CHEMICAL BONDING

I. ATOMIC STRUCTURE REVIEW

A. Nucleus

1. Protons

2. Neutrons

B. Electronic Configuration - review on your own

C. Electronic Configuration & Periodic Chart

1. Atomic Number = total number of electrons

2. Period number (n) = number of occupied shells
= valence shell

3. Group (family) number: all elements belonging to the same group have the same number of electrons in the valence shell

a. "A" groups: group number = number of valence electrons (except He).

b. "B" groups: assume there are two electrons in valence shell.

II. PERIODIC TRENDS IN ATOMIC SIZE

A. Down a group: atomic size increases

B. Across a period from left to right: atomic size decreases

III. LEWIS ELECTRON DOT SYMBOLS FOR THE REPRESENTATIVE ELEMENTS (A groups)

A. Lewis dot symbols are used to represent the valence electrons of atoms

B. Rules for drawing Lewis dot symbols (for "A" group elements only):

1. Write the symbol for the element.
2. Think of the symbol as having 4 sides (like a square).
3. Distribute the electrons (dots) one at a time. It doesn't matter on which side you begin placing the dots, but do not pair electrons until you have distributed the first four dots.

C. Examples

IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIII
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar

IV. OCTET RULE: an atom is most stable when it has 8 electrons in its valence shell. Therefore, in forming compounds, atoms try to get 8 electrons (a stable octet) in their valence shell.

Two ways to get 8 electrons in the valence shell:

V. COVALENT BOND

A. In a covalent bond two NONMETAL atoms SHARE electrons.

1. One covalent bond is a pair of electrons.

2. EXCEPTIONS to the Octet Rule: In a molecule (a particle composed of nonmetal atoms held together by covalent bonds) each atom will have 8 electrons around it (in its valence shell), except for the following:

a. Those which have less than 8:

H

B

b. Those which can have more than 8: any element in the third or lower period on the periodic chart.

B. LEWIS STRUCTURES for Molecules and Polyatomic Ions

1. Choosing the central atom:

   a. Water

   b. Oxy-Acids:

   c. Organic Compounds

   d. All others:

2. Rules for Drawing Lewis Structures

   a. Count and add all the valence electrons of the atoms in the molecule or polyatomic ion. If it is a polyatomic ion, add an additional electron for each negative charge or subtract an electron for each positive charge.
   b. Divide the total number of valence electrons by two to get thenumber of electron pairs.
   c. Place one pair of electrons in each bond.
   d. Complete the octets of the atoms bonded to the central atom. (Don't forget exceptions to the octet rule.)
   e. Place any additional electrons on the central atom in pairs.
   f. If the central atom still has less than an octet, you must form multiple bonds so that each atom has an octet.

3. Examples

   H_2

   H_2O

CO_2

N_2

C_2H_2

HNO_3

H_2SO_4

CO_3^{2-}

NH_4^+

PCl_5

ICl_3

C. Coordinate Covalent Bond

D. Bond Energy: the minimum amount of energy required to separate two covalently bonded atoms.

E. Bond Length: the distance between the two nuclei of covalently bonded atoms when energy is at its minimum.

VI. IONIC BOND

A. An ionic bond is an electrostatic attraction between ions of opposite charge.

B. Ions are formed by the transfer of an electron from the valence shell of one atom to the valence shell of another.

1. Metals

2. Nonmetals

VII. ELECTRONEGATIVITY AND BONDING

A. Electronegativity is a measure of an atom's attraction for the electrons in a bond.

B. Periodic Trends in Electronegativity:

1. Down a group:

2. Across a period from left to right:

C. Using electronegativities of bonding atoms to determine bond type:

1. If the difference in electronegativities of the two bonding atoms is greater than 1.7, the bond is ionic.

2. If the difference in electronegativities of the two bonding atoms is less than 1.7, the bond is covalent.

a. If the difference in electronegativities of the two bonding atoms is zero (0) the bond is a NONPOLAR covalent bond.

b. If the difference in electronegativities of the two bonding atoms is greater than zero (0) but less than 1.7, the bond is a POLAR covalent bond.

VIII. POLAR & NONPOLAR MOLECULES

A. Polar Molecules

1. H_2O

2. HCl

3. NH_3

B. Nonpolar Molecules

1. CO_2

2. CCl_4

3. CH_4

IX. FACTORS THAT AFFECT IONIC BOND FORMATION

A. IONIZATION ENERGY (Ionization Potential)

1. Ionization energy is the energy required to remove an electron from an isolated gaseous atom in the ground state.

$$Na_{(g)} \text{ sd } Na^{+}_{(g)} + e^{-}$$

2. Endothermic process

3. Successive ionization energies (kJ/mole)

Group Number	Element	First	Second	Third	Fourth	Fifth	Sixth	Seventh
IA	H	1312						
IIA	He	2372	5250					
IA	Li	520	7298	11,815				
IIA	Be	899	1757	14,848	21,006			
IIIA	B	801	2427	3660	25,025	32,826		
IVA	C	1086	2353	4620	6222	37,829	47,276	
VA	N	1402	2857	4578	7475	9445	53,265	64,358
VIA	O	1314	3388	5300	7469	10,989	13,326	71,333
VIIA	F	1681	3374	6020	8407	11,022	15,164	17,867
VIII	Ne	2081	3952	6122	9370	12,177	15,238	19,998

4. Periodic Trend in First Ionization Energy

   a. Down a group:

   b. Across a period from left to right:

5. Size of cation versus size of neutral atom

B. ELECTRON AFFINITY

1. Electron affinity is the energy released (or absorbed) when an electron is added to an isolated gaseous atom in the ground state.

$$Cl_{(g)} + e^- \longrightarrow Cl^-_{(g)}$$

2. Exothermic process (usually)

Electron Affinities of some A Group elements (kJ/mole)

Period	IA	IIIA	IVA	VA	VIA	VIIA
1	H *−73*					
2	Li *−60*	B *−27*	C *−122*	N 0	O *−141*	F *−328*
3	Na *−53*	Al *−44*	Si *−134*	P *−72*	S *−200*	Cl *−349*
4	K *−48*	Ga *−30*	Ge *−120*	As *−77*	Se *−195*	Br *−325*
5	Rb *−47*	In *−30*	Sn *−121*	Sb *−101*	Te *−190*	I *−295*
6	Cs *−45*	Tl *−30*	Pb *−110*	Bi *−110*	Po *−180*	At *−270*

3. Periodic Trends in Electron Affinity

   a. Down a group:

   b. Across a period from left to right:

4. Size of anion versus size of neutral atom.

C. IONIC RADIUS (Ion Size)

1. Periodic Trends

   a. Down a group:

   b. Across a period from left to right:

2. Cations vs. Anions

X. INTRA- AND INTERPARTICLE FORCES (in order of decreasing strength)

A. Covalent Bond

B. Ionic Bond

C. Ion-Dipole

B. Hydrogen Bond

E. Dipole-Dipole

F. Dipole-Induced Dipole

G. London Forces