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CHAPTER 5 STUDY GUIDE

Electrons in Atoms

Section 5.1 Light and Quantized Energy

In your textbook, read about the wave nature of light.

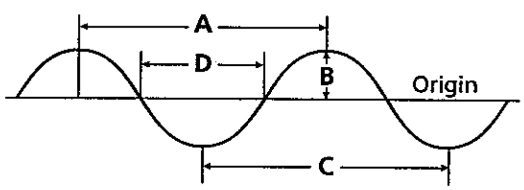
Use each of the terms below just once to complete the passage.

amplitude	energy	frequency	hertz
light	wave	wavelength	speed

Electromagnetic radiation is a kind of **(1)** Energy that behaves like a(n) **(2)** wave as it travels through space. **(3)** Light is one type of electromagnetic radiation. Other examples include X rays, radio waves, and microwaves.

All waves can be characterized by their wavelength, amplitude, frequency, and **(4)** Speed. The shortest distance between equivalent points on a continuous wave is called a(n) **(5)** wavelength. The height of a wave from the origin to a crest or from the origin to a trough is the **(6)** Amplitude. **(7)** Frequency is the number of waves that pass a given point in one second. The SI unit for frequency is the **(8)** Hertz, which is equivalent to one wave per second.

Use the figure to answer the following questions.



9. Which letter(s) represent one wavelength?

Both A + C

10. Which letter(s) represent the amplitude? B

11. If twice the length of A passes a stationary point every second, what is the frequency of the wave?

The frequency is 2 waves or 2 HZ

Section 5.1 *continued*

In your textbook, read about the particle nature of light.

Circle the letter of the choice that best completes the statement or answers the question.

12. A(n) c is the minimum amount of energy that can be lost or gained by an atom.

a. valence electron **b.** electron **c.** quantum **d.** Planck's constant

13. According to Planck's theory, for a given frequency, ν, matter can emit or absorb energy only in

a. units of hertz.
b. whole-number multiples of $h\nu$.
c. entire wavelengths.
d. multiples of $\frac{1}{2}h\nu$, $\frac{1}{4}h\nu$, and so on.

14. The ______ is the phenomenon in which electrons are emitted from a metal's surface when light of a certain frequency shines on it.

a. quantum **b.** Planck concept **c.** photon effect **d.** photoelectric effect

15. Which equation would you use to calculate the energy of a photon?

a. $E_{photon} = h\nu \times$ Planck's constant
b. $E_{photon} = h\nu$
c. $E_{photon} = \frac{1}{2}h\nu$
d. $c = \lambda\nu$

In your textbook, read about atomic emission spectra.

For each statement below, write ***true*** **or** ***false*****.**

False **16.** Like the visible spectrum, an atomic emission spectrum is a continuous range of colors.

True **17.** Each element has a unique atomic emission spectrum.

True **18.** A flame test can be used to identify the presence of certain elements in a compound.

True **19.** The fact that only certain colors appear in an element's atomic emission spectrum indicates that only certain frequencies of light are emitted.

False **20.** Atomic emission spectra can be explained by the wave model of light.

False **21.** The neon atoms in a neon sign emit their characteristic color of light as they absorb energy.

True **22.** When an atom emits light, photons having certain specific energies are being emitted.

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Section 5.2 Quantum Theory and the Atom

In your textbook, read about the Bohr model of the atom.

Use each of the terms below to complete the statements.

atomic emission spectrum	electron	frequencies	ground state
higher	energy levels	lower	

1. The lowest allowable energy state of an atom is called its Ground State.
2. Bohr's model of the atom predicted the Frequencies of the lines in hydrogen's atomic emission spectrum.
3. According to Bohr's atomic model, the smaller an electron's orbit, the lower the atom's energy level.
4. According to Bohr's atomic model, the larger an electron's orbit, the higher the atom's energy level.
5. Bohr proposed that when energy is added to a hydrogen atom, its electron moves to a higher-energy orbit.
6. According to Bohr's atomic model, the hydrogen atom emits a photon corresponding to the difference between the energy Levels associated with the two orbits it transitions between.
7. Bohr's atomic model failed to explain the Atomic emission spectrum of elements other than hydrogen.

In your textbook, read about the quantum mechanical model of the atom.

Answer the following questions.

8. If you looked closely, could you see the wavelength of a fast-moving car? Explain your answer.
   No: the wavelength is far too small to be seen or even measured with the most sensitive scientific Instrument
9. Using de Broglie's equation, $\lambda = \frac{h}{mv}$ which would have the larger wavelength, a slow-moving proton or a fast-moving golf ball? Explain your answer.
   The proton would have the larger wavelength because wavelength increases with decreasing mass + Velocity

Section 5.2 *continued*

In your textbook, read about the Heisenberg uncertainty principle.

For each item in Column A, write the letter of the matching item in Column B.

Column A	Column B
C **10.** The modern model of the atom that treats electrons as waves	**a.** Heisenberg uncertainty principle
A **11.** States that it is impossible to know both the velocity and the position of a particle at the same time	**b.** Schrödinger wave equation
D **12.** A three-dimensional region around the nucleus representing the probability of finding an electron	**c.** quantum mechanical model of the atom
B **13.** Originally applied to the hydrogen atom, it led to the quantum mechanical model of the atom	**d.** atomic orbital

Answer the following question.

14. How do the Bohr model and the quantum mechanical model of the atom differ in how they describe electrons?

The quantum mechanical model treats electrons as waves & doesn't describe the electron's path around the nucleus. The Bohr model treats electrons as particles traveling in specific circular orbits.

In your textbook, read about hydrogen's atomic orbitals.

In the space at the left, write the term in parentheses that correctly completes the statement.

do not **15.** Atomic orbitals (do, do not) have an exactly defined size.

two **16.** Each orbital may contain at most (two, four) electrons.

spherically shaped **17.** All s orbitals are (spherically shaped, dumbbell shaped).

n **18.** A principal energy has (n, n^2) energy sublevels.

electrons **19.** The maximum number of (electrons, orbitals) related to each principal energy level equals $2n^2$.

three **20.** There are (three, five) equal energy p orbitals.

2s & 2p **21.** Hydrogen's principal energy level 2 consists of (2s and 3s, 2s and 2p) orbitals.

nine **22.** Hydrogen's principal energy level 3 consists of (nine, three) orbitals.

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Section 5.3 Electron Configuration

In your textbook, read about ground-state electron configurations.

Use each of the terms below just once to complete the passage.

Aufbau principle	electron configuration	ground-state electron configuration	Hund's rule
lowest	Pauli exclusion principle	spins	stable

The arrangement of electrons in an atom is called the atom's **(1)** electron Configuration. Electrons in an atom tend to assume the arrangement that gives the atom the **(2)** lowest possible energy. This arrangement of electrons is the most **(3)** Stable arrangement and is called the atom's **(4)** ground-state electron configuration

Three rules define how electrons can be arranged in an atom's orbitals. The **(5)** Aufbau principle states that each electron occupies the lowest energy orbital available. The **(6)** Pauli exclusion principle states that a maximum of two electrons may occupy a single atomic orbital, but only if the electrons have opposite **(7)** Spins. **(8)** Hund's Rule states that single electrons with the same spin must occupy each equal-energy orbital before additional electrons with opposite spins occupy the same orbitals.

Complete the following table.

Element	Atomic Number	Orbitals					Electron Configuration
		1s	2s	$2p_x$	$2p_y$	$2p_z$	
9. Helium	2		↑↓				$1s^2$
10. Nitrogen	7	↑↓	↑↓	↑	↑	↑	$1s^2 2s^2 2p^3$
11. Neon	10	↑↓	↑↓	↑↓	↑↓	↑↓	$1s^2 2s^2 2p^6$

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Section 5.3 *continued*

Answer the following questions.

12. What is germanium's atomic number? How many electrons does germanium have?

Atomic # 32 32 electrons

13. What is noble-gas notation, and why is it used to write electron configurations?

Noble-gas notation uses the bracketed symbol of the nearest preceding noble-gas atom in the periodic Table in the elec configurations.
It is the Short hand Form!

14. Write the ground-state electron configuration of a germanium atom, using noble-gas notation.

$[Ar]4s^2 3d^{10} 4p^2$

In your textbook, read about valence electrons.

Circle the letter of the choice that best completes the statement or answers the question.

15. The electrons in an atom's outermost orbitals are called

a. electron dots. **b.** quantum electrons. **c.** valence electrons. **d.** noble-gas electrons.

16. In an electron-dot structure, the element's symbol represents the

a. nucleus of the noble gas closest to the atom in the periodic table.

b. atom's nucleus and inner-level electrons.

c. atom's valence electrons.

d. electrons of the noble gas closest to the atom in the periodic table.

17. How many valence electrons does a chlorine atom have if its electron configuration is $[Ne]3s^2 3p^5$?

a. 3 **b.** 21 **c.** 5 **d.** 7

18. Given boron's electron configuration of $[He]2s^2 2p^1$, which of the following represents its electron-dot structure?

a. •Be• **b.** •Ḃ• **c.** B̈: **d.** B̈e

19. Given beryllium's electron configuration of $1s^2 2s^2$, which of the following represents its electron-dot structure?

a. •Be• **b.** •Ḃ• **c.** B̈: **d.** B̈e

20. Which electrons are represented by the dots in an electron-dot structure?

a. valence electrons **c.** only s electrons

b. inner-level electrons **d.** both a and c