

Chapter 7: Quantum Theory

Quantum View

Electromagnetic Spectrum: based on wavelength and frequency.

Continuous spectrum: white light is continuous spectrum, even when passing through a prism it is continuous as one each color merges into the other.

Line spectrum: obtained by exciting a gas, passing the energy released through a slit and prism. Only certain lines are observed. Also called emission spectrum. Each element has a unique line spectrum hence also known as fingerprint of element.

1900	Max Planck	Derived relationship between energy and frequency. All matter absorbs or emits electromagnetic energy only in discrete quantity called quanta. $E = hv$ (energy = Planck's constant x frequency)
1905	Albert Einstein	Photoelectric effect. He called the bundles of energy Photons.
1913	Niels Bohr	Proposed atom has energy levels. Electrons stay in these energy levels and when energy is absorbed there is transition to another level thus energy is given off when electron comes back to the original level. Thus energy is always given off in certain quantities. $E_n = -B/n^2$
1923	Louis de Broglie	Proposed that particles have a wave nature. (led to development of electron microscope) $\lambda = h/mv$
1926	Erwin Schrödinger	Developed the wave function equation which determines the position of an electron in the atom.
1927	Werner Heisenberg	Said that one cannot determine the location of the electron.

Calculations

E = energy

h = Planck's constant

ν = frequency

B = constant

n = shell number

m = mass

v = velocity

λ = wavelength

Planck's equation $E = h\nu$

Bohr's equation for calculation of energy of electron in an energy level, energy transition and line spectrum: $E_n = -B/n^2$

De Broglie's equation for calculating wavelength of electron: $\lambda = h/mv$

Concepts

1) Ground state and excited state

2) Quantum numbers

- a. Principal quantum number (n) – shell/orbit number (1, 2, 3, 4 etc)
- b. Orbital angular momentum quantum number (l) – orbital location (s, p, d, f)
- c. Magnetic quantum number (m_l) – shape of orbitals (1 for s, 3 for p, 5 for d and 7 for f)
- d. Electron spin quantum number (m_s) – electron spin

Table 1.

Quantum Number	Allowed Values	Name and Meaning
n	$n = 1, 2, 3, \dots$	<i>Principal quantum number</i> : orbital energy and size.
l	$l = (n-1), (n-2), \dots, 0$	<i>Azimuthal (or orbital) quantum number</i> : orbital shape (and energy in a multi-electron atom), letter name for subshell (s, p, d, f)
m_l	$m_l = l, (l-1), \dots, 0, \dots, (-l+1), -l$	<i>Magnetic quantum number</i> : orbital orientation
m_s	$m_s = 1/2, -1/2$	<i>Electron spin quantum number</i> : spin up (\uparrow) or spin down (\downarrow).

Table 2

l Value	Letter Equivalent to l Value	No. of Orbitals in Set	Approximate Shape of Orbitals with Specific l Values
0	s	1	spherical
1	p	3	p_x, p_y, p_z are dumbbells along x, y, and z axes
2	d	5	mostly cloverleaf shapes
3	f	7	very complicated shapes!

Table 3

Shell (n)	Subshell (l)	Orbital Name (nl)	Orientations (m_l)	No. of Orbitals	Maximum Occupancy
$n = 1$	$l = 0$	1s	$m_l = 0$	1	2 e ⁻
$n = 2$	$l = 0$	2s	$m_l = 0$	1	2 e ⁻
	$l = 1$	2p	$m_l = 1, 0, -1$ (or p_x, p_y, p_z)	3	6 e ⁻
$n = 3$	$l = 0$	3s	$m_l = 0$	1	2 e ⁻
	$l = 1$	3p	$m_l = 1, 0, -1$ (or p_x, p_y, p_z)	3	6 e ⁻
	$l = 2$	3d	$m_l = 2, 1, 0, -1, -2$ (or $d_{xy}, d_{yz}, d_{xz}, d_{x^2-y^2}, d_{z^2}$)	5	10 e ⁻