Atomic Structure

Electromagnetic Spectrum

Frequency x wavelenth = c

$$\nu\lambda = c$$

C = speed of light = 3.00 x 10^8 m s^{-1}

http://www.colorado.edu/physics/2000/waves_particles/lightspeed-1.html

$\Delta E = h\nu$

- h = Plank's constant 6.63×10^{-34} J-s
- ν = frequency

- Therefore, you can find energy with these:
  - If $\nu\lambda = c$ and $\Delta E = h\nu$
  - Then $\nu = c/\lambda$ and $\Delta E = hc/\lambda$

DeBroglie's Hypothesis

- The characteristic wavelength of the electron or of any other particle depends on its mass, m, and velocity, v.
- $\lambda = \frac{h}{mv}$
- mv = momentum
- $1J = 1kg\ m^2/s^2$

Rydberg equation

- $\Delta E = 2.18 \times 10^{-18}\ J\left(1/n^2_{initial} - 1/n^2_{final} \right)$

$R_H = 1.10 \times 10^7 / m$

- When $n_f > n_i$ energy is absorbed
- When $n_i > n_f$ energy is emitted

- Note: Can change R_H to negative if changed to final – initial
- The more negative R_H the more stable

- Hydrogen
  - Most stable if n=1
  - If n > 1 the atom is said to be electronically excited

Quanta

- An object can gain or lose energy by absorbing or emitting radiant energy
- The smallest increment of energy at a given frequency, $h\nu$, is called a QUANTUM of ENERGY
- A photon is a quantum of energy
- Light is packaged in units called QUANTA

Spectra

- Radiation made of one wavelength is called MONOCHROMATIC
- Most common radiation sources produce a number of different wavelengths which can be separated into a SPECTRUM
- A rainbow is a CONTINUOUS SPECTRUM
- Atomic Absorption and Emission Spectra

Line Spectra

- Some sources produce only a line spectrum of specific wavelengths

- Elements have characteristic LINE SPECTRA

- BLUE – higher energy - shorter wavelength
- RED – lower energy - longer wavelength
- ROYGBIV goes low to high energy

Wave functions

- Standing Wave

It is important to note that electrons behave as BOTH wave and particles.

This is called a wave-particle duality.

BOHR's MODEL

- An electron in a permitted orbit is said to be in an "allowed" energy state
- Without radiant energy an electron remains in an allowed energy state
- With radiant energy the electron changes from one allowed state to another
- $\Delta E = h\nu$

HEISENBERG's UNCERTAINTY PRINCIPLE

- It is impossible to know BOTH the momentum and location of the electron simultaneously
- How is this a contradiction to Bohr's model?

Quantum mechanics

- Don’t be AFRAID

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- Describes mathematically the wave properties of the electron
- ψ Psi – wave functions

ψ^2 = zero is called a NODE

# of nodes increases with increasing n

ψ^2 Probability that an electron will be found at that location

Electron clouds or electron density

Principle quantum number

- n = allowed orbit
- n is the principle quantum number and it is the same as the energy level
- The radius of the electron orbit in a particular energy state varies as n^2
- The number of electrons in an energy level $=2n^2$
- RADIUS = n^2 (5.30×10^{-11}m) Bohr's radius

Quantum numbers

- (1) **Principle quantum number**
  = n (energy level)

- (2) Azimuthal or angular quantum number
  - = **l**
  - Defines the shape of the orbital
  - Cannot be larger than n-1

- (3)Mangetic quantum number = m_l
  - Describes the orietation in space
  - Along the x, y, or z axis

- (4)Magnetic spin
  + ½ or – ½

4 numbers to identify the specific electron

Example: 4, 0, 0, 1/2

Stuff

- Energy increases with increasing n
- Pauli Exclusion Principle
  - No two electrons in an atom can have the same set of four quantum numbers n, l, m_l, m_s

  - none have the same address
- Electron Spin
- Hund's Rule
- Hund's Rule
  - School bus
  - Electrons fill degererate orbitals (same energy) one at a time and then go back to put in second electron
- Electron spin
  - Designated by positive or negative one/half
  - Spin in opposite directions
  - Give balance

Electron Configurations and the Periodic Table

- Review ending electron configurations
- Families
- Groups
- Exceptions in d-block