Electronegativity & Molecular Polarity Homework Answers

Mrs. Giovannone

p.221 # 1, 3, 5, 6, 7, 10

1. (a) $EN_C = 2.5$, $EN_N = 3.0$, $EN_O = 3.5$ The elements, listed in order of increasing electronegativity, are C, N, O. **(b)** $EN_S = 2.5$, $EN_{Se} = 2.4$, $EN_{Cl} = 3.0$ The elements, listed in order of increasing electronegativity, are Se, S, Cl. **(c)** $EN_{Si} = 1.8$, $EN_{Ge} = 1.8$, $EN_{Sn} = 1.8$ All 3 elements have the same electronegativity. **(d)** $EN_{Tl} = 1.8$, $EN_S = 2.5$, $EN_{Ge} = 1.8$ The electronegativity of sulfur is higher than that of thallium and germanium, which both have the same electronegativity.

3. (a) The C–O bond is polar covalent because the electronegativity of the elements increases rapidly as you move across Period 2, so the electronegativity difference between carbon and oxygen will be greater than 0.5.

(b) The F–I bond is polar covalent because the electronegativity of the halogens decreases rapidly as you move down Group 17, so the electronegativity difference between fluorine and iodine will be greater than 0.5.

(c) The Li–F bond is ionic because lithium has a low electronegativity and fluorine has the highest, so the electronegativity difference will be greater than 1.7.

(d) The Ge–Sn bond is non-polar covalent because these two elements are right beside one another in Group 14, in Periods 4 and 5, and after Period 2, the negativity of elements decreases or increases only gradually as you move down a group, so the electronegativity difference will be less than 0.5.

(e) The Al–Cl bond will be polar covalent because these two elements are in the same period, and electronegativity increases quickly as you move right from Group 13 (aluminum) to Group 17 (chlorine).

5. The bonds K–Br, C–Br, and Br–Br differ as follows: K has a low electronegativity (EN = 0.8) and creates an ionic bond with the highly electronegative Br (EN = 2.8). The electronegativity of C (EN = 2.5) is closer to the electronegativity of Br, and C creates a polar covalent bond with Br. Since the elements in the Br–Br bond are identical, the electronegativity difference between the bromine atoms is zero, resulting in a non-polar covalent bond.

6. (a) $O\delta^+$, $F\delta^-$ **(b)** non-polar **(c)** $C\delta^+$, $Br\delta^-$ **(d)** $C\delta^+$, $O\delta^-$

7. (a) The bond polarity $\delta^+H–Br\delta^-$ is correct.

(b) The bond polarity is shown incorrectly; the correct bond polarity is $\delta^+I–Cl\delta^-$.

(c) The bond polarity $\delta^+Si–S\delta^-$ is correct.

(d) The bond polarity is shown incorrectly; the F–F bond is non-polar.

(e) The bond polarity is shown incorrectly; the correct bond polarity is $\delta^+P–O\delta^-$.

10. (a) $\Delta EN_{Cl\text{-}Rb} = 3.0 - 0.8 = 2.2$

Since $\Delta EN > 1.7$, the bond between Rb and Cl is ionic.

(b) The bonding atoms, S and S, are identical, so they have the same electronegativity values. The bond between these two atoms is therefore non-polar covalent.

(c) $\Delta EN_{F–C} = EN_F - EN_C = 4.0 - 2.5 = 1.5$

Since $0.5 < \Delta EN < 1.7$, the bond between C and F is polar covalent.

(d) $\Delta EN_{C-B\alpha} = 2.5 - 0.9 = 1.6$

Since $0.5 < \Delta EN < 1.7$, the bond between Ba and C is polar covalent.

(e) $\Delta EN_{Se-B} = 2.4 - 2.0 = 0.4$

Since $\Delta EN < 0.5$, the bond between B and Se is non-polar covalent.

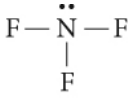
(f) $\Delta EN_{Br-Cs} = 2.8 - 0.7 = 2.1$

Since $\Delta EN > 1.7$, the bond is ionic.

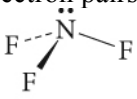
practice pg. 227# 1-2ab

1. (a) Polarity of NF_3 molecule:

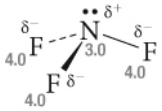
Draw the simplified Lewis structure of nitrogen trifluoride.



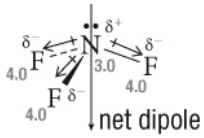
Use VSEPR theory to predict the three-dimensional structure, then draw a diagram. Since nitrogen is the central atom and it is surrounded by 4 electron pairs (3 bonding pairs and 1 lone pair), a molecule of nitrogen trifluoride will have a trigonal pyramidal shape.



Identify the electronegativity of each atom, and determine the partial charges in the molecule. N = 3.0, +; F = 4.0, –

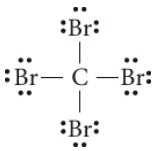


Draw the bond dipoles, and determine whether the molecule has a net dipole. The bond dipoles do not cancel each other out. Therefore, nitrogen trifluoride is a polar molecule with a net dipole pointing away from the lone pair on the nitrogen atom, in the direction of the fluorine atoms.

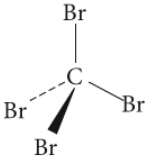


(b) Polarity of CBr_4 molecule:

Draw the simplified Lewis structure of tetrabromomethane.

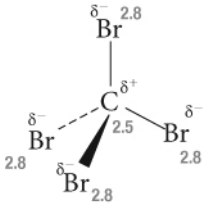


Use VSEPR theory to predict the three-dimensional structure, then draw a diagram.

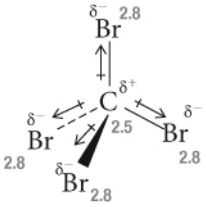


Since carbon is the central atom and it is bonded to 4 bromine atoms, a molecule of tetrabromomethane will have a tetrahedral shape.

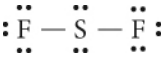
Identify the electronegativity of each atom, and determine the partial charges in the molecule. C = 2.5, +; Br = 2.8, –

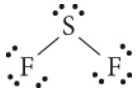


Draw the bond dipoles, and determine whether the molecule has a net dipole. In this tetrahedral shape, all of the atoms bonded to the central carbon atom are identical and all the directions are equivalent. Therefore, the 4 bond dipoles cancel each other. The overall net dipole is zero, so tetrabromomethane is a non- polar molecule.

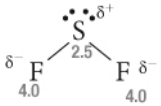


(c) Polarity of SF_2 molecule:

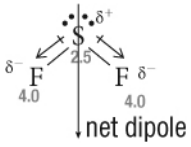
Draw the simplified Lewis structure of sulfur difluoride. 

Use VSEPR theory to predict the three-dimensional structure, then draw a diagram. The central sulfur atom is surrounded by 4 pairs of electrons (2 bonding pairs and 2 lone pairs), so SF_2 has a bent structure. 

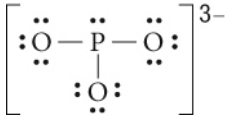
Identify the electronegativity of each atom, and determine the partial charges in the molecule. S = 2.5, +; F = 4.0, –



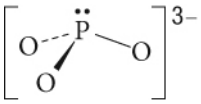
Draw the bond dipoles, and determine whether the molecule has a net dipole. In this bent shape, the bond dipoles do not cancel each other out. Therefore, sulfur difluoride is a polar molecule with a net dipole pointing away from the lone pairs on the nitrogen atom, in the direction of the fluorine atoms.



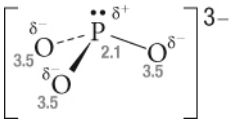
2. (a) Polarity of PO_3^{3-} ion:

Draw the simplified Lewis structure of the phosphite ion. 

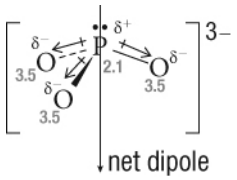
Use VSEPR theory to predict the three-dimensional structure, then draw a diagram. Because the central phosphorus atom has 3 bonding pairs of electrons and 1 lone pair, the PO_3^{3-} ion has a trigonal pyramidal arrangement of atoms

around the phosphorus atom. 

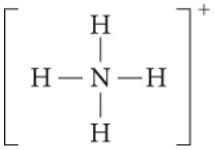
Identify the electronegativity of each atom, and determine the partial charges in the molecule. P = 2.1, +; O = 3.5, –



Draw the bond dipoles, and determine whether the molecule has a net dipole. In this trigonal pyramidal shape, the bond dipoles do not cancel each other out. Therefore, sulfur difluoride is a polar molecule with a net dipole pointing away from the lone pair on the phosphorus atom, toward the oxygen atoms.

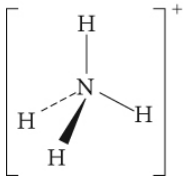


(b) Polarity of NH_4^+ ion:

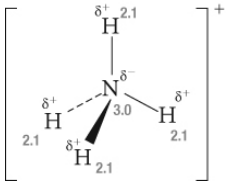
Draw the simplified Lewis structure of the ammonium ion. 

Use VSEPR theory to predict the three-dimensional structure, then draw a diagram.

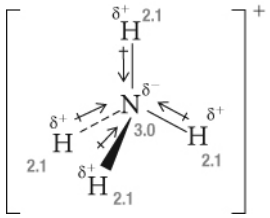
Since nitrogen is the central atom and it is bonded to 4 hydrogen atoms, the ammonium ion will have a tetrahedral shape.



Identify the electronegativity of each atom, and determine the partial charges in the molecule. N = 3.0, –; H = 2.1, +



Draw the bond dipoles, and determine whether the molecule has a net dipole. In this tetrahedral shape, all of the atoms bonded to the central nitrogen atom are identical and all the directions are equivalent. Therefore, the 4 bond dipoles cancel each other. The overall net dipole is zero, so the ammonium ion is non-polar.



p. 229 # 1-9 (exclude 6a)

1. A molecule of diboron tetrafluoride, B_2F_4, is non-polar because the molecule has a trigonal planar shape around both B atoms, resulting in a symmetrical molecule. Since the B–F bonds will have the same difference in electronegativity but their dipoles will point in opposite directions, they will cancel each other out and the molecule will be non- polar.

2. A molecule of methanol, CH_3OH, is polar because it is tetrahedral in shape about the carbon atom and angular or bent around the oxygen atom. The C–O bond has a greater difference in electronegativity than the C–H bonds, and since the O has a negative partial charge, this causes the molecule to be polar with a net dipole pointing in the direction of the O.

3. PH_3 is non-polar, unlike PF_3, because P and H have identical electronegativities.

4. Using partial charges to represent the dipole in a bond:

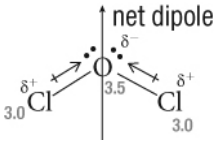
(a) δ– F–B δ+

(b) There is no partial charge in the N–Cl bond, as $EN_N = EN_{Cl} = 3.0$.

(c) δ– C–H δ+

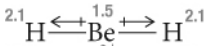
(d) There is no partial charge in the C–C bond because the atoms are identical.

5. (a) Polarity of OCl_2: $\Delta EN_{O-C} = EN_O - EN_C = 3.5 - 2.5 = 1.0$



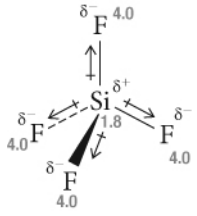
The molecule has more than 1 polar covalent bond. The molecular structure is bent, so the molecule is not symmetrical. OCl_2 is polar.

(b) Polarity of BeH_2: $\Delta EN_{H-Be} = 2.1 - 1.5 = 0.6$



The molecule has more than 1 polar covalent bond. The molecular structure is linear, so the molecule is symmetrical. The atoms bonded to the central atom are the same. BeH_2 is non-polar.

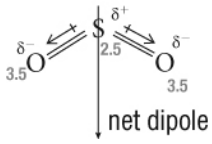
(c) Polarity of SiF_4: $\Delta EN_{F-Si} = 4.0 - 1.8 = 2.2$



The difference in electronegativity suggests the bonding in this molecule should be ionic bonds. However, experimental evidence suggests that the bonds are still polar covalent in nature. The molecular structure is tetrahedral, so the molecule is symmetrical.

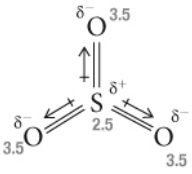
The atoms bonded to the central atom are the same. SiF_4 is non-polar.

(d) Polarity of SO_2: $\Delta EN_{O-S} = 3.5 - 2.5 = 1.0$



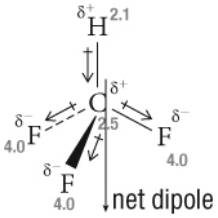
The molecule has more than 1 polar covalent bond. The molecular structure is bent, so the molecule is not symmetrical. SO_2 is polar.

(e) Polarity of SO_3: $\Delta EN_{O-S} = 3.5 - 2.5 = 1.0$



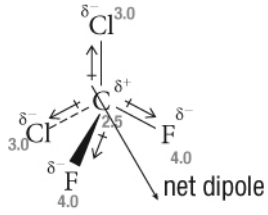
The molecule has more than 1 polar covalent bond. The molecular structure is trigonal planar, so the molecule is symmetrical. The atoms bonded to the central atom are the same. SO_3 is non-polar.

(f) Polarity of CHF_3: $\Delta EN_{F-C} = 4.0 - 2.5 = 1.5$



The molecule has more than 1 polar covalent bond. The molecular structure is tetrahedral, but since the atoms attached to the central carbon atom are not identical, the molecule is asymmetrical. Therefore, CHF_3 is polar.

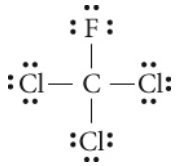
(g) Polarity of CCl_2F_2: $\Delta EN_{F-C} = 4.0 - 2.5 = 1.5$



The molecule has more than 1 polar covalent bond. The molecular structure is tetrahedral, but since the atoms attached to the central carbon atom are not identical, the molecule is asymmetrical.

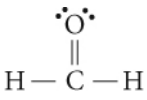
Therefore, CCl_2F_2 is polar.

6. (b) Simplified Lewis structure for CF_3Cl:



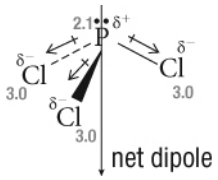
The central atom has 4 bonding electron pairs and no lone pairs, so the molecule is tetrahedral. The atoms bonded to the C are different, so the CF_3Cl molecule is polar.

(c) Simplified Lewis structure for H_2CO:

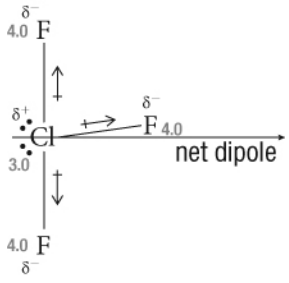


The central atom has 3 groups of electrons and no lone pairs, so the molecule is trigonal planar. The atoms bonded to the C are different, so the H_2CO molecule is polar.

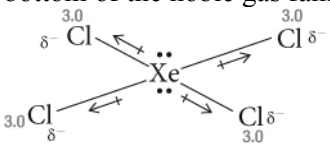
7. (a) Three-dimensional diagram indicating that PCl_3 is polar:



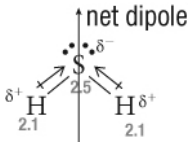
(b) Three-dimensional diagram indicating that ClF_3 is polar:



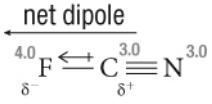
(c) Three-dimensional diagram indicating that $XeCl_4$ is non-polar: Note that the electronegativity of the noble gases is not given on the periodic table because these elements rarely form compounds. However, since Xenon is near the bottom of the noble gas family, we can assume that it has a lower electronegativity than Cl.



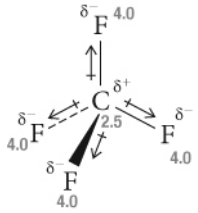
(d) Three-dimensional diagram indicating that H_2S is polar:



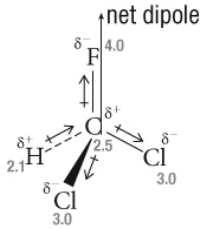
(e) Three-dimensional diagram indicating polarity for FCN:



8. (a) Three-dimensional diagram indicating that tetrafluoromethane, CF_4, is non-polar:



Three-dimensional diagram indicating that dichlorofluoromethane, $CHFCl_2$, is polar:



(b) Both dichlorofluoromethane and tetrafluoromethane are tetrahedral, so the difference between the molecules is in the distribution of charges in the bond dipoles. $CHFCl_2$ has multiple bond dipoles, which are not equivalent due to different electronegativities of the atoms bonded to the central atom. This results in a net dipole and thus the molecule is polar. CF_4 has an even distribution of charges in the bond dipoles, which cancel each other out, and is therefore non-polar.

(c) Answers may vary. Sample answer: Since dichlorofluoromethane is polar, there would be stronger intermolecular forces between its molecules than between those of CF_4, due to dipole–dipole interactions. Some physical properties that depend on intermolecular forces include melting point, boiling point, viscosity, solubility, binding affinity, miscibility, surface tension, adhesion, hydrophobicity, heat of vaporization, heat of fusion, elasticity, tensile strength, and capacitance. The boiling point and heat of vaporization are important considerations in the use of each compound as a refrigerant.

9.