Chemical Bonding Worksheet #1 (Ionic Bonding) ANSWERS

1. . How many valance electrons do the atoms below have?

a. nitrogen atom — $5e^-$

b. sodium atom — $1e^-$

c. sulfur atom — $6e^-$

d. bromine atom — $7e^-$

2. Draw the Lewis valance electron dot structures for the following atoms:

a. nitrogen atom — $\cdot\ddot{N}\cdot$

b. sodium atom — $\dot{Na}$

c. sulfur atom — $\cdot\ddot{S}\cdot$

d. bromine atom — $\cdot\ddot{Br}:$

e. magnesium atom — $\dot{Mg}$

f. aluminum atom — $\dot{Al}\cdot$

g. carbon atom — $\cdot\dot{C}\cdot$

h. neon atom — $:\ddot{Ne}:$

3. Why do atoms become ions?

In order to obtain a stable octet.

4. How many valance electrons are needed in order for an atom or ion to be stale? (What are the "magic" numbers?) 2, 8, 8

5. Which type of atom will always lose valance electrons to become an ion? metals

6. Which type of atom will always gain valance electrons to become an ion? non-metals

7. When the following become IONS what will the charge be?

a. nitrogen — $[:\ddot{N}:]^{-3}$

b. sodium — $[Na]^{+1}$

c. sulfur — $[:\ddot{S}:]^{-2}$

d. bromine — $[:\ddot{Br}:]^{-1}$

e. magnesium — $[Mg]^{+2}$

f. aluminum — $[Al]^{+3}$

8. Define Ionic Bond:

9. Draw the electron dot structures showing the transfer of electrons that occurs when the following form ionic compounds:

a. calcium and bromine

$\dot{Ca} + \cdot\ddot{Br}: + \cdot\ddot{Br}: \rightarrow [Ca]^{+2}\,2[:\ddot{Br}:]^{-1}$

b. potassium and bromine

$\dot{K} + \cdot\ddot{Br}: \rightarrow [K]^{+1}[:\ddot{Br}:]^{-1}$

c. strontium and sulfur

$\dot{Sr} + \cdot\ddot{S}\cdot \rightarrow [Sr]^{+2}[:\ddot{S}:]^{-2}$

d. potassium and chlorine

$\dot{K} + \cdot\ddot{Cl}: \rightarrow [K]^{+1}[:\ddot{Cl}:]^{-1}$

e. magnesium and fluorine

$\dot{Mg} + \cdot\ddot{F}: \rightarrow [Mg]^{+2}\,2[:\ddot{F}:]^{-1}$

f. barium and iodine

$\dot{Ba} + \cdot\ddot{I}: + \cdot\ddot{I}: \rightarrow [Ba]^{+2}\,2[:\ddot{I}:]^{-}$

g. sodium and nitrogen

$\dot{Na} + \dot{Na} + \dot{Na} + \cdot\ddot{N}\cdot \rightarrow 3[Na]^{+1}[:\ddot{N}:]^{-3}$

h. rubidium and phosphorus

$\dot{Rb} + \dot{Rb} + \dot{Rb} + \cdot\ddot{P}\cdot \rightarrow 3[Rb]^{+1}[:\ddot{P}:]^{-3}$

10. What is the net charge on each of the above compounds? 0 **Why?**

Elements in compounds are then stable ∵ they now each have an e^- valence shell like a noble gas.

Chemical Bonding Worksheet #2 (Types of Bonds)

1. Determine the type of bond (ionic, polar covalent, or nonpolar covalent) that will form between the following atoms.

i pc npc

a. Ca and Cl
(1.00 3.16) → ionic
= 2.16

b. C and S
(2.55 2.58) → n.p.c
= 0.03

c. Mg and F
(1.31 3.98) → ionic
= 2.67

d. N and O
(3.04 3.44) → p.c.
= 0.40

e. H and O
(2.20 3.44) → p.c.
= 1.24

f. S and O
(2.58 3.44) → p.c.
= 0.86

g. Br and Cl
(2.96 3.16) → n.p.c
= 0.20

h. F and O
(3.98 3.44) → p.c.
= 0.54

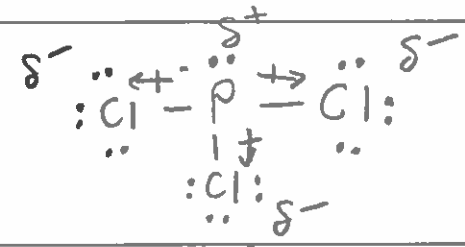
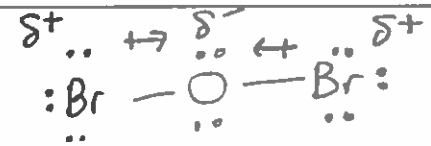
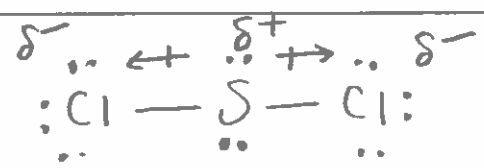
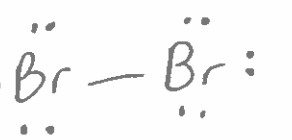
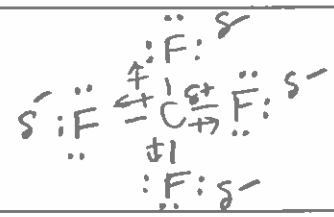
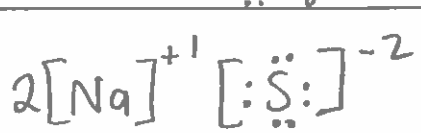
i. P and S
(2.19 2.58) → n.p.c
= 0.39

j. H and Cl
(2.20 3.16) → p.c.
= 0.94

k. C and H
(2.55 2.20) → n.p.c.
= 0.35

l. H and H
(2.20 2.20) pure covalent
= 0

2. Using the table of electronegativity values, calculate the EN difference for the atoms that bonded in the following table. Then state whether the bond is nonpolar covalent, polar covalent, or ionic. State which atom has the greater EN, and in your drawing show which atom is partially positive or partially negative and include any dipole moments or square brackets with charges where appropriate.

Formula	EN difference	Type of bond (ionic, polar covalent, non-polar covalent, pure covalent)	Drawing of molecule
PCl_3	Cl 3.16 P 2.19 0.97	polar covalent	
Br_2O	O 3.44 Br 2.96 0.48	polar covalent	
SCl_2	Cl 3.16 S 2.58 0.58	polar covalent	
Br_2	Br 2.96 Br 2.96 0	pure covalent	
CF_4	F 3.98 C 2.55 1.43	polar covalent	
Na_2S	S 2.58 Na 0.93 1.65	ionic (metal + non-metal)	
CaO	O 3.44 Ca 1.00 2.44	ionic