4. Structure of Atom

Exercises

1. Choose correct option.

A. The energy difference between the shells goes on ..... when moved away from the nucleus.

Ans. decreasing

B. The value of Plank's constant is -

Ans. 6.626×10^{-34} Js

C. p-orbitals are..... in shape.

Ans. dumb-bell

D. "No two electrons in the same atoms can have identical set of four quantum number". This statement is known as -

Ans. Pauli's exclusion principle

E. Principal Quantum number describes–

Ans. size of the orbital

2. Make the pairs. (Answer are given directly)

Ans.

Column A	Column B(Ans)
Neutrons	Chadwick
p-orbital	— 1.6×10^{-19} C
Charge on electron	Six elections
Lyman series	Ultraviolet region

3. Complete the following information about the isotopes in the chart given below:

Ans.

Substance	Mass Number	Number of		
		Protons	Neutrons	Electron
Carbon–14	14	6	14 – 6 = 8	6
Lead–208	208	82	208 – 82 = 126	82
Chlorine–35	35	17	35 – 17 = 18	17
Uranium–238	238	92	238 – 92 = 146	92
Oxygen–18	18	8	18 – 8 = 10	8
Radium–223	223	88	223 – 88 = 135	88

4. Match the following :

Ans.

Element	No. of Neutron (Ans)
$^{40}_{18}Ar$	22
$^{14}_{6}C$	8
$^{40}_{19}K$	21
$^{14}_{7}N$	7

5. Answer in one sentence :

A. If an element 'X' has mass number 11 and it has 6 neutrons, then write its representation.

Ans. Atomic number = 11 – 6 = 5
∴ Element is $^{11}_{5}$B

B. Name the element that shows simplest emission spectrum.

Ans. Since hydrogen atom has one electron it shows the simplest emission spectrum.

C. State Heisenberg uncertainty principle.

Ans. Heisenberg uncertainty principle: This principle states that it is not possible to deter mine simultaneously the position and momentum of a moving microscopic particle like electron with absolute certainty.

D. Give the names of quantum numbers.

Ans. (1) Principal quantum number denoted by n.

(2) Azimuthal quantum number denoted by l.

(3) Magnetic quantum number denoted by
m_l.

(4) Spin quantum number denoted by
m_s.

E. Identify from the following the isoelectronic species:

Ans.

Species:	Number of electrons:
Ne	10
O^{2-}	10
Na^{+}	10
Ar	18
Cl^{2-}	18
$K^{\oplus}$	18

∴ Isoelectronic species: (i) Ne, O^{2-}, Na^{+}

6. Answer the following questions.

A. Differentiate between Isotopes and Isobars.

Ans.

Isotopes :	Isobars
1. Isotopes are the atoms of the same element.	1. Isobars are the atoms of different elements.
2. Isotopes have same atomic number but different mass numbers.	2. Isobars have different atomic numbers but same mass number.
3. They have same number of electrons.	3. They have different number of electrons.
4. They have same number of protons but different number of neutrons.	4. They have different number of protons and neutrons.
5. They have same chemical properties.	5. They have different chemical properties.
6. For example: $^{35}_{17}Cl$ and $^{37}_{17}Cl$	6. For example: $^{40}_{18}Ar$ and $^{40}_{19}K$

B. Define the terms:
i. Isotones

Ans. Isotones: The atoms of different elements having same number of neutrons in their nuclei are called Isotones.

Isotones	Atomic Number Z	Mass Number A	Number of Neutrons N = A – Z

$^{11}_{5}B$	5	11	6
$^{12}_{6}C$	6	12	6

ii. Isoelectronic species

Ans. Isoelectronic species are defined as atoms and ions having the same number of electrons.
e.g. Ar, $Ca^{2}+$, and K containing 18 electrons each.

iii. Electronic configuration

Ans. Electronic configuration of an atom is defined as the distribution of its electrons in orbitals.

C. State and explain Pauli's exclusion principle. (Rotate your phone)

Ans. (1) Statement of Pauli's principle: No two electrons in an atom can have all the four quantum numbers, (n. 1, m and s) same.
Or
Only two electrons may exist in the given orbital having three quantum numbers same but fourth quantum number being different with opposite spins.

(2) This principle describes the capacity of a sub- shell or orbital to accommodate maximum number of electrons.

(3) Consider helium atom, which has two electrons. The four quantum numbers of two electrons in He atom will be,

Electron No.	Quantum numbers			
	n	*l*	*m*	*s*
1st electron	1	0 (*s*)	0	$+\frac{1}{2}$
2nd electron	1	0 (*s*)	0	$-\frac{1}{2}$

$1s^2$ 

Hence 1s-orbital (n+1=3+2=5) is filled prior to 4p-orbital (n+1=4+1=5).

D. State Hund's rule of maximum multiplicity with a suitable example.

Ans. (1) Statement of Hund's rule of maximum multiplicty: It states that pairing of electrons in the orbitals belonging to the same subshell does not occur unless each orbital belonging to that subshell has accommodated one electron each.

(2) Consider filling of p-subshell which has three degenerated orbitals namely

$p_x p_y p_z$

(3) After filling three electrons, one in each with same spins, the next electrons enter with pairing.

4 electrons
And

6 electrons

(4) It is observed that half filled and completely filled set of degenerate orbitals have extra stability.

E. Write the drawbacks of Rutherford's model of an atom.

Ans. Rutherford's atomic model has following drawbacks:

(1) According to the classical electromagnetic theory, a revolving charged particle like electron should emit radiation and lose energy. Due to this, electron should come closer to the nucleus by following a spiral path and finally fall into the nucleus, giving unstable atom. But in practice this does not happen.

(2) The orbital motion is an accelerated motion accompanied by continuous change in the velocity of electron due to changing direction.

(3) Orbital motion of an electron dos not explain the change in energy and hence atomic spectrum.

(4) The electron would follow a spiral path and fall into the nucleus.

(5) This model does not explain the distribution of electrons around the nucleus and their energies.

F. Write postulates of Bohr's Theory of hydrogen atom.

Ans. Niels Bohr proposed an atomic model for H atom or hydrogen like atom with

one electron. The postulates of Bohr's theory are as follows:

(1) The electron in the hydrogen atom can move around the nucleus in certain permitted circular orbits arranged concentrically in increasing order of energy.

(2) The energy of an electron in the orbit does not change with time. Hence orbits are called stationary orbits. On absorption of required energy electron moves from lower orbit to higher orbit. The transition from higher energy orbit to lower energy orbit is accompanied by emission of energy in the form of electromagnetic radiation.

(3) The frequency v of radiation absorbed or emit ted on transition between two stationary orbits differing by energy
ΔE is given by,
$v = \frac{\Delta B}{h} = \frac{E_2 - E_1}{h}$
where E_1 and E_2 are the energies of the lower and higher energy states respectively. Above equation represents Bohr's frequency rule.

(4) An electron can occupy only those orbits in which its angular momentum is integral (n)
multiple of h $\frac{h}{2\pi}$
$\text{Angular momentum} = \frac{nh}{2\pi}$

G. Mention demerits of Bohr's Atomic model.

Ans. (1) Bohr's atomic model fails to account for finer details of hydrogen atom spectrum obtained from high resolving sophisticated spectroscope.

(2) It can not explain the spectra of atoms other than hydrogen or multielectron atoms.

(3) It can not explain the splitting of spectral lines in the presence of a magnetic field (Zeeman effect) or electric field (stark effect).

(4) It fails to explain the ability of atoms to combine and form molecules by chemical bonds.

H. State the order of filling atomic orbitals following the Aufbau principle.

Ans. (1) Aufbau is a German word which means building up. This principle explains the sequence of filling up of orbitals with electrons.

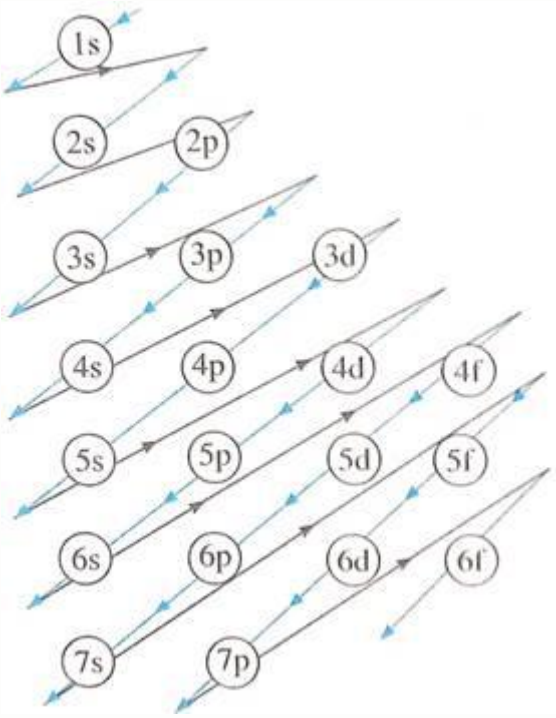
(2) Aufbau principle: It states that in the ground state of an atom, the orbitals are

filled with electrons in order of the increasing energies.

(3) The orbitals are filled in order of increasing value of (n+1). For example, 4s-orbital (n+1=4+0=4) is filled prior to 3d-orbital (n+1=3+2=5).

(4) Among two orbitals having same (n+1) value, that orbital with lower value of n will be filled first. For example, 3d-orbital (n+1=3+2=5) is filled prior to 4p-orbital (n+1=4+1=5).

(5) The increasing order of energy of different orbitals is as follows:
1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s <4d < 5p < 6s < 4f <5d < 6p < 7s <5f < 6d



I. Explain the anomalous behavior of copper and chromium.

Ans. Copper:

a. Copper (Cu) has atomic number 29.

b. Its expected electronic configuration is 2 $1s^2$ $2s^2$ 2p $3s^2$ 3p6 $4s^2$ $3d^2$.

c. The 3d orbital is neither half-filled nor fully filled. Hence, it has less stability.

d. Due to inter electronic repulsion forces, one 4s electron enters into 3d orbital. This makes 3d orbital completely filled and 4s orbital half-filled which gives extra stability and the electronic configuration of Cu becomes, $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^1$ $3d^{10}$.

Chromium:

a. Chromium (Cr) has atomic number 24.

b. Its expected electronic configuration is $1s^2\ 2s^2 2p6\ 3s^2\ 3p6\ 4s^2\ 3d^{\underline{a}}$.

c. The 3d orbital is less stable as it is not half-filled.

d. Due to inter electronic repulsion forces, one 4s electron enters into 3d orbital. This makes 4s and 3d orbitals half filled which gives extra stability and the electronic configuration of Cr becomes $1s^2\ 2s^2\ 2p\ 3s^2\ 3p6\ 4s^1\ 3d5$.

J. Write orbital notations for electrons in orbitals with the following quantum numbers.
a. n = 2, 1 = 1

Ans. the orbital is 2p

b. n = 4, 1 = 2

Ans. the orbital is 4d

c. n = 3, 1 = 2

Ans. the orbital is 3d.

K. Write electronic configurations of Fe, Fe^{2+}, Fe^{3+}

Ans. $_{26}Fe$ [Ar] $3d^6\ 4s^2$
Fe^{2+} [Ar] $3d^6$
Fe^{3+} [Ar] $3d^5$

L . Write condensed orbital notation of electronic configuration of the following elements:
a. Lithium (2-3)

Ans. $_3Li$ [He] $2s^1$

b. Carbon (Z=6)

Ans. $_6C$ [He] $2s^2\ 2p_x^1\ 2p_y^1$ OR [He] $2s^2\ 2p^2$

c. Oxygen (Z-8)

Ans. $_8O$ [He] $2s^2\ 2p_x^2 2p_y^1\ 2p_z^1$ OR [He] $2s^2\ 2p^4$

d. Silicon (Z=14)

Ans. $_{14}$Si [Ne] $3s^2\ 3p_x^1\ 3p_z^1$ OR [Ne] $3s^2\ 3p^2$

e. Chlorine (Z=17)

Ans. $_{17}$C1 [Ne] $3p_x^2\ 3p_y^2\ 3p_z^1$ OR [Ne] $3s^2\ 3p^5$

f. Calcium (Z=20)

Ans. $_{20}$Ca [Ar] $4s^2$

M. Draw shapes of 2s and 2p orbitals.

Ans. 2s orbitals:

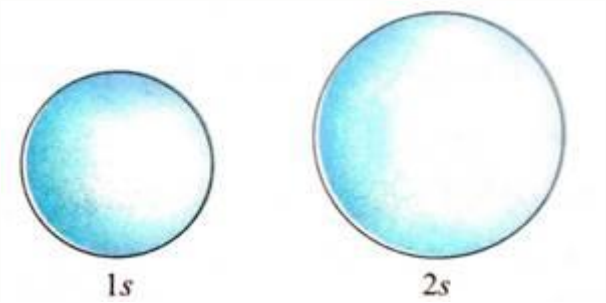


Figure shows the boundary surface diagram of atomic orbitals 1s and 2s, which are spherical in shape. Here, a boundary surface is drawn in space for an orbital such that the value of probability density 2 is constant and encloses a region where the probability of finding electron is typically more than 90%. Such a boundary surface diagram is a good representation of shape of an orbital.

2p orbitals:

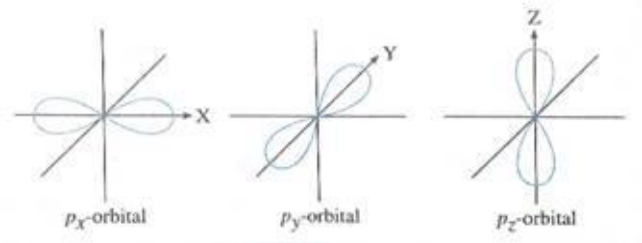
(1) This orbital has Azimuthal quantum number, 1=1 and magnetic quantum numbers m = + 1, 0,- 1.

(2) Hence p-orbital has three orientations along three coordinates namely, x, y and z. Therefore, p-orbitals are designated as
p_x, p_y *and* p_z
they determine the geometry of molecules.

(3) The shape of p-orbital resembles a dumb-bell. Hence p-orbital has two lobes separated by a nodal plane having zero electron density.

(4) These three p-orbitals are degenerate, i.e., they are equivalent in energy in the absence of an external magnetic field.

(5) The size and energy of p-orbitals increase with the increase in principal energy level or n, in the order of 2p < 3p < 4p.



N. Explain in brief, the significance of azimuthal quantum number.

Ans. i. Azimuthal quantum number is also known as the subsidiary quantum number and is represented by the letter l.

ii. It represents the subshell to which the electron belongs. It also defines the shape of the orbital that is occupied by the electron.

iii. Its value depends upon the value of principal quantum number 'n'. It can have only positive values between 0 and (n − 1).

iv. Atomic orbitals with the same value of 'n' but different values of 'l' constitute a subshell belonging to the shell for the given 'n'. The azimuthal quantum number gives the number of subshells in a principal shell. The subshells have I to be 0, 1, 2, 3 ... which are represented by symbols s, p, d, f,... respectively.

Number of subshells in shell.	Number of subshells in shell.	Number of subshells in shell.	Number of subshells in shell.	Number of subshells in shell.
K	1	0	s	1
L	2	0, 1	s, p	2
M	3	0, 1, 2	s, p, d	3
N	4	0, 1, 2, 3	s, p, d, f	4

O. If n=3, what are the quantum number l and m?

Ans. For n = 3,l = 0, 1, 2.
For $l = 0, m_l = 0$
For $l = 1, m_l = -1, 0, +1$
For $l = 2, m_l = -2, -1, 0, +1, +2$

P. The electronic configuration of oxygen is written as
$1s^2 2s^2\ 2p_x^2\ 2p_y^1\ 2p_z^1$ and not as

$1s^2 2s^2\ 2p_x^2, 2p_y^2\ 2p_z^o$, Explain.

Ans. By Hund's rule,
$^{16}_{8}O\ 1s^2\ 2s^2\ 2p_x^2\ 2p_y^1\ 2p_z^1$
The electronic configuration,
$1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^o$ violates Hund's rule.

Q. Write note on 'Principal Quantum number.

Ans. (1) Principal quantum number (n): This describes the orbit or shell of an atom to which the electron belongs. It is represented by 'n' which has integral values. The energy of an electron depends on the value of n.

n	1	2	3	4
orbit	1st	2nd	3rd	4th
shell	K	L	M	N

R. Using concept of quantum numbers, calculate the maximum numbers of electrons present in the 'M' shell. Give their distribution in shells, subshells and orbitals.

Ans. 'M' shell has,
n =3
l =0, 1, 2
m_l: When $l = 0, m_l = 0, m_s = \pm \frac{1}{2}$
m_l: When $l = 1, m_l = -1, 0, +1$
$m_l = \pm\frac{1}{2}, \pm\frac{1}{2}, \pm\frac{1}{2}$
m_l: When $l = 2,\ m_l = -2, -1, 0, +1, +2$
$m_l = \pm\frac{1}{2}, \pm\frac{1}{2}, \pm\frac{1}{2}, \pm\frac{1}{2}, \pm\frac{1}{2}$

S. Indicate the number of unpaired electrons in :
a. Si (Z=14)

Ans. $_{14}Si$ [Ne] $3s^2\ 3_x^1\ 3_y^1$:
Number of unpaired electrons = 2

b. Cr (Z=24)

Ans. $_{24}Cr$ [Ar] $3d^5\ 4s^1$
Number of unpaired electrons = 6

T. An atom of an element contains 29 electrons and 35 neutrons.
Deduce-
a. the number of protons

Ans. Since atom has 29 electrons, it has 29 protons.
The atomic number of the element is 29.
It is copper.

b. the electronic configuration of that element

Ans. ∴ Electronic configuration:
$_{29}Cu$ [Ar] $3d^{10}$ $4s^{1}$.