**CHAPTER 4 REVIEW**

**Arrangement of Electrons in Atoms**

**Teacher Notes and Answers**

**Chapter 4**

**SECTION 1**

**SHORT ANSWER**

1. In order for an electron to be ejected from a metal surface, the electron must be struck by a single photon with at least the minimum energy needed to knock the electron loose.

2. The ground state is the lowest energy state of the atom. When the atom absorbs energy, it can move to a higher energy state, or excited state.

3. A photon is emitted when an atom moves from an excited state to its ground state or to a lower-energy excited state.

4. When an atom loses energy, it falls from a higher energy state to a lower energy state. The frequency of the emitted light, observed in an element’s line-emission spectrum, may be measured. The energy of each transition is calculated using the equation $E = hv$, where $v$ is the frequency of each of the lines in the element’s line-emission spectrum. From the analysis of these results, the energy levels of an atom of each element may be determined.

5. Energy is proportional to frequency, and ultraviolet radiation has a higher frequency than infrared radiation. To produce ultraviolet radiation, electrons must drop to lower energy levels than they do to produce infrared radiation.

6. Wave B has the higher frequency. Wavelength is inversely proportional to frequency, so as the wavelength decreases, its frequency increases.

7. Six different photons were emitted. Each time an excited helium atom falls back from an excited state to its ground state or to a lower energy state, it emits a photon of radiation that shows up as this specific line-emission spectrum. There are six lines in this helium spectrum.

8. $9.7 \times 10^{14}$ Hz

9. $9.4 \times 10^9$ m

**SECTION 2**

**SHORT ANSWER**

1. d

2. a

3. a

4. c

5. c

6. c

7. c

8. Scientists knew that any wave confined to a space could have only certain frequencies. De Broglie suggested that electrons should be considered as waves confined to the space around an atomic nucleus; in this way, electron waves could exist only at specific frequencies. According to the relationship $E = hv$, these frequencies correspond to the specific quantized energies of the Bohr orbitals.

9. The principal quantum number refers to the main energy level. The angular momentum quantum number refers to the shape of the orbital. The magnetic quantum number refers to the orientation of an orbital around the nucleus. The spin quantum number indicates the spin state of an electron in an orbital.

10. The Heisenberg uncertainty principle states that it is impossible to determine simultaneously both the position and velocity of an electron (or any other particle). Because measuring the position of an electron actually changes its position, there is always a basic uncertainty in trying to locate an electron. Thus, the exact position of the electron cannot be found. An electron cloud or orbital represents the region that is the probable location of an electron.

11.
SECTION 3

SHORT ANSWER

1. The Pauli exclusion principle states that no two electrons in an atom may have the same set of four quantum numbers. If both electrons in the same orbital had the same spin state, each electron would have the same four quantum numbers. If one electron has the opposite spin state, the fourth quantum number is different and the exclusion principle is obeyed.

2. This orbital notation is possible if the helium atom is in an excited state.

3. $1s^2 2s^2 2p^6 3s^2 3p^3$; \[ \begin{array}{c|c|c|c|c|c|c} \text{1s} & \text{2s} & \text{2p}_x & \text{2p}_y & \text{2p}_z & \text{3s} & \text{3p}_x \text{3p}_y \text{3p}_z \\ \hline \uparrow & \uparrow & \downarrow & \downarrow & \uparrow & \uparrow & \downarrow \\ \end{array} \]

4. $1s^2 2s^2 2p^3$; \[ \begin{array}{c|c|c|c|c} \text{1s} & \text{2s} & \text{2p}_x & \text{2p}_y & \text{2p}_z \\ \hline \uparrow & \uparrow & \uparrow & \uparrow & \uparrow \\ \end{array} \]

5. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$; \[ \begin{array}{c|c|c|c|c|c|c} \text{1s} & \text{2s} & \text{2p}_x & \text{2p}_y & \text{3s} & \text{3p}_x \text{3p}_y \text{3p}_z & \text{4s} \\ \hline \uparrow & \uparrow & \downarrow & \downarrow & \downarrow & \downarrow & \uparrow \\ \end{array} \]

6. $1s^2 2s^2 2p^6 3s^2 3p^1$; \[ \begin{array}{c|c|c|c|c|c|c} \text{1s} & \text{2s} & \text{2p}_x & \text{2p}_y & \text{3s} & \text{3p}_x \text{3p}_y \text{3p}_z \\ \hline \uparrow & \uparrow & \downarrow & \downarrow & \uparrow & \uparrow & \downarrow \\ \end{array} \]

7. $1s^2 2s^2 2p^6 3s^2 3p^6$; \[ \begin{array}{c|c|c|c|c|c|c} \text{1s} & \text{2s} & \text{2p}_x & \text{2p}_y & \text{3s} & \text{3p}_x \text{3p}_y \text{3p}_z \\ \hline \uparrow & \uparrow & \downarrow & \downarrow & \uparrow & \uparrow & \downarrow \\ \end{array} \]

8. $1s^2 2s^2 2p^1$; \[ \begin{array}{c|c|c|c|c|c|c} \text{1s} & \text{2s} & \text{2p}_x & \text{2p}_y & \text{2p}_z \\ \hline \uparrow & \uparrow & \uparrow & \uparrow & \uparrow \\ \end{array} \]

9. a. Pauli exclusion principle
b. Hund’s rule
CHAPTER 4 REVIEW

Arrangement of Electrons in Atoms

SECTION 1

SHORT ANSWER Answer the following questions in the space provided.

1. In what way does the photoelectric effect support the particle theory of light?

2. What is the difference between the ground state and the excited state of an atom?

3. Under what circumstances can an atom emit a photon?

4. How can the energy levels of the atom be determined by measuring the light emitted from an atom?

5. Why does electromagnetic radiation in the ultraviolet region represent a larger energy transition than does radiation in the infrared region?
SECTION 1 continued

6. Which of the waves shown below has the higher frequency? (The scale is the same for each drawing.) Explain your answer.

[Diagram of two waves with labels: Wave A and Wave B]

7. How many different photons of radiation were emitted from excited helium atoms to form the spectrum shown below? Explain your answer.

[Spectrum for helium]

PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.

8. ________________ What is the frequency of light that has a wavelength of 310 nm?

9. ________________ What is the wavelength of electromagnetic radiation if its frequency is $3.2 \times 10^{-2}$ Hz?
CHAPTER 4 REVIEW

Arrangement of Electrons in Atoms

SECTION 2

SHORT ANSWER Answer the following questions in the space provided.

1. _____ How many quantum numbers are used to describe the properties of electrons in atomic orbitals?
   (a) 1  (c) 3
   (b) 2  (d) 4

2. _____ A spherical electron cloud surrounding an atomic nucleus would best represent
   (a) an s orbital.  (c) a combination of two different p orbitals.
   (b) a p orbital.  (d) a combination of an s and a p orbital.

3. _____ How many electrons can an energy level of \( n = 4 \) hold?
   (a) 32  (c) 8
   (b) 24  (d) 6

4. _____ How many electrons can an energy level of \( n = 2 \) hold?
   (a) 32  (c) 8
   (b) 24  (d) 6

5. _____ Compared with an electron for which \( n = 2 \), an electron for which \( n = 4 \)
   has more
   (a) spin.  (c) energy.
   (b) particle nature.  (d) wave nature.

6. _____ According to Bohr, which is the point in the figure below where electrons cannot reside?
   (a) point A  (c) point C
   (b) point B  (d) point D

7. _____ According to the quantum theory, point D in the above figure represents
   (a) the fixed position of an electron.
   (b) the farthest position from the nucleus that an electron can achieve.
   (c) a position where an electron probably exists.
   (d) a position where an electron cannot exist.
8. How did de Broglie conclude that electrons have a wave nature?

9. Identify each of the four quantum numbers and the properties to which they refer.

10. How did the Heisenberg uncertainty principle contribute to the idea that electrons occupy “clouds,” or “orbitals”?

11. Complete the following table:

<table>
<thead>
<tr>
<th>Principal quantum number, ( n )</th>
<th>Number of sublevels</th>
<th>Types of orbitals</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 1</td>
<td>( s )</td>
<td></td>
</tr>
<tr>
<td>2 2</td>
<td>( s, p )</td>
<td></td>
</tr>
<tr>
<td>3 3</td>
<td>( s, p, d )</td>
<td></td>
</tr>
<tr>
<td>4 4</td>
<td>( s, p, d, f )</td>
<td></td>
</tr>
</tbody>
</table>
CHAPTER 4 REVIEW

Arrangement of Electrons in Atoms

SECTION 3

SHORT ANSWER  Answer the following questions in the space provided.

1. State the Pauli exclusion principle, and use it to explain why electrons in the same orbital must have opposite spin states.

   _________________________________________________________________
   _________________________________________________________________
   _________________________________________________________________
   _________________________________________________________________
   _________________________________________________________________

2. Explain the conditions under which the following orbital notation for helium is possible:

   \[
   \begin{array}{c}
   \uparrow \\
   1s \\
   \uparrow \\
   2s \\
   \end{array}
   \]

   Write the ground-state electron configuration and orbital notation for each of the following atoms:

3. Phosphorus

4. Nitrogen

5. Potassium
6. Aluminum

7. Argon

8. Boron

9. Which guideline, Hund’s rule or the Pauli exclusion principle, is violated in the following orbital diagrams?

   a. ____________________________
   b. ____________________________