Chem 11 Practice Questions for Ch. 8

1. Atoms having equal or nearly equal electronegativities are expected to form
   A) no bonds
   B) polar covalent bonds
   C) nonpolar covalent bonds
   D) ionic bonds
   E) covalent bonds

2. For the elements Cs, F, and Cl, the order of increasing electronegativity is:
   A) F < Cl < Cs
   B) Cs < Cl < F
   C) Cl < Cs < F
   D) F < Cs < Cl
   E) none of these

3. Based on electronegativity differences, which of the following is most likely to be ionic?
   A) CaF_2
   B) Br_2
   C) BH_3
   D) NO
   E) CF_4

4. Which of the following bonds would be the most polar without being considered ionic?
   A) Mg-O
   B) C-O
   C) O-O
   D) Si-O
   E) N-O

5. Which of these is an isoelectronic series?
   A) Na^+, K^+, Rb^+, Cs^+
   B) K^+, Ca^{2+}, Ar, S^{2-}
   C) Na^+, Mg^{2+}, S^{2-}, Cl^-
   D) Li, Be, B, C
   E) none of these (A-D)

6. Using the following bond energies:

Bond	Bond Energy (kJ/mol)
C≡C	839
C–H	413
O=O	495
C=O	799
O–H	467

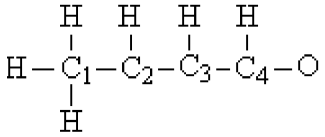
estimate the heat of combustion for one mole of acetylene:

$C_2H_2(g) + {}^{5}\!/_{2}\,O_2(g) \rightarrow 2CO_2(g) + H_2O(g)$

A) 1228 kJ
B) –1228 kJ
C) –447 kJ
D) +447 kJ
E) +365 kJ

7. In the Lewis structure for elemental nitrogen there is (are)
A) a single bond between the nitrogens
B) a double bond between the nitrogens
C) a triple bond between the nitrogens
D) three unpaired electrons
E) none of the above

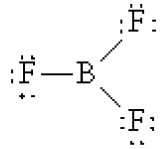
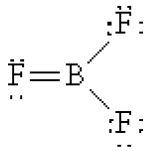
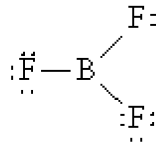
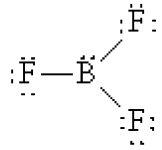
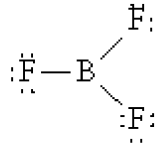
Consider the compound crotonaldehyde, whose skeleton is:



8. How many electrons must be shown (as bonding or nonbonding electrons) in the Lewis structure of this molecule?
A) 12
B) 18
C) 24
D) 28
E) 32

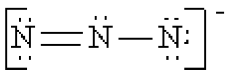
9. How many nonbonding electrons appear in the Lewis structure of this molecule?
A) 2
B) 4
C) 6
D) 8
E) 10

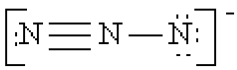
10. Which carbon in this molecule has tetrahedral bonding?
    A) 1
    B) 2
    C) 3
    D) 4
    E) all

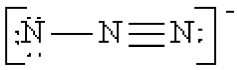
11. Which of the following Lewis structures best describes BF_3?
    A) 
    B) 
    C) 
    D) 
    E) 

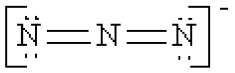
12. Which of the following has an incomplete octet in its Lewis structure?
    A) SO_2
    B) ICl
    C) CO_2
    D) F_2
    E) NO

13. Which of the following is not a valid resonance structure for N_3^-?

A) 

B) 

C) 

D) 

E) all are correct

14. Select the molecule from the following that has a dipole moment.
A) CO_2
B) SeO_3
C) XeF_4
D) SF_4
E) $BeCl_2$

15. If a compound has a number of individual dipoles, then:
I. It is polar overall.
II. There is an electronegativity difference between the bonded atoms.
III. it is ionic.
IV. It doesn't have resonance.

A) II only
B) II, IV
C) I, II, IV
D) I, III
E) All of the above statements are correct.

16. The Cl–Kr–Cl bond angle in $KrCl_4$ is closest to
A) 90°
B) 109°
C) 120°
D) 150°
E) 360°

17. The bond angles about the carbon atom in the formaldehyde molecule, $H_2C=O$, are about:
    A) 120°
    B) 60°
    C) 109°
    D) 180°
    E) 90°

18. Which of the following species has a trigonal bipyramid structure?
    A) NH_3
    B) IF_5
    C) I_3^-
    D) PCl_5
    E) none of these

19. Which ion is planar?
    A) NH_4^+
    B) CO_3^{2-}
    C) SO_3^{2-}
    D) ClO_3^-
    E) all are planar

20. Which ion is larger in each pair? i) O^{2-} or S^{2-} ii) Fe^{2+} or Fe^{3+} iii) S^{2-} or K^+
    A) S^{2-}, Fe^{2+}, S^{2-}
    B) S^{2-}, Fe^{3+}, S^{2-}
    C) O^{2-}, Fe^{3+}, K^+
    D) S^{2-}, Fe^{2+}, K^+
    E) O^{2-}, Fe^{2+}, S^{2-}

21. Calculate the lattice energy of the ionic compound MCl_2 given the information below:

    $\Delta H°_f$ $MCl_2(s)$ = -342 kJ/mol
    IE_1 of M = +600 kJ/mol
    IE_2 of M = +1150 kJ/mol
    $Cl_2(g) \Rightarrow 2\ Cl(g)$ ΔH = +244 kJ
    $Cl(g) + 1\ e^- \Rightarrow Cl^-(g)$ ΔH = -349 kJ
    $M(s) \Rightarrow M(g)$ ΔH = +150 kJ
    A) -2136 kJ
    B) -2015 kJ
    C) -1446 kJ
    D) -1788 kJ
    E) -1666 kJ

Answer Section

1. ANS: C
2. ANS: B
3. ANS: A
4. ANS: D
5. ANS: B
6. ANS: B
7. ANS: C
8. ANS: D
9. ANS: B
10. ANS: A
11. ANS: A
12. ANS: E
13. ANS: A

14. ANS: D

15. ANS: A

16. ANS: A

17. ANS: A

18. ANS: D

19. ANS: B

20. ANS: A

21. ANS: D

Solutions

(1) If the electronegativity difference is negligible, the atoms attract the bonding electrons equally and cannot develop a surplus or deficit of electrons. The bond is nonpolar. Of course when we talk about the polarity of a bond, we are talking about a covalent bond. For an ionic bond, the polarization is complete and we have two ions with fully developed charges. A non-polar bond is necessarily a non-polar covalent bond.

(2) Electronegativity increases as we go up and to the right in the periodic table. Thus, Cs is the least electronegative, followed by Cl, and F is the most electronegative.

$$Cs < Cl < F$$

(3) Metals have low electronegativity, and except for special cases, when we see a binary compound of a metal and a nonmetal we can assume that their bonds are ionic, and that the compound is an ionic compound.

(4) Nonmetal-nonmetal bonds are covalent, and metal-nonmetal bonds are ionic. Mg-O bond is ionic. All of the bonds involve oxygen, so we focus on the other atom. Since O is quite electronegative, we are looking for the least electronegative "other" atom for maximum bond polarity (except Mg, which is a metal). Si is the least electronegative among C, O, Si, N, and is a nonmetal (metalloid, but not a metal). Si-O bond is the most polar of the covalent bonds listed.

(5.) "Isoelectronic" means "same electronic configuration". We can eliminate a series of neutral atoms like Li, Be, B, C because they are different elements that must have different configurations. We are looking for anions of elements right before a noble gas and cations of elements right after a noble gas in the periodic table, having gained or lost just enough electrons to make their configurations the same as the nearby noble gas. K^+, Ca^{2+}, Ar, S^{2-} fits the bill.

(6.) We need to know the bonding of C_2H_2, which means we need to know its Lewis structure.

* no. of valence e^- = (2)(4) + (2)(1) = 10

* skeletal structure: H – C ⁝ C – H

$$H-C\equiv C-H$$

We also technically need to know if O_2 is $:\ddot{O}-\ddot{O}:$ or $:\ddot{O}=\ddot{O}:$, but O=O bond is the only one listed, so obviously it is the latter. Likewise, only C=O bond is given, suggesting that CO_2 is O=C=O. Likewise, H_2O is H–O–H. So we write

$$H-C\equiv C-H + \frac{5}{2}\,O=O \rightarrow 2\ O=C=O + H-O-H$$

(C–H and C≡C–H bonds broken — not seen in products; O=C=O and H–O–H formed — not seen in reactants)

$$\Delta H = (\text{Sum of bond energies})_{\text{broken}} - (\text{Sum of bond energies})_{\text{formed}}$$

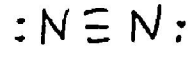
$$\Delta H = [839 + (2)(413) + \tfrac{5}{2}(495)] - [(2)(2)(799) + (2)(467)] = -1228\ \text{kJ}$$

(2 C–H bonds per H–C≡C–H; 2 moles; 2 C=O bonds per O=C=O; 2 O–H bonds per H–O–H)

(7.)

# of valence e^- = (2)(5) = 10

skeletal structure: 

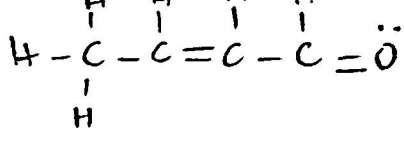


(8.) The subscripts on carbons are just labels. Normally we would use those labels to distinguish them in our Lewis structure calculations. Don't be tempted to go through the whole procedure! The question is asking how many electrons would be in bonding or non-bonding pairs. That leaves nothing out! However many valence e^- are brought by the atoms will be either in a bond or a lone pair. Just count the no. of valence electrons!

6(1) + (4)(4) + (1)(6) = 28 e^-

6H 4C 1 Oxygen

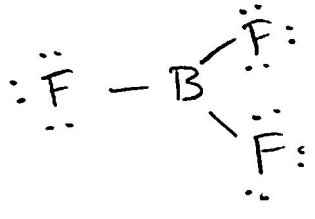
(9.) Carbon very rarely has any lone pairs in molecules. (:C≡O: is an exception). If it has three neighbors, that means it has a double bond with one of them (and that can't be H because it can't make more than one bond). Oxygen normally makes 2 bonds. If it has only one neighbor, it either has a double bond or is negatively charged. But this is a neutral molecule. So we can guess:



where only O has its usual 2 lone pairs, which corresponds to 4 nonbonding electrons

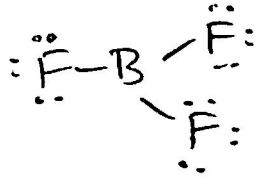
(10) The carbon with 4 bonds (i.e. 4 e^- groups) has tetrahedral bonding.

(11) Boron violates the octet rule by needing only 6 e^- in a molecule. Fluorine conforms to the octet rule (always). It makes one bond to satisfy the octet rule, which means it always has 3 lone pairs when bonded. Eliminate any structures where B has more than 6 e^- (3 bonds) or F has more than one bond or less than 3 lone pairs. So we have



The long way to do it:

* no of valence e^- = (3) + (3)(7) = 24
* draw the skeleton



remember B needs only 6 e^-

no. of e^- used = (12)(2) = 24 ✓

done

Well, maybe it wasn't that much longer!

(12) NO molecule has an <u>odd number</u> of <u>electrons</u> because N brings 5 e^- while O brings 6 e^-. We can <u>not</u> satisfy the octet rule when we start with an odd no. of valence e^-.

no. of val. e^- = 5 + 6 = 11

(13) N_3^- has $(3)(5)+1 = 16$ valence e^-

$\ddot{N}=\ddot{N}-\ddot{N}:$ has $(9)(2) = 18$ valence e^-, so it is not valid.

The other structures are valid because they have 16 e^-

(14) We need to determine the molecular geometries and see if the bond dipoles cancel. If the molecule is symmetric, they cancel. If not, they don't.

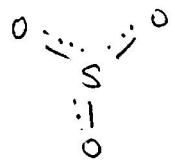
CO_2: $\ddot{O}=C=\ddot{O}$ symmetric, no dipole moment

SeO_3: $:\ddot{O}-\ddot{Se}-\ddot{O}:$ (with $:\ddot{O}:$ below Se)

no. of val. e = (4)(6) = 24

no. of e^- used = (10+3)(2) = 26

Form a double bond

$\ddot{O}=S-\ddot{O}:$ (with $:O:$ below S) which actually is 

No lone pairs on Se, 3 bonds, trigonal planar

$\Rightarrow$ symetric $\Rightarrow$ no dipole

XeF_4 no. of lone pairs on Xe = $\frac{1}{2}(8-4-0) = 2$

4 bonds + 2 lone pairs $\rightarrow$ 6 e^- groups $\rightarrow$ octahedral e^- geometry

Two lone pairs on octahedral e^- geometry $\Rightarrow$ square planar

$\Rightarrow$ symetric $\Rightarrow$ no dipole

$BeCl_2$ Be needs only 2 pairs of e^-: $:\ddot{Cl}-Be-\ddot{Cl}:$

Linear, symmetric $\Rightarrow$ no dipole

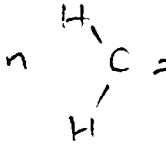
It must be SF_4

(16) no. of lone pairs $= \frac{1}{2}(8-4-0) = 2$

no. of e^- groups = 4 bonds + 2 lone pairs = 6 $\Rightarrow$ octahedral e^- geometry

2 bonds missing from octahedral e^- geometry $\Rightarrow$ square planar

Cl – Kr – Cl angle = 90°

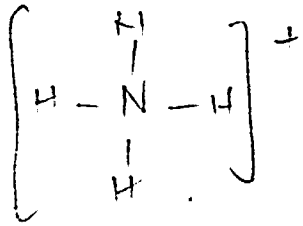
(17) C in  has 3 e^- groups $\Rightarrow$ trigonal planar

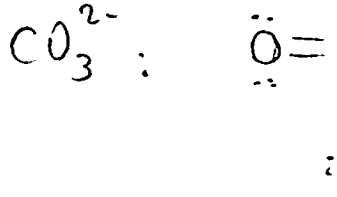
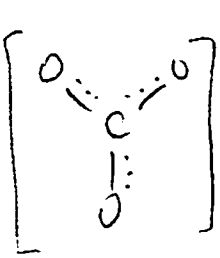
$\Rightarrow$ bond angles = 120°

(18) For a trigonal bipyramidal structure (not just e^- geometry, but the molecular geometry) we need 5 bonds around the central atom. That leaves only IF_5 and PCl_5 as viable candidates. Furthermore, we need the e^- geometry to be the same as molecular geometry, with no lone pairs. Trigonal bipyramidal molecular structure has no lone pairs; it uses all 5 e^- groups for bonding.

IF_5: no. of lone pairs on I $= \frac{1}{2}(7-5-0) = 1$ not zero

PCl_5: no. of lone pairs on P $= \frac{1}{2}(5-5-0) = 0 \Rightarrow$ trigonal bipyramidal

(19) NH_4^+:  tetrahedral

CO_3^{2-}:  actually  trigonal planar

Since choice (A) is not planar, choice (E) can't be true and since we already have a planar choice, we don't worry about SO_3^{2-} or ClO_3^-

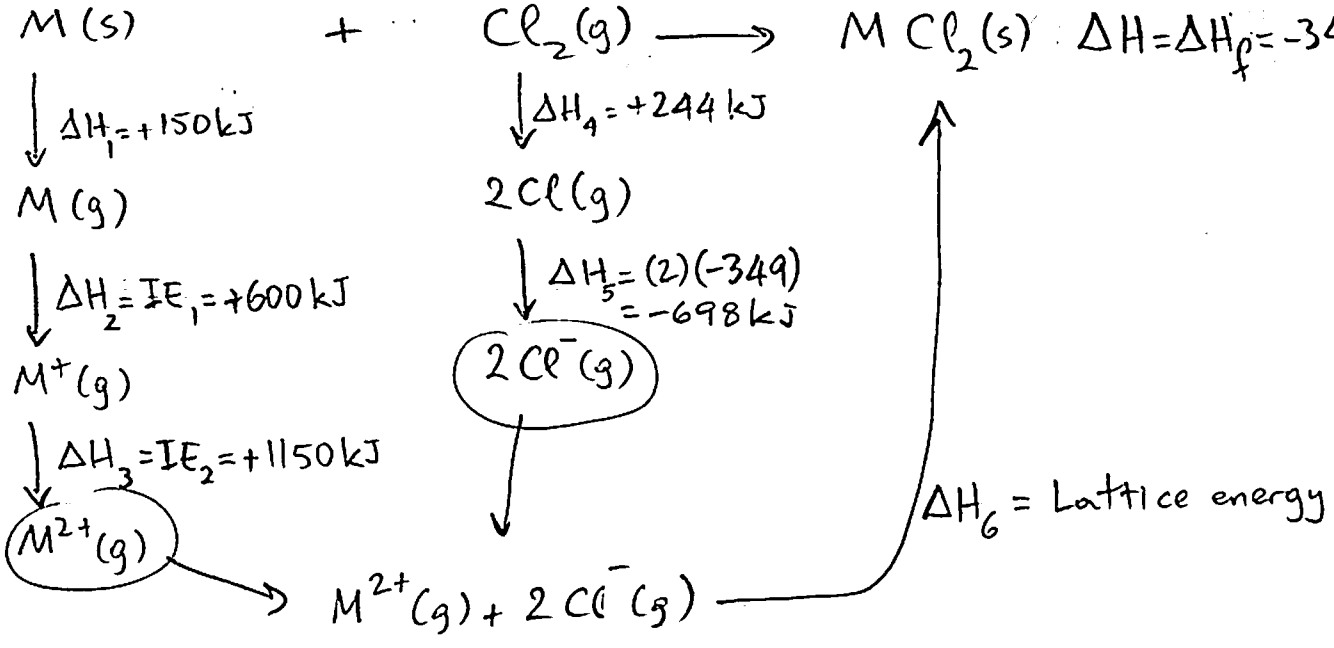
(20.) S is below O in the same group. For the same charge, it is larger.

Fe^{2+} has one more e^- than Fe^{3+}, so there is more e^- repulsion (and more nuclear shielding), so it's larger.

S^{2-} and K^+ are isoelectronic, but K^+ has 3 more protons, so it is smaller than S^{2-}.

Correct answer: S^{2-}, Fe^{2+}, S^{2-}

(21.)



$\Delta H_1 + \Delta H_2 + \Delta H_3 + \Delta H_4 + \Delta H_5 + (\text{Lattice energy}) = \Delta H_f$

Lattice energy $= (-342) - 150 - 600 - 1150 - 244 - (-698) = -1788$ kJ