

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

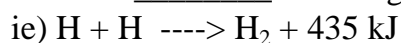
4. **Quantum mechanics** Orbital Model

C. Properties of Ionic and molecular compounds

	<u><i>Ionic</i></u>	<u><i>Molecular</i></u>
State		
Color		
Solubility		
Conductivity		
Bond		
Ions Present		
Components		
Melting point		
Naming Rules	<ol style="list-style-type: none"> 1. 2. 	<ol style="list-style-type: none"> 1. 2. Common: water, ozone, peroxide, ammonia, methane, ethane, methanol, ethanol, sucrose, glucose
Formula Rules	<ol style="list-style-type: none"> 1. 2. 3. 4. 5. 	<ol style="list-style-type: none"> 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 _____



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

Breaking a chemical bond is _____



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

D. Electron Configurations (optional)

- 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)				
fluoride (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

_____ 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
<u>Cathodic protection</u>	*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	<u>Galvanization:</u> 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:

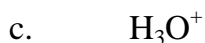
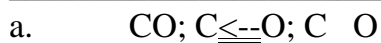
c. Double bond

d. Triple bond:

e. Cyclic carbon compounds:

4. Problems with the lewis diagram model

Coordinate Covalent Bond: _____



Resonance: _____

Hybrid: _____

ie) ozone

Exceptions to the octet rule (pg. 326)



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) Tetrahedral- _____ . e.g.) CH₄

2) Pyramidal- _____ e.g.) NH₃

3) V-shape - _____ e.g.) H₂O

4) Trigonal Planar - _____ e.g.) C₂H₄

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

* _____ e.g.) Cl₂

* _____ e.g.) C₂H₂

* _____ e.g.) CO₂

6) Combinations: e.g.) CH₃OH(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 0 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 δ^-) & the elements with smaller electronegativities become positive regions (delta δ^+).

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

Step 3) _____

Non-polar covalent bond: _____.

e.g.) Cl₂ --> _____ CO₂ --> _____

CONCLUSION: Polarity is based on

D. Saturated vs. Unsaturated fats (optional)

	<u>EXAMPLES</u>	<u>EFFECT ON BODY</u>	<u>BONDS</u>
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. Br₂ is a liquid & Cl₂ is not because Br₂ 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($\text{C}_{25}\text{H}_{52}$) are not

☞ Likes dissolve likes ☞ - polar dissolves polar so anything that dissolves in water must be polar. Grease, $\text{C}_{14}\text{H}_{30(\text{s})}$ (non-polar) dissolves in benzene, $\text{C}_6\text{H}_{6(\text{l})}$ (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 \hookrightarrow ionic \hookrightarrow metallic(vary) \hookrightarrow dispersion \hookrightarrow hydrogen \hookrightarrow 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

<u>INTERMOLECULAR BONDS</u>	<u>INTRAMOLECULAR BONDS</u>
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
$> \text{ or } = 1.7$	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