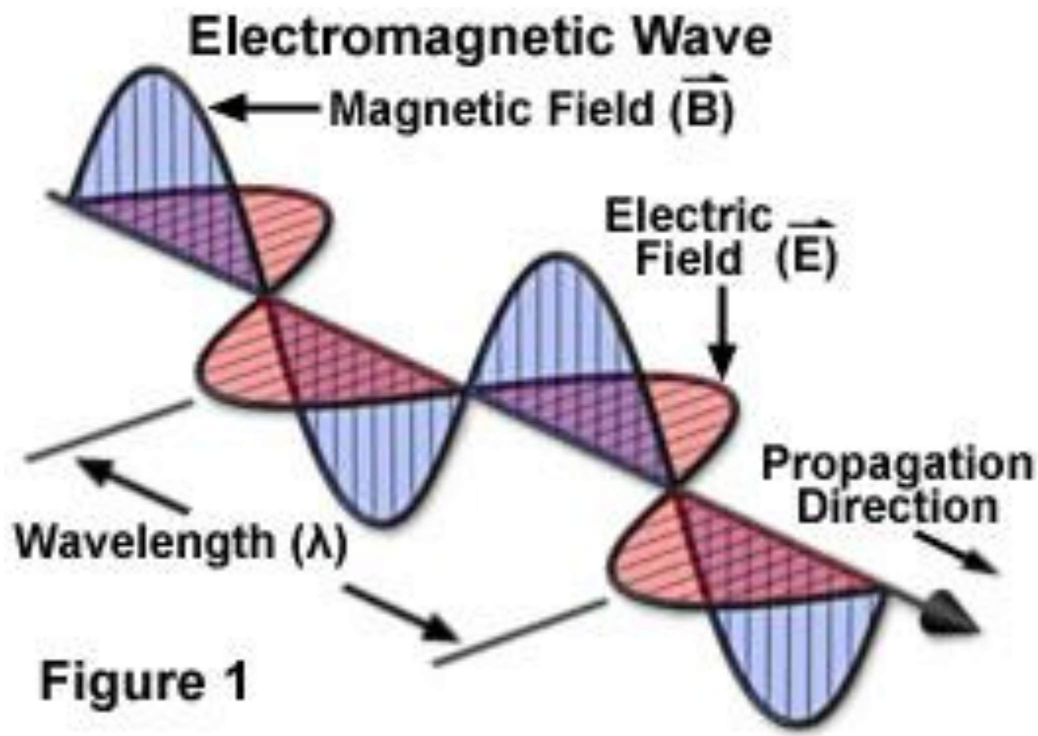
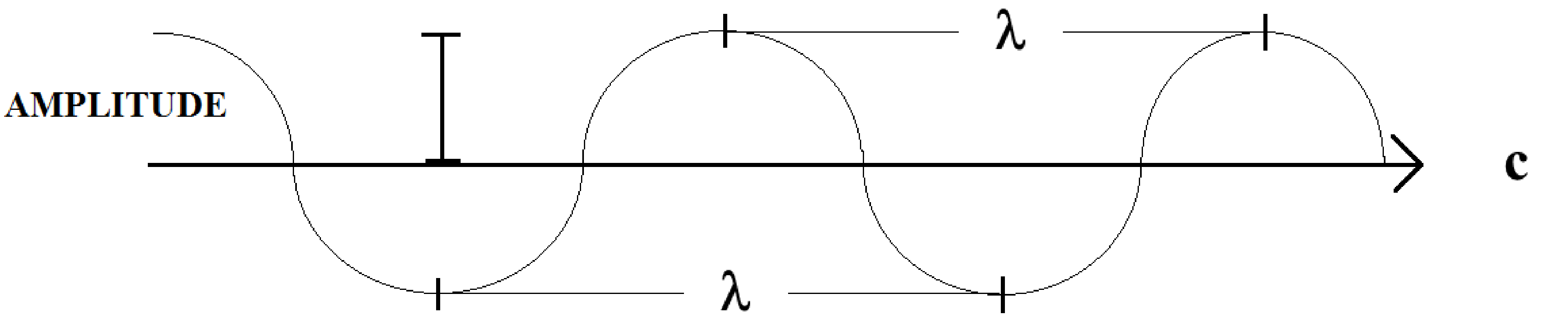
Electromagnetic Radiation:

Electromagnetic(EM) radiation is the transmission of energy in the form of a wave. Consists of an electric and magnetic component.



Wave Parameters of Measurement:

λ: wavelength(in metres or nanometers)
u: speed(in m/s)
c= speed of light in vacuum(3.00×10^{8}m/s)
ν: frequency(in Hz or s^{-1})



In vacuum.

$$\lambda = \frac{c}{\nu}$$

$$c = \lambda \cdot \nu$$

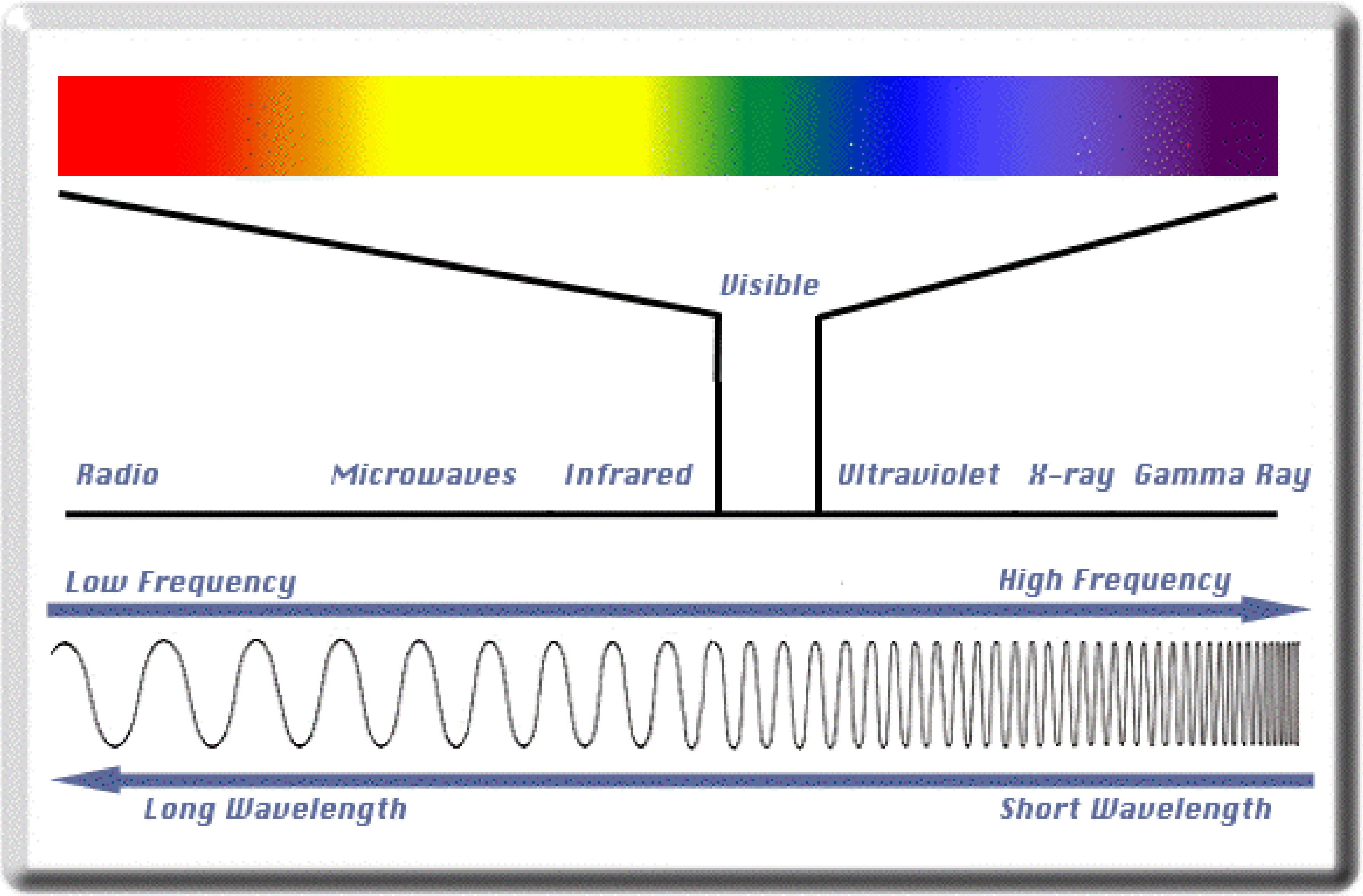
Ex:

What is the frequency of light emitted by a sodium vapor lamp(λ = 589 nm)?

$$1\ \text{Å} = 1\times10^{-10}\ \text{m}$$

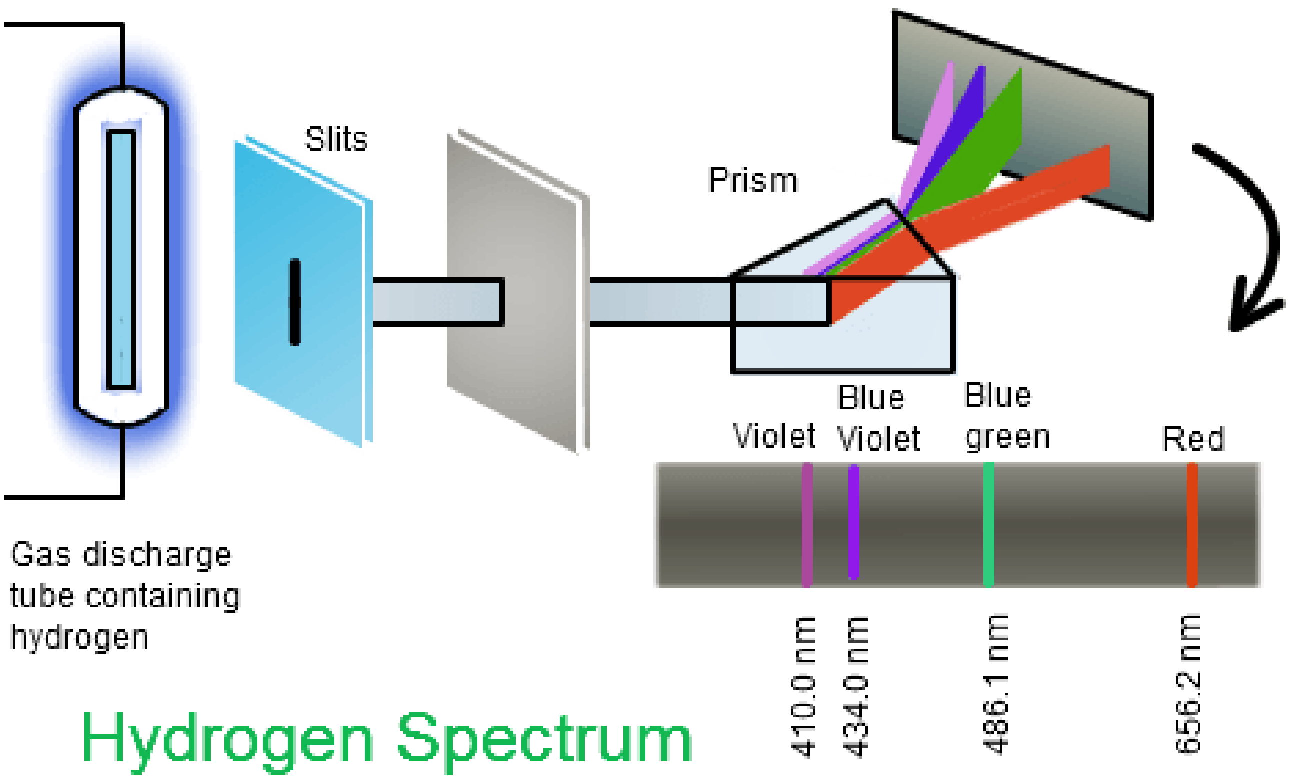
$$1\ \text{nm} = 1\times10^{-9}\ \text{m}$$

Electromagnetic Spectrum:



Atomic Spectra:

When elements are exposed to energy(heat, light, electrical energy) certain wavelengths of light are emitted and is known as a line spectrum.



Reference: https://chemistry.tutorvista.com/inorganic-chemistry/spectral-lines.html

Planck's Equation:

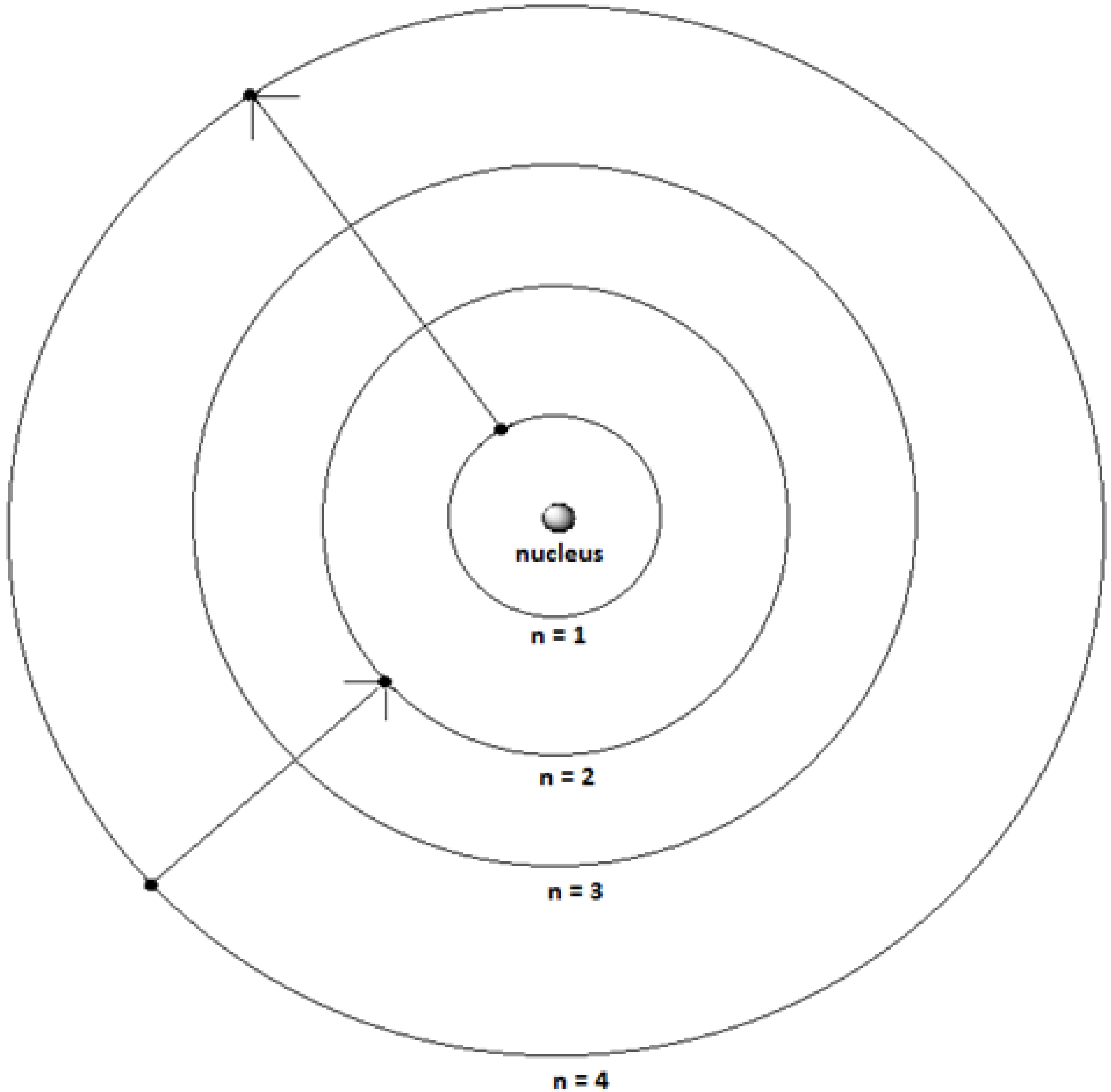
$$E = h\nu$$

E: energy(in joules)
h: Planck's constant($h = 6.626 \times 10^{-34}$ J·s)

Ex:
Calculate the energy of a single photon of blue light with a wavelength of 435 nm.

Bohr's Model of the Hydrogen Atom:

Atom absorbs energy. Electron excited from n = 1 to n = 4.



Electron moves from n = 4 to n = 2. Light emitted. 486 nm(Blue-green).

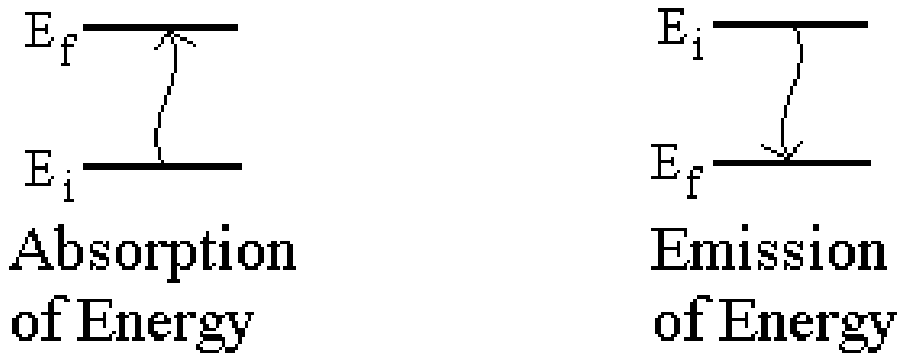
Bohr's Model of the Hydrogen Atom:

The electron travels around the nucleus in well defined fixed orbits.

Electrons in orbits closest to the nucleus are the lowest in energy.

For an electron to be excited from a low orbit to a high orbit it must absorb a specific amount of energy. When the electron goes from a high orbit to a lower orbit a specific amount of energy is given off.

Bohr's Model of the Hydrogen Atom:



$$\Delta E = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

ΔE: energy absorbed or emitted

R_H: Rydberg Constant(2.179×10^{-18}J)

Ex: Calculate the energy absorbed or given off when an electron travels from the n =5 to the n = 2 level.

$$\Delta E = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$\mathbf{R_H = 2.179\times10^{-18} J}$$

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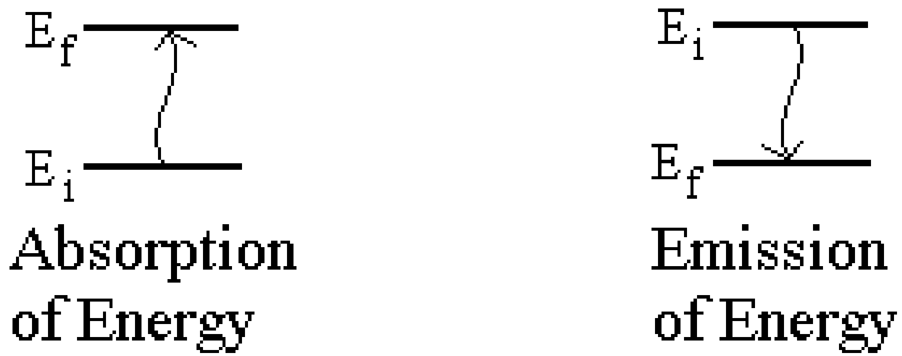

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Wave-Particle Duality:

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

u: velocity
m: mass
λ: wavelength
$h = 6.626\times10^{-34} kg\cdot m^{2}\cdot s^{-1}$

Ex: Calculate the wavelength of a beam of electrons travelling at a speed of 3.00×10^{7} m/s.
mass of electron = 9.109×10^{-31} kg

The Uncertainty Principle:

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

m: mass

Δx: uncertainty in position

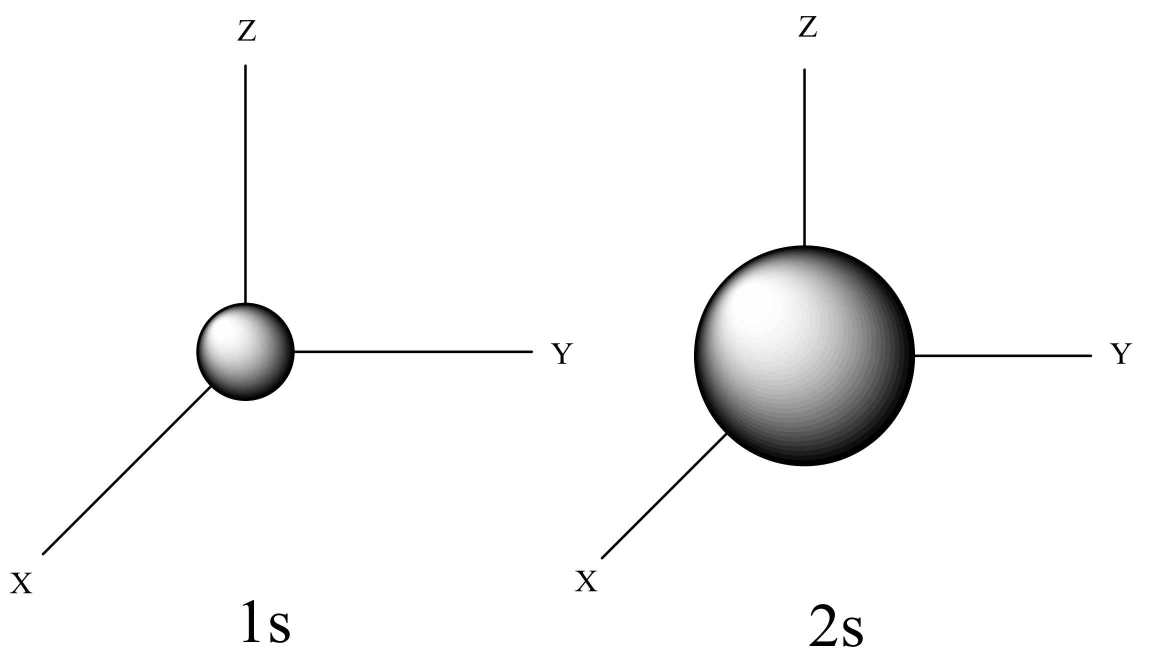
Δv:uncertainty in velocity

$h = 6.626 \times 10^{-34} kg \cdot m^2 \cdot s^{-1}$

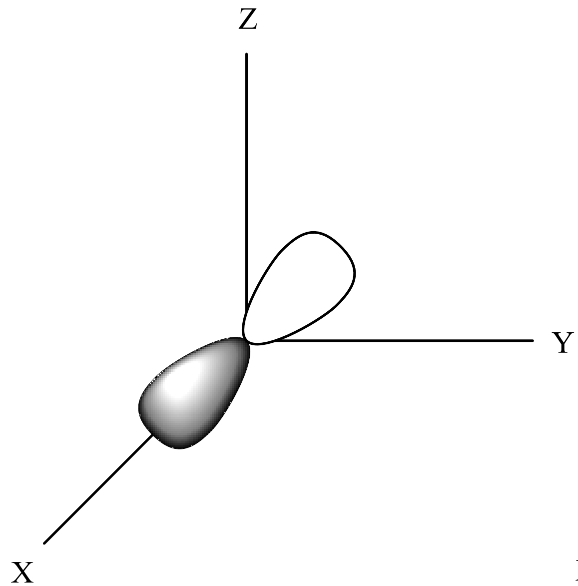
Atomic Orbitals:

Atomic orbital is a region of space where there is a high probability of finding an electron.

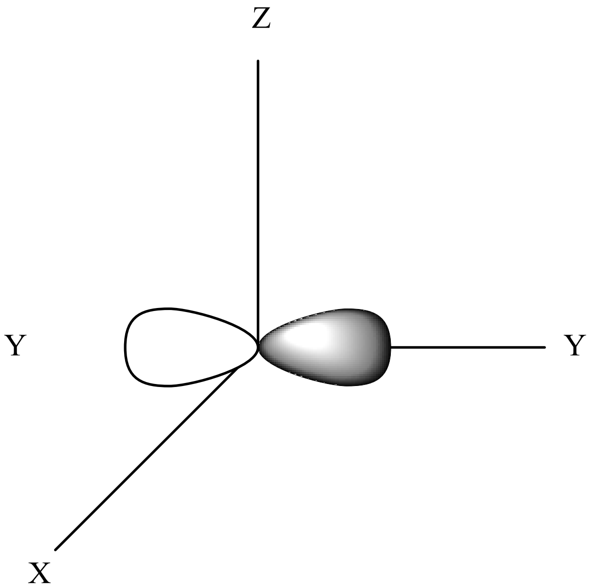
s orbital



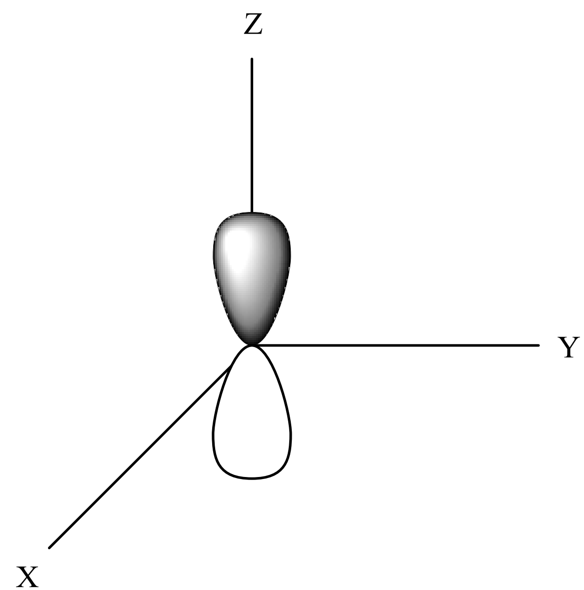
p orbitals



p_x

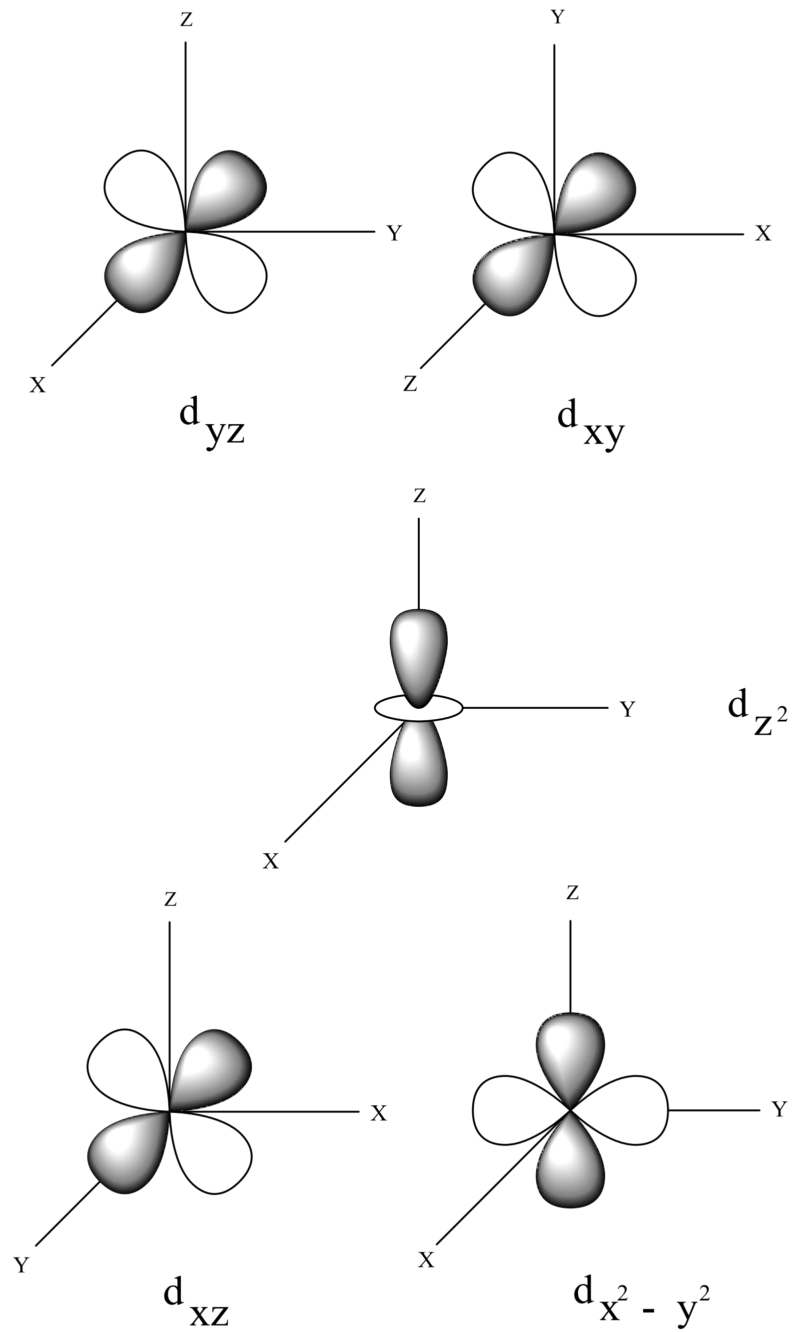


p_y



p_z

d orbitals



Quantum Numbers:

Principle Quantum Number(n):

Assigns the level or shell to which an electron belongs. Indicates relative distance from the nucleus.

n= 1, 2, 3, … **(Positive Integers)**

Orbital(Angular-Momentum) Quantum Number(l):

Assigned to each of the subshells in a shell. Indicates the shape of the orbital. l is a positive integer including zero but no larger than n-1.

$l = 0, 1, 2, \ldots, n-1$

Also denote subshells by letter.

l =	**0,**	**1,**	**2,**	**3,**	**4**
notation =	**s,**	**p,**	**d,**	**f,**	**g**

Magnetic Quantum Number(m_l):

Assigned to each orbital in a subshell.

Describes the relative orientation of the orbital.

Negative or positive integer and range from -l to +l.

$$m_l = -l, -l+1, -l+2, \dots, 0, \dots +l$$

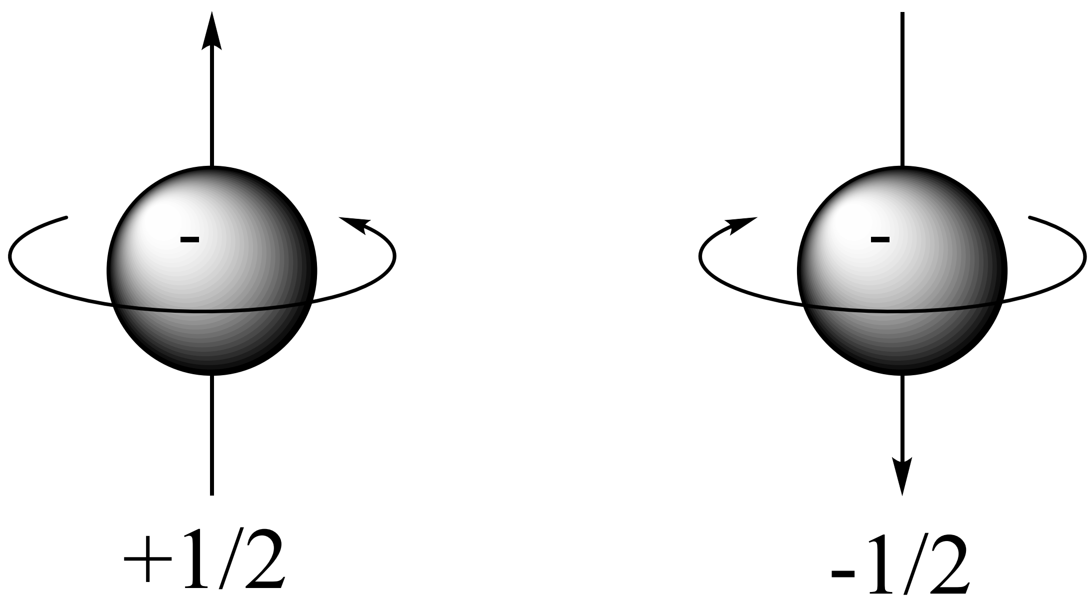
Thus for	**l = 0(s orbital)**	$m_l = 0$
	l = 1(p orbital)	$m_l = -1, 0, +1$

Corresponds to
p_x, p_y, and p_z.

Magnetic Spin Quantum Number(m_s)

m_s describes the spin of an electron

m_s can be +1/2(denoted by ↑) or-1/2(denoted by ↓)



Electron Configuration:

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d

Hund's Rule: When filling electrons into orbitals of identical energy, electrons occupy these orbitals singly before pairing up.

Pauli's Exclusion Principle: For a single atom, no two electrons have the same four quantum numbers.

Aufbau Process: "Building Up Method." Determining the electron configuration of an atom is achieved by the succesive adding of electrons until the desired configuration is obtained.

Ex: Nitrogen atomic number = 7 (7 electrons)

N $1s^2\ 2s^2\ 2p^3$ or $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$

or

N $\underline{\uparrow\downarrow}$ $\underline{\uparrow\downarrow}$ $\underline{\uparrow\ \uparrow\ \uparrow}$

1s 2s 2p