<u>Arrangement of Electrons in Atoms</u> <u>Refinements of the Atomic Model</u>

- I. Wave-Particle Nature of Light
 - A. Electromagnetic Radiation
 - 1. Form of energy that exhibits wavelike behavior as it travels through space.
 - 2. Wavelength (λ) -- Distance between corresponding points on adjacent waves
 - 3. Frequency(v) -- Number of waves that pass a given point in a specific amount of time, usually one second.
 - 4. Formula:

 $c = \upsilon \lambda$

 $c = speed of light (3 \times 10^8 m/s)$

- $\lambda =$ wavelength (m)
- υ = frequency (Hertz (Hz) or /s)
- 5. Continuous spectrum -- Spectrum in which all wavelengths within a given range.
- 6. Electromagnetic spectrum -- Consists of all electromagnetic radiation, arranged according to increasing wavelength
- B. Light as particles
 - 1. Photoelectric effect -- Emission of electrons by certain metals when light shines on them.
 - 2. Quantum -- Finite quantity of energy that can be gained or lost by an atom
 - 3. Photon -- Quantum of light
 - 4. Formula:

E = hv

E = Energy of photon

- h = Planck's Constant (6.626 x 10^{-34} Js)
- υ = Frequency (Hertz (Hz) or /s)
- 5. Ground state -- State of lowest energy of an atom
- 6. Excited state
 - a. State in which it has a higher potential energy than it has in its ground state
 - b. Objects may give off light
- II. Bohr Model of the Hydrogen Atom
 - A. Line spectrum
 - 1. Produced when an electron drops from a higher- energy orbit to a lower energy orbit
 - 2. Photon is emitted when an electron drops from an excited state to a ground state
 - B. Believed all atoms were made up of spheres within spheres
 - C. Most atoms are not spherical

III. Spectroscopy

- A. Spectroscope -- Instrument that separates light into a spectrum that can be examined
- B. Origin of spectral lines
 - 1. Light rays of different frequencies are bent different amounts when passed through a prism
 - 2. Blue light (short wavelength, high frequency) is bent the most
 - 3. Red light (long wavelength, low frequency) is bent the least
 - 4. When substances are heated, the electrons are raised to higher energy levels
 - 5. Electrons fall back to ground state
 - 6. Electrons emit a wavelength
- IV. Quantum Model of the Atom
 - A. Quantum theory -- describes mathematically the wave properties of electrons and other very small particles
 - B. Oribital -- 3-D region about the nucleus in which a particular electron can be located

Quantum Numbers and Atomic Orbitals

- I. Quantum Numbers
 - A. Numbers that specify the properties of atomic orbitals and of their electrons
 - B. Indicate:
 - 1. distance from the nucleus
 - 2. orbital shape
 - 3. orbital position
 - C. Principal Quantum Number
 - 1. Symbolized by n
 - 2. Indicates main energy levels surrounding a nucleus
 - 3. Referred as shells
 - 4. Ranges from 1-7
 - D. Orbital Quantum Number
 - 1. Indicates shape of an orbital
 - 2. Referred as sublevels or subshells
 - 3. Ranges from s, p, d, f
 - 4. s spherical
 - 5. p two lobes
 - 6. d four lobes, one with two lobes with donut
 - 7. f very complex
 - E. Magnetic Quantum Number
 - 1. Indicates the orientation of an orbital about the nucleus
 - 2. s-orbital -- one orientation
 - 3. p-orbital -- three orientations (x, y, z)
 - 4. d-orbital -- five orientations
 - 5. f-orbital -- seven orientations
 - F. Spin Quantum Number
 - 1. Indicates two possible states of an electron in an orbital
 - 2. +1/2 and -1/2 (opposite spins)

- G. Magnetism
 - 1. Diamagnetism -- weakly repels any magnetic field, but only when another magnetic field is in its presence.
 - 2. Paramagnetism -- weak attraction to a magnetic field, but only when another magnetic field is in its presence.
 - 3. Ferromagnetism -- strong form of paramagnetism (bar magnet), but also maintains a constant magnetic field.

Electron Configurations

- I. Rules Governing Electron Configurations
 - A. Electron Configuration -- Arrangement of electrons in atoms
 - B. Aufbau Principle
 - 1. Electron occupies the lowest-energy orbital that can receive it.
 - 2. 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p...
 - C. Hund's Rule
 - 1. Orbitals of equal energy are each occupied by one electron before any is occupied by a second electron
 - 2. p_x, p_y, p_z
 - D. Pauli Exclusion Principle
 - 1. No two electrons in the same atom can have the same four quantum numbers
 - 2. two electrons of opposite spin can occupy the same orbital
- **II.** Representing Electron Configurations
 - A. Orbital Notation
 - 1. Unoccupied orbital is represented by a line ()
 - 2. Orbital containing one electron is represented by an arrow (
 - 3. Orbital containing two electrons is represented by two arrows in opposite directions (\uparrow \downarrow)
 - 4. Lines are labeled with principal quantum number and subshell letter $(\uparrow \downarrow)$
 - B. Electron-Configuration Notation
 - 1. Lines and arrows are not used
 - 2. Number of electrons in a sublevel is shown by adding superscripts to the sublevel designations

 - 3. $1s^{1} =$ one electron in the 1s orbital 4. $1s^{2} =$ two electrons in the 1s orbital
 - C. Electron-Dot Notation
 - 1. Shows only electrons in the highest, or outermost, main energy level
 - 2. Highest Occupied Energy Level -- Electron-containing main energy level with the highest principal quantum number.
 - 3. Dots representing number of electrons in the highest occupied energy levels are placed around the symbol of the element
 - 4. Inner-Shell Electrons -- Electrons not in the highest occupied energy level

5.

Number of electrons	Electron-dot notation
1	Х
2	Х
3	Х
4	Х
5	Х
6	Х
7	Х
8	Х

- III. Elements of the Second and Third Periods
 - A. Second-period Elements
 - 1. After 1s is filled, 2s is then filled
 - 2. After 2s is filled, 2p is then filled
 - a. Each p orbital is filled with one electron
 - b. After last p orbital is filled, the p orbitals receive a second electron, but with opposite spins
 - 3. Octet s and p sublevels of their highest main energy level filled eight electrons

B. Third-period Elements

- 1. After the first octet is filled, the third main energy level is then filled.
- 2. After an octet is filled, square brackets are placed around the symbol

Ne =
$$1s^{2} 2s^{2} 2p^{6} = [Ne]$$

Na = $1s^{2} 2s^{2} 2p^{6} 3s^{1} = [Ne] 3s^{1}$

- 3. Noble gas
 - a. Group 18 elements (He, Ne, Ar, Kr, Xe, Rn)
 - b. Contains an octet in its outer shell (except He)
- 4. Noble gas configuration Outer main energy level fully occupied by eight electrons

- IV. Elements of the Fourth and Fifth Periods
 - A. Cr and Cu seems to violate the normal electron configuration
 - B. 3d orbital is filled
 - C. d orbital contains 5 orbitals in its sublevel
 - D. d orbital can hold 10 electrons
- V. Elements of the Sixth and Seventh Periods
 - A. 4f orbital is filled
 - B. f orbital contains 7 orbitals in its sublevel
 - C. f orbital can hold 14 electrons

VI. Diagram ("Electron Convention"):

Rules that electrons follow (with a few exceptions, such as Cu and Cr)

- 1. Electrons are cheap.
- 2. They want to be as close to the convention (nucleus) as possible.
- 3. They prefer a room to themselves. (Each room can only occupy 2 electrons.)
- 4. If they must room with someone (in order to save money), they will room with one other, but they must be of an opposite gender (opposite spin).



VII. Diagram (Order of filling):

Orbital	Shape	Total # of orientations (spaces)	Total # electrons (occupants)	1 st Shell	2 nd Shell	3 rd Shell	4 th Shell	5 th Shell	6 th Shell	7 th Shell
	0.1.1	_	14			~		~~~~~~~~~~~~~~~~~~~~~~~~~~~~~~~~~~~~~~~		
t	8-lobes, 2-lobes with 2- rings		14		$\sum_{i=1}^{n}$		41	5t	6t	
d	4-lobes, 2-lobes with ring	5	10	\bigcirc		3d	4d	5d	6d	7d
р	2-Lobes	3	6		2p	3p	4p	5p	6p	7p
S	Spherical	1	2	ls	2s	3s	4s	5s	6s	7s

Order in which electron configurations are written

1s	2s 2p	3s 3p 3d	4s 4p 4d 4f	5s 5p 5d 5f	6s 6p 6d 6f	7s 7p 7d 7f
1 st	2 nd	3 rd	4 th	5 th	6 th	7 th
Shell						

IX. Examples

1. Write both the complete and shorthand electron configurations, as well as the electron-dot symbols, for each of the following:

a. Na

b. Al

- c. P
- d. Ar
- e. Br

f. Sr

- 2. Identify each atom on the basis of its electron configuration:
- a. $1s^2 2s^2 2p^1$
- b. $1s^2 2s^2 2p^5$
- c. [Ne] 3s²
- d. [Ne] $3s^2 3p^2$
- e. [Ne] $3s^2 3p^5$
- f. [Ar] $4s^{1}$
- g. [Ar] $3d^{6} 4s^{2}$

3. If the speed of light is 3×10^8 m/s, calculate the wavelength of light whose frequency is 7.5×10^2 Hz

- 4. Determine the energy, in joules, of a photon whose frequency is 3.55×10^{17} Hz.
- 5. What is the frequency of a radio wave with an energy of 1.55×10^{-24} J?

<u>Periodic Law</u> <u>Electron Configuration and the Periodic Table</u>

- I. Groups, Periods, and the Blocks of the Periodic Table
 - A. Inner-shell electrons have little influence on properties of an element because they do not come into contact with their surroundings
 - B. Outermost electrons play the largest role in determining chemical properties
 - C. Each period is determined by the sublevels being filled with electrons
 - 1. First energy level holds two electrons in its 1s sublevel
 - 2. Second energy level holds eight electrons total in the 2s and 2p
 - D. Period of an element can be determined from its electron configuration
 - E. 4 blocks of periodic table
 - 1. s-block
 - 2. p-block
 - 3. d-block
 - 4. f-block
 - 5. Name of each block is based on whether and s, p, d, or f sublevel is being filled in successive elements of that block.

Electron Configuration and Periodic Properties

I. Atomic Radii

A. One-half the distance between the nuclei of identical atoms joined in a molecule

- B. Period Trends
 - 1. Gradual decrease in atomic radii across the periods
 - 2. Caused by increasing positive charge of the nucleus
 - 3. Electrons are pulled closer to the nucleus
 - 4. Can be off set by large number of electrons in the same outer energy level
- C. Group Trends
 - 1. Atomic radii increases going down the column
 - 2. New outer shells are added going down the column

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- II. Ionization Energy
 - A. A + energy \rightarrow A + e
 - A = atom
 - A^{+} = atom with positive charge
 - B. Ion -- Atom or group of atoms that has a positive or negative charge
 - C. Ionization -- Process that results in the formation of an ion
 - D. Ionization energy (First ionization energy) -- Energy required to remove one electron from an atom of an element
 - E. Period Trends
 - 1. Group 1 elements (Alkali metals)
 - a. Lowest first ionization energies
 - b. Loose electrons most easily
 - c. Main reason for high reactivity
 - 2. Group 18 elements (Noble gases)
 - a. Very high first ionization energies
 - b. Does not lose electrons easily
 - c. Main reason for low reactivity
 - 3. Going across the period (left to right)
 - a. Ionization energy increases
 - b. Decreasing radii
 - 4. Nonmetals have higher ionization energies than metals
 - F. Group Trends
 - 1. Ionization energy decreases down the groups
 - 2. Electrons are further away from the nucleus
 - 3. Inner-shell electrons shield outer-shell electrons from the positive charge of the nucleus
 - G. Formation of positive ions
 - 1. "Second ionization energy" -- removal of the second electron (requires more energy than the first ionization energy).
 - 2. "Third ionization energy" -- removal of the third electron (requires even more energy than the second ionization energy).
 - 3. Each successive electron must be removed from an ion with a larger positive charge
 - 4. Large jumps in ionization energy occurs when an element has a noble gas configuration

III. Electron Affinity

- A. Energy change that occurs when an electron is acquired by a neutral atom
- B. Exothermic
 - 1. $A + e \rightarrow A + energy$
 - 2. Expressed as a negative number
- C. Endothermic
 - 1. $A + e^{-} + energy \rightarrow A^{-}$
 - 2. Expressed as a positive number
 - 3. Usually unstable and loses its electron spontaneously

- D. Period Trends
 - 1. Halogens (Group 17)
 - a. Gain electrons most easily
 - b. Large negative numbers
 - 2. Electron affinities generally become more negative across each period.
 - 3. Group 2 and 18
 - a. Have positive values
 - b. Must be forced to gain electrons
- E. Group Trends
 - 1. Electron affinity generally becomes more positive going down a column
 - 2. More difficult to add electrons to larger atoms
- F. Formation of negative ions
 - 1. Second Electron Affinity -- Addition of second electron to a negative ion
 - 2. More difficult to add second electron
 - 3. All values are positive
- IV. Ionic Radii
 - A. One-half the diameter of an ion in a chemical compound
 - B. Cation
 - 1. Positive ion
 - 2. Lose of one or more electrons
 - 3. Electron cloud becomes smaller
 - 4. Remaining electrons are drawn closer to the nucleus
 - C. Anion
 - 1. Negative ion
 - 2. Addition of one or more electrons
 - 3. Electron cloud becomes larger
 - a. Smaller attractive force to the nucleus
 - b. Greater repulsion from the electrons
 - D. Periodic Trends
 - 1. Metals form cations
 - 2. Nonmetals form anions
 - 3. Regular decrease of cation radii with increasing atomic number
 - a. Positive charge of nucleus increases (atomic number)
 - b. Negative charge of electron clouds decreases
 - E. Group Trends
 - 1. Increase of ionic radii down the columns
 - 2. Outer electrons are at higher levels
- V. Valence Electrons
 - A. Electrons available to be lost, gained, or shared in the formation of chemical compounds
 - B. Electrons in the highest energy levels
 - C. Electrons most influenced by nearby atoms or ions
 - D. Often located in incompletely filled main-energy levels (outer electrons)

- VI. Electronegativity
 - A. Measure of the power of an atom in a chemical compound to attract electrons
 - B. Ranges from 0 4
 - C. Fluorine is the most electronegative
 - D. Cs, Fr have the smallest electronegativity
 - E. Periodic Trends
 - 1. Across each period, electronegativity increases
 - 2. Nonmetals tend to be more electronegative
 - 3. Halogens are the most electronegative
 - F. Group Trends
 - 1. Electronegativity decreases down each group
 - 2. Larger atoms attract electrons less easily than small atoms
- VII. Periodic Properties of the d- and f-block Elements
 - A. d-block elements vary less and with less regularity
 - B. d-block elements contain 0 2 electrons in the highest occupied energy level
 - C. Electrons in d-sublevel are responsible for many characteristic properties
 - D. Atomic radii
 - 1. Generally decrease across the periods
 - 2. Radii varies because of electron repulsion
 - E. Ionization Energy
 - 1. Generally increase across the periods
 - 2. Some elements are harder to remove their electrons
 - F. Ion formation and ionic radii
 - 1. Order in which electrons are removed from all atoms is exactly the reverse of the order given by the electron-configuration notation
 - 2. First electron to be removed for the d-block elements are those in the highest occupied s-sublevels
 - 3. Group 3 elements form ions with +3 charge
 - 4. Most d-block elements commonly form +2 ions
 - 5. Some d-block elements may have more than one charge (ie. Cu forms +1 and +2)
 - 6. Cations have smaller radii than the atoms
 - G. Electronegativity
 - 1. Electronegativity increases across the periods as radii decrease
 - 2. Small atoms attract electrons more strongly than large atoms

VIII. Examples:

1. Without looking at the periodic table give the group, period, and block in which the element with the following shorthand electron configuration is located: $[Kr] 5s^{1}$

Group _____

Period _____

Block _____

2. Without looking at the periodic table, write both the group and complete electron configurations for the element in the fourth period in Group 2.

Group _____

Electron configuration _____

3. Without looking at the Periodic Table, identify the period, block, and group in which the element with the following electron configuration is located: [Ar] $3d^8 4s^2$

Period _____

Block

Group _____

4. Without looking at the Periodic Table write the outer electron configuration for the fifth period element in Group 12

a. Electron configuration

b. Refer to the table to identify that element and to write its shorthand electron configuration.

Element _____

Shorthand Electron configuration

5. Without looking at the Periodic Table, write the outer electron configuration for the third period element in Group 17

a. Outer Electron Configuration

b. Name the element and identify as a metal, nonmetal, or metalloid.

Element _____

Metal, Nonmetal, Metalloid

6.a.	Without looking at the Periodic Table, identify the period, block, and group of an element
with	the electron configuration: $[Ar]3d^{10}4s^24p^3$

Period _____

Block _____

Group _____

b. Name this element and identify it as a metal, nonmetal, or metalloid.

Element

Metal, Nonmetal, Metalloid

7. Identify the block, period, group, group name (where appropriate), element name, element type (metal, nonmetal, or metalloid), and relative reactivity (high or low) for the element with the electron configuration of: $[He]2s^2 2p^5$

Period _____

Group _____

Group name	
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Element	
---------	--

Element type	
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Reactivity	
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8. Identify the block, period, group, group name (where appropriate), element name, element type (metal, nonmetal, or metalloid), and relative reactivity (high or low) for the element with the electron configuration of: $[Ar]3d^{10}4s^{1}$

Block	
Period	-
Group	-
Group name	
Element	
Element type	
Reactivity	

9. Identify the block, period, group, group name (where appropriate), element name, element type (metal, nonmetal, or metalloid), and relative reactivity (high or low) for the element with the electron configuration of: $[Kr]5s^{1}$

Block	
Period	
Group	
Group name	
Element	
Element type	
Reactivity	

10. Without looking at the Periodic Table identify the period, block, and group of the element: $[Ar] 3d^{7} 4s^{2}$

Period _____

Group	
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<u>Chemical Bonding</u> <u>Introduction to Chemical Bonding</u>

I. Types of Chemical Bonds

- A. Chemical Bond
 - 1. Link between atoms that result from the mutual attraction of their nuclei for electrons.
 - 2. Held together by electrostatic attraction between positively charged nuclei and negatively charged electrons
- B. Types
 - 1. Ionic Bonds
 - 2. Covalent Bonds
 - a. Nonpolar-covalent bond
 - b. Polar-covalent bond
 - 3. Metallic Bonds
 - 4. Intermolecular Forces
 - a. Dipole-dipole forces
 - b. Hydrogen bonding
 - c. London dispersion forces
- C. Ionic Bond
 - 1. Chemical bond resulting from electrostatic attraction between positive and negative ions
 - 2. One atom has given up one or more electrons, and another atom has gained them
- D. Covalent Bond
 - 1. Chemical bond resulting from the sharing of electrons between two atoms
 - 2. Represented by a pair of electron dots
- E. Bonds are never completely ionic and rarely completely covalent
- F. Bonds can range between the two extremes
- G. Degree of ionic or covalent bonding is determined by comparing electronegativities of the bonded atoms.
- H. The greater the difference in electronegativity
 - 1. The more ionic the bond
 - 2. Electrons spend more time close to the bonded atom that attracts them more strongly.
- I. Electronegativity of one atom is subtracted from the electronegativity of the other atom.
- J. Range
 - 1. 4.0 through 1.7 = Ionic bond (100% 50% Ionic character)
 - 2. 1.7 through 0.3 = Polar covalent bond (50% 5% Ionic character)
 - 3. 0.3 through 0.0 = Nonpolar covalent bond (5% 0% Ionic character)
- K. Bond between identical atoms is completely covalent

- L. Nonpolar-covalent bond
 - 1. Covalent bond in which the bonding electrons are shared equally by the bonded atoms, with a resulting balanced distribution of electrical charge
 - 2. 0% 5% Ionic character
 - 3. Electronegativity difference of 0 to 0.3
 - 4. Contains very little to no polarity
 - 5. Equal sharing of electrons between two atoms
- M. Polar -- Uneven distribution of charge
- N. Polar-covalent bond
 - 1. Covalent bond in which the united atoms have an unequal attraction for the shared electrons
 - 2. 5% 50% Ionic character
 - 3. Electronegativity difference of 0.3 to 1.7
 - 4. Contains some polarity
 - 5. Unequal sharing of electrons between atoms
 - 6. Electrons spend more time around the more electronegative element
 - 7. Atom with negative charge is symbolized with δ
 - 8. Atom with positive charge is symbolized with δ^{+}
- O. Metallic Bonds
 - 1. Metals give up their electrons
 - 2. Electrons move about freely to other same metal atoms
- P. Atoms of nonmetals form covalent bonds with each other
- Q. Atoms of metals form metallic bonds with each other
- R. Atoms of metals form ionic bonds with atoms of nonmetals
- S. Polar-covalent bonds form between metals and nonmetals that do not differ greatly in electronegativity
- II. Why Chemical Bonding Occurs
 - A. Most atoms have lower potential energy in bonds than as independent atoms
 - 1. Main reason why most elements are found to be chemically combined to other atoms
 - 2. Rare for atoms to occur uncombined
 - B. Chemical-bond formation is often an energy-releasing process
 - C. Breaking chemical bonds is often an energy-absorbing process
 - D. Atoms separated by breaking a chemical bond have a higher total potential energy than when they are bonded

Covalent Bonding and Molecular Compounds

I. Molecules

- A. Group of two or more atoms held together by covalent bonds and able to exist independently
- B. Diatomic molecule
 - 1. Molecule containing two atoms
 - 2. ex: Hydrogen hydrogen; hydrogen chlorine
- C. Chemical formula
 - 1. Shorthand representation of the composition of a substance using atomic symbols and numerical subscripts
 - 2. ex: Chemical formula for hydrogen is H₂

Chemical formula for hydrogen chloride is HCl Chemical formula for water is H_OO

- 3. Molecular formulas -- Shows the types and numbers of atoms combined in a single molecule
- II. Formation of Covalent Bonds
 - A. Repulsion between like charges (protons-protons, electrons-electrons)
 - B. Attraction between unlike charges (protons-electrons)
 - 1. Stronger than repulsion force
 - 2. If two atoms become too close, strong repulsion occurs
 - C. Bond length
 - 1. Average distance between two bonded atoms
 - 2. Distance of minimum potential energy
 - D. Bond energy
 - 1. Energy required to break a chemical bond and form neutral atoms
 - 2. Measured in KJ/mole
 - E. Bond energies become larger as bond lengths become shorter
 - F. Electrons spend more time in the space between the two nuclei
 - 1. Region of increased electron density
 - 2. Attraction of the nucleus to the high electron density is what holds the two atoms together
 - G. Sharing of electrons allows the atoms to achieve the stable, noble gas configuration.
 - H. All halogens and hydrogen form naturally with another same atom (diatomic molecule)
 - 1. ex: H_2 , F_2 , Cl_2 , Br_2 , I_2 , At_2
 - 2. Boxes are drawn around the bonding electron pair to represent the sharing of those electrons



III. Octet Rule

- A. Each noble gas atom has eight electrons in its highest occupied energy level (exept helium)
- B. Electrons are shared to achieve a noble gas configuration with other elements



C. Octet Rule

- 1. Chemical compounds tend to form so that each atom, by gaining, losing, or sharing electrons, has an octet of electrons in its highest occupied energy level
- 2. Valence electrons are in the outermost s and p orbitals
- 3. ex: Nitrogen
 - a. Group 15
 - b. Outer electron configuration $2s^2 2p^3$
 - c. 5 valence electrons
 - d. Needs three more electrons to complete an octet
 - e. Forms 3 covalent bonds in many molecules





- 4. Exceptions to Octet Rule (p. 449)
 - a. Some elements of the third and higher periods of the periodic table can be surrounded by more than four pairs of valence electrons when combined with atoms of the highly electronegative elements (F, O, Cl)
 - b. Bonding may take place in the d-orbital (along with the s and p orbitals)
 - c. ex: PCl_5 , SF_6 , XeF_4
 - d. Octet rule is also violated when the total number of valence electrons in a molecule is odd
 - e. ex: NO

Nitrogen has 5 valence electrons Oxygen has 6 valence electrons Total of eleven

IV. Lewis Structure

- A. Electron dot symbols are used to represent molecules
- B. Shared pair electrons appear between the two atomic symbols
- C. Unshared pair (lone pair) electrons
 - 1. Pair of electrons that is not involved in bonding, but instead belongs exclusively to one atom
 - 2. Appear around atomic symbol, but not between atomic symbols
 - 3. ex:

```
Н:Н
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: Cl : Cl :
...
H : Cl :
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- 1. Formulas in which atomic symbols represent nuclei and inner-shell electrons, dot-pairs or dashes between two atomic symbols represent electron pairs in covalent bonds, and dots adjacent to only one atomic symbol represent unshared electrons
- 2. Dots representing covalent bond is replaced by a dash
- 3. ex:

- E. Structural Formula
 - 1. Indicates the kind, number, arrangement, and bonds of the atoms in a molecule 2. ex:

- a. Shows that the molecule (ethane) contains 2 carbons and 6 hydrogens.
- b. Shows the 2 carbons are bonded to each other.
- c. Shows 3 hydrogens are bonded to each carbon.
- F. Single Bond Covalent bond produced by the sharing of one pair of electrons between two atoms
- G. Steps for writing Lewis structures
 - 1. Determine the type and number of atoms in the molecule
 - 2. Draw the electron-dot symbols for each type of atom
 - 3. Determine the total number of valence electrons in the atoms to be combined
 - 4. Arrange the atoms to form a skeleton structure for the molecule and connect the atoms by electron-pair bonds
 - 5. Add unshared pairs of electrons so that each nonmetal atom (except hydrogen) is surrounded by eight electrons (hydrogen has 2 electrons)
 - 6. Count the electrons in the structure to be sure that the number of valence electrons used equals the number available

V. Multiple Covalent Bonds

- A. Double Bond
 - 1. Covalent bond between two atoms produced by the sharing of two pairs of electrons between two atoms
 - 2. Two atoms that share two electron pairs
 - 3. Shown as two side-by-side pairs of dots
 - 4. ex.

Н Н		Н Н
•• ••		
C : : C	or	C :: C
		1 1
нн		 H H
		11 11

- 5. May also be show as two parallel dashes
- 6. ex.

$$\begin{array}{c|c} H & H \\ | & | \\ C = C \\ | & | \\ H & H \end{array}$$

7. All 4 electrons in a double bond "belong" to both atoms

B. Triple Bond

- 1. Covalent bond produced by the sharing of three pairs of electrons between two atoms
- 2. ex.

N	<u>ls</u>	<u>2s</u>	<u>2p</u>
N	1s	2s	<u>2p</u>

C. Double bonds tend to be shorter and stronger than single bonds

 $: N \stackrel{.}{:} N :$ or $: N \equiv N :$

- D. Triple bonds tend to be shorter and stronger than double bonds
- E. Steps
 - 1. Determine the type and number of atoms
 - 2. Draw the electron-dot symbols for each type of atom
 - 3. Determine the total number of valence electrons in the atoms to be combined
 - 4. Arrange the atoms to form a skeleton structure for the molecule and connect the atoms by electron-pair bonds
 - 5. Add unshared pairs of electrons so that each nonmetal atom (except hydrogen) is surrounded by eight electrons

- 6. Count the electrons in the Lewis structure to be sure that the number of valence electrons used equals the number available
- 7. If too many electrons have been used, subtract one or more lone pairs from existing bonds (between C, N, or O atoms) until the total number of valence electrons is correct
- VI. Polyatomic Ions
 - A. Charged group of covalently bonded atoms
 - B. Charge results from an excess of electrons (negative charge) or shortage of electrons (positive charge)

Ionic Bonding and Ionic Compounds

I. Formation of Ionic Bonds

- A. Ionic Compounds -- Composed of positive and negative ions combined so that the positive and negative charges are equal
- B. Formula Unit -- Simplest unit indicated by the formula of any compound
- C. ex.

1s

$$Na^{+} + :Cl^{+} \rightarrow Na^{+} + :Cl^{+}$$

Na

Cl

D. Lattice energy -- Energy released when one mole of an ionic crystalline compound is formed from gaseous ions

3s

3p

II. Comparison of Ionic and Molecular Compounds

2s

2p

- A. Covalent bonds
 - 1. Bonds are strong, but not as strong between atoms with ionic bonding
 - 2. Tend to have lower melting or boiling points
- B. Ionic bonds
 - 1. Strong attraction between positive and negative charges of its atoms
 - 2. Usually has higher melting and boiling points
 - 3. Does not vaporize easily
 - 4. Tend to be hard and brittle
 - 5. Many are soluble in water
 - 6. Molten state, or dissolved in water, ionic compounds can conduct an electric current

Metallic Bonding

- I. Formation of Metallic Bonds
 - A. Metallic Bond
 - 1. Chemical bond resulting from the attraction between positive ions and surrounding mobile electrons
 - 2. Electron sea -- Electrons are free to move around from one metal ion to another metal ion
- II. Nature of Metals
 - A. Good conductors of electricity
 - 1. Free electrons through the solid
 - 2. Electrons are free to move about
 - B. Good conductor of heat
 - 1. Heated end of metal, electrons gain kinetic energy
 - 2. High kinetic electrons move through the solid transferring the high kinetic energy
 - C. Malleability -- Able to be shaped or extended by hammering, rolling, or increased pressure
 - D. Ductility -- Able to be drawn, pulled, or extruded through a small opening to produce a wire
 - 1. Metal ions can slide past one another
 - 2. Makes metals easy to shape
 - E. Metallic luster
 - 1. Able to absorb and remit light of many wavelengths
 - 2. Gives shiny appearance

The Properties of Molecular Compounds

I. VSEPR Theory

- A. "<u>Valence-shell</u>, <u>electron-pair</u> <u>r</u>epulsion"
- B. Electrostatic repulsion between the valence-level electron pairs surrounding an atom causes these pairs to be oriented as far apart as possible
- C. Geometry of molecules
- D. Determined by how many atoms and electrons surround a central atom

S.	VSEPR Geometries				
Steric No.	Basic Geometry 0 Ione pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	XEX Linear				
3	X E X Trigonal Planar	E X <120° Bent or Angular			
4	XIIIIIIIIEEE XXIIIIIIIIEEEEEEEEEEEEEEEE	X/IIIIE X X < 109° Trigonal Pyramid	X Sent or Angular		
5	X 120° X X X X X X X X X X X X X	< 90° X X < 120° E X X Sawhorse or Seesaw	×	Linear	
6	$\begin{array}{c} X & g_{0} \circ \\ X / \dots & & \\ X / \dots & & \\ X / & \\ X \\ Octahedral \end{array}$	Square Pyramid	90° Kanta Square Planar	X L L X X X X X X X X X X X X X	X 180°

II. Hybridization

- A. Mixing of two or more atomic orbitals of similar energies on the same atom to give new orbitals of equal energies
- B. S and p orbitals are combined into one orbital called sp



III. Intermolecular Forces

- A. Forces of attraction between molecules
- B. Weaker than ionic and metallic bonding
- C. Dipole dipole forces and molecular polarity
 - 1. Dipole -- Equal but opposite charges separated by a short distance
 - 2. Represented by an arrow pointing toward the negative pole with a tail crossed to indicate that it lies in the direction of the positive pole
 - 3. ex.

$$H - Cl$$

D. Dipole-Dipole Forces

- 1. Forces of attraction between polar molecules
- 2. Negative region in one polar molecule attracts the positive region in adjacent molecules
- E. Hydrogen Bonding
 - 1. Intermolecular attraction between a hydrogen atom bonded to a strongly electronegative atom (ie: F, O, N) and an unshared pair of electrons on a strongly electronegative atom (ie: F, O, N)
 - 2. Usually represented by dotted lines connecting the hydrogen-bonded hydrogen to the unshared electron pair of the electronegative to which it is attracted
 - 3. Accounts for surface tension in water and shapes of ice crystals (snow flakes)
- F. London Dispersion Forces (similar to Van der Waal forces)
 - 1. Intermolecular attractions resulting from the constant motion of electrons and the creation of instantaneous dipoles and induced dipoles
 - 2. Stronger at lower temperatures
 - 3. Caused by uneven distribution of electrons in an atom or molecule
 - a. Creates regions of positive and negative dipoles
 - b. May induce dipoles in nearby atoms or molecules
 - 4. Strength increases as the number of electrons in an element increases
 - 5. Only bonding force between noble gases

IV	Examples:		
1.	Draw the electron configurations	for	the following:
a.	Na	b.	Sn
c.	0	d.	S
2.	Draw the electron configurations	for	the following:
a.	Na ⁺¹	b.	S ⁻²
c.	Al ⁺³	d.	P-3
3.	Write the short-hand configuration	n f	or the following:
a.	Cs	b.	Pb
c.	Au	d.	Ne
4.	The following covalent molecules	ha	ve only single covalent bonds. Draw an electron dot
str	ucture for each.		
a.	H ₂ O	b.	H ₂ O ₂
c.	PCl ₃	d.	NH ₃
5.	Classify the following compounds	s as	ionic or covalent.
a.	MgCl	b.	HCl
c.	Na ₂ S	d.	H ₂ S
e.	CaO	f.	H ₂ O
			-

6. Draw the electron dot structures of these molecules and also draw the structures using the stick figures.

a.	F ₂	b. HCl
c.	НССН	d. HCN
e.	I ₂	f. OF ₂
g.	H ₂ S	h. NI ₃

7. Use VSEPR theory to predict the shapes of the following species.

- a. CO₂ b. SiCl₄
- c. SO₃ d. SCl₂
- e. CO

8. Determine what type of bond (ionic, polar covalent, or nonpolar covalent) will form between atoms of the following elements and show the polarity of the bond if it is polar covalent.

- a. Ca and Cl b. C and S
- c. Mg and F d. N and O
- e. H and O f. S and O

9. Draw the electron dot structure for each of the following molecules. Then identify polar covalent bonds by assigning δ^+ and δ^- to the appropriate atoms.

- a. HOOH b. HBr c. BrCl d. H₂O
- $\mathbf{d} \cdot \mathbf{n}_2 \mathbf{0}$

10. Based on the information about molecular shapes, which of these molecules would you expect to be polar?

a. SO_2 c. CO_2 b. H_2S d. BF_3