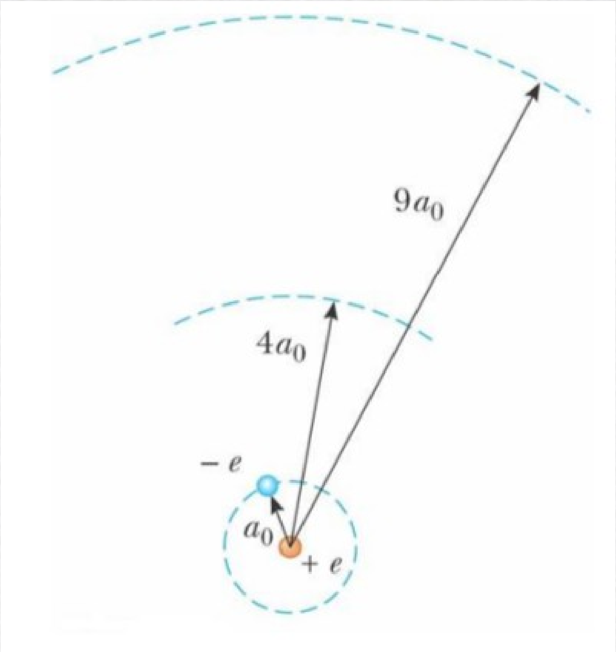
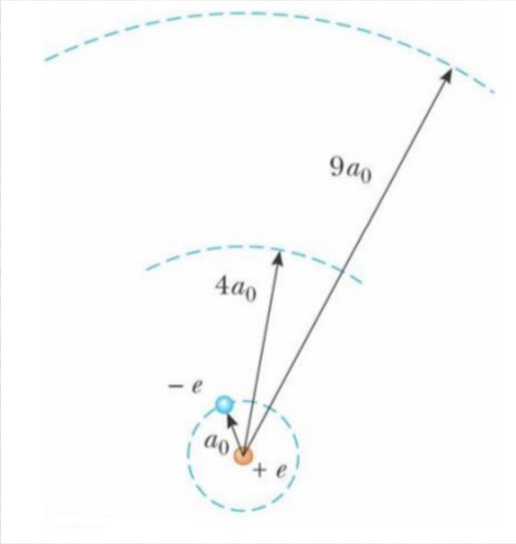
Radios permitidos



$$r_n = n^2 a_o$$

Radios permitidos y energía total
Cuantización de la radiación

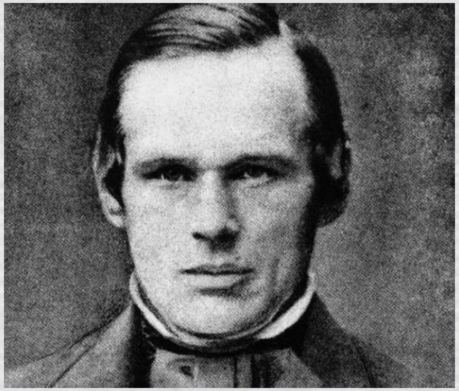


$$E_n = -\ 13.6\ \frac{Z^2}{n^2}\ eV$$

$$\boxed{r_n = n^2 a_o}$$

$$\boxed{\frac{1}{\lambda_{ik}} = RZ^2 \left| \frac{1}{n_i^2} - \frac{1}{n_k^2} \right|}$$

Un poquitín de historia …



A. **Ångström** (1853)

4 líneas del H

(656.279, 486.133, 434.047, 410.174 nm)

Un poquitín de historia …

J. **Balmer** (1885)

656.279, 486.133, 434.047, 410.174 n

$$\frac{9b}{5}, \frac{4b}{3}, \frac{25b}{21}, \frac{9b}{8} \qquad b = 364.60 \text{ nm}$$

Un poquitín de historia …

J. **Rydberg** (1890)

$$\frac{9b}{5}, \frac{4b}{3}, \frac{25b}{21}, \frac{9b}{8} \qquad b = 364.60 \text{ nm}$$

$$\boxed{\frac{1}{\lambda} = R\left(\frac{1}{n^2} - \frac{1}{m^2}\right)}$$

Recordemos que Bohr aún NO **había presentado su modelo atómico**

(violet, Hδ). In 1885 a Swiss school teacher, Johann Balmer (1825–1898), had found, by trial and error, a formula to reproduce the wavelengths of these spectral lines of hydrogen, today called the **Balmer series** or **Balmer lines**:

$$\frac{1}{\lambda} = R_H \left(\frac{1}{4} - \frac{1}{n^2} \right), \tag{5.7}$$

where $n = 3, 4, 5, \ldots$, and $R_H = 1.09677583 \times 10^7 \pm 1.3\ \mathrm{m}^{-1}$ is the experimentally determined Rydberg constant for hydrogen.[7] Balmer's formula was very accurate, to within

Bohr (1913)

Solving this equation for the radius r shows that the only values allowed by Bohr's quantization condition are

$$\boxed{r_n = \frac{4\pi\epsilon_0\hbar^2}{\mu e^2}n^2 = a_0 n^2,} \tag{5.13}$$

where $a_0 = 5.291772083 \times 10^{-11}$ m $= 0.0529$ nm is known as the **Bohr radius**. Thus the electron can orbit at a distance of a_0, $4a_0$, $9a_0$, . . . from the proton, but no other separations are allowed. According to Bohr's hypothesis, when the electron is in one of these orbits, the atom is stable and emits no radiation.

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Inserting this expression for r into Eq. (5.11) reveals that the allowed energies of the Bohr atom are

Thus the total energy $E = K + U$ of the atom is

.4)

$$E = K + U = K - 2K = -K = -\frac{1}{8\pi\epsilon_0}\frac{e^2}{r}. \qquad (5.11)$$

Bohr (1913)

Solving this equation for the radius r shows that the only values allowed by Bohr's quantization condition are

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Inserting this expression for r into Eq. (5.11) reveals that the allowed energies of the Bohr atom are

$$\boxed{E_n = -\frac{\mu e^4}{32\pi^2\epsilon_0^2\hbar^2}\frac{1}{n^2} = -13.6\text{ eV}\frac{1}{n^2}.} \tag{5.14}$$

If the electron does not radiate in any of its allowed orbits, then what is the origin of the spectral lines observed for hydrogen? Bohr proposed that a photon is emitted or absorbed when an electron makes a transition from one orbit to another. Consider an electron as it "falls" from a higher orbit, n_{high}, to a lower orbit, n_{low}, without stopping at any intermediate orbit. (This is *not* a fall in the classical sense; the electron is *never* observed between the two orbits.) The electron loses energy $\Delta E = E_{\text{high}} - E_{\text{low}}$, and this energy is carried away from the atom by a single photon. Equation (5.14) leads to an expression for the wavelength of the emitted photon,

$$\boxed{E_{\text{photon}} = E_{\text{high}} - E_{\text{low}}}$$

$$\boxed{E_{\text{photon}} = E_{\text{high}} - E_{\text{low}}}$$

$$\frac{hc}{\lambda} = \left(-\frac{\mu e^4}{32\pi^2\epsilon_0^2\hbar^2}\frac{1}{n_{\text{high}}^2}\right) - \left(-\frac{\mu e^4}{32\pi^2\epsilon_0^2\hbar^2}\frac{1}{n_{\text{low}}^2}\right),$$

$$\boxed{E_{\text{photon}} = E_{\text{high}} - E_{\text{low}}}$$

$$\frac{hc}{\lambda} = \left(-\frac{\mu e^4}{32\pi^2\epsilon_0^2\hbar^2}\frac{1}{n_{\text{high}}^2}\right) - \left(-\frac{\mu e^4}{32\pi^2\epsilon_0^2\hbar^2}\frac{1}{n_{\text{low}}^2}\right),$$

$$\frac{1}{\lambda} = \frac{\mu e^4}{64\pi^3\epsilon_0^2\hbar^3 c}\left(\frac{1}{n_{\text{low}}^2} - \frac{1}{n_{\text{high}}^2}\right). \tag{5.15}$$

$$\boxed{E_{\text{photon}} = E_{\text{high}} - E_{\text{low}}}$$

$$\frac{hc}{\lambda} = \left(-\frac{\mu e^4}{32\pi^2\epsilon_0^2\hbar^2} \frac{1}{n_{\text{high}}^2} \right) - \left(-\frac{\mu e^4}{32\pi^2\epsilon_0^2\hbar^2} \frac{1}{n_{\text{low}}^2} \right),$$

$$\frac{1}{\lambda} = \boxed{\frac{\mu e^4}{64\pi^3\epsilon_0^2\hbar^3 c}} \left(\frac{1}{n_{\text{low}}^2} - \frac{1}{n_{\text{high}}^2} \right). \tag{5.15}$$

this term is exactly the Rydberg constant for hydrogen:

$$R_H = \frac{\mu e^4}{64\pi^3\epsilon_0^2\hbar^3 c} = 10967758.3 \text{ m}^{-1}.$$

$$\boxed{E_{\text{photon}} = E_{\text{high}} - E_{\text{low}}}$$

$$\frac{hc}{\lambda} = \left(-\frac{\mu e^4}{32\pi^2\epsilon_0^2\hbar^2}\frac{1}{n_{\text{high}}^2}\right) - \left(-\frac{\mu e^4}{32\pi^2\epsilon_0^2\hbar^2}\frac{1}{n_{\text{low}}^2}\right),$$

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Ångström → Balmer → Rydberg → Bohr ...

After the quantum revolution, the physical processes responsible for Kirchhoff's laws (discussed in Section 5.1) finally became clear.

- A hot, dense gas or hot solid object produces a continuous spectrum with no dark spectral lines. This is the continuous spectrum of blackbody radiation emitted at any temperature above absolute zero and described by the Planck functions $B_\lambda(T)$ and $B_\nu(T)$. The wavelength $\lambda_{\max}$ at which the Planck function $B_\lambda(T)$ obtains its maximum value is given by Wien's displacement law, Eq. (3.15).
- A hot, diffuse gas produces bright emission lines. Emission lines are produced when an electron makes a downward transition from a higher orbit to a lower orbit. The energy lost by the electron is carried away by a single photon. For example, the hydrogen Balmer emission lines are produced by electrons "falling" from higher orbits down to the $n = 2$ orbit; see Fig. 5.6(a).

. (a) Emission lines.

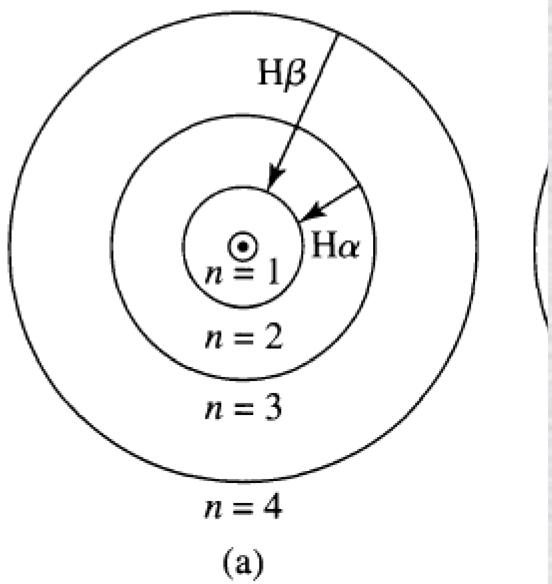
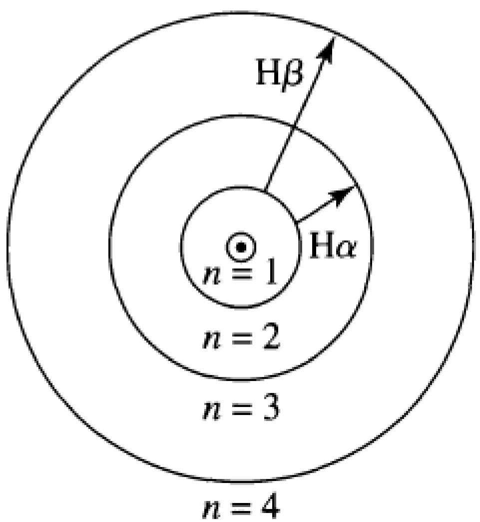


FIGURE 5.6 Balmer lines produced by the Bohr hy
lines.

- A cool, diffuse gas in front of a source of a continuous spectrum produces dark absorption lines in the continuous spectrum. Absorption lines are produced when an electron makes a transition from a lower orbit to a higher orbit. If an incident photon in the continuous spectrum has exactly the right amount of energy, equal to the difference in energy between a higher orbit and the electron's initial orbit, the photon is absorbed by the atom and the electron makes an upward transition to that higher orbit. For example, the hydrogen Balmer absorption lines are produced by atoms absorbing photons that cause electrons to make transitions from the $n = 2$ orbit to higher orbits; see Figs. 5.6(b) and 5.7.



(b)

(b) Absorption

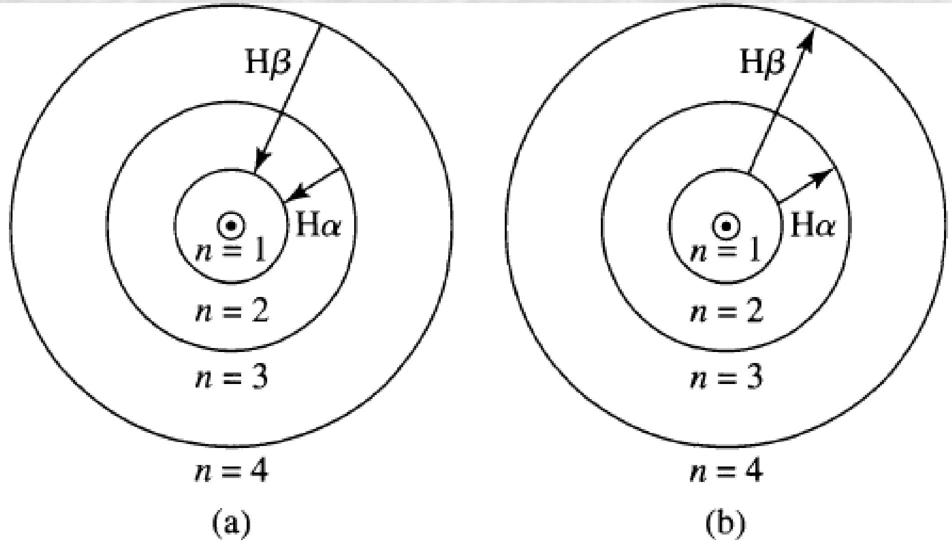


FIGURE 5.6 Balmer lines produced by the Bohr hydrogen atom. (a) Emission lines. (b) Absorption lines.

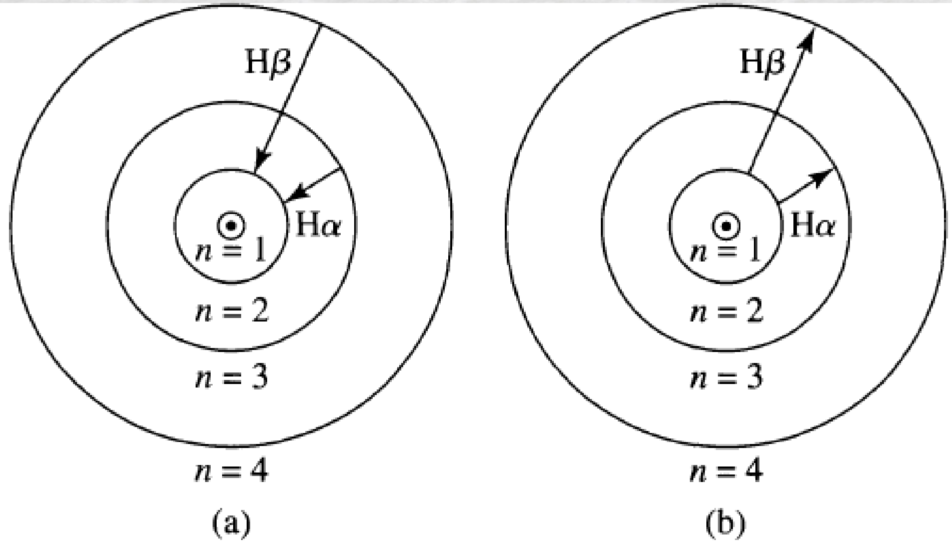


FIGURE 5.6 Balmer lines produced by the Bohr hydrogen atom. (a) Emission lines. (b) Absorption lines.

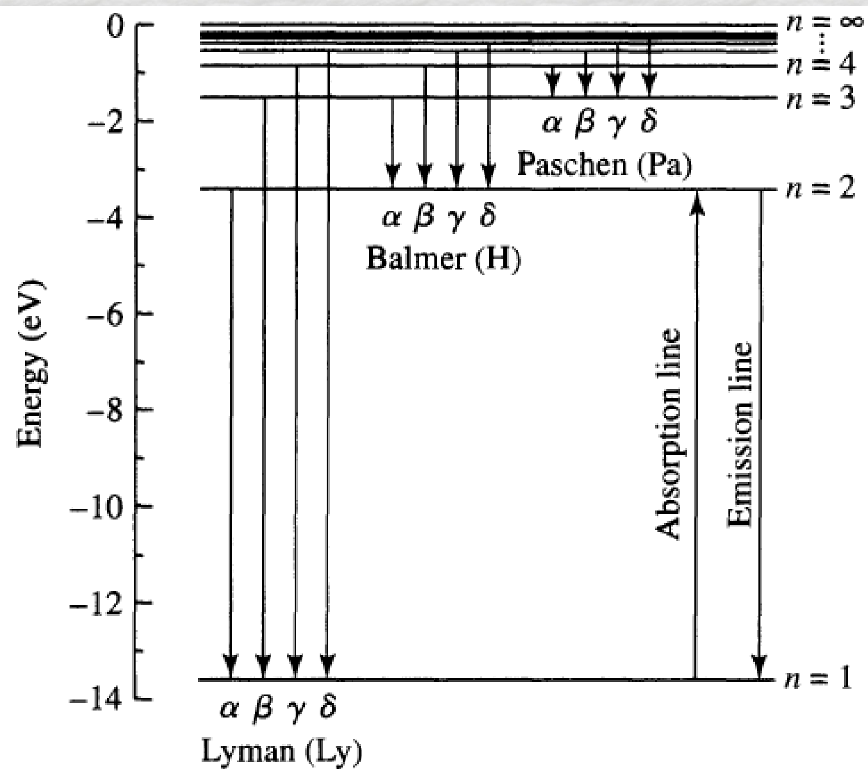
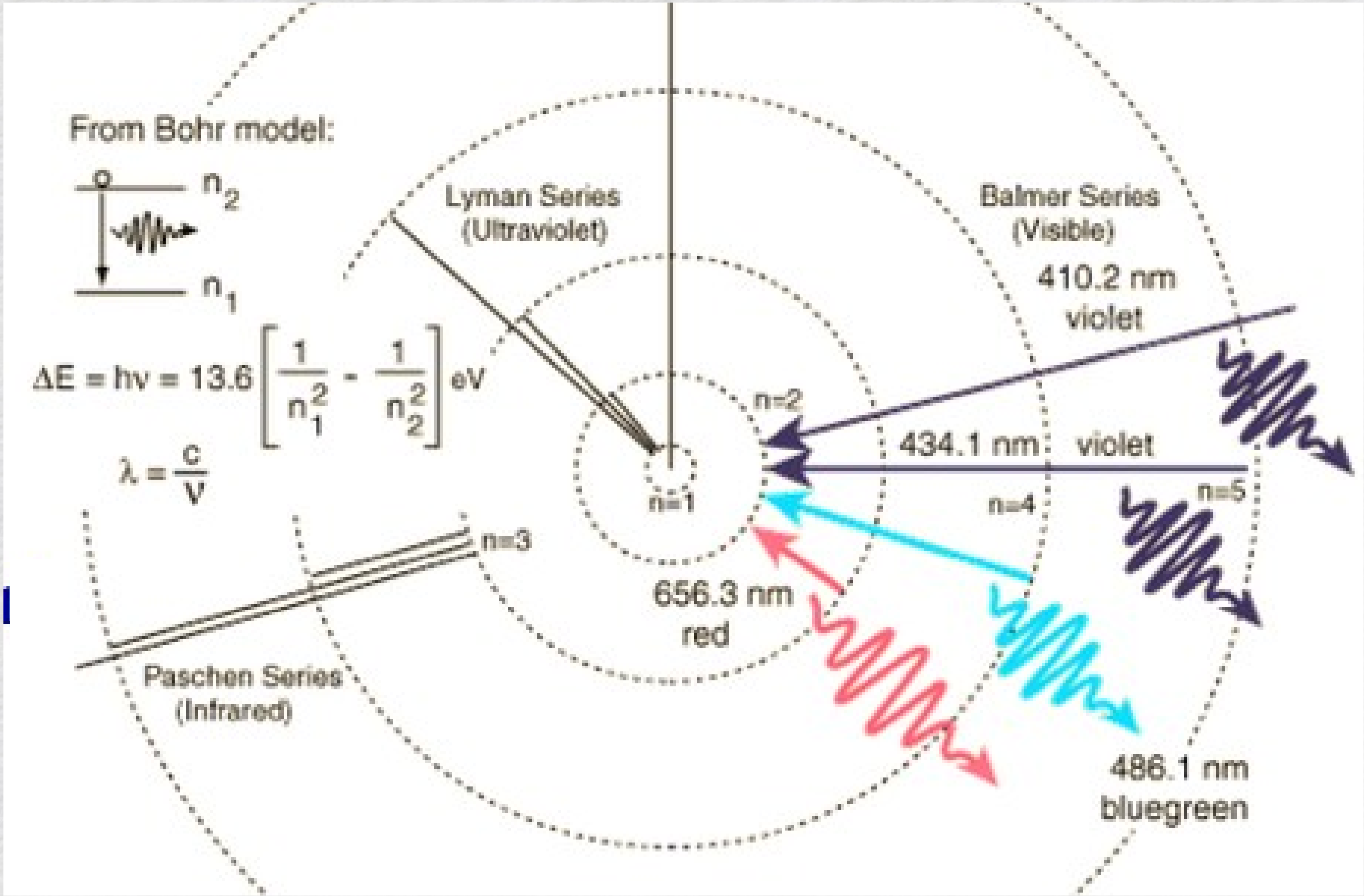


FIGURE 5.7 Energy level diagram for the hydrogen atom showing Lyman, Balmer, and Paschen lines (downward arrows indicate emission lines; upward arrow indicates absorption lines).

Series espectrales del H



Series espectrales del H

Series m	Lyman $(n=1)$	Balmer $(n=2)$	Paschen $(n=3)$	Brackett $(n=4)$	Pfund $(n=5)$	Humphreys $(n=6)$
2	1,215.67					
3	1,025.72	6,562.80				
4	972.537	4,861.32	18,751.0			
5	949.743	4,340.46	12,818.1	40,512.0		
6	937.803	4,101.73	10,938.1	26,252.0	74,578	
7	930.748	3,970.07	10,049.4	21,655.0	46,525	123,680
8	926.226	3,889.05	9,545.98	19,445.6	37,395	75,005
9	923.150	3,835.38	9,229.02	18,174.1	32,961	59,066
10	920.963	3,797.90	9,014.91	17,362.1	30,384	51,273
11	919.352	3,770.63	8,862.79	16,806.5	28,722	46,712
12	918.129	3,750.15	8,750.47	16,407.2	27,575	43,753
13	917.181	3,734.37	8,665.02	16,109.3	26,744	41,697
14	916.429	3,721.94	8,598.39	15,880.5	26,119	40,198
15	915.824	3,711.97	8,545.39	15,700.7	25,636	39,065
16	915.329	3,703.85	8,502.49	15,556.5	25,254	38,184
17	914.919	3,697.15	8,467.26	15,438.9	24,946	37,484
18	914.576	3,691.55	8,437.96	15,341.8	24,693	36,916
19	914.286	3,686.83	8,413.32	15,260.6	24,483	36,449
20	914.039	3,682.81	8,392.40	15,191.8	24,307	36,060
21	913.826	3,679.35				
∞	911.5	3,646.0	8,203.6	14,584	22,788	32,814

Series espectrales del H

Lyman ($n=1$)	Balmer ($n=2$)	Paschen ($n=3$)	Brackett ($n=4$)	Pfund ($n=5$)	Humphreys ($n=6$)
1,215.67					
1,025.72	6,562.80				
972.537	4,861.32	18,751.0			
949.743	4,340.46	12,818.1	40,512.0		
937.803	4,101.73	10,938.1	26,252.0		
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919.352	3,770.63	8,862.79	16,806.5		
918.129	3,750.15	8,750.47	16,407.2		
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