# Review #1 Practice Test/Chemical Bonding/AP Chemistry

1. Write the Lewis Dot Diagrams for the Following

A) $Li^+$ B) $Cl^-$ C) $S^{2-}$ D) $Sr^{2+}$ E) $I^-$

2. Define electronegativity and explain the difference between electronegativity and electron affinity. Describe the general trend for electronegativity on the Periodic Table.
3. List the following bonds in order of increasing ionic character.

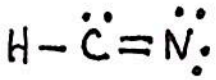
Cs to F, Cl to Cl, Br to Cl, Si to C

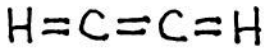
4. Write the Lewis structures for the following molecules:

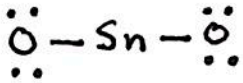
A) ICl B) $PH_3$ C) $H_2S$ D) $N_2H_4$ E) $HClO_3$

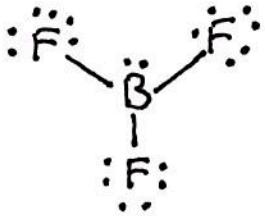
5. Explain the concept of formal charge and how it is calculated.

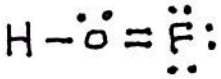
6. The following Lewis structures are incorrect. Explain what is wrong with each one and provide a correct Lewis structure.

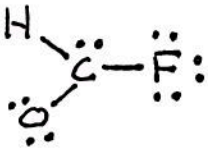
A) 

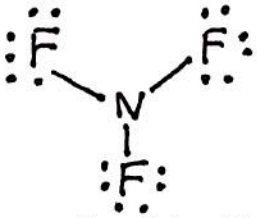
B) 

C) 

D) 

E) 

F) 

G) 

7. Draw three resonance structures for the chlorate ion. Show formal charges and explain which structure would be favored, based on formal charges and EXPLAIN why.

8. Which of the following are covalent compounds and which are ionic?

RbCl, $PF_5$, $BrF_3$, $KO_2$, $CI_4$

9. Explain why the bond energy of a molecule is usually defined in terms of a gas phase reaction. Why are bond breaking processes always endothermic and bond forming processes always exothermic?

10. Describe what is meant by the term Lattice Energy.

# AP Chem: Key For Chpt. 8 Review Sheet.

1) (A) $[Li]^+$ (B) $[:\ddot{\underset{..}{Cl}}:]^-$ (C) $[:\ddot{\underset{..}{S}}:]^{2-}$ (D) $[Sr]^{2+}$ (E) $[:\ddot{\underset{..}{I}}:]^-$

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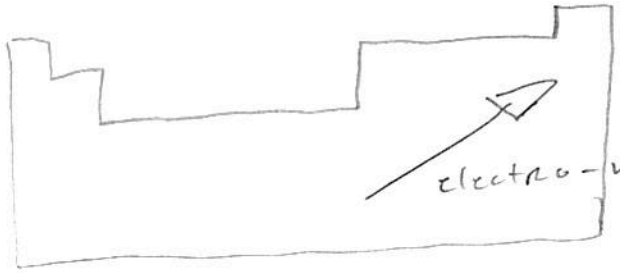
(2) electronegativity describes how evenly/unevenly electrons are shared between two atoms that are bonded together. Electronegativity values are unitless numbers between 0 and 4 that are used to describe the difference between each atom. A zero has no affinity for electrons and a 4 has the highest Affinity. Bond type can be ascribed by comparing the electronegativity values for each atom in the bond.

Ex: H-F the difference is 1.9.
2.1 4.0

- Differences from 0 - 0.4 are non-polar covalent
- values greater than 0.4, but less than 2.0 are polar covalent
- values of 2.0 and above are ionic.

Electron affinity is much different than electronegativity. Electron A. describes the stability that an atom gains or loses by accepting an $e^-$.




when atoms become stable upon accepting an electron they lose energy. Atoms w/ high E. Affinity therefore release heat in KJ/mol (larger - values / - means leaving the exothermic system). Atoms that are unstable upon accepting an electron are small + values (endothermic). Energy must be continuously supplied to the system in order for the $e^-$ to remain in the atom's domain.

(3) Cs–F – 3.3 (0.7 – 4.0)  Cl–Cl – 0 (3.0 – 3.0)  Br–Cl – 0.2 (2.8 – 3.0)

Si–C (1.8 – 2.5) 0.7

most ionic: Cs–F, Si–C, Br–Cl, Cl–Cl : least ionic

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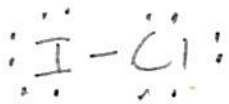
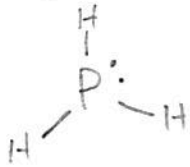
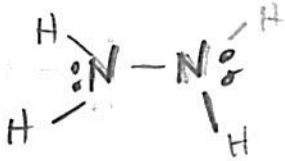
(4) (A) ICl (14e⁻)  (B) $PH_3$ (8e⁻)  (C) $H_2S$ (8e⁻)

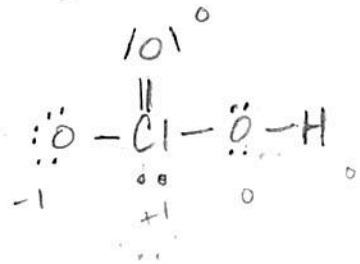




(D) $N_2H_4$ (14e⁻)



(E) $HClO_3$ (26e⁻)  18 +7 +1 = 26e⁻



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(5) Formal charge is used to describe the charge that remains on each atom in a molecule after the non-bonding and bonding electrons have been removed from the original valence electrons for that atom.

Ex:

$$\text{valence e's} - \text{non bonded e's} - \frac{1}{2}(\text{bonded e's}) = \text{f.c.}$$

Formal charge should be low (ideally zeros or ±1) in a properly represented Lewis Structure. If there are multiple Lewis structures, f.c's can be used to choose the most likely candidate. negative values should ideally correspond w/ the most electro-negative atoms because they want electrons!

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(A) H − C ≡ N:

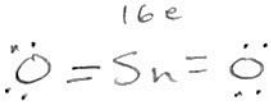
HCN has 10e⁻. The review sheet shows 12e⁻. Also avoid free pairs of e's on carbon

(B) $HC_2H$ 10e⁻

H − C ≡ C − H

the review sheet shows 12e⁻ and double bonded H's (this is not possible)

(C) 16e⁻



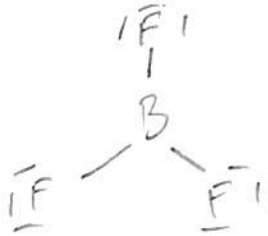
$SnO_2$ has 16e's

the review sheet shows 12e⁻!

(D) $BF_3$ has 24e⁻

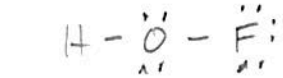
the review sheet shows 26e⁻

Boron can have only 3 bonds (it is an exception)



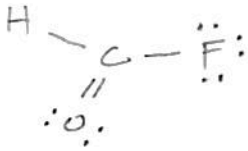
(E) HOF 14e⁻

the review sheet shows 14e⁻ (correct) but Fluorine has too many e's associated w/it (more than 8)



(F) HCOF has 18e⁻

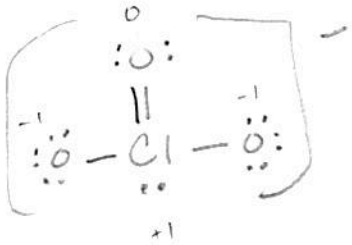
the review sheet shows 18e⁻, but carbon has a free pair of e's (avoid)



(G) $NF_3$ has 26e⁻

the review sheet shows 24e⁻ (add lone pair)

(7)



7
+3(6)
+1(m)
26e's

F.C's
Cl = 2
All O's = −1

the oxygens should be most negative (w/respect to formal charges) because they are most electronegative (want e's)

(8) Rb − Cl −2.2 ionic
0.8 3.0

PF5 P − F −1.9
2.1 4.0
strongly polar covalent

$BrF_3$
Br − F −1.2
2.8 4.0
polar-covalent

$KO_2$
K − O −2.7 ionic
0.8 3.5

$CI_4$
C − I nonpolar covalent.
2.5 2.5

(9) · Bond E is described with gas phase values b/c it avoids all of the energy Δ's that accompany Δ's of phase (ie. solid → liquid → gas, etc). This way the bond energies are easier to relate to each other.

- Bonds are stable. To break them energy needs to be supplied to the system (endothermic) until the bond is overcome
- Bond formation is stable (especially if spontaneous) when bonds are formed the energy of instability ("stress") is relieved as energy leaving the system (exothermic). Think of electron affinity. Atoms that "want" electrons have high negative values (exothermic) because they are more stable once they get an $e^-$.

(10) <u>Lattice Energy</u>: the energy required to completely separate 1.0 mol of solid compound into its gaseous ions. This energy is a sum of all thermochemical steps involved to achieve this transition from solid → gaseous ions.