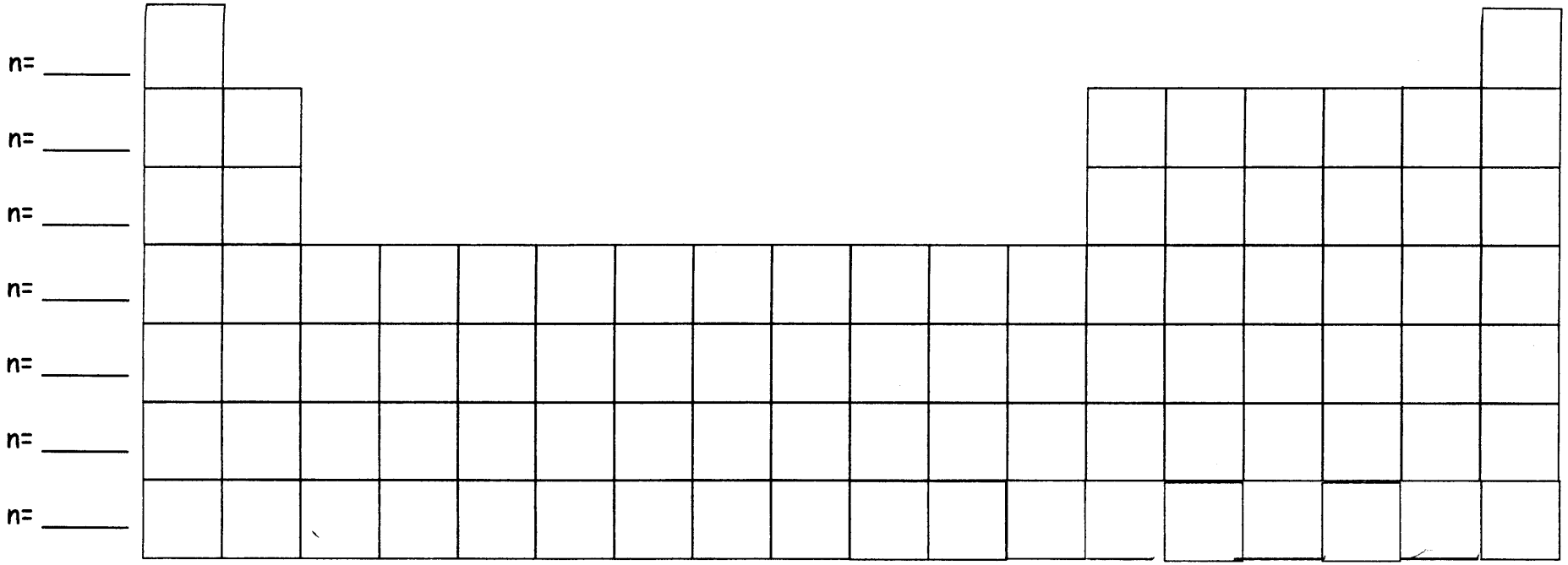
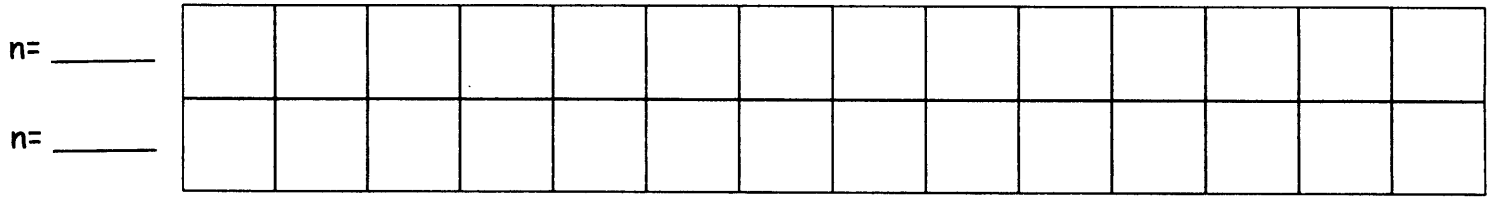
PERIODIC TABLE FOR ELECTRON CONFIGURATION





WS Quantum Numbers and Orbitals

Name _____

1. Indicate which of the following orbital destinations are possible.
   a. 7s b. 1p c. 5d d. 2d e. 4f f. 5g g. 7i

2. Without referring to a text, periodic table or handout, deduce the maximum number of electrons that can occupy an:
   a. s orbital _____ b. the subshell of p orbitals _____ c. the subshell of d orbitals _____
   d. the subshell of f orbitals _____ e. the subshell of g orbitals _____

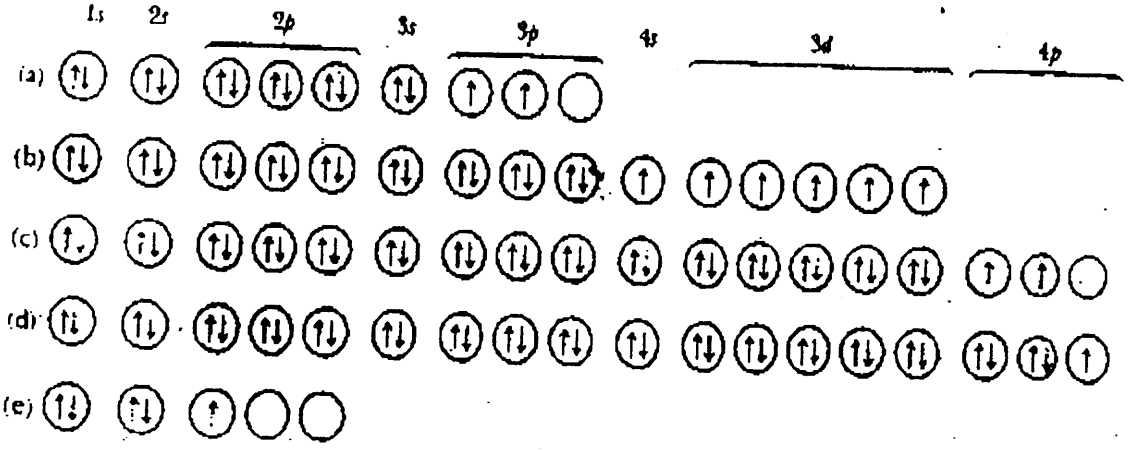
3. Explain why there are 10 members of each d transition metal series. _____

4. Explain why there are 14 members of each f inner-transition metal series. _____

5. Indicate which of the following electron configurations is ruled out by the Pauli exclusion principle.
   a. $1s^2 2s^2 2p^7$ b. $1s^2 2s^2 2p^6 3s^3$ c. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{12}$ d. $1s^2 2s^2 2p^6 3s^2 3p^6$

6. Explain why the following ground-state electron configurations are not possible:
   a. $1s^2 2s^3 2p^3$ b. $1s^2 2s^2 2p^3 3s^6$ c. $1s^2 2s^2 2p^7 3s^2 3p^8$ d. $1s^2 2s^2 2p^6 3s^2 3p^1 4s^2 3d^{14}$

7. Write the corresponding electron configurations for each of the following pictorial representations. Name the element assuming that the configuration describes a neutral atom.



Name
a) _____
b) _____
c) _____
d) _____
e) _____

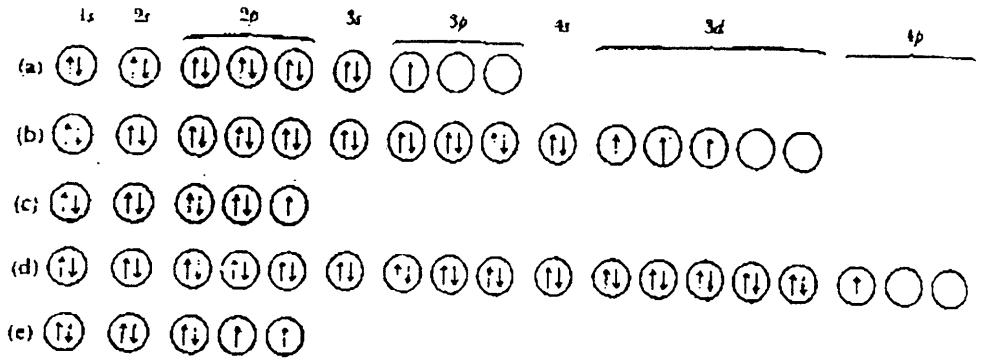
8. Write the electron configuration for the first excited state of:
   a. Be b. Be^{2+} c. He d. F^- e. O f. O^{2-}

9. Which of the following electron configurations of neutral atoms represents an excited state?
   a. $1s^2 2s^2 2p^5 3s^1$ b. $1s^2 2s^2 2p^5 3s^2$ c. $1s^1 2s^1$ d. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^1$

10. Write the ground state electron configurations for the following elements:
    a. Ti b. K c. Fe d. As e. Ag f. Zn

11. Write the ground state core notations for the following elements:
    a. P b. Ni c. Sc d. Cd e. Cr f. Mn

12. Write the corresponding electron configuration for each of the following pictorial representations. Name the element that each represents assuming they are neutral atoms.



Name

a)_____

b)_____

c)_____

d)_____

e)_____

13. Give two examples of:
    a. an atom with a half-filled subshell
    b. an atom with a completely filled outer shell
    c. an atom with its outer electrons occupying a half-filled subshell and a filled subshell.

14. Indicate which groups of elements have an outer:
    a. s electron configuration
    b. p electron configuration
    c. d electron configuration
    d. f electron configuration

    Some chemists call these various elements the s-block, p-block, d-block and f-block elements.

15. Determine the element of lowest atomic number whose ground state contains:
    a. three d electrons
    b. a complete d subshell
    c. ten total p electrons
    d. an f electron
    e. 13 d electrons
    f. 23 p electrons
    g. 7 s electrons

16. How many unpaired electrons are there in the ground state of each of the following atoms?
    a. Ge b. Se c. V d. Fe e. Si f. Mo g. Ag

17. How many unpaired electrons are there in the ground state of each of the following ions?
    a. Cl^- b. O^{2-} c. Al^{3+} d. Ca^{2+} e. Na^+ f. P^{3-} g. Xe^+

18. Write the ground-state core notation for the following ions:
    a. Cl^- b. P^{3-} c. Br^- d. Se^{2-} e. Na^+ f. Ba^{2+} g. Fe^{3+} h. Ag^+ i) Ni^{2+} j) Cr^{3+}

19. Apply Hund's rule as you write the ground-state electron configuration for:
    a. O^+ b. C^- c. F^+ d. Ar^+

20. Determine the number of unpaired electrons in the ground-state of the following species:
    a. F^+ b. Sn^{2+} c. Bi^{3+} d. Ar^+

Part I:
Determine the element whose outermost electron is being defined by the following quantum numbers.

	Element	Quantum Number Code
1.	_____	$n = 1, l = 0, m_l = 0, m_s = -\frac{1}{2}$
2.	_____	$n = 2, l = 1, m_l = -1, m_s = +\frac{1}{2}$
3.	_____	$n = 2, l = 1, m_l = 0, m_s = -\frac{1}{2}$
4.	_____	$n = 2, l = 1, m_l = 1, m_s = +\frac{1}{2}$
5.	_____	$n = 3, l = 1, m_l = -1, m_s = -\frac{1}{2}$
6.	_____	$n = 3, l = 1, m_l = 0, m_s = +\frac{1}{2}$
7.	_____	$n = 3, l = 1, m_l = 1, m_s = -\frac{1}{2}$
8.	_____	$n = 4, l = 0, m_l = 0, m_s = +\frac{1}{2}$
9.	_____	$n = 3, l = 2, m_l = -2, m_s = +\frac{1}{2}$
10.	_____	$n = 3, l = 2, m_l = -2, m_s = -\frac{1}{2}$
11.	_____	$n = 3, l = 2, m_l = 0, m_s = +\frac{1}{2}$
12.	_____	$n = 3, l = 2, m_l = 1, m_s = -\frac{1}{2}$
13.	_____	$n = 3, l = 2, m_l = 2, m_s = +\frac{1}{2}$
14.	_____	$n = 4, l = 1, m_l = -1, m_s = +\frac{1}{2}$
15.	_____	$n = 4, l = 1, m_l = 1, m_s = -\frac{1}{2}$
16.	_____	$n = 4, l = 1, m_l = 0, m_s = +\frac{1}{2}$

Part II:
Which of the following sets of quantum numbers are not allowed in an atom? For the sets of quantum numbers that are incorrect, state what is wrong.

	√ if not allowed	Explanation of error if applicable
1. $n = 2, l = 1, m_l = -1$	_____	_____
2. $n = 1, l = 1, m_l = 0$	_____	_____
3. $n = 8, l = 7, m_l = 6$	_____	_____
4. $n = 1, l = 0, m_l = 2$	_____	_____
5. $n = 3, l = 2, m_l = 2$	_____	_____
6. $n = 4, l = 3, m_l = 4$	_____	_____
7. $n = 0, l = 0, m_l = 0$	_____	_____
8. $n = 2, l = -1, m_l = 1$	_____	_____

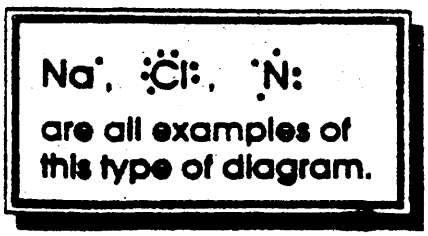
Part III:
What is the maximum number of electrons in an atom that can have these quantum numbers?

1. $n = 4$ _____
2. $n = 5, m_l = +1$ _____
3. $n = 5, m_s = +\frac{1}{2}$ _____
4. $n = 3, l = 2$ _____
5. $n = 2, l = 1$ _____
6. $n = 0, l = 0, m_l = 0$ _____
7. $n = 2, l = 1, m_l = -1, m_s = -\frac{1}{2}$ _____
8. $n = 3$ _____
9. $n = 2, l = 2$ _____
10. $n = 1, l = 0, m_l = 0$ _____

LEWIS DOT DIAGRAMS

Name _____

Lewis diagrams are a way to indicate the number of valence electrons around an atom.



Draw Lewis dot diagrams of the following atoms.

1. calcium
2. potassium
3. argon
4. aluminum
5. bromine
6. carbon
7. helium
8. oxygen
9. phosphorus
10. hydrogen

VALENCE ELECTRONS

Name _____

The valence electrons are the electrons in the outermost principal energy level. They are always "s" or "s and p" electrons. Since the total number of electrons possible in s and p sublevels is eight, there can be no more than eight valence electrons.

Determine the number of valence electrons in the atoms below.

Example: carbon
Electron configuration is $1s^2$ $\boxed{2s^2\ 2p^2}$.
Carbon has 4 valence electrons.

1. fluorine _____
2. phosphorus _____
3. calcium _____
4. nitrogen _____
5. iron _____
6. argon _____
7. potassium _____
8. helium _____
9. magnesium _____
10. sulfur _____
11. lithium _____
12. zinc _____
13. carbon _____
14. iodine _____
15. oxygen _____
16. barium _____
17. aluminum _____
18. hydrogen _____
19. xenon _____
20. copper _____

Name _____ Date _____ Class _____

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?

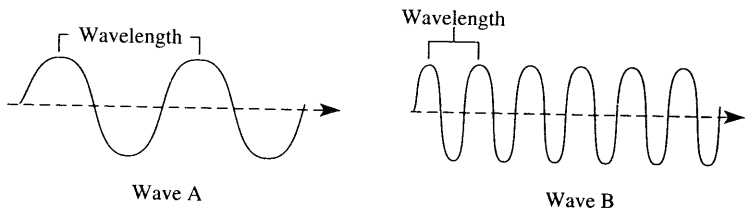
4. How can the energy levels of the atom be determined by measuring the light emitted from an atom?

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

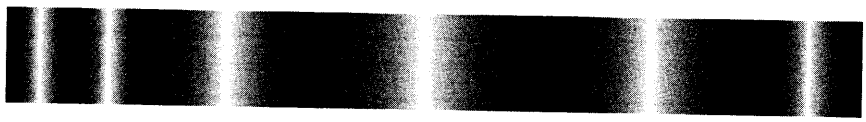
Name _____ Date _____ Class _____

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.



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 3.2×10^{-2} Hz?

Name _____ Date _____ Class _____

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
(b) 2
(c) 3
(d) 4

2. _____ A spherical electron cloud surrounding an atomic nucleus would best represent

(a) an *s* orbital.
(b) a *p* orbital.
(c) a combination of two different *p* orbitals.
(d) a combination of an *s* and a *p* orbital.

3. _____ How many electrons can an energy level of $n = 4$ hold?

(a) 32
(b) 24
(c) 8
(d) 6

4. _____ How many electrons can an energy level of $n = 2$ hold?

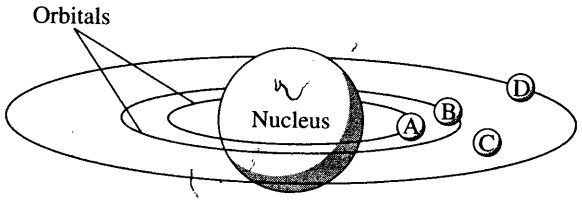
(a) 32
(b) 24
(c) 8
(d) 6

5. _____ Compared with an electron for which $n = 2$, an electron for which $n = 4$ has more

(a) spin.
(b) particle nature.
(c) energy.
(d) wave nature.

6. _____ According to Bohr, which is the point in the figure below where electrons cannot reside?

(a) point A
(b) point B
(c) point C
(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.

Name _____ Date _____ Class _____

SECTION 2 continued

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:

Principal quantum number, *n*	Number of sublevels	Types of orbitals
1		
2		
3		
4		

Name _____ Date _____ Class _____

CHAPTER 4 REVIEW

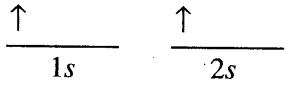
Arrangement of Electrons in Atoms

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:



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

3. Phosphorus

4. Nitrogen

5. Potassium

Name _____ Date _____ Class _____

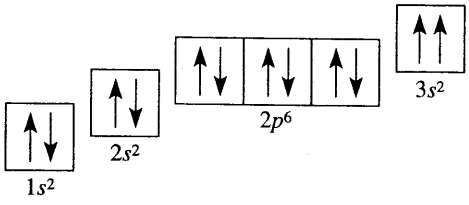
SECTION 3 continued

6. Aluminum

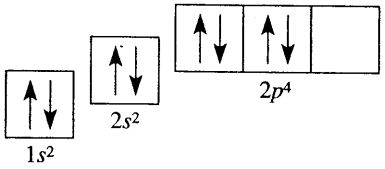
7. Argon

8. Boron

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



a. _____



b. _____

Name _____ Date _____ Class _____

CHAPTER 4 REVIEW

Arrangement of Electrons in Atoms

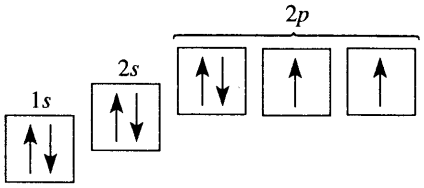
MIXED REVIEW

SHORT ANSWER **Answer the following questions in the space provided.**

1. Under what conditions is a photon emitted from an atom?

2. What do quantum numbers describe?

3. What is the relationship between the principal quantum number and the electron configuration?



4. In what way does the figure above illustrate Hund's rule?

5. In what way does the figure above illustrate the Pauli exclusion principle?

Name _____ Date _____ Class _____

MIXED REVIEW continued

6. Elements of the fourth and higher main-energy levels do not seem to follow the normal sequence for filling orbitals. Why is this so?

_____

_____

_____

7. How do electrons create the colors in a line-emission spectrum?

_____

_____

_____

8. Write the ground-state electron configuration of the following atoms:

a. Carbon

_____

b. Potassium

_____

c. Gallium

_____

d. Copper

_____

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

9. _____ What is the wavelength of light that has a frequency of 3×10^{-4} Hz in a vacuum?

10. _____ What is the energy of a photon that has a frequency of 5.0×10^{14} Hz?