Chapter 6

6.1-The Wave Nature of Light

- **electromagnetic radiation** (also known as radiant energy)
  - types: visible light, radio waves, infrared radiation (heat), UV radiation, X-rays
  - all move at same **speed of light** = $2.998 \times 10^{8}\ ms^{-1}$
  - **wavelength**: distance between identical points on successive waves (such as two adjacent peaks or troughs)
    - **symbol**: λ
    - given usually in nanometers (nm): $1 \times 10^{-9}\ m = 1\ nm$
  - **frequency**: number of times per second that one complete wavelength passes a given point
    - **symbol**: ν
    - given in s^{-1} (or Hz)

inverse relationship

- equation
  - $c = \lambda\nu$

6.2 - Quantized Energy and Photons

- **blackbody radiation**: emission of light from hot objects
- **photoelectric effect**: emission of electrons from metal surfaces on which light shines (min. freq. of light [diff. for diff. metals] required for emission of e^-)
- **emission spectra**: emission of light from electronically excited gas atoms

not on AP exam

know in case for test

- **Max Planck** = proposed energy can either be released or absorbed by atoms only in discrete "chunks" of min. size
  - gave name "quantum" (meaning "fixed amt") to smallest quantity of energy that can be emitted or absorbed as electromagnetic radiation
  - matter can emit & absorb energy only in whole numbers (of $h\nu$)
  - equation

calculated for photon

(multiple by Avogadro's # to get ans. for mol of photon)

- $E = h\nu$

Planck's constant = 6.626×10^{-34} J·s

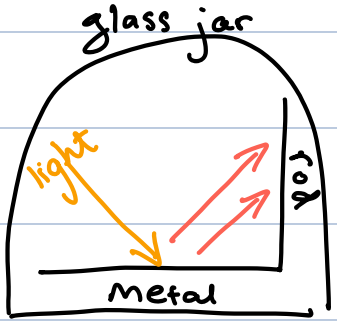
energy (J)

- $E = \frac{hc}{\lambda}$

constants

inverse relationship

- **photoelectric effect drawing**



- Einstein: assumed radiant energy striking metal surface behaves like stream of tiny energy packets
    - each packet ("particle" of energy) = photon
    - energy of photon $E = h\nu$
- work function: certain amount of energy required for electrons to overcome attractive forces, holding them in metal
    - if photons striking metal have less energy than work function, electrons can't acquire sufficient energy to escape metal

6.5-Line Spectra and the Bohr Model

- Bohr model = only works for atoms w/ one e^- (hydrogen)
  - equation
    - $\frac{1}{\lambda} = (R_H)\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$
      - positive constants ($n_2 > n_1$)
      - Rydberg's constant = $1.096776 \times 10^7\ m^{-1}$
    - $E = -2.18 \times 10^{-18}\ J\left(\frac{1}{n^2}\right)$ ← energy associated w/ e^- sitting in rings (not jumping)
    - $\Delta E = -2.18 \times 10^{-18}\ J\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$ ← energy associated when e^- jumps
      - Planck's constant × speed of light × Rydberg's constant

opposites:
- when $n_f > n_i$ = absorption ($\Delta E = \oplus$, energy absorbed)
- when $n_f < n_i$ = emission ($\Delta E = \ominus$, energy released)

- groundstate = most stable = lowest energy ($n=1$)

vs. excited state = not lowest = not stable ($n = 2, 3, 4\ldots$)

- n = is called principle quantum number

6.4 - The Wave Behavior of Matter

- Werner Heisenberg
  - Heisenberg's uncertainty principle: impossible to know simultaneously both exact momentum of e^- and its exact location in space

6.5 - Quantum Mechanics and Atomic Orbitals

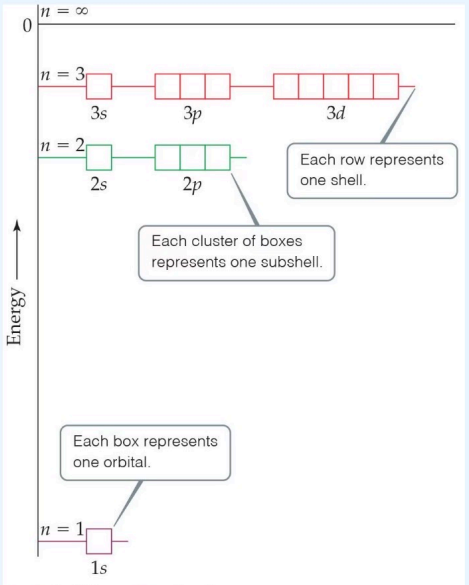
- orbitals and quantum numbers

							Period
n	principal E level						
l	sublevel	0 s	1 p	2 d	3 f		0 to n-1
m_l	orbitals	1	3	5	7		$-l$ to $+l$
m_s	spin	$\pm 1/2$					$+1/2$ $-1/2$

↑ one orbital two electrons

Figure 6.18

Energy levels in the hydrogen atom.



$n = 1$ shell has one orbital
$n = 2$ shell has two subshells composed of four orbitals
$n = 3$ shell has three subshells composed of nine orbitals

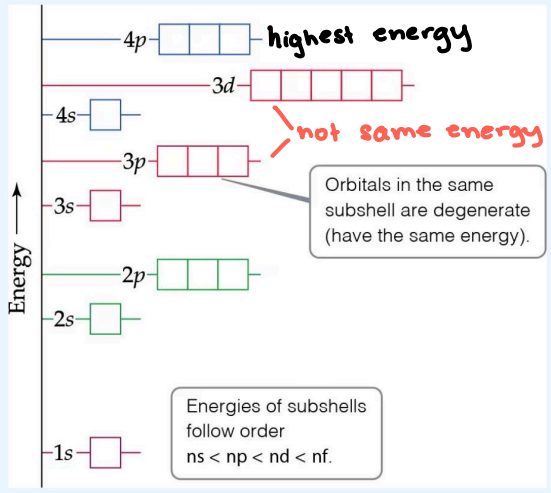
6.6 - Representations of Orbitals

- see book
- just shows how s, p, d, and f build up into respective shapes

6.7 - Many-Electron Atoms

same energy
different energy levels
vs. hydrogen (see 6.5 - Figure 6.18)

Figure 6.23
General energy ordering of orbitals for a many-electron atom.



- **Pauli exclusion principle** = no two e^- in given atom can have same set of four quantum numbers (n, l, m_l, m_s)
- **Aufbau principle** = electrons fill atomic orbitals of lowest available energy levels first
- **Hund's Rule** = put one electron with same spin on each line before going back for the other way ↑↓↑↑

6.8 - Electron Configurations

- Hund's Rule & Examples

- hydrogen: H [1] : $1s^1$ (1s)
- helium: He [1↿⇂] : $1s^2$ (1s)

Table 6.3

Electron Configurations of Several Lighter Elements

Element	Total Electrons	Orbital Diagram								Electron Configuration
		1s		2s		2p			3s	
Li	3	⇅		1						$1s^2 2s^1$
Be	4	⇅		⇅						$1s^2 2s^2$
B	5	⇅		⇅		1				$1s^2 2s^2 2p^1$
C	6	⇅		⇅		1	1			$1s^2 2s^2 2p^2$
N	7	⇅		⇅		1	1	1		$1s^2 2s^2 2p^3$
Ne	10	⇅		⇅		⇅	⇅	⇅		$1s^2 2s^2 2p^6$
Na	11	⇅		⇅		⇅	⇅	⇅	1	$1s^2 2s^2 2p^6 3s^1$

- condensed electron config: [nearest noble gas] config till atom
- transition metals = "d" (# before d = period - 1)