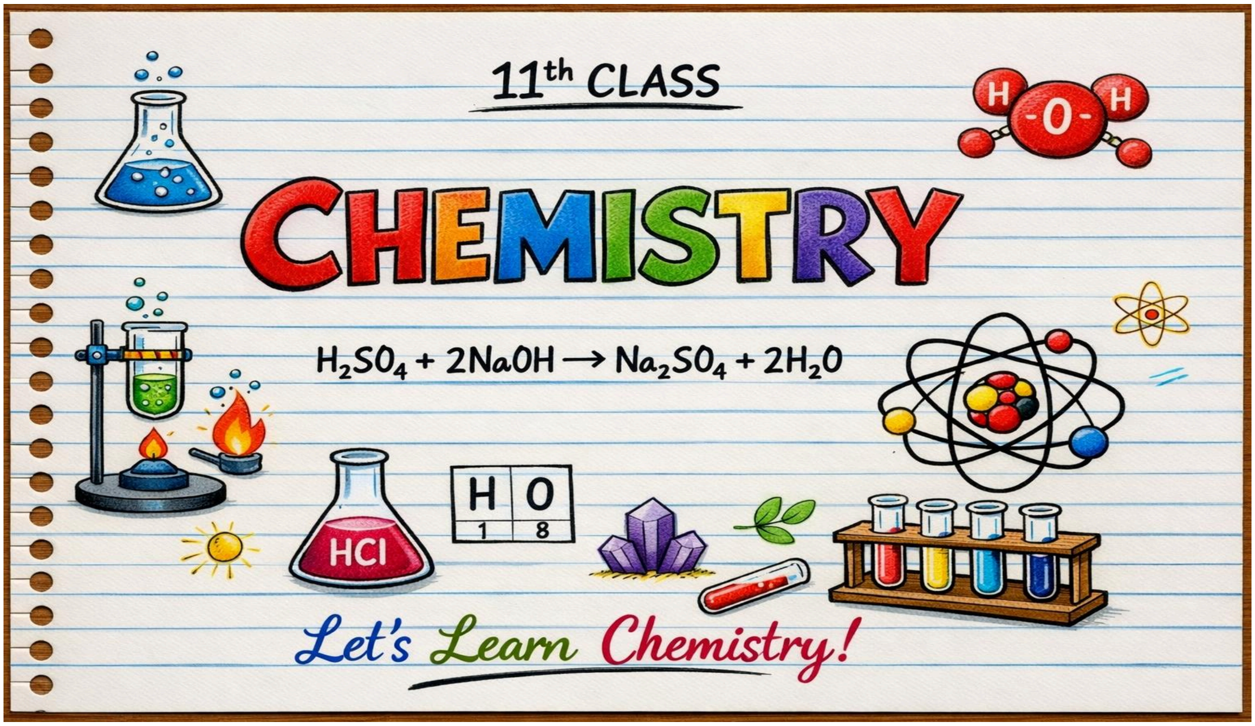

Class: 11th

Subject: Chemistry

Unit 2: ATOMIC STRUCTURE

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❖ Important MCQs:

1. Moseley's experiment proved that the square root of X-ray frequency is directly proportional to:

(a) mass number

(b) atomic number ✅

(c) neutron number

(d) electron number

2. Atomic number (Z) is equal to the number of:

(a) neutrons

(b) protons ✅

(c) nucleons

(d) electrons

3. Nucleon number (A) is the sum of:

(a) electrons + protons

(b) protons + neutrons ✅

(c) protons only

(d) neutrons only

4. The formula relating mass number (A), atomic number (Z), and neutron number (N) is:

(a) A = Z − N

(b) A = Z + N ✅

(c) A = N − Z

(d) A = Z × N

5. Elements with the same atomic number but different mass numbers are called:

(a) isobars

(b) isotopes ✅

(c) isotones

(d) ions

6. Identity of an element depends on:

(a) mass number

(b) neutron number

(c) atomic number ✅

(d) number of electrons

7. In an electric field, neutrons behave as:

(a) positively charged and move towards negative plate

(b) negatively charged and move towards positive plate

(c) neutral and move straight ✅

(d) neutral and move in circular path

8. Protons in an electric field are deflected towards:

(a) positive plate

(b) negative plate ✅

(c) straight path

(d) both plates

9. Electrons are deflected more than protons because:

(a) electrons have greater charge

(b) electrons have smaller mass ✅

(c) electrons are neutral

(d) electrons move slower

10. The angle of deflection of a particle in an electric field is proportional to:

(a) mass/charge

(b) charge/mass ✅

(c) mass × charge

(d) $mass^2$

11. The radius of deflection of a particle in an electric field is:

(a) proportional to charge/mass

(b) proportional to mass/charge ✅

(c) inversely proportional to $mass^2$

(d) independent of mass and charge

12. If a particle has high charge and very low mass, its deflection will be:

(a) small

(b) large ✅

(c) zero

(d) moderate

13. Increasing the mass of a particle (charge constant) will:

(a) decrease radius of deflection

(b) increase radius of deflection ✅

(c) keep radius same

(d) cause no movement

14. Which particle will have least deflection in an electric field among charged particles?

(a) electron

(b) proton ✅

(c) alpha particle

(d) beta particle

15. Moseley's work corrected the periodic table arrangement based on:

(a) atomic mass

(b) atomic number ✅

(c) valency

(d) density

16. Atomic emission spectrum is observed when an element in gaseous state is:

(a) cooled to absolute zero

(b) heated to high temperature or subjected to electrical discharge ✅

(c) dissolved in water

(d) mixed with another element

17. Atomic absorption spectrum is obtained when:

(a) an element emits radiation

(b) white light passes through gaseous atoms and certain wavelengths are absorbed ✅

(c) atoms are completely ionized

(d) elements are in solid state

18. The dark lines in atomic absorption spectrum:

(a) correspond to wavelengths not absorbed

(b) correspond to wavelengths absorbed by the atoms ✅

(c) are random

(d) correspond to nuclear transitions

19. Each element has a unique spectrum because:

(a) all elements have same number of electrons

(b) each element has a unique arrangement of electrons ✅

(c) elements emit same radiation

(d) spectra depend on temperature only

20. Atomic spectra are often called the "fingerprints" of elements because:

(a) they are identical for all elements

(b) they uniquely identify each element ✅

(c) they depend on mass number

(d) they are random

21. Successive ionization energies of an element:

(a) decrease gradually

(b) remain constant

(c) increase as electrons are removed from inner shells ✅

(d) decrease then increase randomly

22. Large jumps in successive ionization energies indicate:

(a) removal of outermost electrons

(b) removal of electrons from inner shells closer to nucleus ✅

(c) electrons are neutral

(d) no pattern in shells

23. First ionization energy of elements generally:

(a) decreases across a period

(b) increases across a period ✅

(c) remains constant

(d) decreases then increases randomly

24. Down a group in the periodic table, ionization energy:

(a) increases

(b) decreases ✅

(c) remains constant

(d) fluctuates randomly

25. The reason noble gases have highest ionization energy in a period is:

(a) they are metals

(b) they have full valence shells and strong nuclear attraction ✅

(c) they are large atoms

(d) they have fewer protons

26. The Bohr model of atom required only one quantum number because:

(a) electrons move in three-dimensional space

(b) it described only the size and energy of the orbit ✅

(c) it included spin of electrons

(d) it explained magnetic properties

27. Schrödinger's model requires how many quantum numbers to describe an electron?

(a) 1

(b) 2

(c) 3 ✅

(d) 4

28. Principal quantum number (n) determines:

(a) shape of orbital

(b) orientation in space

(c) size and energy of orbital ✅

(d) spin of electron

29. Maximum number of electrons in a shell is given by:

(a) n^2

(b) $2n^2$ ✅

(c) n

(d) $4n^2$

30. Azimuthal quantum number (l) describes:

(a) orientation of orbital

(b) shape of orbital ✅

(c) spin of electron

(d) energy only

31. For n = 3, possible values of l are:

(a) 0,1

(b) 0,1,2 ✅

(c) 1,2,3

(d) 0,1,2,3

32. The letters s, p, d, f in quantum numbers represent:

(a) sizes of orbitals

(b) shapes of orbitals ✅

(c) energies of orbitals

(d) spin directions

33. Magnetic quantum number (m) indicates:

(a) shape of orbital

(b) energy of electron

(c) orientation of orbital in space ✅

(d) spin of electron

34. The number of orbitals in a p-subshell (l = 1) is:

(a) 1

(b) 2

(c) 3 ✅

(d) 5

35. Spin quantum number (s) is necessary because:

(a) it describes energy of orbital

(b) it differentiates between two electrons in the same orbital ✅

(c) it determines orientation in space

(d) it determines shape of orbital

36. An atomic orbital is defined as:

(a) a one-dimensional line around nucleus

(b) a two-dimensional plane around nucleus

(c) a three-dimensional region around nucleus with maximum probability of finding an electron ✅

(d) the path of an electron around the nucleus

37. The shape of s-orbital is:

(a) spherical ✅

(b) dumbbell-shaped

(c) double dumbbell

(d) complex

38. A p-orbital has:

(a) spherical symmetry

(b) two lobes along any axis ✅

(c) four lobes

(d) seven lobes

39. How many orientations does a d-subshell have?

(a) 3

(b) 5 ✅

(c) 7

(d) 1

40. The d z^2 orbital is characterized by:

(a) two lobes along z-axis with a doughnut in xy plane ✅

(b) four lobes along x and y axes

(c) spherical shape

(d) single lobe along z-axis

41. Number of orbitals in f-subshell is:

(a) 3

(b) 5

(c) 7 ✅

(d) 9

42. Electronic configuration is defined as:

(a) distribution of protons in nucleus

(b) distribution of neutrons in nucleus

(c) distribution of electrons among shells, subshells, and orbitals ✅

(d) arrangement of atoms in a molecule

43. Maximum number of electrons in M-shell (n = 3) is:

(a) 2

(b) 8

(c) 18 ✅

(d) 32

44. According to Aufbau principle, the order of filling of subshells is determined by:

(a) atomic number

(b) (n + l) value ✅

(c) number of neutrons

(d) energy of nucleus

45. If two subshells have same (n + l) value, which is filled first?

(a) subshell with higher n

(b) subshell with lower n ✅

(c) subshell with more electrons

(d) subshell with fewer electrons

46. The electronic configuration of Na atom is:

(a) $1s^2\ 2s^2\ 2p^6\ 3s^1$ ✅

(b) $1s^2\ 2s^2\ 2p^6\ 3p^1$

(c) $1s^2\ 2s^2\ 3s^2$

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

47. According to Pauli's exclusion principle:

(a) all electrons have same spin

(b) no two electrons in an atom can have same four quantum numbers ✅

(c) electrons occupy orbitals randomly

(d) two electrons in an orbital always have same spin

48. Maximum electrons in one orbital are:

(a) 1

(b) 2 ✅

(c) 3

(d) 4

49. An electron in an orbital is represented by:

(a) circle

(b) arrow ✅

(c) dot

(d) square

50. In box diagram of orbitals, boxes are arranged:

(a) in order of decreasing energy

(b) in order of increasing energy ✅

(c) randomly

(d) according to spin only

51. Hund's rule states that when electrons are to be placed in degenerate orbitals:

(a) they pair up in one orbital first

(b) they occupy separate orbitals with parallel spins first ✅

(c) they occupy orbitals randomly

(d) they occupy orbitals with opposite spins first

52. According to Hund's rule, the correct electronic configuration of three electrons in three p-orbitals is:

(a) ↑↓ , ↑ , ↑

(b) ↑ , ↑ , ↑ ✅

(c) ↑↓ , ↑↓ , ↑

(d) ↑ , ↑↓ , ↑

53. Degenerate orbitals are:

(a) orbitals of different energies

(b) orbitals of same energy ✅

(c) orbitals with opposite spins

(d) orbitals with same shape only

54. Valence electrons are:

(a) electrons in the innermost shell

(b) electrons in partially filled inner subshells

(c) electrons in the outermost shell that participate in chemical reactions ✅

(d) electrons in completely filled shells only

55. Group 1 elements in periodic table have outermost configuration of:

(a) ns^2

(b) ns^1 ✅

(c) np^1

(d) nd^1

56. The electronic configuration of group 13 elements is generally:

(a) $ns^2 np^1$ ✅

(b) $ns^2 np^2$

(c) $ns^1 np^2$

(d) $ns^2 nd^1$

57. Transition elements have their valence electrons in:

(a) s-block orbitals

(b) p-block orbitals

(c) d-block orbitals ✅

(d) f-block orbitals

58. Inner-transition elements belong to which block?

(a) s-block

(b) p-block

(c) d-block

(d) f-block ✅

59. The chemical similarity among elements of a group is due to:

(a) same number of neutrons

(b) same atomic mass

(c) same valence electron configuration ✅

(d) same number of shells

60. The f-block elements can hold maximum how many electrons in their outermost f-subshell?

(a) 2

(b) 6

(c) 10

(d) 14 ✅

61. Which subshell loses electrons first in d-block elements while forming positive ions?

(a) (n–1)d

(b) ns ✅

(c) np

(d) nf

62. In a P-type semiconductor, holes are:

(a) Negative charge carriers

(b) Positive charge carriers ✅

(c) Neutral particles

(d) Unpaired electrons

63. In the yellow-colored area of the periodic table, which subshell is being filled?

(a) ns

(b) np

(c) (n-1)d ✅

(d) (n-2)f

64. Which subshell is being filled in the light golden area of the periodic table?

(a) ns

(b) np

(c) (n-1)d

(d) (n-2)f ✅

65. When moving across a period, which subshell fills immediately after 4s?

(a) 3p

(b) 3d ✅

(c) 4p

(d) 4d

66. Positive ions are formed when atoms:

(a) Gain electrons

(b) Lose electrons ✅

(c) Gain protons

(d) Lose neutrons

67. The electronic configuration of Na^+ ion is:

(a) $1s^2\ 2s^2\ 2p^6$ ✅

(b) $1s^2\ 2s^2\ 2p^5$

(c) $1s^2\ 2s^2\ 2p^6\ 3s^1$

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

68. Negative ions are formed when atoms:

(a) Gain electrons ✅

(b) Lose electrons

(c) Gain protons

(d) Lose neutrons

69. A species with one or more unpaired electrons is called:

(a) Ion

(b) Molecule

(c) Free radical ✅

(d) Isotope

❖ Important Short Questions:

1. Define electron configuration

Ans: Electron configuration is the arrangement of electrons in the orbitals of an atom according to increasing energy. It helps predict chemical behavior and reactivity of elements.

Example: Oxygen (O): $1s^2\ 2s^2\ 2p^4$

2. State Pauli's Exclusion Principle

Ans: No two electrons in an atom can have the same set of four quantum numbers. An orbital can hold a maximum of two electrons with opposite spins.

Example: Helium (He): $1s^2$ ⟶ two electrons with opposite spins.

3. State Hund's Rule

Ans: Electrons occupy degenerate orbitals singly with parallel spins before pairing. This minimizes electron repulsion and increases stability.

Example: Nitrogen (N): $1s^2\ 2s^2\ 2p^3$ ⟶ each 2p orbital gets one electron.

4. What are valence electrons?

Ans: Valence electrons are electrons in the outermost shell. They determine chemical reactivity and the type of bonds an element can form.

Example: Sodium (Na): $1s^2\ 2s^2\ 2p^6\ 3s^1$ ⟶ Valence electron = 1

5. Explain the difference between positive and negative ions

Ans: Positive ions (cations) form when atoms lose electrons; negative ions (anions) form when atoms gain electrons. They achieve stable electron configurations similar to noble gases.

Example: Na ⟶ Na^+, Cl ⟶ Cl^-

6. What is a free radical? Give an example

Ans: A free radical is a species with one or more unpaired electrons. Free radicals are highly reactive.

Example: Chlorine atom: Cl· ⟶ has one unpaired electron

7. Write the electron configuration of Na^+ and explain its similarity with neon

Ans: Na^+ has 10 electrons: $1s^2\ 2s^2\ 2p^6$. Its electron configuration is identical to neon, making it stable and isoelectronic with Ne.

Example: $Na^+ \longrightarrow 1s^2\ 2s^2\ 2p^6$

8. Write the electron configuration of S^{2-} and explain its similarity with argon

Ans: S^{2-} has 18 electrons: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$. This is the same as argon, making S^{2-} isoelectronic with Ar and stable.

Example: $S^{2-} \longrightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

9. Explain why d-block elements lose 4s electrons first when forming ions

Ans: In d-block elements, 4s electrons are higher in energy than 3d electrons when forming ions. Therefore, 4s electrons are removed first to form stable cations.

Example: Ti atom: $[Ar]\ 3d^2\ 4s^2 \longrightarrow Ti^{2+}$: $[Ar]\ 3d^2$

10. Define a semiconductor

Ans: A semiconductor is a material that can conduct electricity under certain conditions. Its conductivity is between that of conductors and insulators, making it useful in electronic devices.

Example: Silicon (Si) and Germanium (Ge) used in computers and smartphones

11. Explain P-type semiconductor formation with an example

Ans: P-type semiconductors are formed by adding trivalent impurity atoms (like Al) to a pure semiconductor (Si). These create "holes" in the crystal lattice which act as positive charge carriers.

Example: Si doped with Al ⟶ Si-Al lattice with holes

12. Explain N-type semiconductor formation with an example

Ans: N-type semiconductors are formed by adding pentavalent impurity atoms (like P) to a pure semiconductor (Si). The extra electron from the impurity becomes a negative charge carrier.

Example: Si doped with P ⟶ Si-P lattice with free electrons

13. What is meant by covalent bonding in semiconductors?

Ans: Covalent bonding in semiconductors occurs when atoms share valence electrons to achieve stability. Each Si atom shares electrons with four neighboring Si atoms forming a strong lattice.

Example: Si-Si bonds in crystalline silicon

14. Name elements that commonly act as semiconductors

Ans: Common semiconductors include silicon (Si), germanium (Ge), and arsenic (As). These have 4 valence electrons, allowing covalent bonding and doping to control conductivity.

15. Classify elements of the periodic table into metals, non-metals, transition, representative, s, p, d, f blocks

Ans: Elements are classified as metals (left & center), non-metals (right), transition elements (d-block), representative elements (s & p-block), and inner transition elements (f-block).

Example: Na ⟶ s-block, Fe ⟶ d-block, U ⟶ f-block

16. Explain the building-up order (Aufbau principle) of electron configuration

Ans: Electrons fill orbitals in order of increasing energy. Lower energy orbitals are filled first before higher ones. This is known as the Aufbau principle.

Example: O ⟶ $1s^2$ $2s^2$ $2p^4$

17. How does electronic configuration determine the position of an element in the periodic table?

Ans: The outer electron configuration (valence electrons) determines the group and block of an element. Period number corresponds to the highest principal quantum number (n).

Example: Na ⟶ $1s^2\ 2s^2\ 2p^6\ 3s^1$ ⟶ Period 3, Group 1

18. What is the difference between a period and a group in the periodic table?

Ans: A period is a horizontal row showing elements with increasing atomic number. A group is a vertical column containing elements with similar valence electron configuration.

Example: Na and K are in the same group (Group 1), but different periods (3 and 4)

19. Explain why elements in the same group have similar chemical properties

Ans: Elements in the same group have the same number of valence electrons, leading to similar bonding and chemical behavior.

Example: Li, Na, K ⟶ all react vigorously with water

20. Why are semiconductors important in electronic devices?

Ans: Semiconductors can control the flow of electricity and act as switches or amplifiers. This makes them essential for devices like smartphones, computers, and solar cells.

Example: Silicon chips in computers

❖ Important Long Questions:

🌟 **Q1. Explain the concept of atomic number, mass number, and nucleon number. How are they related? Also explain how Moseley proved the importance of atomic number.**

❖ **Answer:**

1. Atomic Number (Z):

Atomic number is the number of protons present in the nucleus of an atom. In a neutral atom, it is equal to the number of electrons. It determines the identity of an element.

2. Mass Number (A) / Nucleon Number:

Mass number (also called nucleon number) is the total number of protons and neutrons present in the nucleus. It represents the total mass of the nucleus.

3. Relationship between Atomic Number and Mass Number:

The relation is:

$A = Z + N$

where:

- A = Mass number
- Z = Number of protons
- N = Number of neutrons

✅ **Example:** For chlorine ⟶ Z = 17, A = 35 ⟶ N = 18

4. Moseley's Contribution (1913):

Moseley bombarded elements with high-energy electrons and studied the X-rays produced. He found that the square root of frequency of X-rays is directly proportional to atomic number:

$\sqrt{\nu} \propto Z$

He concluded that atomic number is a fundamental property of an element.

5. Importance of Moseley's Work:

- Elements are arranged according to atomic number
- It corrected defects of Mendeleev's periodic table
- It proved atomic number defines element identity

⭐ **Summary:**

Atomic number (Z) represents protons, while mass number (A) represents protons + neutrons. Their relation is A = Z + N. Moseley proved that atomic number is the true basis of classification of elements, making it fundamental for the modern periodic table.

🌟 Q2. Effect of Electric Field on Electrons, Protons, and Neutrons

❖ **Answer:**

Electrons:

Electrons are negatively charged particles. When placed in an electric field, they experience a force opposite to the direction of the field. Because their mass is very small, they get highly deflected.

Example: In a cathode ray tube, electrons bend sharply toward the positive plate.

Protons:

Protons are positively charged particles. In an electric field, they experience a force along the direction of the field. Due to their larger mass compared to electrons, their deflection is much smaller.

Example: In a mass spectrometer, protons bend slightly toward the negative plate.

Neutrons:

Neutrons are neutral particles with no charge. They experience no force in an electric field, so they do not deflect at all.

Comparison:

Deflection depends on the charge-to-mass ratio (q/m). Electrons have a high q/m ratio, so they deflect the most. Protons have a lower q/m ratio, so their deflection is small. Neutrons have zero charge, so they are not affected by the field.

✦ Summary:

Electrons bend opposite to the field due to negative charge, protons bend slightly along the field due to positive charge, and neutrons are unaffected because they have no charge. The amount of deflection depends on the particle's charge-to-mass ratio.

🌟 Q.3 What are quantum numbers? Explain in detail principal (n), azimuthal (l), magnetic (m), and spin (s) quantum numbers with their significance.

❖ **Answer:**

Quantum numbers are numerical values assigned to electrons in an atom to describe their energy, shape, orientation, and spin. Each electron in an atom has a unique set of quantum numbers. There are four quantum numbers: principal (n), azimuthal (l), magnetic (m), and spin (s).

1. Principal Quantum Number (n)

- **Symbol:** n
- Represents the main energy level or shell of an electron.
- Determines the size and energy of the orbital.
- **Values:** n = 1, 2, 3, ...
- Higher n means the electron is farther from the nucleus and has higher energy.

Significance: Determines which electron shell the electron belongs to and its energy.

2. Azimuthal (Angular Momentum) Quantum Number (l)

- **Symbol:** l
- Determines the shape of the orbital within a shell.
- **Values:** l = 0 to (n-1)

Orbital types:

- l = 0 ⟶ s-orbital (spherical)
- l = 1 ⟶ p-orbital (dumbbell-shaped)
- l = 2 ⟶ d-orbital (clover-shaped)
- l = 3 ⟶ f-orbital (complex shape)

Significance: Determines the subshell and the shape of the electron cloud.

3. Magnetic Quantum Number (m)

- **Symbol:** m
- Determines the orientation of the orbital in space.
- **Values:** m = -l to +l (including 0)
- **Example:** For p-orbital (l = 1), m = -1, 0, +1, so three p-orbitals are oriented differently along x, y, z axes.

Significance: Explains how orbitals are aligned in 3D space.

4. Spin Quantum Number (s)

- **Symbol:** s
- Describes the spin of the electron.

- **Values:** s = +1/2 or -1/2
- Electrons in the same orbital must have opposite spins (Pauli Exclusion Principle).

Significance: Explains the magnetic properties of electrons and ensures that no two electrons in an atom have the same set of four quantum numbers.

✦ **Summary:**

Quantum numbers describe where an electron is, its energy, shape, orientation, and spin.

- n ⟶ size & energy level
- l ⟶ orbital shape
- m ⟶ orientation
- s ⟶ spin

🌟 **Q.4 Explain the shapes of atomic orbitals (s, p, d) with reference to azimuthal and magnetic quantum numbers. Also discuss their orientations in space.**

❖ **Answer:**

Atomic orbitals are regions around the nucleus where electrons are most likely to be found. The shape and orientation of orbitals are determined by the azimuthal quantum number (l) and magnetic quantum number (m). The main types of orbitals are s, p, and d.

1. s-Orbitals

- Azimuthal quantum number (l): 0
- **Shape:** Spherical; electron density is uniform in all directions.
- Magnetic quantum number (m): 0 ⟶ only one orientation.
- **Orientation:** Since it is spherical, it has no directional preference.

Example: 1s, 2s orbitals in hydrogen or lithium.

2. p-Orbitals

- Azimuthal quantum number (l): 1
- **Shape:** Dumbbell-shaped with two lobes on opposite sides of the nucleus.
- Magnetic quantum number (m): -1, 0, +1 ⟶ three orientations along x, y, and z axes.
- **Orientation:** px along x-axis, py along y-axis, pz along z-axis.

Example: 2p orbitals in carbon or nitrogen atoms.

3. d-Orbitals

- Azimuthal quantum number (l): 2
- **Shape:** Mostly cloverleaf-shaped with four lobes; dz^2 orbital has a dumbbell with a doughnut ring.
- Magnetic quantum number (m): -2, -1, 0, +1, +2 ⟶ five orientations.
- **Orientation:** dxy, dyz, dxz lie between axes; dx^2-y^2 lies along x and y axes; dz^2 along z-axis.

Example: 3d orbitals in iron or copper atoms.

✦ **Summary**

- **s-orbitals:** Spherical, 1 orientation.
- **p-orbitals:** Dumbbell-shaped, 3 orientations.
- **d-orbitals:** Cloverleaf or doughnut-shaped, 5 orientations.
- **Key point:** l determines the shape of the orbital, while m determines its spatial orientation.

🌟 **Q.5 Explain the electronic configuration of atoms. Discuss the Aufbau principle, Pauli Exclusion Principle, and Hund's Rule with examples.**

❖ **Answer:**

Electronic configuration describes the arrangement of electrons in the orbitals of an atom. It helps to understand the chemical behavior of elements, their reactivity, and position in the periodic table. Electrons occupy orbitals in a specific order based on energy levels and sublevels.

1. Aufbau Principle:

The Aufbau principle states that electrons fill atomic orbitals in order of increasing energy, starting from the lowest energy level. For example, in sodium (Na, atomic number 11), the electron configuration is:

$1s^2\ 2s^2\ 2p^6\ 3s^1$

This shows that electrons fill 1s first, then 2s, 2p, and finally 3s.

2. Pauli Exclusion Principle:

The Pauli Exclusion Principle states that no two electrons in an atom can have the same set of four quantum numbers. In simpler terms, each orbital can hold a maximum of two electrons, and they must have opposite spins.

Example: In oxygen (O, atomic number 8), the 2p orbital has three orbitals. Each orbital holds one electron first (with same spin) before pairing occurs: $2p^4$ →⬆️⬇️⬆️⬆️

3. Hund's Rule:

Hund's Rule states that electrons occupy degenerate orbitals (orbitals with the same energy) singly first with parallel spins before pairing. This minimizes electron-electron repulsion and stabilizes the atom.

Example: In nitrogen (N, atomic number 7), the $2p^3$ electrons occupy three 2p orbitals as: ⬆️⬆️⬆️

✦ **Summary:**

- Electronic configuration shows electron distribution in orbitals.
- **Aufbau Principle:** Fill orbitals from lowest to highest energy.
- **Pauli Exclusion Principle:** Each orbital holds a maximum of 2 electrons with opposite spins.
- **Hund's Rule:** Degenerate orbitals are singly filled first with parallel spins.
- These rules together explain the stability, chemical properties, and periodic placement of elements.

🌟 **Q.6 Explain the concept of ionization energy.**

❖ **Answer:**

Ionization energy (IE) is the minimum amount of energy required to remove an electron from a gaseous atom or ion in its ground state to form a positive ion. It shows how strongly an atom holds its electrons.

1. First Ionization Energy (I1):

Energy needed to remove the first electron from a neutral atom.

Example:

$Na\ (g) \longrightarrow Na^{+}\ (g) + e^{-}$

I1 = 496 kJ/mol

2. Second Ionization Energy (I2):

Energy required to remove the second electron from a singly charged ion. It is higher than I1 because the positive ion holds the remaining electrons more strongly.

Example:

$Na^{+}\ (g) \longrightarrow Na^{2+}\ (g) + e^{-}$

I2 = 4562 kJ/mol

Factors Affecting Ionization Energy:

- **Atomic size:** Smaller atoms ⟶ higher IE

- **Nuclear charge:** More protons → higher IE
- **Electron shielding:** More inner electrons → lower IE
- **Subshell configuration:** Half-filled or fully filled subshells are more stable → higher IE

Periodic Trends:

- **Across a period:** IE increases → nuclear charge increases
- **Down a group:** IE decreases → electrons are farther from the nucleus

✦ **Summary:**

Ionization energy tells us how easily an atom can lose an electron. It is important for understanding element reactivity and chemical behavior.

🌟 **Q.7 What are semiconductors? Explain the formation of P-type and N-type semiconductors with suitable examples.**

❖ **Answer:**

Semiconductors:

Semiconductors are materials whose electrical conductivity is between conductors and insulators. They conduct electricity under certain conditions and are used in electronic devices. Examples include silicon (Si), germanium (Ge), and arsenic (As).

P-type Semiconductor:

- Formed by adding trivalent impurities (3 valence electrons) like Al to a pure semiconductor (Si).
- Each impurity atom replaces a Si atom but has one less electron, creating a hole that acts as a positive charge carrier.
- **Example:** Si doped with Al ⟶ P-type semiconductor.

N-type Semiconductor:

- Formed by adding pentavalent impurities (5 valence electrons) like P to a pure semiconductor (Si).
- Each impurity atom replaces a Si atom but has one extra electron, which is free to move and acts as a negative charge carrier.
- Example: Si doped with P ⟶ N-type semiconductor.

Importance:

- **P-type:** carries positive charges (holes)
- **N-type:** carries negative charges (electrons)

Both types are crucial in electronic devices like smartphones, computers, and solar cells.

📘 Exercise

Q.1 Four choices are given for each question. Select the correct choice.

I. The quantum number 'm' of a free gaseous atom is associated with:

a) the effective volume of the orbital

b) the shape of the orbital

✅ c) the spatial orientation of the orbital

d) the energy of the orbital in the absence of a magnetic field

II. When 3d subshell is completely filled, the next entering electron goes into:

a) 4f

b) 4s

✅ c) 4p

d) 4d

III. Quantum number values for 2p orbitals are:

✅ a) n = 2, l = 1

b) n = 1, l = 2

c) n = 1, l = 0

d) n = 2, l = 0

IV. An electron having the set of values n = 4, l = 0, m = 0 and s = +1/2 lies in:

a) 2s

b) 3s

✅ c) 4s

d) 4p

V. The quantum number values (n, l, m) for the fourth electron of ^{9}Be atom are:

a) 1, 0, 0

✅ b) 2, 0, 0

c) 2, 1, 0

d) 1, 1, 1

VI. The correct order of first ionization energies is:

a) F > He > Mg > N > O

✅ b) He > F > N > O > Mg

c) Hc > O > F > N > Mg

d) N > F > He > O > Mg

VII. A p orbital has a characteristic shape with how many lobes?

a) 1

✅ b) 2

c) 3

d) 4

VIII. The three p orbitals in a given energy level are oriented:

a) Along the same axis

b) At 45° to each other

✅ c) Mutually perpendicular to each other along the x, y, and z axes

d) In a complex tetrahedral arrangement

IX. How many d orbitals are there in a given energy level?

a) 1

b) 3

✅ c) 5

d) 7

X. What information does the principal quantum number (n) give us about orbitals:

✅ a) Size

b) Shape

c) Size and shape

d) Spin

XI. How many unpaired electrons are present in an atom of cobalt?

a) Two

✅ b) Three

c) Four

d) Five

0.2 Attempt the following short-answer questions:

a) There are three orientations of p-orbital due to three values of magnetic quantum number. Justify it.

Ans: A p-orbital has angular momentum quantum number l = 1, so the magnetic quantum number m = -1, 0, +1. Each m value represents a different orientation of the orbital in space along x, y, z axes. Hence, there are three mutually perpendicular p-orbitals in a given energy level.

b) I_3 of Mg is much bigger than its I_2. Justify.

Ans: I_1 and I_2 remove electrons from the outer shell. I_3 removes an electron from the stable noble gas core ($1s^2\ 2s^2\ 2p^6$). Removing a core electron requires much higher energy, so $I_3 >> I_2$.

c) Among the elements Li, K, Ca, S, and Kr, which one has the lowest first ionization energy? Which has the highest first ionization energy?

Answer:

- **Lowest:** K (Potassium), because the outermost electron is far from the nucleus and easily removed.
- **Highest:** Kr (Krypton), as it is a noble gas with a stable, tightly bound electron configuration.

d) Consider the electronic configuration of the potassium atom (atomic number 19):

(i) Write the full electronic configuration using s, p, d, f notation.

Answer: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

(ii) Explain why the 4s subshell is filled before the 3d subshell in potassium, even though the principal quantum number of the 3d subshell is lower.

Answer: The 4s orbital has lower energy than 3d for potassium. According to the Aufbau principle, electrons occupy orbitals of lowest energy first, so 4s fills before 3d.

e) (i) An atom of element X has an atomic number of 17 and a mass number of 35. Determine the number of protons, neutrons, and electrons in this atom.

Answer:

- Atomic number = 17 ⟶ Protons = 17
- Mass number = 35 ⟶ Neutrons = 35 – 17 = 18
- Neutral atom ⟶ Electrons = 17

Example: This is a chlorine atom (Cl), with 17 protons, 18 neutrons, and 17 electrons.

e) (ii) If this element forms an ion with a charge of -1, how many protons, neutrons, and electrons will be present in the ion?

Answer:

- Protons = 17 (unchanged)
- Neutrons = 18 (unchanged)
- Electrons = 17 + 1 = 18

Example: Cl^- ion has one extra electron compared to neutral chlorine, giving a total of 18 electrons.

Question f(i)

Q: In the ground state of mercury ^{80}Hg, how many electrons occupy atomic orbitals with n = 3?

Answer:

- n = 3 includes $3s^2$, $3p^6$, and $3d^{10}$ ⟶ total electrons = 18

Example: Electrons in the third shell of mercury occupy all orbitals of 3s, 3p, and 3d completely.

Question f(ii)

Q: How many electrons occupy 4d atomic orbitals in mercury?

Answer:

- 4d subshell is fully filled ⟶ 10 electrons

Example: Mercury’s 4d orbitals contain 10 electrons because all five 4d orbitals have 2 electrons each.

Question f(iii)

Q: How many electrons occupy 4pz atomic orbital in mercury?

Answer:

- 4p subshell has 6 electrons in total ⟶ 2 electrons in each orbital ⟶ 4pz = 2 electrons

Example: The 4pz orbital of mercury is fully filled like 4px and 4py orbitals.

Question f(iv)

Q: How many electrons in the valence shell have spin "up" (s = +1/2)?

Answer:

- Mercury valence shell (n = 6) ⟶ $6s^2$
- One electron in 6s has spin "up" ⟶ 1 electron

Example: In the 6s subshell, one electron spins up (+1/2) and the other spins down (−1/2).

g) The successive ionization energies for an unknown element are:

I_1 = 896 kJ/mol

I_2 = 1752 kJ/mol

I_3 = 14,807 kJ/mol

I_4 = 17,948 kJ/mol

To which family in the periodic table does the unknown element most likely belong?

Answer:

There is a large jump between I_2 and I_3.

- This indicates that the element has 2 valence electrons, because the first two electrons are removed easily (I_1 and I_2), and removing the third requires much more energy.
- Elements with 2 valence electrons belong to the alkaline earth metals (Group 2).

Example: Magnesium (Mg) and Calcium (Ca) show similar ionization energy patterns with 2 easily removable electrons.

i) Consider the following ionization energies for aluminium:

$Al(g) \longrightarrow Al^{+}(g) + e^{-}$ I_1 = 580 kJ/mol

$Al^{+}(g) \longrightarrow Al^{2+}(g) + e^{-}$ I_2 = 1815 kJ/mol

$Al^{2+}(g) \longrightarrow Al^{3+}(g) + e^{-}$ I_3 = 2740 kJ/mol

$Al^{3+}(g) \longrightarrow Al^{4+}(g) + e^{-}$ I_4 = 11,600 kJ/mol

(i) Account for the trend in the values of the ionization energies.

Answer:

Ionization energies increase from I_1 to I_4 because after each electron removal, the positive charge on the ion increases. This increases the attraction between the nucleus and remaining electrons. As a result, it becomes progressively more difficult to remove the next electron, so ionization energy increases.

(ii) Explain the large increase from I_3 to I_4.

Answer:

The large jump occurs because after removing three electrons, aluminium attains a stable noble gas configuration (Ne). The fourth electron is removed from an inner shell (core electron), which is much closer to the nucleus. Therefore, a very high energy is required, causing a sharp increase.

(iii) List the four aluminium species in order of increasing size and explain your ordering.

Answer:

Order: $Al^{3+} < Al^{2+} < Al^{+} < Al$

As positive charge increases, the number of electrons decreases while nuclear charge remains the same. This results in stronger attraction and smaller size. Hence, Al^{3+} is smallest and neutral Al atom is largest.

h) (i) State the general order of filling orbitals up to the 4p subshell.

(ii) Explain why the 4s subshell is filled before the 3d subshell, according to the Aufbau principle.

(i) General order of filling orbitals up to 4p

Answer:

The general order of filling orbitals according to the Aufbau principle is:

1s ⟶ 2s ⟶ 2p ⟶ 3s ⟶ 3p ⟶ 4s ⟶ 3d ⟶ 4p

This order is based on increasing energy of subshells using the (n + l) rule. Lower energy orbitals are filled first before higher energy ones.

(ii) Why 4s is filled before 3d

Answer:

Although 3d has lower principal quantum number, the 4s subshell has lower energy than 3d. According to the Aufbau principle, electrons fill orbitals of lowest energy first. Therefore, 4s is filled before 3d due to its lower (n + l) value.

I) Draw the orbital box diagram for the valence electrons of a phosphorus atom (atomic number 15), ensuring that your diagram adheres to Hund's rule and the Pauli Exclusion Principle.

Answer:

Electronic configuration of phosphorus:

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

Valence shell = $3s^2\ 3p^3$

Orbital box diagram:

3s ⟶ [⬆️⬇️]

3p ⟶ [⬆️] [⬆️] [⬆️]

- Each p orbital gets one electron first (Hund's Rule).
- All electrons have parallel spins before pairing.
- No two electrons in the same orbital have the same spin (Pauli Principle).

Example: Nitrogen (Z = 7) also shows similar half-filled p orbitals ($2p^3$).

DESCRIPTIVE QUESTIONS

🌟 Q.3 What are quantum numbers? Describe briefly principal and spin quantum numbers.

❖ **Answer:**

Quantum numbers are a set of numbers that describe the position, energy, shape, and orientation of an electron in an atom. There are four quantum numbers: principal (n), azimuthal (l), magnetic (m), and spin (s). These help in completely describing an electron in an orbital.

Principal Quantum Number (n):

The principal quantum number (n) describes the energy level and size of an orbital. It can have values 1, 2, 3, 4, … (K, L, M, N shells). As the value of n increases, the distance of the electron from the nucleus and its energy also increase. The maximum number of electrons in a shell is given by $2n^2$.

Spin Quantum Number (s):

The spin quantum number (s) describes the spin of an electron on its own axis. It has only two possible values: +1/2 (spin up) and −1/2 (spin down). This shows that an orbital can have a maximum of two electrons with opposite spins.

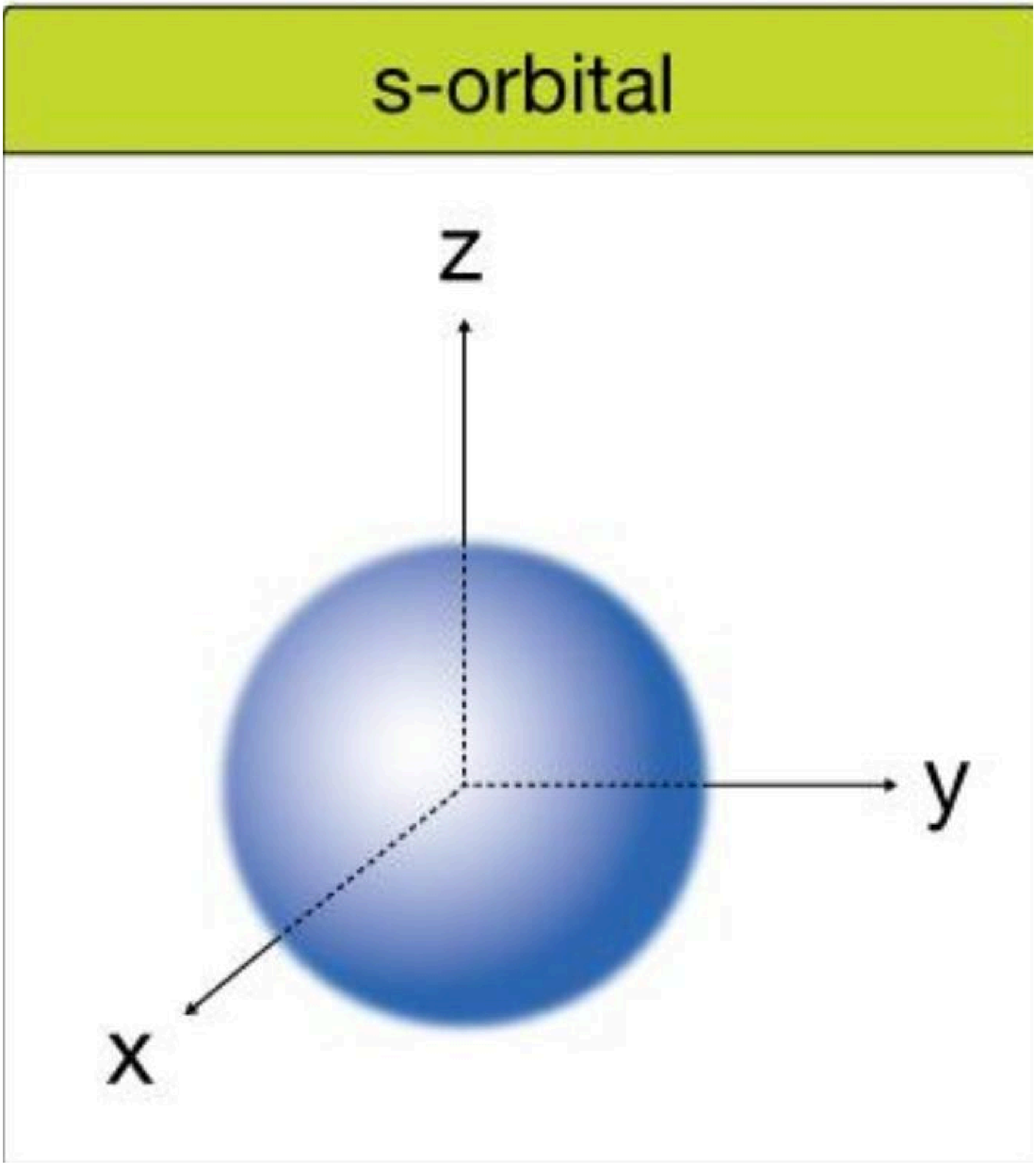
Example: In helium (He), two electrons are present in 1s orbital with opposite spins (+1/2 and −1/2).

🌟 Q.4 Draw the shapes of s, p and d-orbitals. Justify using azimuthal and magnetic quantum numbers.

❖ Answer:

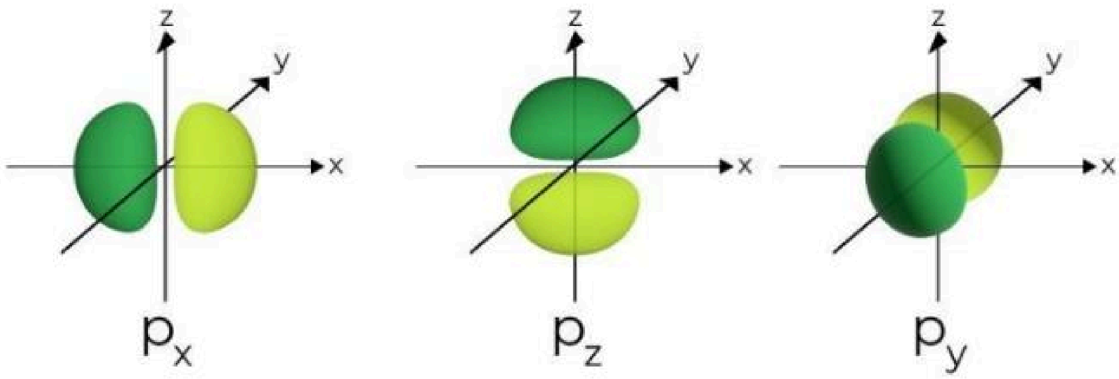
The shapes and orientations of atomic orbitals are determined by quantum numbers, especially the azimuthal quantum number (l) and the magnetic quantum number (m).

1. s-orbital (l = 0)



- For s-orbital, l = 0
- Therefore, m = 0 ⟶ only one orbital
- Shape is spherical, with uniform electron density in all directions

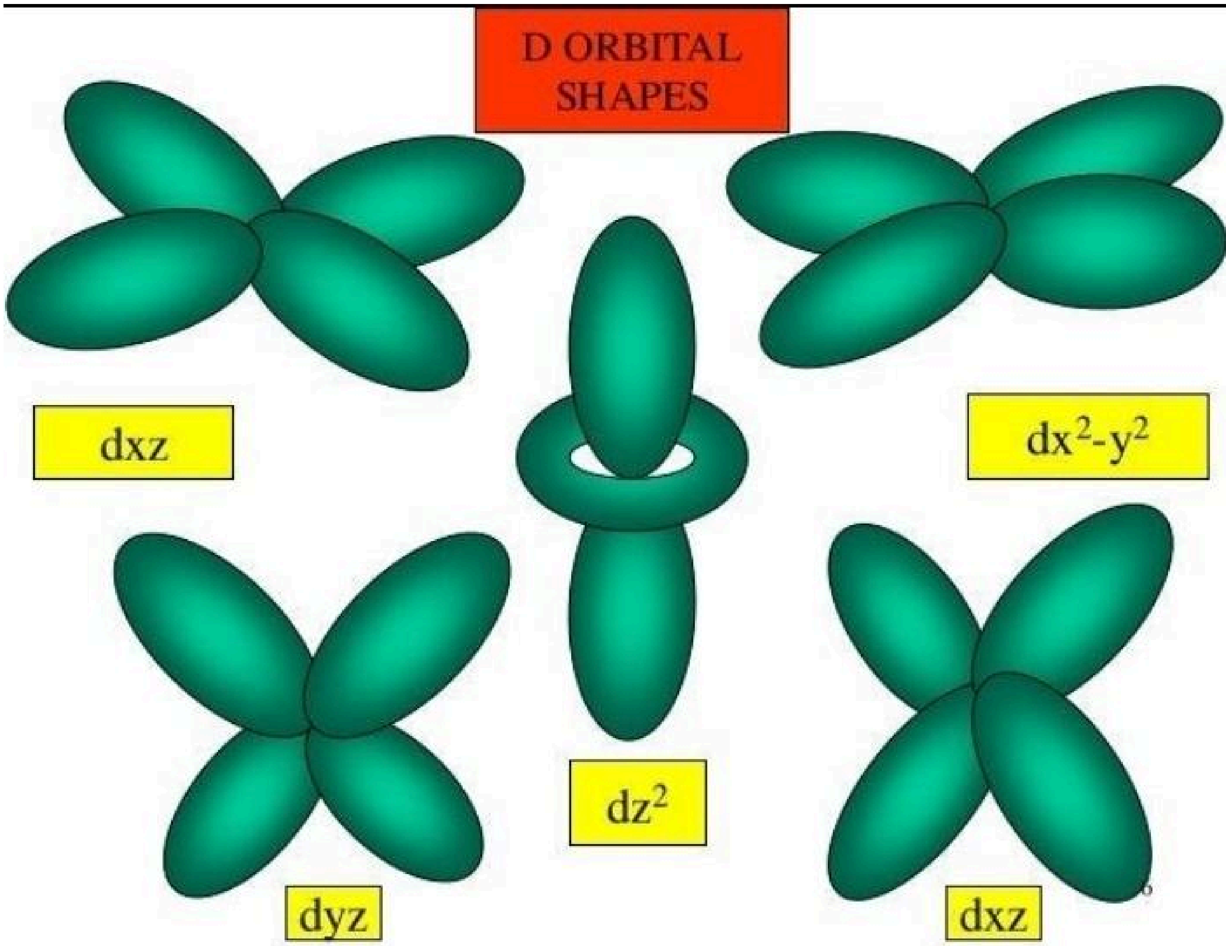
2. p-orbitals (l = 1)



Shape of p-orbital

- For p-orbitals, l = 1
- Therefore, m = -1, 0, +1 ⟶ three orbitals
- These are px, py, pz
- Shape is dumbbell, oriented along x, y, z axes (mutually perpendicular)

3. d-orbitals (l = 2)



- For d-orbitals, l = 2
- Therefore, m = -2, -1, 0, +1, +2 ⟶ five orbitals
- Shapes are complex, mostly cloverleaf (four lobes)
- The dz^2 orbital has a dumbbell with a doughnut ring

Justification:

- Azimuthal quantum number (l) determines the shape of the orbital
- Magnetic quantum number (m) determines the orientation and number of orbitals

Conclusion:

Thus, the shapes and spatial orientations of s, p, and d orbitals are fully explained by the values of l and m quantum numbers.

🌟 **Q.5 What do you mean by successive ionization energies? How is the electronic shell structure of magnesium (Mg) derived from successive ionization energies?**

❖ **Answer:**

Successive ionization energies are the energies required to remove electrons one by one from an isolated gaseous atom. Each successive ionization energy is higher than the previous one because it becomes more difficult to remove an electron from a positively charged ion.

Example (general):

$M (g) \longrightarrow M^+ (g) + e^- \longrightarrow I_1$

$M^+ (g) \longrightarrow M^{2+} (g) + e^- \longrightarrow I_2$

Electronic Shell Structure of Magnesium (Mg):

Magnesium (Mg) has atomic number 12, so it has 12 electrons. When we measure its successive ionization energies, we observe the following trend:

- The first two ionization energies (I_1 and I_2) are relatively low. This shows that the first two electrons are removed easily from the outermost shell (3s orbital).

- A very large jump occurs at the third ionization energy (I_3). This indicates that the third electron is removed from an inner shell, which is closer to the nucleus and more strongly held.
- **After that**, the ionization energies increase gradually, with another big jump when removing electrons from the innermost shell.

Explanation:

- The large jump after removing two electrons proves that Mg has two electrons in its outermost shell.
- These two electrons belong to the third shell ($3s^2$).
- The remaining electrons are in inner shells (2nd and 1st shells).

Thus, the electronic configuration of Mg is:

$1s^2\ 2s^2\ 2p^6\ 3s^2$

This shows the shell structure as:

K = 2, L = 8, M = 2

✦ Summary:

Successive ionization energies help us identify the arrangement of electrons in different shells. In magnesium, the large jump after the second ionization energy confirms that it has two valence electrons in the outermost shell, proving its electronic configuration and shell structure.