The hydrogen atom

The spectrum of Hydrogen:

If we put a hydrogen atom into some stationary state ψ_{nlm}, it should stay there forever.

If we *tickle* it slightly (by collision with another atom, say, or by shining light on it), the atom may undergo a **transition** to some other stationary state:

- by *absorbing* energy, and moving up to a higher-energy state, or
- by *giving off* energy (typically in the form of electromagnetic radiation), and moving down.

Such perturbations are *always* present.

Transitions (**quantum jumps**) are constantly occurring.

A container of hydrogen gives off light (**photons**), whose energy corresponds to the *difference* in energy between the initial and final states:

$$E_\gamma = E_i - E_f = -13.6\,\mathrm{eV}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

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According to the **Planck formula:** $E_\gamma = h\nu$ and we know: $\lambda = c/\nu$,

$$\boxed{\frac{1}{\lambda} = \mathcal{R}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)}$$

Rydberg formula

where:

$$\mathcal{R} \equiv \frac{m_e}{4\pi c\hbar^3}\left(\frac{e^2}{4\pi\epsilon_0}\right)^2 = 1.097 \times 10^7\,\mathrm{m}^{-1}$$

is known as the **Rydberg constant** (Bohr calculated it!).

The spectrum of Hydrogen:

Lyman series: transitions to the ground state (n_f = 1) lie in the ultraviolet.

Balmer series: transitions to the first excited state (n_f = 2) lie in the optical.

Paschen series: transitions to the first excited state (n_f = 3) lie in the infrared.

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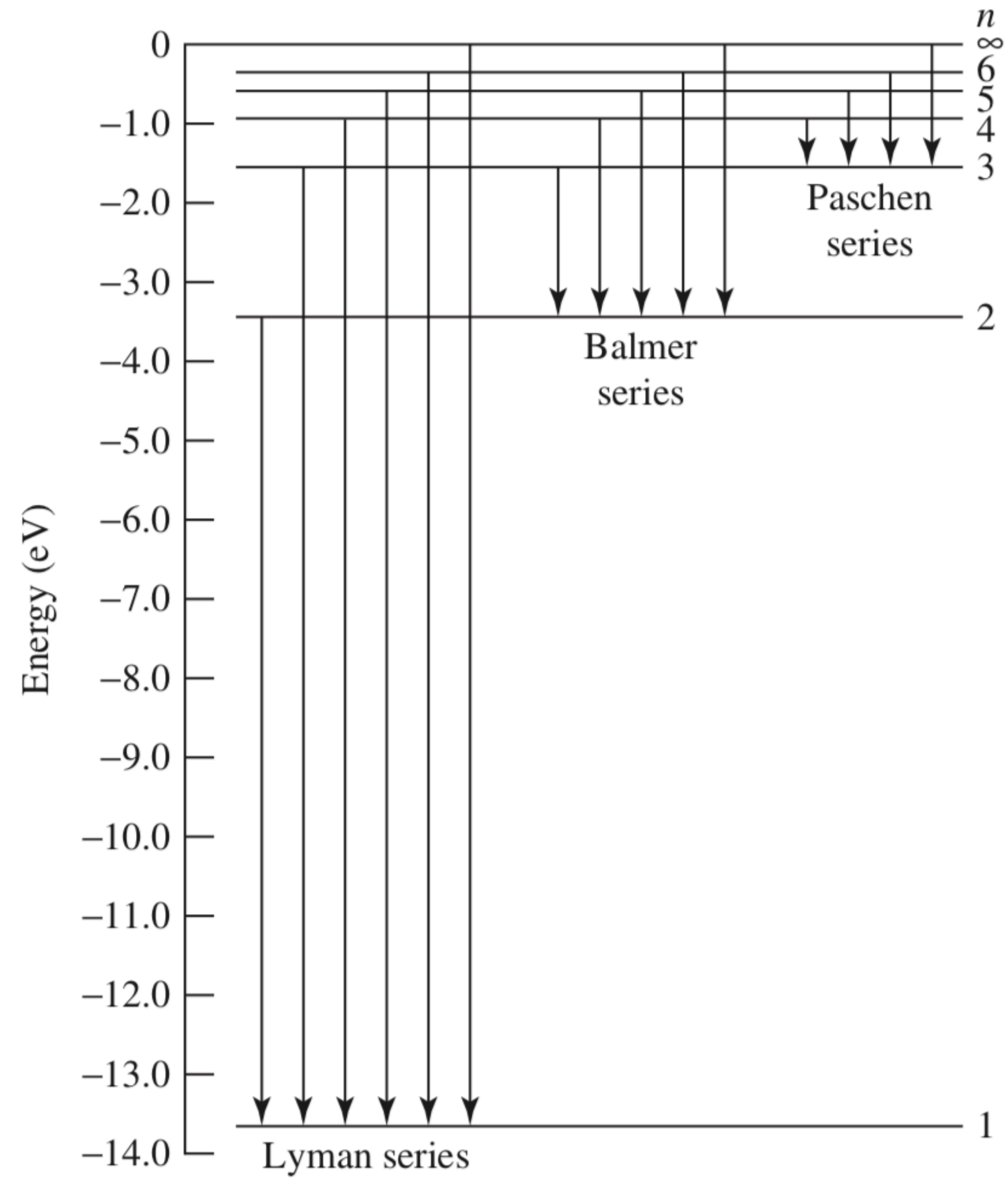
$$\frac{1}{\lambda} = \mathcal{R}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$



At room temperature, most hydrogen atoms are in the ground state.

To obtain the emission spectrum we must populate the excited states.

This is done by passing an electric spark through the gas.