Chapter 4 Practice Problem Key

**4.31** Each O and S atom must get two lone pairs and each N atom must get one lone pair.



**4.35** Follow the steps in Examples 4.2 and 4.3 to draw the Lewis structures.

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**4.40** Follow the steps in Example 4.3.

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**4.53** Name each compound as in Answer 4.17.

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| a. PBr3 = phosphorus tribromide | c. NCl3 = nitrogen trichloride |
| b. SO3 = sulfur trioxide | d. P2S5 = diphosphorus pentasulfide |

* 1. Follow the steps in Example 4.6.

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**4.65** BCl3 is trigonal planar because it has three Cl’s bonded to B but no lone pairs. NCl3 is trigonal pyramidal because it has three Cl’s around N as well as a lone pair.

**4.73** In covalent bonding, atoms share electrons to attain the electronic configuration of the noble gas closest to them in the periodic table. In ionic bonding, one atom donates electrons to the other atom. When the electronegativity difference between the two atoms in a bond is greater than 1.9, the bond is ionic.

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| a. hydrogen and bromine = polar covalent (2.8 – 2.1 = 0.7) | c. sodium and sulfur = polar covalent (2.5 – 0.9 = 1.6) |
| b. nitrogen and carbon = polar covalent (3.0 – 2.5 = 0.5) | d. lithium and oxygen = ionic (3.5 – 1.0 = 2.5) |

**4.77** Calculate the electronegativity difference between the atoms to determine which bond is more polar.

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| a. C–O = (3.5 – 2.5) = 1.0 = more polar C–N = (3.5 – 3.0) = 0.5 | c. Si–C = (2.5 – 1.8) = 0.7 = more polar  P–H = (2.1 – 2.1) = 0 |
| b. C–F = (4.0 – 2.5) = 1.5 = more polar C–Cl = (3.0 – 2.5) = 0.5  |  |

**4.83** No, a compound with only one polar bond must be polar. The single bond dipole is not cancelled by another bond dipole, so the molecule as a whole remains polar.

**4.93** First, give the O atom two lone pairs and each N atom one lone pair. Then add double bonds to the C’s that do not have four bonds.

