Atomic Structure and the Periodic Table

- **nuclear atom** – small, heavy, positive nucleus surrounded by a negative electron cloud
- **electronic structure** – arrangement of the electrons around the nucleus
- **classical mechanics** – fails in describing the electronic motion
- **quantum mechanics** – designed to describe the motion of microscopic particles

Observing Atoms

- **spectroscopy** – studies the interaction of light with matter (emission, absorption, scattering, …)
- **spectrometer** – instrument that separates the different colors of light and records their intensities
- **spectrum** – intensity of light as a function of its color (wavelength or frequency)
- **atomic emission spectrum** – the spectrum emitted by the atoms of an element when they are excited by heating to high temperatures (very characteristic for each element; used for identification of elements)
- **spectral lines** – images of the spectrometer entrance slit produced by the different colors in the spectrum

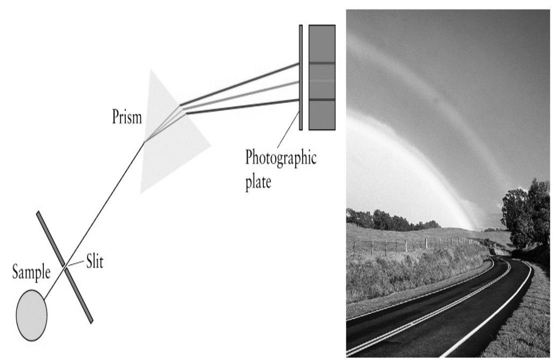

Fig. 7.1

7.1 The Characteristics of Light

- Light is **electromagnetic radiation** – a stream of energy in the form waves
- **Electromagnetic waves** – periodic oscillations of the electric and magnetic fields in space

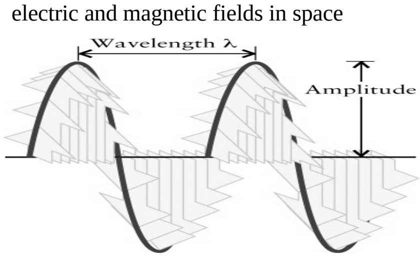

Fig. 7.2

- **Wavelength (l)** – distance between two adjacent minima or maxima of the wave
- **Frequency (n)** – number of oscillations of the electric (or magnetic) field per second
  - units - hertz (Hz) → **1 Hz = 1 s^{-1}**
- **Amplitude** – strength of the oscillation (related to the intensity of the radiation)
- **Speed of light (c)** – rate of travel of all types of electromagnetic radiation (3.00×10^8 m/s)

$$ln = c \qquad \uparrow l \rightarrow \downarrow n$$

- Classification of light – based on the values of *l* and *n*

Table 7.1 Color, frequency, and wavelength of electromagnetic radiation

Radiation type	Frequency, 10^{14} Hz	Wavelength, nm (2 sf)	Energy per photon, 10^{-19} J
x-rays and γ rays	$\geq 10^3$	≤ 3	$\geq 10^3$
ultraviolet	8.6	350	5.7
visible light			
violet	7.1	420	4.7
blue	6.4	470	4.2
green	5.7	530	3.8
yellow	5.2	580	3.4
orange	4.8	620	3.2
red	4.3	700	2.8
infrared	3.0	1000	2.0
microwaves and radio waves	$\leq 10^{-3}$	$\geq 3 \times 10^6$	$\leq 10^{-3}$

Example: What is the wavelength of light with frequency **98.9 MHz**.

98.9 MHz = 98.9×10^{6} Hz = = 98.9×10^{6} s^{-1}

$$l = \frac{c}{n} = \frac{3.00\times10^{8}\ \text{m/s}}{98.9\times10^{6}\ \text{s}^{-1}} = $$
$$= 3.03\ \text{m}$$

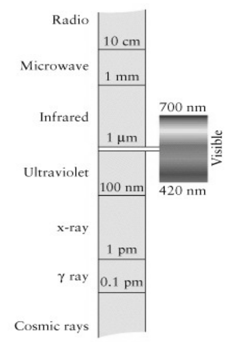

Fig. 7.6

7.2 Quanta and Photons

- Classical mechanics – the energy of an object can be continuously changed
- Quantum mechanics – the energy of an object can be changed (or transferred) only in discrete amounts (**quanta**) – **quantization of energy**
- **Quanta** - the smallest portions of energy
- **Dual nature of light** – light has both wave and particle like properties
  - wave (interference, diffraction)
  - particle (photoelectric effect)

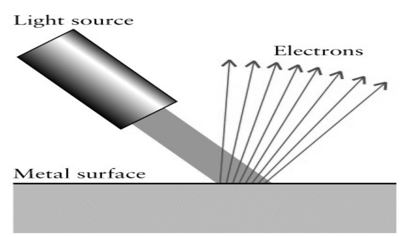

Fig. 7.7

- **Photoelectric effect** – ejection of **e$^-$** from metals by irradiation with light
- Ejection of **e$^-$** begins only above a certain **threshold frequency** (below this frequency, no ejection occurs no mater how intense the light is)

- Explanation – the ejection is caused by particles (**photons**) with energy proportional to the frequency of the radiation (only photons with enough energy can eject electrons)
- Energy of the photon (E_{ph})

$E_{ph} = hn$ $\quad n = c/l \quad$ $E_{ph} = hc/l$

h **– Planck's constant** $\quad$ ***h*** **= 6.626×10^{-34} J·s**

Example: Calculate the energy of a photon of light with wavelength **514 nm**.

$$E_{ph} = h\frac{c}{l} = 6.626\times10^{-34}\,\text{J}\cdot\text{s}\,\frac{3.00\times10^{8}\,\text{m/s}}{514\times10^{-9}\,\text{m}} = 3.87\times10^{-19}\,\text{J}$$

7.3 Atomic Spectra and Energy Levels

- Atomic emission spectra are **line spectra** – consist of discrete frequencies (lines)
- The **Rydberg equation** – fits the observed lines in the hydrogen atomic emission spectrum

$$n = R_H \times \left(\frac{1}{n_1^2} + \frac{1}{n_2^2} \right)$$

n_1 **= 1, 2, 3, …** $\quad$ n_2 **=** n_1 **+ 1,** n_1 **+ 2,** n_1 **+ 3, ...**

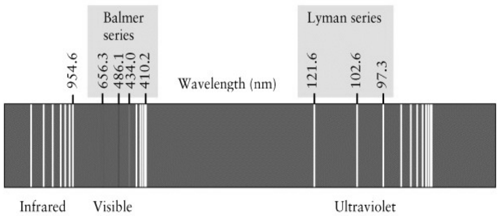

Fig. 7.9

- Lyman series (UV) – $\boldsymbol{n_1}$ **= 1** and $\boldsymbol{n_2}$ **= 2, 3, 4, ...**
- Balmer series (VIS) – $\boldsymbol{n_1}$ **= 2** and $\boldsymbol{n_2}$ **= 3, 4, 5, …**
- Paschen series (IR) – $\boldsymbol{n_1}$ **= 3** and $\boldsymbol{n_2}$ **= 4, 5, 6, ...**

- Explanation:
  - the energy of the electron in the **H** atom can have only certain discrete values (**energy levels**)
  - **transitions** between energy levels generate **photons** with energies equal to the differences between the levels (***DE***)

$$E_{ph} = DE$$

$$Þ\ DE = hn$$

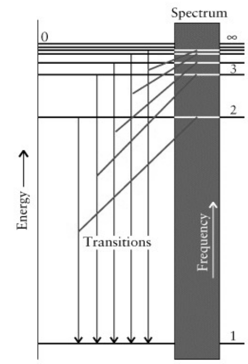

Fig. 7.11