## Planck's Quantum Theory

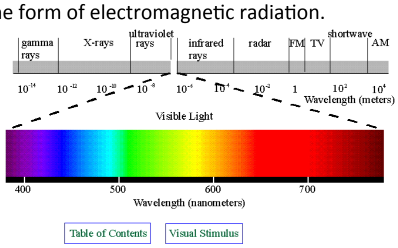
- Observations of radiation from a blackbody and its radiation (measured by spectroscopic lines) provided the first experimental evidence for quantum theory. Every time one sees a neon or sodium light, one is seeing quantum theory in practice. The light from a neon or sodium source is a spectroscopic line. An electric field excites atoms of the neon or sodium atom to a discrete quantum state; the atom then makes a transition by emitting light that is characteristic of the atom, and yields the particular color of light that one sees.

## Planck's Quantum Theory

- Furthermore, semiconductors and electronic chips in general exist due to quantum theory. Electronic devices, from computers, television, to mobile phones are all based on the semiconductor, and aeroplanes, ships, cars all use semiconductors in an essential manner. More complex technologies such as MRI (Magnetic Resonance Imaging), lasers, physical chemistry, fabrication of new drugs, modern materials science and so on all draw on the principles of quantum theory. It is no exaggeration to predict that twenty first century technology will largely be based on the principles of quantum physics.

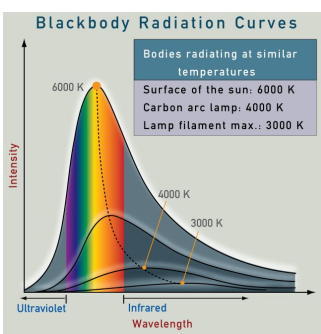
## Planck's Quantum Theory

- A black body which is maintained at a constant temperature T steadily loses energy from its surface in the form of electromagnetic radiation.




## Planck's Quantum Theory

- Since the atoms composing the black body are in contact with a heat bath at temperature , each atom has approximately amount of energy, where k = Boltzmann's constant. Since the atoms are jiggling around due to thermal motion, classical electromagnetic theory then predicts that **all** wavelength's of radiation, in particular up to infinitely short wavelengths, should be emitted by a black body.




## Planck's Quantum Theory

- This classical prediction for the spectrum of radiation that is emitted by such a black-body is contradicted by experiment. Max Planck, a German physicist, correctly explained the experimentally measured black- body spectrum by making the epoch-making conjecture in 1900 that electromagnetic waves are the macroscopic manifestations of packets of wave-energy called photons. Planck further made the quantum hypothesis that the energy of photons is **quantized** in the sense that the energy of the photons only comes in discrete packets, the smallest packet called a quantum.

## Planck's Quantum Theory

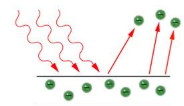
- Photons can have wavelength from zero to infinity. For a wave of frequency ν, or equivalently, of wavelength , the quanta of energy are given by ( is the velocity of light):

$E = h\nu = N\,hc/\lambda$, with N = 1, 2…

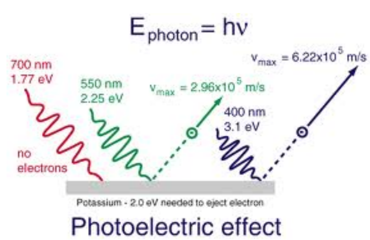
$h = 6.62618 \times 10^{-34}$ J·s

## The Photoelectric Effect

- Albert Einstein showed that light was a stream of particles (which would later be named *photons*). This phenomenon is called the *photoelectric effect*. When you shine a light upon certain metals, a stream of particles (later found to be electrons) is emitted from that metal. The emission has been found to have certain properties.



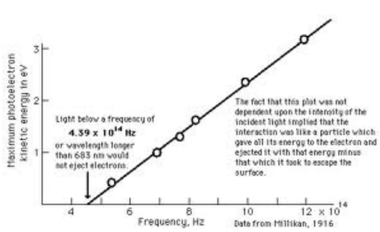
## The Photoelectric Effect




## The Photoelectric Effect

- The *number* of electrons emitted by the metal depends on the *intensity* of the light beam applied on the metal; more intense the beam, higher the number of electrons emitted.
- The emitted electrons move with greater *speed* if the applied light has a higher *frequency*.
- No electron is emitted until the light has a *threshold* frequency, *no matter how intense the light is*.

## The Photoelectric Effect




## The Photoelectric Effect

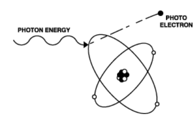
- These observations baffled physicists for many decades, since they cannot be explained if light is thought of only as a wave. If light were to be a wave, *both* the energy and the number of the electrons emitted from the metal should increase with an increase in the intensity of light. Observations contradicted this prediction; only the number, and not the energy, of the electrons increased with the increase of the intensity of the light.

## The Photoelectric Effect

- What Einstein showed was that the photoelectric effect as it had been observed could be explained if individual particles (or *quanta*) of light were penetrating the metal and knocking electrons loose from atoms. *According to Einstein's classic 1905 paper, increasing the intensity of the light increased the number of photons, while the energy of each individual photon remained the same, as long as the frequency of the light remained the same*. Therefore the number of electrons emitted would increase, but the energy transmitted to them by the particles of light would remain the same.

## The Photoelectric Effect

- In one stroke Einstein showed that light is a stream of particles, and also that there was solid evidence for the existence of quanta. His theory could satisfactorily explain all the known properties of the photoelectric effect, and was the first result derived from quantum theory of the interaction between radiation and matter.




## Bohr Model of Atom

- In 1913 Niels Bohr came to work in the laboratory of Ernest Rutherford. Rutherford, who had a few years earlier, discovered the planetary model of the atom asked Bohr to work on it because there were some problems with the model: According to the physics of the time, Rutherford's planetary atom should have an extremely short lifetime. Bohr thought about the problem and knew of the emission spectrum of hydrogen. He quickly realized that the two problems were connected and after some thought came up with the Bohr model of the atom. Bohr's model of the atom revolutionized atomic physics.

## Bohr Model of Atom

- The Bohr model consists of four principles: 1) Electrons assume only certain orbits around the nucleus. These orbits are stable and called "stationary" orbits. 2) Each orbit has an energy associated with it. For example the orbit closest to the nucleus has an energy E1, the next closest E2 and so on. 3) Light is emitted when an electron jumps from a higher orbit to a lower orbit and absorbed when it jumps from a lower to higher orbit. 4) The energy and frequency of light emitted or absorbed is given by the difference between the two orbit energies, e.g.,

## Bohr Model of Atom

- E(light) = $E_f - E_i$
- where "f" and "i" represent final and initial orbits.
- $E = h\nu = N\,hc/\lambda$, with N = 1, 2…
- $h = 6.62618 \times 10^{-34}$ J·s
- With these conditions Bohr was able to explain the stability of atoms as well as the emission spectrum of hydrogen. According to Bohr's model only certain orbits were allowed which means only certain energies are possible. These energies naturally lead to the explanation of the hydrogen atom spectrum.

## Bohr Model of Atom

- Bohr's model was so successful that he immediately received world-wide fame. Unfortunately, Bohr's model worked only for hydrogen. Thus, the final atomic model was yet to be developed.

## Bohr Model of Atom

- An electron in a given stationary state of a hydrogen atom, characterized by the quantum numbers , n, l and $m_l$, and , should, in principle, remain in that state indefinitely. In practice, if the state is slightly perturbed--*e.g.*, by interacting with a photon--then the electron can make a transition to another stationary state with different quantum numbers.

## Bohr Model of Atom

- Bohr found that the allowed E's of the electron are restricted by n according to:
- $E = - (1/n^2)\ 2.179 \times 10^{-18}$ J, n = 1, 2, 3,…
- $R = 2.179 \times 10^{-18}$ J or Rydberg's constant
- For n = 1, $E = 2.179 \times 10^{-18}$ J.

## Bohr Model of Atom

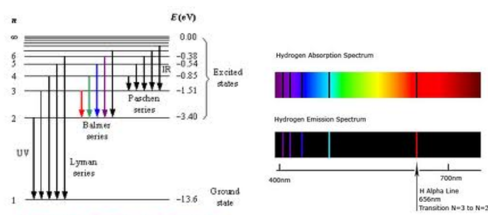
- Suppose that an electron in a hydrogen atom makes a transition from an initial state whose radial quantum number is $n_i$ to a final state whose radial quantum number is $n_f$. The energy of the electron will change by

$\Delta E = E_f - E_i = - 2.179 \times 10^{-18}\ J\ (1/n_f^2 - 1/n_i^2)$.

$E_{photon} = h\nu$ and $\nu = E_{photon} / h$;

$\lambda = c / \nu$

## Bohr Model of Atom




Energy levels of the hydrogen atom with some of the transitions between them that give rise to the spectral lines indicated

## Problem

- Calculate the wavelength of the spectral line when the electron in the hydrogen atom undergoes a transition from $4^{th}$ energy level to $2^{nd}$ energy level. What is the color of the radiation?