6.1 – The Wave Nature of Light

Electromagnetic Radiation – A form of heat that material objects may exchange

Wavelength (λ) – Distance between two adjacent peaks of a wave

Frequency (ν) – Cycles per second of a wave

\[ c = \lambda \nu \]

Electromagnetic Spectrum:

- **Gamma Rays**: < .01 nm
- **X-Rays**: .01nm-10nm
- **UV**: 10nm-400nm
- **Visible**: Violet (420nm), Blue (470nm), Green (530nm), Yellow (580nm), Orange (620nm), Red (700nm)
- **Infrared**: 700nm-1mm
- **Microwaves**: 1mm – 1dm
- **Radio**: >1dm

6.2 – Quantized Energy and Photons

Quantum – Smallest quantity of energy that can be emitted or absorbed as EMR

Planck proved this because the theory of incandescence showed that light was emitted in portions (not continuous)

Wave properties of EMR were observed in Young’s double slit experiment

\[ E = h\nu \]

*Photoelectric Effect* – Light shining on a metal surface causes the surface to emit electrons. A minimum frequency of light is required to emit them

Einstein explained it by saying that EMR is not only generated in packets but also travels in packets

*Work function* (Φ) – Certain amount of energy required for electrons to overcome attractive forces holding them in the metal

\[ E_{\text{photon}} = \frac{1}{2}m_ev_e^2 + \Phi \]

Alternatively: \( KE = h\nu_{\text{photon}} - h\nu_0 \) or \( KE = E_{\text{photon}} - E_0 \) where the naught values are for the threshold frequency/energy

Photon – A quantum of light

6.3 – Line Spectra and the Bohr Model

Spectrum – Produced when radiation is separated into its component wavelengths

*Continuous Spectrum*– Rainbow of colors containing all wavelengths of light
Andrew Rosen

**Line Spectrum** – A spectrum containing radiation of only specific wavelengths

Bohr explained the line spectrum of hydrogen by proposing that only certain orbits for the electron are possible with certain energies

\[ E = (-\hbar c R_H) \left( \frac{1}{n^2} \right) \] (for hydrogen)

**Rydberg Equation:** \[ \frac{1}{\lambda} = R_H \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right) \] (for hydrogen)

Balmer series – Drop of an electron to n=2

Ground State – Lowest-energy state of the atom

Excited State – When the electron is in a higher-energy state

If an atom in the ground state absorbed enough energy to move the electron infinitely far away from the nucleus (n=∞), then the atom is ionized

### 6.4 – The Wave Behavior of Matter

Debroglie came up with matter-waves: \[ \lambda = \frac{\hbar}{mv} \]

**Uncertainty Principle** – It is impossible to know simultaneously both the exact momentum of an electron and its exact location

\[ \Delta x \Delta p \geq \frac{\hbar}{4\pi} \]

### 6.5 Quantum Mechanics and Atomic Orbitals

Quantum mechanics – Mathematical treatment of the motion of very small particles as waves developed by Schrodinger

Standing waves do not travel in space like the electron in a hydrogen atom

Nodes are points where the amplitude of the wave is zero

**Wave Function (ψ)** – Describes the electron in an atom

**Probability density/electron density** = \( \psi^2 \)

Orbitals are described by n, l, and m quantum numbers

**Principal Quantum Number (n)** – Defines energy of an orbital and average distance of an electron from the nucleus

\[ n \geq 1 \]

Orbitals with the same value of n form a shell

Different orbital types within a shell are subshells

**Azimuthal/Angular Momentum Quantum Number (l)** – Defines the shape of the orbital

\[ l = 0 \text{ to } n - 1 \]

S orbital (l=0), p orbital (l=1), d orbital (l=2), f orbital (l=3)
Andrew Rosen

**Magnetic Quantum Number ($m_l$)** – Describes the three-dimensional orientation of the orbital

$$-l \leq m_l \leq l$$

### 6.6 – Representations of Orbitals

S orbitals are always spherically symmetrical with a radius that increases with n

- Boundary surface includes region with 90% electron probability

**Radial Probability Function** – Probability that we will find the electron at a specific radius on the sphere ($r$)

$$P(r) = 4\pi r^2 |\psi(r)|^2$$

**Radial Node** – Spherical surface with zero probability of finding an electron

- **Total Number of nodes in orbital = n-1**
- **Angular Nodes = l**
- **Radial Nodes = n – 1 – l**

P orbital shapes:

D orbital shapes:

### 6.7 – Many-Electron Atoms

Orbitals with the same energy are said to be **degenerate**

- The hydrogen atom is the only one that has the same orbital energies if they have the same n value (4s=4p energy)

All other atoms have increasing energies with increasing $l$ values

**Spin-Magnetic Quantum Number ($m_s$)** – Can be +1/2 or -1/2 and is the spin of the electron

Pauli Exclusion principles states that no two electrons in an atom can have the same set of four quantum numbers
An orbital can hold a maximum of 2 electrons, and they must have opposite signs.

Orbital shapes for other atoms are approximated using orbitals from the hydrogen atom.

### 6.8 – Electron Configurations

The electron configuration is the way electrons are distributed among the various orbitals of an atom.

Ground state is the most stable electron configuration.

Orbitals are filled in order of increasing energy.

Hund’s rule – For degenerate orbitals, the lowest energy is attained when the number of electrons having the same spin is maximized.

Electron configurations can be condensed using the noble gas that precedes it.

**Core electrons** - Inner-shell electrons oftentimes abbreviated with the configuration of the noble gas but can also include full d and f orbitals (for s and p block) or full f orbitals (for the d block).

**Valence electron** – Outer-shell electrons that are involved with chemical bonding that can be from different n values.

Eg: Zr is [Kr] 5s^2 4d^2 and this has 4 valence electrons.

Transition metals fill the d orbitals.

Lanthanides and actinides fill the f orbitals.

### 6.9 – Electron Configurations and the Periodic Table

**Representative elements/main group elements** – the s and p block elements

![Diagram of electron configurations and the periodic table](image)

- **s-block elements**
- **p-block elements**
- **Transition metals**
- **f-block metals**

The d block has n=period-1 and the f block has n=period-2.

The number of columns that each block has is the maximum amount of electrons that can occupy that subshell.

When using the periodic table to find electron configurations, **be careful and make sure that, if applicable, the f block electron configurations are included**.

The number of valence electrons is equal to the group number and the second digit for all groups greater than 10.

The closeness of the 3d and 4s orbitals can create anomalous electron configurations, especially in the transition metals.