CHAPTER 12

Chemical Bonding

CHAPTER ANSWERS

1. bond
2. bond energy
3. ionic
4. covalent
5. In Cl_2 the bonding is pure covalent with the electron pair shared equally between the two chlorine atoms. In HCl there is a shared pair of electrons, but the shared pair is drawn more closely to the chlorine atom, and this makes the bond polar.
6. In H_2 and HF, the bonding is covalent in nature with an electron pair being shared between the atoms. In H_2, the two atoms are identical (the sharing is equal). In HF, the two atoms are different (the sharing is unequal) and as a result the bond is polar. Both of these are in marked contrast to the situation in NaF: NaF is an ionic compound—an electron has been completely transferred from sodium to fluorine, producing separate ions.
7. electronegativity
8. A bond is polar if the centers of positive and negative charge do not coincide at the same point. The bond has a negative end and a positive end. Polar bonds will exist in any molecule with nonidentical bonded atoms (although the molecule as a whole may not be polar if the bond dipoles cancel each other). Two simple examples are HF and HCl; in both cases, the negative center of charge is closer to the halogen atom.
9. from top to bottom: covalent, ionic, polar covalent
10. The level of polarity in a polar covalent bond is determined by the difference in electronegativity of the atoms in the bond.
11. In general, an element farther to the right in a given period or an element closer to the top of a given group is more electronegative. The elements below are ranked in order of increasing electronegativity.
    a. Li < C < F
    b. I < Cl < F
    c. Cs < Rb < Li
12.
    a. I is most electronegative, Rb is least electronegative
    b. Mg is most electronegative, Ca and Sr have similar electronegativities
    c. Br is most electronegative, K is least electronegative

13. Generally, covalent bonds between atoms of *different* elements are *polar*.
    a. ionic
    b. ionic
    c. polar covalent
    d. polar covalent
14. Generally, covalent bonds between atoms of *different* elements are *polar*.
    a. ionic
    b. polar covalent
    c. covalent
15. For a bond to be polar covalent, the atoms involved in the bond must have different electronegativities (must be of different elements).
    a. nonpolar covalent (atoms of the same element)
    b. nonpolar covalent (atoms of the same element)
    c. nonpolar covalent (atoms of the same element)
    d. polar covalent (atoms of different elements)
16. For a bond to be polar covalent, the atoms involved in the bond must have different electronegativities (must be of different elements).
    a. nonpolar covalent (atoms of the same element)
    b. nonpolar covalent (atoms of the same element)
    c. polar covalent (atoms of different elements)
    d. polar covalent (atoms of different elements)
17. The *degree* of polarity of a polar covalent bond is indicated by the magnitude of the difference in electronegativities of the elements involved; the larger the difference in electronegativity, the more polar the bond. Electronegativity differences are given in parentheses below:
    a. H–F (1.9); H–Cl (0.9); the H–F bond is more polar
    b. H–Cl (0.9); H–I (0.4); the H–Cl bond is more polar
    c. H–Br (0.7); H–Cl (0.9); the H–Cl bond is more polar
    d. H–I (0.4); H–Br (0.7); the H–Br bond is more polar
18. The atom with the larger electronegativity will be more negative relative to the other atom.
    a. F
    b. neither (similar electronegativity)
    c. Cl
    d. Cl
19. The larger the difference in electronegativity between two atoms, the more polar will be the bond between those atoms. Electronegativity differences are given in parentheses below:
    a. H–S (0.4); H–F (1.9); the H–F bond is more polar

b. O–S (1.0); O–F (0.5); the O–S bond is more polar

c. N–S (0.5); N–Cl(0); the N–S bond is more polar

d. C–S (0); C–Cl (0.5); the C–Cl bond is more polar

20. The greater the electronegativity difference between two atoms, the more ionic will be the bond between those two atoms.

a. Ca–Cl

b. Ba–Cl

c. Fe–I

d. Be–F

21. A dipole moment is an electrical effect that occurs in a molecule that has separate centers of positive and negative charge. The simplest examples of molecules with dipole moments would be diatomic molecules involving two different elements. For example:

$\delta+$ C $\rightarrow$ O $\delta-$

$\delta+$ N $\rightarrow$ O $\delta-$

$\delta+$ Cl $\rightarrow$ F $\delta-$

$\delta+$ Br $\rightarrow$ Cl $\delta-$

22. The presence of strong bond dipoles and a large overall dipole moment in water make it a polar substance overall. Among the properties of water dependent on its dipole moment are its freezing point, melting point, vapor pressure, and its ability to dissolve many substances.

23. In a diatomic molecule containing two different elements, the more electronegative atom will be the negative end of the molecule, and the *less* electronegative atom will be the positive end.

a. chlorine

b. oxygen

c. fluorine

24. In a diatomic molecule containing two different elements, the more electronegative atom will be the negative end of the molecule, and the *less* electronegative atom will be the positive end.

a. H

b. Cl

c. I

25. In the figures, the arrow points toward the more electronegative atom.

a. $\delta+$ C $\rightarrow$ F $\delta-$

b. $\delta+$ Si $\rightarrow$ C $\delta-$

c. $\delta+$ C $\rightarrow$ O $\delta-$

d. $\delta+$ B $\rightarrow$ C $\delta-$

26. In the figures, the arrow points toward the more electronegative atom.

a. $\delta+$ P $\rightarrow$ F $\delta-$

b. $\delta+$ P $\rightarrow$ O $\delta-$

c. $\delta+$ P $\rightarrow$ C $\delta-$

d. P and H have similar electronegativities.

27. In the figures, the arrow points toward the more electronegative atom.

a. $\delta+$ Si $\rightarrow$ H $\delta-$

b. P–H The atoms have very nearly the same electronegativity, so there is a very small, if any, dipole moment.

c. $\delta+$ H $\rightarrow$ S $\delta-$

d. $\delta+$ H $\rightarrow$ Cl $\delta-$

28. In the figures, the arrow points toward the more electronegative atom.

a. $\delta+$ H $\rightarrow$ C $\delta-$

b. $\delta+$ N $\rightarrow$ O $\delta-$

c. $\delta+$ S $\rightarrow$ N $\delta-$

d. $\delta+$ C $\rightarrow$ N $\delta-$

29. When forming ionic bonds, the element forming the positive ion loses enough electrons to have the same configuration as the previous noble gas. The element forming the negative ion gains enough electrons to have the same configuration as the next noble gas. When forming covalent compounds, electrons are shared in such a way that as many atoms as possible in the molecule have a configuration analogous to a noble gas.

30. preceding

31. gaining

32. Atoms in covalent molecules gain a configuration like that of a noble gas by sharing one or more pairs of electrons between atoms. Such shared pairs of electrons "belong" to each of the atoms of the bond at the same time. In ionic bonding, one atom completely gives over one or more electrons to another atom, and the resulting ions behave independently of one another.

33.

a. Na $1s^2\,2s^2\,2p^6\,3s^1$

Na^+ $1s^2\,2s^2\,2p^6$

Ne has the same configuration as Na^+.

b. I $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^5$

I^- $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6$

Xe has the same configuration as I^-.

c. Ca $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2$

Ca^{2+} $1s^2\,2s^2\,2p^6\,3s^2\,3p^6$

Ar has the same configuration as Ca^{2+}.

d. N $1s^2\ 2s^2\ 2p^3$

N^{3} $1s^2\ 2s^2\ 2p^6$

Ne has the same configuration as N^{3} .

e. F $1s^2\ 2s^2\ 2p^5$

F $1s^2\ 2s^2\ 2p^6$

Ne has the same configuration as F^{-}.

34.

a. Br^{-}, Kr (Br has one electron less than Kr.)

b. Cs^{+}, Xe (Cs has one electron more than Xe.)

c. P^{3-}, Ar (P has three fewer electrons than Ar.)

d. S^{2-}, Ar (S has two fewer electrons than Ar.)

35.

a. Mg^{2+} (Mg has two electrons more than the noble gas Ne.)

b. Al^{3+} (Al has three electrons more than the noble gas Ne.)

c. I^{-} (I has one electron less than the noble gas Xe.)

d. Ca^{2+} (Ca has two electrons more than the noble gas Ar.)

36. Atoms or ions with the same number of electrons are said to be *isoelectronic.*

a. F^{-}, O^{2-}, N^{3-}

b. Cl^{-}, S^{2-}, P^{3-}

c. F^{-}, O^{2-}, N^{3-}

d. Br^{-}, Se^{2-}, As^{3-}

37.

a. Al_2S_3: Al has three electrons more than a noble gas; S has two fewer electrons than a noble gas.

b. RaO: Ra has two electrons more than a noble gas; O has two fewer electrons than a noble gas.

c. CaF_2: Ca has two electrons more than a noble gas; F has one electron less than a noble gas.

d. Cs_3N: Cs has one electron more than a noble gas; N has three fewer electrons than a noble gas.

e. RbP_3: Rb has one electron more than a noble gas; P has three fewer electrons than a noble gas.

38.

a. Na_2S: Na has one electron more than a noble gas; S has two fewer electrons than a noble gas.

b. BaSe: Ba has two electrons more than a noble gas; Se has two fewer electrons than a noble gas.

c. $MgBr_2$: Mg has two electrons more than a noble gas; Br has one electron less than a noble gas.

d. Li_3N: Li has one electron more than a noble gas; N has three fewer electrons than a noble gas.

e. KH: K has one electron more than a noble gas; H has one electron less than a noble gas.

39.

a. Ba^{2+}, [Xe]; S^{2-}, [Ar]

b. Sr^{2+}, [Kr]; F^{-}, [Ne]

c. Mg^{2+}, [Ne]; O^{2-}, [Ne]

d. Al^{3+}, [Ne]; S^{2-}, [Ar]

40.

a. Sr^{2+}, [Kr]; O^{2-}, [Ne]

b. Ca^{2+}, [Ar]; H^{-}, [He]

c. K^{+}, [Ar]; P^{3-}, [Ar]

d. Ba^{2+}, [Xe]; Se^{2-}, [Kr]

41. The formula of an ionic compound represents only the smallest whole number ratio of positive and negative ions present (the empirical formula).

42. An ionic solid such as NaCl consists of an array of alternating positively and negatively charged ions; that is, each positive ion has as its nearest neighbors a group of negative ions, and each negative ion has a group of positive ions surrounding it. In most ionic solids, the ions are packed as tightly as possible.

43. Positive ions are always smaller than the atoms from which they are formed because, in forming the ion, the valence electron shell (or part of it) is "removed" from the atom.

44. In forming an anion, an atom gains additional electrons in its outermost (valence) shell. Additional electrons in the valence shell increase the repulsive forces between electrons, so the outermost shell becomes larger to accommodate this.

45. Relative ionic sizes are given in Figure 12.9. Within a given horizontal row of the periodic chart, negative ions tend to be larger than positive ions because the negative ions contain a larger number of electrons in the valence shell. Within a vertical group of the periodic table, ionic size increases from top to bottom. In general, positive ions are smaller than the atoms from which they come, whereas negative ions are larger than the atoms from which they come.

a. H

b. N

c. Al^{3+}

d. F

46. Relative ionic sizes are given in Figure 12.9. Within a given horizontal row of the periodic chart, negative ions tend to be larger than positive ions because the negative ions contain a larger number of electrons in the valence shell. Within a vertical group of the periodic table, ionic size increases from top to bottom. In general, positive ions are smaller than the atoms from which they come, whereas negative ions are larger than the atoms from which they come.

    a. F^-

    b. Cl^-

    c. Ca

    d. I^-

47. Relative ionic sizes are given in Figure 12.9. Within a given horizontal row of the periodic chart, negative ions tend to be larger than positive ions because the negative ions contain a larger number of electrons in the valence shell. Within a vertical group of the periodic table, ionic size increases from top to bottom. In general, positive ions are smaller than the atoms from which they come, whereas negative ions are larger than the atoms from which they come.

    a. Fe^{3+}

    b. Cl

    c. Al^{3+}

48. Relative ionic sizes are given in Figure 12.9. Within a given horizontal row of the periodic chart, negative ions tend to be larger than positive ions because the negative ions contain a larger number of electrons in the valence shell. Within a vertical group of the periodic table, ionic size increases from top to bottom. In general, positive ions are smaller than the atoms from which they come, whereas negative ions are larger than the atoms from which they come.

    a. I

    b. F^-

    c. F^-

49. Valence electrons are those found in the outermost principal energy level of the atom. The valence electrons effectively represent the outside edge of the atom and are the electrons most influenced by the electrons of another atom.

50. When atoms form covalent bonds, they try to attain a valence electronic configuration similar to that of the following noble gas element. When the elements in the first few horizontal rows of the periodic table form covalent bonds, they will attempt to gain configurations similar to the noble gases helium (two valence electrons, duet rule), and neon and argon (eight valence electrons, octet rule).

51. noble gas electronic configuration

52. These elements attain a total of eight valence electrons, making the valence electron configurations similar to those of the noble gases Ne and Ar.

53. When two atoms in a molecule are connected by a double bond, the atoms share two pairs of electrons (four electrons) in completing their outermost shells. A simple molecule containing a double bond is ethene (ethylene), C_2H_4 ($H_2C::CH_2$).

54. When two atoms in a molecule are connected by a triple bond, the atoms share three pairs of electrons (six electrons) in completing their outermost shells. A simple molecule containing a triple bond is acetylene, C_2H_2 (H:C:::C:H).

55.

a.

:Ï:

b.

Ȧl·

c.

:Ẍe:

d.

:Sr

56.

a.

Rb·

b.

:Ċl:

c.

:K̈r:

d.

Ba:

e.

·P̈·

f.

:Ȧt·

57.

a. each N provides 5; O provides 6; total valence electrons = 16
b. each B provides 3; each H provides 1: total valence electrons = 12
c. each C provides 4; each H provides 1; total valence electrons = 20
d. N provides 5; each Cl provides 7; total valence electrons = 26

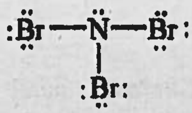
58.

a. phosphorus provides 5; each hydrogen provides 1; total valence electrons = 8
b. each N provides 5; oxygen provides 6; total valence electrons = 16
c. each carbon provides 4; each hydrogen provides 1; total valence electrons = 26

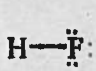
d. each carbon provides 4; each hydrogen provides 1; bromine provides 7; total valence electrons = 26

59.

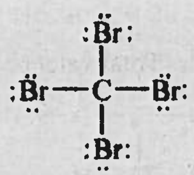
a. Nitrogen provides 5 valence electrons; each bromine provides 7; total valence electrons = 26.



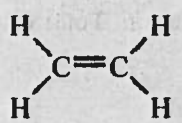
b. Hydrogen provides 1 valence electron; fluorine provides 7; total valence electrons = 8.



c. Carbon provides 4 valence electrons; each bromine provides 7; total valence electrons = 32.



d. Each carbon provides 4 valence electrons; each hydrogen provides 1; total valence electrons = 10.

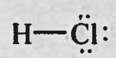


60.

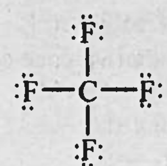
a. Each hydrogen provides 1 valence electron; total valence electrons = 2

H—H

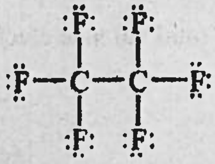
b. Hydrogen provides 1 valence electron; chlorine provides 7 valence electrons; total valence electrons = 8



c. Carbon provides 4 valence electrons; each fluorine provides 7 valence electrons; total valence electrons = 32

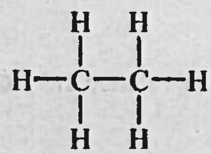


d. Each carbon provides 4 valence electrons; each fluorine provides 7 valence electrons; total valence electrons = 50

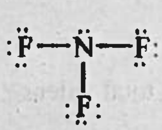


61.

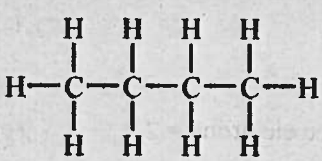
a. Each C provides 4 valence electrons. Each H provides 1 valence electron. Total valence electrons = 14



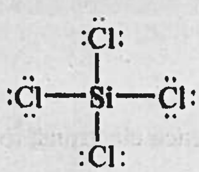
b. N provides 5 valence electrons. Each F provides 7 valence electrons. Total valence electrons = 26



c. Each C provides 4 valence electrons. Each H provides 1 valence electron. Total valence electrons = 26

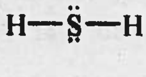
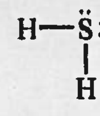


d. Si provides 4 valence electrons. Each Cl provides 7 valence electrons. Total valence electrons = 32

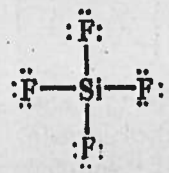


62.

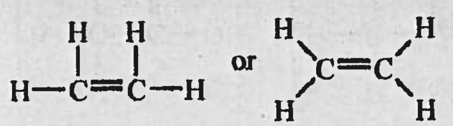
a. Each H provides 1 valence electron. S provides 6 valence electrons. Total valence electrons = 8

 or sometimes drawn as follows to indicate the shape 

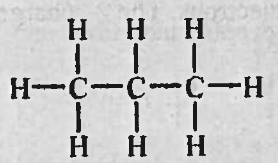
b. Si provides 4 valence electrons. Each F provides 7 valence electrons. Total valence electrons = 32



c. Each C provides 4 valence electrons. Each H provides 1 valence electron. Total valence electrons = 12

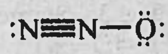


d. Each C provides 4 valence electrons. Each H provides 1 valence electron. Total valence electrons = 20



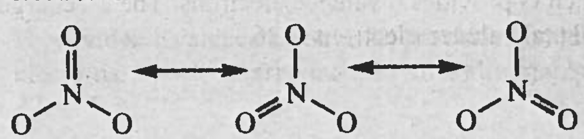
63.

a. no resonance



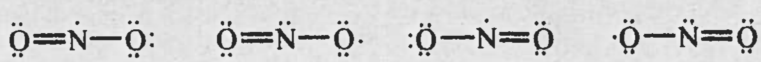
b. The fact that there is an odd number of electrons (17) guarantees resonance structures. Here are two possible structures.

c. Resonance is possible in the nitrate ion. Here are three structures showing only the bonding electrons.

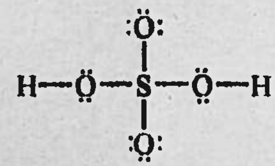


64.

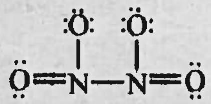
a. N provides 5 valence electrons. Each O provides 6 valence electrons. Total valence electrons = 17 (an odd number). Note that there are several Lewis structures possible because of the odd number of electrons (the unpaired electron can be located on different atoms in the Lewis structure).

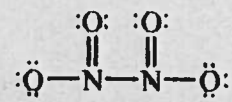


b. Each H provides 1 valence electron. S provides 6 valence electrons. Each O provides 6 valence electrons. Total valence electrons = 32



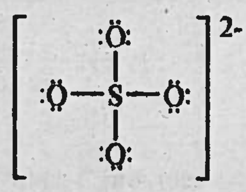
c. Each N provides 5 valence electrons. Each O provides 6 valence electrons. Total valence electrons = 34



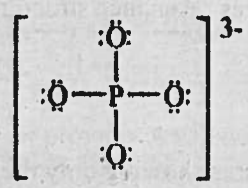


65.

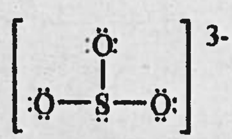
a. S provides 6 valence electrons. Each O provides 6 valence electrons. The 2– charge means two additional valence electrons. Total valence electrons = 32



b. P provides 5 valence electrons. Each O provides 6 valence electrons. The 3– charge means three additional valence electrons. Total valence electrons = 32.

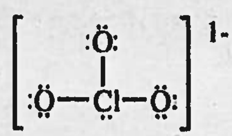


c. S provides 6 valence electrons. Each O provides 6 valence electrons. The 2– charge means two additional valence electrons. Total valence electrons = 26

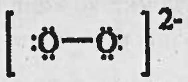


66.

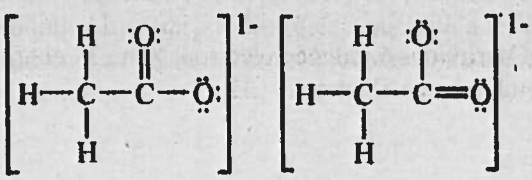
a. Cl provides 7 valence electrons. Each O provides 6 valence electrons. The 1– charge means 1 additional electron. Total valence electrons = 26



b. Each O provides 6 valence electrons. The 2– charge means two additional valence electrons. Total valence electrons = 14

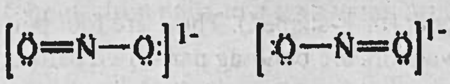


c. Each C provides 4 valence electrons. Each H provides 1 valence electron. Each O provides 6 valence electrons. The 1– charge means 1 additional valence electron. Total valence electrons = 24

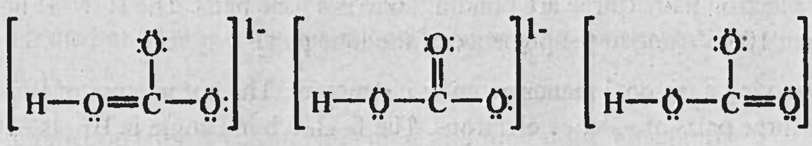


67.

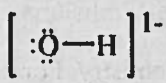
a. N provides 5 valence electrons. Each O provides 6 valence electrons. The 1– charge means one additional valence electron. Total valence electrons = 18



b. H provides 1 valence electron. C provides 4 valence electrons. Each O provides 6 valence electrons. The 1– charge means one additional valence electron. Total valence electrons = 24

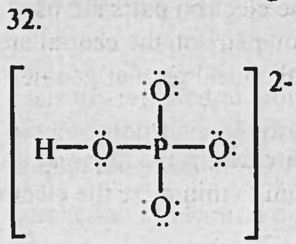


c. O provides 6 valence electrons. H provides 1 valence electron. The 1– charge means one additional valence electron. Total valence electrons = 8

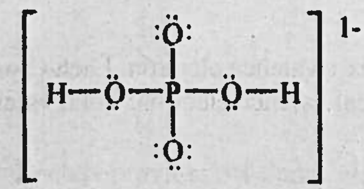


68.

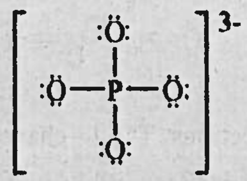
a. H provides 1 valence electron. P provides 5 valence electrons. Each O provides 6 valence electrons. The 2– charge means two additional valence electrons. Total valence electrons = 32.



b. Each H provides 1 valence electron. P provides 5 valence electrons. Each O provides 6 valence electrons. The 1– charge means one additional valence electron. Total valence electrons = 32



c. P provides 5 valence electrons. Each O provides 6 valence electrons. The 3– charge means three additional valence electrons. Total valence electrons = 32



69. The geometric structure of the water molecule is bent (or *V*-shaped). There are four pairs of valence electrons on the oxygen atom of water (two pairs are bonding pairs; two pairs are non-bonding lone pairs). The H–O–H bond angle in water is approximately 106°.

70. The geometric structure of NH_3 is that of a trigonal pyramid. The nitrogen atom of NH_3 is surrounded by four electron pairs (three are bonding; one is a lone pair). The H–N–H bond angle is somewhat less than 109.5° (due to the presence of the lone pair).

71. BF_3 is described as having a trigonal planar geometric structure. The boron atom of BF_3 is surrounded by only three pairs of valence electrons. The F–B–F bond angle in BF_3 is 120°.

72. The geometric structure of SiF_4 is that of a tetrahedron. The silicon atom of SiF_4 is surrounded by four bonding electron pairs. The F–Si–F bond angle is the characteristic angle of the tetrahedron, 109.5°.

73. The geometric structure of a molecule plays a very important part in its chemistry. For biological molecules, a slight change in the geometric structure of the molecule can completely destroy the molecule's usefulness to a cell or can cause a destructive change in the cell.

74. The general molecular structure of a molecule is determined by (1) *how many electron pairs* surround the central atom in the molecule and (2) which of those electron pairs are used for *bonding* to the other atoms of the molecule. Nonbonding electron pairs on the central atom do, however, cause minor changes in the bond angles compared to the ideal regular geometric structure.

75. For a given atom, the positions of the other atoms bonded to it are determined by maximizing the angular separation of the valence electron pairs on the given atom to minimize the electron repulsions.

76. You will remember from high school geometry that two points in space are all that is needed to define a straight line. A diatomic molecule represents two points (the nuclei of the atoms) in space.

77. One of the valence electron pairs of the ammonia molecule is a *lone* pair with no hydrogen atom attached. This lone pair is part of the nitrogen atom and does not enter the description of the molecule's overall shape (aside from influencing the H–N–H bond angles between the remaining valence electron pairs).

78. In NF_3, the nitrogen atom has *four* pairs of valence electrons whereas in BF_3 there are only *three* pairs of valence electrons around the boron atom. The nonbonding electron pair on nitrogen in NF_3 pushes the three F atoms out of the plane of the N atom.

79. If you draw the Lewis structures for these molecules, you will see that each of the indicated atoms is surrounded by *four pairs* of electrons with a *tetrahedral* orientation of the electron pairs (the arrangement in H_2S is distorted because some electron pairs on the S atom are bonding and some are nonbonding).

80.

a. four electron pairs arranged in a tetrahedral arrangement with some distortion due to the nonbonding pair

b. four electron pairs in a tetrahedral arrangement

c. four electron pairs in a tetrahedral arrangement

81.

a. trigonal pyramidal (there is a lone pair on N)

b. nonlinear, *V*-shaped (four electron pairs on Se, but only two atoms are attached to Se)

c. tetrahedral (four electron pairs on Si and four atoms attached)

82.

a. tetrahedral (four electron pairs on C and four atoms attached)

b. nonlinear, *V*-shaped (four electron pairs on S, but only two atoms attached)

c. tetrahedral (four electron pairs on Ge and four atoms attached)

83.

a. tetrahedral

b. tetrahedral

c. tetrahedral

84.

a. basically tetrahedral around the P atom (The hydrogen atoms are attached to two of the oxygen atoms and do not affect greatly the geometrical arrangement of the oxygen atoms around the phosphorus.)

b. tetrahedral (4 electron pairs on Cl and 4 atoms attached)

c. trigonal pyramidal (4 electron pairs on S and 3 atoms attached)

85.

a. <109.5°

b. <109.5°

c. 109.5°

d. 109.5°

86.

a. approximately 109.5° (the molecule is *V*-shaped or nonlinear)

b. approximately 109.5° (the molecule is trigonal pyramidal)

c. 109.5°

d. approximately 120° (the double bond makes the molecule flat)

87. There would be four electron pairs around the sulfur atom, so the arrangement of the electron pairs connecting the sulfur atom to the four oxygen atoms would basically be tetrahedral. Assuming the influence of the hydrogen atoms can be neglected, the molecule overall would give the impression of being basically tetrahedral.

88. The ethylene molecule contains a double bond between the carbon atoms. This makes the molecule planar (flat) with H–C–H and H–C–C bond angles of approximately 120°. The 1,2-dibromoethane molecule would *not* be planar, however. Each carbon would have four bonding pairs of electrons around it, and so the orientation around each carbon atom would be basically tetrahedral with bond angles of approximately 109.5° (assuming all bonds are similar).

89. Resonance is said to exist when more than one valid Lewis structure can be drawn for a molecule. The actual bonding and structure of such molecules is thought to be somewhere "in between" the various possible Lewis structures. In such molecules, certain electrons are delocalized over more than one bond.

90. double

91. repulsion

92. The bond with the larger electronegativity difference will be the more polar bond. See Figure 12.3 for electronegativities.

a. S–F

b. P–O

c. C–H

93. The bond with the larger electronegativity difference will be the more polar bond. See Figure 12.3 for electronegativities.

a. Br–F

b. As–O

c. Pb–C

94. The bond energy of a chemical bond is the quantity of energy required to break the bond and separate the atoms.

95. covalent

96. In each case, the element *higher up* within a group on the periodic table has the higher electronegativity.

    a. Be

    b. N

    c. F

97. For a bond to be polar covalent, the atoms involved in the bond must have different electronegativities (must be of different elements).

    a. polar covalent

    b. covalent

    c. covalent

    d. polar covalent

98. For a bond to be polar covalent, the atoms involved in the bond must have different electronegativities (must be of different elements).

    a. polar covalent (different elements)

    b. *non*polar covalent (two atoms of the same element)

    c. polar covalent (different elements)

    d. *non*polar covalent (atoms of the same element)

99. Electronegativity differences are given in parentheses.

    a. N–P (0.9); N–O (0.5); the N–P bond is more polar.

    b. N–C (0.5); N–O (0.5); the bonds are of the same polarity.

    c. N–S (0.5); N–C (0.5); the bonds are of the same polarity.

    d. N–F (1.0); N–S (0.5); the N–F bond is more polar.

100. In a diatomic molecule containing two different elements, the more electronegative atom will be the negative end of the molecule, and the *less* electronegative atom will be the positive end.

    a. oxygen

    b. bromine

    c. iodine

101. In the figures, the arrow points toward the more electronegative atom.

    a. N–Cl: The atoms have very nearly the same electronegativity, so there is a very small, if any, dipole moment.

    b. $\delta+$ P $\rightarrow$ N $\delta-$

    c. $\delta+$ S $\rightarrow$ N $\delta-$

    d. $\delta+$ C $\rightarrow$ N $\delta-$

102.

a. Al $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

Al^{3+} $1s^2\ 2s^2\ 2p^6$

Ne has the same configuration as Al^{3+}.

b. Br $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$

Br^- $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$

Kr has the same configuration as Br^-.

c. Ca $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

Ca^{2+} $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Ar has the same configuration as Ca^{2+}.

d. Li $1s^2\ 2s^1$

Li^+ $1s^2$

He has the same configuration as Li^+.

e. F $1s^2\ 2s^2\ 2p^5$

F^- $1s^2\ 2s^2\ 2p^6$

Ne has the same configuration as F^-.

103.

a. Na^+

b. I^-

c. K^+

d. Ca^{2+}

e. S^{2-}

f. Mg^{2+}

g. Al^{3+}

h. N^{3-}

104.

a. Na_2Se: Na has one electron more than a noble gas; Se has two fewer electrons than a noble gas.

b. RbF: Rb has one electron more than a noble gas; F has one electron less than a noble gas.

c. K_2Te: K has one electron more than a noble gas; Te has two fewer electrons than a noble gas.

d. BaSe: Ba has two electrons more than a noble gas; Se has two fewer electrons than a noble gas.

e. KAt: K has one electron more than a noble gas; At has one electron less than a noble gas.

f. FrCl: Fr has one electron more than a noble gas; Cl has one electron less than a noble gas.

105.

a. Ca^{2+} [Ar]; Br^{-} [Kr]

b. Al^{3+} [Ne]; Se^{2-} [Kr]

c. Sr^{2+} [Kr]; O^{2-} [Ne]

d. K^{+} [Ar]; S^{2-} [Ar]

106. Relative ionic sizes are indicated in Figure 12.9.

a. Na^{+}

b. Al^{3+}

c. F^{-}

d. Na^{+}

107.

a. He

He:

b. Br

:B̤̈r·

c. Sr

Sr:

d. Ne

:N̤̈e:

e. I

:Ï̤·

f. Ra

Ra:

108.

a. H provides 1; N provides 5; each O provides 6; total valence electrons = 24

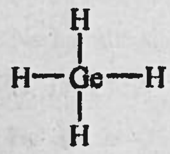
b. each H provides 1; S provides 6; each O provides 6; total valence electrons = 32

c. each H provides 1; P provides 5; each O provides 6; total valence electrons = 32

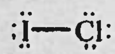
d. H provides 1; Cl provides 7; each O provides 6; total valence electrons = 32

109.

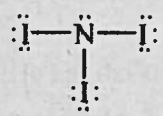
a. GeH_4: Ge provides 4 valence electrons. Each H provides 1 valence electron. Total valence electrons = 8



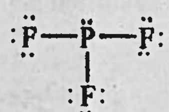
b. ICl: I provides 7 valence electrons. Cl provides 7 valence electrons. Total valence electrons = 14



c. NI_3: N provides 5 valence electrons. Each I provides 7 valence electrons. Total valence electrons = 26

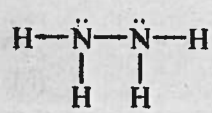


d. PF_3: P provides 5 valence electrons. Each F provides 7 valence electrons. Total valence electrons = 26

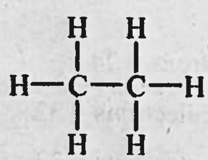


110.

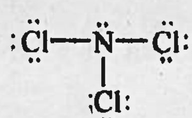
a. N_2H_4: Each N provides 5 valence electrons. Each H provides 1 valence electron. Total valence electrons = 14



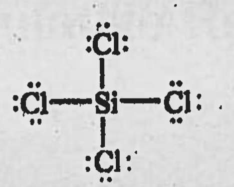
b. C_2H_6: Each C provides 4 valence electrons. Each H provides 1 valence electron. Total valence electrons = 14



c. NCl_3: N provides 5 valence electrons. Each Cl provides 7 valence electrons. Total valence electrons = 26

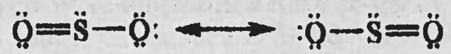


d. **$SiCl_4$: Si provides 4 valence electrons. Each Cl provides 7 valence electrons. Total valence electrons = 32**

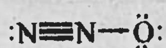


111.

a. **SO_2: S provides 6 valence electrons. Each O provides 6 valence electrons. Total valence electrons = 18**



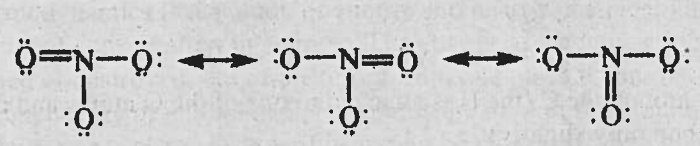
b. **N_2O: Each N provides 5 valence electrons. O provides 6 valence electrons. Total valence electrons = 16**



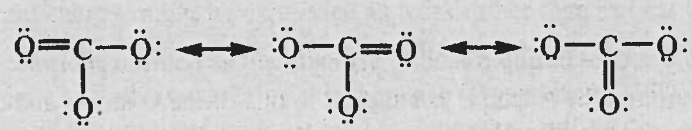
c. **O_3: Each O provides 6 valence electrons. Total valence electrons = 18**

112.

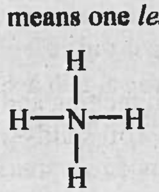
a. **NO_3^-: N provides 5 valence electrons. Each O provides 6 valence electrons. The 1– charge means one additional valence electron. Total valence electrons = 24**



b. **CO_3^{2-}: C provides 4 valence electrons. Each O provides 6 valence electrons. The 2– charge means two additional valence electrons. Total valence electrons = 24**



c. **NH_4^+: N provides 5 valence electrons. Each H provides 1 valence electron. The 1+ charge means one *less* valence electron. Total valence electrons = 8**



113. Beryllium atoms only have *two* valence electrons. In BeF_2, there are single bonds between the beryllium atom and each fluorine atom. The beryllium atom of BeF_2 thus has two electron pairs around it that lie 180° apart from one another. For the water molecule, in addition to the bonding pairs of electrons that attach the hydrogen atoms to the oxygen atoms, there are two nonbonding pairs of electrons that affect the H–O–H bond angle.

114.

a. four electron pairs arranged tetrahedrally about C

b. four electron pairs arranged tetrahedrally about Ge

c. three electron pairs arranged trigonally (planar) around B

115.

a. nonlinear (*V*-shaped due to lone pairs on O)

b. nonlinear (*V*-shaped due to lone pairs on O)

c. tetrahedral

116.

a. ClO_3^-, trigonal pyramid (lone pair on Cl)

b. ClO_2^-, nonlinear (*V*-shaped, two lone pairs on Cl)

c. ClO_4^-, tetrahedral (all pairs on Cl are bonding)

117.

a. $< 109.5°$ (molecule is nonlinear, *V*-shaped)

b. 109.5° (molecule is tetrahedral)

c. 180° (molecule is linear)

d. 120° (molecule is trigonal planar)

118.

a. nonlinear (*V*–shaped)

b. trigonal planar

c. basically trigonal planar around the C (the H is attached to one of the O atoms and distorts the shape around the carbon only slightly)

d. linear

119.

a. trigonal planar

b. basically trigonal planar around the N (the H is attached to one of the O atoms and distorts the shape around the nitrogen only slightly)

c. nonlinear (*V*-shaped)

d. linear

120. Ionic compounds tend to be hard, crystalline substances with relatively high melting and boiling points. Covalently bonded substances tend to be gases, liquids, or relatively soft solids with much lower melting and boiling points.

121. In a covalent bond between two atoms of the same element, the electron pair is shared equally and the bond is nonpolar; with a bond between atoms of different elements, the electron pair is unequally shared, and the bond is polar (assuming the elements have different electronegativities).