Unit II. Electrons in Atoms

Electromagnetic Radiation   Bohr Model   Quantum Model   Atomic Spectra   Back to Notes

PDF form: Advanced discussions of the topics are highlighted in yellow.

READING ASSIGNMENT 1: Read Section 13.3, pg 372-379. Elements can be characterized by certain colors of light they produce when heated or exposed to electricity. Eight elements and their colors were identified from the reading, list the eight and then research at least 3 other elements and add these to your list.

I. Electromagnetic Radiation- energy that is emitted from the sun which has perpendicular electric and magnetic fields.

A. Electromagnetic Radiation spectrum- the compilation of electromagnetic waves emitted from the sun

1. All have a defined wavelength (λ- lambda), the distance from one point of a wave to the similar point on the next adjacent wave.
2. All have a defined frequency (ν- nu), the number of wave cycles that occur in 1 second.
   measured in hertz- (cycles/sec)

APPLET: ELECTROMAGNETIC WAVES. This shows the different types of waves based on wavelengths
TABLE: Wavelengths Frequencies & Energies of common electromagnetic waves.

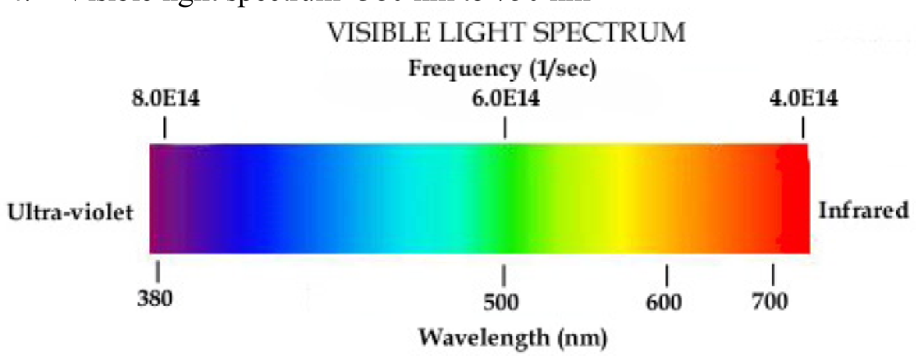
3. The product of wavelength and frequency is a constant,
   (c = constant for light = 2.998×10^8 m/s ~ 3.0×10^8 m/s)

Solving for wavelength & Frequency
$c = \lambda \bullet \nu$ so: $3.0 \times 10^8 \text{ m/s} = \lambda \bullet \nu$

QUESTION: What is the relationship between wavelength and frequency? Click here to see.
** wavelength and frequency are inversely proportional**

QUESTION: How was c determined? Michelson-Morley Experiment. Experimental evidence for the speed of light.

4. Visible light spectrum- 380 nm to 750 nm



B. Photons- a characteristic of the energy associated with electromagnetic radiation

1. Max Planck. (1858-1947). Developed the concept of the Quantum/Photon.
   Quantum ("fixed amount")- the smallest, discrete packet of energy that is either released or absorbed by atoms.
   a. Released energy is always an integer product of (h ν)
   b. A third characteristic of waves (frequency & wavelength) is amplitude, the intensity of a wave. This is measured as the vertical height of a wave peak.

the vertical height of a wave peak.

Calculating the Energy of a Photon
$E = h\ \nu\ = (h\ c) / \lambda$
where h is Planck's constant (6.626×10^{-34} Js). What is a Js?

IN-CLASS PRACTICE: Calculating wavelength, frequency & energy.
ON YOUR OWN-PRACTICE: Electromagnetic Radiation Quiz:
ASSIGNMENT 2: Read p. 378 and answer questions A-D at the bottom.

C. Photoelectric Effect.
Albert Einstein developed the concept of electrons being liberated from atoms by absorbing photons.
(The link to Albert Einstein contains many pictures, quotes and some audio of his speeches)

READING ASSIGNMENT 2: Read p361-370. Answer questions 1-10 pgs 366 & 370.

II. Modern Theories of the Atom

A. Neils Bohr-(1885-1962)- 1913- Bohr Model of the Atom - Planetary Model
Discussion of Bohr Model
-Used Hydrogen to determine position of electrons around nucleus. (single electron)
Stated that energy in electrons is quantized- discrete individualized values. Therefore electrons were restricted to certain orbits about the nucleus, where orbits of greatest distance contain greater energy.
Quantum- the amount of energy needed for an electron to "jump" from one orbit to another.
Energy Level- Another name for orbit.

Calculating the energy of an electron in the (n) energy level.
$E = (-R_H Z^2) / n^2$ Where R_H is the Rydberg constant, Z is the atomic number, and n is the energy level

Ex. Find the energy of 1 electron in 1st energy level of Hydrogen: [$R_H = 2.179 \times 10^{-18}$ J and $Z = 1$]
typically- Electron-volts are used calculating energies of electrons: 1 eV = 1.602×10^{-19} Joules

Calculating the radius of an energy level
Radius of orbit = $(n^2 a_o) / Z$ Where n= energy level, a_o=radius when n=1 (0.529 A) and Z is the atomic number

Energies and Radii for subsequent energy levels:			
N	Energy Level	Energy (eV)	Distance (A)
1	K	-13.595	0.529
2	L	-3.399	2.116
3	M	-1.511	4.761
4	N	-0.850	8.464
5	O	-0.544	13.225
infinity	--	0.000	infinity

B. Quantum Model of the Atom

- Based on the mathematical theory of Erwin Schrodinger- Austrian (1887-1961)
- Describes the relative position of electrons based on probabilities- orbitals- 3 dimensional
- The orbitals are less defined than Bohr's orbits, but still remain quantized.
- The shape of the orbital is described where a high probability of finding an electron exists. 90%

.Example: 1st orbital of Hydrogen- electron can exist up to 1 angstrom away from nucleus but has the highest probability at 0.529 angstroms. This describes a spherical cloud.

1. Electron density- high probabilities create a greater density whereas lower probabilities create a lower density. Nodes are regions that separate electron dense areas.
2. Principles of the Quantum Theory- (Development of Quantum Numbers)

a. The location of an electron cannot be determined exactly. Orbitals describe a volume where there exists a higher probability of finding an electron

b. Orbitals can be defined by a ***Principal Quantum Number*** (n)- which describes the relative distance of the orbital from the nucleus (integer values)

-Energy level or Shell- terms used to describe (n)

Energy Level Number	1	2	3	4	5	6	7
Symbol	K	L	M	N	O	P	Q

-Maximum number of electrons that can fit into an energy level is $2n^2$.

c. Orbitals with the same (n) may have different shapes. ***Subshell Quantum Number*** (l)- Also called Subsidiary/Azimuthal Quantum Number.

-include all integer values from l = 0 to l = n-1.

-The l value defines the number of nodal planes- a plane which has a probability of zero for finding an electron

Different shapes of subshells/ orbitals		
l value (# of nodal planes)	orbital	shape
0	s-type	spherical
1	p-type	dumbbell
2	d-type	4-lobed / other
3	f-type	8-lobed / others
4	g-type	(???????)
5	h-type	(???????)

Comparative energies of the orbitals: s-types < p-type < d-type < f-types < …….

Possible orbitals for each Shell			
Shell	n-value	Possible orbitals (l)	Types of orbitals
K	1	0	1s
L	2	0,1	2s, 2p
M	3	0,1,2	3s, 3p, 3d
N	4	0,1,2,3	4s, 4p, 4d, 4f
O	5	0,1,2,3,4	5s, 5p, 5d, 5f, 5g

d. For l values greater than zero. Orbitals with same l value have the same energy but different orientations- ***Magnetic Quantum Number*** (m)- where m = -l to +l.

-degenerate orbitals- orbitals with the same l value- Subshell.

e. Electrons in orbitals will have ***Spin Quantum Number*** (s)- which defines either a clockwise or counterclockwise spin of the electrons (s = + ½ or – ½)

- Parallel spins- electrons with same spin
- Only 2 electrons can exist in the same orbit- need to have opposite spins.

A table describing quantum numbers
Schroedinger's Model of the atom. Electron's are more wave-like in motion than particle like.

APPLET: David's Whizzy Periodic Table. Shows the electrons in their appropriate energy levels and orbitals
IN CLASS PRACTICE: Quantum Numbers

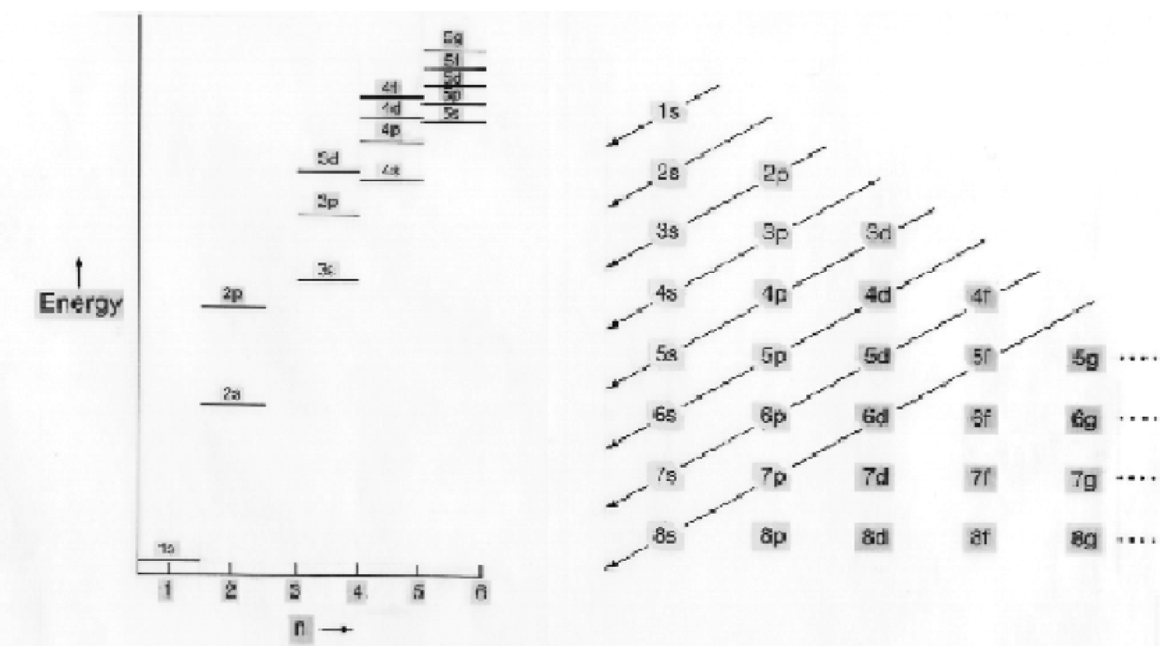
ON YOUR OWN-PRACTICE: Quantum numbers & orbitals. There are 6 quizzes (Quantum Numbers through Shells, Subshells & Orbitals).

QUESTION: How does this apply to atoms other than Hydrogen? Click here to begin the tutorial

3. Electron Configuration – the arrangement of electrons in the orbitals of an atom.
   a. Described by 3 different terms
      1. Number describing the principle
      2. Letter describing orbital type (subshell)
      3. Superscript describing the number of electrons it the subshell.

TUTORIAL. Electron Configurations of the elements.
IN-CLASS PRACTICE: Electron Configurations
ON YOUR OWN-PRACTICE: Electron configurations quiz.
ON YOUR OWN-PRACTICE: Electron configurations & the periodic table quiz.

b. Rules governing where an electron can fill
   1. Aufbau- (building up)- electrons enter the orbitals with the lowest possible energies.
   Building up the Periodic Table



2. Pauli's Exclusion Principle- (Wolfgang Pauli) No two electrons can have the same set of 4 quantum numbers. For 2 electrons to share the same orbital they must have opposite spins.

3. Hund's Rule- When filling a set of degenerate orbitals, the number of unpaired electrons will be maximized and these electrons will have parallel spins.

4. Exceptions to Electron Configurations- [Cr, Cu, Nb, Mo, Ru, Rh, Pd, Ag, Pt, Au]
   - Half and full-filled d orbitals are preferred over full s-type and partially filled d-filled orbitals
   - Caused by interplay between pairing energy and promotion energy. Usually between s & d, and s, d, & f types.

ON YOUR OWN-PRACTICE:: Electron configurations exceptions quiz. Print off the results and include your answers/work

C. Atomic Spectra

1. States:
   a. based on potential energies
      1. Ground state- lowest energy an electron can have
      2. Excited state- an electron at a higher energy level due to quantization

Tutorial on Spectral Lines

b. Series
   1. Lyman- where n1 = 1 ultraviolet spectrum
   2. Balmer- where n1 = 2 visible light spectrum. Balmer Formula
   3. Paschen- where n1 = 3 infrared spectrum

   Each series has a maximum frequency : additional absorbed energy would liberate the electrons from the atom

Line Spectra of Hydrogen Atom

Robert Bunsen (namesake of the Bunsen Burner) is credited with being the father of modern spectroscopy.

Balmer's & Kirchof's experiments on the reversibility of atomic emission specta
Emission Spectra is derived when electrons move from a higher energy level to a lower and release energy
Absorption Spectra is derived when electrons absorb specific wavelengths, moving electrons to higher energy levels, while other wavelengths are allowed to pass through the atoms

APPLET: Atomic Emission & Absorption Spectra for the elements.

c. Calculating the energy of spectral lines is dependent upon n_i and n_f, initial and final energy levels, respectively.

IN-CLASS PRACTICE: Calculating energy for spectral emissions

Calculations of Atomic Spectral Lines (Hydrogen specific)
$\nu = (R_H/h)(1/n_i^2 - 1/n_f^2)$
$E = (R_H)(1/n_i^2 - 1/n_f^2)$

APPLET: Finding ground state & excited states (by orbitals) for spectral lines.

D. Heisenberg's Uncertainty Principle

It is impossible to determine accurately both the velocity and position of a particle simultaneously.
Gives rise to calculating probabilities of location….. orbitals

Ch 13 Prac Quests: 386: 22, 26, 28, 30, 31, 35, 36, 38, 40, 46, 49, 51 & 53

ASSIGNMENT 3: Ch. 13 Rev. Ques. p386: 22, 26, 28, 30, 31, 35, 36, 38, 40, 46, 49, 51, & 53.

Other Interesting Links:
Link to Nick Strobel's page on Electromagnetic radiation and the Bohr Model from a astronomers point of view.
The Quantum Concept: An in-depth analysis of the quantum concept and the questions that had to be asked.
What are the questions that can be asked? Here are questions to check for understanding.
The tutorials in this site are part of the Physics 2000 program at the University of Colorado at Boulder
ON YOUR OWN-PRACTICE: Practice Games over Atomic Structures.