6

Electronic Structure and the Periodic Table

I. Light

A. Wave Nature of Light

1. Wavelength

a. Symbol: ______________________________

b. Definition: ______________________________

____

c. Units: ______________________________

2. Frequency

a. Symbol: ______________________________

b. Definition: ______________________________

____

c. Units: ______________________________

3. Mathematical relationship between frequency and wavelength

Exercises

a. Microwaves have a wavelength of 1.000 cm. Calculate their frequency in s^{-1}.

Solution: This problem asks you to relate frequency and wavelength. Hence the equation to use is

$$\nu = \frac{c}{\lambda}$$

Since c is 2.998×10^8 m/s, ν, the wavelength, should be in the same unit of length, m. Thus,

$$\nu = \frac{2.998 \times 10^8 \frac{\text{m}}{\text{s}}}{1.000 \times 10^{-2}\ \text{m}} = 2.998 \times 10^{10}\ \text{s}^{-1}$$

b. Calculate the wavelength in nm of a band on your AM radio with a frequency of 4.56×10^5 Hz. **(E1)**

4. Amplitude

a. Symbol: ______________________

b. Definition: ______________________

____

B. Particle nature of light
Two scientists who started to consider light as a stream of particles (photons)

______________ and ______________

C. Energy (E) and wave characteristics

1. Mathematical relationship between energy and

a. wavelength

b. frequency

2. Constants in the mathematical equations

a. h

(1) What does it stand for? ______________

(2) What is its value? ______________

b. c

(1) What does it stand for? ______________

(2) What is its value? ______________

3. Exercises

a. Calculate the energy of microwaves in joules, using the information obtained in the preceding worked exercise.

Solution: The microwaves in the preceding problem had a frequency of $2.998 \times 10^{10}\ s^{-1}$. We therefore use the relation between frequency and energy

$$E = h\nu$$

where h is Planck's constant. Substituting, we get

$$E = (6.626 \times 10^{-34}\ J \cdot s)(2.998 \times 10^{10}\ s^{-1}) = 1.986 \times 10^{-23}\ J$$

b. Calculate the energy in kJ/mol of an AM radio wave using the information in Exercise E1. **(E2)**

II. The Hydrogen Atom

A. Atomic spectrum of hydrogen

1. Device that breaks up light into its component colors: ____________________

2. Series of lines in the ultraviolet range: ____________________

3. Series of lines in the visible range: ____________________

4. Series of lines in the infrared range: ____________________

B. Bohr model of the hydrogen atom

1. Definition of ground state: ____________________

____

2. Definition of excited state: ____________________

____

3. Equation for the energy of the hydrogen atom:

R_H is ____________________

n is ____________________

4. Equation relating energy to the level that the electron is in:

5. *Exercises*

a. Find the energy that is given off in the transition from n = 5 to n = 4.
Solution: The energy is given by the equation

$$E = R_H\left(\frac{1}{(n_{lo})^2} - \frac{1}{(n_{hi})^2}\right)$$

We substitute into this equation 5 for n_{hi} and 4 for n_{lo}.

$$E = 2.180 \times 10^{-18}\ \text{J}\left(\frac{1}{(4)^2} - \frac{1}{(5)^2}\right) = 4.905 \times 10^{-20}\ \text{J}$$

b. Calculate the wavelength of the light given off in an electron transition between two states where the energy difference is 182 kJ/mol. **(E3)**

c. Calculate the frequency of the light that results from the transition from n = 6 to n = 3. **(E4)**

C. Quantum mechanical model

1. Quantum mechanics was developed to describe the motion of small particles confined to tiny regions of space.

2. The quantum mechanical model does not specify the position of the electron or the path that it takes about the nucleus. It only deals with ________________

____.

3. ________________________ developed an equation to calculate the height of the electron wave at various points in space.

4. The wave function is given the Greek letter ______________________.

5. The square of the wave function is directly proportional to ______________________

______________________.

6. In the hydrogen atom the electron is more likely to be found ______________________.

III. Quantum Numbers and Energy Levels

A. Definitions

1. Quantum number: ______________________

____

2. Atomic orbital: ______________________

____

B. First quantum number

1. Symbol: ______________________.

2. Values that can be taken on by the first quantum number: ______________________

3. The value of the first quantum number designates ______________________

______________________.

C. Second quantum number

1. Symbol: ______________________

2. The value of the second quantum number designates ______________________

______________________.

3. The second quantum number determines ______________________

______________________.

4. Possible values for ℓ: ______________________

These are always linked to n. The smallest possible ℓ is always 0, and the largest ℓ is $n - 1$. Therefore, if you want all possible values for ℓ in the energy level 1 ($n = 1$), then you are allowed to go from 0 to $(n - 1)$. In this case, you go from 0 to $(1 - 1)$, which is also 0. Thus, there is only one possible sublevel in $n = 1$ and that is $\ell = 0$. What are the possible ℓ values in the principal energy level $n = 4$? **(E5)**

5. Designating sublevels in letters instead of numbers

$\ell = 0$ can be designated as the s sublevel
$\ell = 1$ can be designated as the p sublevel
$\ell = 2$ can be designated as the d sublevel
$\ell = 3$ can be designated as the f sublevel

These are all the sublevels you have to know for all the elements that occur naturally plus all those that have been created so far. Memorize these letter designations because you will need to be able to interchange the two designations quickly.

6. Combining n and ℓ
A sublevel is usually referred to in conjunction with the principal energy level that it is in. Thus, a sublevel can be designated as 3p, which means that $n = 3$ and $\ell = 1$. If the designation is 4f, then $n = 4$ and $\ell = 3$.
Warning: Not all combinations will work! Remember that ℓ can only go up to $n - 1$. Thus, 3f is not possible since in 3f, $n = 3$ and $\ell = 3$. For $n = 3$, the maximum ℓ possible is 2 or d. For similar reasons, 1d is not allowed.

7. For multi-electron atoms, energy is dependent on ℓ as on n. Within a given principal level (same value of n) sublevels increase in energy in the order

______________________________________.

D. Third quantum number

1. Symbol: ______________________________________

2. Each sublevel contains one or more ______________________ which differ from one another in the value assigned to m_ℓ.

3. m_ℓ tells us how the electron cloud surrounding the nucleus is directed in space.
   a. If $\ell = 0$, then $m_\ell = 0$. This is known as the s orbital. The general appearance of the orbital cloud associated with the s orbital is

   ______________________________________.

   b. If $\ell = 1$, then m_ℓ can be 1, 0, −1. They are more commonly referred to as p_x, p_y, and p_z orbitals.

   Describe a p orbital: ______________________________________

   ____

   Orientation of the p orbitals to each other: ______________________

   ____

4. Possible values for m_ℓ:

$$\ell, \ldots, 0 \ldots, -\ell$$

5. *Exercises*
   a. If $\ell = 2$, what are the possible m_ℓ values?
   Solution: Since $\ell = 2$ and m_ℓ can go from ℓ to 0 to $-\ell$, the possible values are −2, −1, 0, 1, 2.
   b. What are the possible m_ℓ values for the f sublevel? **(E6)**

6. Total number of orbitals per sublevel = $2\ell + 1$.

E. Fourth quantum number
   1. Symbol: ______________________
   2. m_s tells us the spin of the electron about its axis.
   3. m_s is not related to n, ℓ, or m_ℓ. It has only two possible values: $+\frac{1}{2}$ or $-\frac{1}{2}$.

F. Pauli exclusion principle
   1. Statement of the principle: ______________________

   ____
   2. The Pauli exclusion principle and the rules of quantum mechanics fix the capacities of principal levels, sublevels and orbitals. Study the summary given in Table 6.3 of your text.
   3. *Exercises*
      a. Give the quantum numbers for the electrons in the $m_\ell = -2$ orbital of the 3d level.
      Solution: 3d means n = 3 and $\ell = 2$. Thus the quantum numbers are

$$n = 3 \qquad \ell = 2 \qquad m_\ell = -2 \qquad m_s = +\tfrac{1}{2}$$

$$n = 3 \qquad \ell = 2 \qquad m_\ell = -2 \qquad m_s = -\tfrac{1}{2}$$

      As you can see, only the spins distinguish these two electrons.
      b. How many electrons are possible in the 4f sublevel?
      Solution: In the 4f sublevel $\ell = 3$; thus the number of possible orbitals is $2\ell + 1$ or 7. Since there are 2 electrons in each orbital, the number of possible electrons in the 4f sublevel is 2(7) = 14.
      c. How many electrons are possible in n = 3? In n = 2, how many electrons are possible with a set of quantum numbers that includes n = 2 and $\ell = 1$? **(E7)**

IV. Electron configurations in the atom

A. From sublevel energies

1. How to write them

a. Determine the number of electrons in the atom by looking up the element's atomic number (Z).

b. Start with n = 1 and $\ell = 0$ by writing 1s. Put a superscript above s to designate how many electrons you wish to put there. Remember that for s the maximum superscript is 2. Distribute the electrons, increasing sublevels and energy levels as you have electrons.

c. ***Exercises:***

(1) Write the electron configuration for sodium.

Solution: Sodium has atomic number 11. 1s can have 2 for a superscript since you have more than 2 electrons to distribute. Thus, start the electron configuration with $1s^2$, which fills up n = 1, since you can only have $\ell = 0 = s$. Then go to n = 2, with 2 possible ℓ's, $\ell = 0 = s$ and $\ell = 1 = p$. You therefore write

$$1s^2\,2s\,2p$$

2s can again have only a superscript of 2, and since you still have nine electrons to distribute, you can write $1s^2\,2s^2\,2p$. Now you have only seven electrons left. The p sublevel can have $m_\ell = -1, 0, 1$, or three orbitals, each orbital having a capacity for two electrons. Thus, 2p can have a maximum superscript of 6. (Using the same reasoning, all p sublevels in any n level can have a maximum superscript of 6.) Since you still have seven electrons, you can write $1s^2\,2s^2\,2p^6$. Now, you have only one electron left. So you start with the third energy level, which can have $\ell = 0 = s$, $\ell = 1 = p$, and $\ell = 2 = d$. Capacities for the sublevels and therefore maximum superscripts are s = 2, p = 6, and d = 10. Since you have only one electron left, you only need 3s and the superscript in this case is 1 because that is all you have. Putting it all together, the electron configuration for sodium is

$$1s^2\,2s^2\,2p^6\,3s^1$$

(2) Write the electron configuration for silicon. **(E8)**

2. Sequence of levels and sublevels

a. The levels and sublevels are written in increasing order of energy. Thus, n = 1 comes before n = 2, and so on.

b. The sublevels are ordered s < p < d < f.

c. Ordinarily, sublevels are filled to capacity before the next one starts to fill. This works out quite well for elements with atomic numbers 1 through 18.

d. After 3p, sublevels start to overlap. Use Figure 6.8 in your text to determine level and sublevel sequences for elements with atomic numbers up to 36. Figure 6.8 gives the order and populations as

$$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6$$

3. Abbreviated electron configurations

a. The abbreviated electron configuration starts with the preceding noble gas, written between brackets.

b. ***Exercises***

(1) Write the abbreviated electron configuration for cobalt.

Solution: The preceding noble gas is argon, Ar. We write [Ar] to indicate that the first 18 electrons in the cobalt atom have the electron configuration $1s^2\,2s^2\,2p^6\,3s^2\,3p^6$. Since cobalt has 27 electrons, we have nine electrons left. Since 4s comes before 3d and after 3p and since 4s can hold two electrons, our abbreviated electron configuration now looks like this

$$[Ar]\ 4s^2$$

We have seven more electrons to distribute. The next energy level and sublevel available is 3d which can hold a maximum of ten electrons. We therefore put the remaining seven electrons into the 3d sublevel, and the abbreviated electron configuration for cobalt is

$$[Ar]\ 4s^2 3d^7$$

(2) Write the abbreviated electron configuration for the selenium atom. **(E9)**

B. From the periodic table

1. The atoms of elements in a group of the periodic table have the same distribution of electrons in the outermost principal energy level, n.

2. Summary of the order in which electrons fill sublevels using the periodic table.

a. The elements in Groups 1 and 2 fill ______________________________.

b. The elements in Groups 13 through 18 fill ______________________________.

c. The transition metals fill ______________________.

d. The lanthanide metals fill ______________________.

e. The actinide metals fill ______________________.

3. Figure 6.9 in your text can be used to deduce the electron configuration of any element.

4. *Exercises*

   a. Write the electron configuration for antimony (Sb).

   Solution: We look for antimony in the periodic table and find that it is past period 1, so the 1s sublevel is filled: $1s^2$.
   Antimony is past period 2 in which the 2s and 2p sublevels are filled: $2s^2\ 2p^6$.
   Antimony is past period 3 in which the 3s and 3p sublevels are filled: $3s^2\ 3p^6$.
   Antimony is past period 4 in which the 4s, 3d, and 4p (in that order) sublevels are filled: $4s^2\ 3d^{10}\ 4p^6$.
   Antimony is in the fifth period, past Groups 1 and 2, thus the 5s sublevel is filled ($5s^2$), and past the transition metal groups (Groups 3-12) so the 4d sublevel is filled ($4d^{10}$). Antimony is in Group **15**, which means that there are 3 electrons in the 5p sublevel ($5p^3$).
   Putting it all together, the electron configuration for antimony is

$$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^3$$

   Check your work by adding up the superscripts. Your total should equal the atomic number for antimony, which it does.

   b. Write the electron configuration for lead. **(E10)**

V. Orbital Diagrams of Atoms

A. Setting up diagrams

1. You will need to know the electron configuration of the atom for which you want to write an orbital diagram.

2. *Exercise*

   Set up the orbital diagram for sulfur.
   The electron configuration for sulfur is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$. Drawing a pair of parentheses to represent each orbital in the sublevel, you find that the setup will look like this:

$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^4$
( )	( )	( )( )( )	( )	( )( )( )

B. Filling in the orbitals

1. Hund's rule

There is no problem in filling in the orbitals in the s sublevels. Hund's rule comes into play when filling orbitals in the p, d, and f sublevels, when they are not filled to capacity. In the case of sulfur, filling the orbitals is straightforward until we come to the 3p sublevel. Hund's rule says that all orbitals are first filled with one electron each. Then if there are more electrons to go around, the orbitals can be filled with the second electron. The orbital diagram for sulfur is

(↑↓)	(↑↓)	(↑↓)(↑↓)(↑↓)	(↑↓)	(↑↓)(↑)(↑)
1s	2s	2p	3s	3p

2. Exercises

a. Write the orbital diagram for cobalt. **(E11)**

b. Write the orbital diagram for fluorine. **(E12)**

VI. Electron Configurations in Monatomic Ions

A. Ions with noble gas structures.

1. Main-group elements form ions that have the same number of electrons as the noble gas atom.
2. Note that the noble gas atom closest to the element does not necessarily have to be the preceding noble gas. Sodium, for example, loses its $3s^1$ electron to achieve the configuration of Ne, which precedes it, but fluorine gains an electron to achieve the configuration of neon, which follows it in the periodic table.
3. Species with the same electron configuration are said to be ______________.

B. Transition metal cations

1. The transition metals (Groups 3-12) do not form ions with noble gas configurations.
2. Which electrons are lost when transition metal atoms form cations with +1 or +2 charges?

____

3. Which electrons are lost when transition metal atoms form cations with +3 or +4 charges?

____

4. *Exercises*

a. Write the electron configuration for As^{3-} and for Cr^{3+}.

Solution: *(1)* Arsenic is a main-group metal, so it forms the electron configuration of its nearest noble gas, which is Kr. Its electron configuration is

$$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6$$

(2) Chromium is a transition metal. Its cation has a charge of +3 so it loses both of its outermost s electrons first ($4s^2$). It also loses one of its outermost d electrons (3d). The electron configuration for the Cr^{3+} is

$$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^3$$

To check, Cr metal has 24 electrons. Cr^{3+} lost three electrons, so there must be 21 electrons left. Adding the superscripts in the electron configuration for Cr^{3+} gives 21 as a total.

b. Write the abbreviated electron configuration for Fe^{2+}, Fe^{3+}, and Te^{2-}. **(E13)**

VII. Periodic Trends in the Properties of Atoms

A. Atomic radius

1. Atomic radius is a measure of ________________________________.
2. Atomic radius decreases as we move across the periodic table from left to right and increases as we move down within a group.
3. Effective nuclear charge
   Effective nuclear charge is the number of electrons in the outer electron configuration. Sodium's outer electron configuration is $3s^1$. Thus, its effective nuclear charge is 1.
4. As we go down a group, the effective nuclear charge does not change, but the energy levels (or layers) increase, so size increases. Going across the periodic table, the number of layers is constant, but the effective nuclear charge increases. Thus electrons are pulled in toward the nucleus, and size decreases.

B. Ionic radius

1. Just as with their corresponding atoms, ionic radii increase as we move down within a group. The ionic radii increase as we move across a period.
2. Cations are smaller than their corresponding atoms, because in the cation there are more protons than electrons. Therefore, the electrons are attracted and move closer to the nucleus.
3. Anions are larger than their corresponding atoms, because the presence of the extra electrons causes the electrons to repel each other. As a result, they try to move as far away from each other as possible, creating bigger clouds.

C. Ionization energy
1. It is a measure of ____

____.
2. The smaller the ionization energy is, the easier the electron is to remove.
3. As we go across the periodic table from left to right, the distance between the nucleus and the electrons decreases, the electrons become more tightly bound, and the ionization energy increases. As we move down a group, the electrons are separated from the nucleus by more layers. Consequently, the nucleus exerts less attraction on the electrons, and it becomes easier to pull out an electron. So ionization energy gets smaller.

D. Electronegativity
1. It is a measure of ____

____.
2. The greater the electronegativity, the greater affinity the atom has for electrons.
3. Electronegativity ____________________ as one moves from left to right in the periodic table.
4. Electronegativity ____________________ as one moves down a group in the periodic table.

SELF-TEST
A. Multiple Choice:
1. How many unpaired electrons are there in an atom with the electron configuration

$$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6$$

a. 1 **b.** 2 **c.** 3 **d.** 4 **e.** some other number

2. Which one of the following represents a possible *excited* electron configuration, in which an electron has been promoted from the ground state to a higher level?

a. $1s^2\,2s^2\,2p^6\,3s^2$ **b.** $1s^2\,2s^2\,2p^6\,3s^1\,3p^1$ **c.** $1s^2\,2s^2\,2p^6\,3s^3$
d. $1s^2\,2s^2\,2p^5\,2d^1\,3p^2$ **e.** $1s^2\,2s^2\,2p^7\,3s^1$

3. Which of the following statements is correct?
a. Atomic radii increase as one moves from left to right across a period.
b. Atomic radii increase as one moves down a group.
c. Negative ions are smaller than the nonmetals from which they are formed.
d. Positive ions are larger than the metals from which they are formed.

4. In which of the following series of atoms are the first ionization energies decreasing?
a. Sr > Ba > Ra **b.** Li > Be > B
c. O > F > S **d.** As > P > N

5. The maximum number of electrons possible in an f sublevel is
a. 2 **b.** 6 **c.** 14 **d.** 32

6. There are two very intense lines in the atomic spectrum of sodium. The wavelengths are 589.0 nm and 589.6 nm. Which line corresponds to the larger energy?
a. Energy is the same for both lines.
b. Energy is greater for the 589.0 nm line.
c. Energy is greater for the 589.6 nm line.
d. Cannot answer because insufficient information is given.

7. How many elements in the second period of the periodic table (starting with Li) have one or more unpaired electrons?
a. 3 **b.** 4 **c.** 5 **d.** 6 **e.** 7

8. A filled d sublevel and a half-filled f sublevel contain a *total* of how many electrons?
a. 11 **b.** 17 **c.** 19 **d.** 24

9. Given the following orbital diagram

	1s	2s	2p	
—	(↑)	(↑)	( )( )( )	= hydrogen atom – excited state
—	(↑↓)	( )	( )( )( )	= hydrogen atom – ground state
—	(↑↓)	( )	( )( )( )	= helium atom – ground state
—	(↑)	(↑)	( )( )( )	= helium atom – excited state

The number of correct orbital diagrams is
a. 0 **b.** 1 **c.** 2 **d.** 3 **e.** 4

10. The radius of

(1) Br^- > Br **(2)** Li < Na **(3)** Mg > Cs **(4)** K^+ > Cs^+

The true statements above are
a. (1),(2) **b.** (1),(3) **c.** (2),(3) **d.** (2),(4) **e.** (3),(4)

B. True or False

______ **1.** The energy required to make the transition from $n = 1$ to $n = 3$ is twice the energy required to make the transition from $n = 1$ to $n = 2$.

______ **2.** A 3p subshell can have an electron with quantum number $\ell = 2$.

______ **3.** An important factor in determining the relative size of halogen atoms in their ground state is the principal quantum number.

______ **4.** The first two quantum numbers (n, ℓ) of the highest energy electron in the ground state of Sc^{3+} are 3,1.

______ **5.** The electron configuration $1s^2\ 2s^2\ 2p^4\ 2d^1$ may be that of fluorine in an excited state.

______ **6.** Pauli's exclusion principle states that two electrons in the same orbital have the same spin.

______ **7.** It is possible for an electron to have $4,0,1,\frac{1}{2}$ as its set of 4 quantum numbers.

______ **8.** The ionic radius of Na^+ is larger than the atomic radius of Na.

C. Fill in the Blanks

____________________ **1.** How many electrons with quantum number $\ell = 1$ are there in the ground state electron configuration for potassium?

____________________ **2.** How many electrons can have both quantum numbers $n = 2$ and $\ell = 1$?

____________________ **3.** How many unpaired electrons are there for arsenic in its ground state?

____________________ **4.** What is the abbreviated ground state electron configuration for Zr^{3+}?

____________________ **5.** What neutral element is represented by the electron configuration $1s^2\ 2s^2\ 2p^6\ 3s^1\ 3p^1$?

____________________ **6.** Does the electron configuration in #5 represent the element in its ground state?

____________________ **7.** What is the symbol of the most electronegative element in Group 15?

____________________ **8.** Which neutral element in Period 4 has the largest radius?

____________ **9.** Which neutral element in Group 2 has the largest first ionization energy?

____________ **10.** Write the abbreviated ground state electron configuration for nickel.

D. Answer the questions below, using **LT** (for *less than*), **GT** (for *greater than*), **EQ** (for *equal to*), or **MI** (for *more information required*) in the blanks provided.

______ **1.** The wavelength of the photon required to promote an electron in the hydrogen atom from the n = 1 to the n = 3 level is (1) the wavelength of the photon required to promote an electron in the hydrogen atom from the n = 1 to the n = 2 level.

______ **2.** For a scandium atom in the ground state, the energy of the 3d sublevel is (2) the energy of the 4s sublevel.

______ **3.** The number of unpaired electrons for a carbon atom in the ground state is (3) the number of unpaired electrons for an oxygen atom in the ground state.

______ **4.** The ℓ quantum number of the 3p sublevel is (4) the ℓ quantum number of the 5p sublevel.

______ **5.** The atomic radius of the fluorine atom is (5) the atomic radius of the fluoride ion.

______ **6.** The atomic radius of the magnesium atom is (6) the atomic radius of the chlorine atom.

______ **7.** The electronegativity of the halogen atom in the second period of the periodic table is (7) the electronegativity of the halogen atom in the fourth period of the periodic table.

______ **8.** The first ionization energy of potassium is (8) the first ionization energy of calcium.

E. Problems:

Consider the second line of the Paschen series (n = 5 to n = 3) of the hydrogen atom.

1. Calculate ΔE in kJ/mol.

2. Calculate λ.

3. Calculate ν.

Consider technecium (Z = 43).

4. Write its electron configuration.

5. Write its orbital diagram.

6. Write its abbreviated electron configuration.

7. Write the abbreviated electron configuration for Tc^{2+}.

A new element was made and found to have 115 protons and 187 neutrons,

8. Its atomic number is ______________________________.

9. Its outer electron configuration is ______________________________.

10. The element it most resembles in chemical properties is ______________________________.

11. Compare its atomic radius with that of Fr.

12. Compare its ionization energy with that of As.

13. Compare its electronegativity with that of Po.