ELECTRONS IN ATOMS

Practice Problems
In your notebook, solve the following problems.

SECTION 5.1 MODELS OF THE ATOM
1. How many sublevels are in the following principal energy levels?
   a. \( n = 1 \) 1
   b. \( n = 2 \) 2
   c. \( n = 3 \) 3
   d. \( n = 4 \) 4
   e. \( n = 5 \) 5
   f. \( n = 6 \) 6
2. How many orbitals are in the following sublevels?
   a. 1s sublevel
   b. 5s sublevel
   c. 4d sublevel
   d. 4f sublevel
   e. 7s sublevel
   f. 3p sublevel
3. What are the types of sublevels and number of orbitals in the following energy levels?
   a. \( n = 1 \) s
   b. \( n = 2 \) s,p
   c. \( n = 3 \) s,p,d
   d. \( n = 4 \) s,p,d,f

SECTION 5.2 ELECTRON ARRANGEMENT IN ATOMS
1. Write a complete electron configuration of each atom.
   a. hydrogen
   b. vanadium
   c. magnesium
   d. barium
   e. bromine
   f. sulfur
   g. krypton
   h. arsenic
   i. radon

SECTION 5.3 PHYSICS AND THE QUANTUM MECHANICAL MODEL
1. What is the wavelength of the radiation whose frequency is \( 5.00 \times 10^{15} \text{ s}^{-1} \)?
   In what region of the electromagnetic spectrum is this radiation?
   \[ \text{Wavelength} = \frac{c}{\nu} = \frac{2.998 \times 10^8 \text{ m/s}}{5.00 \times 10^{15} \text{ s}^{-1}} = 6.0 \times 10^{-8} \text{ m} \]
   \[ \text{Wavelength} = 6.0 \times 10^{-8} \text{ cm} \]
   \[ \lambda = 6.0 \times 10^{-8} \text{ cm} \]
2. An inexpensive laser that is available to the public emits light that has a wavelength of 670 nm. What are the color and frequency of the radiation?
   \[ \lambda = 670 \text{ nm} \]
   \[ \nu = \frac{c}{\lambda} = \frac{2.998 \times 10^8 \text{ m/s}}{670 \times 10^{-9} \text{ m}} = 4.5 \times 10^{14} \text{ Hz} \]
   \[ \nu = 4.5 \times 10^{14} \text{ Hz} \]
3. What is the energy of a photon whose frequency is \( 2.22 \times 10^{14} \text{ s}^{-1} \)?
   \[ E = h\nu = (6.626 \times 10^{-34} \text{ J s}) \times (2.22 \times 10^{14} \text{ Hz}) = 1.47 \times 10^{-19} \text{ J} \]
4. What is the frequency of a photon whose energy is \( 6.00 \times 10^{-15} \text{ J} \)?
   \[ \nu = \frac{E}{h} = \frac{6.00 \times 10^{-15} \text{ J}}{6.626 \times 10^{-34} \text{ J s}} = 9.05 \times 10^{18} \text{ Hz} \]
5. Arrange the following types of electromagnetic radiation in order of increasing frequency.
   a. infrared
   b. gamma rays
   c. visible light
   d. radio waves
   e. microwaves
   f. ultraviolet

6. Suppose that your favorite AM radio station broadcasts at a frequency of 1600 kHz. What is the wavelength in meters of the radiation from the station?
   \[ \lambda = \frac{c}{\nu} = \frac{2.998 \times 10^8 \text{ m/s}}{1600 \times 10^3 \text{ Hz}} = 1.998 \times 10^{-3} \text{ m} \]
   \[ \lambda = 1.998 \times 10^{-3} \text{ m} \]
Practice Problems
In your notebook, solve the following problems.

SECTION 6.1 ORGANIZING THE ELEMENTS

1. Which element listed below should have chemical properties similar to fluorine (F)?
   a. Li
   b. Si
   c. Br
   d. Ne

2. Identify each element as a metal, metalloid, or nonmetal:
   a. fluorine NM
   b. germanium Metalloid
   c. zinc M
   d. phosphorus NM
   e. lithium M

3. Which of the following is not a transition metal?
   a. magnesium
   b. titanium
   c. chromium
   d. mercury

4. Name two elements that have properties similar to those of the element potassium. Li/Na/Rb/Cs/Fr - any 2

5. Elements in the periodic table can be divided into three broad classes based on their general characteristics. What are these classes and how do they differ?

<table>
<thead>
<tr>
<th>Property</th>
<th>Metal</th>
<th>NonMetal</th>
<th>Metalloid</th>
</tr>
</thead>
<tbody>
<tr>
<td>Conductor</td>
<td>Good</td>
<td>Poor</td>
<td>both</td>
</tr>
<tr>
<td>Sheen</td>
<td>High Luster</td>
<td>Dull</td>
<td>both</td>
</tr>
<tr>
<td>Malleability</td>
<td>Very Malleable</td>
<td>Brittle</td>
<td>both</td>
</tr>
</tbody>
</table>
SECTION 6.2 CLASSIFYING THE ELEMENTS

1. Use the periodic table to write the electron configuration for silicon. Explain your thinking.
   \[1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^2\]

2. Use the periodic table to write the electron configuration for iodine. Explain your thinking.
   \[1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 3d^{10} \ 4s^2 \ 4p^6 \ 4d^{10} \ 5s^2 \ 5p^5\]

3. Which group of elements is characterized by an \(s^2p^3\) configuration?
4. Name the element that matches the following description.
   a. one that has 5 electrons in the third energy level
   b. one with an electron configuration that ends in \(4s^24p^5\)
   c. the Group 6A element in period 4

5. Identify the elements that have electron configurations that end as follows.
   a. \(2s^22p^4\) oxygen
   b. \(4s^2\) calcium

6. What is the common characteristic of the electron configurations of the elements Ne and Ar? In which group would you find them?

7. Why would you expect lithium (Li) and sulfur (S) to have different chemical and physical properties?

8. What characterizes the electron configurations of transition metals such as silver (Ag) and iron (Fe)?

   All in the d-block

SECTION 6.3 PERIODIC TRENDS

1. Explain why a magnesium atom is smaller than atoms of both sodium and calcium.

2. Predict the size of the astatine (At) atom compared to that of tellurium (Te).
   Explain your prediction.

3. Would you expect a \(Cl^-\) ion to be larger or smaller than an \(Mg^{2+}\) ion? Explain.

4. Which effect on atomic size is more significant, an increase in nuclear charge across a period or an increase in occupied energy levels within a group?

5. Explain why the sulfide ion \((S^{2-})\) is larger than the chloride ion \((Cl^-)\).

6. Compare the first ionization energy of sodium to that of potassium.

7. Compare the first ionization energy lithium to that of beryllium.

8. Is the electronegativity of barium larger or smaller than that of strontium?
   Explain.

9. What is the most-likely ion for magnesium to form? Explain.

10. Arrange oxygen, fluorine, and sulfur in order of increasing electronegativity.
IONIC AND METALLIC BONDING

Practice Problems

In your notebook, answer the following.

SECTION 7.1 IONS

1. For each element below, state (i) the number of valence electrons in the atom, (ii) the electron dot structure, and (iii) the chemical symbol(s) for the most stable ion.
   a. Ba 2/ Ba²⁺ / Ba³⁺  
   b. I 7/ 15/ I⁻  
   c. K 1/ K⁺ / K²⁺

2. How many valence electrons does each of the following atoms have?
   a. gallium 3  
   b. fluorine 7  
   c. selenium 6

3. Write the electron configuration for each of the following atoms and ions.
   a. Ca 1s²  
   b. chlorine atom 3p⁵  
   c. Na⁺ (lose 1e) 3p⁶  
   d. phosphide ion (gain 3e) 3p⁶

4. What is the relationship between the group number of the representative elements and the number of valence electrons?

5. How many electrons will each element gain or lose in forming an ion? State whether the resulting ion is a cation or an anion.
   a. strontium 8  
   b. aluminum 3  
   c. tellurium 6  
   d. rubidium 1  
   e. bromine 7  
   f. phosphorus 5

6. Give the name and symbol of the ion formed when
   a. a chlorine atom gains one electron. Cl⁻  
   b. a potassium atom loses one electron. K⁺  
   c. an oxygen atom gains two electrons. O²⁻  
   d. a barium atom loses two electrons. Ba²⁺

7. How many electrons are lost or gained in forming each of the following ions?
   a. Mg²⁺  
   b. Br⁻  
   c. Ag⁺  
   d. Fe³⁺

8. Classify each of the following as a cation or an anion.
   a. Na⁺  
   b. Cu²⁺  
   c. F⁻  
   d. O²⁻  
   e. Ca²⁺  
   f. Cs⁺
SECTION 7.2 IONIC BONDS AND IONIC COMPOUNDS

1. Use electron dot structures to predict the formula of the ionic compounds formed when the following elements combine.
   a. sodium and bromine $\text{NaBr}$
   b. sodium and sulfur $\text{Na}_2\text{S}$
   c. calcium and iodine $\text{CaI}_2$
   d. aluminum and oxygen $\text{Al}_2\text{O}_3$
   e. barium and chlorine $\text{BaCl}_2$

2. Which of these combinations of elements are most likely to react to form ionic compounds?
   a. sodium and magnesium
   b. barium and sulfur
   c. potassium and iodine
   d. oxygen and argon

SECTION 7.3 BONDING IN METALS

1. What is a metallic bond?
   cation surrounded by mobile electrons

2. How is the electrical conductivity of a metal explained by metallic bonds?
   free moving e-

   yes, tight & orderly arranged cations

4. Is copper a metallic crystal?
   yes

5. What is an alloy?
   mixture of metals or metals & nonmetals

6. Name the principal elements present in each of the following alloys.
   a. brass $\text{Cu}/\text{Zn}$
   b. bronze $\text{Cu}/\text{Sn}$
   c. stainless steel $\text{Fe}$
   d. sterling silver $\text{Ag}/\text{Cu}$
   e. cast iron $\text{Fe}/\text{C}$
   f. spring steel $\text{Fe}/\text{Cr}$
Practice Problems
In your notebook, solve the following problems.

SECTION 8.1 MOLECULAR COMPOUNDS
1. Classify each of the following as an atom or a molecule.
   a. Be atom
   b. CO₂ mol.
   c. N₂ mol.
   d. H₂O mol.
   e. Ne atom
2. Which of the following are diatomic molecules?
   a. CO₂
   b. N₂
   c. O₂
   d. H₂O
3. What types of elements tend to combine to form molecular compounds? Nonmetals
4. What information does a molecule's molecular structure give?
5. How do ionic compounds and molecular compounds differ in their relative melting and boiling points?

SECTION 8.2 THE NATURE OF COVALENT BONDING
1. Draw the electron dot structure for hydrogen fluoride, HF.
2. Draw the electron dot structure for phosphorus trifluoride, PF₃.
3. Draw the electron dot structure for nitrogen trichloride, NCl₃.
4. Draw the electron dot configuration for acetylene, C₂H₂.

SECTION 8.3 BONDING THEORIES
1. Predict the shape and bond angle for the compound carbon tetrafluoride, CF₄. Tetrahedral, 109.5°
2. Predict the shape and bond angle for phosphorus trifluoride, PF₃. Pyramidal, 107°
3. Predict the shape and bond angle for boron trichloride, BCl₃. Trigonal planar, 120°
4. Predict the shape and bond angle of fluorine monoxide, F₂O. Bent, 105°
6. Predict the shape of the CH₂Cl₂ molecule. What hybridization is involved in the carbon-carbon bonds?

7. How many sigma and pi bonds are used by each of the carbon atoms in the following compound?

\[
\begin{align*}
\text{H} & \text{O} \\
\text{H:C=Cl}_2 & \text{O-H} \\
\text{H} & \\
\end{align*}
\]

SECTION 8.4 POLAR BONDS AND MOLECULES

1. What type of bond—nonpolar covalent, polar covalent, or ionic—will form between each pair of atoms?
   a. Na and O  
   b. O and O  
   c. P and O

2. Explain why most chemical bonds would be classified as either polar covalent or ionic. Diatomic bonds will be nonpolar bonds.

3. Would you expect carbon monoxide and carbon dioxide to be polar or nonpolar molecules?
   CO = polar  
   CO₂ = nonpolar

4. Draw the structural formulas for each molecule and identify polar covalent bonds by assigning the slightly positive (δ+) and slightly negative (δ−) symbols to the appropriate atoms.
   a. NH₃

5. Which would you expect to have the higher melting point, CaO or CS₂?
   b/c it's ionic
8

INTERPRETING GRAPHICS

Use with Section 8.3

Figure 1 Common Molecular Shapes

Use what you have learned in Chapter 8 to complete the table on the following page.

Table 1 Arrangement of electron pairs about an atom

<table>
<thead>
<tr>
<th>Number of valence electron pairs about the central atom</th>
<th>Arrangement of valence-electron pairs</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>linear</td>
</tr>
<tr>
<td>3</td>
<td>trigonal planar</td>
</tr>
<tr>
<td>4</td>
<td>tetrahedral</td>
</tr>
<tr>
<td>5</td>
<td>trigonal bipyramidal</td>
</tr>
</tbody>
</table>
Table 2 Molecular Geometries

<table>
<thead>
<tr>
<th>Molecule</th>
<th>Electron Dot Structure</th>
<th>Shape</th>
<th>Bond Angle</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. CO₂</td>
<td>O = C = O</td>
<td>linear</td>
<td>180°</td>
</tr>
<tr>
<td>2. CH₄</td>
<td>H – C - H</td>
<td>tetrahedral</td>
<td>109.5°</td>
</tr>
<tr>
<td>3. SO₃</td>
<td>O=S=O</td>
<td>trigonal planar</td>
<td>120°</td>
</tr>
<tr>
<td>4. BeF₂</td>
<td>F – Be – F</td>
<td>linear</td>
<td>180°</td>
</tr>
<tr>
<td>5. PF₃</td>
<td>F – P – F</td>
<td>trig. pyramidal</td>
<td>107°</td>
</tr>
<tr>
<td>6. PCl₅</td>
<td>Cl – Cl – P – Cl – Cl</td>
<td>trig. bipyramidal</td>
<td>90°/120°</td>
</tr>
<tr>
<td>7. H₂O</td>
<td>H – O – H</td>
<td>bent</td>
<td>105°</td>
</tr>
</tbody>
</table>

8. If you have access to a molecular model set, construct three-dimensional models of each of the molecules in the table. Compare your models to the shapes shown in Figure 1. With a protractor, measure all the bond angles in your models. Compare these angles to those predicted by VSEPR theory and label each of the illustrations in Figure 1 with the correct bond angles.
CHEMICAL NAMES AND FORMULAS

Practice Problems
In your notebook, solve the following problems.

SECTION 9.1 NAMING IONS

1. What is the charge on the ion typically formed by each element?
   a. oxygen $\text{O}^2-$
   b. iodine $\text{I}^-$
   c. sodium $\text{Na}^+$
   d. aluminum $\text{Al}^3+$
   e. nickel, 2 electrons lost $\text{Ni}^{+2}$
   f. magnesium $\text{Mg}^{+2}$

2. How many electrons does the neutral atom gain or lose when each ion forms?
   a. $\text{Cr}^{3+}$
   b. $\text{P}^{3-}$
   c. $\text{Li}^+$
   d. $\text{Ca}^{2+}$
   e. $\text{Cl}^-$
   f. $\text{O}^{2-}$

3. Name each ion. Identify each as a cation or an anion.
   a. $\text{Br}^-$, bromide
   b. $\text{K}^+$, potassium ion
   c. $\text{H}^-$, hydride
   d. $\text{K}^+$, potassium ion

4. Write the formula (including charge) for each ion. Use Table 9.3 if necessary.
   a. carbonate ion $\text{CO}_3^{2-}$
   b. nitrite ion $\text{NO}_2^-$
   c. sulfate ion $\text{SO}_4^{2-}$
   d. hydroxide ion $\text{OH}^-$
   e. ammonium ion $\text{NH}_4^+$
   f. calcium ion $\text{Ca}^{2+}$

5. Name the following ions. Identify each as a cation or an anion.
   a. $\text{CN}^-$, cyanide
   b. $\text{HCO}_3^-$, hydrogen carbonate
   c. $\text{PO}_4^{3-}$, phosphate
   d. $\text{Cl}^-$, chloride

SECTION 9.2 NAMING AND WRITING FORMULAS FOR IONIC COMPOUNDS

1. Write the formulas for these binary ionic compounds.
   a. magnesium oxide $\text{MgO}$
   b. aluminum chloride $\text{AlCl}_3$
   c. potassium iodide $\text{KI}$
   d. sodium sulfide $\text{Na}_2\text{S}$
   e. sodium sulfide $\text{Na}_2\text{S}$

2. Write the formulas for the compounds formed from these pairs of ions.
   a. $\text{Ba}^{2+}, \text{Cl}^-$, $\text{BaCl}_2$
   b. $\text{Ag}^+, \text{I}^-$, $\text{AgI}$
   c. $\text{Ca}^{2+}, \text{S}^{2-}$, $\text{CaS}$
   d. $\text{K}^+, \text{Br}^-$, $\text{KBr}$
   e. $\text{Al}^{3+}, \text{O}^{2-}$, $\text{Al}_2\text{O}_3$
   f. $\text{Fe}^{2+}, \text{O}^{2-}$, $\text{FeO}$

3. Name the following binary ionic compounds.
   a. $\text{CaCl}_2$, calcium chloride
   b. $\text{Li}_2\text{N}$, lithium nitride
   c. $\text{CaCl}_2$, calcium chloride
   d. $\text{SrBr}_2$, strontium bromide
   e. $\text{NiCl}_2$, nickel chloride
   f. $\text{K}_2\text{S}$, potassium sulfide
   g. $\text{CuCl}_2$, copper chloride
   h. $\text{SnCl}_4$, tin(II) chloride
4. Write formulas for the following ionic compounds.
   a. sodium phosphate  \( \text{Na}_3\text{PO}_4 \)
   b. magnesium sulfate  \( \text{MgSO}_4 \)
   c. sodium hydroxide  \( \text{NaOH} \)
   d. potassium cyanide  \( \text{KCN} \)
   e. ammonium chloride  \( \text{NH}_4\text{Cl} \)
   f. potassium dichromate  \( \text{K}_2\text{Cr}_2\text{O}_7 \)

5. Write formulas for compounds formed from these pairs of ions.
   a. \( \text{NH}_4^+ , \text{SO}_4^{2-} \)
   b. \( \text{K}^+, \text{NO}_3^- \)
   c. barium ion and hydroxide ion  \( \text{Ba}^{2+} , \text{OH}^{-} \)
   d. lithium ion and carbonate ion  \( \text{Li}^+ , \text{CO}_3^{2-} \)

6. Name the following compounds.
   a. \( \text{NaCN} \) sodium cyanide
   b. \( \text{K}_2\text{CO}_3 \) potassium carbonate
   c. \( \text{Na}_2\text{SO}_4 \) sodium sulfate
   d. \( \text{LiNO}_3 \) lithium nitrate

7. Name and give the charge of the metal cation in each of the following
   ionic compounds.
   a. \( \text{Na}_3\text{PO}_4 \) sodium phosphate
   b. \( \text{Na}_2\text{PO}_4 \)
   c. \( \text{CaS} \) calcium sulfide
   d. \( \text{K}_2\text{S} \) potassium sulfide

SECTION 9.3 NAMING AND WRITING FORMULAS FOR MOLECULAR COMPOUNDS

1. Name the following molecular compounds.
   a. \( \text{PCl}_5 \) phosphorus pentachloride
   b. \( \text{CCl}_4 \) methane tetrachloride
   c. \( \text{NO}_2 \) nitrogen dioxide
   d. \( \text{N}_2\text{F}_2 \) dihydrogen difluoride
   e. \( \text{P}_4\text{O}_6 \) phosphorus hexoxide
   f. \( \text{XeF}_2 \) xenon difluoride
   g. \( \text{SiO}_2 \) silicon dioxide
   h. \( \text{Cl}_2\text{O}_7 \) dichloride heptoxide

2. Write the formulas for the following binary molecular compounds.
   a. nitrogen tribromide  \( \text{NBr}_3 \)
   b. dichlorine monoxide  \( \text{Cl}_2\text{O} \)
   c. sulfur dioxide  \( \text{SO}_2 \)
   d. dinitrogen tetrafluoride  \( \text{N}_2\text{F}_4 \)