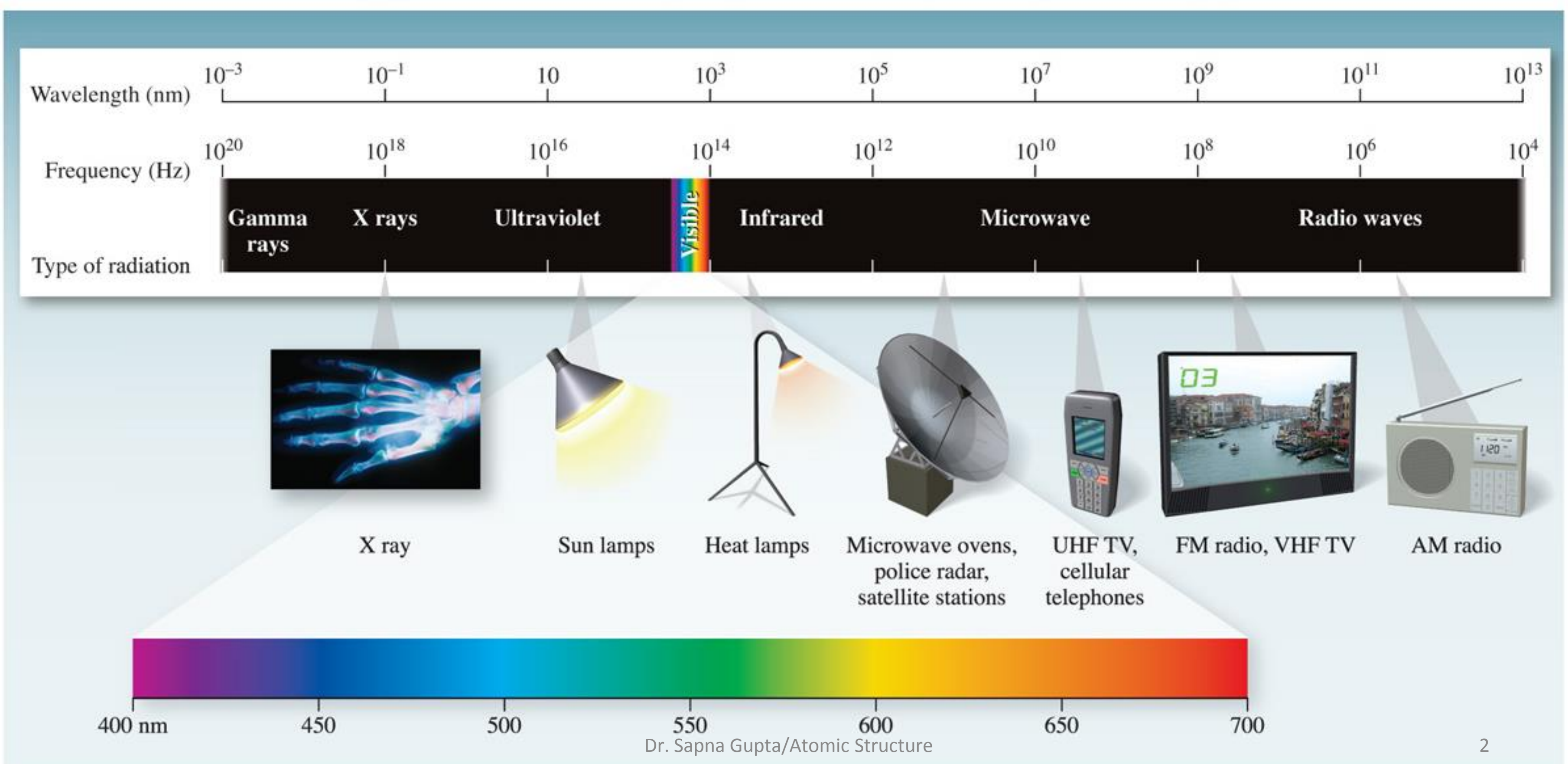
Chapter 7
Atomic Structure -1
Quantum Model of Atom

Dr. Sapna Gupta

The Electromagnetic Spectrum

- The electromagnetic spectrum includes many different types of radiation which travel in waves.
- Visible light accounts for only a small part of the spectrum
- Other familiar forms include: radio waves, microwaves, X rays



Wave Nature

- **Wavelength**: λ (lambda) distance between identical points on successive waves…peaks or troughs
- **Frequency**:ν (nu) number of waves that pass a particular point in one second
- **Amplitude**: the vertical distance from the midline of waves to the top of the peak or the bottom of the trough
- Wave properties are mathematically related as:

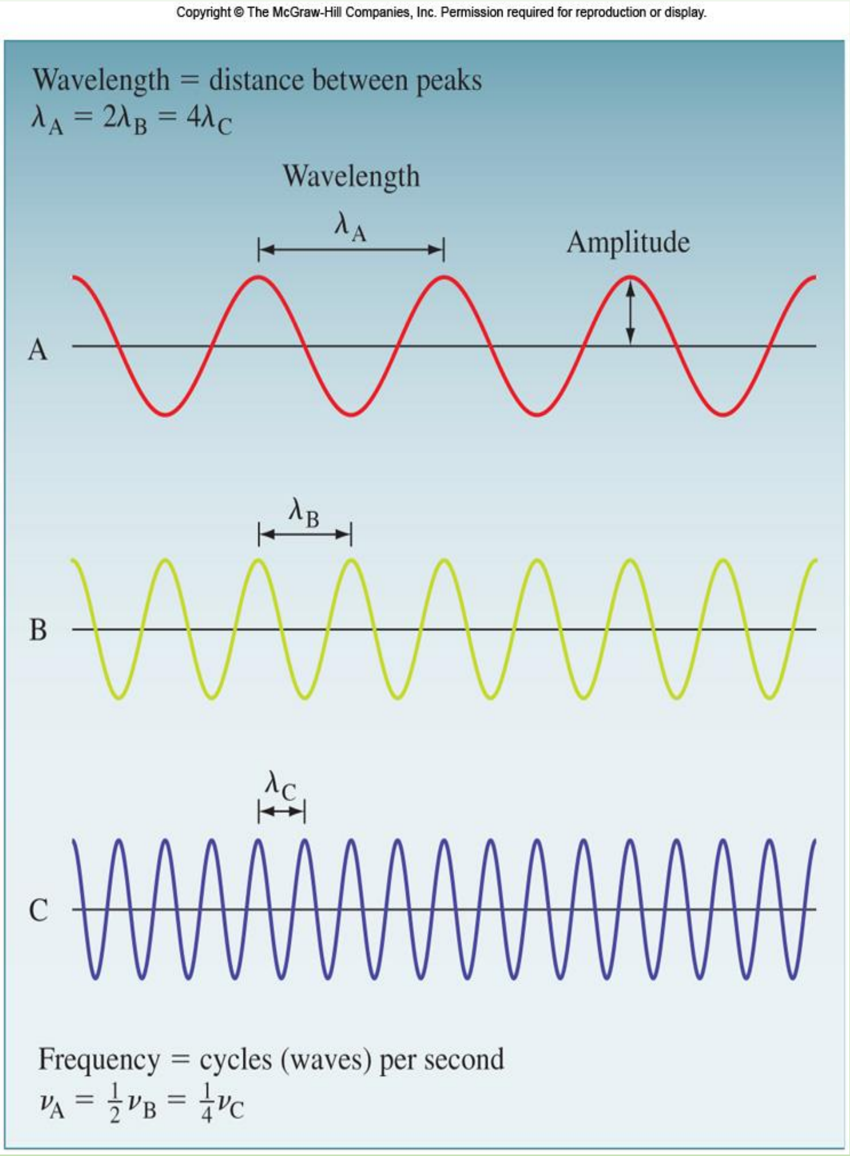
$$c = \lambda\nu$$

where

- c = 2.99792458 x 10^8 m/s (speed of light)

λ = wavelength (in meters, m)

ν = frequency (reciprocal seconds, s^{-1})



Solved Problem:

The wavelength of a laser pointer is reported to be 663 nm. What is the frequency of this light?

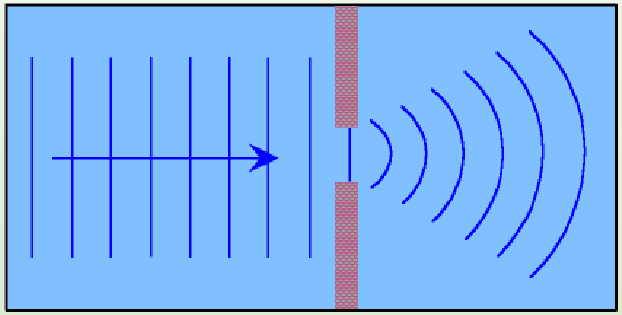
$$\upsilon = \frac{c}{\lambda}$$

$$\lambda = 663\,\text{nm} \times \frac{10^{-9}\,\text{m}}{\text{nm}} = 6.63 \times 10^{-7}\,\text{m}$$

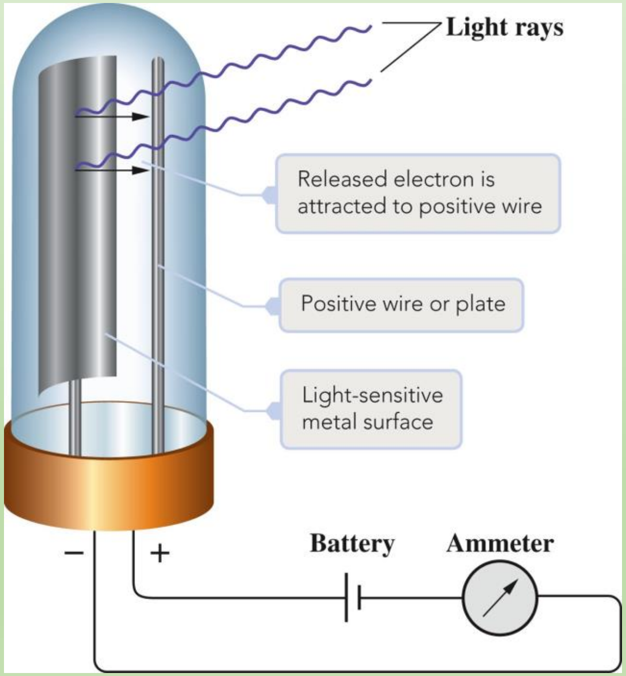
$$\upsilon = \frac{3.00 \times 10^{8}\,\text{m/s}}{6.63 \times 10^{-7}\,\text{m}} = 4.52 \times 10^{14}\,\text{s}^{-1}$$

Nature of Light

- In 1801, Thomas Young, a British physicist, showed that light could be diffracted; which is a wave property.



- The **photoelectric effect** was first observed by Heinrich Hertz1887. It is the ejection of an electron from the surface of a metal or other material when light shines on it.
- The wave theory could not explain the photoelectric effect because this discovery means that light has energy also.



Quantum Theory

- 1900 - Max Planck
- Radiant energy could only be emitted or absorbed in discrete quantities
- Quantum: packets of energy
- Revolutionized way of thinking (energy is quantized)
- Energy of a single quantum of energy

$$E = h\nu$$

where

E = energy (in Joules)

h = Planck's constant 6.63 x 10^{-34} J · s

ν = frequency

Photoelectric Effect

- Electrons ejected from a metal's surface when exposed to light of certain frequency
- Einstein proposed that particles of light are really **photons** (packets of light energy) and deduced that

$$E_{\text{photon}} = h\nu$$

- Only light with a frequency of photons such that $h\nu$ equals the energy that binds the electrons in the metal is sufficiently energetic to eject electrons. (*Threshold frequency*)
- If light of higher frequency is used, electrons will be ejected and will leave the metal with additional kinetic energy.
  - (what is the relationship between energy and frequency?)
- Light of at least the threshold frequency **and** of greater *intensity* will eject *more* electrons.

Solved Problem:

Calculate the energy (in joules) of a photon with a wavelength of 700.0 nm

$$\lambda = 700.0\ \text{nm} \times \frac{10^{-9}\ \text{m}}{\text{nm}} = 7.00 \times 10^{-7}\ \text{m}$$

Using c = λν

$$\upsilon = \frac{3.00 \times 10^{8}\ \text{m/s}}{7.00 \times 10^{-7}\ \text{m}} = 4.29 \times 10^{14}\ \text{s}^{-1}$$

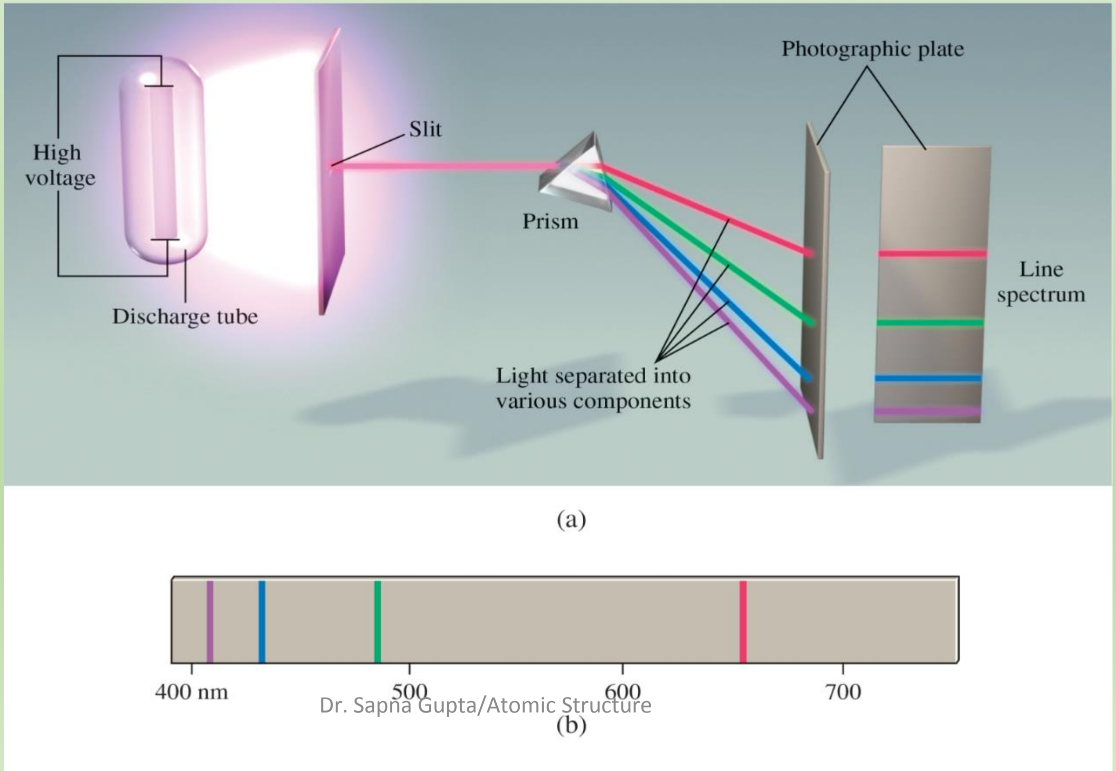
$E_{\text{photon}} = h\nu$

$$E = (6.63 \times 10^{-34}\ \text{J} \cdot \text{s})(4.29 \times 10^{14}\ \text{s}^{-1})$$

$$E = 2.84 \times 10^{-19}\ \text{J}$$

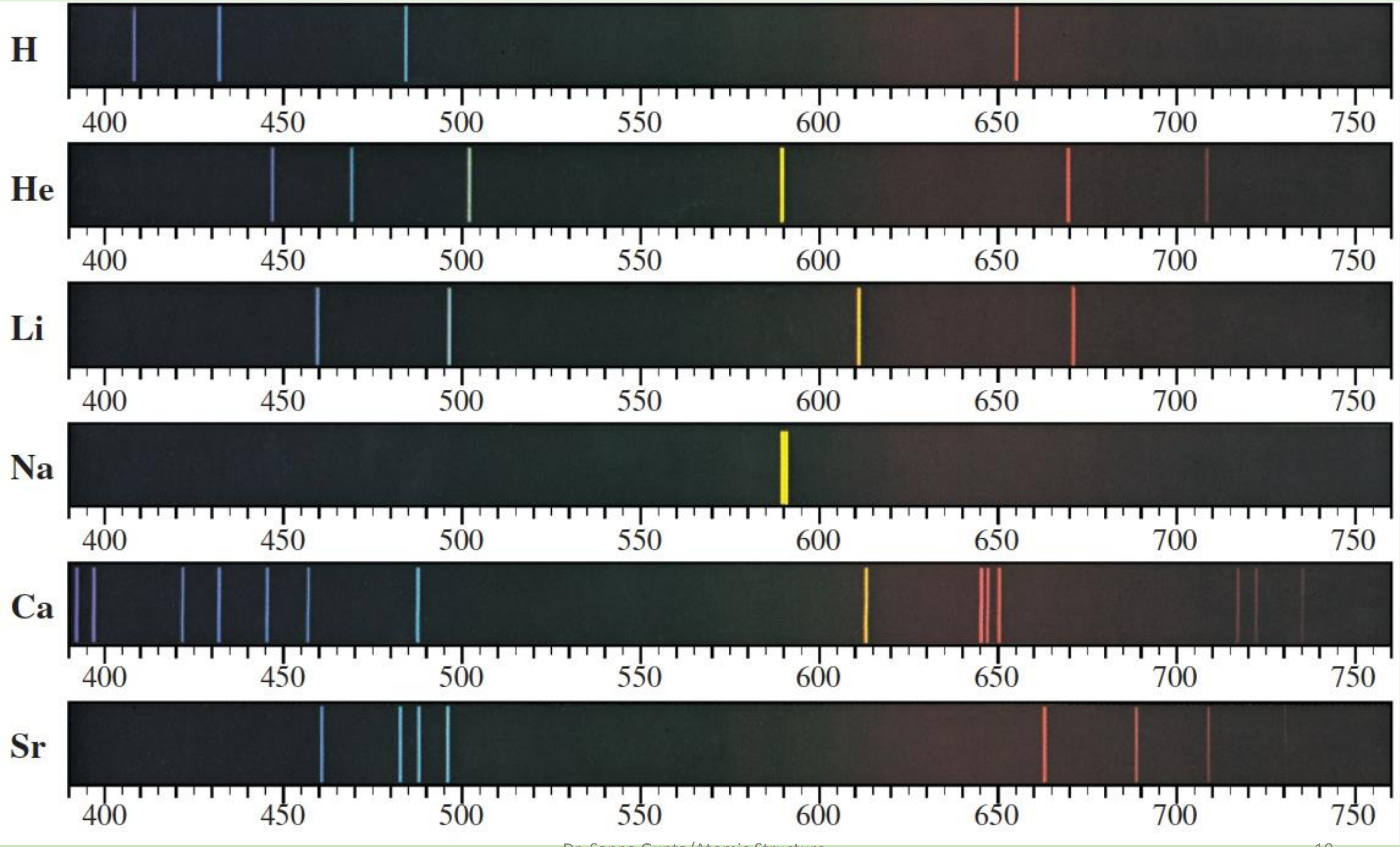
Bohr’s Theory of the Atom

- Planck’s theory along with Einstein’s ideas not only explained the photoelectric effect, but also made it possible for scientists to unravel the idea of atomic line spectra
- ***Line spectra***: emission of light only at specific wavelengths
- Every element has a unique emission spectrum
- Often referred to as “fingerprints” of the element



Line Spectrum

[Line spectra of elements on the web](). (need Java)



Line Spectrum of Hydrogen

- The *Rydberg equation*:

$$\frac{1}{\lambda} = R_\infty\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

- Balmer (initially) and Rydberg (later) developed the equation to calculate all spectral lines in hydrogen
- Bohr's contribution: showed only valid energies for hydrogen's electron with the following equation

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

- As the electron gets closer to the nucleus, E_n becomes larger in absolute value but also more negative.
- Ground state: the lowest energy state of an atom
- Excited state: each energy state in which $n > 1$
- Each spectral line corresponds to a specific transition
- Electrons moving from ground state to higher states require energy; an electron falling from a higher to a lower state releases energy
- Bohr's equation can be used to calculate the energy of these transitions within the H atom

Transition Between Different Levels

- An electron can change energy levels by absorbing energy to move to a higher energy level or by emitting energy to move to a lower energy level.
- For a hydrogen electron, the energy change is given by

$$\Delta E = E_{\mathrm{f}} - E_{\mathrm{i}}$$

$$\Delta E = -R_{\mathrm{H}}\left(\frac{1}{n_{\mathrm{f}}^{2}} - \frac{1}{n_{\mathrm{i}}^{2}}\right)$$

$R_{\mathrm{H}} = 2.179 \times 10^{-8}$ J, Rydberg constant

The energy of the emitted or absorbed photon is related to ΔE:

$$E_{\mathbf{photon}} = \left|\Delta E_{\mathrm{electron}}\right| = h\boldsymbol{\nu}$$

$$h = \text{Planck's constant}$$

We can now combine these two equations: $$h\boldsymbol{\nu} = \left| -R_{\mathrm{H}}\left(\frac{1}{n_{\mathrm{f}}^{2}} - \frac{1}{n_{\mathrm{i}}^{2}}\right) \right|$$

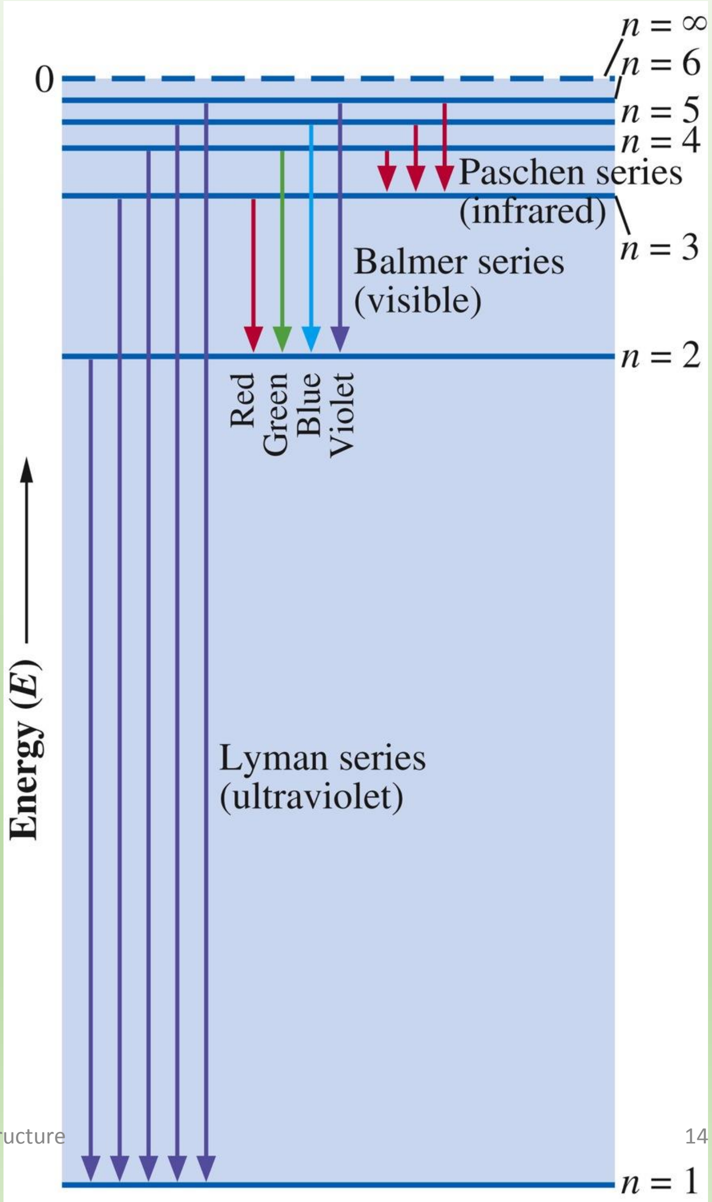
Transition…contd.

Light is absorbed by an atom when the electron transition is from lower n to higher n ($n_f > n_i$). In this case, ΔE will be positive.

Light is emitted from an atom when the electron transition is from higher n to lower n ($n_f < n_i$). In this case, ΔE will be negative.

An electron is ejected when $n_f = \infty$.

Electron transitions for an electron in the hydrogen atom.



Solved Problem:

What is the wavelength of the light emitted when the electron in a hydrogen atom undergoes a transition from n = 6 to n = 3?

$n_i = 6$

$n_f = 3$

$R_H = 2.179 \times 10^{-18}$ J

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

$$|\Delta E| = \frac{hc}{\lambda} \text{ so } \qquad \lambda = \frac{hc}{|\Delta E|}$$

$$\Delta E = \left(-2.179 \times 10^{-18}\,\text{J}\right)\left(\frac{1}{3^2} - \frac{1}{6^2}\right) = -1.816 \times 10^{-19}\,\text{J}$$

$$\lambda = \frac{\left(6.626 \times 10^{-34}\,\text{J} \bullet \text{s}\right)\left(3.00 \times 10^{8}\,\frac{\text{m}}{\text{s}}\right)}{\left|\left(-1.816 \times 10^{-19}\,\text{J}\right)\right|} = 1.094 \times 10^{-6}\,\text{m}$$

Planck

Vibrating atoms have only certain energies:

$$E = h\text{n or } 2h\text{n or } 3h\text{n}$$

Einstein

Energy is quantized in particles called photons:

$$E = h\text{n}$$

Bohr

Electrons in atoms can have only certain values of energy. For hydrogen:

$$E = -\frac{R_H}{n^2}$$

$$R_H = 2.179 \times 10^{-18}\,\text{J},\ n = \text{principal quantum number}$$

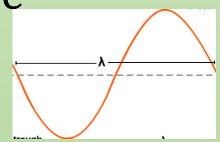
- Light has properties of both waves and particles (matter). But ***what*** about matter?
- In 1923, **Louis de Broglie**, a French physicist, reasoned that particles (matter) might also have wave properties.
- The wavelength of a particle of mass, *m* (kg), and velocity, *v* (m/s), is given by the de Broglie relation:

$$\lambda = \frac{h}{mv}$$

$$h = 6.626 \times 10^{-34}\ \text{J} \bullet \text{s}$$

- Building on de Broglie's work, in 1926, **Erwin Schrödinger** devised a theory that could be used to explain the wave properties of electrons in atoms and molecules.
- In 1927, **Werner Heisenberg** showed how it is impossible to know with absolute precision both the position, *x,* and the momentum, *p,* of a particle such as electron.
- Because $p = mv$ this uncertainty becomes more significant as the mass of the particle becomes smaller.

$$(\Delta x)(\Delta p) \geq \frac{h}{4\pi}$$



- Solving Schrödinger's equation gives us a **wave function**, represented by the Greek letter psi, ψ, which gives information about a particle in a given energy level.
- Psi-squared, ψ^2, gives us the probability of finding the particle in a region of space.

Key Points

- Electromagnetic spectrum
- Wavelength, frequency, energy (calculate)
- Quanta (of light - photon)
- Photoelectric effect
- Emission spectra
- Ground state vs excited state
- Heisenberg uncertainty principle
- Niels Bohr's line spectrum
- Schrodingers wave function