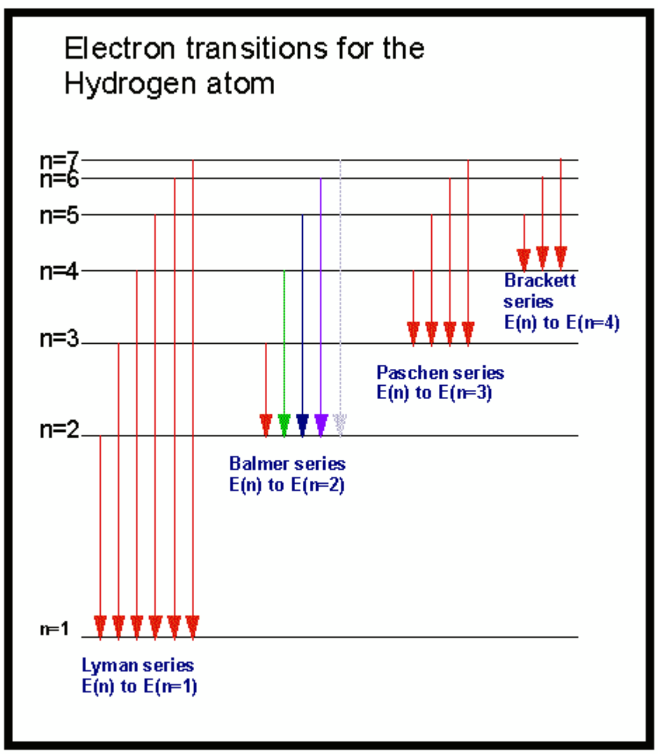
Topic III-2 and 3 Test Study Guide

II. Bohr Model of The Atom

In the Bohr model of the atom, the lone electron in the hydrogen atom can have only certain specific energies

- When the electron has its lowest possible energy, the atom is in its **ground state (Principle quantum # is n=1).**
- Excitation of the electron by absorbing energy raises the atom from the ground state to the **excited state** with ***n=2,3,4,5,6.....*** and so forth.
- A quantum of energy in the form of light is emitted when the electron drops back to a lower energy level.
- The emission occurs in a single abrupt step, called an electron transition
- The quantum energy *E* is related to the frequency of the emitted light by the equation ***E=hv,*** where **h=6.626 x 10^{-34} J·S**
- The light emitted by an electron moving from a higher to a lower energy level has a frequency directly proportional to the energy change in the electron; therefore each transition produces a line of specific frequency in the spectrum
- The figure below shows the explanation of the 3 groups of lines observed in the emission spectrum of hydrogen atoms.



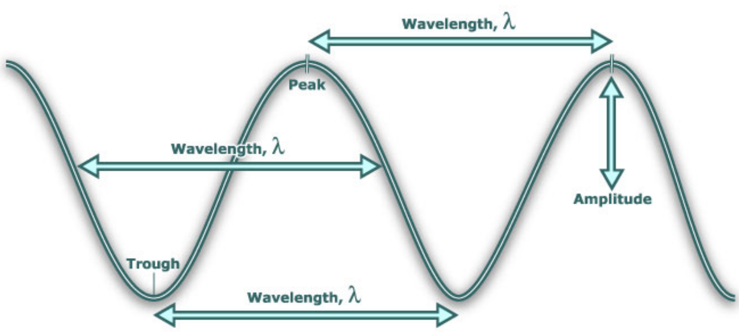
Lyman- ultraviolet (n=1)
Balmer- visible (excited state to n=2)
Paschen - infrared (excited state to

- Lyman series: lines at end of hydrogen spectrum . Transition of higher energy levels to lowest energy level n=1.
- Balmer series: lines in the visible spectrum, lines result from transitions from higher energy levels to the n=2.
- Paschen series: transitions to n=3 from higher energy levels. The energy changes of the electron and therefore the frequencies of emitted light are even smaller and are in the infrared range.
- Spectral lines for the transitions from higher energy levels to n=4 and n=5 also exist
- Bohr's theory only worked for the emission spectrum of hydrogen, but not the emission spectra of other atoms with one or more electron
- Eventually the Bohr model was replaced with the quantum mechanical model, which is based on the description of motion of material objects as waves.

A. The Nature of Light

1. **Types of spectra**
2. **Concepts of light**
   a. **Light as a wave**

Properties of Waves:



1. *Amplitude:* height of wave (from starting point). Shows energy of wave.
2. *Frequency:* # of waves that pass a certain point in a certain amount of time. **Hertz (HZ= $^1/_s$ =s^{-1})**
3. *Wavelength:* (m or nm) 1m=1 x 10^{-9} nm
4. *Speed:* how fast **3.0 x 10^8 m/s**

$$c = \lambda \nu$$

- **c**= speed of light 3.0 x10^8 m/s
- **λ**= wavelength (lambda "λ") m
- **ν** = frequency (nu) Hz or s^{-1}

$$E=h\nu$$

- **E**= energy (Joules) J
- **h=**Planck's constant **(h=6.626 x 10^{-34} J·S)**
- **ν**=frequency (Hz)

b. **light as a particle**

Light can be viewed as a stream of particles known as photons (bundles of light energy), each of which has a particular amount of energy associated with it. (Look @ equation above)

Light has characteristics of waves and particles. We visualize light as a wave that carries energy through space. Sometimes it is a stream of tiny packets of energy called photons. (**Duality of light)**

c. **electromagnetic spectrum**

<u>Electromagnetic radiation:</u> energy that is transferred from one place to another by light (infrared, microwaves, visible light, etc)

- Different types of electromagnetic radiation differ by wavelength & frequency. They all travel at the same speed (speed of light)

B. The Bohr Atomic Model

1. **Explaining the existence of the line spectra**

Look at Power point.

2. **Energy-level diagram and quantum numbers**

$$\mathbf{E_n = -\ 1312/n^2\ kJ/mole}$$

Bohr Atom Energy Level Diagram	
n=5	E=-1312/5^2= 52.48 kJ/mol= 8.72 x 10^{-20} J/atom
n=4	E=-1312/4^2= 82.00 kJ/mol=1.36 x 10^{-19} J/atom
n=3	E=-1312/3^2= 145.7 kJ/mol= 2.42 x 10^{-19} J/atom
n=2	E=-1312/2^2= 328.0 kJ/mol= 5.45 x 10^{-19}J/atom
n=1	E=-1312/1^2=1312 kJ/mol=2.18 x 10^{-17} J/atom

3. **Electron configurations :** arrangement of electrons in the atom

<u>**III. Modern Atomic Structure**</u>

A. Changing Bohr's Model

1. **Duality of mater: DeBroglie**
   - Electrons move in wave-like motion (Bohr- electrons move in circle)

2. **Heisenberg's Uncertainty Principle**
   - Impossible to know <u>both</u> the velocity and position of the electron at the same time.
3. **Matter-Waves: Schrödinger's wave concept**

- Model applies to all atoms (unlike Bohr- only applies to hydrogen)
- Probability of finding an electron in a certain volume of space around a nucleus.
- 3D areas of probability-= atomic orbitals (electron cloud)

B. Describing The Modern Atom

1. **Quantum numbers**

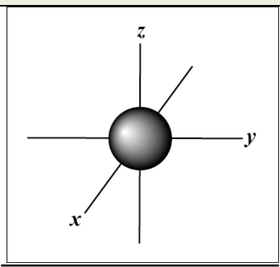
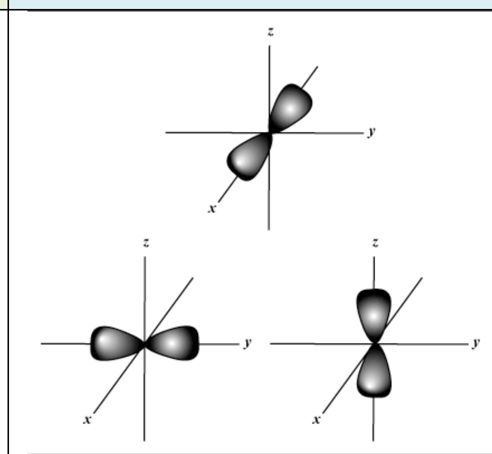
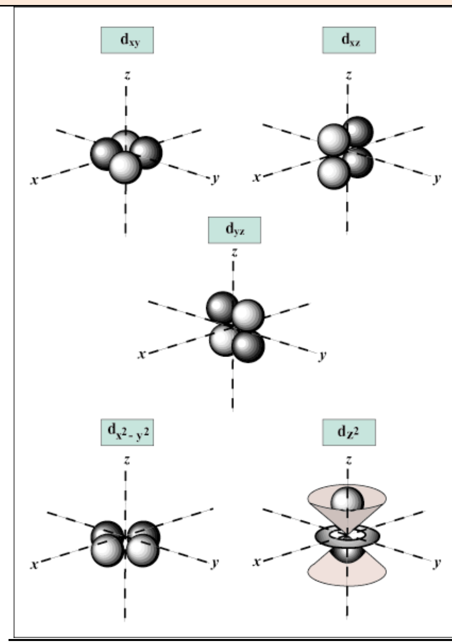
- Principle Quantum Number (n) "Energy Level"
  - n=1,2,3, 4, 5, ...
- Sublevels: describe shape/orientation of the cloud around the nucleus
  - s (sharp)
  - p (principle)
  - d(diffuse)
  - f (fundamental)

Principle Energy Level "n"	Number of Sublevels	Orbital Designation	Number of Orbitals	Number of Electrons
1	1	1s	1	2
2	2	2s 2p	1 3	2 6 **Total= 8**
3	3	3s 3p 3d	1 3 5	2 6 10 **Total=18**
4	4	4s 4p 4d 4f	1 3 5 7	2 6 10 14 **Total=32**

Name	Symbol	Possible Values	What it describes..
Principle Quantum Number	n	N=1,2,3.....	Main energy level
Angular Momentum Quantum Number	L	From 0 to n-1	Shape of electron cloud
Magnetic Quantum Number	m_L	-1 to +1 including 0	orientation of orbitals
Electron spin Quantum number	m_s	+1/2, -1/2	Direction of electron spin

Angular Momentum Quantum Number L	Sublevel	m_L
0	s	0
1	p	-1,0,1
2	d	-2,-1,0,1,2
3	f	-3,-2,-1,0,1,2,3

2. **Atomic orbitals**
   a. **shapes of orbitals**

S Orbital	P Orbital	D Orbital
		

   b. **Pauli Exclusion Principle:**
   1. An orbital can hold a max of 2 electrons
   2. 2 electrons in the same orbital must have opposite spins.

3. **Electron Configurations**

* arrangement of electron in the atom

Exceptions:

Cr

Predicted: [Ar] $4s^2 3d^4$

Actual: **[Ar]$4s^1 3d^5$**

(much more stable--> half filled)

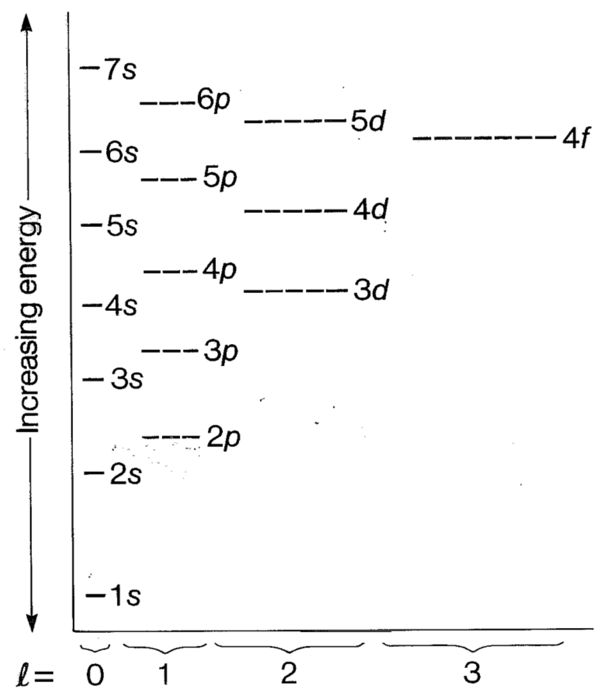
Cu

Predicted: [Ar] $4s^2 3d^9$

Actual**:** **[Ar] $4s^1 3d^{10}$**

Cu^+ [Au] $3d^{10}$
Cu^{2+} [Au] $3d^9$
Transition metals-1st electron is taken from s.

a. **Aufbau Principle:** electrons occupy orbitals of lowest possible energy first



b. **Hund's rule:** Electrons occupy orbitals of equal energy, each orbital must contain 1 electron before pairing up can occur

c. **electron blocks and the periodic table**