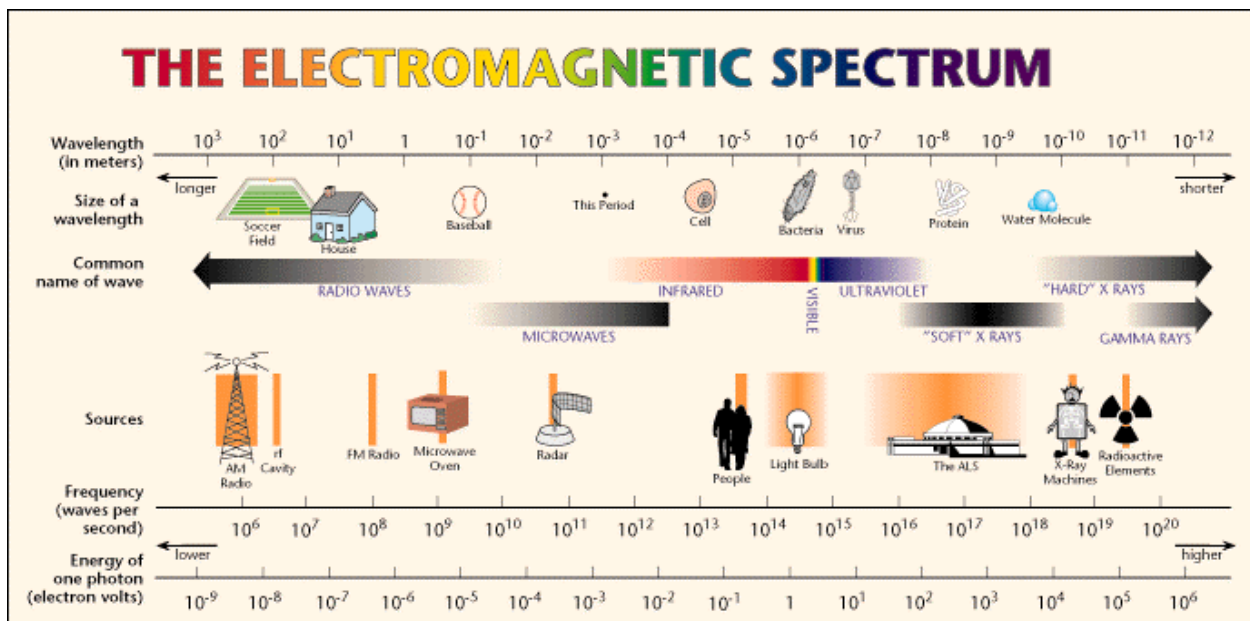# Lecture 1: Overview; History of Radiation

Radiation can be defined as the propagation of energy through matter or space. It can be in the form of electromagnetic waves or energetic particles.

Source: Lawrence Berkeley Laboratories, "MicroWorlds: Electrmagnetic Spectrum." [cited 26 March 2003.]
http://www.lbl.gov/MicroWorlds/ALSTool/EMSpec/EMSpec2.html

Ionizing radiation has the ability to knock an electron from an atom, i.e. to ionize.

- alpha particles
- beta particles
- neutrons
- gamma rays
- x-rays

Non-ionizing radiation does not have enough energy to ionize atoms in the material it interacts with.

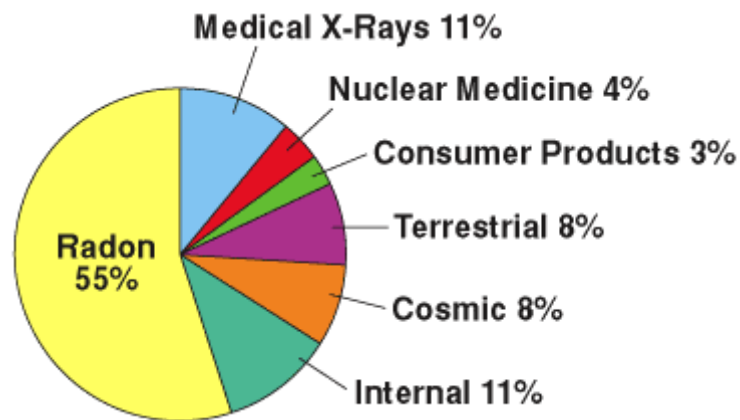- microwaves
- visible light
- radio waves

- TV waves
- Ultraviolet radiation (except for the very shortest wavelengths)

Health effects depend on how radiation interacts with biological material at the microscopic level.

Radiation delivers energy in "small packets", not uniformly distributed throughout the entire mass.

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Fig. 1.4 in Hall, Eric J. *Radiobiology for the Radiologist*, 5th ed.
Philadephia PA: Lippincott Williams & Wilkins, 2000.

Sources of Radiation Exposure to the US Population
(from U.S. NRC, *Glossary: Exposure*. [updated 21 July 2003, cited 26 March 2004]
http://www.nrc.gov/reading-rm/basic-ref/glossary/exposure.html

1895: The Discovery of the X-ray

Radiation was discovered by Wilhelm Conrad Roentgen on November 8, 1895.

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Crookes/Hittorf Tube

Image removed.

First "Roentgen Ray" image: December 22 1895

Image removed.
One of the first images obtained by Rontgen using x-rays, showing the bones of a hand.

1896: > 1000 papers published on X Rays.

Roentgen received the first Nobel Prize in physics in 1901

What Happens in the Crookes/Hittorf Tube?

1. Phosphorescence/ Fluorescence: <u>**The glow from the tube**</u>

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Fluorescence: rapid (10^{-8} – 10^{-9} sec)

Phosphorescence: slow

Bremsstrahlung (X-rays): The Mysterious Rays:

Image removed.

Roentgen actually discovered "Bremsstrahlung" (braking radiation)

An electron hitting a target can do one of two things:

- **Collision with an orbital electron**

- **Deceleration in the field of the nucleus**

Charged particle changing acceleration emits electromagnetic radiation: x rays.

The nature of the electron

J.J. Thompson (1897)
- An object in the tube casts a shadow.
- A paddle wheel in the cathode ray tube turned, the ***electron has momentum***.
- These experiments represented the discovery of the ***electron as a particle***.
- Thompson measured the charge to mass ratio of cathode rays.
  - Measured the deflection of the cathode rays in an electric field.
  - $e/m = 1.76 \times 10^{11}$ C/kg
  - Ratio ~ 1700 times greater than that of the proton, i.e., the electron is much less massive than hydrogen.
- Thompson develops a model of the atom as a matrix of positive charge with embedded negative electrons (Plum pudding model).

Image removed.

1896: The Discovery of Radioactivity: Henri Becquerel

Image removed.

Becquerel working on fluorescence/phosphorescence

Begins testing other substances:
uranyl potassium sulfate on photographic plate in black paper placed in the sun…
….the sun was not necessary.

1903 Nobel Prize in physics

1897: The Discoveries of Marie Sklodowska Curie and Pierre Curie

Marie Curie: PhD project, repeated Becquerel's expt with U and Th

- Purifies substances from pitchblende

- More radioactive than uranium or thorium:…..polonium

- Continues to purify a chemically different substance (sulfides acid insoluble; calcium barium, soluble)

- Purifies sample of radium (the activity of 1.00 g pure radium will later be defined as 1 "curie".

Marie and Pierre Curie share the 1903 Nobel Prize in physics with H. Becquerel

Puzzling properties….

1900: Crookes
Iron hydroxide precipitated in a uranyl salt solution, the activity was in the ppt.

After a few days, the precipitate lost its activity, and the uranium reacquired it!!

1899/1900 Alpha, Beta and Gamma

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Rutherford: working with uranium, 2 kinds of radiation
    (α)   nature not clear at the time
    (β)   showed to be "cathode rays" electrons

    (γ) Villard in France, also working with uranium, discovers a very penetrating radiation, different from α or β.

Becquerel's plates were exposed by the gamma radiation from uranium

The Structure of the Atom

Hydrogen emission spectrum:

- White light entering an optical spectrometer produces the familiar rainbow of colors.
- Light emitted from the combustion of single elements produced a series of discrete lines. The line spectrum was characteristic of each element.

> Image removed.
> Fig 2.1 in Turner J. E. *Atoms, Radiation, and Radiation Protection*, 2nd ed. New York: Wiley-Interscience, 1995.

Johann Balmer (1885)

Balmer developed an empirical formula to describe the hydrogen emission spectrum.

$$\frac{1}{\lambda} = R_\infty \left(\frac{1}{2^2} - \frac{1}{n^2} \right)$$

$R_\infty = 1.09737 \times 10^7 \ m^{-1}$ the Rydberg constant,
n represents any integer greater than 2

e.g., when $n = 3$, $\lambda = 6562$ Å
 when $n = 4$, $\lambda = 4861$ Å
 the series converges to a limit in the ultraviolet.

Balmer (correctly) speculated that additional series would exist in the infrared and the ultraviolet.

> Image removed.
> Fig 4.2(b) in [Hall]

Ernest Rutherford (1909)

- By this time alpha, beta, and gamma radiation were known.
- Rutherford used 7.69 MeV alpha particles from a ^{214}Po isotopic source.
- Collimated alpha particles were directed at a thin metal foil (60 μm).
- Scintillation screen observed with microscope at different deflection angles.
- Occasional deflections at a large angle (1 in 8000 > 90º for platinum, 1 in 20,000 for gold)

An enormously strong electric or magnetic field would be required to reverse the direction of the relatively massive alpha particle.

"It was about as credible as if you had fired a 15-in shell at a piece of tissue paper and it came back and hit you" (E. Rutherford, 1911)

Image removed.
Fig 2.7 in [Hall]

The positively charged nucleus is heavy, dense and very small.

Rutherford developed the ***planetary model of the atom***: a small nucleus with electrons in orbits around the nucleus.

[Today we know that the nuclear radius is ~ $R = 1.3 \ A^{\frac{1}{3}} \times 10^{-15}$ m
e.g., gold: nuclear radius = $1.3 \ (197)^{\frac{1}{3}} \times 10^{-15}$ m = 7.56×10^{-15} m
atomic radius = 1.79×10^{-10} m, the ratio = 4.22×10^{-5} (the nucleus is a very small part of the atomic cross section).]

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Problems with the Rutherford model
- An object that does not move in a straight line is under acceleration
- An accelerated charge emits radiation and, thus, loses energy.

- An electron orbiting the nucleus should lose energy and spiral into the nucleus.

Max Planck (1900)

- Light as a wave could explain most, but not all, experimental observations. Notable exceptions were the *photoelectric effect*, and *blackbody radiation*.
- To explain blackbody radiation, Planck proposed that light energy is quantized.
- The energy of the light wave is an integral number of quanta.

$$E = h\upsilon$$

E = energy of 1 quantum,
υ is the frequency (s^{-1}),
h = 6.62 x 10^{-34} J-s (Planck's constant)

Albert Einstein (1905)

Einstein used Planck's quantum theory to explain the photoelectric effect,

$$E = h\upsilon, \quad E = h\frac{c}{\lambda}$$

the shorter the wavelength, the greater the energy of the quanta.

- These are the beginnings of the quantum theory and the wave-particle duality of light.

- The successful explanation of two experimental observations was strong support for this radical new approach.

- Plank received the Nobel Prize in 1918 for the theory that light energy is quantized.

- Einstein received the Nobel Prize in 1921 for his work on the photoelectric effect.

Niels Bohr (1913)

- Bohr puts forth a revolutionary theory than eventually leads to quantum mechanics.

- Bohr reasoned from the hydrogen spectral data that, if specific lines were observed, the atom is only able to emit certain wavelengths, which relate to changes only in allowed steps.

Bohr postulates that:
- The electron can exist only in certain orbits.
- Electrons in certain orbits are stable and do not emit energy.
- The energy levels arise because the electrons in the fixed orbits have different energies in each successive orbit.
- Photons are emitted or absorbed when electrons move from one energy level to another.
- The photon energy equals the difference between the two energy levels.
- Wavelength is related to energy by the Planck equation.

$$E_{photon} = E_{level\,x} - E_{level\,y} = h\upsilon = h\frac{c}{\lambda}$$

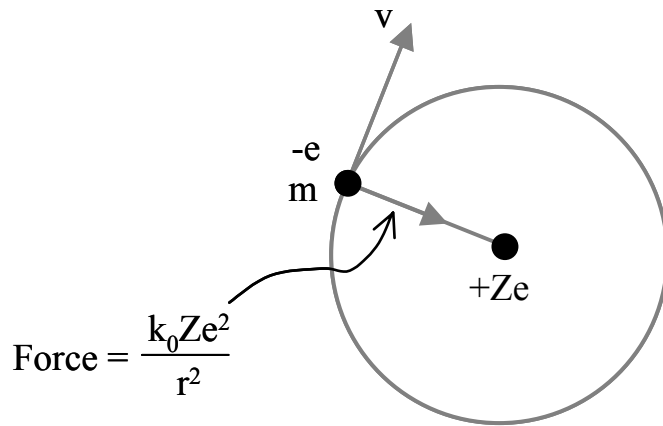- Stable states involve motion of electrons in circular orbits.

In order to produce the observed hydrogen spectral lines, it was required that the electron angular momentum be an integral multiple of $\frac{h}{2\pi}$.

$$mvr = n\frac{h}{2\pi}$$

$$\left(\hbar = \frac{h}{2\pi}\right)$$

n, which represents the electron energy levels, will become the principle quantum number.

The Bohr Model of the Hydrogen Atom

$$\text{Force} = \frac{k_0 Z e^2}{r^2}$$

Equation of motion for the electron

Forces on the electron

Circular motion: $\dfrac{mv^2}{r}$

Coulombic: $\dfrac{k_0 Z e^2}{r^2}$

The Bohr model gives the allowed radii of electron orbitals:

$$r_n = 5.29\,x10^{-11}\,\frac{n^2}{Z}\,m$$

The Bohr model gives the electron orbital velocities:

$$v_n = 2.19\,x10^6\,\frac{Z}{n}\,m\,s^{-1}$$

The Bohr model can be used to calculate the energy levels of the hydrogen electrons:

$$E = -13.6\,\frac{Z^2}{n^2}\,eV$$

The Bohr approach reproduced the Balmer empirical description of the hydrogen spectral lines, and provided an independent calculation of the Rydberg constant.

$$\frac{1}{\lambda} = 1.009737\,x10^7\,Z^2\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)m^{-1}$$