

Atomic Structure

1. Idea of de Broglie Matter Waves

Concept:

Louis de Broglie proposed that matter (like electrons) exhibits wave-like properties. This concept bridges the gap between particles and waves, introducing the wave-particle duality of matter.

Derivation of de Broglie Wavelength:

1. From Einstein's energy equation for photons:

$$E = h\nu$$

Where h is Planck's constant, and ν is the frequency.

2. For a photon, energy is also related to momentum p :

$$E = pc \quad (\text{where } c \text{ is the speed of light}).$$

3. Equating these two expressions for energy:

$$pc = h\nu$$

4. Substituting $\nu = \frac{c}{\lambda}$, we get:

$$p = \frac{h}{\lambda}$$

5. Extending this concept to particles:

- Momentum $p = mv$ for a particle of mass m and velocity v .
- Thus, de Broglie wavelength is:

$$\lambda = \frac{h}{mv}$$

Significance:

For macroscopic objects, the wavelength is negligibly small, making wave properties undetectable. For subatomic particles (e.g., electrons), λ becomes significant and measurable.

Example:

An electron ($m = 9.1 \times 10^{-31} \text{ kg}$) moving at 10^6 m/s :

$$\lambda = \frac{6.626 \times 10^{-34}}{(9.1 \times 10^{-31})(10^6)} \approx 7.27 \text{ nm}$$

2. Heisenberg Uncertainty Principle

Statement:

It is impossible to simultaneously measure the exact position (x) and momentum (p) of a particle with absolute certainty. The principle is mathematically expressed as:

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

Where:

- Δx = Uncertainty in position
- $\Delta p = m \cdot \Delta v$, uncertainty in momentum
- h = Planck's constant

Derivation:

1. From wave-particle duality, momentum p is related to wavelength:

$$p = \frac{h}{\lambda}$$

2. A wave packet representing a particle has a spread in wavelengths ($\Delta\lambda$) due to wave superposition:

$$\Delta p = \frac{h}{\Delta\lambda}$$

3. The spread in position (Δx) is inversely proportional to the spread in wavelength:

$$\Delta x \cdot \Delta\lambda \approx 1$$

4. Combining the above equations:

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

Significance:

The uncertainty principle limits the precision of measurements, implying that electrons do not follow deterministic paths (as suggested by Bohr's model).

3. Schrödinger Wave Equation

The Schrödinger equation provides a mathematical framework for quantum mechanics. It describes the wave behavior of particles in terms of their energy and position.

Time-Independent Schrödinger Equation:

$$\hat{H}\psi = E\psi$$

Where:

- $\hat{H}$ = Hamiltonian operator (total energy operator)
- ψ = Wave function (describes the quantum state of a system)
- E = Total energy

The Hamiltonian operator is expressed as:

$$\hat{H} = -\frac{\hbar^2}{2m}\nabla^2 + V(r)$$

Where:

- $-\frac{\hbar^2}{2m}\nabla^2$ = Kinetic energy operator
- $V(r)$ = Potential energy

Interpretation:

- ψ gives the probability amplitude of finding an electron in a region.
- $|\psi|^2$ gives the probability density.

4. Quantum Numbers

Principal Quantum Number (n):

Defines the energy level of an electron and its average distance from the nucleus.

- Example: For $n = 1$, the electron is in the first energy level.

Azimuthal Quantum Number (l):

Determines the shape of the orbital:

$$l = 0, 1, 2, \dots, (n - 1)$$

- Example: $l = 0$ (s-orbital), $l = 1$ (p-orbital).

Magnetic Quantum Number (m_l):

Specifies the orientation of the orbital:

$$m_l = -l, \dots, 0, \dots, +l$$

Spin Quantum Number (m_s):

Represents the electron's spin:

$$m_s = +\frac{1}{2} \text{ or } -\frac{1}{2}$$

5. Shapes of Orbitals

s-Orbitals:

- Spherical in shape.
- Example: Hydrogen's $1s$ orbital.

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p-Orbitals:

- Dumbbell-shaped.
- Three orientations (p_x , p_y , p_z).

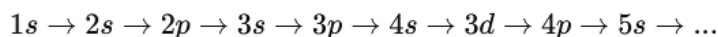
d-Orbitals:

- Complex shapes (cloverleaf and donut).
- Five orientations (d_{xy} , d_{yz} , etc.).

6. Aufbau Principle

Electrons occupy orbitals in the order of increasing energy.

Order of Filling:



Example:

- Carbon ($6e^-$): $1s^2 2s^2 2p^2$.
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7. Effective Nuclear Charge (Z_{eff})

The net positive charge experienced by an electron:

$$Z_{\text{eff}} = Z - S$$

Where:

- Z = Atomic number (total protons)
- S = Shielding constant (due to inner electrons).

Trends:

- Increases across a period (due to higher nuclear charge).
 - Decreases down a group (due to increased shielding).
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Disclaimer:

These notes are intended to serve as a quick reference or revision guide for recalling important concepts. They summarize key topics to help users understand and retain information efficiently.

However, for a more comprehensive understanding, detailed study, and in-depth learning, it is highly recommended to consult subject matter experts or refer to standard textbooks and authoritative sources. While we have worked meticulously to ensure accuracy and relevance in these notes, we cannot guarantee they are free of errors. Users are advised to cross-check the information and use these notes at their own discretion. We take no responsibility for any inaccuracies or their consequences.

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