Chapter 7: Quantum Theory of the Atom

These Notes are to SUPPLIMENT the Text, They do NOT Replace reading the Text Material. Additional material that is in the Text will be on your tests! To get the most information, READ THE CHAPTER prior to the Lecture, bring in these lecture notes and make comments on these notes. These notes alone are NOT enough to pass any test!

The author is providing these notes as an addition to the students reading the text book and listening to the lecture. Although the author tries to keep errors to a minimum, the student is responsible for correcting any errors in these notes.

Drop a stone in a quite pond and watch the up / down wave motion.

Electromagnetic radiation includes visible light, x-rays and radio waves.

A **Wave** is a continuously repeating change or oscillation in matter or in a physical field.

Visible light, x-rays, radio waves are forms of **Electromagnetic Radiation**

Wavelength, the Greek letter **lambda** λ is the distance between any two adjacent identical points of a wave

Frequency is the Greek letter **nu** ν is the number of wavelengths that pass in one second.

Unit of freq is /s or s^{-1} called hertz (Hz).

Relation between Wavelength and Frequency is: $c = \lambda * \nu$

Where c = speed of light = 3.00 x 10^8 m/s **Speed of light = Wavelength * Frequency**

Electromagnetic Spectrum is the range of frequencies or wavelengths of electromagnetic radiation

1 Nanometer = 1 nm = 10^{-9} meter

Exercise 7.1 What is wavelength of yellow sodium emission if the frequency is 5.09 x 10^{14} /s

$$c = \lambda * \nu \qquad \lambda = c / \nu = \frac{3.00 \times 10^{8} \text{ m/s}}{5.09 \times 10^{14} \text{ /s}} = 5.89 \times 10^{-7} \text{ m or } \mathbf{589\ nm}$$

Exercise 7.1b The frequency of the red line in the spectrum of Potassium is 3.91 x 10^{14} /s. What is its wavelength?

$$\nu = c / \lambda$$

$$\lambda = \frac{c}{\nu} = \frac{3.00 \times 10^{8} \text{ m/s}}{3.91 \times 10^{14} \text{ /s}} = 7.67 \times 10^{-7} \text{ m} = 767 \text{ nm}$$

Example 7.2 The color violet has the wavelength of 408 nm. What's it's frequency?

$$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^{8} \text{ m/s}}{408 \times 10^{-9} \text{ m}} = 7.35 \times 10^{14} \text{ / sec}$$

Example 7.2 Cesium has two bright blue spectral lines. One has a wavelength of 456 nm, What is it's frequency? **DO IN CLASS**

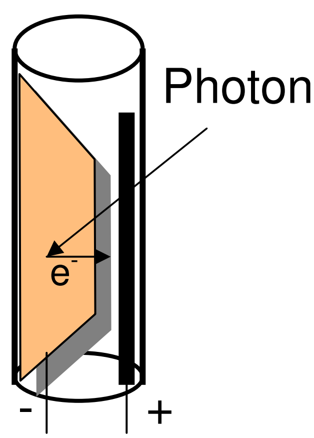
Description	Frequency	Wavelength Start	Wavelength End	
Gamma Rays		*1 ppm*	*10 ppm*	
X-Rays		*10 ppm*	*10 nm*	*1.0×10^{-10} m*
Far UV		*10 nm*	*100 nm*	
Near UV		**100 nm**	**400 nm**	**1.0×10^{-8} m**
VISIBLE		**400 nm [purple]**	**750 nm [red]**	
Near Infrared		**1 μ m**	**10 μ m**	**1.0×10^{-6} m**
Far Infrared		*100 μ m*	*1 mm*	
Microwave		*1 mm*	*100 + mm*	
Radar				
TV FM AM Radio Waves				

Color	Wavelength of light	Comment	Energy	
Ultra-Violet	*350 – 200 nm*			
Blue	**440 – 490 nm**	**Shortest Wavelength**	70 – 58 kcal/mole	**Most E**
Green	**490 – 570 nm**		58 – 50 kcal/mole	
Yellow	**570 – 585 nm**		50 – 49 kcal/mole	
Orange	**585 – 620 nm**		49 – 46 kcal/mole	
Red	**620 – 780 nm**	**Longest Wavelength**	46 – 37 kcal/mole	**Least E**
Infrared				

7.2 Quantum Effects & Photons

Diffraction is a property of waves in which the waves spread out when they encounter an obstruction or small hole about the size of the wavelength.

Photoelectric effect: light has both wave and particle properties. PE is the ejection of electrons from the surface of a metal when light shines on it. Photon strikes the metal plate. If it has the proper energy, it ceases to exist as a particle, it's absorbed and an electron is ejected. The electron is attracted to the + pole and thus conducts and electric current. Light is said to have both a wave and a particle characteristics, it consists of quanta or photons and particles of electromagnetic radiation.



Atoms oscillate at a definite frequency and can have only certain energies as defined by:

Planks Constant: $\mathbf{E = n\,h\,\nu}$
n = integer, ν = frequency, **h = Planks Constant = 6.63×10^{-34} J s**

Einstein stated that light consists of quanta or photons or particles of electromagnetic energy and has an energy E proportional to the observed frequency of the light.

$$E = h\,\nu$$

Example 7.3 The red spectrum of lithium is at 671 nm, what is it's energy?

From $c = \lambda * \nu$ we get:

$$\textbf{Frequency} = \nu = c / \lambda = \frac{3.00 \times 10^{8}\ \text{m/s}}{6.71 \times 10^{-7}\ \text{m}} = 4.47 \times 10^{14}\ \text{/s}$$

$$E = h\,\nu = 6.63 \times 10^{-34}\ \text{J s} * 4.47 \times 10^{14}\ \text{/s} = \mathbf{2.96 \times 10^{-19}\ J}$$

Exercise 7.3 What is the energy of the following:

Infrared	1.0×10^{-6} m
UltraViolet	1.0×10^{-8} m
X-Ray	1.0×10^{-10} m

Atomic Line Spectrum

Continuous Spectrum is a spectrum containing light of all wavelengths

Line Spectrum is a spectrum showing only certain colors or specific wavelengths of light.

The Line Spectrum of Hydrogen. J.J. Balmer showed the wavelengths λ in the visible spectrum of Hydrogen is calculated from:

$$1 / \lambda = 1.097 \times 10^{7}\ /\text{m} \left[\frac{1}{2^2} - \frac{1}{n^2} \right]$$

N = 3 = 656 nm   Find n = 4,5,6 [Possible Test Question]

Bohr's Postulates:

To get to a higher energy level, an electron must gain energy or be excited.
Bohr's Postulates also explains absorption of light – why materials have color [good test question]

1. Energy-level: an electron can have only specific energy values in an atom and these are called energy levels.

Bohr's rule for quantization = $E = R_H / n^2$ n = 1,2,3,… [for the Hydrogen atom]
n = the principal quantum number
$R_H = 2.179 \times 10^{-18}$ J

2. Transitions between energy levels: An electron in an atom can change energy only by going from one energy level to another energy level via an **electron transition.**

When an electron looses energy, it emits a photon.

$$\Delta E = -R_H \left(1/n_f^2 - 1/n_i^2 \right)$$

Example. 7.4 An electron goes from n = 4 to n = 2 energy levels, how much energy is emitted:

$\Delta E = -R_H (1/n_f^2 - 1/n_i^2)$ = -2.179×10^{-18} J [$1/2^2 - 1/4^2$] = $- 4.086 \times 10^{-19}$ J

Negative sign = energy is emitted.

Since Energy = **E = h ν**

then Frequency = $\nu = E / h = \dfrac{- 4.086 \times 10^{-19}\ J}{6.63 \times 10^{-34}\ J\ s} = 6.17 \times 10^{14} / s$

and since $\nu = c / \lambda$, then $\lambda = c / \nu = \dfrac{3.00 \times 10^{8}\ m/s}{6.17 \times 10^{14} / s}$ = **4.86×10^{-7} m or 486 nm**

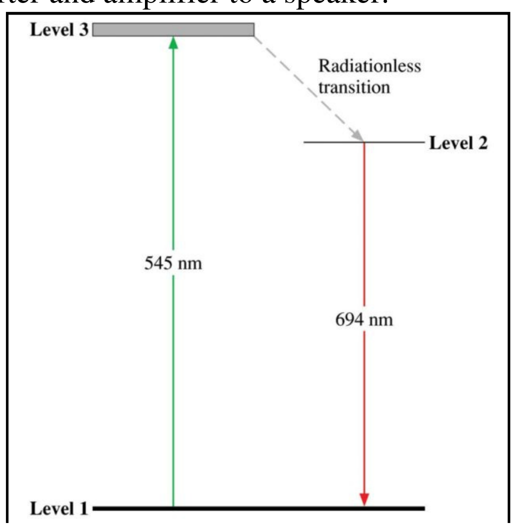
and Color = ?

To gain energy, and electron is excited. It gains energy by collision of two hydrogen atoms.

Exercise 7.4 Calculate the wavelength of light emitted from a Hydrogen Atom for the transition of n=3 to n=1. HAVE SOMEONE IN THE CLASS DO THIS ONE.

Laser Disk Players.

The ruby laser consists of aluminum oxide with a small amount of Chromium III Cr^{+3} ions. The laser [Light Amplification by Stimulated Emission of Radiation] emits an intense beam of monochromatic light. [description BSOC]. The laser is aimed at the CD and hits pits of digital data. The back reflected light is then merged via a prism with light that did not hit the CD. The light is either in or out of sync, depending on a pit or no pit. This signal goes to a A/D converter and amplifier to a speaker.



<u>De Broglie Relation</u> [BSOC]

Light has wave particle properties. So light, as a particle, can have wave properties depending on the mass of the object:

$$\lambda = h / m v$$

An electron will have a wavelength of 100×10^{-12} m, which due to the size of the electron is appreciable.
A baseball has a wavelength of 10^{-34} m which is too small to see.
A 145 g baseball is moving at 60 mph [27 m/s].

$\lambda = h / m v = \dfrac{6.63 \times 10^{-34}\ J\ s}{145\ g * 27\ ms} = 1.69 \times 10^{-34}$ m Note: $1\ J = 1\ kg\ m^2 / s^2$

Example 7.5 Calculate the wavelength in meters of a 1.00 kg mass moving at 1.00 km/hr?

Convert: 100 km / hr * 1 hr / 3600 x * 10^3 m / 1 km = 0.278 m/s

$$\lambda = h / m\nu = \frac{6.63 \times 10^{-34} \text{ J s}}{1.00 \text{ kg} * 0.278 \text{ m/s}} = 2.38 \times 10^{-33} \text{ m}$$

Execise 7.6 Calculate the wavelength in picometers of an electron moving at 2.19 x 10^6 m/s?
The mass of an electron, m = 9.11 x 10-31 kg, and Planck's constant (h = 6.63 x 10-34 J•s, or 6.63 x 10-34 kg•m2/s2

Wave Functions

Erwin Schrödinger worked in the branch of physics that mathematically describes the wave properties of submicroscopic particles called **Quantum Mechanic**s or **Wave Mechanics**.
He stated that at any moment, you cannot know the precise position and the precise momentum of an electron. [good test question]

An electron does not have a precise orbit.

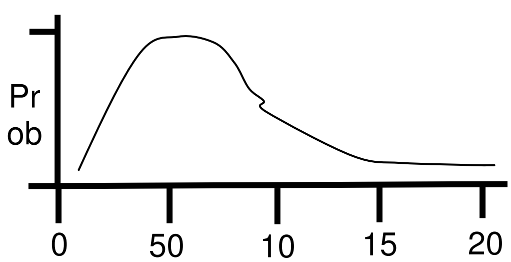
Heisenberg's Uncertainty Principle: The product of the uncertainty in position and the uncertainty I momentum of a particle can be no smaller than Plank's constant divided by 4 π. You can know very well where a particle is, but you cannot know where it is going. [good test question]

Or, if you really want to know the math: $\Delta x / \Delta p >= h / 4$ pi

The more precise you know the position, the less well you know the momentum of a particle

Quantum Numbers and the Atomic Orbitals

The **Schrödinge**r the wave function says an electron is likely to be at a position [psi = ψ] where ψ^2 = probablility of finding an electron within a region of space.

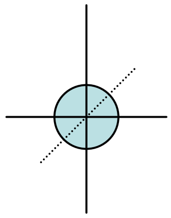


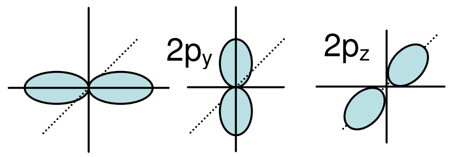
Per **Schrödinge**r the wave function in an atom is called an atomic orbital.

1. **Principal Quantum Number:** This quantum number describes the energy of an electron. n = 1, 2, 3 ..
   The larger the value of n, the larger the orbital.

2. **Angular Momentum Quantum Number:** l
   The Quantum Number distinguishes orbitals of **given n having different shapes.**
   L is an integer value of 0 → n – 1
   l = s, p, d, f, g
   Discuss shapes of s, p, d, f,

s Orbital: **p Orbital**





d Orbital **BSOC**

3. Magnetic Quantum Number: (which shape is it) m_l = from –l to 0 to +l

There are 2l + 1 orbital's in each subshell of quantum number l.
The Magnetic Quantum Number describes which of the p or d orbital's we have

4. Spin Quantum Number: m_s = - ½ or + ½ spin
Permissible Values of Quantum Number for the Atomic Orbitals

n	l	m_l	Subshell	# of Orbitals
1	0	0	1s	1
2	**0**	**0**	**2s**	**1**
2	**1**	**-1, 0, +1**	**2p**	**3**
3	0	0	3s	1
3	1	-1, 0, +1	3p	3
3	2	-2, -1, 0, +1, +2	3d	5
4	**0**	**0**	**4s**	**1**
4	**1**	**-1, 0, +1**	**4p**	**3**
4	**2**	**-2, -1, 0, +1, +2**	**4d**	**5**
4	**3**	**-3, -2, -1, 0, +1, +2, +3**	**4f**	**7**

Exercise 7.6 Which are permissible:

a. n = 1 l = 1 m_l = -2 m_s = + ½

b. n = 3 l = 1 m_l = 0 m_s = - ½

c. n = 2 l = 1 m_l = 0 m_s = + ½

d. n = 2 l = 0 m_l = 0 m_s = 1

Exercise 7.7 Which are permissible:

a. n = 0 l = 1 m_l = 0 m_s = + ½

b. n = 2 l = 3 m_l = 0 m_s = - ½

c. n = 3 l = 2 m_l = 3 m_s = + ½

d. n = 4 l = 2 m_l = 2 m_s = 0

Practice Questions:

Review Questions: All Example Problems in the chapter

Concept Questions: 7.19, 7.23, 7.27

Practice Problems: 7.33, 7.35, 7.41

7.43, 7.45, 7.51

7.57, 7.59, 7.61, 7.63