CHAPTER 8 – Basic Concepts of Chemical Bonding

Section 8.1 – Lewis Symbols and the Octet Rule

(a) Complete the Lewis electron-dot symbols for each of the following elements by drawing the valence electrons in an appropriate manner.

Na Mg Al Si P S Cl Ar

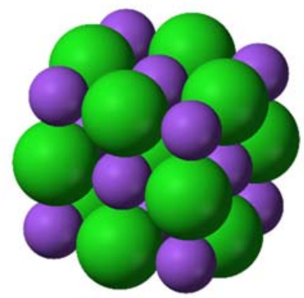
(b) Write a definition of the octet rule. Try to avoid using the word “eight” in your definition.

Section 8.2 – Ionic Bonding

(a) This is a diagram of the crystal lattice of NaCl.

How can you identify which ion is Na^+

and which ion is Cl^-?



(b) Fill in the valence electrons on both sides of the chemical equation shown below.
Draw an arrow to indicate the direction of electron transfer in this reaction.
Draw the appropriate charges on the ions in the product.

$$Na + Cl \longrightarrow Na \quad Cl$$

(c) Describe several properties of ionic substances.

(d) Look at Figure 12.25 on page 481. Explain why ionic solids tend to be brittle and can be cleaved along well-defined planes when a stress is applied to them.

(e) When sodium loses an electron, this is an (endothermic exothermic) process.

When chlorine gains an electron, this is an (endothermic exothermic) process.

$$Na(g) \rightarrow Na^+(g) + e^- \qquad \Delta H = +496 \text{ kJ}$$

$$Cl(g) + e^- \rightarrow Cl^-(g) \qquad \Delta H = -349 \text{ kJ}$$

(f) If we combine the two reactions shown above, we might think that the formation of sodium chloride is an endothermic process. Why is this view incorrect?

(g) The principal reason ionic compounds are stable is the ______________________________

____. This can be examined with Coulomb's law.

(h) Define lattice energy.

(i) Label each of the following reactions with a ΔH value of either +788 kJ or –788 kJ.

$$NaCl(s) \rightarrow Na^+(g) + Cl^-(g) \qquad \Delta H =$$

$$Na^+(g) + Cl^-(g) \rightarrow NaCl(s) \qquad \Delta H =$$

(j) Write Equation 8.4 on page 293.

For a given arrangement of ions, the lattice energy increases as the ______________________

______________ increase and as their ___________________ decrease.

(k) Explain why MgO has a higher lattice energy than NaF.

(l) Explain why NaF has a higher lattice energy than KCl.

(m) Without consulting Table 8.2, arrange these three compounds in order of increasing lattice energy: LiCl, KBr, and $MgCl_2$. Justify your answer.

(n) Even though lattice energy increases with increasing ionic charge, we never find ionic compounds that contain Na^{2+} ions. Explain.

(o) Write a balanced equation for each step of the Born Haber cycle. Include phases of matter for each substance.

Step	Description	Balanced equation	ΔH°
1	sublimation of 1 mole of Na		108 kJ
2	formation of 1 mole of gaseous Cl atoms from ½ mole of gaseous Cl_2 molecules		122 kJ
3	ionization energy for gaseous Na atoms		496 kJ
4	electron affinity for gaseous Cl atoms		–349 kJ
5	formation of solid NaCl from gaseous sodium and chloride ions		lattice energy

(p) Write the balanced equation for the standard enthalpy of formation of sodium chloride.
ΔH_f^o = –411 kJ

(q) Show how the six reactions mentioned in parts (o) and (p) can be used to calculate the lattice energy for sodium chloride.

(r) In forming an ion, a transition metal atom loses electrons from the _________ subshell first.

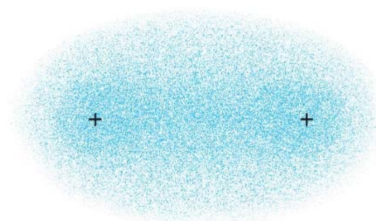
Then it will lose electrons from the ________ subshell until it reaches the charge on the ion.

(s) Write the electron configuration for each of the following ions:

Fe^{2+} _______________________ Fe^{3+} _______________________

Section 8.3 – Covalent Bonding

(a) The diagram at right represents the electron distribution in a molecule of hydrogen. Label the region on the diagram that represents the electrons involved in covalent bonding.



(b) The atoms in the H_2 molecule are held together because the two ________________________ are attracted to the concentration of negative charge between them.

(c) In writing Lewis structures, we usually show each ________________ electron pair as a line, and any ________________ electron pairs as dots.

(d) Write the Lewis structure for the diatomic molecules N_2, O_2, and F_2

(e) As a general rule, the length of the bond between two atoms _________________ as the number of shared electron pairs increases.

Section 8.4 – Bond Polarity and Electronegativity

(a) Bond polarity is a measure of how equally or unequally the electrons in any covalent bond are shared. Give two examples of each of the following bond types.

nonpolar covalent bond ___________ ___________

polar covalent bond ___________ ___________

ionic bond ___________ ___________

(b) Define electronegativity.

(c) If an atom has a large negative value for its electron affinity and a rather high value for its 1st ionization energy, it should have a rather ____________ value for electronegativity.

(d) The periodic trend is that electronegativity tends to ___________________ from left to right across a period, and it tends to ___________________ from top to bottom down a group.

(e) Use electronegativity values to explain why the F–F bond is nonpolar and the H–F bond is polar.

(f) Each of the following bonds is polar. Use the periodic trends in electronegativity to help you to label each atom with either a partial positive charge ($\delta+$) or a partial negative charge ($\delta-$).

H—O N—C I—F Si—Br Cl—As

(g) The greater the difference in electronegativity between two atoms, the ____________________ their bond is.

(h) If we use an arrow with a crossed end like this ($\leftarrow\!\!\!\!+$ reversed: $+\!\!\!\longrightarrow$) to denote the charge separation in a polar bond, the arrow always points toward the atom that has a ___________________ electronegativity value.

(i) Whenever two electrical charges of equal magnitude but opposite sign are separated by a distance, a ___________________ is established.

(j) The quantitative measure of the magnitude of a dipole is called its ________________________

(k) If a molecule is nonpolar, it has a dipole moment of ______________.

(l) Assuming that bond lengths are fairly constant, we can say that as the electronegativity difference between two atoms increases, the dipole moment of the bond between them _______________________

(m) Even though there is a continuum between the extremes of ionic and covalent bonding, the simplest approach is to assume that the bond between a metal and a nonmetal is _____________ and that the bond between two nonmetals is ______________________.

(n) $SnCl_4$ is a colorless liquid at room temperature. It freezes at –33°C and boils at 114°C. Do these properties suggest ionic or covalent bonding in this substance? ________________

Section 8.5 – Drawing Lewis Structures

(a) Draw Lewis electron dot structures for each of the following molecules. There should be only single bonds in each structure.

$SiCl_4$ PF_3 SBr_2 CH_3NH_2 C_2H_5Cl

(b) Draw Lewis electron dot structures for each of the following molecules. There may be a double or a triple bond in each structure.

$COCl_2$ C_2H_4 C_2H_2 HCN SO_2

(c) Draw Lewis electron dot structures for each of the following polyatomic ions.

ClO_3^- NO_3^- NO_2^+ SO_4^{2-} C_2^{2-}

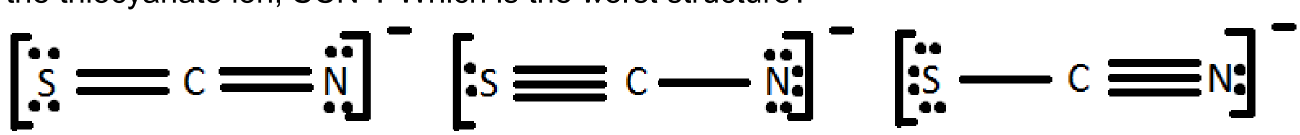
(d) The formal charge of any atom in a molecule or ion is the charge the atom would have if each bonding pair in the Lewis structure were ______________________________________

(e) Formal charge = _______________ minus ______________________________

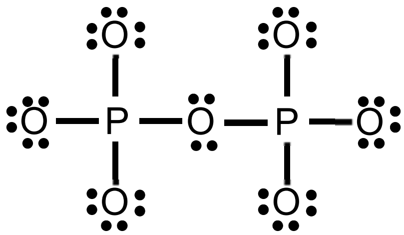
(f) Calculate the formal charges on each atom in the polyatomic ions that were drawn in part (c).

(g) If more than one Lewis structure is possible, we can use formal charges to predict the dominant structure. The dominant Lewis structure is generally the one in which the atoms bear formal charges closest to _____________. A Lewis structure in which any negative charges reside on the more __________________ atoms is generally more preferred.

(h) Based on formal charges, which of these Lewis structures is the most dominant structure for the thiocyanate ion, SCN^-? Which is the worst structure?



(i) Use formal charges to determine the overall charge of the following polyatomic ion.



Remember that formal charges do not represent real charges on atoms. These charges are just a convenient bookkeeping method that helps us decide which structure is preferred.

<u>Section 8.6 – Resonance Structures</u>

(a) Based on Figure 8.12 on page 309, does the ozone molecule contain one (shorter) double bond and one (longer) single bond? Explain.

(b) When a double-headed arrow is drawn between two resonance structures, does it mean that the two forms of the molecule are oscillating back and forth? Explain.

(c) How does “blue + yellow = green” help us understand the concept of resonance structures?

(d) Two or more resonance structures will have ________________ arrangement of atoms, but a different ____

(e) Draw two different resonance forms for the acetate ion, $CH_3CO_2^-$

(f) Draw three different resonance forms for the carbonate ion.

(g) Draw the two resonance structures for benzene.

(h) If a C–C single bond has a length of 1.54 Å and a C=C double bond has a length of 1.34 Å, what is the expected bond length in benzene,

C_6H_6, based on Figure 8.14 on page 311? ______________________

Bond Order is defined as the number of chemical bonds between a pair of atoms. For example, the bond order in F_2 is 1, and the bond order in O_2 is 2. The bond order in N_2 is 3. In molecules that have two or more resonance forms, the bond order does not have to be an integer. In benzene, the bond order between the carbon atoms is 1.5, based on resonance. Where does the 1.5 value come from? It represents the fact that each C–C bond in benzene can be either a single bond or a double bond in two different resonance structures. The average bond order based on these two resonance forms is $\frac{1+2}{3} = 1.5$

(i) What is the bond order for each of the following, based on resonance?

The O–O bond in O_3 _____________ The C–O bond in $CH_3CO_2^-$ ___________

The S–O bond in SO_3 ___________ The N–O bond in NO_3^- ______________

Section 8.7 – Exceptions to the Octet Rule

(a) Draw the Lewis structure of NO, NO_2, and ClO_2. Each of these Lewis structures has an odd number of electrons.

(b) Draw the Lewis structure of BF_3. The central atom has less than an octet.

(c) Why is a Lewis structure for BF_3 with a double bond an unfavorable structure?

(d) Boron trifluoride can react with molecules that have an unshared pair of electrons. This will allow boron to complete its octet. Draw the Lewis structure for NH_3BF_3, and determine the formal charge on nitrogen and boron.

(e) There are many molecules that have central atoms with more than an octet around them. Such molecules (or ions) are called hypervalent. Hypervalent molecules (or ions) are formed only for central atoms from period ________ and below in the periodic table. The principal reason for their formation is the ____ of the central atom.

(f) Circle the molecules (or ions) that should exist, based on these rules that apply to hypervalent molecules (or ions).

OF_4 SF_4 NF_5 PF_5 OF_6 SeF_6 XeF_2 IF_4^- Br_3^-

(g) Draw the Lewis structures for each of the molecules or ions that you circled in part (f).

(h) There are Lewis structures where you might have to choose between satisfying the octet rule and obtaining the most favorable formal charges in a hypervalent structure. Draw two different Lewis structures for the phosphate ion, PO_4^{3-}

obeys the octet rule | minimizes the formal charges on the atoms

(i) Draw two different Lewis structures for the chlorate ion, ClO_3^-

obeys the octet rule | minimizes the formal charges on the atoms

Section 8.8 – Strengths of Covalent Bonds

(a) The bond enthalpy is the energy required to ____

____

(b) The greater the bond enthalpy, the _____________________ the bond.

(c) Breaking a bond is always an ______________________ process, and forming a bond is always an _____________________ process.

(d) Use the bond enthalpies on page 316 and Equation [8.12] to estimate the ΔH for the following reaction:

$$N_2 + 3\,F_2 \rightarrow 2\,NF_3$$

(e) Use the bond enthalpies on page 316 and Equation [8.12] to estimate the ΔH for the following reaction:

$$2\ C_2H_6 \ + \ 7\ O_2 \ \rightarrow \ 4\ CO_2 \ + \ 6\ H_2O$$

(f) In general, as the number of bonds between two atoms increases, the bond grows _______________________ and _____________________. This trend is illustrated in Figure 8.17 for N–N single, double, and triple bonds.