AP Chemistry
Kotz - Chapter 9 Practice Problem Answers

1.

a. $[Ar]3d^3$ The first two electrons come out of the 4s sublevel then one out of the 3d
b. $[Ne]3s^2 3p^6$ noble gas configuration
c. $[Ar]3d^{10}$ The two electrons come out of the 4s sublevel (highest energy level first)
d. $[Kr]4d^{10}5s^2 5p^6$ noble gas configuration
e. $[Ar]3d^{10}4s_2 4p^6$ noble gas configuration
f. $[Ar]3d^{10}$ Remember the normal configuration for Cu is $[Ar]4s^1 3d^{10}$, it's one of the ones that migrate an electron from the 4s sublevel to complete the higher energy 3d sublevel. The first electron out for a cation is the lone electron in the 4s.

2.

a. KI potassium iodide

$K^+ + [:\ddot{\underset{..}{I}}:]^-$

b. barium fluoride BaF_2

$Ba^{2+} + 2[:\ddot{\underset{..}{F}}:]^-$

c. Rb_2S rubidium sulfide

$2Rb^+ + [:\ddot{\underset{..}{S}}:]^{2-}$

d. aluminum oxide, Al_2O_3

$2Al^{3+} + 3[:\ddot{\underset{..}{O}}:]^{2-}$

These are all examples of ionic compounds (metal/nonmetal), so the electronegativity difference is large enough that the electrons are completely transferred from the metal to the nonmetal. The metal loses electrons exposing the filled lower energy s and p sublevels, and the nonmetal completes the octet of its current s and p sublevel. Note that even though no dots are shown around the cation, it is considered to have a full octet of electrons (from the next lower energy level as stated above).

3.

Lattice energy for sodium fluoride is -914kJ/mol. Determine the enthalpy of formation for $NaF_{(s)}$

Endothermic steps:

a. Enthalpy of sublimation for Na = **+107kJ/mol**
b. Dissociation of ½ mole of F_2 to produce one mole of F atoms = ½ the bond dissociation energy = ½ (+159kJ/mol) = **+79.5kJ/mol**
c. Conversion of Na to Na^+ through the first ionization energy of sodium is **+496kJ/mol**

Exothermic steps:

a. Conversion of F to F^- through the electron affinity of fluorine is **-328kJ/mol**
b. Lattice energy given in the problem for sodium fluoride is **-914kJ/mol**

Adding everything up you get **ΔH^o_f = -560kJ/mol** (Appendix value = -573.7kJ/mol)

Remember enthalpy of formation is for one mole of that substance from its elements in their standard states.

4. In this problem, the missing piece of information is the lattice energy of cesium bromide $CsBr_{(s)}$
We will add the positive value for the enthalpy of formation to the other endothermic values, subtract the electron affinity for the bromine and we should end up with the lattice energy. Note that in this problem we also have the added enthalpy of turning liquid bromine into gaseous bromine.
Endothermic steps:

a. Enthalpy of sublimation for Cs = **+76.1kJ/mol**

b. Enthalpy of vaporization of liquid bromine is 30.9kJ/mol (of which we need ½ a mole) = **15.45kJ/mol**

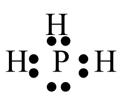
c. Dissociation of ½ mole of Br_2 to produce one mole of Br atoms = ½ the bond dissociation energy = ½ (+193kJ/mol) = **+96.5kJ/mol**

d. Conversion of Cs to Cs^+ through the first ionization energy of cesium is **+376kJ/mol**

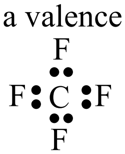
e. Now we'll add to all of this the positive value for the enthalpy of formation **+405.9kJ/mol**. This total should equal the sum of the lattice energy + the electron affinity of bromine.
Total = **970kJ/mol**

Adding the electron affinity for bromine = **-325kJ/mol** we get
645kJ/mol, but since this it is exothermic our final answer is **-645kJ/mol**, which is the lattice energy of cesium bromide.

5. a. Since the valence of phosphorus is 5, it is three electrons short of an octet. Therefore we expect that it will want to bond with three other hydrogens, each contributing and electron

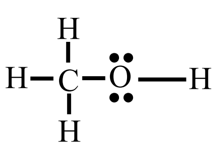


b. Carbon has a valence of 4, so it will bond with four other fluorines to complete the octet.

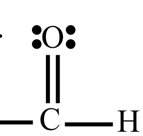


6. Various ways can be used to determine the Lewis structures for these molecules. Things to keep in mind: The least EN atom is *usually* the central atom; hydrogen can never be a central atom; formal charge can be used to determine the best possible arrangement; knowledge of organic functional groups can help here as well.

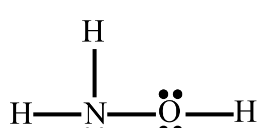
a.



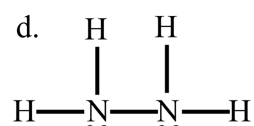
b.



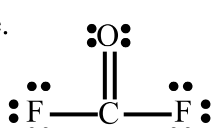
c.



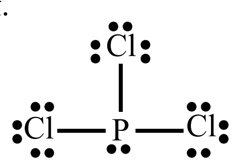
d.



e.



f.



7. a. B<N<F
b. Ca<As<Br
c. Ga<C<O

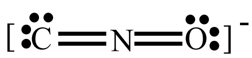
In each case, electronegativity increases to the right and up on the periodic table. This means that the atoms with the most nonmetallic character are in the upper right, and those with the most ionic character are in the lower left.

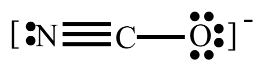
8.

a. F-F (no δ) < F-Cl < F-Br < F-I < F-H (δ^- towards fluoride except for F-F. Remember the F-H bond is responsible for one type of hydrogen bonding in intermolecular attractions.) The bigger the difference in electronegativity the more polar the bond.

b. H-H (no δ) < H-I < H-Br < H-Cl < H-F (δ^+ towards hydrogen except for H-H)

9.

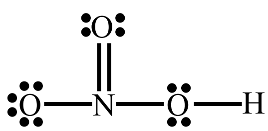
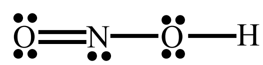
a. 

b. 

Formal charge for

a.	b.
C = 4 - 4 - 2 = -2	N = 5 - 2 - 3 = 0
N = 5 - 0 - 4 = +1	C = 4 - 0 - 4 = 0
O = 6 - 4 - 2 = 0	O = 6 - 6 - 1 = -1

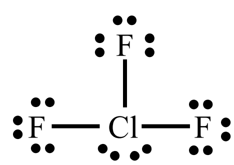
Structure b is preferred because the numbers are lower and the negative number is on the more electronegative atom.

10. Nitric acid is  and nitrous acid is 

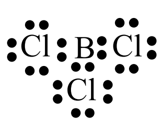
In nitric acid, there is a resonance between the two oxygens that are bonded to the nitrogen but not the hydrogen. Often, oxygen won't make more than two bonds so that making a double bond between the nitrogen and oxygen for the oxygen that is also bonded to the hydrogen isn't much of a factor since this would create three bonds the oxygen. For the same reason, the other resonance form for nitrous acid is also not much of a factor. Therefore resonance is more important for the nitric acid.

11. a. 

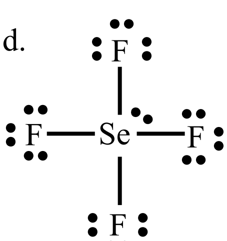
11e^- total so nitrogen has a single electron and is a free radical

b. 

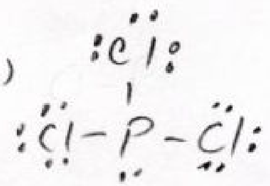
28e^- total so Cl must "see" 10 electrons Cl has an expanded valence.

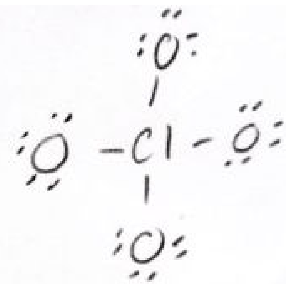


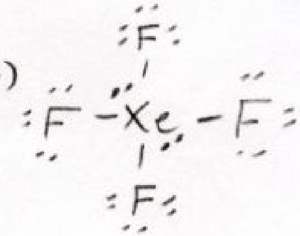
c. 24e^- total. Boron has an incomplete octet.

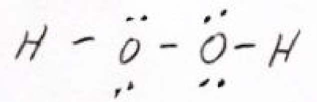
d. 

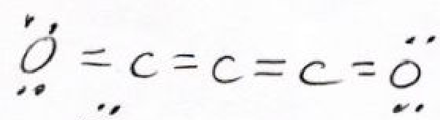
34e^- total so Se "sees" 10 electrons and has an expanded valence.

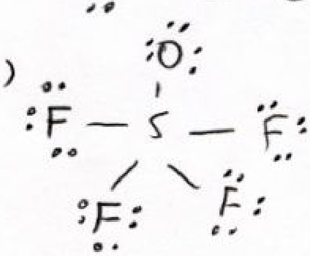
12. a)  trigonal Pyramidal

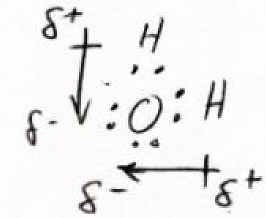
b)  tetrahedral

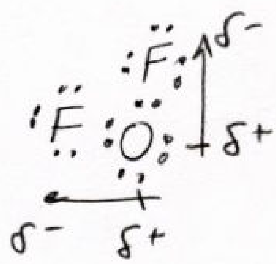
c)  Square planar

13. a)  bent on either O atom

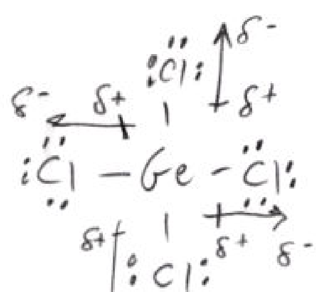
b)  linear sp hybridization

c)  trigonal bipyramidal

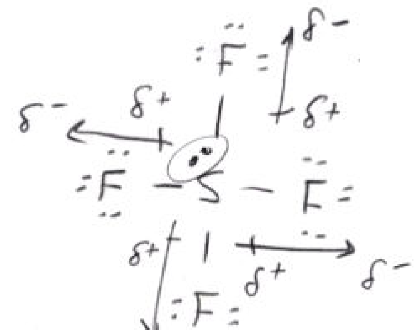
14. 



H_2O is more polar as ΔEN is bigger between H–O than H–F

15. 

Symmetrical, No net dipole



molecular geometry = See Saw
the nonbonding pair make this unsymmetrical so it has a net dipole.

16.

a. I - Cl we expect the bond length to be about half of an I - I bond and half of a Cl - Cl bond.

½ (266pm) + ½ (199pm) = **232.5pm**

b. C - F

½ (154pm) + ½ (143pm) = **148.5pm**

We expect that these bonds will probably be shorter than this because electronegativity differences predict that these bonds will be polar. Polar bonds tend to pull the atoms closer together.

17. Estimate the enthalpy change for

$H_{2(g)} + F_{2(g)} \rightarrow 2HF_{(g)}$

Using the formula that $\Delta H = \Sigma BE_{(reactants)} - \Sigma BE_{(products)}$

Where sigma is the sum of all the bond energies for the compounds involved.

H_2 = 436kJ/mol

F_2 = 159kJ/mol

2HF = 2(565kJ/mol) = 1130kJ/mol

436 + 159 - 1130 = **-535kJ/mol**