UNIT 7 ATOMIC AND NUCLEAR PHYSICS

PHYS:1200 LECTURE 33 — ATOMIC AND NUCLEAR PHYSICS (1)

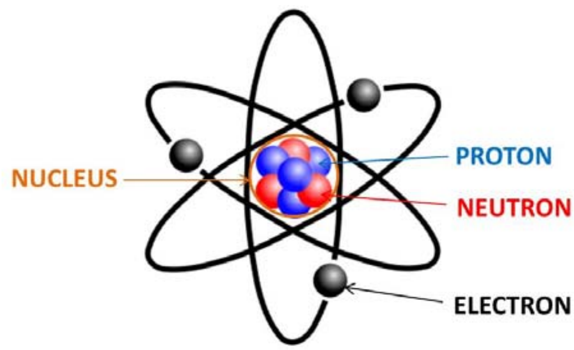
The physics that we have presented thus far in this course is classified as **Classical Physics**. Classical physics encompasses all of the physics developed prior to the 20th century. This includes all of the work of Galileo, Newton, Faraday, and Maxwell, and represents a tremendous triumph in human history. Classical physics includes the laws of mechanics, electricity and magnetism, thermodynamics, optics, and fluids. Classical physics provided a firm foundation for our understanding of the Universe in terms of Newton's law of gravity. The foundations for the great conservation laws of energy and momentum were laid during the heyday of classical physics. Classical physics also enabled the development of the industrial revolution through the introduction of the steam engine and electrical motor.

By the end of the 19th century, there was a general belief that the laws of physics were essentially known. It is important to understand that classical physics dealt mostly with the behavior of large (macroscopic) objects. However, at the end of the 19th century, the concept that matter was composed of microscopic entitles – atoms, became accepted. It was logical for physicists to attempt to understand the behavior of atoms in terms of the laws of classical physics. It would soon become apparent that the laws of classical physics led to predictions that were wrong when applied to atomic scale phenomena. Also, some of the most basic ideas of space and time on which Newtonian mechanics was based, came under question with the introduction of Einstein's theory of special relativity in 1905.

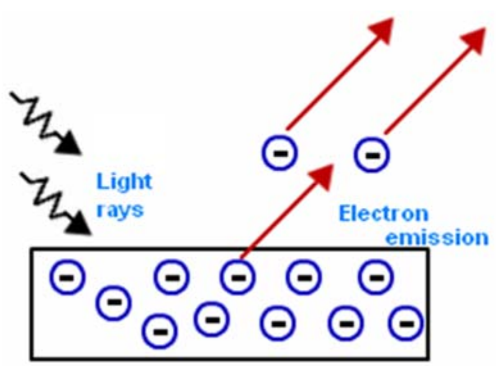
The first quarter of the 20th century was a period of great upheaval in physics when it became necessary to establish new physical principles for dealing with atomic-level phenomenon. 20th century physics is generally referred to a **Modern Physics** and includes the study of atomic and nuclear physics. We will devote the final 4 lectures of this course to topics in modern physics. Note that the introduction of the concepts of Modern Physics does not require that we abandon the principles of Classical Physics. The principles of Classical Physics continue to be valid in their restricted range of applicability. For example, the principles of Classical Physics were the foundation of our ability to put a man on the Moon. The principles of Modern Physics extend

Classical Physics into the microscopic world of atoms and nuclei. As a general principle, we require that the new laws include the old ones within their range of applicability.

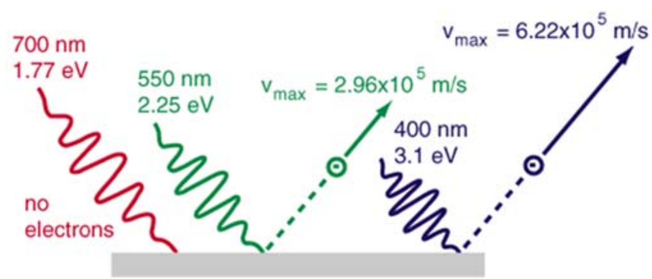
33-1. The Problem of the Atom in Classical Physics.—An **atom** is the smallest constituent unit of ordinary matter that has the properties of a chemical element. The typical picture of the atom (right) showing electrons orbiting around the nucleus results from the analogy with the planets orbiting around the sun. However, this simple picture turns out to be in serious contradiction with the laws of classical physics. Electrons in orbit around the nucleus would experience a centripetal acceleration. According to the laws of electromagnetism, any charged particle that experiences acceleration radiates electromagnetic waves. Thus an orbiting electron would continuously radiate energy, and as a result would collapse into the nucleus. This would occur on a very rapid timescale, making the existence of stable atoms impossible – classically, there could be no atoms! The resolution of this, and other problems in classical physics led to a revolution in our thinking about light and matter, particles and waves. Another problem, arising from the application of classical ideas to new problems at the atomic level was the photoelectric effect, which we will now discuss in detail.



33-2. The Photoelectric Effect.—The figures below illustrate the experimental facts of this phenomenon. When light waves hit a metal surface, electrons may be ejected from the surface



(a)



(b)

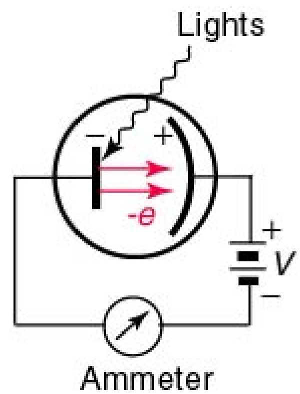
if the wavelength of the light is shorter than some critical value. The ejected electrons are referred to as ***photoelectrons*** because they are associated with light hitting the surface. As shown in (b) red light does not produce photoelectrons, while green and blue light do not produce photoelectrons. Measurements of the velocities of the photoelectrons indicate that higher energy photoelectrons are ejected by shorter wavelengths of light. The fact that photoelectrons are emitted only when light of a wavelength shorter than some critical value falls on a metal is at odds with classical physics. The resolution of this problem changed our thinking about how light interacts with matter in a profound way. (This was done by Einstein, for which he won the Nobel Prize in 1921.)

According to electromagnetic wave theory, if the intensity of light (the amount of light energy per unit time per unit area) falling on the surface were sufficiently high, photoelectrons would be produced regardless of the wavelength of the light. Einstein proposed that when light interacted with matter, instead of behaving like a wave, light behaves like a particle that has an amount of energy that depends on its wavelength (or frequency – recall that frequency and wavelength are not independent but are related by $\lambda f = c.$). When light interacts with matter, it can be thought of as a beam of particles called photons. **Photons are packets (called *quanta*) of electromagnetic energy moving at the speed of light.** The energy of a photon is given by

Photon Energy $$E_{photon} = hf = \frac{hc}{\lambda}, \text{ since } f = \frac{c}{\lambda}, \quad [1]$$

where h is a constant, called Planck's constant. (since h and c are constants hc is a constant also). Thus, light having a shorter wavelength has more energy.

How does the photon concept explain the photoelectric effect? An electron must absorb energy from the photon beam to escape from a metal surface. This can only occur if the photon has sufficient energy, so that when it interacts with the electron it transfers its energy to the electron. Since the photon energy is $E = hf = hc/\lambda$, the photon must have a short enough wavelength (or high enough frequency). Multiple photons of lesser energy do not cause the electron to pop out of the metal. An electron will only pop out when a single photon of sufficient energy interacts with it. If the photon has more than the minimum energy needed, the extra energy is given to the electron as kinetic energy --- the velocity of the ejected electron is higher.

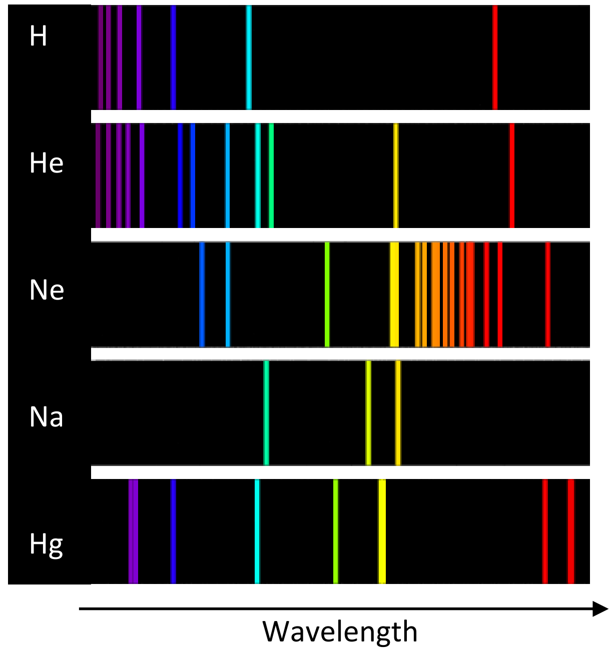


The photoelectric effect is used in **photocells** (slide 8) as safety devices. An example of a photocell is shown on the right. When light of appropriate wavelength falls on the photo-sensor, electrons are emitted and current flows to a positively biased electrode. As long as the light beam falls on the sensor the current will flow and is detected by an electronic circuit. If the light beam is interrupted, the current stops, and this condition is sensed by the circuitry. This device is used to prevent a garage door from closing if a person or animal is blocking the beam. (slide 9) Photocells are also used as devices to detect light. For example, photocells are used to turn outside lighting on and off. During day when the sun shines, the photocell sense current and uses this to prevent the outside lights from turning on. During the evening and night, the lack of photocurrent is used to indicate that the outside lights should be on. The photoelectric effect essentially converts photons to electrons and is the basis for the cameras used in cell phones. (slide 16).

33-3. The Quantum Concept.—The correct explanation of the photoelectric effect in terms of the concept of the photon was a radical departure from classical ideas. Prior to this, light was considered as a wave not a particle. The photon concept showed that in some circumstances, **when light interacts with matter, it must be considered as a particle.** The radical concept is that we now must think of **energy in quantized packets**. Classically, energy can be found in continuous amounts. In the realm of Modern Physics, energy is concentrated in definite, discreet amounts called ***quanta***, and we say that **energy is quantized**. In the photoelectric effect, energy is absorbed by the electrons only in **discreet amounts** given by the photon energy E = hf.

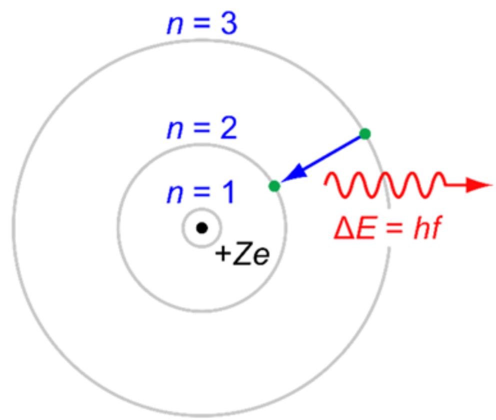
33-4. The Modern Physics of the Atom

a. Atomic emission spectra.—At the beginning of this lecture, we discussed the fact that the concept of the atom was at odds with classical physics. In the classical picture, atoms would have extremely short lifetimes because the electrons would quickly lose all of their energy in the



emission of electromagnetic radiation. Now, **atoms do emit electromagnetic radiation** (the light produced by fluorescent lamps is emitted by excited atoms in a gas discharge), but the details of the light emission from atoms as observed experimentally indicates that it is **not due to electrons spiraling inward to the nucleus**. If this were the case, we would find that the light from atoms would be emitted with a continuous spectrum of wavelengths. (Recall from the discussion of diffraction, that diffraction gratings are used to separate light into its wavelength components. Diffraction grating are used to obtain the spectrum (intensity vs. wavelength) of light emission.) However, when the spectra of light from various gases are obtained, we find that the **light is emitted in discreet spectral lines**. An example of the line spectra of various atoms is shown on the right. **The spectra consist of very narrow lines at specific wavelengths**. The black spaces indicate the absence of light at those wavelengths. Each element produces a unique spectra that is essentially its fingerprint. Notice that the spectrum sodium exhibits two strong yellow lines. This emission from sodium is responsible for the characteristic yellowish glow of low pressure sodium lights used for street illumination. A sample of an unknown material can be analyzed spectrally to determine its composition. This is a common technique of analytical chemistry and forensic science.

b. The Bohr model of the atom.—**Neils Bohr**, a Danish physicist, proposed a theory of the atom in 1913 based on the quantum concept (energy in discreet quantities). The observation of **discreet spectral lines** was a key element of Bohr's model of the atom. The problem with the classical picture of the atom is that orbiting electrons would quickly radiate away all of their



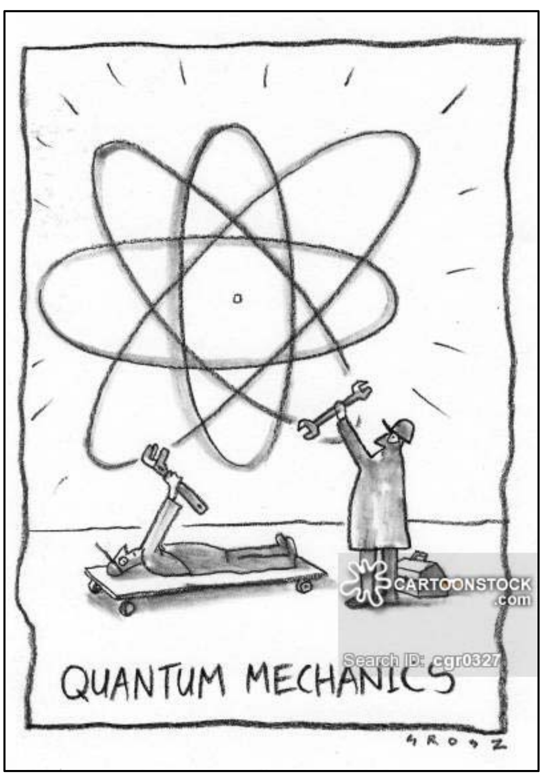
energy and collapse into the nucleus. Bohr made a bold assumption that the electrons in an atom can only be in certain **allowed stationary orbits or states**, and in these allowed states they **would not radiate**. The model proposed by Bohr is illustrated in the diagram below. The nucleus of the atom contains Z protons each having a charge +e. Three of the stationary states of the electron are indicated and are labeled as n = 1, 2, and 3. Electrons in the states of lower n values have lower energies. The states further from the nucleus (high n values) are higher energy states. To account for the discreet spectra of light emission from the atom, Bohr introduced another assumption. If an atom is in a high energy stationary state (high n), it can spontaneously make a transition to a lower energy (lower n) stationary, by **emitting a photon whose energy is exactly the energy difference between the high energy state and low energy state.** If the **energy difference** between the two states is $\Delta E,$ then the **frequency and wavelength** of the emitted radiation is given by $\Delta E = hf = hc/\lambda,$ where $\Delta E = E_i - E_f$, and E_f is the energy of the final state, and E_i is the energy of the initial state. The spectrum of light emitted by electrons making transitions from high to low energy states is called an **emission spectrum**. An electron in a low energy state could make a transition to a higher energy state if it absorbed a photon with an energy at lease as high as the difference between the low and high energy states. This process is called **absorption** and is an important part of the production of laser light that will be discussed in the next lecture.

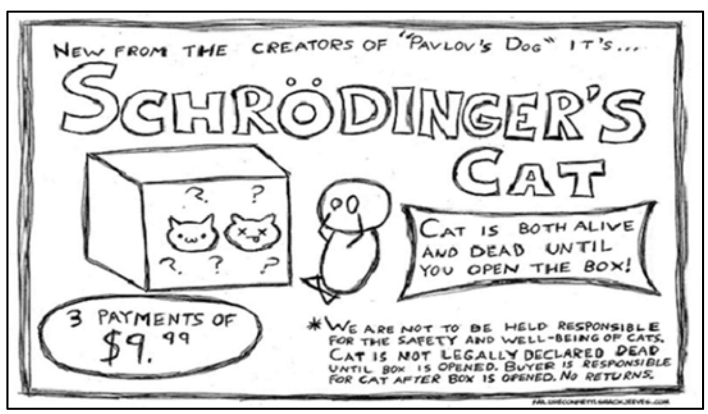
Bohr's model was successful in reproducing the **exact wavelengths of all of the spectral lines of hydrogen.** His model also could predict the spectrum of singly ionized helium (helium atoms with one electron removed. Bohr's model cannot be applied to atoms with more than one electron, because it does not take into account the electrostatic repulsion between the negatively charged electrons. Also, electrons have another property, spin, which could not be included in this model. A more comprehensive model of atomic scale systems, **Quantum**

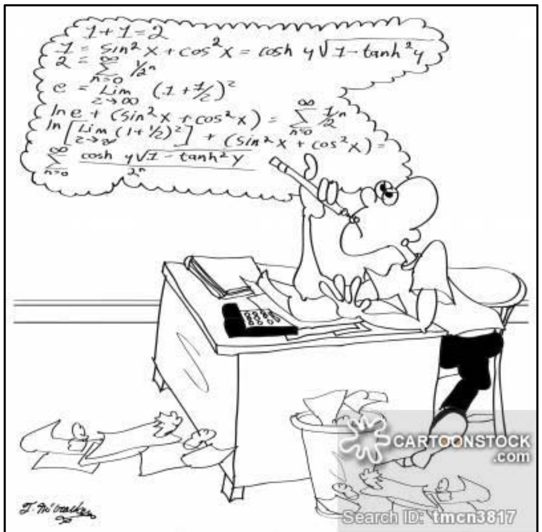
Mechanics, was developed in the late 1920’s by Schrodinger and Heisenberg. Quantum Mechanics replaces Classical Mechanics as the correct theory to explain atomic level phenomena.

c. *Quantum mechanics*.—The modern ideas of the atom is a radical departure from classical physics in that it requires that certain quantities, like energy are quantized, i.e., can take on only certain values. **Quantum mechanics** is the new theory of modern physics that replaced classical physics as the correct set of rules for microscopic phenomenon. Quantum mechanics was developed in the 1920’s by Schrodinger, Heisenberg, and Dirac.

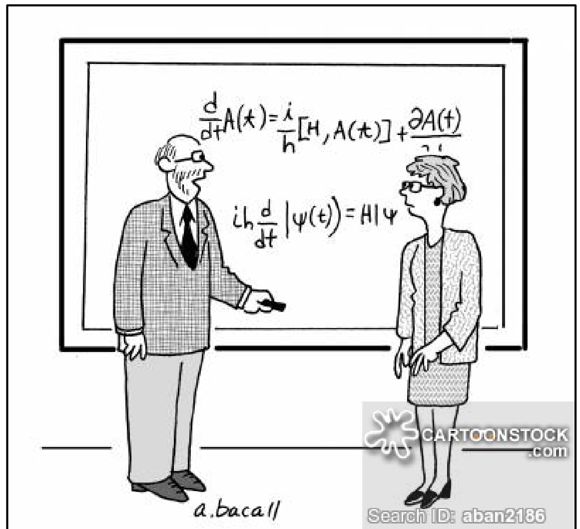
Quantum Mechanics contains a fundamental concept that not all variables can be measured simultaneously with arbitrary accuracy – this is known as the **uncertainty principle** and is discussed in the next lecture. Many of the ideas of Quantum Physics are non-intuitive based on the classical ways of thinking about things. However, Quantum Mechanics (the rules of Quantum Physics) are accepted since it leads to predictions that are in agreement with observations – this is, of course, the ultimate requirement of a theory. Some examples of quantum physics humor are given below.







WHILE TRYING TO FIGURE OUT IF INSURANCE WOULD COVER HER MEDICAL BILLS, FLO'S MIND WANDERED TO SIMPLER TOPICS LIKE QUANTUM PHYSICS.



"I can understand Heisenberg's equation and Schrodinger's equation for quantum mechanics but I cannot understand derivative trading."