Name: **Key** *(Complete and keep as a study aid)*

| Chemistry 20 | Unit A Learning Outcome - Review Outline | GCCHS |
|---|---|---|

**Review Blueprint**

| # | | Learning Outcome |
|---|---|---|
| 1. | A1.1 | recall principles for assigning names to ionic compounds |
| 2. | A1.2 | explain why formulas for ionic compounds refer to the simplest whole-number ratio of ions that result in a net charge of zero |
| 3. | A1.3 | define valence electron(s), electronegativity, ionic bond and intra-molecular force |
| 4. | A1.4 | use the periodic table and electron dot diagrams to support and explain ionic bonding theory |
| 5. | A1.5 | explain how an ionic bond results from the simultaneous attraction of oppositely charged ions |
| 6. | A1.6 | explain that ionic compounds form lattices and that these structures relate to the compounds' properties (*melting point, solubility, reactivity, etc*) |
| 7. | A2.1 | recall principles for assigning names to molecular substances |
| 8. | A2.2 | explain why formulas for molecular substances refer to the number of atoms of each constituent element |
| 9. | A2.3 | relate electron pairing to multiple and covalent bonds |
| 10. | A2.4 | draw electron dot diagrams of atoms and molecules, writing structural formulas for molecular substances and using Lewis structures to predict bonding in simple molecules |
| 11. | A2.5 | apply valence-shell electron-pair repulsion (VSEPR) theory to predict molecular shapes for linear, angular (V-shaped, bent), tetrahedral, trigonal pyramidal and trigonal planar molecules |
| 12. | A2.5<br>A2.6 | apply valence-shell electron-pair repulsion (VSEPR) theory to predict molecular shapes for linear, angular (V-shaped, bent), tetrahedral, trigonal pyramidal and trigonal planar molecules<br>illustrate, by drawing or by building models, the structure of simple molecular substance |
| 13. | A2.7 | explain intermolecular forces, London (dispersion) forces, dipole-dipole forces and hydrogen bonding |
| 14. | A2.8 | relate properties of substances (e.g., melting and boiling points, enthalpies of fusion and vaporization) to the predicted intermolecular bonding in the substances |
| 15. | A2.9 | determine the polarity of a molecule based on simple structural shapes and unequal charge distribution |
| 16. | A2.10 | describe bonding as a continuum ranging from complete electron transfer to equal sharing of electrons |

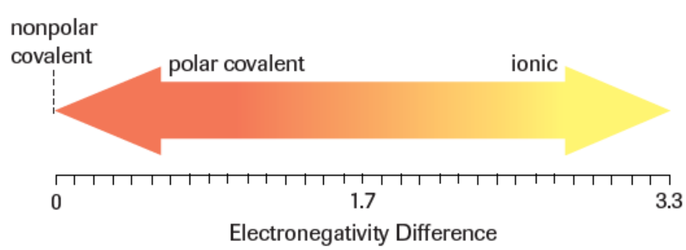



**Figure 12**
This model of the bonding type continuum shows the relationship of bonded atom electronegativity to bond polarity.

1) Complete the following chart:

| | **Name** | **Formula** |
|---|---|---|
| **a)** | **aluminum chloride** | **$AlCl_3$** |
| **b)** | **Lithium carbonate** | **$Li_2CO_3$** |
| **c)** | **aluminum sulfide** | **$Al_2S_3$** |
| **d)** | **Potassium Chloride** | **KCl** |
| **e)** | **copper (II) Nitrate** | **$Cu(NO_3)_2$** |
| **f)** | **Copper Sulfate** | **$CuSO_4$** |

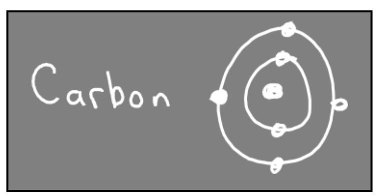
2) Explain why **aluminum sulfide** has the formula **$Al_2S_3$** instead of **AlS.**

**The aluminum ion has a 3+ charge and the sulfide ion has a 2- charge. To balance these to zero, we need to give them each a charge of 6 … ….. .. 2 x 3+ = 6+ while 3 x 2- = 6-. (***formulas for ionic compounds refer to the simplest whole-number ratio of ions that result in a net charge of zero)*

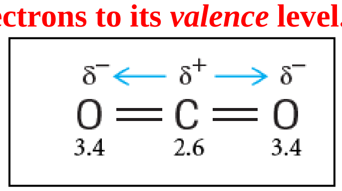
3) Define the following terms, (*provide examples and diagrams*).

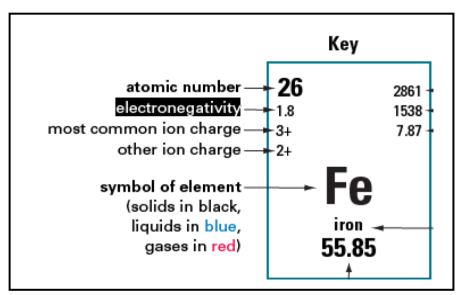
a) valence electrons – **are electrons that are present in the outermost energy level of an atom. These are the electrons that will be active in covalent bonding or electron exchange.**

**Carbon has 4 valance electrons**

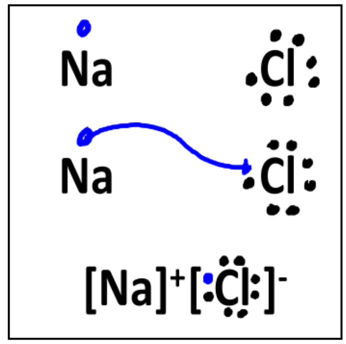



b) Electronegativity - **is a number that describes the relative ability of an atom to attract a pair of bonding electrons to its *valence* level.**

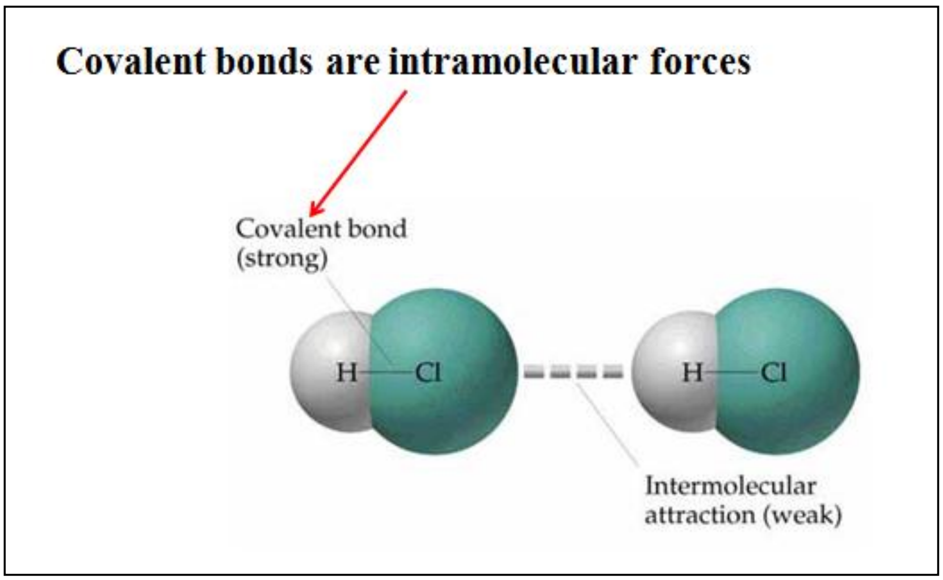


c) ionic bond - **refers to (electrostatic) force of attraction that exists between positive & negative ions (cations and anions). Ionic compounds like NaF and NaCl contain ionic bonds.**

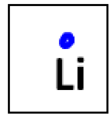



d) intra-molecular force - **refers to the relatively strong bonds or forces of attraction & repulsion *within* a molecule – typically covalent bonds.**

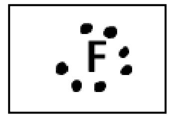


4) Complete each of the following directives regarding the compound ***lithium fluoride***:

a) Draw the Lewis dot diagram for a **lithium *atom*.**




b) Draw the Lewis dot diagram for a **fluorine *atom*.**




c) Use your periodic table to find the electronegativity of **lithium.**

**1.0**

d) Use your periodic table to find the electronegativity of **fluorine.**

**4.0**

e) Calculate the electronegativity difference between **lithium & fluorine.**

**$\Delta EN = EN_2 - EN_1 = 4.00 - 1.00 = 3.00$**

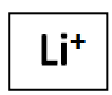
f) Use figure 12 on page 100 to predict if the bond between lithium and fluorine would be ionic, polar covalent or nonpolar covalent. Explain your answer.

**With an electronegitivity value of 3.00 it would be classified as ionic on the electronegitivity spectrum**

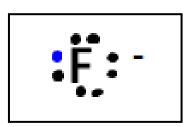
g) Use the location of each element, (lithium and fluorine) on the periodic table to explain another way that allowed you to predict that this was the type of bond present.

1) **On the periodic table Li is clearly a metal and Fluorine is a nom-metal. Compounds formed between metals and non-metals are generally classified as ionic compounds.**
2) **From the position on the periodic table, Li is an alkali metal. Alkali metals generally want to lose 1 electron. From its position, F is a Halogen. Halogens want to gain one electron. Compounds that result after the process of losing and gaining electrons are generally classified as ionic compounds.**

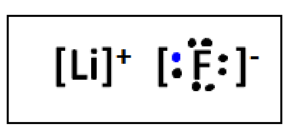
h) Draw the Lewis dot diagram for a **lithium *ion*.**




i) Draw the Lewis dot diagram for a **fluoride *ion*.**




j) Draw the Lewis dot diagram for a **lithium fluoride.**



**5)** Explain the force of attraction that takes place between the lithium ion and the fluoride ion in the compound that you drew above. **It is the electrostatic force of attraction ($F_e$) that occurs between unlike charges.**

**6)** Ionic compounds like NaCl do not form molecules. Explain what structure that they do form and list some of the properties that are determined by that structure. **Instead of molecules these ionic compounds form a 3-D crystal lattice. These structures affect properties like: melting point, solubility, reactivity etc.**

7) Complete the following chart:

| | Name | Formula |
|---|---|---|
| **a)** | **tetraphosphorus decaoxide** | $P_4O_{10}$ |
| **b)** | **carbon disulfide** | $CS_2$ |
| **c)** | **carbon dioxide** | $CO_2$ |
| **d)** | **nitrogen gas** | $N_2$ |
| **e)** | **nitrogen triiodide** | $NI_3$ |
| **f)** | **hydrogen peroxide** | $H_2O_2$ |

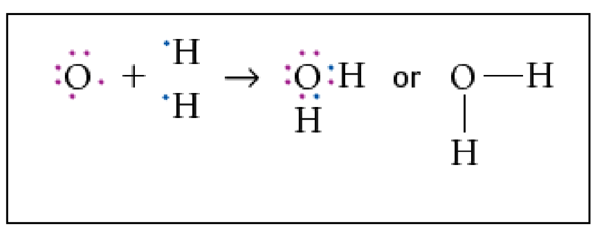
**8)** Explain why the molecular compound **$CO_2$** is called **carbon dioxide** instead of **carbon oxide.** – **This is a molecular compound, the IUPAC rules mandate that we include prefixes (except mono for the 1st element). Since there are two oxygen, we say dioxide.**

9) Explain each of the following:

a) Lone pairs - **A full valance orbital has two electrons, repels nearby electrons and is not available to share. Two electrons in the same orbital are called a *lone pair*.**

b) Bonding pairs – **Bonding pairs are composed of 2 bonding electrons from different elements that share an orbital in a covalent bond.**

c) Covalent bonds and electron pairs (see page 86) **A covalent bond forms between two elements when they share bonding electrons that form a bonding pair in a previously unfilled orbital to produce a stable octet.**



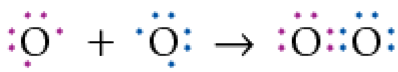
10) Complete each of the following directives regarding the compounds ***chlorine gas*** ($Cl_2$) and ***oxygen gas*** ($O_2$):

a) Draw the Lewis dot diagram for two **chlorine *atoms*** reacting to form ***chlorine gas.***

b) Describe the bond between the two chlorine atoms.

**Single covalent bond**

c) Draw the Lewis dot diagram for two **oxygen *atoms*** reacting to form ***oxygen gas.***



d) Describe the bond between the two oxygen atoms.

**Double covalent bond**

11) Complete each of the following directives regarding the VSEPR Theory:

a) What does VSEPR stand for?

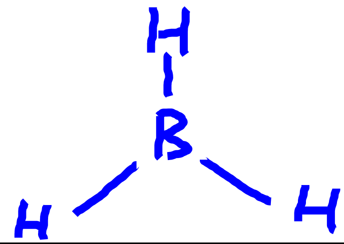
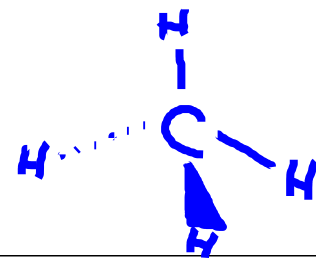
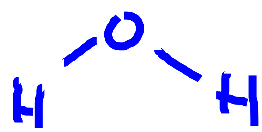
**Valance Shell Electron Pair repulsion**

b) Explain the VSEPR theory- **The VSEPR theory says that pairs of valance electrons will stay as far apart as possible due to the repulsion of their negative charges. This is what determines the shape of a particular molecule.**

**According to VSEPR:**

- **Only the valence electrons of the central atoms affect the molecular shape**
- **Valence electrons are paired in a molecule or polyatomic ions.**
- **Bonded pairs of electrons and lone pairs are treated the same.**
- **Valence electron pairs repel each other (like charges repel)**
- **The molecule's shape is determined by electron pairs at their max distance apart**

12) Complete this chart:

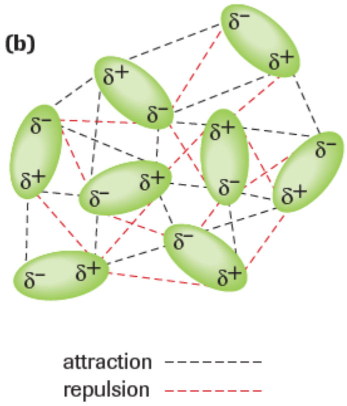
| Formula | Bond Pairs | Lone Pairs | Total Pairs | Molecular Shape | |
|---|---|---|---|---|---|
| | | | | Geometry | Stereochemical Formula |
| $CO_2$ | **2** | **0** | **2** | **linear** |  |
| $BH_3$ | **3** | **0** | **3** | **triagonal planar** |  |
| $CH_4$ | **4** | **0** | **4** | **tetrahedral** |  |
| $H_2O$ | **2** | **2** | **4** | **Angular (V-shaped)** |  |

13) Explain each of the following terms:

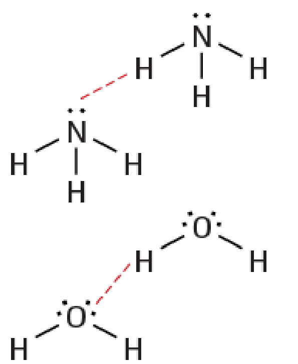
a) intermolecular forces – **these are the relatively weak forces of attraction & repulsion *between* molecules. Examples include *dipole-dipole and London dispersion forces*.**

b) dipole-dipole forces – **Polar molecules (dipoles) are molecules that have a distinctive positive abd a distinctive negative end. When the negative end of one molecule is attracted to the positive of another molecule, it is called dipole-dipole force.**

c) London (dispersion) forces – **in a liquid, polar molecules can move and rotate to maximize attraction and minimize repulsion for a net greater overall attration.**




d) hydrogen bonding – **This force of attraction of the single proton in a hydrogen atom (small radius) .. which is involved in a covalent bond to an element with a high electronegativity… for the lone pair of electrons in an adjacent molecule.**



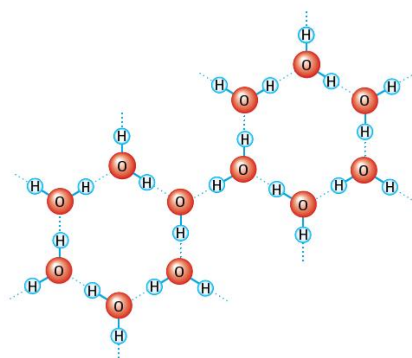


**Figure 11**
In ice, hydrogen bonds between the molecules result in a regular hexagonal crystal structure. The ···H– represents a hydrogen (proton) being shared unequally between two pairs of electrons.

14) Study Table 1, page 107 and then:

a) Make a rule to relate the total number of electrons in a hydrogen compound to its boiling point. – **The more electrons in a hydrogen compound, the higher the boiling point. (Due to greater attraction of the electrons present, the more energy that must be applied to change ot from a liquid to a gas).**

b) Would you expect these molecules to be polar or nonpolar? Explain.- **Based on VSEPR & molecular polarity theories, we would expect it to be non-polar.**

c) What would you predict their shape to be? – **Based on the four equivalent bonds, the shape is tetrahedral.**

15) Classify each of the following bonds as ionic, polar covalent or nonpolar covalent:

**a)** H and Cl – **3.2 - 2.2 = 1.0 so it is probably polar covalent**

b) I and Br – **3.0 - 2.7 = 0.3 so it is probably polar covalent**

c) C and O – **3.4 - 2.6 = 0.8 so it is probably polar covalent**

d) C and H – **2.6 - 2.2 = 0.4 so it is probably polar covalent**

16) Copy table 3, page 129 to help understand that bonding is a continuum ranging from complete electron transfer to equal sharing of electrons.

**Table 3** Properties of Ionic, Metallic, Molecular, and Covalent Network Crystals

| Crystal | Entities | Force/Bond | Properties | Examples |
|---|---|---|---|---|
| ionic | cations<br>anions | ionic | hard; brittle; medium to high melting point; liquid and solution conducts | $NaCl(s)$, $Na_3PO_4(s)$, $CuSO_4{\cdot}5H_2O(s)$ |
| metallic | cations<br>electrons | metallic | soft to very hard; solid and liquid conducts; ductile; malleable; lustrous | $Pb(s)$, $Fe(s)$, $Cu(s)$, $Al(s)$ |
| molecular | molecules | London<br>dipole–dipole<br>hydrogen | soft; low melting point; nonconducting solid, liquid, and solution | $H_2O(s)$ (ice), $CO_2(s)$ (dry ice), $I_2(s)$ |
| covalent network | atoms | covalent | very hard; very high melting point; nonconducting | $C(s)$, $SiC(s)$, $SiO_2(s)$ |