Chapter 7 Periodicity

Electronic structure--refers to # of electrons in atoms or ions
as well as their distribution around the nucleus

understanding of electronic structure comes from light
emitted or absorbed by substances

Electronic radiation:
- carries energy through space so is known as radiant energy
- visible light only one can see bec. of chem rxns in our eyes
- all move through vacuum at speed of light (3.0×10^8 m/s or 3.0×10^{10} cm/s)
- all have wavelike characteristics
  - have wavelength (λ)
  - have frequency (Hz or s^{-1})
  - λ and s^{-1} are inversely proportional

$c = \lambda \times \nu$

Speed = wavelength x frequency

Max Planck said energy can be absorbed or released in "chunks"
called the chunks quanta

proposed that energy (E) of single quantum equals a constants times its frequency

$E = h\nu$ $h = 6.626 \times 10^{-34}$ J*s

J·s

Einstein used Planck's quantum theory to explain the photoelectric effect (light shining on a metal surface causes the surface to emit electrons)

assumed that energy stiking metal surface is stream of timy energy packets (photon) behave like particles

deduced each photon must have energy propotional to frequency of light

$E = mc^2$

certain amt. of energy is required for an electron to overcome attractive forces that hold it w/in the metal

if photons have less energy the electrons don't have sufficient energy to escape from metal surface

if photons have more than enough energy to free electrons from metal surface the excess appears as KE of emitted electrons

Dual nature of light:

1. light behaves macroscopically like a wave but it consists of a collection of photons

2. when you look at light at atomic level see particle like properties

de Broglie equation:

$\lambda = \frac{h}{mc}$ → Planck's 6.626 x 10^{-34} J·s

mass of part. mc → speed of particle

said electron in its movement around nucleus has a specific wavelength

wavelength of the particle depends on its mass and velocity

wave characterisitcs of electron led to development of electron microscope

electron micrographs demonstrate tiny particles of matter can behave as waves

Planck & Einstein helped with understanding how e^- are arranged in an atom

Neils Bohr-used quantum model of H atoms-

bright line spectra- contains only radiation of specific wavelength with blank areas in between

continuous spectrum-all colors of light shown with no dark areas in between

like a rainbow

assumed e^- moved in orbits around nucleus & have certain amt. of energy

$$E = -2.178 \times 10^{-18} J(z^2/n^2)$$

calculate energy of e^-

E = energy of electron

z = nuclear charge – # p^+ - atomic #

n = corresponds to the energy level

the larger the value > orbit radius

neg. sign means energy of e^- bound to nucleus is lower than it would be if e^- were at an infinite distance from nucleus where would be no interactions

at infinite distance E = 0

also means that atom has lost energy and is more stable

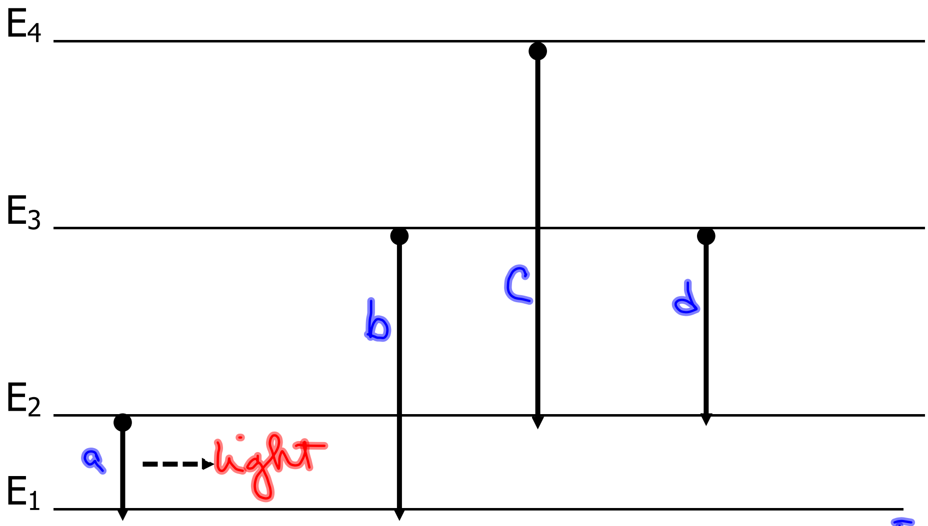
energy is removed by production of a photon

ground state--lowest energy state

Light is emitted when electrons move from higher energy level to a lower energy level

energy of emitted light is

$E = E_{final} - E_{initial}$



$E_a = E_2 - E_1$

$E_b = E_3 - E_1$

$E_c = E_4 - E_2$

$E_d = E_3 - E_2$

energy level picture of atomic energy
lower energy level is preferred

Bohr's model important-- introduces 2 ideas

1. electrons exist only in discrete energy levels which are discribed by quantum numbers
2. energy is involved in moving electron from one level to another