10/21/05

CHM 123- Lecture (Friday 10:30)

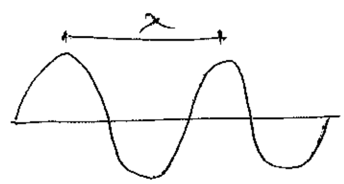
Average on Exam 2- 74%

Chapter 7- Electronic Configuration and properties

Electromagnetic Radiation

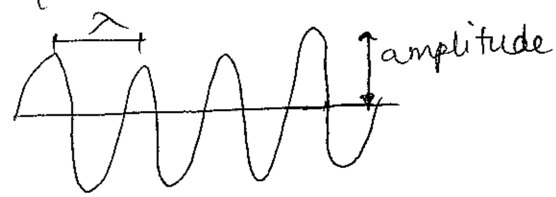
- visible light, ultraviolet, infrared, x-rays etc.
- oscillating perpendicular electric and magnetic fields.
- travel through space at 2.998×10^8 m/s.
- described in terms of frequency and wavelength.

(a)



Two complete cycles of wavelength λ.

(b)



Frequency twice as much as (a).

- $\nu\lambda = c$
- λ = wavelength

  c = speed of light

  ν = frequency ⇒ units are Hz (Hertz).

Refer to table 6.1 in text.

Planck's Quantum Theory

Observation: heat a solid - emits light with a wavelength distribution that depends on the temperature.

- vibrating atoms in hot wire caused light emission whose color changed with T.
- Classical physics could not explain this – the wavelength should be independent on T.

- Max Planck (1900)
  - Assume energy can be released or absorbed by atoms in packets only.
  - these packets have a minimum size.
  - Packet of energy = quantum
  - Energy of quantum ∝ frequency of radiation ($E = h\nu$)
  - h = Planck's constant = 6.626×10^{-34} Js
  - Energy is emitted or absorbed in whole # multiples ($1h\nu$, $2h\nu$ etc).

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

The Photoelectric Effect

- Einstein assumed light striking metal was stream of tiny energy particles.
- Each energy packet behaves like tiny particle of light (photon).
- $E = h\nu$ radiant energy is quantized.
- Dilemma: is light a wave or a ~~[illegible]~~ particle?

Answer: Both. Behavior of Electrons, atoms can be explained only if they are considered to have the properties of both waves & particles.

Line Spectra

- Most radiation sources emit light in continuous spectrum (e.g. white light).
- When voltage applied to gaseous elements – not continuous spectrum but line spectrum.
  - spectrum containing radiation of specific λ.

Electronic transitions in H atom

Line spectrum of H atom contains 4 lines (Balmer Series).

$$\nu = c\left(\frac{1}{2^2} - \frac{1}{n^2}\right) \qquad n = 3, 4, 5, 6 \qquad c = 3.289 \times 10^{15}$$

Bohr Model

- ~~orbits~~ original picture of atom – electrons orbiting nucleus.
  - classical physics: such an e^- would continuously lose energy by emitting radiation and ultimately spiral into nucleus (doesn't happen).
- Bohr – only orbitals of certain radii (and therefore certain energies) are permitted.
- e^- in a permitted orbit has specific energy ("allowed" energy state).
- Allowed orbits have specific energies.

$$E_n = (-R_H)\left(\frac{1}{n^2}\right) \qquad n = 1, 2, 3, 4, \ldots \qquad R_H = 2.18 \times 10^{-18}\ \text{J}$$

(R_H → Rydberg constant)

- Each orbit – different n value
- Radius increases as n increases $(r \propto n^2)$
- 1st allowed orbit – n = 1, radius = 0.529 Å (Bohr radius)

- Energies of electrons negative for all values of n.

$n \rightarrow \infty$: (electron is no longer bound to atom, free electron)

$$E_n = (-2.18 \times 10^{-18} J)\left(\frac{1}{\infty^2}\right) = \underline{\underline{0}}$$

- Bohr e^- "jump" from one energy state to another by absorbing or emitting photons of certain ν.
- This ν corresponds to E difference between the 2 states.
- $\Delta E = h\nu$
- $$\nu = \frac{\Delta E}{h} = \left(\frac{R_H}{h}\right)\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

'n' initial → n_i; final state → n_f.