COPPERBELT UNIVERSITY
CHEMISTRY DEPATMENT
CH110/CH120/FO130 TUTORIAL SHEET 7- CHEMICAL BONDING

1. Indicate the number of valence electrons in the following:
   a) oxygen b) Group 1
   c) an element with an electron configuration of $1s^22s^22p^63s^23p^1$

   ANSWER:
   a) Oxygen is a Main Group element in Group 6 it has 6 valence electrons
   b) A Group 1 element has 1 valence electron
   c) This is a Main Group element of Group 3, it has 3 valence electrons

2. Define each of the following:
   a) Ionic bond
   b) Covalent bond
   c) Electronegativity

   ANSWER:
   a) An ionic bond ***is a force of attraction between oppositely charged ions a compound***
   b) A covalent bond is a ***type of bonding in which electrons are shared by atoms of a compound***
   c) Electronegativity *is the tendency of an atom in a molecule to attract shared electrons to itself*

   See Glossary Zumdhal and Zumdahl, 9th Edition.

3. a) List the following elements in order of increasing electronegativity: O, Ge, C

   ANSWER:
   Ge (1.8), C (2.5) and O (3.5). See page 357, Zumdahal and Zumdahl, 9th Edition.

   b) Choose the atom or ion in each set with the smallest atomic radius.

   i. Li, Li^+, H^-

   ANSWER:
   Li^+: The explanation is that Li^+ and H^- are isoelectronic with a $1s^2$ electronic configuration and smaller than Li with an $1s^22s^1$ electronic configuration. However, Li^+ is smaller than H^- because it has a higher nuclear charge drawing the electrons closer to the ion's nucleus.

   ii. Na^+, Cl^-, K^+

   ANSWER:
   Na^+: The explanation is that K^+ and Cl^- are isoelectronic with a [Ar or $1s^22s^22p^63s^23p^6$] electronic configuration but Na^+ with [Ne or $1s^22s^22p^6$] is smaller than.

iii. F, O^{2-}, F^{-}

ANSWER:

F: The explanation is that O^{2-} and F^{-} are isoelectronic with a [Ne] or electronic configuration have more electrons in the valence shell than F.

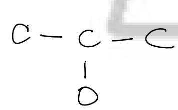
4. a) List at least 2 ions with each of the following electron configurations:
   i. $1s^22s^22p^6$
   ii. $1s^22s^22p^63s^23p^6$
   iii. $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^6$

   ANSWER:
   i. O^{2-}, F^{-}
   ii. Cl^{-}, K^{+}
   iii. $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^6$

   b) What the term is used to describe two ions with the same electron configuration?

5. Draw Lewis structures for the following compounds: (Follow the Octet Rule.)
   a) H_2S
   b) HCN
   c) NH_2F
   d) CH_2I_2
   e) CO_3^{2-}
   f) SO_3^{2-}
   g) diphosphorus dichloride
   h) dinitrogen tetroxide
   i) $C_3H_4Cl_2$
   j) C_3H_4

6. a) Add hydrogen atoms and electrons in order to complete Lewis structures of the following compound: C_3H_6O (acetone; nail polish remover)



b) Write Lewis structures for the following. Show all resonance structures where applicable.
   i. NO_2^-
   ii. NO_3^-
   iii. N_2O_4

7. a) Explain the basis of a covalent bond. What makes a bond polar? What makes a molecule polar?
   b) Indicate the molecular **<u>geometry</u>** of compounds (a)-(d) in question 5 above. For each of these compounds, indicate whether they are polar or nonpolar.

   b) Indicate the bond polarity (show the partial positive and partial negative ends) in the following bonds.
   i. C – O
   ii. P – H
   iii. H – Cl
   iv. Br – Te
   v. Se – S
   vi. Cl – I
   vii. Si – S
   viii. O – P
   ix. Br – Br

   c) For each set below, indicate which bond would be the most polar:
   i. C – F, N – F, or O – F
   ii. Si – F or C – F

8. Indicate the molecular geometry of the following compounds (which disobey the octet rule):
   a) SF_4 b) XeF_4 c) PF_5 d) SeF_6

9. a) Explain the following intermolecular forces:
   i. London forces
   ii. Dipole-dipole interaction
   iii. Hydrogen bonding

   b) Explain the difference in boiling points for each of the following
   i. HF (20^0C) and HCl (-85^0C)
   ii. HCl (-85^0C) and LiCl (1360^0C)

10. The equations for the heats of formation of copper (I) oxide, Cu_2O and copper (II) oxide, CuO are:

$$2Cu(s) + \tfrac{1}{2}O_2(g) \rightarrow Cu_2O(s);\quad \Delta H^0_f = -166.7 kJ/mol$$

$$Cu(s) + \tfrac{1}{2}O_2(g) \rightarrow CuO(s);\quad \Delta H^0_f = -155.2 kJ/mol$$

The first and second ionization energies of copper are 750 and 2000kJ/mol respectively; its atomization energy is 339.3kJ/mol. The atomization energy of oxygen is 249.2kJ/mol. The first and second electron affinity for oxygen are -141.4kJ/mol and 790.8kJ/mol respectively. Draw the Born-Haber cycle and calculate the lattice energies (ΔH_{latt}) of the two oxides.