Atomic Theory

1. Electromagnetic Radiation

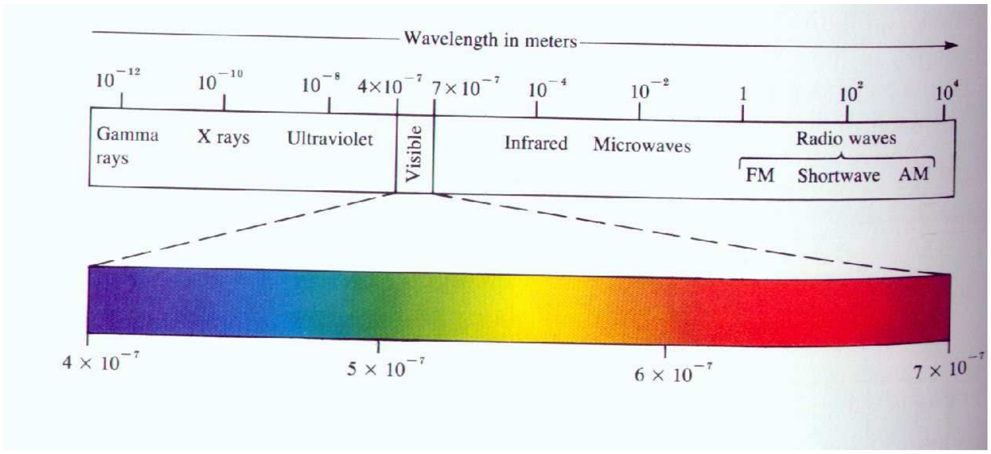
3 characteristics of waves:

a.) Wavelength, (λ): distance between two consecutive peaks. Units of meters, m.

b.) Frequency, (ν): number of waves (cycles) per second that pass a given point. Units of Hertz, Hz, or 1/s = s^{-}

c.) Velocity, (c): at or near the speed of light which is 2.9979 $x10^{8}$ m/s

Electromagnetic spectrum shows the relationship of wavelength to frequency at constant velocity. This relationship can be expressed as: **$c = \lambda \nu$**



Inspect and memorize the spectrum.

2. The Nature of Matter

The dual nature of matter suggests that there is no division between the macroscopic particle world of ordinary objects at ordinary speeds with the mass-less light-energy wave properties.

Max Planck (black body radiation) stated that energy can not be absorbed or emitted in any amount but in discrete whole numbers. *Planck's constant = 6.626x10^{-34} J/Hz = h*. Energy can only occur in discrete quantities "quanta", $h\nu$. A packet of energy is called a quantum (or Einstein called it photon).

Energy can only come in multiples of a ***quanta***, it has a particulate nature. $\Delta E = h\nu$

Example:

Calculate and compare the energy of a photon of blue light (λ = 430 nm) light with red light (λ =680 nm).

Example for blue light:

Convert nm to m (n = 10^{-9})

Find the frequency: $\nu = c/\lambda$, units of Hz

Calculate the Energy: E = $h\,\nu$ = (6.626x10^{-34} J/s^{-}) (3.0x10^{8} m/s/ 430 x10^{-9}m) = 4.6x10^{-19} J

Energy has mass: **$m = h/\lambda v$** from deriving ($E=mc^2$), here v=velocity.

Does matter exhibit wave properties?

The de Broglie's Equation: **$\lambda = h/m\,v$** allows for the calculation of the wavelength of a particle.

Derived from $E=mc^2 = h\nu$, replace c=speed with v for velocity

Substitute n with v/λ from c=$\lambda \cdot \nu$

$h\frac{v}{\lambda} = mv^2$ and $\lambda = \frac{hv}{mv^2} = \frac{h}{mv}$ h= Planck's Constant, m in kilograms, v is speed (m/s)

Example:

Compare the wavelength of an electron (mass = 9.11x10^{-31} kg) traveling at 1.0x10^{7} m/s with that of a ball (mass = 0.10 kg) traveling at 35 m/s.

$\lambda = h/$ m v

electron = 7.27 x10^{-11} m

ball = 1.9x10^{-34} m

Conclusion: macroscopic world events have extremely short wavelengths. Impossible to verify experimentally. Classical Newtonian physics explain these phenomena's, but this type of treatment is inadequate for describing the atom. Small particles of matter (i.e. electrons) at or near the speed of light require quantum mechanics to understand their behavior. All matter exhibits both wave and particulate properties.

3. The Atomic Spectrum of Hydrogen

Continuous Spectrum: all colors of the spectrum are visible. Source white light.
Emission Spectrum: Only certain lines are visible, those energies corresponding to allowed energies within the (H) atom. As electrons change energies from Excited state to a lower or ground state (exothermic). Contrast steps (quanta) vs. ramp (continuous). *Closer to the nucleus, the electron is more tightly bound and is more negative energy, more stable. At an infinite distance, ∞, the energy is zero. As the electron is brought closer to the nucleus, energy is released from the system*.

4. The Bohr Model

Electrons move around the nucleus only in allowed circular orbits. Classical physics would then predict this atom to release energy and eventually collapse on itself.

Equation to describe energy levels (n) available to the H atom:

$$E = -2.178x10^{-18}\,\mathrm{J}\left(\frac{Z^2}{n^2}\right)$$

Z= nuclear charge, Ex. H =1

Can be used to describe the change in energy from one orbit to another.

$$\Delta E = -\,2.178x10^{-18}\,J\left(\frac{Z_2^2}{n_2^2} - \frac{Z_1^2}{n_1^2}\right)$$

Example: Calculate the amount of energy required to excite the hydrogen electron from its 1st to 2nd energy level. Also calculate the wavelength of light associated with this energy increase.

$$E_1 = -2.178x10^{-18}\,J\left(\frac{1^2}{1^2}\right) = -2.178x10^{-18}\,J$$

$E_2 - E_1 =$ **$1.63x10^{-18}$ J**

$$E_1 = -2.178x10^{-18}\,J\left(\frac{1^2}{2^2}\right) = -5.445x10^{-19}\,J$$

To calculate the wavelength of light: $E/h = \nu$, $\lambda = c/\nu$ **= 1.22x10-7 m**

The Bohr atom is fundamentally incorrect and cannot be used for other atoms. It does show the quantized energy of the atom.

Homework for Study: (7.1-4) P. 321 #33, 34, 35, 37, 38, 40, 42, 45, 47, 50, 53, 54, 111, 113, 115