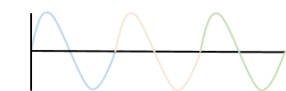


In-class Review (14th Sept – 16th Sept) – This Review is not graded.

(ANSWERS WILL BE POSTED AFTER BOTH SECTIONS HAVE COMPLETED THE REVIEW)

Concepts Covered

Concept Video 1: Properties of a wave, and properties of light (electromagnetic radiation)



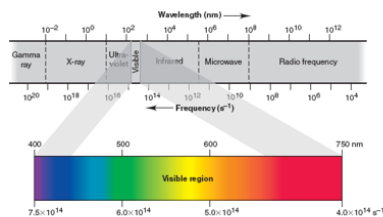
Wavelength (λ) : The **distance** the wave travels in one cycle (unit of length)

Frequency (ν) : **Number** of cycles a wave undergoes **per second**

Amplitude : The height of the crest – relates to the intensity or brightness

Speed of a wave = **wavelength** x **frequency**

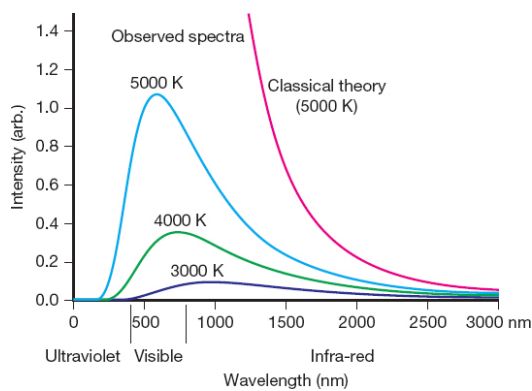
$$c = \lambda \nu$$



$$c = 3.00 \times 10^8 \text{ ms}^{-1}$$

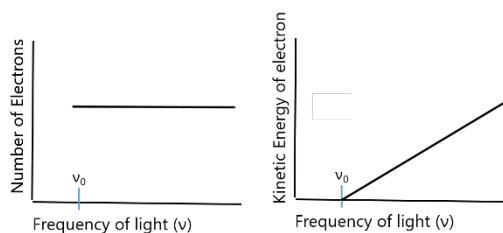
Wavelength and frequency are *inversely proportional*

Concept Video 2: Blackbody Radiation and Photoelectric Effect – why the wave theory was not enough to describe the nature of light.

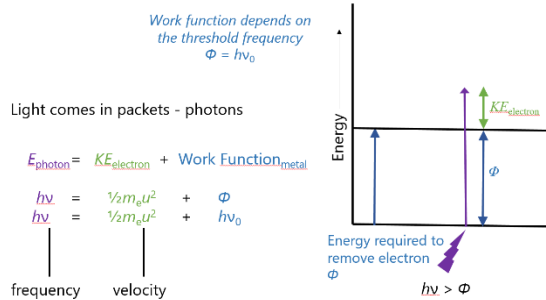


Blackbody Radiation observation explained by Planck.

Energy being released/absorbed in quanta, and each quanta has $E = h\nu$



Photoelectric effect – related energy of a photon of light to the Kinetic energy of an ejected electron.

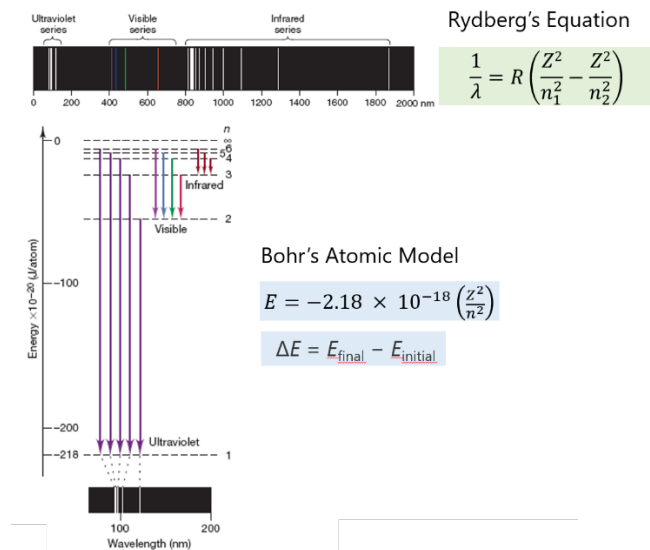


The photon must have a minimum threshold frequency (ν_0) to overcome the work potential of the metal and eject the electron.

Concept Video 3/4: Atomic Spectra

Rydberg- Mathematical equation that explained the observed spectral lines for H atom spectra

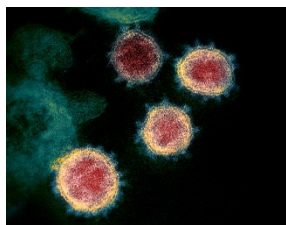
Bohr – Concept of orbits. An electron can only move from one orbit to another by emitting/absorbing specific (quantized) energy



Concept Video 4: Wave and Particle Nature of Light

Light and matter both have wave and particle like nature.

deBroglie wavelength of matter: $\lambda_{\text{deBroglie}} = h/p$



Electrons behave like waves (Davisson and Germer experiment shows that!)

Electron Microscopy Image of the SARS-CoV2 virus:

Image Courtesy – NIAID RML

Uncertainty principle: Both the position and the momentum of a particle cannot be determined precisely. There is always a minimum uncertainty.

Uncertainty Principle: $\Delta x \cdot \Delta p \geq h/4\pi$

Concept Video 5: Schrodinger Equation

Describe electron distribution as a wave equation and provided solutions for it (that gave energy values related to the wave equations)

Wavefunctions: A mathematical function that relates the location of an electron at a given point in space to its energy. Contains a radial (r) and angular component (θ, ϕ)

Orbitals : Mathematically derived regions of space with different *probabilities* of containing an electron.

Each wavefunction is defined by characteristic **quantum numbers** (n, l, m)

1. Principal Quantum Number (n)

1. Positive integer (1,2,3,...)
2. Indicates the relative size of the orbital – relative distance
3. Specifies the energy level (higher n indicated higher energy)

2. Angular Quantum Number (l)

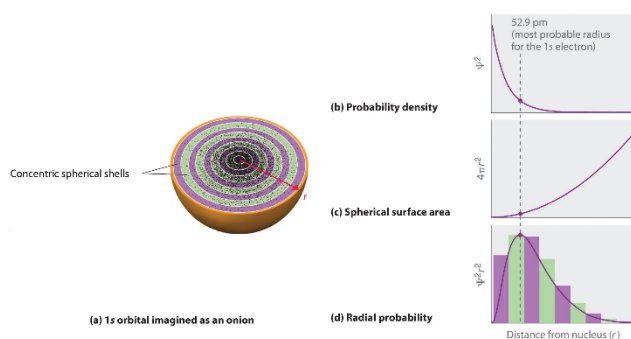
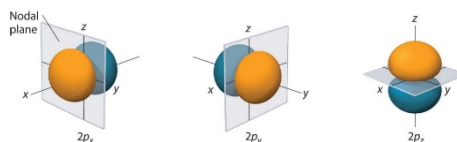
1. Positive Integer (0 to $n-1$)
2. Shape of the orbital
3. The value of n limits l
if $n=1$, l can only have the value 0; if $n=2$, l can have the values 0 and 1
When $l=0$ (s orbital); $l=1$ (p orbital); $l=2$ (d orbital); $l=3$ (f orbital)

3. Magnetic Quantum Number (m_l)

1. Integer ($-l$ to $+l$)
2. Orientation of the orbital around the nucleus
3. The value of l limits m_l ; For $l=1$, values of m_l can be $-1, 0$, and 1

Shapes of Orbitals

IMAGE COURTESY: [UCDAVIS CHEMWIKI](https://chemwiki.ucdavis.edu),
[CC BY-NC-SA 3.0 US](https://creativecommons.org/licenses/by-nc-sa/3.0/us/)



The square of the wave function Ψ^2 is the probability density, and gives us the **probability of finding an electron** at a distance from the nucleus. The probability density can be multiplied by volume to obtain the probability.

The **radial distribution** gives the **probability** for an electron to be found anywhere on the surface of a sphere located

at a distance r from the nucleus.

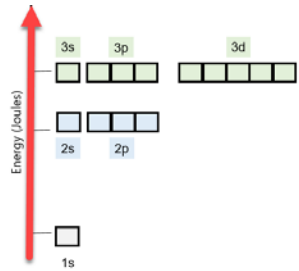
IMAGE COURTESY: [UCDAVIS CHEMWIKI](https://chemwiki.ucdavis.edu), [CC BY-NC-SA 3.0 US](https://creativecommons.org/licenses/by-nc-sa/3.0/us/)

Nodes: Regions where there is no probability of finding an electron (Both the wave function, Ψ , and the probability density, Ψ^2 , are 0 at nodes)

Radial Node = $n-1-l$

Angular Node = l

Concept Video 6: Why does Bohr's Model work for H atom



Energy of orbitals in a H atom depend on only principal quantum number, n .

Not true for other multi-electron species (we will see this in the next set of videos)

Review Questions

Try to work on these problems on your own (maybe after Monday's class). Bring your questions to the review. You will work on these within small groups during your review.

Feel free to ask your peers for help! TAs, TEAM Mentors, and I will also be there to help you.

If you know the concepts and had no trouble with these problems, help your peers.

Answers will be posted but after both Sections have already completed the review.

- 1) Covalent bonds in a molecule absorb radiations in the infra red region and vibrate. Specific bonds vibrate at specific frequencies and this information can be used to identify functional groups in molecules (especially important in Organic Chemistry!). The C—O bond absorbs radiation of wavelength 9.6 μm (micrometer). What frequency (s^{-1}) corresponds to this wavelength? What is the energy of the photon required to cause this bond vibration? (3 significant figures)

From wavelength, calculate frequency using the data below:

$$\lambda \text{ (m)} = \frac{c}{\nu} = 9.6 \times 10^{-6} \text{ m}$$

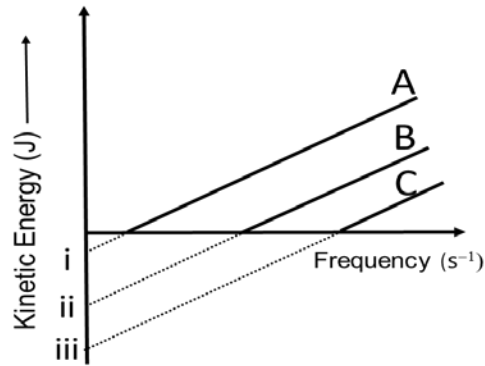
$$\text{So, } \nu = c / \lambda = 3.1229 \times 10^{13} \text{ s}^{-1}$$

$$E = h \nu = 2.07 \times 10^{-20} \text{ J}$$

- 2) A quantum dot emits yellow light with a wavelength of 537 nm.
If 1.00 mol of photons are emitted at this wavelength, what is energy emitted?
Give your answer to 3 significant figures.

$$\begin{aligned} E_{\text{photon}} &= h\nu = hc/\lambda = \frac{6.626 \times 10^{-34} \text{ J}\cdot\text{s} \times 2.998 \times 10^8 \frac{\text{m}}{\text{s}}}{5.37 \times 10^{-7} \text{ m}} = 3.70 \times 10^{-19} \text{ J/photon} \\ E_{\text{one mol photons}} &= N_A \times E_{\text{photon emitted}} = 6.022 \times 10^{23} \frac{\text{photons}}{\text{mol}} \times 3.70 \times 10^{-19} \frac{\text{J}}{\text{photon}} \\ &= 2.23 \times 10^5 \text{ J/mol} \end{aligned}$$

- 3) The plot of kinetic energy (J) of the ejected electron versus frequency (s^{-1}), for three metals A, B, and C is given.
- What do the values i, ii, and iii represent?
 - Why do the lines (A, B, and C) not begin at the origin?
 - Which metal would require the shortest wavelength to eject an electron?
 - Which metal would require the longest wavelength to eject an electron?



- The values (i), (ii), and (iii) represent the (–ve) of the work potential ($-\phi$).

From photoelectric effect equation:

$$h\nu = \text{KE} + \phi$$

KE (represented on y axis) = $h\nu$ (frequency is the variable on x axis) - ϕ

$$y = mx + b; \text{ where } x = \nu; y = \text{KE}; \text{ and } b = -\phi$$

So (i), (ii), and (iii) represent $-\phi$ for A, B, and C, respectively.

- The lines do not begin at the origin because the electrons must absorb a minimum (threshold) energy before it has enough energy to overcome the nuclear attraction.
- Metal C requires the shortest wavelength (highest energy) to eject an electron
- Metal A required the longest wavelength (lowest energy) to eject an electron

- 4) a. Find the wavelength (nm) of the photon emitted when an H atom undergoes a transition from $n = 5$ to $n = 2$.
 b. Without doing any calculations, compare this to a transition from $n = 6$ to $n = 3$. Which wavelength is longer? Explain your answer.

a. Using Rydberg's equation:

$$\frac{1}{\lambda} = R \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right) \quad R = 1.096776 \times 10^7 \text{ m}^{-1}$$

$$n_1 = 2 \quad n_2 = 5$$

$$\frac{1}{\lambda} = R \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right) = (1.096776 \times 10^7 \text{ m}^{-1}) \left(\frac{1}{2^2} - \frac{1}{5^2} \right) = 2,303,229.6 \text{ m}^{-1}$$

$$\lambda \text{ (nm)} = \left(\frac{1}{2,303,229.6 \text{ m}^{-1}} \right) \left(\frac{1 \text{ nm}}{10^{-9} \text{ m}} \right) = 434.1729544 \text{ nm} = \mathbf{434.17 \text{ nm}}$$

- b. The transition from $n = 6$ to $n = 3$ will require a longer wavelength. In hydrogen atom spectra the energy difference between subsequent levels keeps decreasing, so the transition from $n = 6$ to $n = 3$ will require lesser energy and hence a longer wavelength.

- 5) How fast must a 48.5 g tennis ball travel to have a de Broglie wavelength equal to that of a photon of wavelength 540 nm?

$$\text{Mass (kg)} = 0.0485 \text{ kg}$$

$$\text{Wavelength (m)} = (5400 \text{ nm}) \left(\frac{10^{-9} \text{ m}}{1 \text{ nm}} \right) = 5.4 \times 10^{-7} \text{ m}$$

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

$$u = \frac{h}{m\lambda} = 2.52997 \times 10^{-26} \text{ m/s} = \mathbf{2.5 \times 10^{-26} \text{ m/s}}$$

- 6) Give the total number of orbitals with the following quantum numbers:
- a. $n=2$
 - b. $n=3; l=1$
 - c. $n=4; l=3$
 - d. $n=4; l=2; m_l = -1$

$$n=2$$

4 total orbitals (one 2s and three 2p orbitals)

$$n=3; l=1$$

3 total orbitals (three 3p orbitals; $m_l = -1, 0, +1$)

$$n=4; l=3$$

7 total orbitals (seven 4f orbitals $m_l = -3, -2, -1, 0, +1, +2, +3$)

$$n=4; l=2; m_l = -1$$

one orbital

7) In the space provided below, name the orbital and draw the radial probability distribution for the following orbitals with the following quantum numbers

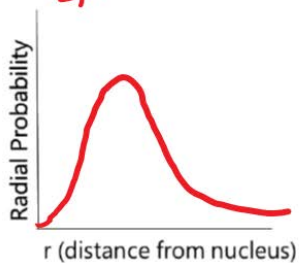
a. $n=2; l=1; m_l=-1$

b. $n=3; l=1; m_l=0$

c. $n=4; l=1; m_l=0$

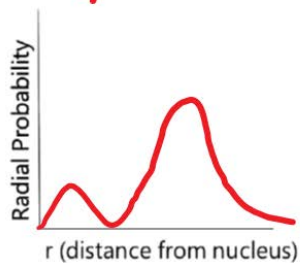
a.

2p



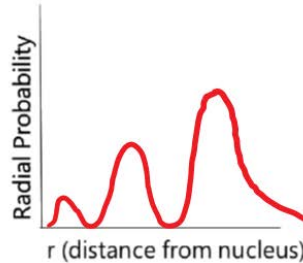
b.

3p



c.

4p



Radial nodes = $n - l - 1$

2p \Rightarrow 0 radial nodes

3p \Rightarrow 1 radial node

4p \Rightarrow 2 radial nodes