CHEMISTRY 20 -- UNIT 3: CHEMICAL BONDING

INTRODUCTION

A. ***Periodicity***: __________ that can be obtained from the periodic table (Chapter 12)

1. Groups & Period Characteristics
   - a. 3 MAIN GROUPS:
   - b. 5 GROUP NAMES:
   - c. PERIODS:
2. Reactivity on the periodic table
   - * Most reactive metal: ___________; Most reactive nonmetal: ____________
   - * Halogens react vigorously with _______ metals because they transfer electrons
   - * Alkali metals and Alkaline metals react with _________ to form ______ coats
   - * Alkali metals also react violently with _________
   - * Nobel gases are very ________ because their orbitals are full
3. State: solid, liquid or gas at room temperature

B. Electrons in Atoms (Chapter 11)

History:

1. J. J. Thomson's Billiard Ball Model:
2. E. Rutherford's Raisin Bun Model:
3. Bohr's Energy level Model
   - a. ***Energy levels***:
   - b. Energy level diagrams for atoms

   - a. Energy level diagrams for ions

4. ***Quantum mechanics*** Orbital Model

C. Properties of Ionic and molecular compounds

	Ionic	***Molecular***
State		
Color		
Solubility		
Conductivity		
Bond		
Ions Present		
Components		
Melting point		
Naming Rules	1. 2.	1. 2. Common: water, ozone, peroxide, ammonia, methane, ethane, methanol, ethanol, sucrose, glucose
Formula Rules	1. 2. 3. 4. 5.	1. 2.

D. What is a chemical ***Bond***

- A chemical bond is ____
- Atoms try to achieve the electron structure (not proton structure) or the nearest _________
- When a chemical bond is formed, the atoms form a more stable energy level. They will have a net _________ of energy to obtain this. Forming a chemical bond is _____________

ie) H + H ----> H_2 + 435 kJ

* When a chemical bond is broken, the atoms will have a net _________ of energy. Breaking a chemical bond is ______________________

ie) H_2 + 435 kJ ----> H + H

- The energy required to break a bond is called the ________________________________
- GRAPH: Potential Energy vs distance for a Hydrogen bond

* There are two main types of bonding forces

a. ***Intermolecular forces***: ____

____

ie)

The amount of force depends upon many factors like size, polarity, number and type of electrons etc...

b) I***ntramolecular forces:*** ____

____

ie)

* Approximate Bond Force strength..

Strongest ---> Weakest

_____________->_____->_______->_______->_______->___________->______________

In the flow diagram, place **INTRA** below the bonding force that is intramolecular and **INTER** below the force that is intermolecular.

<u>VALENCE ORBITALS</u>

A. Bohr said that electrons exist in energy ___________________ outside the nucleus. These _________are regions around the nucleus where you would most likely find the electrons.

B. ***<u>AUFBAU'S PRINCIPLE:</u>*** ____

<u>Valence electrons</u>____
___________________________________(free electrons) Group # = # of valence electrons.

<u>Valence orbital</u> ____.

(***<u>Valere</u>*** in Latin means to be strong.)

C. The rules of valence electrons are as follows

1. ***<u>PAUL'S EXCLUSION PRINCIPLE</u>***: ____

____

2. ____

____

____

3. ***<u>HUND'S RULE</u>***: ____

____

<u>lone pairs</u>. ____

<u>bonding electron</u> ____

4. ***<u>OCTET RULE</u>***: ____

Exceptions: ____

EXAMPLE: sulphur atom

$6e^-$--- 2e,2e,1e,1e   Valence orbital with 2 lone pairs & 2 bonding electrons

$8e^-$ --- 2e,2e,2e,2e

$2e^-$ --- 2e

$16p^+$

S

D. Electron Configurations (optional)

1. There are principle energy levels with sublevels. Each sublevel has a different cloud shape and can hold a different number of electrons.

s orbital	holds 2 e	spherical
p orbital	holds 6 e	dumbbell shaped
d orbital	holds 10 e	complex
f orbital	holds 14 e	complex

The principle energy levels fill the sublevels as follows:

$1s^2$			
$2s_2$	$2p^6$		
$3s^2$	$3p^6$	$3d^{10}$ (period 4 atoms)	
$4s^2$	$4p^6$	$4d^{10}$ (period 5 atoms)	$4f^{14}$ (period 6)
$5s^2$	$5p^6$	$5d^{10}$ (period 6 atoms)	$5f^{14}$ (period 7)
$6s^2$	$6p^6$	$6d^{10}$ (period 7 atoms)	
$7s^2$	$7p^6$ (future elements)		

Electrons wish to find the lowest possible energy state

The s orbitals fill the alkali metals and alkaline earths

The p orbitals are the last 6 elements on the right of each row. The very last is the Noble gas.

The "d" orbitals form the transition metal part of the table

The "f" orbitals form the rare earths (lanthanide and actinide series)

LEWIS DIAGRAM OR ELECTRON DOT DIAGRAM FOR ELEMENTS & IONS

Gilbert Lewis (1875 - 1946) developed a scheme for drawing particles with valence electrons shown as dots: Lewis Dot Diagram. (see page 300 for chart of beginning elements in the periodic table)

A. ***Electron dot*** ____

B. Rules for Drawing Lewis Diagrams for elements and simple ions

1. Write the ______________________ to represent the nucleus & the innermost energy levels. Determine the number of valence electrons by the elements ________________.
2. Place a dot to represent _________________. Start by placing one dot by each side of the element symbol.
3. If necessary, start filling in the second dot to a maximum of _____ positions (octet)
4. Remove electrons for cations or add electrons for anions & enclose both in bracket

EXAMPLES: (pg 300)

N (Group 5 = 5 valence electrons)

C. Examples

Element or Ion	Number of valence electrons	Electron Dot Diagram	Number of bonding electrons	Number of lone pairs
calcium (Ca^{+})				
calcium (Ca)				
germanium (Ge)				
potassium (K)				
sulphur (S)				
nitrogen (N)				
aluminum (Al)				
neon (Ne)				
flouride (F^{-})				
fluorine (F)				
helium (He)				

D. Incorrect Lewis diagrams:

1. C 2. P 3. B 4. Na 5. P 6. S

E. Lewis Diagrams for Molecular Compounds

* Sum the valence electrons for all the atoms in the molecule. Only this number of valence electrons can be used.

* Determine if necessary which atom can form the most number of bonds. ☝ This is the ___________________

* Valence electrons obey the ________________ rule (except hydrogen)

EXAMPLES:

1) Cl_2

2) CH_4

F. Lewis Diagrams for multiple bonds (F & G) above

* Same rules as above

* ____________ covalent bonds have two pairs of electrons shared between atoms

* ____________ covalent bonds have three pairs of electrons shared between atoms

1) O_2

2) CO_2

3) N_2

G. Lewis Diagrams for Ionic Compounds

* Electrons are ___________________ from metal to nonmetal

* The electrons around the metal are __________________

* Charged species are enclosed in ____________________

EXAMPLES:

1) NaCl

2) MgS_2

H. Lewis Diagrams for Complex Ions (OPTIONAL)

* ____________ the valence electrons for all the atoms in the molecule. Add electrons for negative charges and subtract for positive charges.

* Arrange the electrons around the atom so the sum is achieved. Use double and triple bonds if necessary.

* Charged species are enclosed in ________

EXAMPLES:

1) Cyanide

2) Sulphate

3) Ammonium

BONDS

I. ***Ionic Bonding***

A. Review

* ***Ions***: ____

Metallic ions: ___________________________________ (______electronegativity)

Non-metallic ions: ___________________________________ (________ electro)

Complex ions: ____

* ***Ionic bond***:____

* Ionic bonds produce a __________________________ with a repeating pattern.

* Ionic compounds characteristics (Look at beginning of your notes)

B. Electron dot diagrams to show ionic compounds (Review)

Na + Cl -----> $[Na]^{+}[\ Cl\ ]^{-}$(brackets show that they are not sharing)

C. Oxidation - Reduction (LEO goes GER; OIL RIG) (Optional)

1. ***Oxidation***: ___________of electrons (metal) Na ----> $Na^{+} + 1e^{-}$(half reaction)

The lower the electronegativity, the easier it is oxidized

2. ***Reduction***: ____________ of electrons (non-metal) $Cl + 1e^{-}$ ----> Cl^{-}

. Stopping Corrosion (optional)

METHOD	EXAMPLES	EXPLANATION
Plating	* w/gold or silver or tin	* Covers the reactive metal
Cathodic protection	*blocks of Zn or Mg in an Fe/Be ship.	Attach a more reactive metal that corrode/oxidize easier. This saves the other metal
Oxide coating	chromium(III) oxide, aluminum oxide, zinc oxide, tin(IV) oxide	***Galvanization***: Covers the steel with ZnO
Creating alloys	stainless steel	
Painting or covering with plastic		

D. Ionic models (optional diagrams)

II. ***Metallic bonding***: ____
____
PROPERTIES:
* ____
* ____
* Diagram

ie) Al, Fe, Cu, Zn, Ca, Na, Ag, Pb, Hg

III. ***Network covalent bonding***:____
____
PROPERTIES:
*____
*____
* Diagram

ie) diamonds (C_n), Carborundum (SiC),

III. ***Covalent bonds***:

A. Introduction

1. Summary:

* occur between _______________ atoms and result in the mutual _______ of e-

* unpaired electrons between two atoms that collide, are shared until they have an octet (except Hydrogen) & have the same electron structure as a Noble gas. ie) Cl + Cl ------ Cl Cl

* Covalent bonds may be single, double or triple

2. ***Bonding Capacity***: The maximum number of bonds a molecular atom can form.

Covalent bonds are represented by a dash. (- single, = double , ≡ triple)

ATOM	**# OF VALENCE ELECTRONS**	**# OF BONDING ELECTRONS**	**BONDING CAPACITY & TYPES OF BONDS**
hydrogen (H)			
carbon (C)			
nitrogen (N)			
oxygen (O)			
fluorine(halogen)(F)			
neon(noble) (Ne)			

3. Draw the ***structural diagram*** by replacing the shared pairs with a line (bond) & omitting the lone pairs. The ***molecular formula*** is written with only the symbols.

Examples **Molecular** **Lewis** **Structural**

a. Diatomic:

b. One Lone pair:

c. Double bond

d. Triple bond:

e. Cyclic carbon compounds:

4. Problems with the lewis diagram model

Coordinate Covalent Bond: ____

____

a. CO; C<u><--</u>O; C   O

b. NH_4^+

c. H_3O^+

Resonance: ____

____

Hybrid: ____

ie) ozone

Exceptions to the octet rule (pg. 326)

a. NO_2

b. P and S may sometimes expand the octet to ten or twelve electrons

SUMMARY

* Lewis diagrams tell us: 1) electrons involved in bonding 2) how many bonds
* Lewis diagrams do not tell us: 1) what an atom looks like 2) where electrons are
* Bonds are formed by _______ electrons
* ____ electrons are involved in each bond, called a _______ bond
* Diagrams and structural formula’s usually satisfy the _______ rule

B. ***Stereochemistry***: ____

* ***VSEPR Theory***: ____

1. ASSUMPTIONS:
   * all molecules have a 3-D shape
   * number of valence electrons, number of bonding electrons and the central atom must be known to determine the shape.
2. RULES
   1) Draw: ____
   2) Count ____
   3) Minimize ____
3. SHAPES & POLARITY

1) ***Tetrahedra***l- ____. e.g.) CH_4

2) ***Pyramidal***- ____ e.g.) NH_3

3) ***V-shape*** - ____ e.g.) H_2O

4) ***Trigonal Planar*** -____ e.g.) C_2H_4

5) ***Linear***: * ____ e.g.) HCl

* ____e.g.)Cl_2

* ____e.g.) C_2H_2

* ____e.g.) CO_2

6) Combinations: e.g.) CH_3OH(tetrahedral around C & V-shaped around O)

C. Electronegativity: ____
(electron attractivity or an atoms desire for an octet)

* Principles of electronegativity developed by Linus Pauling (on periodic table)
  1. As you go left and down the periodic table(metals), e- are held more ___________.
  Therefore the element's reactivity to lose e- increases & electronegativity _________.
  ________________________________ (Least electronegative: Fr,Ce = 0.7)
  2. As you go right and up the periodic table(non-metals), e- are held more ___________.
  Therefore the element's reactivity to gain increases & electronegativity _____________.
  _____________________________________ (Most electronegative: F = 4.0)
  3. NOBLE GAS EXCEPTION: The noble gases hold their electrons most tightly,
  but have ____________________ and are hardly reactive because they have an ______.
* Bond type based on electronegativity
  1. If the electronegative difference between two bonded atoms is equal or greater than 1.7 the bond is ______.
  2. If the electronegativity difference between two bonded atoms is less than 1.7 than the bond is polar ___________ .
  3. If the electronegativity difference between two bonded atoms is O than the bond is nonpolar covalent or metallic

- Polarity based on electronegativity

Polar covalent bond: ____

e.g.) HCl--> __________________

Affected by: 1) electronegativity of the atoms; 2) bond dipoles & 3) shape

Polar molecule: ____

____

electric poles: ____

bond dipole: ____

Step 1) ____

Step 2) ____

POLAR: Elements with larger electronegativities become negative regions (delta) & the elements with smaller electronegativities become positive regions (delta).

NON POLAR: If the electronegativities are equal on the ends, than they cancel each other out.

Step 3) ____

Non-polar covalent bond: ____.

e.g.) Cl_2 --> __________________ CO_2 --> ________________________

CONCLUSION: Polarity is based on

D. Saturated vs. Unsaturated fats (optional)

	EXAMPLES	**EFFECT ON BODY**	**BONDS**
SATURATED			
POLYUNSATURATED			

INTERMOLECULAR BONDING AND FORCES

Intermolecular bonding: ____

- The size of the molecules and the molecular polarity determine the strength (weak)
- electron & nuclear attractions between molecules cause intermolecular bonding
- intermolecular forces affect ____

(Like substances (polar) dissolve like (polar) substances.)

- There are three main types:

The first two are called ______________________after Dutch Physicist Johannes van der Waals

1. __________________________: the electrostatic attraction between ***all*** molecules when they form ***temporary*** dipoles. These are the __________ intermolecular forces.

* LD hold molecules together when they are in solid and liquid states
* LD are weaker than covalent bonds because electrons are farther from nuclei
* If two molecules have the same # of electrons they are called _________________

ie) Ar = 18e & HCl = 18e BUT HCl has higher boiling point due to other forces.

* _______________ are an indirect measure of these forces. Br_2 is a liquid & Cl_2 is not because Br_2 has stronger dispersion forces.

2. _____________________: when electrostatic attraction between two polar molecules

ie)

3. ____________________: special, relatively strong dipole-dipole forces between molecules containing O-H, F-H, & N-H bonds.

* These bonds form because of the __________________ between hydrogen and _____
* The small size of hydrogen atom means _____________________________________ and this positive pole exerts a strong attraction to the negative pole of a nearby molecule.

- V shaped molecules with hydrogen bonds, like water, have stronger hydrogen bonds than other molecules, like ethanol because _________________________________

ie) H 👍 F ----H 👍 F (Hydrogen bonds are represented by dotted lines because ___ ____)

ie) water

ie) ethanol

* PHYSICAL PROPERTIES Affected by Intermolecular bonds

1. ____
____

ie) H_2O is a liquid at room temperature and boils at 100 C because of extra strong hydrogen bonds

HCl (18 e) bp = -83.7 C

HBr (36 e) bp = -67.0 C

HI (54 e) bp = -35.4 C

2. ____

ie) H_2O and methanol(CH_3OH) are soluble BUT H_2O and wax($C_{25}H_{52}$) are not

✌Likes dissolve likes✍ - polar dissolves polar so anything that dissolves in water must be polar. Grease, $C_{14}H_{30(s)}$ (non-polar) dissolves in benzene, $C_6H_{6(l)}$ (non-polar)

3. ____

ie) water ---> ice - patterns form & it becomes less dense due to hydrogen bonds

SUMMARY

* Strength of all the bonds from strongest to weakest

network covalent 👍 ionic 👍 metallic(vary) 👍 dispersion 👍 hydrogen 👍 dipole-dipole

* Limitations of bonding model
    1. Structures of common gases such as NO2 & O3 cannot be explained easily
    2. Special properties of gases like oxygen are difficult to explain
    3. Formation of complex ions like nitrate, nitrite, carbonate, phosphate and sulphate are difficult to explain
    4. Graphite and diamond are different structural forms of carbon. Reasons for the variety are difficult to explain
* Summary of two bond types

INTERMOLECULAR BONDS	**INTRAMOLECULAR BONDS**
Types: Van der Waals(London dispersion & dipole-dipole), Hydrogen	Types: Ionic, covalent, metallic, network covalent
Between molecules	Within molecules
Small energy change when broken	Large energy change when broken
Creates physical properties (melting point, boiling point, surface tension, solubility)	Creates chemical properties
Distance between charges is larger	Distance between charges is smaller
Polarity affects strength	Electronegativity affects type & structure
Generally weaker bonds	Generally stronger bonds

* Electronegativity difference

Electronegativity difference	Type of bond	Description
> or = 1.7	Ionic	Transfer of electrons between metal & nonmetal
< 1.7	Polar covalent	Electrons shared between unlike nonmetal atoms
0	Nonpolar covalent	Electrons shared between like nonmetal atoms
0	Metallic	Electrons move freely between metal ions