CHEM 4 Worksheet #12

Chemical Bonding

Kevin Barnese

Name: _____ **Date:** _____

Part 1: Lewis Structures for Atoms and Ions

1. Draw the Lewis dot structure for each atom. Include the number of valence electrons.

a) Na Valence electrons: ____

b) O Valence electrons: ____

c) N Valence electrons: ____

d) Cl Valence electrons: ____

e) Ca Valence electrons: ____

f) C Valence electrons: ____

2. Draw the Lewis dot structure for each ion. Indicate which noble gas has the same Lewis structure.

a) Na^+ Same as: _____

b) Cl^- Same as: _____

c) Ca^{2+} Same as: _____

d) O^{2-} Same as: _____

Part 2: Identifying Bond Types

Use the following reference information for problems 3 and 4:

Element	EN	Element	EN	Element	EN
Li	1.0	N	3.0	Cl	3.0
Mg	1.2	O	3.5	Br	2.8
P	2.1	F	4.0	Sr	1.0
H	2.1				

ΔEN	Bond Type	Example
0–0.4	Pure covalent	Cl_2
0.4–2.0	Polar covalent	HF
2.0+	Ionic	NaCl

3. For the compound MgO, determine if it is ionic or covalent.

a) What is the electronegativity of Mg?

b) What is the electronegativity of O?

c) Calculate the electronegativity difference.

d) Is this bond ionic, polar covalent, or pure covalent?

4. Classify each compound as ionic or covalent:

a) LiBr

b) NH_3

c) SrF_2

d) PCl_3

Part 3: Lewis Structures for Ionic Compounds

5. Draw the Lewis structure for each ionic compound:

a) Li_2S

b) $AlBr_3$

c) MgO

d) K_3N

Part 4: Lewis Structures for Covalent Compounds

6. The following structures show only the bonds. Add all necessary lone pairs to complete the octets:

a) H–O–H

b) F–F

7. Add lone pairs to complete these structures:

a) O=C=O

b) N≡N

c) H–S–H

8. Draw the Lewis structure for each molecule.

a) CH_4

c) CCl_4

b) H_2O

d) NF_3

9. Draw the Lewis structure for each molecule. Show the total number of valence electrons.

a) O_2

c) HCN

b) C_2H_4 (H_2CCH_2)

d) N_2

10. Draw the Lewis structure for HNO_3 (H connects to an O). Show the total number of valence electrons. Total valence electrons: _____

Part 6: Lewis Structures for Polyatomic Ions

11. Draw the Lewis structure for each ion. Show the total number of electrons in the Lewis structure (remember to add electrons for negative charges and subtract for positive charges). Include resonance structures if necessary.

a) CO_3^{2-} Total electrons: _____

b) NO_3^- Total electrons: _____

c) SO_4^{2-} Total electrons:_____

d) NH_4^+ Total electrons: _____

Part 7: Exceptions to the Octet Rule

12. Draw the Lewis structure for BF_3 (boron only has 6 electrons).

13. Draw the Lewis structure for SF_6 (sulfur has 12 electrons — expanded octet).

14. Draw the Lewis structure for NO (odd number of electrons — has 1 unpaired electron).

Part 8: VSEPR Theory and Molecular Geometry

15. For each molecule, determine the electron geometry and molecular geometry:

a) CH_4 Lewis structure:

Electron geometry: _____ Molecular geometry: _____

b) H_2O Lewis structure:

Electron geometry: _____ Molecular geometry: _____

c) BF_3 Lewis structure:

Electron geometry: _____ Molecular geometry: _____

d) NH_3 Lewis structure:

Electron geometry: _____ Molecular geometry: _____

e) PCl_5 Lewis structure:

Electron geometry: _____ Molecular geometry: _____

f) SF_6 Lewis structure:

Electron geometry: _____ Molecular geometry: _____

Part 9: Bond Angles and VSEPR

16. Predict the bond angle in each molecule:

a) CH_4 (tetrahedral) Expected bond angle: _____

b) NH_3 (trigonal pyramidal) Expected bond angle: _____ Why is it slightly less than CH_4?

c) H_2O (bent) Expected bond angle: _____ Why is it even smaller than NH_3?

d) CO_2 (linear) Expected bond angle: _____

17. Compare BH_3 and NH_3:

a) What is the molecular geometry of BH_3?

b) What is the molecular geometry of NH_3?

c) Why are they different despite both having 3 hydrogen atoms?

Part 10: Molecular Polarity

18. Classify each molecule as polar or nonpolar:

a) H_2O

b) CCl_4

c) NH_3

d) BF_3

e) CH_3Cl

f) CO_2

Answers

1. a) Na: 1 valence electron b) O: 6 c) N: 5 d) Cl: 7 e) Ca: 2 f) C: 4 **2.** a) Na^+: same as Ne b) Cl^-: same as Ar c) Ca^{2+}: same as Ar d) O^{2-}: same as Ne **3.** a) 1.2 b) 3.5 c) 2.3 d) Ionic **4.** a) Ionic b) Covalent c) Ionic d) Covalent **5.** Appropriate ionic Lewis structures with brackets and charges **6–7.** Structures with all lone pairs added to satisfy octets **8.** a) CH_4: 8 e^- b) H_2O: 8 e^- c) CCl_4: 32 e^- d) NF_3: 26 e^-; appropriate Lewis structures with correct lone pairs **9.** a) O_2: 12 e^- (double bond) b) C_2H_4: 12 e^- (double bond) c) HCN: 10 e^- (triple bond) d) N_2: 10 e^- (triple bond) **10.** HNO_3: 24 e^-; H connected to O, double bond between N and one O, lone pairs on O atoms **11.** a) CO_3^{2-}: 24 e^-, 3 resonance structures b) NO_3^-: 24 e^-, 3 resonance structures c) SO_4^{2-}: 32 e^- d) NH_4^+: 8 e^- **12–14.** Appropriate structures showing octet rule exceptions (BF_3 6e^-, SF_6 12e^-, NO odd e^-) **15.** a) Tetrahedral/Tetrahedral b) Tetrahedral/Bent c) Trigonal planar/Trigonal planar d) Tetrahedral/Trigonal pyramidal e) Trigonal bipyramidal/Trigonal bipyramidal f) Octahedral/Octahedral **16.** a) 109.5° b) ∼107° (lone pair repulsion) c) ∼104.5° (two lone pairs) d) 180° **17.** a) Trigonal planar b) Trigonal pyramidal c) BH_3 has no lone pairs; NH_3 has one lone pair **18.** a) Polar b) Nonpolar c) Polar d) Nonpolar e) Polar f) Nonpolar