

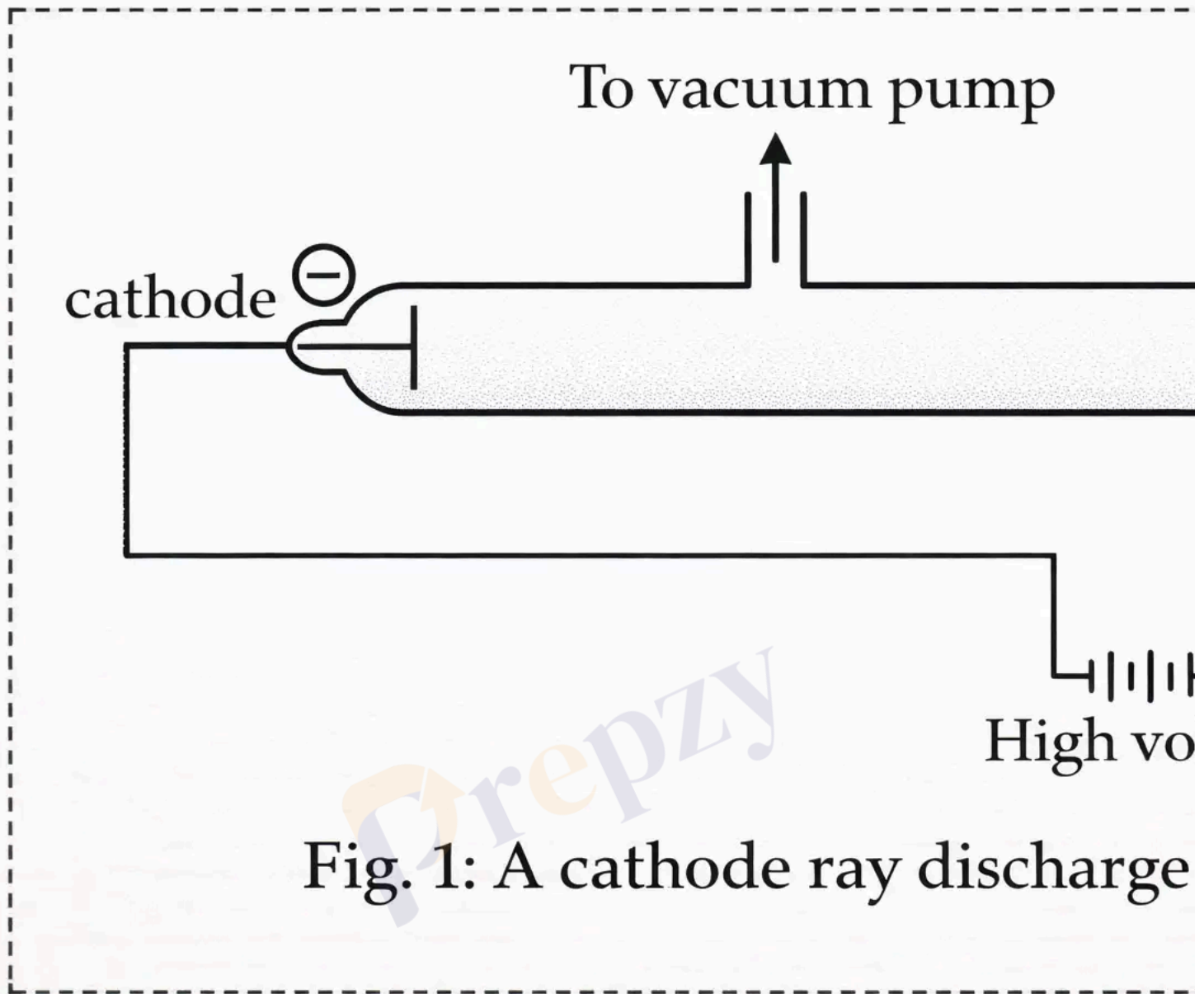
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## Structure of Atom

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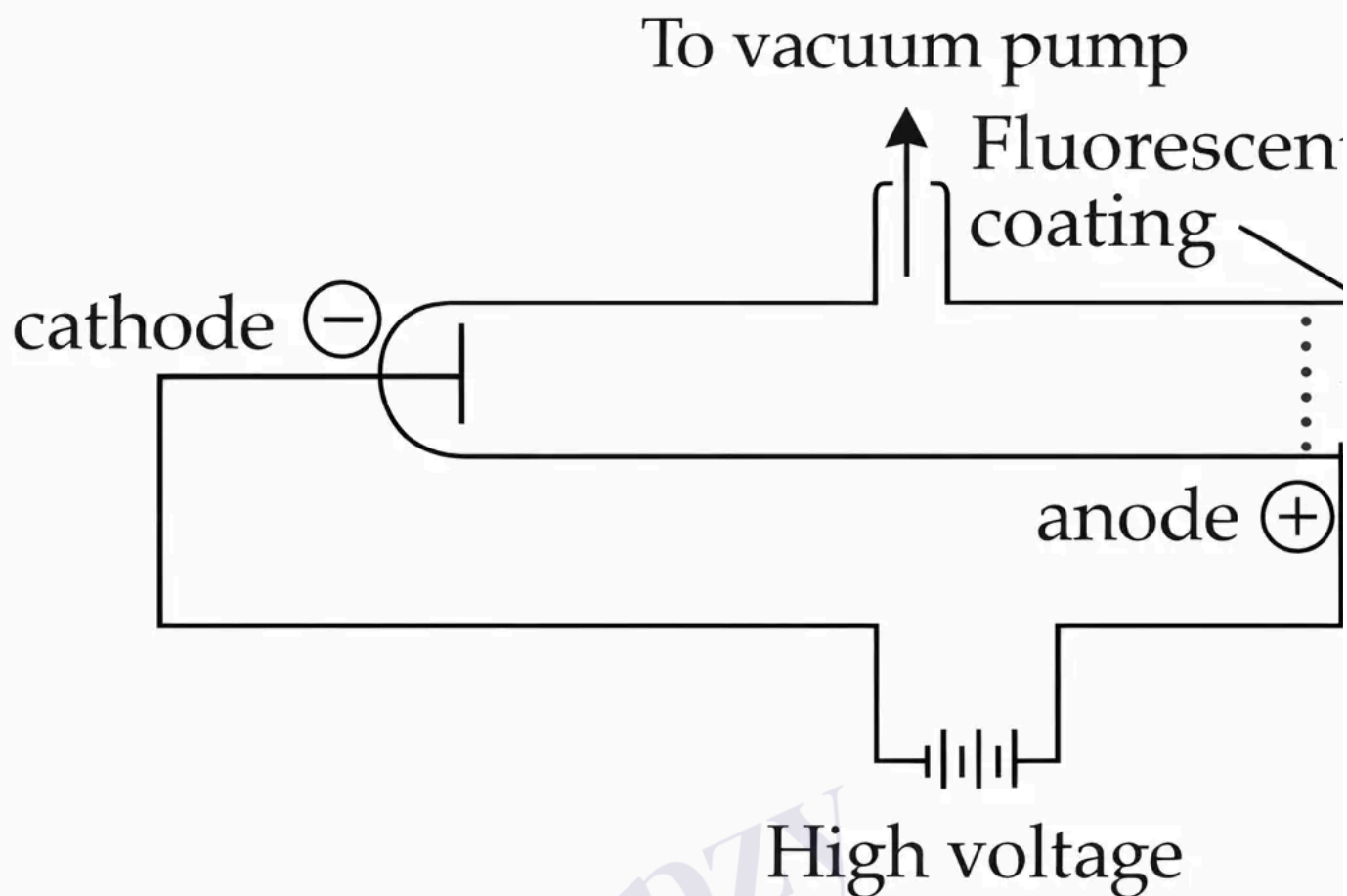
### Discovery of Electron

In the mid-1850s, scientists like Faraday studied electrical discharge in partially evacuated tubes called cathode ray discharge tubes. These electrodes sealed inside. When a high voltage is applied, electrons are emitted from the cathode and travel towards the anode, forming cathode rays. The behavior of cathode rays is observed by their effect on fluorescent materials inside the tube, which glow when struck by the rays. The behavior of cathode rays in electricity is due to negatively charged particles called electrons.



**Fig. 1: A cathode ray discharge**

A cathode ray tube consists of a cathode (negative electrode), an anode (positive electrode), a vacuum pump to remove air, and a high voltage source. Cathode rays are produced by the travel of electrons in straight lines towards the anode.



**Fig. 2: A cathode ray discharge tube with a fluorescent coating on the anode and zinc sulphide coating for the cathode.**

When cathode rays strike the fluorescent coating inside the tube, the coating glows, demonstrating the presence and path of electrons. The properties of cathode rays do not depend on the material of the electrodes or the gas inside the tube.

### Atomic Models

Several models have been proposed to explain atomic structure:

- **Thomson Model:** Proposed in 1898, it describes the atom as a sphere of positive charge with electrons embedded like plums in a pudding.
- **Rutherford Nuclear Model:** Based on gold foil experiments, it concluded that most of the atom is empty space with a small, dense, positively charged nucleus. Electrons are moving in circular orbits.

Rutherford's model explained the deflection of alpha particles but could not explain atomic stability or electronic structure.

### Fundamental Particles of an Atom

The atom consists of three fundamental particles:

- **Electron ( $e^-$ ):** Negatively charged particle with charge  $-1.6022 \times 10^{-19}$  C and very small mass.
- **Proton ( $p^+$ ):** Positively charged particle with charge  $+1.6022 \times 10^{-19}$  C and mass approximately 1 atomic mass unit (u).
- **Neutron ( $n^0$ ):** Neutral particle with no charge and mass approximately 1 u.

**Table 2.1 Properties of Fundamental Particles**

Name	Symbol	Absolute charge/C	Relative charge	Mass/kg
Electron	e	$-1.602176 \times 10^{-19}$	-1	$9.109382 \times 10^{-31}$
Proton	p	$+1.602176 \times 10^{-19}$	+1	$1.676216 \times 10^{-27}$
Neutron	n	0	0	$1.674927 \times 10^{-27}$

Electrons and the nucleus are held together by electrostatic forces of attraction.

### Atomic Number and Mass Number

**Atomic Number (Z):** Number of protons in the nucleus, equal to the number of electrons in a neutral atom.

**Mass Number (A):** Total number of protons and neutrons in the nucleus.

**Isotopes:** Atoms with the same atomic number but different mass numbers.

**Isobars:** Atoms with the same mass number but different atomic numbers.

### Dual Nature of Electromagnetic Radiation

Electromagnetic radiation exhibits both wave and particle nature.

- **Wave Nature:** Described by wavelength ( $\lambda$ ), frequency ( $\nu$ ), and velocity ( $c$ ). Electromagnetic waves do not require a medium and can travel through a vacuum.
- **Particle Nature:** Explained by the photoelectric effect where light ejects electrons from metal surfaces, indicating light behaves as particles (photons).

Planck's quantum theory states that energy is emitted or absorbed in discrete packets called quanta, with energy  $E = h\nu$ , where  $h$  is Planck's constant.

### Atomic Spectra

Atoms emit or absorb light at specific wavelengths, producing line spectra unique to each element. The hydrogen atom's spectral lines are given by the Rydberg formula:

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

where  $n_1$  and  $n_2$  are integers with  $n_2 > n_1$ .

## Bohr Model of Atom

Bohr proposed that electrons revolve in fixed circular orbits with quantized energies. Key postulates include:

- Electrons move in definite orbits without radiating energy.
- Each orbit corresponds to a specific energy level.
- Electrons can jump between orbits by absorbing or emitting energy.
- Angular momentum is quantized:  $m_e v r = n(h/2\pi)$ , where  $n = 1, 2, 3, \dots$

Energy of the  $n$ th orbit:  $E_n = -R_H / n^2$ , where  $R_H$  is the Rydberg constant.

Bohr's model explains hydrogen spectra but has limitations such as inability to explain Zeeman effect and molecular bonding.

## Solved Examples

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**Example 1:** Calculate the wavelength of light emitted when an electron in a hydrogen atom falls from the  $n=3$  to  $n=2$  energy level.

**Solution:**

Using Rydberg formula for wave number:

$$\bar{\nu} = 109,677 (1/2^2 - 1/3^2) \text{ cm}^{-1} = 109,677 (1/4 - 1/9) = 109,677 (5/36) = 15,243 \text{ cm}^{-1}$$

$$\text{Wavelength } \lambda = 1/\bar{\nu} = 1/15,243 \text{ cm}^{-1} = 6.56 \times 10^{-5} \text{ cm} = 656 \text{ nm}$$

This corresponds to red light in the visible spectrum.

**Example 2:** Find the de Broglie wavelength of an electron moving with velocity  $3 \times 10^6$  m/s. (Mass of electron =  $9.11 \times 10^{-31}$  kg, Planck's constant

**Solution:**

$$\lambda = h / mv = 6.626 \times 10^{-34} / (9.11 \times 10^{-31} \times 3 \times 10^6) = 2.42 \times 10^{-10} \text{ m}$$

This wavelength is in the X-ray region.

## Practice Set

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- **Level 1:** Define atomic number and mass number.
- **Level 2:** Explain the photoelectric effect and its significance.

- **Level 3:** Calculate the radius of the third Bohr orbit for hydrogen. (Given  $a_0 = 52.9 \text{ pm}$ )

## Answer Key

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- **Level 1:** Atomic number is the number of protons in the nucleus; mass number is the total number of protons and neutrons.
- **Level 2:** Photoelectric effect is the emission of electrons from a metal surface when light of suitable frequency falls on it. It demonstrates
- **Level 3:** Radius  $r_n = n^2 a_0 = 3^2 \times 52.9 \text{ pm} = 9 \times 52.9 \text{ pm} = 476.1 \text{ pm}$ .

## de-Broglie and Quantum Principles

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### Dual Behaviour of Matter

Einstein suggested light has both wave and particle nature. de Broglie extended this to matter, proposing particles like electrons also have wave nature. The wavelength is given by  $\lambda = h / mv$ .

### Heisenberg's Uncertainty Principle

It is impossible to simultaneously know the exact position and momentum of an electron. Mathematically,  $\Delta x \times \Delta p_x \geq h / 4\pi$ .

### Quantum Mechanical Model

Developed by Heisenberg and Schrödinger, this model treats electrons as wave functions ( $\psi$ ). The probability of finding an electron is given by  $\psi^2$  and is determined simultaneously; only probabilities can be calculated.

## Solved Examples

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**Example:** Calculate the uncertainty in velocity if the uncertainty in position of an electron is  $1 \times 10^{-10} \text{ m}$ . (Mass of electron =  $9.11 \times 10^{-31} \text{ kg}$ ,  $h = 6.626 \times 10^{-34} \text{ J s}$ )

**Solution:**

$$\Delta x \times \Delta v \geq h / 4\pi m$$

$$\Delta v \geq h / (4\pi m \Delta x) = 6.626 \times 10^{-34} / (4\pi \times 9.11 \times 10^{-31} \times 1 \times 10^{-10}) \approx 5.8 \times 10^3 \text{ m/s}$$

## Practice Set

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- **Level 1:** State de Broglie's hypothesis.
- **Level 2:** Explain Heisenberg's uncertainty principle.
- **Level 3:** Calculate the de Broglie wavelength of a proton moving at  $2 \times 10^6 \text{ m/s}$ . (Mass of proton =  $1.67 \times 10^{-27} \text{ kg}$ )

## Answer Key

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- **Level 1:** Matter particles have wave properties with wavelength  $\lambda = h / mv$ .
- **Level 2:** It is impossible to know exact position and momentum of a particle simultaneously.
- **Level 3:**  $\lambda = h / mv = 6.626 \times 10^{-34} / (1.67 \times 10^{-27} \times 2 \times 10^6) = 1.98 \times 10^{-13} \text{ m}$ .

## Orbitals and Quantum Numbers

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### Atomic Orbitals

Atomic orbitals are regions in space where electrons are most likely to be found. They differ in size, shape, and orientation.

### Quantum Numbers

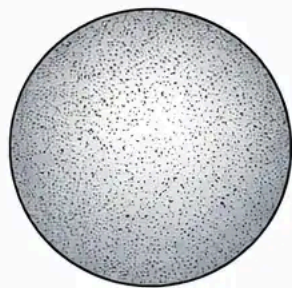
Four quantum numbers describe electron states:

- **Principal quantum number (n):** Energy level and size of orbital (n=1,2,3,...)
- **Azimuthal quantum number (l):** Shape of orbital (l=0 to n-1)
- **Magnetic quantum number (m<sub>l</sub>):** Orientation of orbital (m<sub>l</sub> = -l to +l)
- **Spin quantum number (m<sub>s</sub>):** Electron spin (+½ or -½)

### Shapes of Orbitals

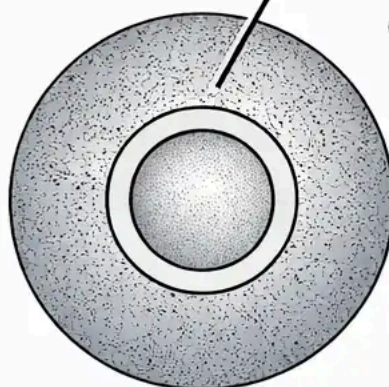
**s-orbitals:** Spherical shape.

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1s-orbital

(a)



Region of zero probability called node.

2s-orbital

(b)

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## Shape of 1s and 2s orbitals

**p-orbitals:** Dumbbell-shaped with three orientations ( $p_x$ ,  $p_y$ ,  $p_z$ ).

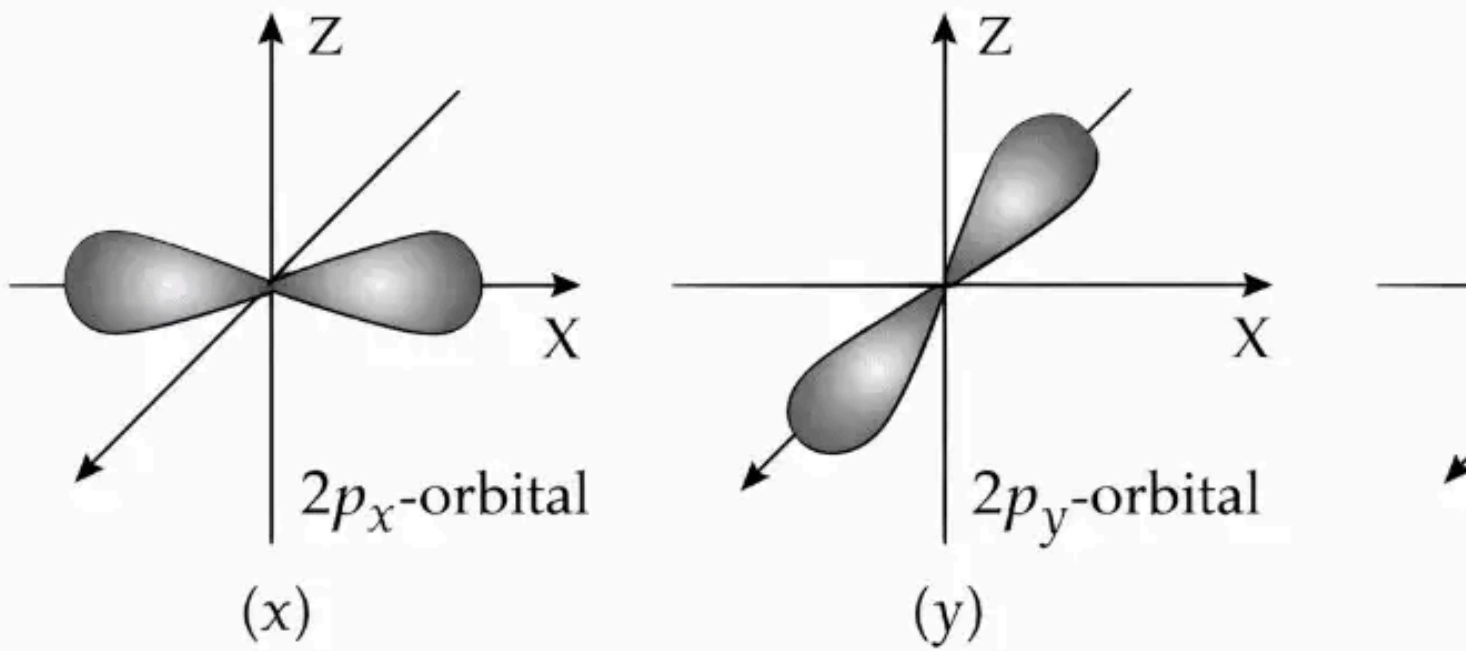


Fig. 2.2: Shape of  $p$ -orbitals

**d-orbitals:** Five shapes, four cloverleaf-shaped and one donut-shaped.

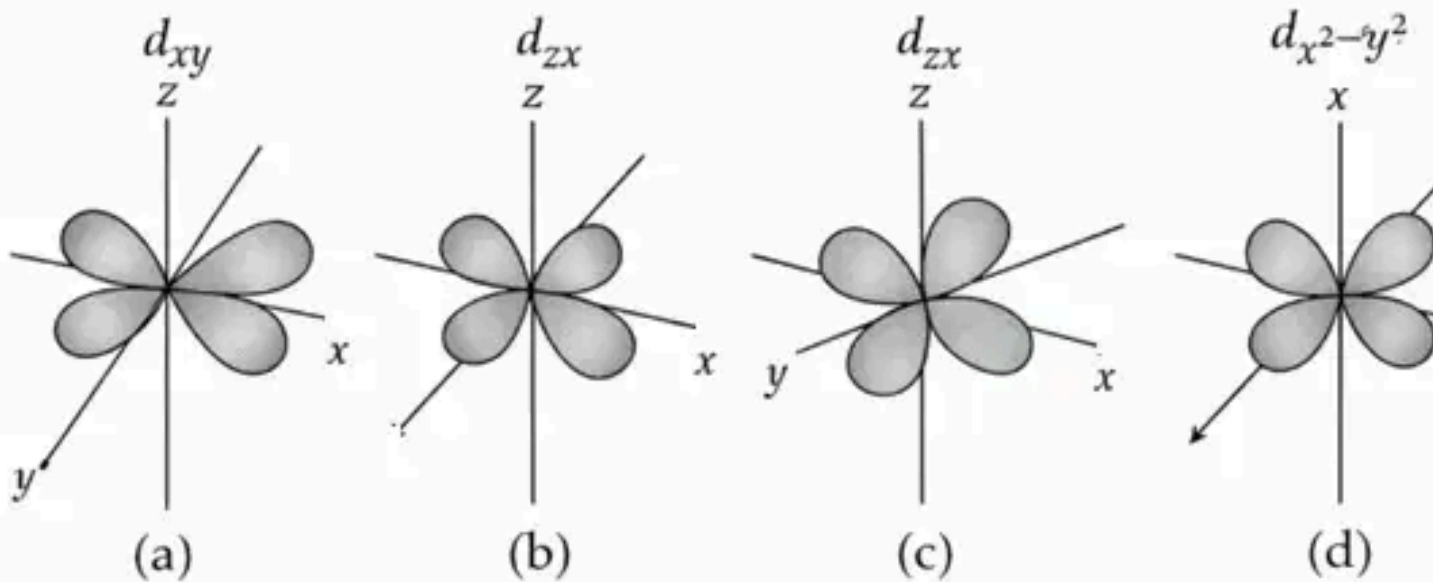
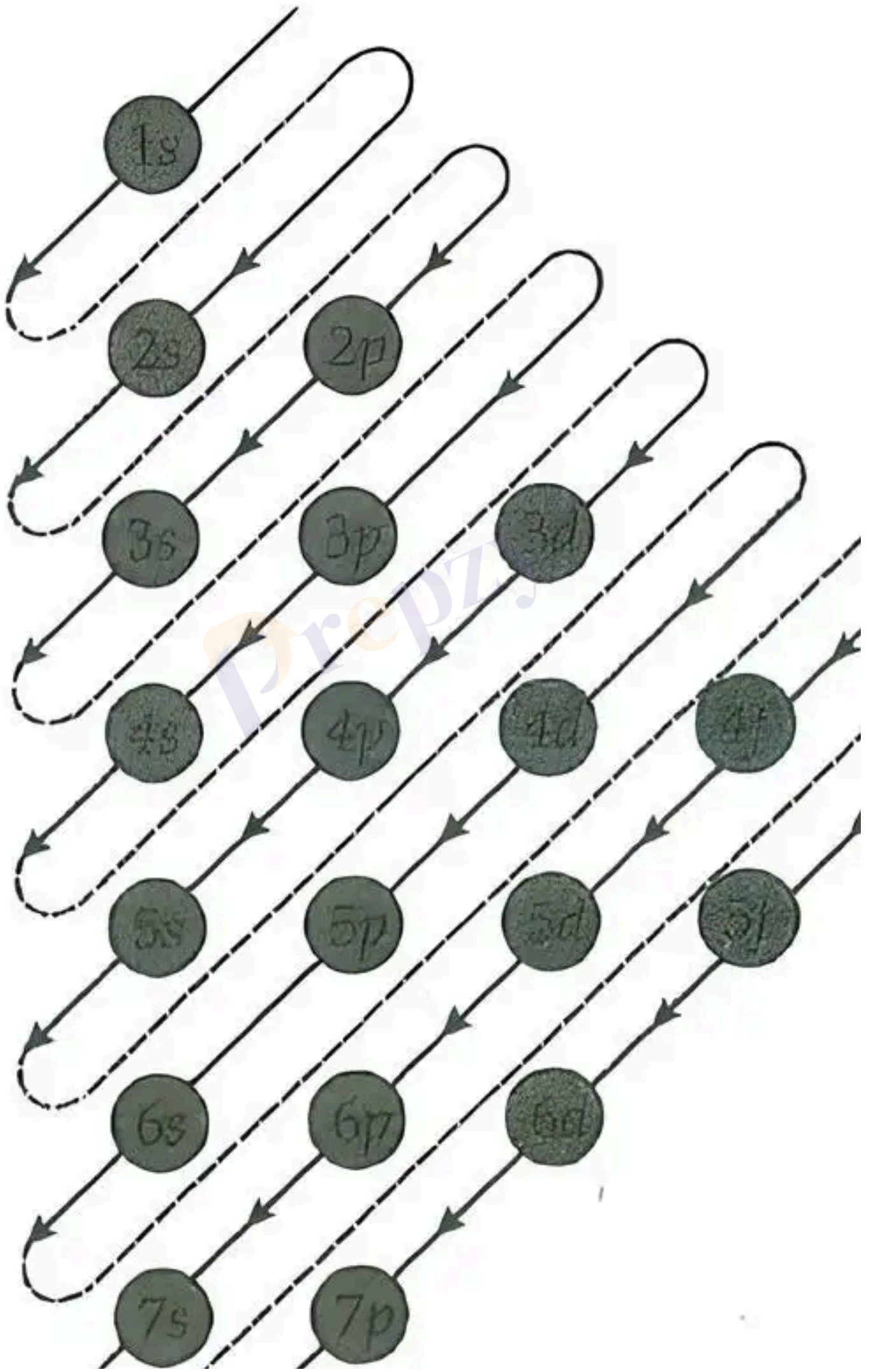


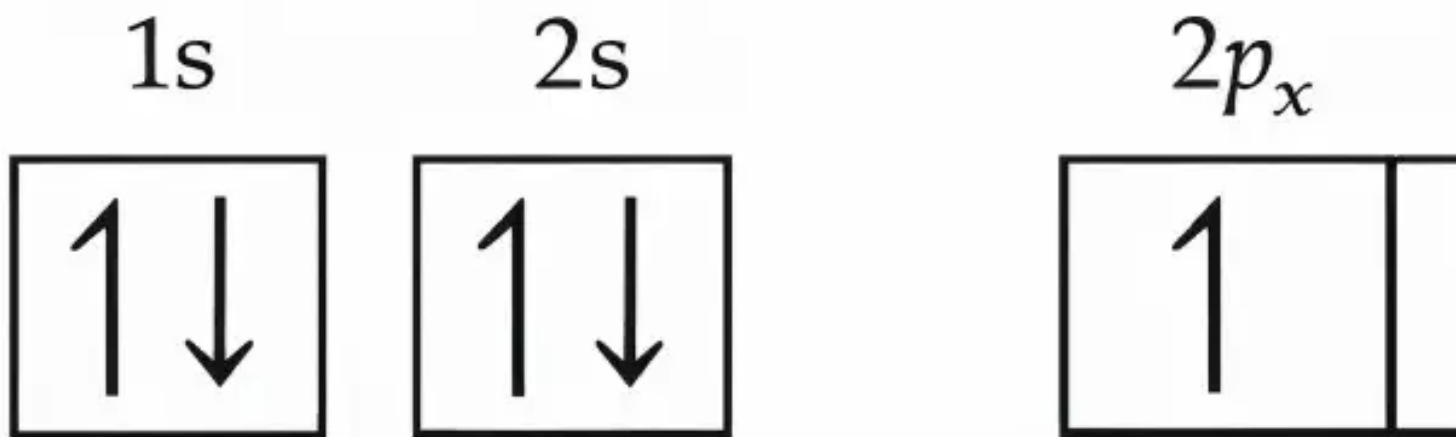
Fig. 2.3: Shape of five  $d$ -orbitals

## Electron Configuration Rules

- **Aufbau Principle:** Electrons fill orbitals from lowest to highest energy.
- **Pauli's Exclusion Principle:** No two electrons have the same set of four quantum numbers; maximum two electrons per orbital with opposite spins.
- **Hund's Rule:** Electrons occupy degenerate orbitals singly with parallel spins before pairing.

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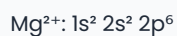




**Fig. 2.4: Order of filling of orbitals**

### Electronic Configuration

Distribution of electrons in orbitals determines chemical properties. For example:



### Solved Examples

**Example:** Write the electronic configuration of nitrogen (atomic number 7).

**Solution:**  $1s^2 2s^2 2p^3$

### Practice Set

- **Level 1:** Define principal quantum number.
- **Level 2:** State Pauli's exclusion principle.
- **Level 3:** Write the electronic configuration of oxygen (atomic number 8) and explain Hund's rule.

### Answer Key

- **Level 1:** Principal quantum number (n) indicates the energy level and size of the orbital.
- **Level 2:** No two electrons in an atom can have the same set of four quantum numbers.
- **Level 3:** Oxygen:  $1s^2 2s^2 2p^4$ . Hund's rule states electrons fill degenerate orbitals singly with parallel spins before pairing.

### Quick Reference Table

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