6.6 - Notice that this system is not typical, in that the energy differences between the levels increases as n increases. Compare with fig 6.12 (p. 214)

a) Given the limits of the diagram, the greatest energy change would involved a transition from n=1 to n=4.

b) The quantum numbers involved in the lowest energy transition are 1 and 2

c) Longer wavelength light represents lower energy changes. As such, putting them in order of increasing wavelength requires putting them in order of decrease distance (energy). Therefore, the order is iii, iv, ii, i

6.33

The existence of line spectra is consistent with Bohr's theory of quantized energies in that each line represents an electron transition between an excited and a ground state. In that not all energy changes are allowed, not all colors are represent in a spectra.

6.34

a) When an excited hydrogen atom emits light of only certain wavelengths, due to the fall of electrons from higher to lower energy positions, we are experiencing strong evidence for the existence of allowed energy levels, with spaces in between which electrons can not occupies

b) b) When an electron moves from a ground state to an excited state it moves to a higher energy position further from the nucleus. As such, the atom expands.

6.37

a) Energy is emitted as an electron transitions from n=4 to n=2

b) Energy is absorbed as an electron moves from a 2.12Å orbit to an 8.46Å orbit.

c) Energy is emitted as an electron adds to a H^+ ion.

6.39a) $E = -2.18 \times 10^{-18} J / n^2$

$E_2 = -2.18 \times 10^{-18} J / 2^2$

$E_2 = -5.45 \times 10^{-19} J$

$E = -2.18 \times 10^{-18} J / n^2$

$E_6 = -2.18 \times 10^{-18} J / 6^2$

$E_6 = -6.06 \times 10^{-20} J$

$\Delta E = E_f - E_i$

$\Delta E = -5.45 \times 10^{-19} J - (-6.06 \times 10^{-20} J)$

$\Delta E = -4.84 \times 10^{-19} J$

$E = h\nu$

$4.84 \times 10^{-19} J = 6.626 \times 10^{-34} J{\cdot}s \cdot \nu$

$x = 7.30 \times 10^{14}$ hz

$C = \lambda\nu$

$3.00 \times 10^{8} ms^{-1} = \lambda \cdot 7.30 \times 10^{14}$ hz

$E = 4.11 \times 10^{-7}$ meter (411 nm)

6.39

b) 411 nm light is in the visible region; it is violet.

6.38a)

$$E = -2.18 \times 10^{-18} J / n^2$$
$$E_1 = -2.18 \times 10^{-18} J / 1^2$$
$$E_1 = -2.18 \times 10^{-18} J$$

$$E = -2.18 \times 10^{-18} J / n^2$$
$$E_\infty = -2.18 \times 10^{-18} J / \infty^2$$
$$E_\infty \approx 0 \text{ J}$$

To move an electron from n=1 to n=∞, completely removing it from the atom, 2.18×10^{-18} J of energy is needed, or

$$\frac{2.18 \times 10^{-18} J}{\text{electron}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} \times \frac{6.02 \times 10^{23} \text{ electron}}{1 \text{ mole electron}} = 1310 \text{ kJ / mol}$$

6.38b) The experimentally determined ionization energy for hydrogen is the same as the value determined through calculation, using energy values for n=1 and n=∞.