Name ________________________________**Chem 162, Section**: _______**Group Number:** ______

ALE 8. Lewis Structures of Elements and Ionic Bonding

(Reference: Sections 9.1 and 9.2 - *Silberberg 5th edition*)

Important!! For answers that involve a calculation you must show your work neatly using dimensional analysis with correct significant figures and units to receive full credit. No work, no credit. Report numerical answers to the correct number of significant figures. **CIRCLE ALL NUMERICAL RESPONSES**.

<u>Section 9.1 Atomic Properties and Chemical Bonding</u>: Read section 9.1 and then answer the following questions.

1. <u>Problem 9.1 (modified):</u> In Chem 161 you learned that metals tend to lose electrons when they react chemically. In general terms, how does each of the following atomic properties influence the ***metallic character*** (i.e. the ability to lose electrons) of the *<u>main-group elements within a period</u>*?
   a.) Ionization energy

   b.) Atomic Radius

   c.) Number of valence electrons

   d.) Effective nuclear charge

2. <u>Problem 9.3 (modified):</u>
   a.) What is the relationship between the tendency of a main-group element (i.e. A-group element) to form a monatomic cation as and its position in the periodic table?

   b.) In what part of the periodic table are the main-group elements that typically form monatomic cations with 1+ and 2+ charges?

   c.) In what part of the periodic table are the main-group elements that typically form monatomic anions with 1- and 2- charges?

3. Problems 9.7 and 9.9: Using only the periodic stable as a guide, state the type of bonding—ionic, covalent, or metallic—you would expect in the following substances. Briefly explain your reasoning in each case.

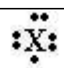
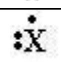
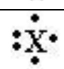
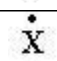
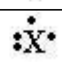
Substance	**Type of Bonding**	**Reasoning**
a.) $ICl_{3(g)}$		
b.) $N_2O_{(g)}$		
c.) $LiCl_{(s)}$		
d.) $Cr_{(s)}$		
e.) $H_2S_{(g)}$		
f.) $CaO_{(s)}$		

4. Problems 9.11 and 9.13: Using only the periodic table, draw the Lewis structure (electron dot structure) for each of the following.

a.) Ba b.) Kr c.) Br

d.) As e.) Se f.) Ga

5. Problems 9.15 (modified): Using only the periodic table, determine the group number and the condensed electron configuration of an element with the following Lewis structures.

Lewis Structure	**Group Number**	**Condensed Electron Configuration**
a.) 		
b.) 		
c.) 		
d.) 		
e.) 		

<u>Section 9.2 Ionic Bonding</u>: Read section 9.2 and then answer the following questions.

6. Problem 9.16 (modified): In order for any chemical bond to form, the overall process of bond formation must be exothermic. If energy is required to form monatomic cations from metals and monatomic anions from nonmetals (i.e. these processes are endothermic), why do ionic compounds exist?

7. Problem 9.21: Using only the periodic table as a guide, write condensed electron configurations and Lewis structures to depict the monatomic ions formed from each of the following reactants, and predict the formula of the compound the ions produce.
   For example, for Rb and O:

$$2\ Rb\ ([Kr]5s^1)\ +\ O\ ([He]2s^22p^4)\ \rightarrow\ 2\ Rb^+\ ([Kr])\ +\ O^{2-}\ ([He]2s^22p^6)\ =\ Rb_2O$$

   a.) Cs and S

   b.) O and Ga

   c.) N and Mg

   d.) Br and Li

8. Problems 9.23 and 9.25: Using only your knowledge of polyatomic ions and the periodic table as guides, identify the main group to which "X" belongs in each ionic compound formula.

   a.) X_3PO_4 b.) $X_2(SO_4)_3$ c.) $X(NO_3)_2$

   d.) CaX_2 e.) Al_2X_3 f.) XPO_4

9. Problem 9.27: For each pair, circle the compound with the higher (i.e. more negative, hence, more exothermic) lattice energy. *Explain your reasoning in each case*.

   a.) CaO or CaS

   b.) BaO or SrO

10. Problem 9.31:

   a.) Use the following to calculate $\Delta H^o_{lattice}$ of magnesium fluoride, MgF_2. *Show your work!!*

   $Mg(s) \rightarrow Mg(g)$ $\Delta H° = 148$ kJ

   $F_2(g) \rightarrow 2\,F(g)$ $\Delta H° = 159$ kJ

   $Mg(g) \rightarrow Mg^+(g) + e^-$ $\Delta H° = 738$ kJ

   $Mg^+(g) \rightarrow Mg^{2+}(g) + e^-$ $\Delta H° = 1450$ kJ

   $F(g) + e^- \rightarrow F^-(g)$ $\Delta H° = -328$ kJ

   $Mg(s) + F_2(g) \rightarrow MgF_2(s)$ $\Delta H_f° = -1123$ kJ

   $Mg^{2+}(g) + 2F^-(g) \rightarrow MgF_2(s)$ $\boldsymbol{\Delta H^o_{lattice}} = ?$

   b.) Compared with the lattice energy of LiF (-1050 kJ/mol) and the lattice energy of NaCl (−788 kJ/mol), does the relative magnitude of the lattice energy of MgF_2, calculated above, surprise you? *Explain*.

11. Without exception, all ionic compounds have very high (in the hundreds of ^oC) melting points and are, therefore, solids at room temperature. Use your knowledge of ionic bonding to explain why the melting point of ionic compounds is so high.