Analysis of Hydrogen Emission Spectrum

https://gemini.google.com/app/22633f49b9890137

Response: The incorrect statement is **(3)**.

Analysis of the Options:

- **Understanding the Spectrum:** The arrow shows wavelength (λ) increasing from left to right. Since energy ($E = \frac{hc}{\lambda}$) is inversely proportional to wavelength, energy increases from right to left ($a > b > c > d$). The question states that line (d) appears in the visible range. Since line (d) is the lowest energy line in this visible set (the Balmer series), line (d) corresponds to the $n = 3 \rightarrow n = 2$ transition (H_α line, red light). The lines c, b, a correspond to $n = 4 \rightarrow n = 2$, $n = 5 \rightarrow n = 2$, and $n = 6 \rightarrow n = 2$, respectively.

- **(1) Correct:** Line (a) has the shortest wavelength, so it has the highest energy among the given lines.

- **(2) Correct:** Line (d) has the longest wavelength (lowest energy) in the Balmer series, corresponding to the lowest energy electronic transition in the visible spectrum ($n = 3 \rightarrow n = 2$).

- **(3) Incorrect:** Passing white light through a gas sample at room temperature produces an **absorption spectrum**, not an emission spectrum. Furthermore, hydrogen exists as diatomic molecules (H_2) at $27^\circ\mathrm{C}$; an atomic hydrogen spectrum requires high temperature or electric discharge to dissociate H_2 into atomic H and excite the electrons.

- **(4) Correct:** Line (c) corresponds to $n = 4 \rightarrow n = 2$ (H_β), which falls in the blue-green / green region of the visible spectrum (~486 nm).

- **(5) Correct:** The Lyman series corresponds to transitions down to $n = 1$ (ultraviolet region), which have higher energy and shorter wavelengths than the Balmer series ($n = 2$). Thus, the Lyman series appears to the left of line (a).
