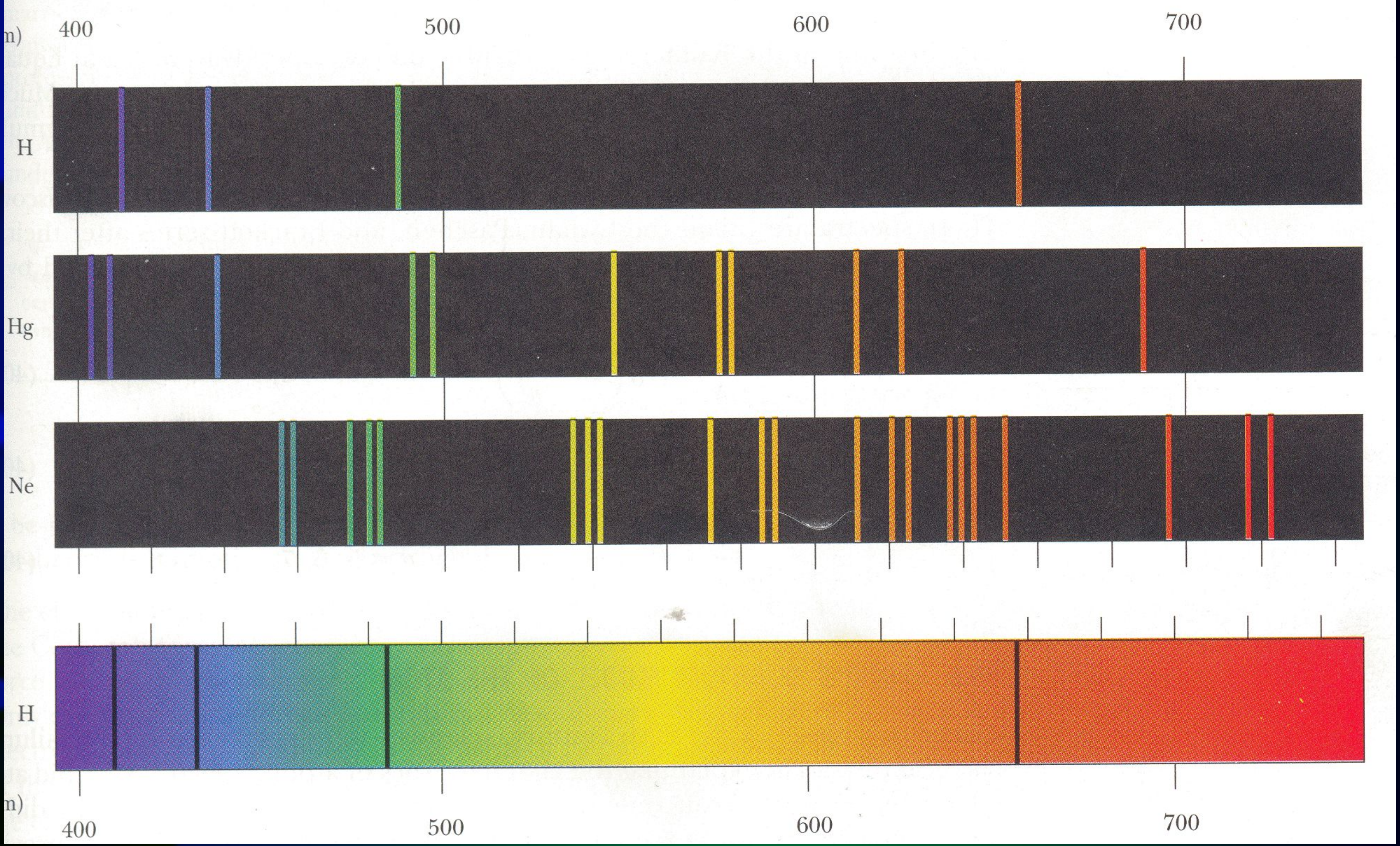
Orbit quantization rule

Lecture №3

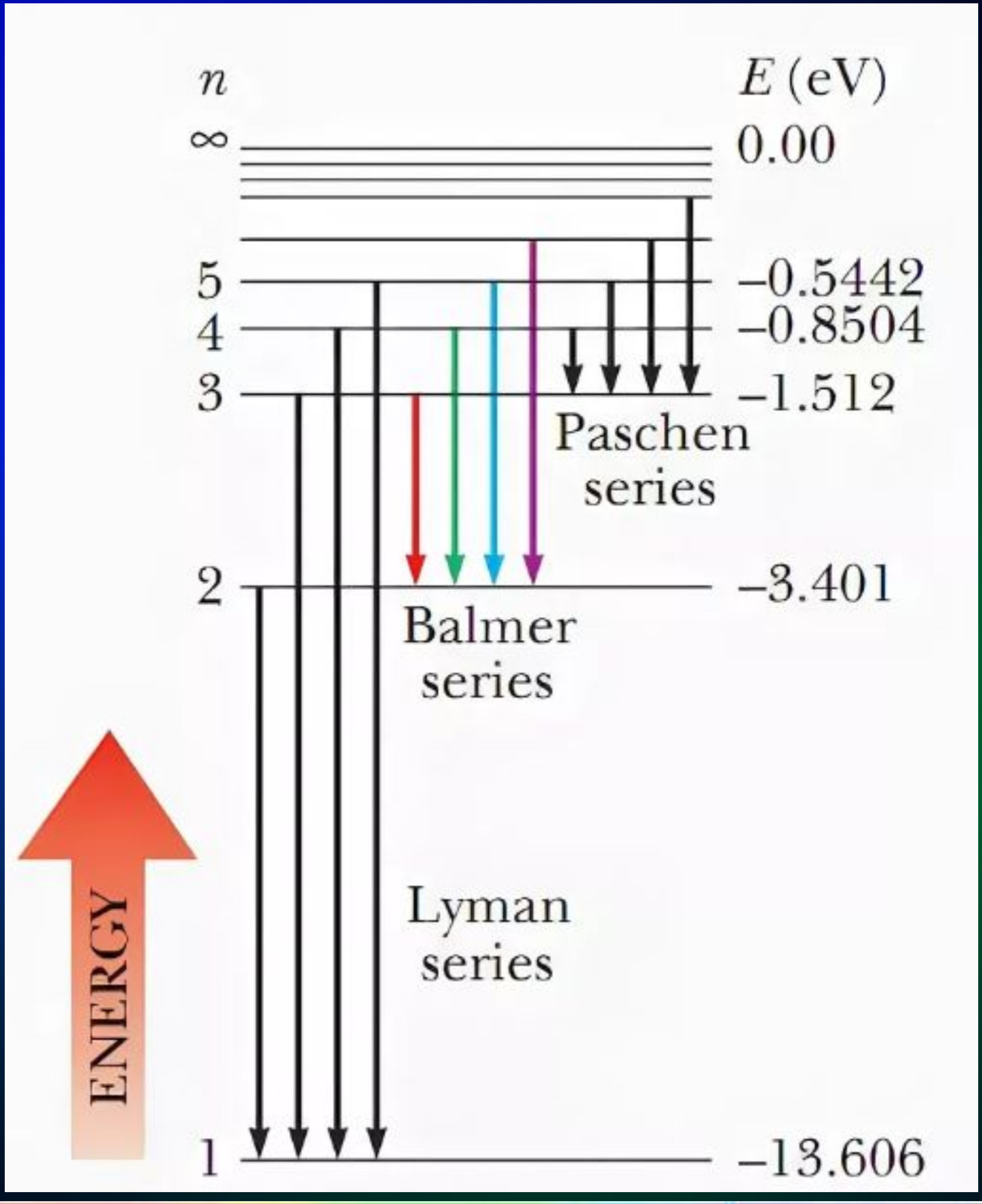
-Isolated atoms in the form of rarefied gas or metal vapors emit a spectrum consisting of separate spectral lines (line spectrum).

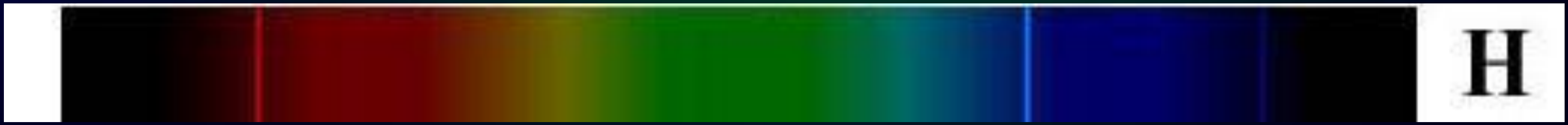
-Lines in the spectra are not randomly distributed, they located in series.

-The distance between the lines in the series decreases as the transition from long waves to short waves.



Line spectra of radiation in the visible region: hydrogen, mercury, neon. The spectrum of hydrogen absorption


n
E (eV)
∞
0.00
5
–0.5442
4
–0.8504
3
–1.512
Paschen series
2
–3.401
Balmer series
ENERGY
Lyman series
1
–13.606


H

$R' = 1{,}09 \cdot 10^{7} m^{-1}$ – Rydberg constant.

$R = R' \cdot c.$

$$R = 3{,}29 \cdot 10^{15}\, s^{-1}$$

The rule for quantizing orbits: from all the orbits of an electron only those are possible for which the angular momentum is equal to an integral multiple of the Planck constant

$$m_e v r = n\hbar$$

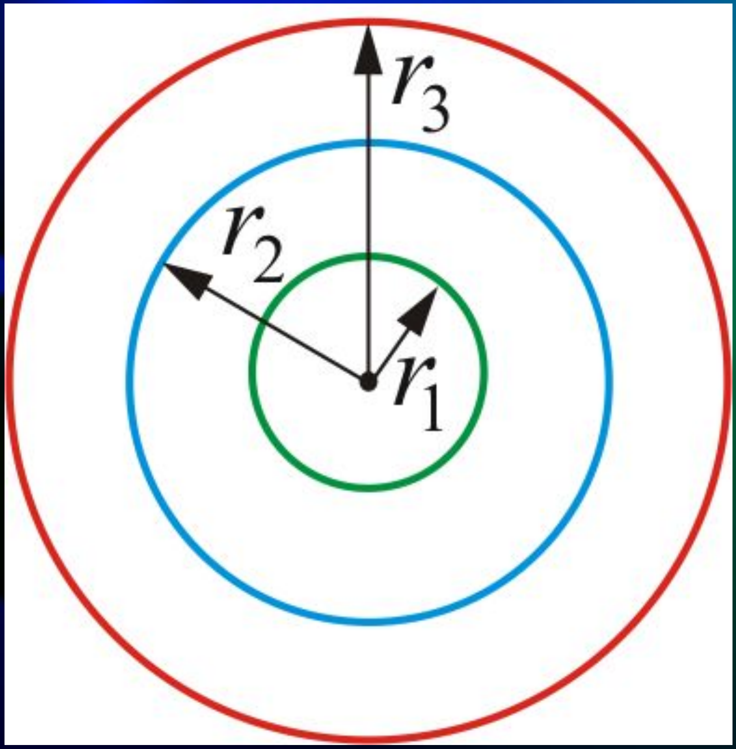
where n = 1, 2, 3,… principal quantum number.

The equation of motion of the electron is obtained from the equality of the centrifugal force to the Coulomb force:

$$m_e \frac{v^2}{r} = \frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{r^2} \Rightarrow r_n = \frac{h^2 n^2 4\pi\varepsilon_0}{m_e Z e^2}$$

The radius of the first orbit of the hydrogen atom is called the Bohr radius.

At n =1, Z = 1 for Hydrogen:



$$r_1 = 4\pi\varepsilon_0 \frac{h}{m_e e^2} = 0{,}529\ \mathring{A} =$$

$$= 0{,}529 \cdot 10^{-10}\ m.$$

The internal energy of the atom is made up of the kinetic energy of the electron (the nucleus is motionless) and the potential energy of the interaction of the electron with the nucleus.

$$E = \frac{m_e \upsilon^2}{2} - \frac{Ze^2}{4\pi\varepsilon_0 r}$$

From the equation of motion of the electron it follows that:

$$\frac{m_e \upsilon^2}{2} = \frac{Ze^2}{2r}$$

– the kinetic energy is equal to the potential energy.

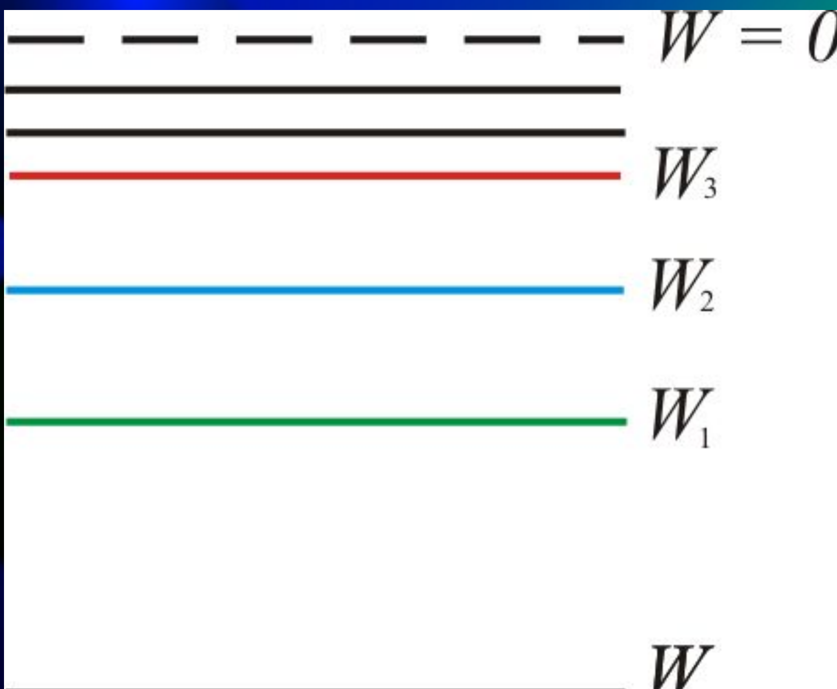
For hydrogen atom

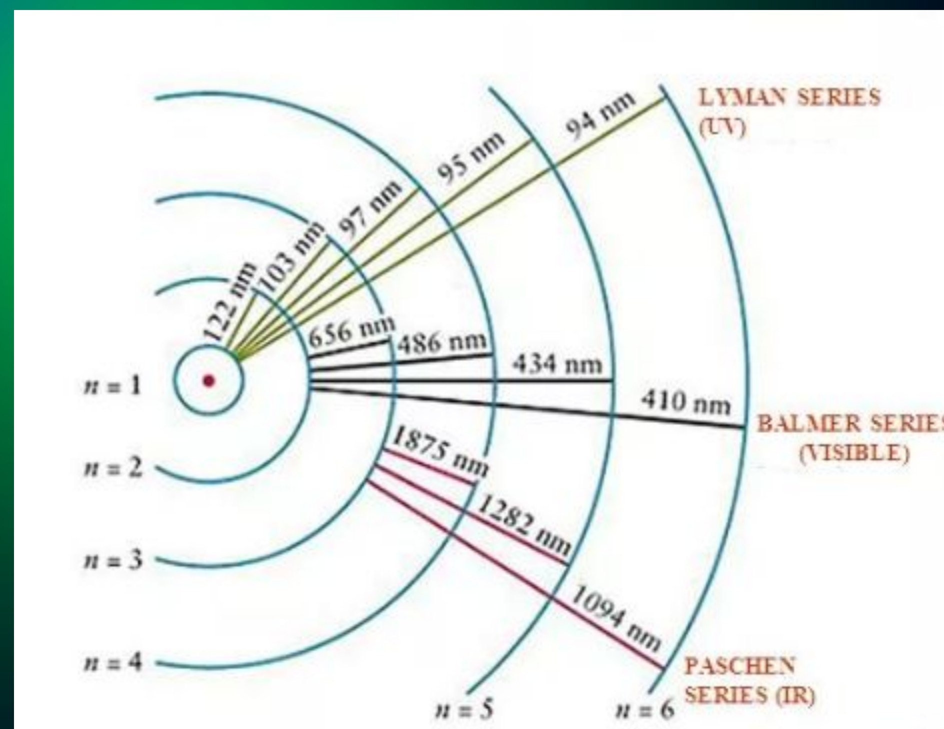
$$W_n = -\frac{m_e e^4}{8h^2\varepsilon_0^2}\frac{1}{n^2}$$

W_n takes only discrete values of energy n = 1, 2, 3….

The scheme of energy levels determined

$$\left(W_n = -\frac{Z^2 m_e e^4}{8h^2\varepsilon_0^2}\frac{1}{n^2}\right)$$ shown in figure





At the transition of an electron in a hydrogen atom from state n to state k, a photon with energy:

$$h\nu = -\frac{m_e e^4}{8h^2\varepsilon_0^2}\left(\frac{1}{n^2}-\frac{1}{k^2}\right)$$

is emitted and radiation frequency

$$\nu = \frac{m_e e^4}{8h^3\varepsilon_0^2}\left(\frac{1}{k^2}-\frac{1}{n^2}\right)$$

We obtain a generalized Balmer formula, which agrees well with experiment, where the Rydberg constant

$$R = \frac{me^4}{8\varepsilon_0^2 h^3}$$

The success of Bohr's theory:

-calculation of the Rydberg constant for hydrogen-like systems;
-explanation of the structure of their line spectra.

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Disadvantages of Bohr's theory:

1) the internal contradiction of the theory: **the mechanical combination of classical physics with quantum postulates.**

2) the theory could not explain the question of the intensities of the spectral lines.

3) the absolute inability to apply the theory to explain the spectra of helium (He) (two electrons in orbit, and Bohr's theory does not cope).