Unit 7. Atomic Structure

Upon successful completion of this unit, the students should be able to:

7.1 List the eight regions of the electromagnetic spectrum in the designated order and perform calculations involving frequency, wavelength, and energy (note: students should memorize \( E = h\nu \) and \( c = \nu\lambda \) and be able to apply Planck’s constant and the speed of light if given)

1. List the eight forms of electromagnetic radiation in order of increasing wavelength.

2. What is the wavelength (in meters) of photons that have an energy of \( 1.83 \times 10^3 \) kJ/mol?

3. A photon of visible light has a frequency of \( 5.45 \times 10^{14} \) s\(^{-1}\). What is the energy (in joules) of this photon, and what is the wavelength (in nanometers) of this photon?

4. Mars is roughly 60 million km from earth. Show how you would calculate how long it takes a radio signal from earth to reach Mars. Include units on all numbers.

5. Calculate the frequency of blue light of wavelength \( 4.5 \times 10^2 \) nm.

6. a. Calculate the energy of a photon that is emitted at wavelength of \( 5.69 \times 10^3 \) nm.
   b. Calculate the energy of a mole of photons in part “a”.

7.2 Define photon and relate the photon to the dual nature of electromagnetic radiation.

1. Define photon.

2. Einstein was the first to study the photoelectric effect. That is, when light is shone on a metal plate, electrons are only emitted from the plate when the light has a certain minimum energy. If the light is below the minimum energy required, electrons will not be emitted from the metal plate, even if the intensity of the light is increased. The photoelectric effect can be used to complete an electric circuit and is used in such things as automatic door openers and burglar alarms. See diagram below.

Explain how the photoelectric effect provides evidence that light is more than just waves but must also be quantized into photons.
7.3 Describe the relation between electronic transitions and line spectra.

1. What causes the emission spectrum of an element (that is, why do the elements emit light when they are heated or a voltage is applied to them)? Why isn’t the emission spectrum continuous?

2. The visible lines in the hydrogen emission spectrum correspond to the following four electronic transitions: \( n=3 \rightarrow n=2 \), \( n=4 \rightarrow n=2 \), \( n=5 \rightarrow n=2 \), \( n=6 \rightarrow n=2 \). The color of the lines are violet, red, blue, and blue-violet. Which electronic transition corresponds to the red line in the hydrogen emission spectrum? How do you know?

7.3a Solve quantitative problems related to electronic transitions in the hydrogen atom.

1. Calculate the energy change corresponding to the excitation of an electron from the \( n = 1 \) to \( n = 3 \) electronic state in the hydrogen atom.

7.4 Explain in a qualitative way the Bohr model of the hydrogen atom.

1. When an electron falls to lower energy levels, it (absorbs, emits) ____________ energy.

2. For a hydrogen atom, which electronic transition would result in the emission of a photon with the highest energy?
   a. \( n = 2 \rightarrow n = 3 \)
   b. \( n = 3 \rightarrow n = 6 \)
   c. \( n = 5 \rightarrow n = 3 \)
   d. \( n = 6 \rightarrow n = 4 \)

3. Describe two main conclusions arising from Bohr’s experiments with the hydrogen atom, one of which is still widely accepted and another which is not.

7.5 List the four quantum numbers and relate them to electronic structure by using them to predict relevant information about shells, subshells, orbitals, and electrons.

1. a. How many subshells exist in the sixth shell \( (n = 6) \)?
   b. What is the maximum number of orbitals in the fifth shell \( (n = 5) \)?
   c. What is the number of orbitals in a sublevel with \( l = 3 \)?
   d. Give the complete set of three quantum numbers for a \( 4p_x \) orbital.

2. The number of orbitals in a given subshell, such as a \( 5d \) subshell, is determined by the number of possible values of:
   a. \( n \)
   b. \( l \)
   c. \( m_l \)
   d. \( m_s \)

3. What are all the possible values for \( l \) if \( n = 5 \)?

4. How many \( g \) orbitals are allowed in a given shell?
5. Which of the following is NOT a valid set of quantum numbers?
   a. \( n = 2, l = 1, m_\ell = 0, m_s = -\frac{1}{2} \)
   b. \( n = 2, l = 1, m_\ell = -1, m_s = -\frac{1}{2} \)
   c. \( n = 3, l = 0, m_\ell = 0, m_s = \frac{1}{2} \)
   d. \( n = 3, l = 2, m_\ell = 3, m_s = \frac{1}{2} \)

7.6 Differentiate between a Bohr orbit and a quantum mechanical orbital.

1. The Niels Bohr model of the atom differs from the current quantum mechanical model in that the current model:
   a. assumes that the electron travels in circular orbits
   b. considers all electron shells to be of equal energy
   c. has considerably higher ground state energies
   d. makes use of statistical probabilities

2. The figures below represent two different ways of showing the relative probability of finding the hydrogen electron in its ground state at various distances from the nucleus. In (a), the depth of the color is proportional to the probability of finding an electron at that point. In (b), the probability of finding the electron at a given distance along the x-axis is plotted.

Explain in your own words the physical meaning of the two graphs.

7.7 Sketch any s, p, or d orbital.

1. On the axes provided, sketch the indicated orbitals.
   a. \( 3p_x \)
   b. \( 6d_{xy} \)
7.8 Write electronic configurations (i.e. \( \text{Mg} = 1s^22s^22p^63s^2 \)) for all elements which strictly follow the Aufbau principle, Pauli exclusion principle, and Hund's rule and solve related problems.

1. Indicate if the following electron configuration is ground state, excited state, or impossible:
   \( 1s^22s^22p^62d^{10}3s^23p^6 \)

2. Write the complete electron configurations (don’t use noble gas abbreviations) for the following:
   a. Si
   b. Sr
   c. Ti
   d. Al
   e. Pb

3. Indicate if the following electron configuration is ground state, excited state, or impossible:
   \( 1s^22s^22p^63s^2 \)

4. a. Write the complete ground state electron configuration of element 74, tungsten (W) (don’t use the noble gas abbreviated form).
   b. How many unpaired electrons does tungsten have?

7.9 Write electron configurations for the anomalies Cr, Mo, Cu, Ag, Au and write electron configurations for all other anomalies in which adequate hints are provided and solve related problems.

1. Write the electron configuration of Ag.

7.10 Use the Periodic Table to write noble gas abbreviated electron configurations (i.e. \( \text{Mg} = [\text{Ne}]3s^2 \)).

1. Write the noble gas abbreviated electron configuration of the following:
   a. Mn
   b. Se
   c. Ag
   d. Cs

2. Describe the following electron configurations as ground state (G), excited state (E), or impossible (I):
   \[
   [\text{Ne}]3s^23p^7 \quad [\text{Ar}]4s^23d^6 \\
   [\text{He}]2s^12p^4 \quad [\text{Ar}]4s^23d^{10}4p^63f^2
   \]

3. Write the noble gas abbreviated electron configuration of uranium (hint: uranium has one 6d electron in the ground state).
7.11 Draw orbital diagrams based on electron configurations (i.e. \( \text{Mg} = 1s^2 2s^1 2p^3 3s^1 \)) and use them to predict paramagnetism and diamagnetism.

1. Draw the orbital diagram for the \( \text{Cr}^+ \) ion. Then tell if the \( \text{Cr}^+ \) ion is paramagnetic or diamagnetic?

7.12 Define paramagnetic and diamagnetic.

1. Define paramagnetic.

7.13 Explain the concept of effective nuclear charge and the trend which exists across a period.

1. a. Complete the following statement. The effective nuclear charge felt by outermost electrons generally (increases, decreases, remains the same) _________________ from left to right across a period of the Periodic Table (for representative elements).

   b. Explain in detail your answer to part “a”. Make sure to discuss ALL the relevant factors.

7.14 Define ionization energy and electron affinity and be able to write equations related to each.

1. Define electron affinity.

2. Write the equation representing the third ionization energy of barium.

7.15 Use the Periodic Table to predict trends in ionization energies, electron affinities, and atomic sizes and to provide adequate explanations for these trends.

1. Consider the following elements: \( \text{Cs, Na, P} \).

   a. Arrange in order of increasing atomic radius.
   b. Arrange in order of increasing first ionization energy.

2. Which has a more negative electron affinity: \( \text{Cl} \) or \( \text{P} \)?

3. Which element of the following has the largest fifth ionization energy: \( \text{Mg, Al, Si, or P} \)?

4. Which has a higher first ionization energy: \( \text{Na} \) or \( \text{S} \)? ______

5. Which is larger: \( \text{Ca} \) or \( \text{Sr} \)? ______

6. Consider the following elements: \( \text{Sr, Sn, Te} \).

   a. Arrange in order of increasing atomic radius.
   b. Explain your answer to part “a” using the concept of effective nuclear charge.
7. Consider the following elements: Ca, Ba, Sr.
   a. Arrange in order of increasing first ionization energy:
   b. Which of the following would offer an explanation for your choice in part “a”?
      A. From top to bottom down a group, the outermost electron becomes one in a higher principal energy level therefore making removing an electron more difficult.
      B. From top to bottom down a group, the outermost electron becomes one in a higher principal energy level therefore making removing an electron easier.
      C. From top to bottom down a group, the outermost electron becomes one in a lower principal energy level therefore making removing an electron more difficult.
      D. From top to bottom down a group, the outermost electron becomes one in a lower principal energy level therefore making removing an electron easier.

8. a. Which has a more negative electron affinity: N or N³⁻?
   b. Which has a more negative electron affinity: Sb or I?

9. Which has a higher electronegativity: S or Si?

10. Which element has the largest sixth (6th) ionization energy?

11. In the blanks to the right of the question, arrange the following atoms or ions from SMALLEST to LARGEST.
    A) Al, Ba and Cl          Order: __________ < _____________ < __________
    B) Kr, Se and Se²⁻        Order: __________ < _____________ < __________
    C) F⁻, Na⁺ and Ne        Order: __________ < _____________ < __________

12. A certain element, has a third ionization energy very much higher than the first two ionization energies. This element has no d electrons and is larger than sodium. What is the element? Explain your reasoning to receive full credit.

Additional Unit 7 Sample Questions:

1. Determine which of the following statements is FALSE. Then rewrite this false statement so that it would be a TRUE statement.
   a. The f subshell can hold a maximum of 14 electrons.
   b. A s orbital has a spherical shape.
   c. The d subshell has 10 orbitals in it.
   d. An electron in a 3d orbital is higher in energy than an electron in a 3p orbital.
   e. The 1p subshell does not exist, but the 2p subshell does exist.

2. a. List four atoms or ions that are isoelectronic with Sc⁺³.
   b. Put your answers to part a in increasing (smallest to largest) order.

3. Describe the relationship between the first ionization energy of sodium and the electron affinity of Na⁺.