

NAME:

HONORS CHEMISTRY

SECTION:

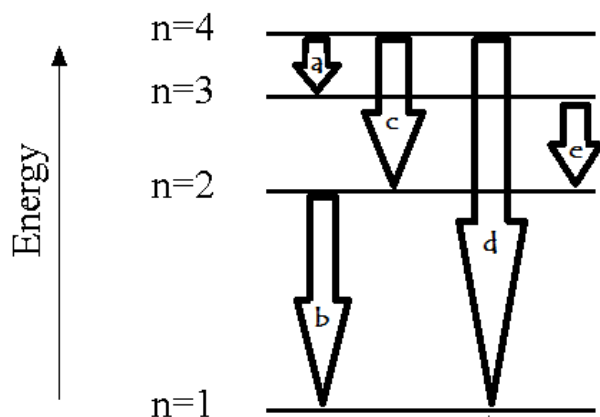
Chapter 11 Review Sheet

After studying chapter 11.1-11.10, you should be able to:

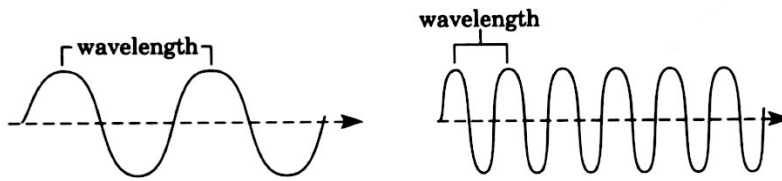
- Discuss the dual wave-particle nature of light.
- Explain and calculate the relationship between frequency, wavelength, and energy of light.
- Compare and contrast the Bohr model and the quantum model of the atom.
- Characterize the four quantum numbers, and use quantum numbers to describe electrons in atoms.
- Relate the number of sublevels corresponding to each of an atom's main energy levels, the number of orbitals per sublevel, and the number of orbitals per main energy level.
- Apply the Aufbau principle, the Pauli Exclusion Principle, and Hund's rule to write the electron configurations of the elements.
- Explain the origin of the atomic emission spectrum of an element.
- Write Lewis dot diagrams to represent the valence electrons of main-block elements.
- Identify the s, p, d and f blocks of the periodic table.

Problems for you to try

1. What do we mean when we say that light and electrons exhibit a wave-particle duality?
2. What is the frequency of radiation whose wavelength is 4.50×10^{-5} cm? What is the energy of a single photon of this radiation? In what region of the electromagnetic spectrum would it be located?
3. Which quantum has the greatest energy?



4. Which wave shown below resulted from the larger quantum jump (i.e., electron transition between allowed energy levels)? Explain how you arrived at your answer.



5. Explain how electrons change energy levels to produce emission spectra. Include a discussion of why only certain colors of light are observed. You should include a diagram of energy levels with your answer.

6. How many different energy quanta were released by the element that produced the spectrum shown below? How do you know this?



7. Define the four quantum numbers n , l , m , and s , explain what information is given by each, and describe the range of values each may take.

8. Generate a list of all the allowed values of quantum numbers l , m and s for energy level 3 ($n = 3$).

$l =$	$l =$	$l =$
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9. Sketch an s orbital, a p orbital, and a d orbital.

10. Complete the following table.

Sublevel	Number of orbitals	Maximum # of electrons in sublevel
s		
p		
d		
f		

11. Explain the three rules for writing electron configurations.

12. Using arrows and boxes, write electron configurations for:

a) P

b) Y

13. Using noble gas notation, write electron configurations for:

a) F

b) Sr

14. Identify the elements that have the following electron configurations:

a) $1s^2 2s^2 2p^6 3s^2 3p^3$

b) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

c) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$

15. Explain why the 4s sublevel fills before the 3d sublevel begins to fill as electrons are added.

16. Write Lewis dot diagrams for the following elements:

a) Ca

d) Ge

b) As

e) Kr

c) Cl

f) Rb

17. Which of the following statements of the modern model of the atom are true, and which are false?

- The probability of finding an electron is the same in any location within an orbital.
- An electron will probably spend most of its time close to the nucleus.
- Electrons travel in circular orbits around the nucleus.

18. Which of the following statements about an orbital are true and which are false?

- An electron will be found inside an orbital 90% of the time.
- An electron travels around the surface of an orbital.
- An electron cannot be found outside an orbital.