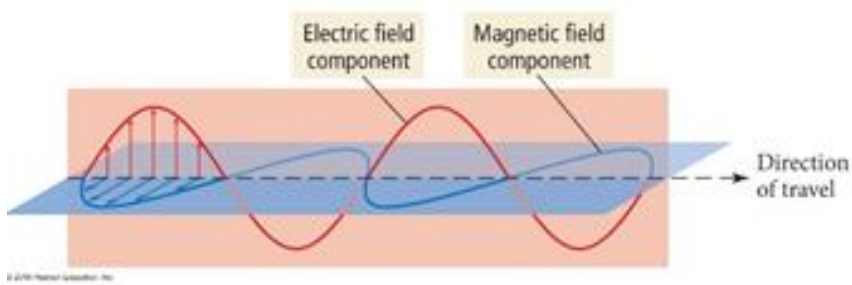
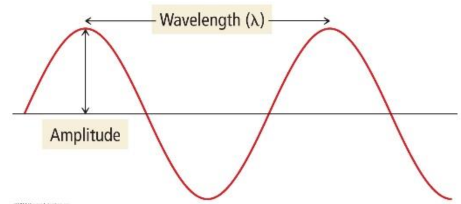
WAVE NATURE OF LIGHT

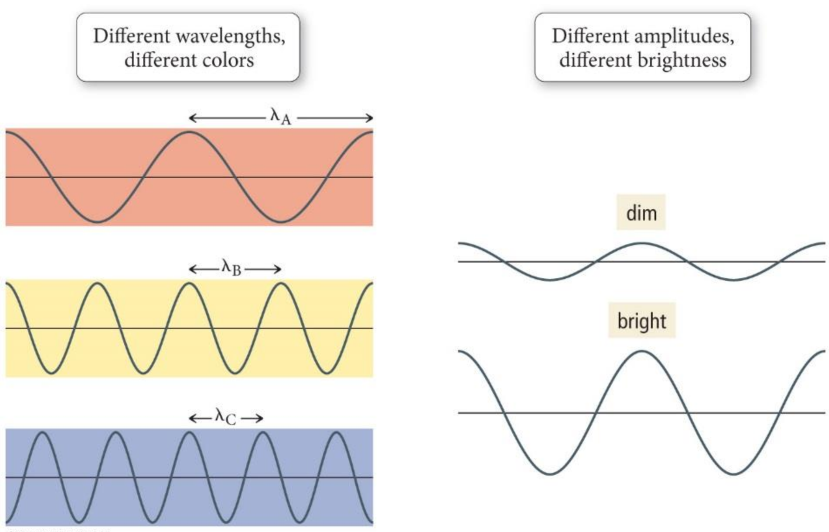
- Light is ***electromagnetic radiation***, a type of energy composed of oscillating electric and magnetic fields. The fields oscillate perpendicular to each other. In vacuum, these waves travel at a speed of 3.00×10^8 m/s.



- Waves can be characterized by their ***amplitude*** or ***wavelength*** (λ). Wavelength is the distance between adjacent crests. Wavelength is measured in units such as meters, micrometers or nanometers.



- Wavelength and amplitude are independent properties of light. Amplitude determines light's intensity, while wavelength determines light's color.



- Light is also characterized by its ***frequency*** (ν), the number of waves crests that pass through a stationary point in a given period of time. The units of frequency are waves/second (s^{-1}) or hertz (Hz).

WAVE NATURE OF LIGHT

- Frequency and wavelength of light are inversely proportional. Therefore,

$$\nu = \frac{c}{\lambda} \quad \text{where c = speed of light in vacuum}$$

Examples:

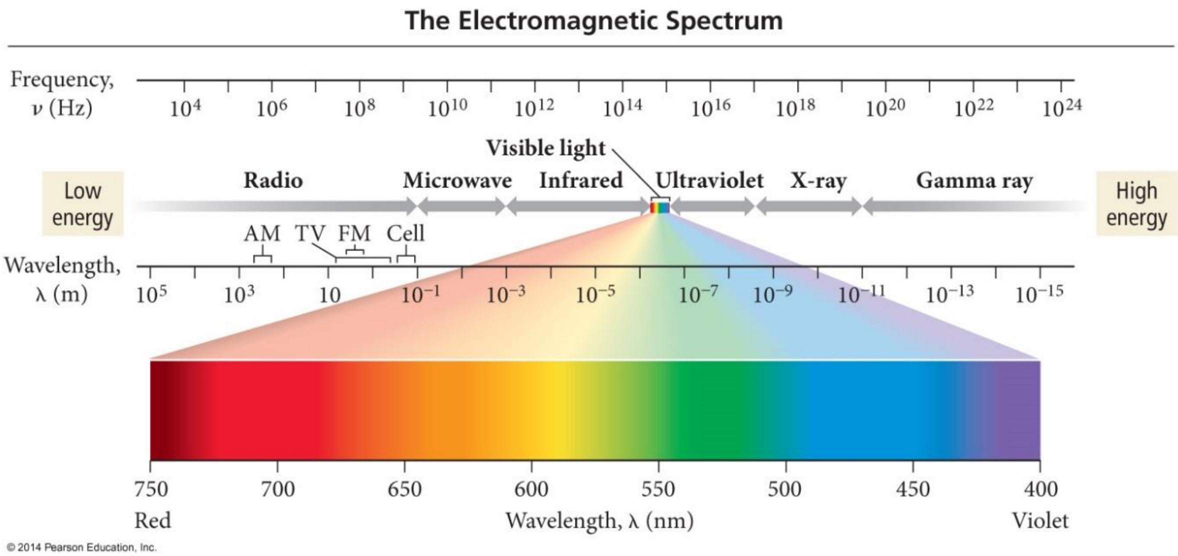
1. A barcode scanner emits a red light with frequency of 4.62×10^{14} s^{-1}. What is the wavelength of this light in nm?

2. A laser emits a green light with wavelength of 515 nm. Calculate the frequency of this light.

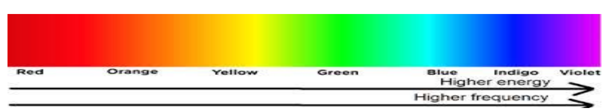
3. How many minutes would it take a light wave to travel from the planet Venus to Earth? (Average distance from Venus to Earth = 28 million miles).

ELECTROMAGNETIC SPECTURM

- The classification of electromagnetic waves according to their frequency is called ***electro-magnetic spectrum***. These waves include long-wavelength, low-frequency waves (radio waves) and short-wavelength, high-frequency waves (gamma rays).



- Visible light is only a small portion of the electromagnetic spectrum and ranges from red (long λ) to violet (short λ).



Examples:

1. Arrange the 3 types of radiation–visible light, X-ray and microwave–in order of increasing:

   a) wavelength

   b) frequency

   c) energy

2. Arrange the 3 colors of visible light–green, red and blue–in order of increasing:

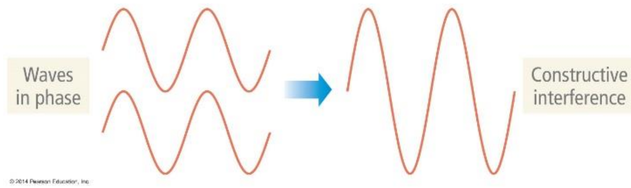
   a) wavelength

   b) frequency

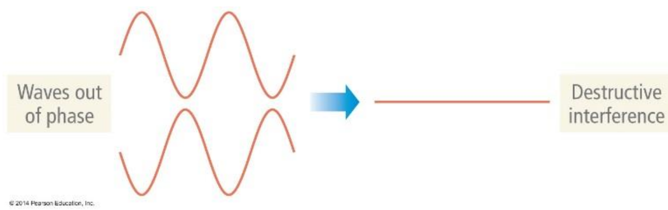
   c) energy

INTERFERENCE & DIFFRACTION

- Waves interact with one another in a characteristic way, called ***interference***, by adding or cancelling one another depending on their alignment upon interaction. For example, when two waves of equal amplitude align with each other ***in phase***–align with overlapping crests–a wave of twice the amplitude results. This is called ***constructive interference***.

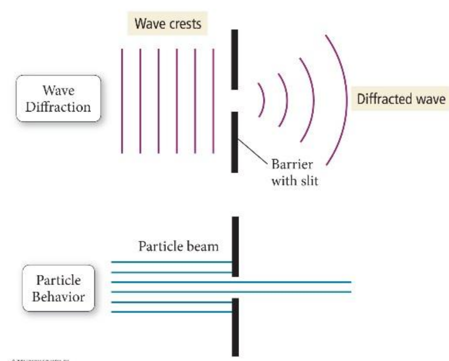


- On the other hand, if the two waves are completely out of phase when they interact–align so that crest of one aligns with trough of the other–the waves cancel by ***destructive interference***.

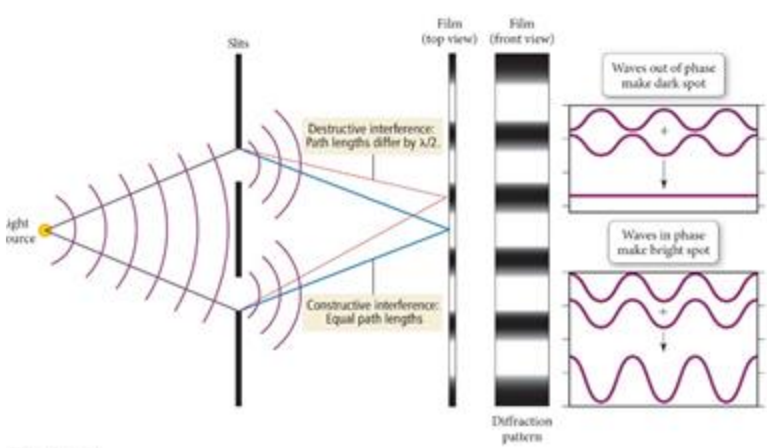


- Another characteristic behavior waves exhibit is called diffraction. When waves encounter an obstacle or a slit that is comparable in size to its wavelength, it bend (or diffracts) around it.

- Encountering similar situation, particles, by contrast, do not diffract and pass through the slit.



- The diffraction of light through two slits separated by a small distance, coupled with interference, results in a very characteristic pattern, called ***interference pattern***.



QUANTUM THEORY

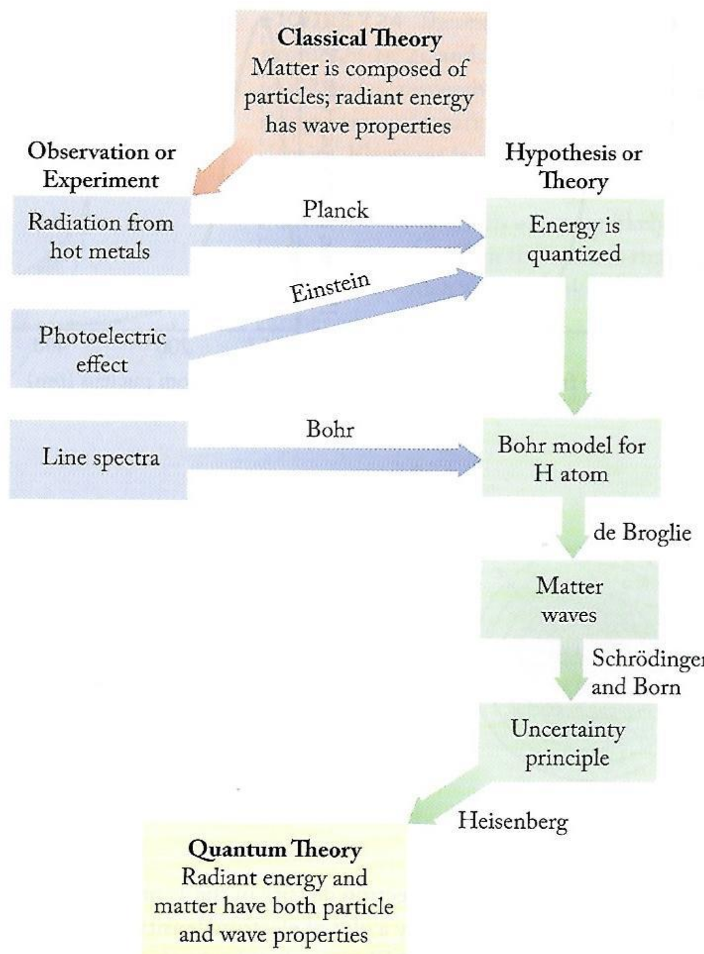
- At the end of the nineteenth century, physicists thought matter and energy to be distinct. Matter was thought to consist of particles, whereas energy in the form of light (electromagnetic radiation) was described as a wave.

- Particles were things that had mass and their position in space could be specified. Waves were described as massless and their positions in space could not be specified. It was also thought that matter and light could not intermingle.

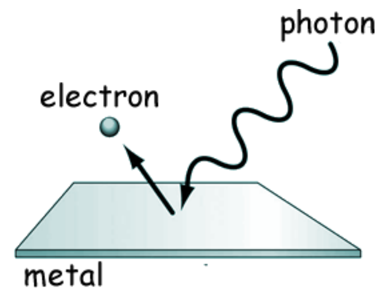
- Although the wave nature of radiation was well understood, scientists discovered limitation to this model when describing radiation.

- The development of quantum theory in the next 30 years revolutionized our understanding of the nature of light and radiation and transformed classical theories of matter and light into a modern theory.

- The diagram on the right partially traces the development of this theory and various contributions by the different scientists.



PARTICLE NATURE OF LIGHT



- At the beginning of the twentieth century, certain experimental results suggested that the understanding above was incorrect. Two such results were the ***black body radiation*** and ***photoelectric effect***, that led to two important advances in understanding of the nature of light and energy.

- These two advances are summarized below:

  1. Max Planck's studies of black body radiation led to his proposal that energy could be ***quantized***–that is have fixed values that are whole number multiples of hv (where h is a constant called Planck's constant). Based on Planck's proposal, the change in energy of a system, ΔE, can be represented by the equation:

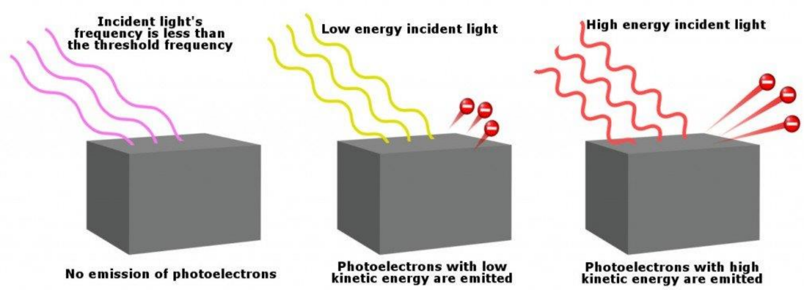
$$\Delta E = nhv$$

  2. Albert Einstein's analysis of photoelectric effect led him to propose that light energy must come in fixed amounts (***quanta***). According to Einstein, the energy of a quantum of light is proportional to its frequency, as shown:

$$E = h\nu \qquad \text{where h, called Planck's constant} = 6.626\text{x}10^{-34}\text{ Js}$$

- A quantum of light is also called a ***photon***. The energy of a photon of light can also be expressed in terms of its wavelength:

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



Examples:

1. An electromagnetic wave has a wavelength of 625 nm. What is the energy of one photon of this wave?

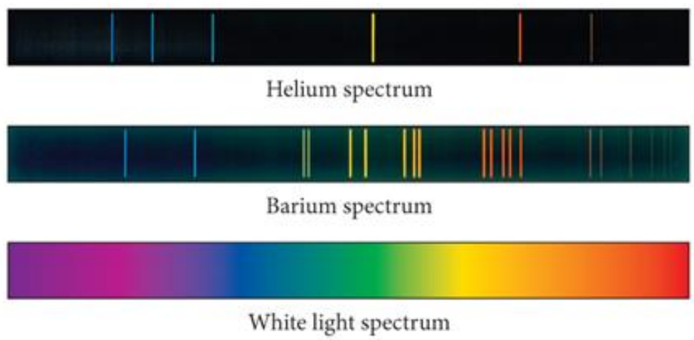
2. A nitrogen gas laser pulse with a wavelength of 337 nm contains 3.83 mJ of energy. How many phones does it contain?

3. It takes 7.21×10^{-19} J of energy to remove an electron from an iron atom. What is the maximum wavelength of the light that can do this?

ATOMIC LINE SPECTRA

- When an atom absorbs energy–in the form of light, heat or electricity–it re-emits the energy as light. A neon light is an example of this effect.

- Close inspection of the light emitted by various atoms reveals that it contains several distinct wavelengths and colors. The components of the light emitted by the atom can be separated by passing the light through a prism. The result is a series of bright lines called a ***line spectrum.*** The line spectrum of a particular element is always the same.



- Note that the white light spectrum is ***continuous***, while the atomic spectra contain only certain wavelengths. This indicates that only certain energies are allowed for the electron in an atom, or in other words, the energy of the electron in the atom is quantized. This observation ties in perfectly with the postulates of Max Planck that state that were discussed earlier.

- Johannes Rydberg, a Swedish mathematician, analyzed many atomic spectra and developed an equation that predicts the wavelengths of the hydrogen emission spectrum. However, this equation (shown below), did not explain why atomic spectra are discrete.

$$\frac{1}{\lambda} = R\left(\frac{1}{m^2} - \frac{1}{n^2}\right) \qquad \text{where } R=1.097\text{x}10^{7}\ \text{m}^{-1}$$

<u>Examples:</u>

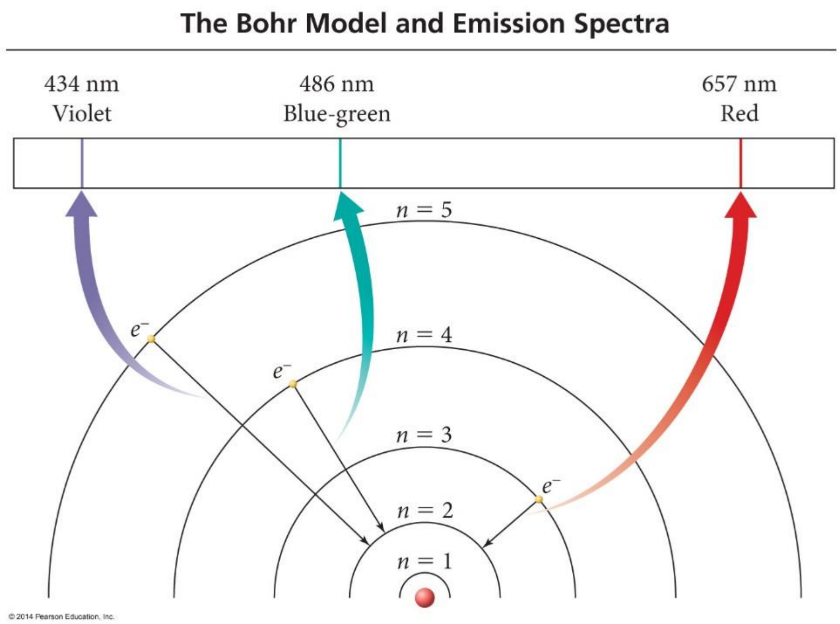
1. Calculate the wavelength of light produced when m=2 and n=4.

BOHR MODEL OF ATOM

- ***Neils Bohr***, a Danish physicist, studied the hydrogen atom extensively, and developed a model for the atom that was able to explain the atomic spectra.

- ***Bohr's model*** of the atom consisted of electrons orbiting in circular orbits (analogous to planetary orbits around the sun). However, unlike the planetary orbits–which can exist at any distance from the sun–Bohr's orbits exist only at specific, fixed distances from the nucleus. Bohr called these orbits ***stationary states***.

- Bohr further proposed that electrons only emit energy when they transition–jump–between orbits. The emission spectrum of an atom consists of discrete lines because the energy of the photon emitted when an electron makes a transition can only be of certain values depending on the difference in the energy of the stationary states. Three of these transitions for the hydrogen atom are shown below:

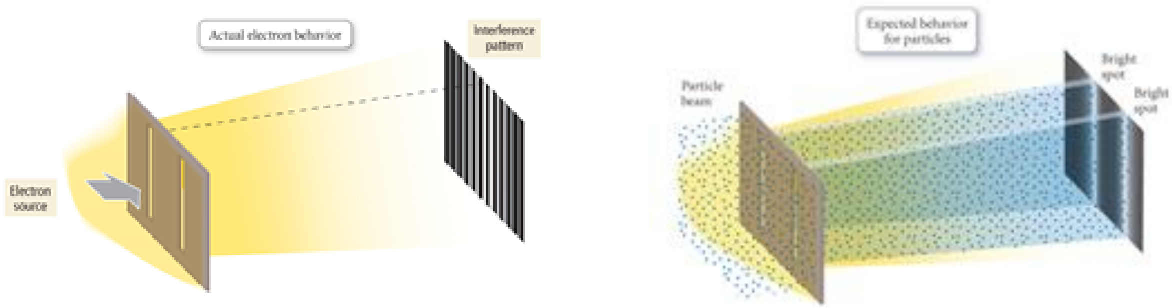


- In spite of its initial success in explaining the line spectrum of the hydrogen atom, the Bohr model left many unanswered questions. However, it did serve as a intermediate model between the classical view of the electron (particle) and a quantum-mechanical view (particle and wave).

- Bohr model was later replaced with a more thorough quantum-mechanical model that fully incorporates the wave nature of the electron.

WAVE NATURE OF MATTER

- Earlier we described diffraction and interference as characteristic behaviors of waves as they pass through a slit. Similar experiments with electrons passing through two slits resulted in the same interference patterns as observed with light.



- The interference pattern observed with the electrons in these experiments were similar to that expected of waves and not particles. Furthermore, the interference pattern observed with electrons is not caused by pairs of electrons interfering with each other, but rather single electrons interfering with themselves.

- These experimental results led to the proposal by Louis de Broglie in 1924 that electrons–originally thought of as particles–also possesses a wave nature, and its wavelength is related to its kinetic energy by the ***de Broglie equation***:

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

where, h = Planck's constant

m = mass of electron

v = velocity of electron

- Based on the de Broglie equation, the velocity of a moving electron is related to its wavelength–knowing one is equivalent to knowing the other.

- If de Broglie's proposal is correct and electrons travel in waves, then they should also exhibit the wave properties of diffraction and interference. A fast moving electron has a wavelength of about 10^{-10} m, so the "perfect" slit would be the natural spacing between crystals in an atom.

- In 1927, C. Davisson and L. Germer guided a beam of electrons at a nickel crystal and obtained an electron diffraction pattern. This experimental observation proved that electrons–particles with mass and charge–create diffraction patterns, just as electromagnetic waves do.

- Although electrons do not have orbits of fixed radius, as de Broglie thought, the energy levels of atoms are related to the wave nature of the electron.

WAVE NATURE OF MATTER

- The table below lists the calculated de Broglie wavelength of several objects.

Substance	***Mass (g)***	***Velocity (m/s)***	***λ (m)***
Slow electron	9×10^{-28}	1.0	7×10^{-4}
Fast electron	9×10^{-28}	5.9×10^{6}	1×10^{-10}
Alpha particle	6.6×10^{-24}	1.5×10^{7}	7×10^{-15}
Baseball	142	25.0	2×10^{-34}
Earth	6.0×10^{27}	3.0×10^{4}	4×10^{-63}

<u>Examples:</u>

1. If an electron has a velocity of 5.0×10^{5} m/s, what is its wavelength in nm?

2. What is the velocity of an electron that has a wavelength approximately the length of a chemical bond (1.2×10^{-10} m)?

3. Protons can be accelerated to speeds near that of light in particle accelerators. Estimate the de Broglie wavelength (in nm) of such a proton moving at 2.90×10^{8} m/s. (mass of a proton = 1.673×10^{-27} kg).

UNCERTAINTY PRINCIPLE

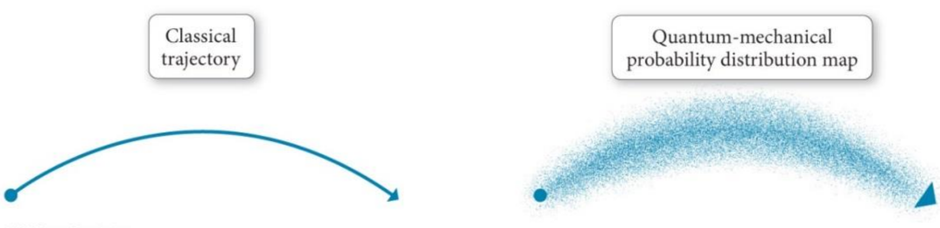
- The wave nature of the electron is difficult to reconcile with its particle nature. How can a single entity behave both as a wave and as a particle?

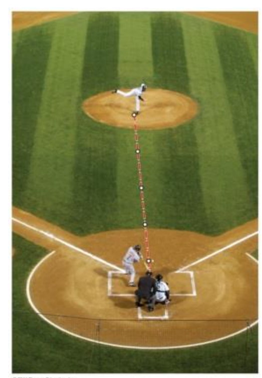
- Results of experiments to study this phenomena indicate a very odd occurrence: no matter how hard one tries or whatever method one uses, one cannot both see the interference pattern and simultaneously determine which slit the electron passes through.

- The wave and particle properties of the electron are said to be complimentary properties– they exclude one another. As a result, the more we know about one of them, the less we know about the other, and which property we observe depends on the type of experiment performed. In quantum mechanics, the observation of an event affects it outcome.

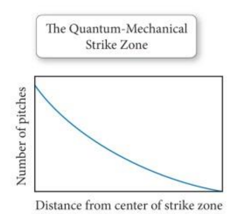
- Based on the de Broglie equation, the velocity of an electron is related to its wave nature. The position of an electron, however, is related to its particle nature. Consequently, the inability to observe the wave and particle nature of electron simultaneously means that we cannot simultaneously measure its velocity and position. This is referred to as the ***Heisenberg's uncertainty principle***.

- Whereas a classical particle moves in a predicted path and has a defined trajectory, the path of an electron can only be defined by a probability map–a statistical picture of where the electron is most likely to be found. The darker shadings in the probability map indicate greater probability of finding an electron in that region.





- To better understand the concept of probability distribution map, consider a baseball thrown by a pitcher to a catcher on home plate. The path of the thrown ball can be predicted by the catcher and he can catch it with certainty by placing his mitt in the correct place.



- Now consider the pitcher throwing an electron to the catcher on home plate. Unlike the baseball, the path of the electron is indeterminate and can only be described statistically. The statistical pattern of electron thrown can be used to draw a quantum mechanical strike zone.

QUANTUM MECHANICS & THE ATOM

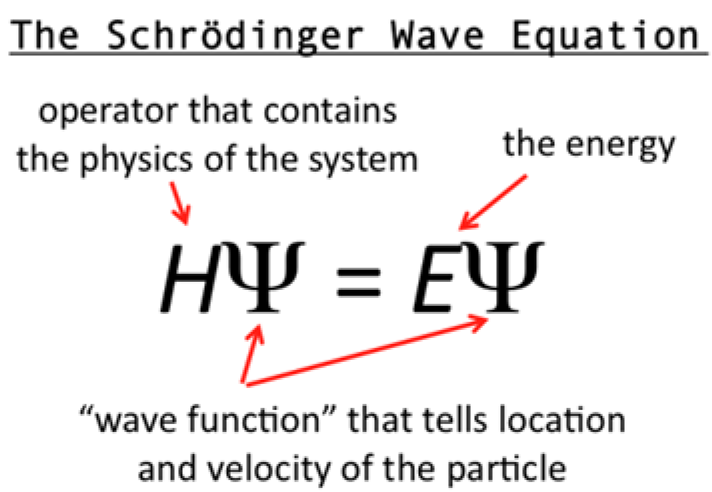
- The probability distribution map that describes the position of an electron is called an **orbital**. Because of the complementary nature of position and energy of an electron, we can specify the energy of an electron precisely, but its position can only be described by the orbital.

- Erwin Shrodinger developed equations, that when solved, give the wave function of the quantum system, that describe the wave properties of a particle in that system.

- Each solution to the equation is associated with a particular wave function, also called an atomic orbital. Each orbital is specified by three interrelated ***quantum numbers***:

  - principal quantum number (n)
  - angular momentum quantum number (l)
  - magnetic quantum number (m_l)

- These quantum numbers all have integer values, and will be discussed in detail individually.



Development of quantum physics

- Blackbody radiation
  (Max Planck, 1900: NP 1918)

- Photoelectric effect
  (Albert Einstein, 1905: NP 1921)

- Atomic spectra and structure
  (Niels Bohr, 1913: NP 1922)

Development of quantum physics

- Interference of particles
  (Louis de Broglie, 1924: NP 1929)

- Wave mechanics
  (Erwin Schrödinger, 1926: NP 1933)

- Uncertainty principle and matrix mechanics (Werner Heisenberg, 1927: NP 1932)

QUANTUM NUMBERS

- An atomic orbital is specified by three quantum numbers that are associated, respectively, with orbital's size, shape and orientation in space.

- The quantum numbers have a hierarchical relationship: the size-related number limits the shape-related number, which limits the orientation-related number.

<u>The Principal Quantum Number (n):</u>

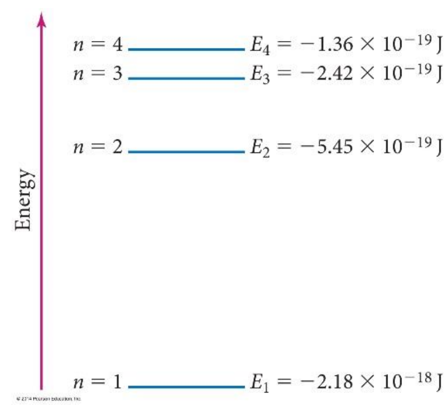
- This quantum number is an integer that determines the overall size and energy of an orbital. Its possible values are n= 1, 2, 3….and so on.

- For hydrogen, the energy of an electron in an orbital with quantum number n is given by the equation:

$$E_n = -2.18\text{x}10^{-18}\text{ J }\left(\frac{1}{n^2}\right) \qquad (n = 1, 2, 3...)$$

- The energy is negative, because the electron's energy is lowered by its interaction with the nucleus. The constant $2.18\text{x}10^{-18}$ J is called the Rydberg constant for hydrogen (R_H).

- Orbitals with higher values of n have greater energies (less negative), as shown in the energy level diagram. Note also, that as n increases, the spacing between the energy levels becomes smaller.



<u>The Angular Momentum Quantum Number (*l*):</u>

- This quantum number (also called azimuthal) is an integer that determines the shape of the orbital. The possible values of l are 0, 1, 2….(n–1). For example, if n=1, then the only possible value of l is 0; if n=2, the possible values of l are 0 and 1. To avoid confusion between n and l, values of l are assigned letters, shown below:

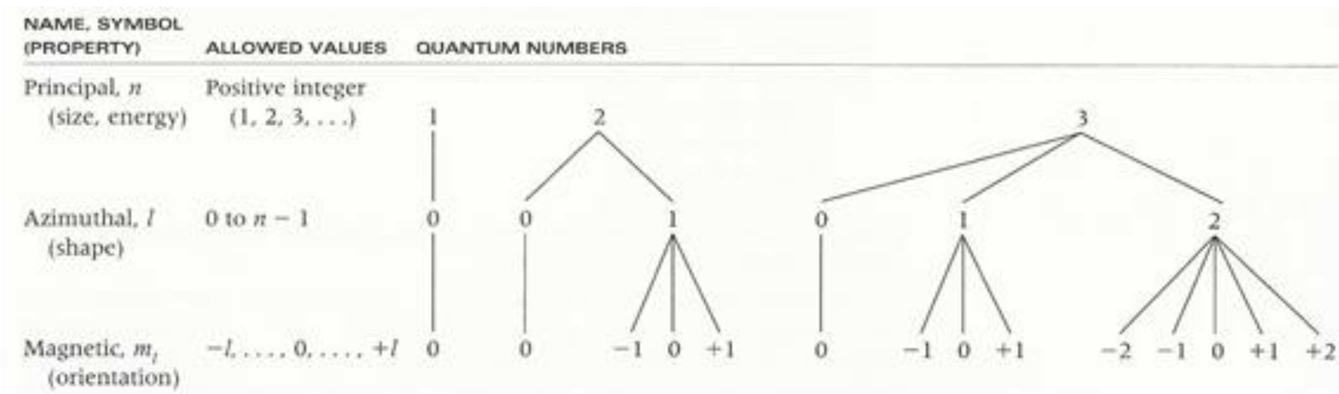
Value of l	Letter Designation
$l = 0$	s
$l = 1$	p
$l = 2$	d
$l = 3$	f

QUANTUM NUMBERS

The Magnetic Quantum Number (m_l):

- This quantum number is an integer that specifies the orientation of the orbital. The possible values of m_l are integer values (including zero) ranging from $-l$ to $+l$. For example, if $l = 1$, then m_l are –1, 0 and +1.

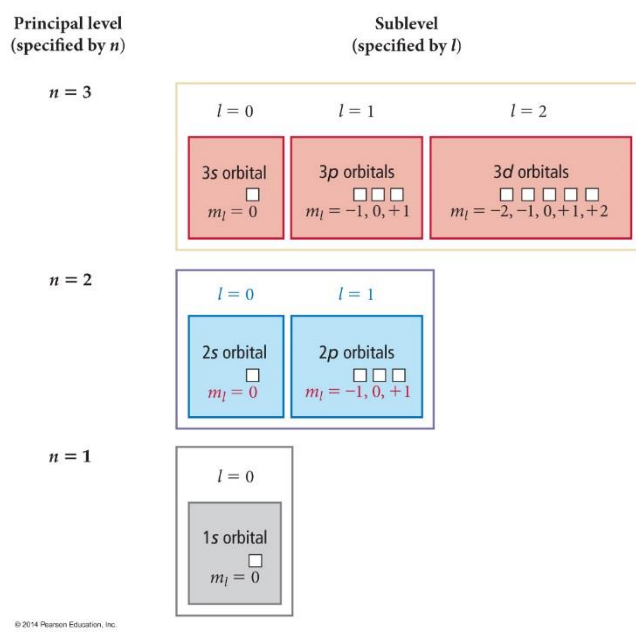
- The chart below shows the hierarchy of quantum numbers for atomic orbitals:



- Each specific combination of the 3 quantum numbers specifies one atomic orbital. For example, the orbital with n=1, l=0 and m_l =0 is known as 1s orbital.

- Orbitals with the same values of n are said to be in the same principal level. Orbitals with the same value of n and l are said to be in the same sublevel.

- The orbitals (and their corresponding quantum numbers) in the first three principal levels are shown on the right.



QUANTUM NUMBERS

Examples:

1. Write l and m_l values for n=4.

2. Supply the missing quantum number(s) or sublevel names below:

n	l	m_l	Name
?	?	0	4p
2	1	0	?
3	2	–2	?
?	?	?	2s

3. Each set of quantum numbers below is supposed to specify an orbital. However, each set contains one quantum number that is not allowed. Replace the quantum number that is not allowed with one that is allowed.

   a) n =3; l =3; m_l = +2

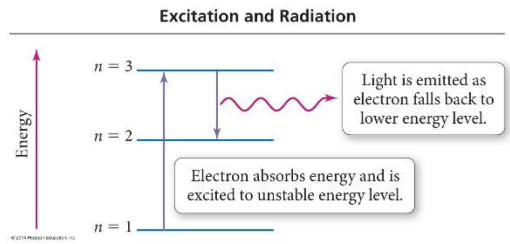
   b) n =2; l =1; m_l = –2

   c) n =1; l =1; m_l = 0

4.

ATOMIC SPECTRA & QUANTUM THEORY

- Quantum theory explains the atomic spectra that was discussed earlier. When an atom absorbs energy, an electron in a lower energy orbital is excited to a higher energy orbital.



- As a result, the atom becomes unstable, and quickly returns to a lower energy orbital, releasing a photon of light containing an amount of energy equal to the energy difference between the two energy levels.

- The difference in the energy between the two levels is given by:

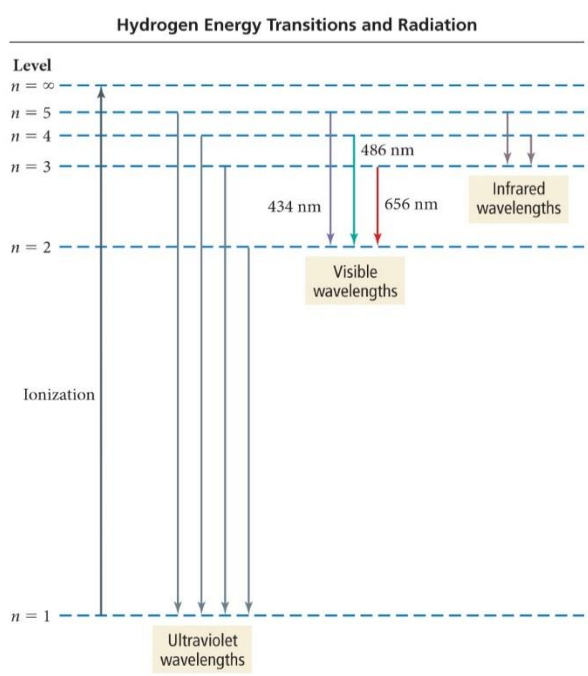
$$\Delta E = E_{final} - E_{initial}$$

- Using the equation introduced earlier for calculating the energy of an electron in any orbital, we can arrive at the following relationship:

$$\Delta E = -2.18\text{x}10^{-18}\ \text{J}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

- The Rydberg equation can be derived from the relationship above and the relationship discussed earlier between the wavelength and the energy of light.

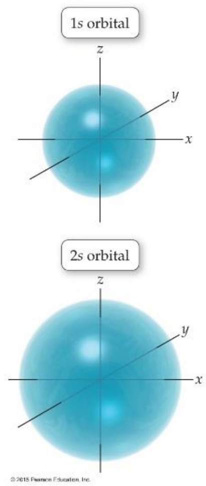
- Several of the transitions in the hydrogen atom and their corresponding wavelengths are shown below. Note that transitions between orbitals that are further apart in energy produce light that is higher in energy and shorter in wavelength than transitions between orbitals closer together.

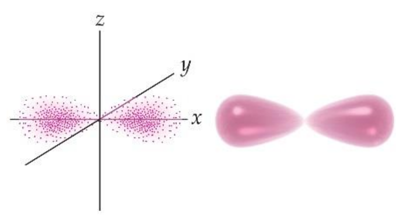


ORBITAL SHAPES

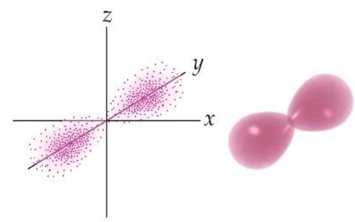
- Each orbital type can be represented by a geometric shape that encompasses the volume where the electron in likely to be found.

- These shapes for the s, p and d orbitals are shown below:

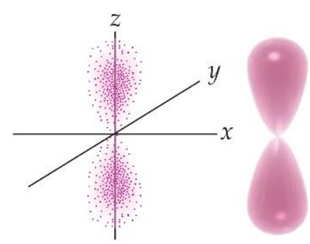




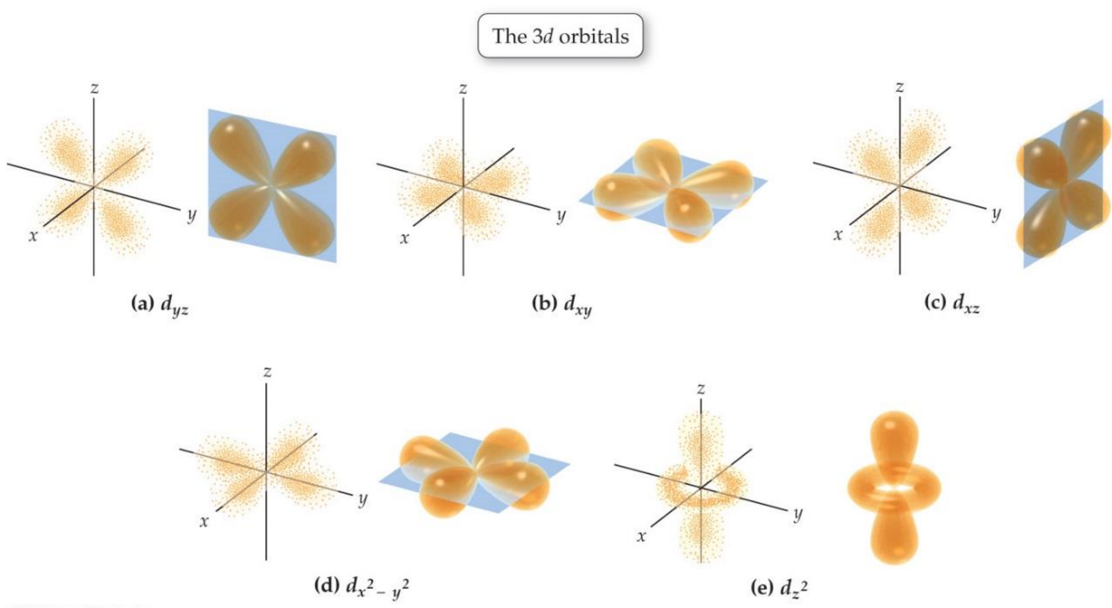
(a) p_x



(b) p_y



(c) p_z



ANSWERS TO IN-CHAPTER PROBLEMS:

Page	***Example No.***	***Answer***
2	1	649 nm
	2	5.83×10^{14} s^{-1}
	3	2.5 min
3	1a	x-ray < visible < microwave
	1b	microwave < visible < x-ray
	1c	microwave < visible < x-ray
	2a	blue < green < red
	2b	red < green < blue
	2c	red < green < blue
6	1	435 nm
	2	3.18×10^{-19} J
	3	541 nm
7	1	486 nm
10	1	1.47 nm
	2	6.1×10^{6} m/s
	3	1.37×10^{-6} nm
16	1	Discussed in class
	2	Discussed in class
	3	Discussed in class