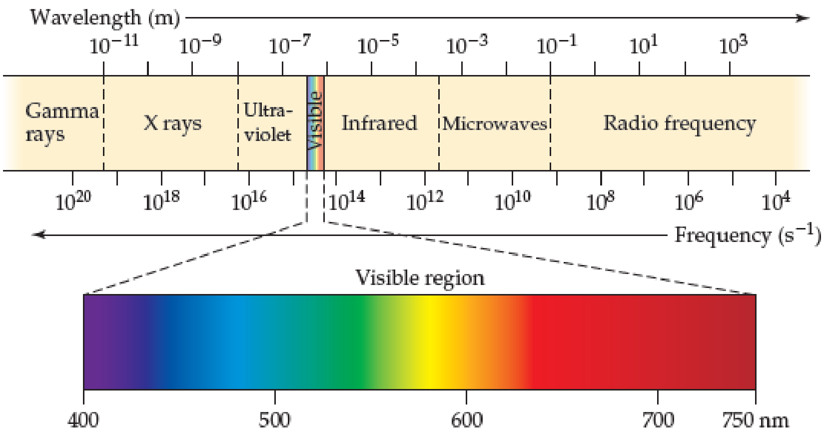
General Chemistry

Chapter 6 • Electronic Structure of Atoms

Review

Definitions:

- **Electromagnetic radiation:** the radiant energy released by certain electromagnetic processes. Visible light is electromagnetic radiation, as is invisible light, such as radio, infrared, and X-rays.
- **The speed of light (c):** 2.998×10^8 m/s. All electromagnetic radiation moves at the same speed.
- **Periodic:** the pattern of peaks and troughs repeats itself at regular intervals.
- **Wave length (λ):** the distance between two adjacent peaks (or between two adjacent troughs). (unit: m)
- **Frequency (ν):** the number of complete wavelengths or cycles that pass a given point each second. (unit: s^{-1})
- $\lambda\nu = c$
- **Electromagnetic spectrum:**



- **Quantum (E):** the smallest quantity of energy that can be emitted or absorbed as electromagnetic radiation.
  - ✓ **E = hν;** Plank constant (h): 6.626×10^{-34} J-s
- **Photoelectric effect:** Light shining on a clean metal surface causes electrons to be emitted from the surface. A minimum frequency of light, different for different metals, is required for the emission of electrons.
  - ✓ **Photon:** energy of photon = E = hν
- **Work function:** A certain amount of energy that required for the electrons to overcome the attractive forces holding them in the metal.
- **dual wave particle nature:** Light possesses both wave-like and particle-like characteristic and, depending on the situation, will behave more like waves or more like particles.
- **Spectrum:** is produced when radiation from polychromatic is separated into its component wavelengths.
  - ✓ **Monochromatic:** the radiation composed of a single wavelength.
  - ✓ **Polychromatic:** the radiation composed of many different wavelength.

- ✓ **Continuous spectrum:** a spectrum consists of a continuous range of colors.
- ✓ **Line spectrum:** a spectrum consists of only specific wavelengths.

- **Bohr's model:** The energies corresponding to the allowed orbits for the electron in the hydrogen atom are:

$$\mathbf{E = (-hc}\boldsymbol{R_H}\mathbf{)\left(\frac{1}{n^2}\right) = (-2.18 \times 10^{-18}}\boldsymbol{J}\mathbf{)\left(\frac{1}{n^2}\right)}$$

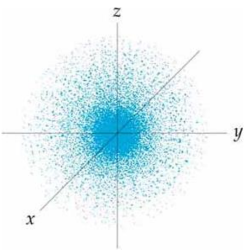
- ✓ **Ground state:** The lowest-energy state (n=1)
- ✓ **Excited state:** When the electron is in a higher-energy state (n=2 or higher)
- ✓ The energy change for the transition of electron jumping from an initial state to a final state:

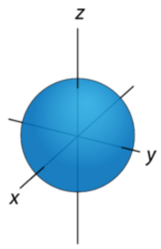
$$\Delta E = E_f - E_i = (-2.18 \times 10^{-18} J)\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

- **Momentum:** mv
- **Matter waves:** $\lambda = \frac{h}{mv}$
- **Uncertainty principle:** $\Delta x \cdot \Delta(mv) \geq \frac{h}{4\pi}$
- **Wave function:** a series of mathematical functions that describe the electron in the atom.
- **Probability density (electron density):** ψ^2, the probability that the electron will be found at a given point in space.
- **Orbital:** a set of wave functions that yields by the solution to Schrodinger's equation. Each orbital has a characteristic shape and energy.
- **The principal quantum number (n):** energy level, **n = 1, 2, 3, 4, ...**
- **The angular momentum quantum number (*l*):** shape (of orbital), ***l* = 0, 1, 2, 3, ...n-1**
- **The magnetic quantum number (m_l):** orientation, **m_l = interval of [-*l*, +*l*]**

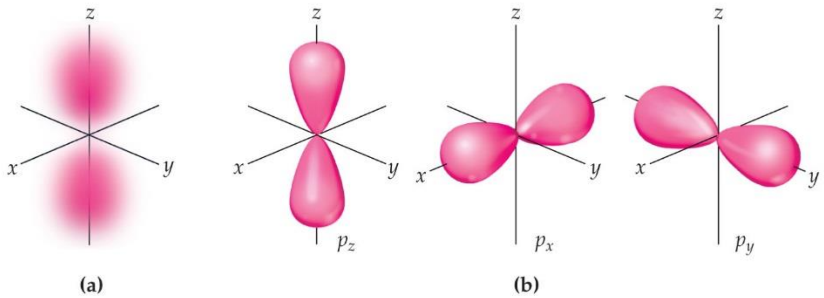
Values of n	1	2	3	4
Letter used	K	L	M	N
Values of *l*	0	1	2	3
Letter used	*s*	*p*	*d*	*f*

- **Electron shell:** the collection of orbitals with the same value of n.
- **Subshell:** the set of orbitals that have the same n and *l* values.
- ***s, p, d* Orbitals:**

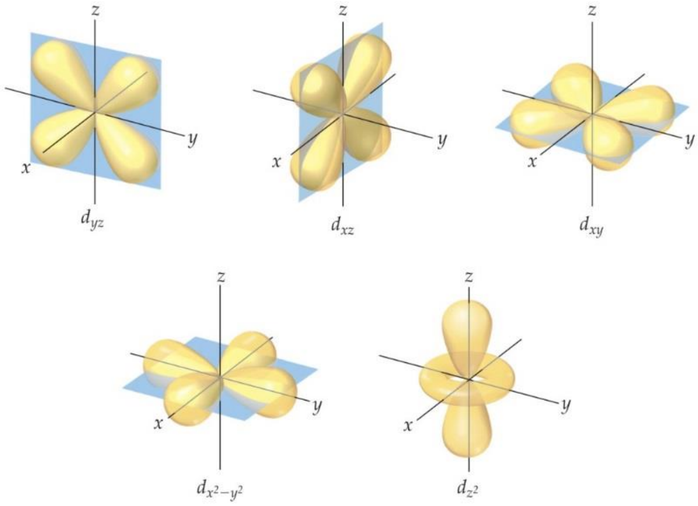




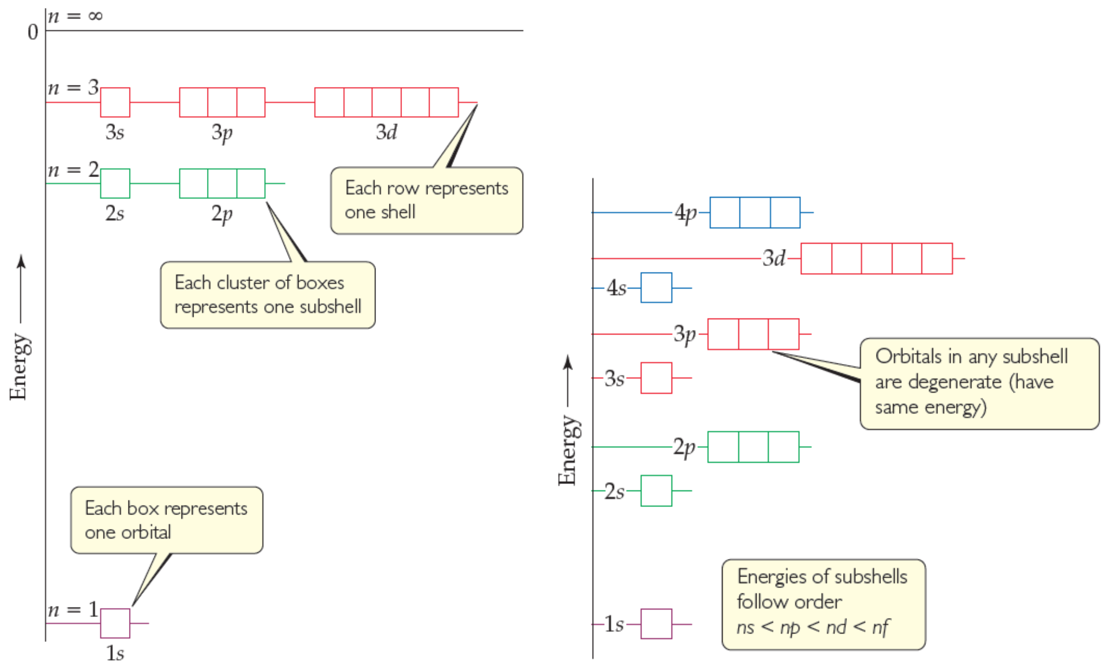
s orbitals (spherical)



p orbitals (dumbbell-shaped, the subscript *x, y, z* denotes axis along which the orbital is aligned)



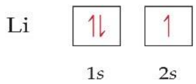
d orbitals (cloverleaf-shaped, the subscript *xy, xz, yz* denotes the four lobes lie in *xy, xz, yz* planes)



General energy ordering of orbitals for the **hydrogen atom** (left) and
a **many-electron atom** (right)

- **Electron spin:** that causes each electron to behave as if it were a tiny sphere spinning on its own axis.

- **The spin magnetic quantum number (m_s):** independent of other three quantum numbers because m_s is always = –½ or +½
- **Electron configuration:** the way electrons are distributed among the various orbitals of an atom.
  - ✓ The most stable electron configuration—the ground state—is that in which the electrons are in the lowest possible energy states.
  - ✓ The orbitals are filled in order of increasing energy, with no more than two electrons per orbital.
- **Orbital diagram:** Each orbital is denoted by a box and each electron by a half arrow**.**



- **The Pauli exclusion principle:** an orbital can hold a maximum of two electrons and they must have opposite spins.
- **Hund's rule:** for degenerate orbitals, the lowest energy is attained when the number of electrons having the same spin is maximized.

The rule: Filling up of orbitals is dependent on orbital energy while removal of electrons from orbitals is dependent on orbital location.