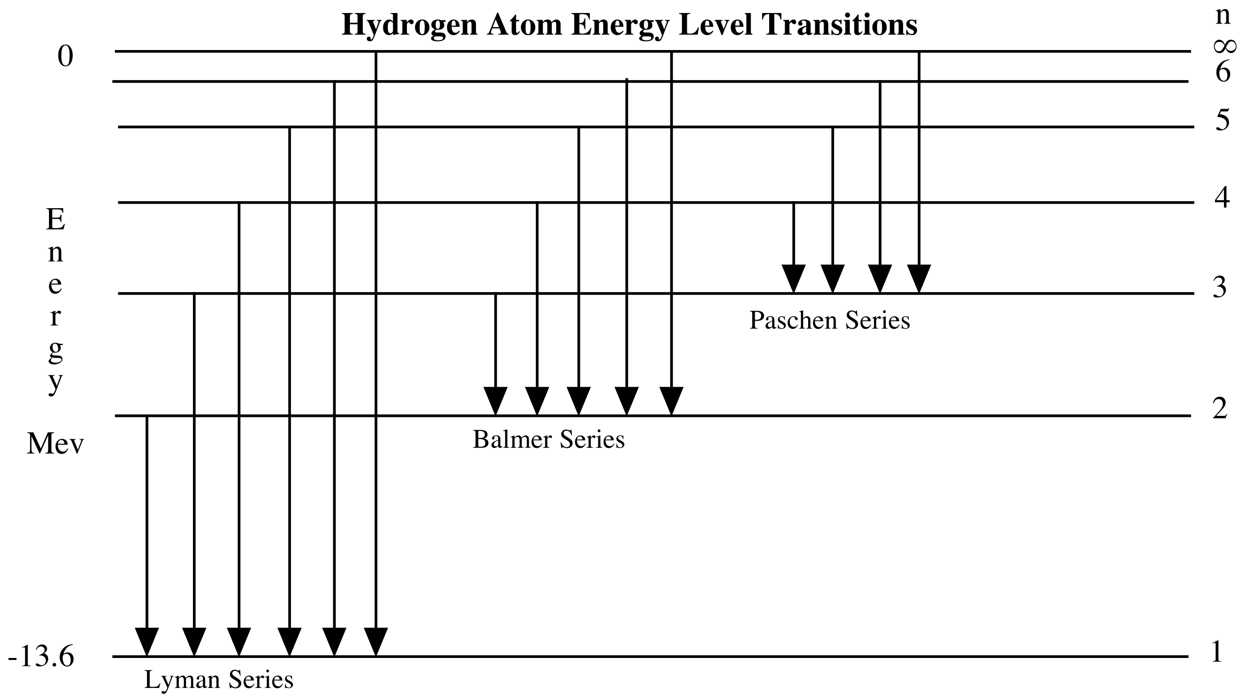
Hydrogen Atom Energy Level Transitions



Energy levels and transitions in the spectrum of atomic hydrogen. The state of lowest energy, called the ground state, is found by putting n = 1 in the equaton $E = \frac{-13.6 \text{ ev}}{n^2}$

The set of light wavelengths emitted by an atom is called the atom's emission spectrum. The emission spectrum of an atom can b seen by looking at the light emitted by an excited atom through a diffraction grating. Bohr's model of the hydrogen atom showed that a definite amount of energy is released when an electron moves from a higher to a lower energy level. The energy released i each transition corresponds to a definite bright line (photon emitted) in the hydrogen spectrum.

Example #1 : For the hydrogen atom, determine the energy difference between the first and second energy levels.

$E_1 = -13.6 \text{ ev} /(1)^2 = -13.6 \text{ ev}$

$E_2 = -13.6 \text{ ev}/(2)^2 = -3.4 \text{ ev}$

$\Delta E = E_2 - E_1 = -3.4 \text{ ev} - (-13.6) = 10.2 \text{ ev}$ This is the energy that must be added to the atom to raise the electron from its ground state to its first energized state.

Example #2 : An electron in an excited hydrogen atom drops from the second energy level to the first energy level.

Using the equations **$E = h f$ and $c = f \lambda$**

E = energy in ev or joules, h = Planck's constant (6.63×10^{-34} J · s), f = frequency (Hz),

c = speed of light (3×10^8 m/s), λ = wavelength (m), 1 ev = 1.6×10^{-19} J

a) Determine the energy of the photon emitted.
b) Calculate the frequency of the photon emitted.
c) Calculate the wavelength of the photon emitted.

a) $E1 = -13.6 \text{ ev} /(1)^2 = -13.6 \text{ ev}$

$E_2 = -13.6 \text{ ev}/(2)^2 = -3.4 \text{ ev}$

$\Delta E = E_2 - E_1 = -3.4 \text{ ev} - (-13.6) = 10.2 \text{ ev}$

b) $E = (10.2 \text{ ev}) (1.6 \times 10^{-19} \text{ J/ev}) = 1.632 \times 10^{-18}$ J

$E = h f$

$f = E/h = 1.632 \times 10^{-18} \text{ J}/ 6.63 \times 10^{-34} \text{ J} \cdot \text{s} = 2.46 \times 10^{15}$/sec or 2.46×10^{15} Hz

c) $c = f \lambda$

$\lambda = c/f = (3 \times 10^8 \text{ m/s})/(2.46 \times 10^{15} \text{ Hz}) = 1.22 \times 10^{-7}$ m.

1. Find the energy in a quantum of light that has a frequency of 1.7×10^{14} Hz.

2. Find the wavelength of a photon of light with energy 2.04 eV. Express your answer in meters.
($h = 6.63 \times 10^{-34}$ J·s; $c = 3 \times 10^{8}$ m/s; 1 eV = 1.6×10^{-19} J)

3. An electron in the n = 1 state in an atom of hydrogen jumps to the n = 3 state. What energy in eV is needed to make the transition?

4. An electron in an excited hydrogen atom falls from the n = 5 state to the n = 4 state. Find the wavelength (in meters) of the light produced.
($h = 6.63 \times 10$ J·s; $c = 3 \times 10^{8}$ m/s; 1 eV = 1.6×10^{-19} J)

KEY

1) 1.1×10^{-19} J

2) 6.09×10^{-7} m

3) 12.1 eV

4) 4.03×10^{-6} m