Honors Chemistry – Unit 3 Review

Chapter 7 – Ionic & Metallic Bonding

1. The electrons in the highest occupied energy level of an atom are called the VALENCE electrons.

2. The OCTET rule states that atoms in compounds tent to have the electron configuration of a noble gas.

3. Oxygen atoms attain a stable electron configuration by GAINING two electrons.

4. Ionic compounds are composed of METALS and NONMETALS which are arranged in a repeating 3D crystal structure. This structure makes these compounds BRITTLE. When DISSOLVED or MELTED, ionic compounds can conduct electricity.

5. Determine the number of valence electrons in each of the following and then draw a Lewis do structure:
   a. Magnesium 2 Mg
   b. Chlorine 7 Cl
   c. Hydrogen 1 H
   d. Neon 8 Ne

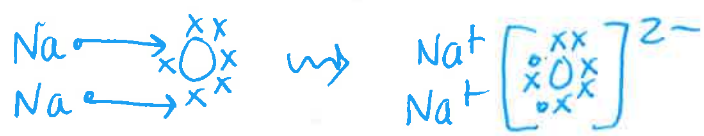
6. Write the electron configuration for the following:
   a. Calcium ion Ca^{+2} $1s^2 2s^2 2p^6 3s^2 3p^6$
   b. Strontium ion Sr^{+2} $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
   c. Sulfur ion S^{2-} $1s^2 2s^2 2p^6 3s^2 3p^6$
   d. Fluorine ion F^- $1s^2 2s^2 2p^6$

7. Which of these is not an ionic compound?
   a. KF
   b. Na_2SO_4
   c. SiO_2 2 nonmetals
   d. Na_2O

8. Using Lewis Dot diagrams, show how an ionic bond of sodium oxide is formed.

   Na^+ O^{2-}

   Na_2O.



9. Describe how a metallic bond is formed.

   metal atoms get close to each other so "d" orbitals can overlap. This allows e⁻ to become delocalized creating ⊕ metal cation & free "sea of ⊖ electrons." This opposite attraction of ⊕ & ⊖ makes the bond.

Chapter 8 – Covalent Bonding

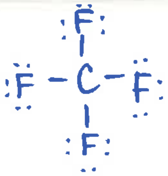
1. Covalent bonds occur between NONMETALS and NONMETALS due to the SHARING of electrons.

2. How many electrons are shared in the following bonds?
   a. Single covalent bond 2e⁻
   b. Double covalent bond 4e⁻
   c. Triple covalent bond 6e⁻

3. For the following compounds – CF_4, CO_2, NH_3, N_2, CO, SF_6, BF_3, CH_2Cl_2, H_2O – do the following:
   a. Draw the Lewis Dot structure
   b. Determine the number of lone pair electrons on the central atom
   c. Determine the number of atoms bonded to the central atom
   d. Indicate the VSEPR geometry for each molecule
   e. Determine if the bonds are nonpolar covalent or polar covalent
   f. Determine if the molecule is nonpolar or polar
   g. Determine the type(s) of intermolecular attractions (dispersion, dipole, hydrogen bonding)
   h. Calculate the bond dissociation energy for all bonds in the molecules

See next page for answer to all questions

CH_4 a)

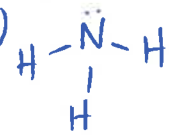


b) Ø
c) 4
d) Tetrahedral
e) $4.0-2.5=1.5$ polar covalent
f) non polar
g) dispersion
h) $488\times 4=1952$ $\frac{kJ}{mol}$

CO_2 a) $:\ddot{O}=C=\ddot{O}:$

b) Ø
c) 2
d) linear
e) $3.5-2.5=1.0$ polar covalent
f) non polar
g) dispersion
h) $736\times 2=1472$ $\frac{kJ}{mol}$

NH_3 a)



b) 1
c) 3
d) trigonal pyramidal
e) $3.0-2.1 \cong 0.9$ polar covalent
f) polar
g) All 3
h) $391\times 3=1173$ $\frac{kJ}{mol}$

N_2 a) $:N\equiv N:$

b) Ø
c) Ø
d) linear
e) $3.0-3.0=Ø$ non polar covalent
f) non polar
g) dispersion
h) 945 kJ/mol

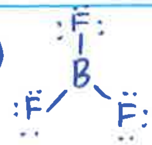
CO a) $:C\equiv O:$

b) Ø
c) Ø
d) linear
e) $3.5-2.5=1.0$ polar covalent
f) polar
g) dispersion, dipole
h) 1075 kJ/mol

SF_6 a)

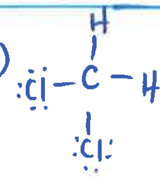
b) Ø
c) 6
d) octahedral
e) $4.0-2.5=1.5$ polar covalent
f) non polar
g) dispersion
h) $343\times 6=2058$ $\frac{kJ}{mol}$

BF_3 a)



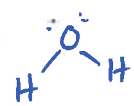
b) Ø
c) 3
d) trigonal planar
e) N/A (makes it look ionic)
f) non polar
g) dispersion
h) $766\times 3=2298$ $\frac{kJ}{mol}$

CH_2Cl_2 a)



b) Ø
c) 4
d) tetrahedral
e) CH $2.5-2.1=0.4$ nonpolar covalent
CCl $3.0-2.5=0.5$ polar covalent
f) polar
g) dispersion, dipole
h) $393\times 2=786$
$330\times 2=660$
$+$ 1446 $\frac{kJ}{mol}$

H_2O a)



b) 2
c) 2
d) bent
e) $3.5-2.1=1.4$ polar covalent
f) polar
g) all 3
h) $464\times 2=928$ $\frac{kJ}{mol}$

4. Order the types of bonds from strongest to weakest: London dispersion forces, hydrogen bonds, dipole interactions, ionic bonds, covalent bonds.

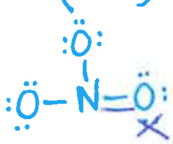
ionic bonds, covalent bonds, hydrogen bonds, dipole interactions, London dispersion forces

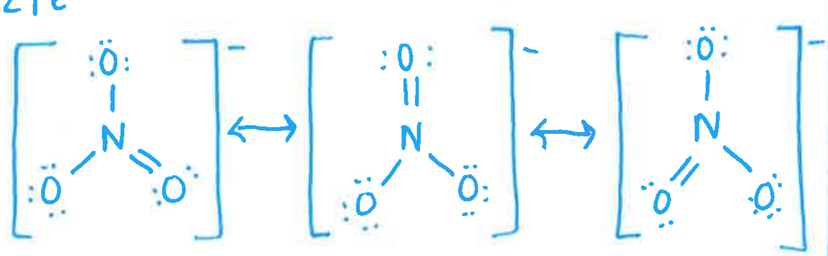
5. Which of these molecules can form a hydrogen bond with a water molecule?

a. N_2 non polar
b. NH_3 polar → must be a polar molecule and have another H w/ another highly electronegative element (N, O, Cl, F)
c. O_2 non polar
d. CH_4 non polar

6. Draw all forms of the Lewis structure of the nitrate polyatomic ion.

NO_3^-
$5+(3\times 6)+1=24e^-$





Chapter 9 – Chemical Names & Formula

M– Molecular, I – Ionic, A – Acid

M 1. nitrogen trifluoride NF_3

I 2. barium phosphide Ba_3P_2
Ba^{+2} P^{3-}

M 3. P_4O_{10} tetraphosphorus decoxide

M 4. SCl_2 sulfur dichloride

I 5. $Cu(OH)_3$ copper (III) hydroxide
Cu^{+3} OH^-

I 6. ammonium carbonate $(NH_4)_2CO_3$
NH_4^+ CO_3^{2-}

A 7. carbonic acid H_2CO_3
ic → ate
H^+ CO_3^{2-}

A 8. HCl hydrochloric acid
halogen

M 9. phosphorus triiodide PI_3

M 10. disulfur decafluoride S_2F_{10}

I 11. K_2S potassium sulfide

I 12. $NiSO_4$ nickel (II) sulfate
Ni^{+2} SO_4^{2-}

I 13. aluminum phosphate $AlPO_4$
Al^{+3} PO_4^{3-}

I 14. magnesium perchlorate $Mg(ClO_4)_2$
Mg^{+2} ClO_4^-

I 15. iron (III) sulfide Fe_2S_3
Fe^{+3} S^{2-}

M 16. dinitrogen monoxide N_2O

A 17. H_2SO_3 sulfurous acid
SO_3^{2-}
sulfite → ous

M 18. CCl_4 carbon tetrachloride

I 19. calcium iodide CaI_2
Ca^{+2} I^-

A 20. hydrobromic acid HBr
halogen
H^+ Br^-

A 21. bromic acid $HBrO_3$
ic → ate
H^+ BrO_3^-

I 22. $SrCl_2$ Strontium chloride

I 23. PbS lead (II) sulfide
Pb^{+2} S^{2-}

M 24. dinitrogen tetroxide N_2O_4