

Solutions to Suggested Problems From Chapter 4

2. Li is $[\text{He}]2s^1$ so it has 1 valence electron

C is $[\text{He}]2s^22p^2$ so it has 4 valence electrons

Mg is $[\text{Ne}]3s^2$ so it has 2 valence electrons

Ar is $[\text{Ne}]3s^23p^6$ so it has 8 valence electrons

4. C^{4-} is $[\text{He}]2s^22p^6$ so it has 8 valence electrons

N^{3-} is $[\text{He}]2s^22p^6$ so it has 8 valence electrons

S^{2-} is $[\text{Ne}]3s^23p^6$ so it has 8 valence electrons

I is $[\text{Kr}]5s^25p^6$ so it has 8 valence electrons

6. Filled d and f orbitals are significantly lower in energy (more negative ionization energy) than the s and p orbitals in the next higher shell.

13. Placing electrons between two nuclei helps to minimize the repulsive interaction between two nuclei, thus stabilizing the bond.

14. Based on our earlier understanding of atomic orbitals we should expect them to be paired (that is, to have opposite spins); we will return to this later when we build a quantum model of covalent bonding.

23. Bonding domains are areas where bonds are found between atoms. A single bond, a double bond, and a triple bond between any two atoms represents one bond domain. BF_3 and NH_3 must have three bonding domains since there are 3BF bonds and 3NH bonds, respectively. O_3 , therefore, must be the one with two bonding domains.

30. CO_2 has $1(4) + 2(6) = 16$ valence electrons

N_2O has $2(5) + 1(6) = 16$ valence electrons

CNO^- has $1(4) + 1(5) + 1(6) + 1 = 16$ valence electrons

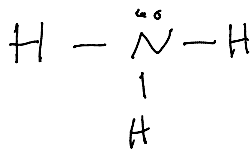
NO_2^+ has $1(5) + 2(6) - 1 = 16$ valence electrons

SO_2 has $1(6) + 2(6) = 18$ valence electrons

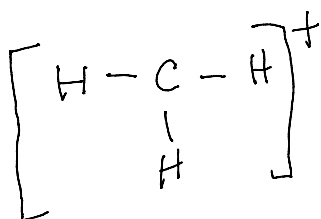
O_3 has $3(6) = 18$ valence electrons

NO_2^- has $1(5) + 2(6) + 1 = 18$ valence electrons

32. NH_3 has $1(5) + 3(1) = 8$ valence electrons

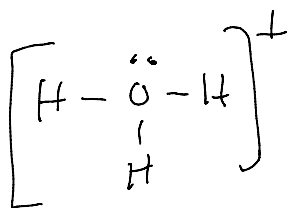


CH_3^+ has $1(4) + 3(1) - 1 = 6$ valence electrons

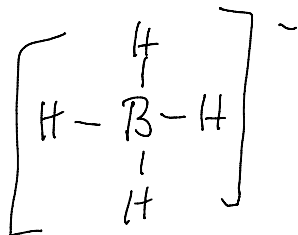


Now, this should bother you because carbon does not have an octet of electrons. However, we have no lone pairs of electrons to work with so this is the Lewis structure we are left with. Compounds such as this are very reactive because the atom without the octet would like to grab onto another pair of electrons.

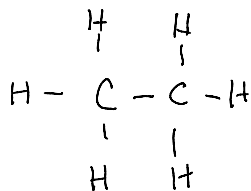
H_3O^+ has $3(1) + 1(6) - 1 = 8$ valence electrons



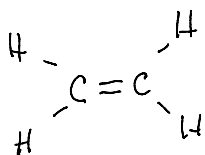
BH_4^- has $1(3) + 4(1) + 1 = 8$ valence electrons



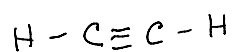
34. C_2H_6 has $2(4) + 6(1) = 14$ valence electrons



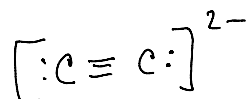
C_2H_4 has $2(4) + 4(1) = 12$ valence electrons



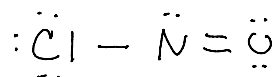
C_2H_2 has $2(4) + 2(1) = 10$ valence electrons



C_2^{2-} has $2(4) + 2 = 10$ valence electrons

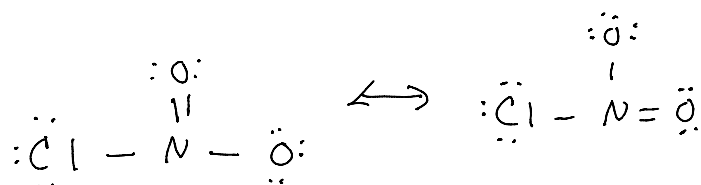


36. ClNO has $1(7) + 1(5) + 1(6) = 18$ valence electrons

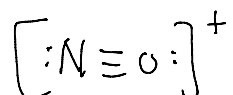


For reasons to be discussed later, this is the proper location of the double bond.

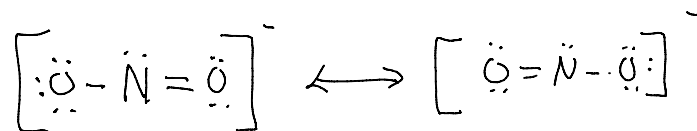
ClNO_2 has $1(7) + 1(5) + 2(6) = 24$ valence electrons



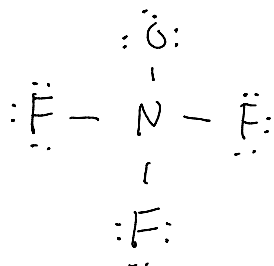
NO^+ has $1(5) + 1(6) - 1 = 10$ valence electrons



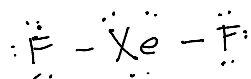
NO_2^- has $1(5) + 2(6) + 1 = 18$ valence electrons



ONF₃ has 1(6) + 1(5) + 3(7) = 32 valence electrons

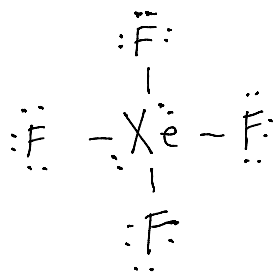


40. XeF₂ has 1(8) + 2(7) = 22 valence electrons

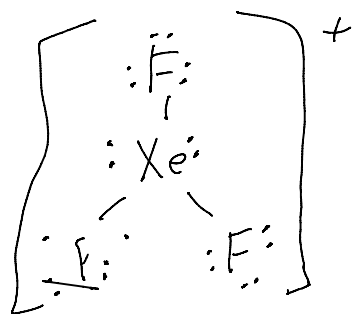


Note that fluorine only forms single bonds.

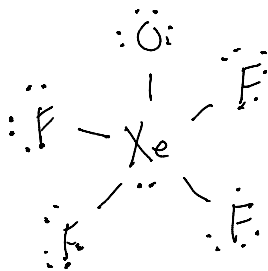
XeF₄ has 1(8) + 4(7) = 36 valence electrons



XeF₃⁺ has 1(8) + 3(7) - 1 = 28 valence electrons

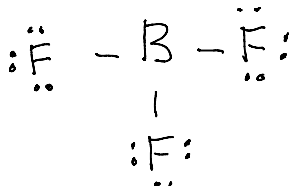


OXeF₄ has 1(6) + 1(8) + 4(7) = 42 valence electrons



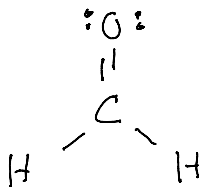
41. N₂ has 10 valence electrons. Here are the others: CO has 10 valence electrons; NO has 11 valence electrons; CN⁻ has 10 valence electrons; NO⁺ has 10 valence electrons; and NO⁻ has 12 valence electrons.

43. BF₃ has 1(3) + 3(7) = 24 valence electrons

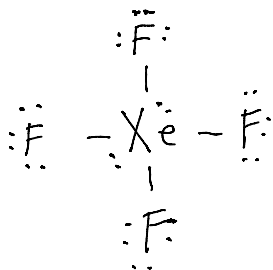


Note that boron doesn't satisfy the octet rule.

H₂CO has 2(1) + 1(4) + 1(6) = 12 valence electrons

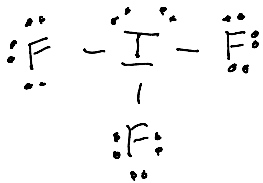


XeF₄ has 1(8) + 4(7) = 36 valence electrons



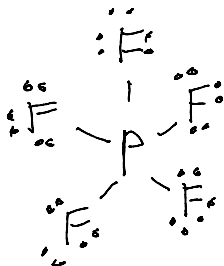
Note that this has an expanded octet and is an exception to the octet rule.

IF₃ has 1(7) + 3(7) = 28 valence electrons

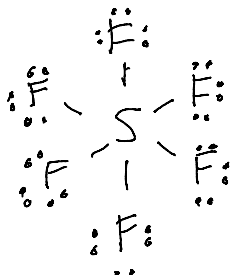


Note that there is an expanded octet on iodine, violating the octet rule.

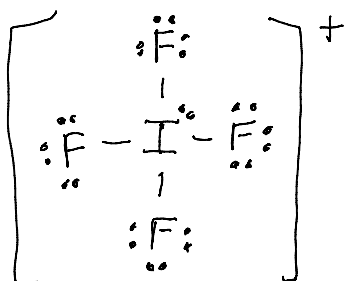
45. PF_5 has $1(5) + 5(7) = 28$ valence electrons



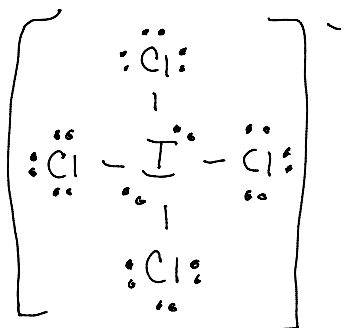
SF_6 has $1(6) + 6(7) = 48$ valence electrons



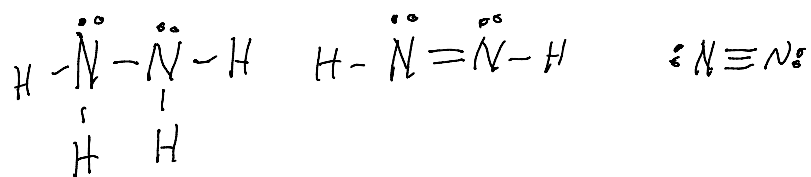
IF_4^+ has $1(7) + 4(7) - 1 = 34$ valence electrons



ICl_4^- has $1(7) + 4(7) + 1 = 36$ valence electrons



47. Longer bonds will have smaller bond orders; thus we expect H_2NNH_2 to have the single bond, HNNH to have the double bond, and N_2 to have the triple bond. This is confirmed by the following Lewis structures



48. For H_2S we have $0.104 + 0.037 = 0.141$ nm for an S-H bond

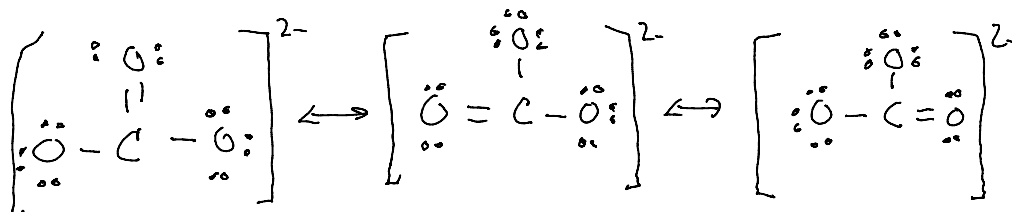
For OF_2 we have $0.066 + 0.64 = 0.130$ nm for an O-F bond

For NH_3 we have $0.070 + 0.037 = 0.107$ nm for a N-H bond

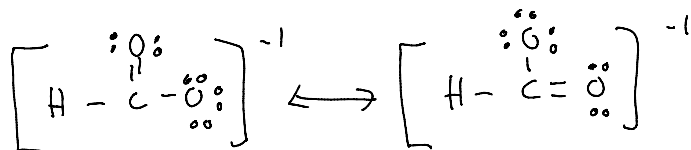
For BF_3 we have $0.088 + 0.064 = 0.152$ nm for a B-F bond

49. If the S-O bonds lengths are similar, then we also expect the bond orders to be similar and the bond strengths also should be similar.

50. CO_3^{2-} has $1(4) + 3(6) + 2 = 24$ valence electrons



52. Compounds b, c, and d cannot have resonance hybrids because they each have no more than only single bonds. For HCO_2^- we have $1(1) + 1(4) + 2(6) + 1 = 18$ valence electrons



57. Electronegativity, as with AVEE, increases from the bottom to the top of the periodic table and from the left side to the right side of the periodic table. Of the choices provide, only (b) fits the bill.

60. A Lewis structure for IBr is shown to the right. The partial charges are

$$\delta_{\text{I}} = 7 - 6 - 2 \left(\frac{2.36}{2.36 + 2.69} \right) = +0.0653 \quad \text{:}\ddot{\text{I}}\text{-}\ddot{\text{Br}}\text{:}$$

$$\delta_{\text{Br}} = 7 - 6 - 2 \left(\frac{2.69}{2.36 + 2.69} \right) = -0.0653$$

A Lewis structure for ICl is the same as that shown above for IBr. The partial charges are

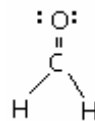
$$\delta_{\text{I}} = 7 - 6 - 2 \left(\frac{2.36}{2.36 + 2.87} \right) = +0.0975$$

$$\delta_{\text{Cl}} = 7 - 6 - 2 \left(\frac{2.87}{2.36 + 2.87} \right) = -0.0975$$

63. A Lewis structure for CH₂O is shown to the right. The partial charges are

$$\delta_{\text{O}} = 6 - 4 - 4 \left(\frac{3.61}{3.61 + 2.54} \right) = -0.35$$

$$\delta_{\text{H}} = 1 - 0 - 2 \left(\frac{2.30}{2.30 + 2.54} \right) = +0.05$$



65. CH₄: the EN for H is smaller than that for C; thus, H has the more positive partial charge

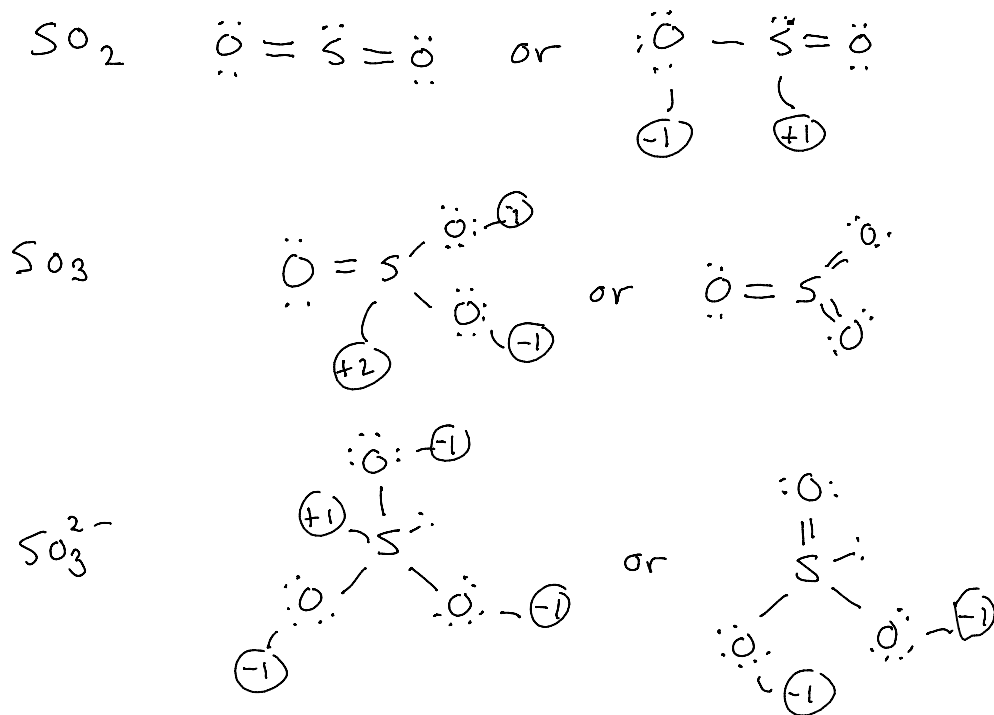
CCl₄: the EN for C is smaller than that for Cl; thus, C has the more positive partial charge

BH₃: the EN for B is smaller than that for H; thus, B has the more positive partial charge

NO₂: the EN for N is smaller than that for O; thus, N has the more positive partial charge

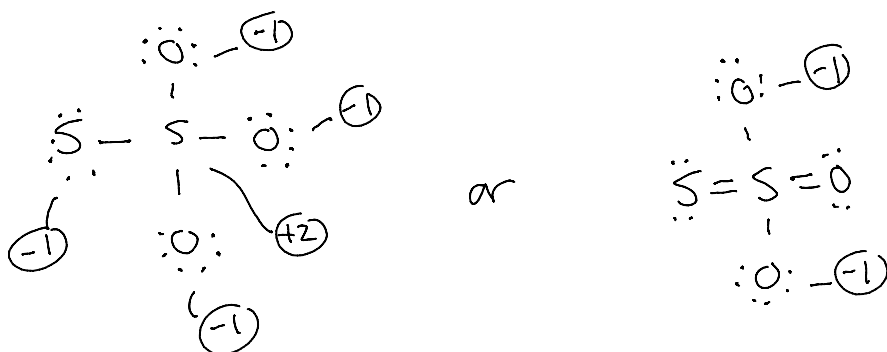
CINO: the EN for Cl is smaller than that for N or O; thus, Cl has the more positive partial charge

66. The three Lewis structures are shown below with the formal charges indicated.



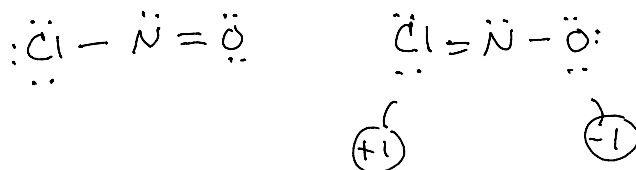
All three species show a dilemma – is it better to preserve the octet of electrons around sulfur or better to minimize formal charges. As discussed on page 139 preserving the octet is more consistent with available experimental evidence.

71. The Lewis structure and the formal charges are shown below.

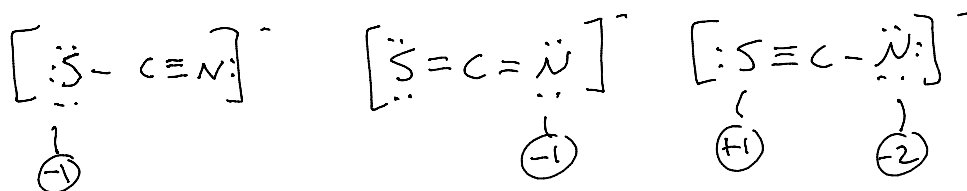


Once again, we have drawn a structure that preserves the octet and one that minimizes formal charges.

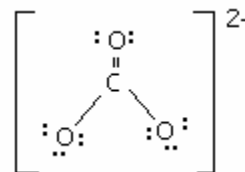
72. Two possible Lewis structures are shown below along with the formal charges. The best choice is the one with no formal charges. An experimental determination of the N-O bond length could help demonstrate this.



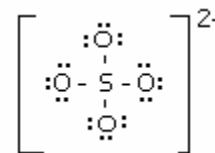
76. The three possibilities are shown below along with the formal charges. The best structure is the one in the middle as it minimizes the formal charges and puts the negative formal charge on the atom with the largest AVEE (or electronegativity).



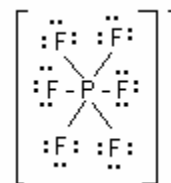
- 79: For CO_3^{2-} the Lewis structure is shown to the right. There are three electron domains and three bonding domains; thus, the geometry is trigonal planar with bond angles of approximately 120° .



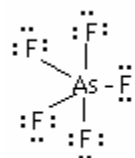
For SO_4^{2-} the Lewis structure is shown to the right. There are four electron domains and four bonding domains; thus, the geometry is tetrahedral with bond angles of approximately 109.5° .



For PF_6^- the Lewis structure is shown to the right. There are six electron domains and six bonding domains; thus, the geometry is octahedral and the bond angles are approximately 90° .



For AsF_5 the Lewis structure is shown to the right. There are five electron domains and five bonding domains; thus, the geometry is trigonal bipyramidal with bond angles of approximately 120° in the trigonal plane and 90° from the trigonal plane to the axial positions.



82. In each of these compounds there will be single bonds between each I and the atoms to which it is bound.

For I_2 there are 14 total electrons. The single bond between the two iodine atoms leaves 12 electrons, or three pairs of non-bonding electrons on each.

For I_3^- there are 22 total electrons. The two single bonds connecting the three iodine atoms in a linear chain accounts for four electrons, leaving 18 to account for. Placing three pairs on each terminal iodine accounts for 12 electrons, leaving six on the central iodine. Note that the central iodine has an three non-bonding pairs and two bonding pairs.

For IF_3 there are 28 total electrons. The three single I-F bonds account for six electrons and the three sets of lone-pairs on each fluorine account for 18 additional electrons. This leaves four more electrons as two lone-pairs on the central iodine atom.

For ICl_4^- there are 36 total electrons. The four I-Cl bonds account for eight electrons and the three sets of lone-pairs on each chlorine account for 24 additional electrons. This leaves four more electrons as two lone-pairs on the central iodine atom.

87. The Lewis structure for each of these ions and compounds has 32 total electrons with four bonding domains and four electron domains. This means that each has a tetrahedral bonding geometry.

89. A Lewis structure for SF_3^+ has 26 electrons in four electron domains around the central sulfur. Of these four electron domains, three are bonding. The geometry, therefore, is trigonal pyramidal.

A Lewis structure for SF_4 has 34 electrons in five electron domains around the central sulfur. Of these five electron domains, four are bonding. The geometry, therefore, is see-saw.

A Lewis structure for SF_5^- has 42 electrons in five electron domains around the central sulfur. Of these five electron domains, four are bonding. The geometry, therefore, is square pyramidal.

A Lewis structure for SF_6 has 48 electrons in six electron domains around the central sulfur. Of these six electron domains, all are bonding. The geometry, therefore, is octahedral.

100. For $O - X - O$ to be linear it must have two bonding domains. Possible electron domains are two, five and six.

For an electron domain of 2, the bonds between X and O must be double (to satisfy the octet on X) and two lone-pairs on O (to satisfy its octet). This gives a total of 16 electrons. The two oxygen atoms account for 12 of the electrons, leaving six for X, which must be C.

For an electron domain of 5, there must be single bonds between X and O and three lone-pairs on X. Each oxygen has three lone-pairs (to satisfy its octet). This gives a total of 22 electrons. The two oxygen atoms account for 12 of the electrons, leaving 10 for X, which does not match any of the choices.

For an electron domain of 6, there must be single bonds between X and O and four lone-pairs on X. Each oxygen has three lone-pairs (to satisfy its octet). This gives a total of 24 electrons. The two oxygen atoms account for 12 of the electrons, leaving 12 for X, which does not match any of the choices.

103. The missing values are in **bold**

Bonding Domains	Non-Bonding Domains	Bond Angle
2	0	180°
5	0	90° and 120°
3	0	120°
2	1	120°
2	3	180°
4	0	109°
3	2	109°
2	3	180°
3	2	90°
4	2	90°

106. BCl_3 is an unusual compound because boron does not have an octet. A Lewis structure for BCl_3 leaves only three bonding electrons around boron. This is an ideal trigonal planar bonding geometry of 120°.

For H_2CCH_2 there is a double-bond between the two carbon atoms. The bonding geometry around each carbon is trigonal planar, but the three substituents are not the same and the H-C-H bond angle can flex outwards.

For H_2O the electron domain around the oxygen is four, with two bonding domains and two non-bonding domains. The ideal tetrahedral geometry has 109.5° bond angles. Here, with two bonding domains, the bent geometry is smaller due to the size of the non-bonding electron domain relative to that for the bonding domains.

113. The Lewis structure for SOCl_2 has four electron domains around the central atom of sulfur, with three bonding domains and one non-bonding domain. Each bond is polar and the trigonal pyramidal geometry does not allow them to cancel out each other; the result is a polar molecule.

The Lewis structure for SO_2Cl_2 has four electron domains, all of which are bonding. The tetrahedral structure does not allow the individual polar bonds to cancel each other; thus, the molecule is polar.