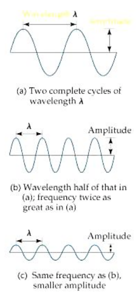
Electron Configurations & the Periodic Table

Chapter 7

Electromagnetic Radiation

visible light, ultraviolet, infrared, X-rays, etc.

- oscillating perpendicular electric and magnetic fields
- travel through space at 2.998 x 10^8 m/s
- described in terms of frequency and wavelength
- $\nu\lambda = c$

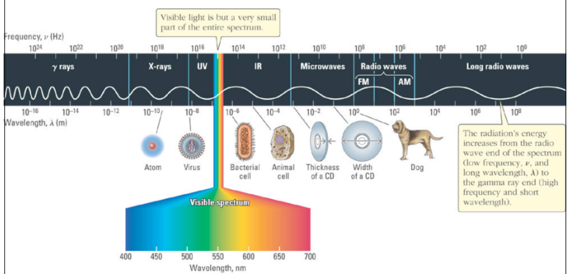


Wavelength Units

TABLE 6.1 Common Wavelength Units for Electromagnetic Radiation

Unit	Symbol	Length (m)	Type of Radiation
Angstrom	Å	10^{-10}	X ray
Nanometer	nm	10^{-9}	Ultraviolet, visible
Micrometer	m m	10^{-6}	Infrared
Millimeter	mm	10^{-3}	Infrared
Centimeter	cm	10^{-2}	Microwave
Meter	m	1	TV, radio

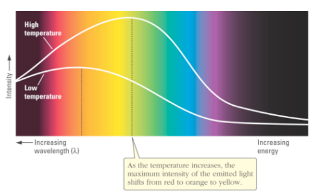
Electromagnetic Spectrum



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 of T



(b)

Planck's Quantum Theory

- 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 a quantum is proportional to the frequency of radiation
  - ($E = h\nu$)
  - h = Planck's constant = 6.626×10^{-34} Js
  - energy is emitted or absorbed in whole # multiples (1 $h\nu$, 2 $h\nu$, etc.)

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

Example 1

What is the energy (in kilojoules per mole) of photons of radar waves with $\nu = 3.35 \times 10^8$ Hz?

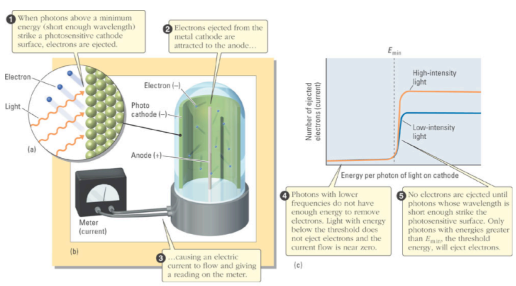
$E = h\nu$

$E = (3.35 \times 10^8 /s)(6.626 \times 10^{-34}\ J\ s) = 2.22 \times 10^{-25}\ J$

$E = (2.22 \times 10^{-28}\ kJ)(6.022 \times 10^{23}\ photons/mol)$

$E = 1.34 \times 10^{-4}\ kJ/mol$

The Photoelectric Effect

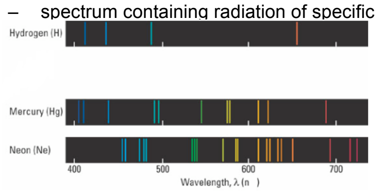


The Photoelectric Effect

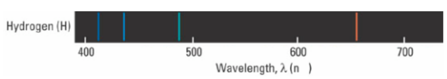
- Einstein assumed light striking metal was stream of tiny energy packets
- Each energy packet behaves like tiny *particle* of light (photon)
- $E = h\nu$ **radiant energy is quantized**
- Dilemma – is light a wave or a particle???

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)

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

How could this very simple formula be explained??

Bohr Model

Original picture of atom – e^- orbiting nucleus

- classical physics – such an e- would continuously lose energy by emitting radiation and ultimately spiral into nucleus (doesn't happen)
- Bohr – only orbits 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:

$$\mathbf{E_n} = (\mathbf{-R_H})\left(\frac{\mathbf{1}}{\mathbf{n^2}}\right)$$

n = 1, 2, 3, 4, . . .
$R_H = 2.18 \times 10^{-18}$ J

Bohr Model (cont'd)

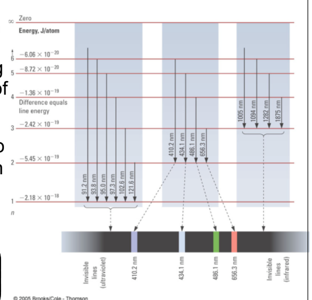
- each orbit – different n value
- radius increases as n increases ($r \propto n^2$)
- 1st allowed orbit – n = 1, radius = 0.529 Å
- Energies of electrons negative for all values of n.

  n → ∞:

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

Bohr Model (cont'd)

- 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 two states



$\Delta E = h\nu$

$$\nu = \frac{\Delta E}{h} = \left(\frac{R_H}{h}\right)\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$