} = \frac{Z^2(H)}{Z^2(He^+)}$$

$$\frac{\lambda(He^+)}{91.2} = \frac{1}{4} \quad \text{or} \quad \lambda(He^+) = \frac{91.2}{4} = 22.8 \text{ nm}$$

9. $\frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right], \frac{1}{\lambda} = 1.097 \times 10^7 \text{m}^{-1}\left[\frac{1}{1^2} - \frac{1}{\infty^2}\right]$

$\lambda = 91 \times 10^{-9}\text{m} = 91 \text{nm}$

10. $r = \frac{0.529n^2}{Z}$;

$$\therefore \quad \frac{r_2}{r_1} = \frac{(2)^2}{(1)^2} = \frac{4}{1}$$

Area of orbital $= \pi r^2$

$$\therefore \quad \frac{A_2}{A_1} = \frac{\pi r_2^2}{\pi r_1^2} = \left(\frac{r_2}{r_1}\right)^2 = \left(\frac{4}{1}\right)^2 = \frac{16}{1}$$

11. Bohr's theory is applicable to H-like species (e.g. Li^{2+}).

12. Angular momentum of an electron $= mvr = \frac{nh}{2\pi}$

(n is the orbit number)

For 5th orbit, $mvr = \frac{5h}{2\pi} = \frac{2.5h}{\pi}$

13. The energy of second Bohr orbit of hydrogen atom (E_2) is -328 kJ mol^{-1}, therefore,

$$E_2 = -\frac{1312}{2^2} \text{ kJ mol}^{-1}, \; E_n = -\frac{1312}{n^2} \text{ kJ mol}^{-1}$$

If $n = 4$

$$E_4 = -\frac{1312}{4^2} \text{ kJ mol}^{-1} = -82 \text{ kJ mol}^{-1}$$

13. The energy of second Bohr orbit of hydrogen atom (E_2) is $-328\ \text{kJ mol}^{-1}$, therefore,

$$E_2 = -\frac{1312}{2^2}\ \text{kJ mol}^{-1},\quad E_n = -\frac{1312}{n^2}\ \text{kJ mol}^{-1}$$

If $n = 4$

$$E_4 = -\frac{1312}{4^2}\ \text{kJ mol}^{-1} = -82\ \text{kJ mol}^{-1}$$

14. Atomic emission spectra of an element cannot be used to determine the mass number of the nucleus of atom.

15. According to hydrogen spectrum series, Lyman series is present in UV field.

16. Transition energy,

$$\Delta E = E_3 - E_2$$

$$= [-2.41 - (-5.42)] \times 10^{-12}\text{erg} = 3.01 \times 10^{-12}\text{erg}$$

$$= 3.01 \times 10^{-19}\ \text{J} \qquad [\because 1\ \text{erg} = 10^{-7}\ \text{J}]$$

$$\because \quad \Delta E = \frac{hc}{\lambda}$$

$$\therefore \quad \lambda = \frac{hc}{\Delta E} = \frac{6.625 \times 10^{-34} \times 3 \times 10^8}{3.01 \times 10^{-19}}$$

$$= 660 \times 10^{-9}\text{m} = 660\ \text{nm}$$

17. Ionisation energy of H $= 2.18 \times 10^{-18}$ J atom^{-1})

$\therefore \quad E_1 =$ (energy of 1s-orbit of H-atom)

$$= 2.18 \times 10^{-18} \text{ J atom}^{-1}$$

$$\Rightarrow E_n = \frac{-2.18 \times 10^{-18}}{n^2} \text{ J atom}^{-1} \qquad (\because Z = 1 \text{ for H-atom})$$

Now, $$\Delta E = E_4 - E_1 = \frac{-2.18 \times 10^{-18}}{4^2} - \frac{-2.18 \times 10^{-18}}{1^2}$$

$$= -2.18 \times 10^{-18} \times \left[\frac{1}{4^2} - \frac{1}{1^2}\right]$$

$$= -2.18 \times 10^{-18} \times \left(-\frac{15}{16}\right)$$

$$= +2.0437 \times 10^{-18} \text{ J atom}^{-1}$$

$$\nu = \frac{\Delta E}{h} = \frac{2.0437 \times 10^{-18} \text{ J atom}^{-1}}{6.625 \times 10^{-34} \text{ Js}}$$

$$= 3.084 \times 10^{15} \text{ s}^{-1}\text{atom}^{-1}$$

18. $$\lambda = \frac{h}{\sqrt{2m\text{KE}}} = \frac{6.626 \times 10^{-34}}{\sqrt{2 \times 1 \times 0.5}} = 6.626 \times 10^{-34}\text{m}$$

19. $$\lambda_A = \frac{h}{m_A v_A} \text{ and } \lambda_B = \frac{h}{m_B v_B}$$

$$\frac{\lambda_A}{\lambda_B} = \frac{m_B v_B}{m_A v_A}$$

Here, $m_B = 25\%$ of $m_A = 0.25 m_A$

$v_B = 75\%$ of $v_A = 0.75 v_A$

$$\therefore \quad \frac{1}{\lambda_B} = \frac{0.25 m_A \times 0.75 v_A}{m_A \times v_A}$$

or $\lambda_B = 5.33$ Å

20. $\Delta x \cdot \Delta v \geq \dfrac{h}{4\pi m}$

$$\Delta x = \frac{6.63 \times 10^{-34}}{4 \times 3.14 \times 9.1 \times 10^{-31} \times 300 \times 0.001 \times 10^{-2}}$$
$$= 0.01933$$
$$= 1.93 \times 10^{-2} \text{m}$$

21. Correct configuration should be $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^1$ for the copper which has atomic number 29 ($_{29}$Cu). Due to extra stability of fully filled orbital of d-subshell, the last electron enter into d-orbital instead of s-orbital.

22. All statements are correct.

23. Number of nodal plane = 1.

24. Energy order $n = 3 < n = 4$

$$l = 0 < l = 1 < l = 2$$

Therefore, the order is $e_3 > e_1 > e_2 > e_4$.

25. $(Z = 24) = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1, 3d^5$

For 19th electron (i.e. $4s^1$)

$$n = 4, l = 0, m = 0, s = +\frac{1}{2}$$

26. The two electrons will have opposite spins.

27. $m = -1$ is not possible for s-orbital (e.g. $l = 0$).

28. Correct sets of quantum numbers are:

	n	l	m	s
(ii)	2	1	1	$+\frac{1}{2}$
(iv)	1	0	0	$-\frac{1}{2}$
(v)	3	2	2	$+\frac{1}{2}$

29. For $n = 4$, $l = 3$ (f-subshell), number of values of $m = 2l + 1 = 7$ values.
Number of orbitals in f-subshell is 7.

30. $4l + 2$

31. 1. Cr = [Ar] = $3d^5 4s^1$, an exception to Aufbau principle.

2. For a given value of l, m can have any value from $-l$ to $+l$, so it can have negative value.

3. Ag is in copper group with $d^{10}s^1$ configuration, i.e., 46 electrons are spin paired.

4. EC of K (19) = $1s^2, 2s^2, 2p^6, 3s^2, 3p^6\ 4s^1$
The valence electron is present in $4s^1$. Hence, the value of principal quantum number (n) is 4.

32. (A) According to Bohr, among the infinite number of possible circular orbitals, an electron can revolve only in those orbitals whose angular momentum (mvr) is an integral multiple of the factor $\frac{h}{2\pi}$, i.e. $mvr = \frac{nh}{2\pi}$

This equation is known as Bohr's equation.

(B) $\lambda = \frac{h}{p}$ is known as de-Broglie equation.

(C) Total energy of an electron,

$$E = \text{KE} + \text{PE} = \frac{e^2}{2r} + \left(\frac{-e^2}{r}\right) = -\frac{e^2}{2r}$$

(D) In infrared region, Paschen series of lines is obtained.

33. (A) According to Bohr's theory,

Total energy, $E_n = K_n + V_n$

where, kinetic energy, $K_n = \frac{1}{2}\frac{ZKe^2}{r_n}$

and potential energy, $V_n = -\frac{ZKe^2}{r_n}$

$$\therefore \qquad \frac{V_n}{K_n} = \frac{-ZKe^2}{r_n} \times \frac{2r_n}{ZKe^2} = -2$$

(B) $\because \qquad E_n \propto \frac{1}{n^2} \Rightarrow E_n \propto \frac{1}{r_n^2}$

$\because$ Radius of nth orbit $r_n \propto E_n^x$

$\therefore \qquad x = -1$

(C) Angular momentum $= \frac{h}{2\pi}\sqrt{l(l+1)}$

where, $\qquad l = 0, 1, 2 \ldots$

For lower orbit $\qquad n = 1, l = 0, m = 0$

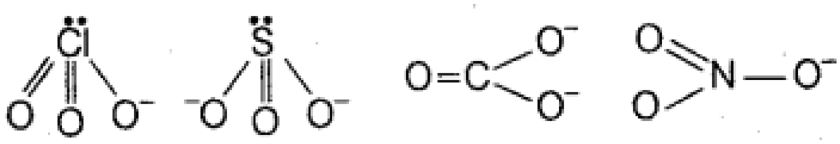
Hence, angular momentum of lowest orbit

$$= \frac{h}{2\pi}\sqrt{0\,(0+1)} = 0$$

(D) Given, $\frac{1}{r^n} = Z^y$

We know that $r_n \propto \frac{1}{Z}$, $\therefore y = 1$

34.



ClO_3^- and SO_3^{2-} are isoelectronic and pyramidal.

35. Electronic configuration of Fe^{2+}. $[Ar]3d^6 4s^0$

Number of $de^- = 6$

$Ne = 1s^2, 2s^2, 2p^6 (6 - pe^-)$

$Mg = 1s^2, 2s^2, 2p^6, 3s^2 (6 - se^-)$

$Cl = 1s^2, 2s^2, 2p^6, 3s^2, 3p^5 (11 - pe^-)$

$Fe = 3d^6, 4s^2 (6 - de^-)$

36. Both Be^{2+} and Li^+ have 2 electrons.

37. The orbital of the electron having $n = 3, l = 1$ and $m = 1$ is $3p_z$ (as nl_m) and an orbital can have a maximum of two electrons with opposite spins,

$\therefore 3p_z$ orbital contains only two electrons or only 2 electrons are associated with $n = 3, l = 1, m = -1$.

38. n represents the main energy level and l represents the subshell.

If $n = 4$ and $l = 3$, the subshell is $4f$.

In f subshell, there are 7 orbitals and each orbital can accommodate a maximum of two electrons, so, maximum number of electrons in $4f$ subshell $= 7 \times 2 = 14$

39. ${}_{37}Rb = {}_{36}[Kr]\ 5s^1$

Its valence electron is $5s^1$.

So, $n = 5, l = 0$ (For s-orbital)

$m = 0$ (As $m = -l$ to 0 to $+l$)

$s = +\frac{1}{2}$

40. According to the rules for quantum number the possible values of n, l, m and s are

$n = 1$ to ∞, any whole number,

$l = 0$ to $(n-1)$ for every value of n.

$m = -l$ to 0 to $+l$, for every value of l,

and $s = \frac{1}{2}$ or $-\frac{1}{2}$

(a) $n = 4, l = 3, m = 0$

All the values are according to rules.

(b) $n = 4, l = 2, m = 1$

All the values are according to rules.

(c) $n = 4, l = 4, m = 1$

$\because$ The value of l can have maximum $(n-1)$ value i.e. 3 in this case.

$\therefore$ This set of quantum numbers is non-permissible.

(d) $n = 4, l = 0, m = 0$

All the values are according to rules.

$\therefore$ Choice (a), (b) and (d) are permissible.

41. $E_1 - E_2 = 1312 \times Z^2\left(\frac{1}{1^2} - \frac{1}{2^2}\right) = 1312 \times Z^2\left(\frac{3}{4}\right)$

$E_2 - E_3 = 1312 \times Z^2\left(\frac{1}{2^2} - \frac{1}{3^2}\right) = 1312 \times Z^2\left(\frac{5}{36}\right)$

From Eqs. (i) and (ii), we get

$$\frac{E_1 - E_2}{E_2 - E_3} = \frac{3 \times 36}{4 \times 5} = \frac{27}{5}$$

42. Energy required to break one Cl—Cl bond

$$= \frac{\text{Bond energy per mole}}{\text{Avogadro's number}} = \frac{243 \times 10^3}{6.023 \times 10^{23}}\ \text{J}$$

Let, the wavelength of the photon required to break one Cl—Cl bond be λ.

$$\lambda = \frac{hc}{E} = \frac{6.6 \times 10^{-34} \times 3 \times 10^8 \times 6.023 \times 10^{23}}{243 \times 10^3} = 4.91 \times 10^{-7}\ \text{m}$$

43. $E_1 = 25$ eV, $E_2 = 50$ eV

$$E_1 = \frac{hc}{\lambda_1} \text{ and } E_2 = \frac{hc}{\lambda_2} \text{ or } \frac{E_1}{E_2} = \frac{\lambda_2}{\lambda_1}$$

or $\frac{25}{50} = \frac{\lambda_2}{\lambda_1}$ or $\lambda_1 = 2\lambda_2$

44. Number of atomic orbitals in an orbit $= n^2 = 4^2 = 16$.

45. $6s \rightarrow 4f \rightarrow 5d \rightarrow 6p$ for $n = 6$

46. Electronic configuration of hydrogen atom is $1s^1$ and its $3s$, $3p$ and $3d$-orbitals will have same energy wrt $1s$-orbital.

47. For Lyman series,

$$\frac{1}{\lambda} = R\left[\frac{1}{1^2} - \frac{1}{n_2^2}\right] \Rightarrow \frac{15R}{16} = R\left[\frac{1}{1^2} - \frac{1}{n_2^2}\right]$$

$$\frac{15R}{16R} = \left[\frac{n_2^2 - 1}{n_2^2}\right] \Rightarrow \frac{15}{16} = \frac{n_2^2 - 1}{n_2^2}$$

$$15n_2^2 = 16n_2^2 - 16 \Rightarrow n_2^2 = 16, n_2 = 4$$

48. Ni^{3+} (28) = [Ar] $3d^7$ Mn^{3+}(25) = [Ar] $3d^4$

Fe^{3+}(26) = [Ar] $3d^5$ Co^{3+}(27) = [Ar] $3d^6$

49. Deuterium, an isotope of hydrogen, is $_1H^2$.

Thus, it contains 1 proton and 1 (2 – 1) neutron in its nucleus.

50. Energy of photon,

$$E = \frac{hc}{\lambda} = \frac{6.6 \times 10^{-34} \times 3 \times 10^8}{1.5 \times 10^{-10}} = 1.32 \times 10^{-15} \text{ J}$$

Energy of ejected electron,

$$E' = \frac{1}{2}mv^2 = \frac{1}{2} \times 9.1 \times 10^{-31} \times (1.5 \times 10^7)^2$$

$$= 1.024 \times 10^{-16}$$

Total energy of photon

= binding energy of electron + energy of ejected electron

1.32×10^{-15} = binding energy + 1.024×10^{-16}

$\therefore$ Binding energy $= (1.32 \times 10^{-15}) - (1.024 \times 10^{-16})$

$$= 1.2176 \times 10^{-15} \text{ J}$$

$$= \frac{1.2176 \times 10^{-15}}{1.6 \times 10^{-19}} \text{ eV}$$

$$= 7.6 \times 10^3 \text{ eV}$$