Answer to Essential Question 28.6: (a) The photon energy is the difference in energy. Converting the wavelength to energy, we get:

$$E = hc/\lambda = (6.626 \times 10^{-34} \text{ J s})(3.00 \times 10^8 \text{ m/s})/(632.9 \times 10^{-9} \text{ nm}) = 3.141 \times 10^{-19} \text{ J}.$$

Converting this energy to electron volts gives $3.141 \times 10^{-19} \text{ J} \times \dfrac{1 \text{ eV}}{1.602 \times 10^{-19} \text{ J}} = 1.96 \text{ eV}$.

(b) If the $n = 3$ and $n = 1$ levels are separated by 20.66 eV of energy, and the energy difference between the $n = 3$ and $n = 2$ levels is 1.96 eV, the difference between the $n = 2$ and $n = 1$ levels must be 20.66 eV – 1.96 eV = 18.70 eV.

Chapter Summary

Essential Idea: The Atom.
        In our modern view of the atom, the nucleus is modeled as a point particle with a positive charge. The chemical properties of the atom are dominated by the behavior of the atom's electrons. In contrast to the popular representation of an atom being made up of a positive nucleus orbited by electrons in well-defined circular or elliptical orbits, our modern view is that an electron behaves more like a cloud of negative charge, with the density of the charge cloud at a point corresponding to the probability of finding the electron at that point.

Atomic spectra
Exciting a gas causes light to be emitted. The resulting emission spectrum is the fingerprint of the gas – lines are seen at specific wavelengths. A photon produced by an atom comes from an electron that makes a transition from a higher energy level to a lower energy level within the atom. The energy of the photon corresponds to the difference in energy between these electron energy levels.

The four quantum numbers
        One of the early models of the atom, the Bohr model, used a single quantum number to define the quantization of the energy levels, allowed radii, and angular momenta of the electron orbits. Our modern view, the quantum mechanical view of the atom, four quantum numbers are used to completely describe the state of a particular electron. Each quantum number is associated with a physical property of the electron.

        The energy of the electron is quantized by the principal quantum number, n. Allowed values of n are $n = 1, 2, 3, \ldots$ For the hydrogen atom, the energy levels have energies given by:

$$E_n = \frac{-13.6 \text{ eV}}{n^2},$$ (Eq. 28.2: **Energies of the electron levels in the hydrogen atom**)

        The magnitude of the orbital angular momentum of the electron is quantized by the orbital quantum number, ℓ, which can take on values $\ell = 0, 1, \ldots, n-1$. The magnitude of the orbital angular momentum is given by:

$$L = \sqrt{\ell(\ell+1)} \, \frac{h}{2\pi}.$$ (Eq. 28.4: **Orbital angular momentum**)

        The direction of the orbital angular momentum is also quantized. In general, we say that the z-component can take on a limited number of directions, determined by the magnetic quantum

number, m_ℓ, which can take on the values $m_\ell = -\ell, -\ell +1, \ldots, \ell -1, \ell$. The z-component of the orbital angular momentum is given by:

$$L_z = \frac{m_\ell\, h}{2\pi},$$ (Eq. 28.5: **z-component of the orbital angular momentum**)

Finally, the spin angular momentum can take on one of only two values, conventionally referred to as "spin up" and "spin down." The spin angular momentum is characterized by the spin quantum number, which can take on values of +1/2 or –1/2.

Understanding the periodic table of elements
One key to understanding the periodic table is the Pauli exclusion principle – no two electrons in an atom can have the same set of four quantum numbers.

The layout of the periodic table of the elements has to do with the highest-energy filled or partly filled subshell in the ground-state electron configuration of an atom. Elements that have similar ground-state configurations (such as four electrons in a p subshell) are grouped in a column in the periodic table, and generally have similar chemical properties.

Complete ground-state configurations are written with triplets of numbers and letters, in the form $3p^4$. The first number (3, in this case) represents the value of n, the principal quantum number, for the orbital. The letter (p, in this case) stands for the value of the orbital quantum number, ℓ. The second number (4, in this case) is the number of electrons in the subshell in this configuration.

Applications of atomic physics
Two applications of atomic physics include some lasers (such as the helium-neon laser) and the properties of fluorescence and phosphorescence. These applications involve the emission of photons from electrons in atoms dropping down from one energy level to another.

End-of-Chapter Exercises

Exercises 1 – 12 are conceptual questions that are designed to see if you have understood the main concepts of the chapter.

1. In a particular line spectrum, photons are observed that have an energy of 2.5 eV. Choose the phrase below that best completes the following sentence. The element producing the light definitely has …
    (i)… an electron energy level at an energy of 2.5 eV.
    (ii)… an electron energy level at an energy of –2.5 eV.
    (iii)… two electron energy levels that are 2.5 eV apart in energy.

2. Imagine that there is an atom with electron energy levels at the following energies: –88 eV, –78 eV, and –60 eV. Confining ourselves to electron transitions between these levels only, would we expect to see photons emitted from this atom with the following energies? Explain why or why not. (a) 60 eV, (b) –60 eV, (c) 18 eV, (d) 138 eV.

3. To be visible to the human eye, photons must have an energy between about 1.8 eV and 3.1 eV. In a particular atom, one of the electron energy levels is at an energy of –60.0 eV. Can electron transitions associated with this energy level produce photons that are visible to the human eye? Explain why or why not.