Chemistry 2e
7: Chemical Bonding and Molecular Geometry
7.1: Ionic Bonding

1. Does a cation gain protons to form a positive charge or does it lose electrons?

Solution

The protons in the nucleus do not change during normal chemical reactions. Only the outer electrons move. Positive charges form when electrons are lost.

3. Which of the following atoms would be expected to form negative ions in binary ionic compounds and which would be expected to form positive ions: P, I, Mg, Cl, In, Cs, O, Pb, Co?

Solution

P, I, Cl, and O would form anions because they are nonmetals. Mg, In, Cs, Pb, and Co would form cations because they are metals.

5. Predict the charge on the monatomic ions formed from the following atoms in binary ionic compounds:
(a) P
(b) Mg
(c) Al
(d) O
(e) Cl
(f) Cs

Solution

(a) P^{3-}; (b) Mg^{2+}; (c) Al^{3+}; (d) O^{2-}; (e) Cl^{-}; (f) Cs^{+}

7. Write the electron configuration for each of the following ions:
(a) As^{3-}
(b) I^{-}
(c) Be^{2+}
(d) Cd^{2+}
(e) O^{2-}
(f) Ga^{3+}
(g) Li^{+}
(h) N^{3-}
(i) Sn^{2+}
(j) Co^{2+}
(k) Fe^{2+}
(l) As^{3+}

Solution

(a) $[Ar]4s^2 3d^{10} 4p^6$; (b) $[Kr]4d^{10} 5s^2 5p^6$; (c) $1s^2$; (d) $[Kr]4d^{10}$; (e) $[He]2s^2 2p^6$; (f) $[Ar]3d^{10}$; (g) $1s^2$; (h) $[He]2s^2 2p^6$; (i) $[Kr]d^{10} 5s^2$; (j) $[Ar]3d^7$; (k) $[Ar]3d^6$; (l) $[Ar]3d^{10} 4s^2$

9. Write out the full electron configuration for each of the following atoms and for the monatomic ion found in binary ionic compounds containing the element:
(a) Al
(b) Br
(c) Sr
(d) Li
(e) As
(f) S

Solution
(a) $1s^2 2s^2 2p^6 3s^2 3p^1$; Al^{3+}: $1s^2 2s^2 2p^6$; (b) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$; $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$; (c) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^2$; Sr^{2+}:$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$; (d) $1s^2 2s^1$; Li^+ $1s^2$; (e) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^3$; $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$; (f) $1s^2 2s^2 2p^6 3s^2 3p^4$; $1s^2 2s^2 2p^6 3s^2 3p^6$

Chemistry 2e
7: Chemical Bonding and Molecular Geometry
7.2: Covalent Bonding

11. Why is it incorrect to speak of a molecule of solid NaCl?

Solution
NaCl consists of discrete ions arranged in a crystal lattice, not covalently bonded molecules.

13. Predict which of the following compounds are ionic and which are covalent, based on the location of their constituent atoms in the periodic table:
(a) Cl_2CO
(b) MnO
(c) NCl_3
(d) $CoBr_2$
(e) K_2S
(f) CO
(g) CaF_2
(h) HI
(i) CaO
(j) IBr
(k) CO_2

Solution
ionic: (b), (d), (e), (g), and (i); covalent: (a), (c), (f), (h), (j), and (k)

15. From its position in the periodic table, determine which atom in each pair is more electronegative:
(a) Br or Cl
(b) N or O
(c) S or O
(d) P or S
(e) Si or N
(f) Ba or P
(g) N or K

Solution
(a) Cl; (b) O; (c) O; (d) S; (e) N; (f) P; (g) N

17. From their positions in the periodic table, arrange the atoms in each of the following series in order of increasing electronegativity:
(a) C, F, H, N, O
(b) Br, Cl, F, H, I
(c) F, H, O, P, S
(d) Al, H, Na, O, P
(e) Ba, H, N, O, As

Solution
(a) H, C, N, O, F; (b) H, I, Br, Cl, F; (c) H, P, S, O, F; (d) Na, Al, H, P, O; (e) Ba, H, As, N, O

19. Which atoms can bond to sulfur so as to produce a positive partial charge on the sulfur atom?

Solution
N, O, F, and Cl

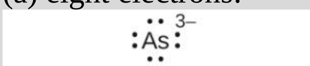
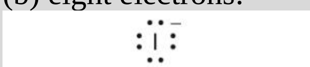
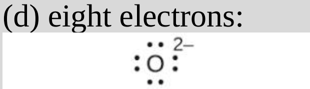
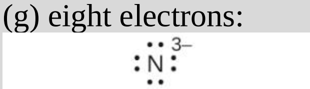
21. Identify the more polar bond in each of the following pairs of bonds:
(a) HF or HCl

(b) NO or CO
(c) SH or OH
(d) PCl or SCl
(e) CH or NH
(f) SO or PO
(g) CN or NN

Solution
(a) HF; (b) CO; (c) OH; (d) PCl; (e) NH; (f) PO; (g) CN

Chemistry 2e
7: Chemical Bonding and Molecular Geometry
7.3: Lewis Symbols and Structures

23. Write the Lewis symbols for each of the following ions:
(a) As^{3-}
(b) I^{-}
(c) Be^{2+}
(d) O^{2-}
(e) Ga^{3+}
(f) Li^{+}
(g) N^{3-}

Solution
(a) eight electrons:

;
(b) eight electrons:

;
(c) no electrons
Be^{2+};
(d) eight electrons:

;
(e) no electrons
Ga^{3+};
(f) no electrons
Li^{+};
(g) eight electrons:


25. Write the Lewis symbols of the ions in each of the following ionic compounds and the Lewis symbols of the atom from which they are formed:
(a) MgS
(b) Al_2O_3
(c) $GaCl_3$
(d) K_2O
(e) Li_3N
(f) KF

Solution
(a)

;
(b)

;
(c)

;

(d)

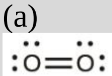
;
(e)

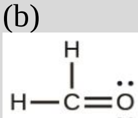
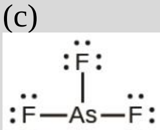
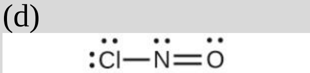
;
(f)

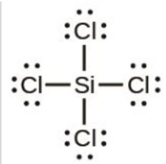

27. Write the Lewis structure for the diatomic molecule P_2, an unstable form of phosphorus found in high-temperature phosphorus vapor.

Solution

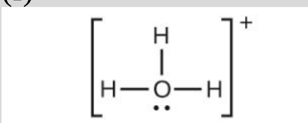

29. Write Lewis structures for the following:
(a) O_2
(b) H_2CO
(c) AsF_3
(d) ClNO
(e) $SiCl_4$
(f) H_3O^+
(g) NH_4^+
(h) BF_4^-
(i) HCCH
(j) ClCN
(k) C_2^{2+}

Solution
(a)


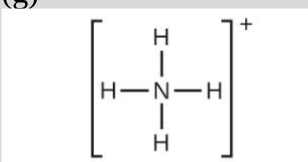
In this case, the Lewis structure is inadequate to depict the fact that experimental studies have shown two unpaired electrons in each oxygen molecule.
(b)

;
(c)

;
(d)

;
(e)

;

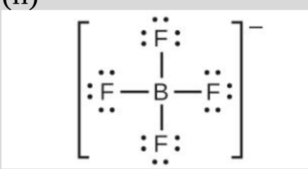
(f)

;

(g)

;

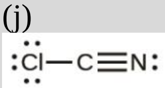
(h)

;

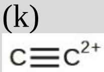
(i)

;

(j)

;

(k)



31. Write Lewis structures for the following:

(a) SeF_6

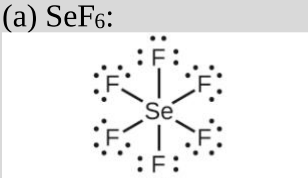
(b) XeF_4

(c) $SeCl_3^+$

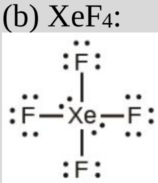
(d) Cl_2BBCl_2 (contains a B–B bond)

Solution

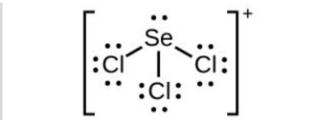
(a) SeF_6:

;

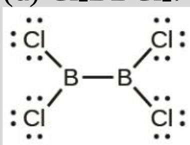
(b) XeF_4:

;

(c) $SeCl_3^+$:


;

(d) Cl_2BBCl_2:



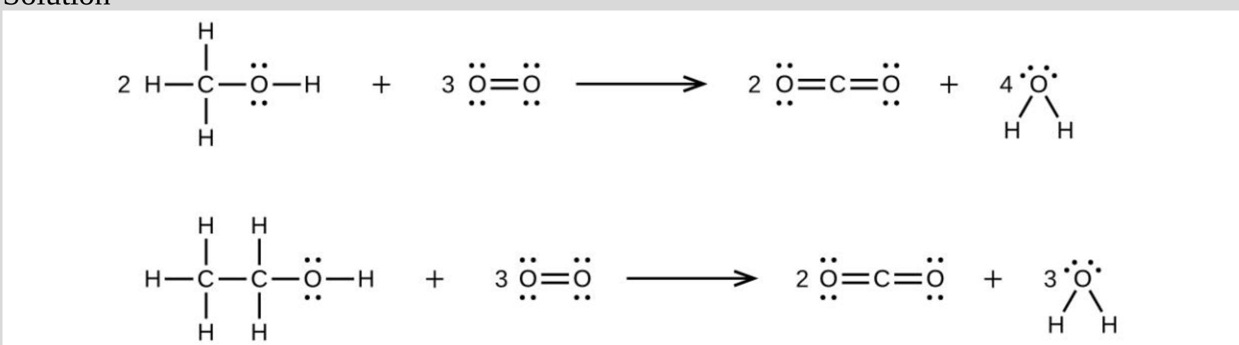
33. Correct the following statement: “The bonds in solid $PbCl_2$ are ionic; the bond in a HCl molecule is covalent. Thus, all of the valence electrons in $PbCl_2$ are located on the $Cl^–$ ions, and all of the valence electrons in a HCl molecule are shared between the H and Cl atoms.”

Solution

Two valence electrons per Pb atom are transferred to Cl atoms; the resulting Pb^{2+} ion has a $6s^2$ valence shell configuration. Two of the valence electrons in the HCl molecule are shared, and the other six are located on the Cl atom as lone pairs of electrons.

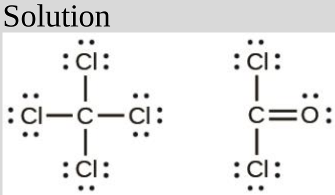
35. Methanol, H_3COH, is used as the fuel in some race cars. Ethanol, C_2H_5OH, is used extensively as motor fuel in Brazil. Both methanol and ethanol produce CO_2 and H_2O when they burn. Write the chemical equations for these combustion reactions using Lewis structures instead of chemical formulas

Solution



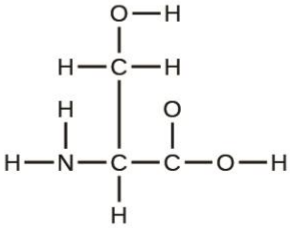
37. Carbon tetrachloride was formerly used in fire extinguishers for electrical fires. It is no longer used for this purpose because of the formation of the toxic gas phosgene, Cl_2CO. Write the Lewis structures for carbon tetrachloride and phosgene.

Solution

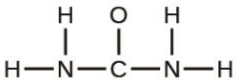


39. The arrangement of atoms in several biologically important molecules is given below. Complete the Lewis structures of these molecules by adding multiple bonds and lone pairs. Do not add any more atoms.

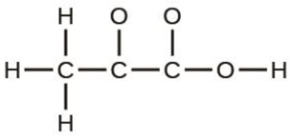
(a) the amino acid serine:



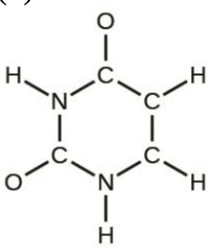
(b) urea:



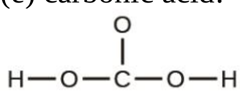
(c) pyruvic acid:



(d) uracil:

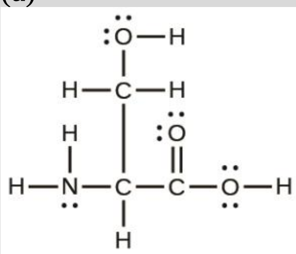


(e) carbonic acid:

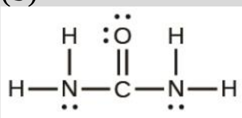


Solution

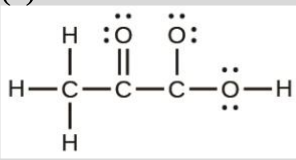
(a)

;

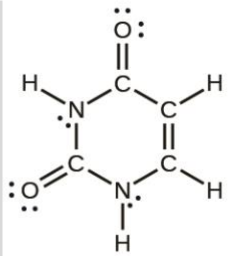
(b)

;

(c)

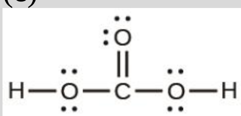
;

(d)



;

(e)



41. A compound with a molar mass of about 42 g/mol contains 85.7% carbon and 14.3% hydrogen by mass. Write the Lewis structure for a molecule of the compound.

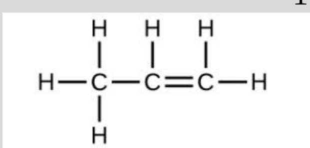
Solution

A 100.0-g sample of this compound would contain 85.7 g C and 14.3 g H:

$$\frac{85.7\text{ g}}{12.011\text{ g mol}^{-1}} = 7.14\text{ mol C}$$

$$\frac{14.3\text{ g}}{1.00794\text{ g mol}^{-1}} = 14.19\text{ mol H}.$$

This is a ratio of 2 H to 1 C, or an empirical formula of CH_2 with a formula mass of approximately 14. As $\frac{42}{14} = 3$, the formula is 3 × CH_2 or C_3H_6. The Lewis structure is:



43. How are single, double, and triple bonds similar? How do they differ?

Solution

Each bond includes a sharing of electrons between atoms. Two electrons are shared in a single bond; four electrons are shared in a double bond; and six electrons are shared in a triple bond.

Chemistry 2e
7: Chemical Bonding and Molecular Geometry
7.4: Formal Charges and Resonance

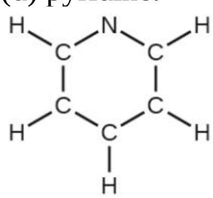
45. Write resonance forms that describe the distribution of electrons in each of the molecules or ions given below:

(a) sulfur dioxide, SO_2

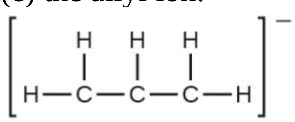
(b) carbonate ion, CO_3^{2-}

(c) hydrogen carbonate ion, HCO_3^{-} (C is bonded to an OH group and two O atoms)

(d) pyridine:



(e) the allyl ion:

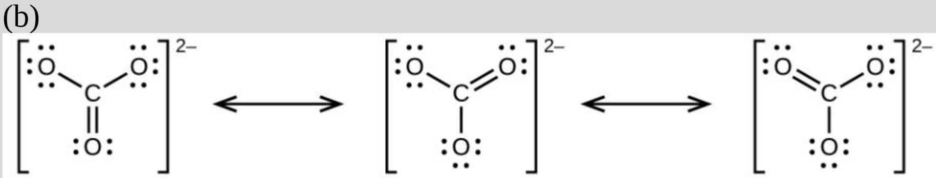


Solution

(a)

;

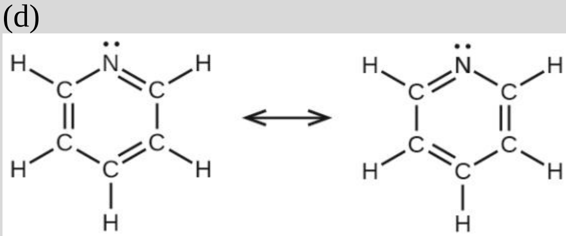
(b)

;

(c)

;

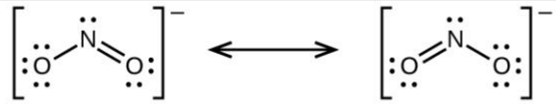
(d)

;

(e)

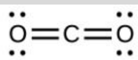
47. Sodium nitrite, which has been used to preserve bacon and other meats, is an ionic compound. Write the resonance forms of the nitrite ion, NO_2^{-}.

Solution

For NO_2^- 

49. Write the Lewis structures for the following, and include resonance structures where appropriate. Indicate which of the three has the strongest carbon-oxygen bond.
(a) CO_2
(b) CO

Solution
(a)



(b)

CO has the strongest carbon-oxygen bond, because there are is a triple bond joining C and O. CO_2 has double bonds, and carbonate has 1.3 bonds.

51. Determine the formal charge of each element in the following:
(a) HCl
(b) CF_4
(c) PCl_3
(d) PF_5

Solution

	Element	Bonding Electrons	Nonbonded Electrons	Valence Electrons	Formal Charge
(a)	H Cl	1 1	0 6	1 7	0 0
(b)	C F	4 1	0 6	4 7	0 0
(c)	P Cl	3 1	2 6	5 7	0 0
(d)	P F	5 1	0 6	5 7	0 0

53. Calculate the formal charge of chlorine in the molecules Cl_2, $BeCl_2$, and ClF_5.

Solution

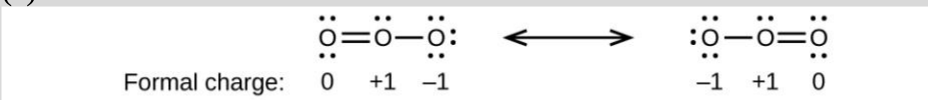
	Element	Bonding Electrons	Nonbonded Electrons	Valence Electrons	Formal Charge
Cl_2	Cl	1	6	7	0
$BeCl_2$	Be Cl	2 1	0 6	2 7	0 0
ClF_5	Cl F	5 1	2 6	7 7	0 0

55. Draw all possible resonance structures for each of the compounds below. Determine the formal charge on each atom in each of the resonance structures:
(a) O_3
(b) SO_2
(c) NO_2^-

(d) NO_3^-

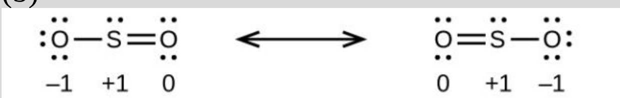
Solution

(a)



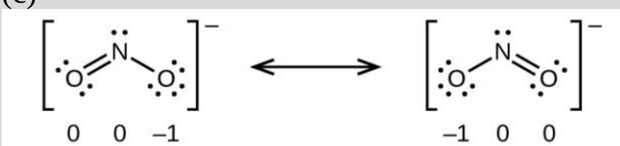
;

(b)



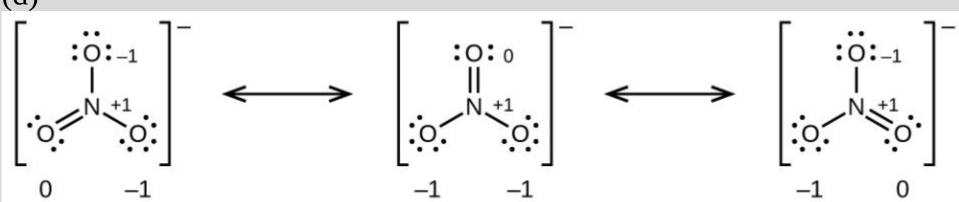
;

(c)



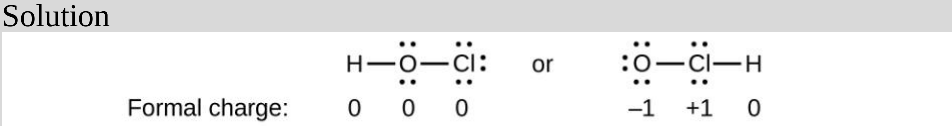
;

(d)



57. Based on formal charge considerations, which of the following would likely be the correct arrangement of atoms in hypochlorous acid: HOCl or OClH?

Solution

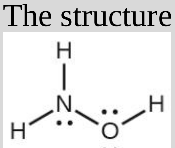


The structure with formal charges of 0 is the most stable and would therefore be the correct arrangement of atoms.

59. Draw the structure of hydroxylamine, H_3NO, and assign formal charges; look up the structure. Is the actual structure consistent with the formal charges?

Solution

The structure that gives zero formal charges is consistent with the actual structure:



61. Write the Lewis structure and chemical formula of the compound with a molar mass of about 70 g/mol that contains 19.7% nitrogen and 80.3% fluorine by mass, and determine the formal charge of the atoms in this compound.

Solution

There are 19.7 g N and 80.3 g F in a 100.0-g sample:

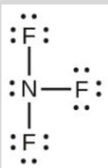
$$\frac{19.7\text{ g}}{14.0067\text{ g mol}^{-1}} = 1.406\text{ mol}$$

$$\frac{1.406\text{ mol}}{1.406\text{ mol}} = 1\text{ N}$$

$$\frac{80.3\text{ g}}{18.9984\text{ g mol}^{-1}} = 4.2267\text{ mol}$$

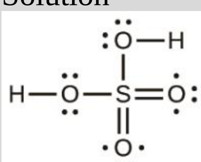
$$\frac{4.2267\text{ mol}}{1.406\text{ mol}} = 3\text{ F}$$

The empirical formula is NF_3 and its molar mass is 71.00 g/mol, which is consistent with the stated molar mass.
Oxidation states: N = +3, F = –1.
Formal charges: N = 0, F = 0:



63. Sulfuric acid is the industrial chemical produced in greatest quantity worldwide. About 90 billion pounds are produced each year in the United States alone. Write the Lewis structure for sulfuric acid, H_2SO_4, which has two oxygen atoms and two OH groups bonded to the sulfur.

Solution



Chemistry 2e
7: Chemical Bonding and Molecular Geometry
7.5: Strengths of Ionic and Covalent Bonds

65. Using the bond energies in Table 7.2, determine the approximate enthalpy change for each of the following reactions:

(a) $H_2(g) + Br_2(g) \longrightarrow 2HBr(g)$

(b) $CH_4(g) + I_2(g) \longrightarrow CH_3I(g) + HI(g)$

(c) $C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(g)$

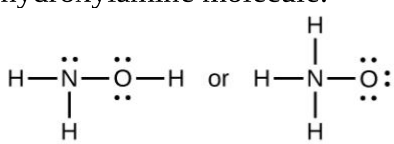
Solution

(a) $$\begin{aligned}\Delta H^\circ_{298} &= \sum D_{\text{bonds broken}} - \sum D_{\text{bonds formed}} \\ &= D_{\text{H−H}} + D_{\text{Br−Br}} - 2D_{\text{H−Br}} \\ &= 436 + 190 - 2(370) = -114\ \text{kJ}\end{aligned};$$

(b) $$\begin{aligned}\Delta H^\circ_{298} &= \sum D_{\text{bonds broken}} - \sum D_{\text{bonds formed}} \\ &= 4D_{\text{C−H}} + D_{\text{I−I}} - 3D_{\text{C−H}} - D_{\text{C−I}} - D_{\text{H−I}} \\ &= 4(415) + 150 - 3(415) - 240 - 295 = 30\ \text{kJ}\end{aligned};$$

(c) $$\begin{aligned}\Delta H^\circ_{298} &= \sum D_{\text{bonds broken}} - \sum D_{\text{bonds formed}} \\ &= D_{\text{C=C}} + 4D_{\text{C−H}} + 3D_{\text{O=O}} - 4D_{\text{C=O}} - 4D_{\text{O−H}} \\ &= 611 + 4(415) + 3(498) - 4(741) - 4(464) \\ &= -1055\ \text{kJ}\end{aligned}$$

67. When a molecule can form two different structures, the structure with the stronger bonds is usually the more stable form. Use bond energies to predict the correct structure of the hydroxylamine molecule:



Solution

(a) $$\begin{aligned}&2\ \text{N−H bonds} = 2(390) \\ &1\ \text{N−O bond} = \quad 200 \\ &1\ \text{O−H bond} = \underline{\quad 464} \\ &\qquad\qquad\qquad 1444\ \text{kJ}\end{aligned}$$; (b) $$\begin{aligned}&3\ \text{N−H bonds} = 3(390) \\ &1\ \text{N−O bond} = \underline{\quad 200} \\ &\qquad\qquad\qquad 1370\ \text{kJ}\end{aligned}$$; the greater bond energy is for (a), and it is more stable

69. Using the standard enthalpy of formation data in Appendix G, show how can the standard enthalpy of formation of HCl(g) can be used to determine the bond energy.

Solution

$$\text{HCl}(g) \longrightarrow \frac{1}{2}\text{H}_2(g) + \frac{1}{2}\text{Cl}_2(g) \qquad \Delta H_1^\circ = -\Delta H^\circ_{\text{f [HCl}(g)]}$$

$$\frac{1}{2}\text{H}_2(g) \longrightarrow \text{H}(g) \qquad \Delta H_2^\circ = \Delta H^\circ_{\text{f [H}(g)]}$$

$$\frac{1}{2}\text{Cl}_2(g) \longrightarrow \text{Cl}(g) \qquad \Delta H_3^\circ = \Delta H^\circ_{\text{f [Cl}(g)]}$$

$$\text{HCl}(g) \longrightarrow \text{H}(g) + \text{Cl}(g) \qquad \Delta H^\circ_{298} = \Delta H_1^\circ + \Delta H_2^\circ + \Delta H_3^\circ$$

$$D_{\text{HCl}} = \Delta H^\circ_{298} = -\Delta H^\circ_{\text{f [HCl}(g)]} + \Delta H^\circ_{\text{f [H}(g)]} + \Delta H^\circ_{\text{f [Cl}(g)]}$$

$$= -(-92.307 \text{ kJ}) + 217.97 \text{ kJ} + 121.3 \text{ kJ}$$

$$= 431.6 \text{ kJ}$$

71. Using the standard enthalpy of formation data in Appendix G, determine which bond is stronger: the S–F bond in $SF_4(g)$ or in $SF_6(g)$?

Solution

$$\text{SF}_4(g) \longrightarrow \frac{1}{8}\text{S}_8(s) + 2\text{F}_2(g) \qquad \Delta H_1^\circ = -\Delta H_{\text{f[SF}_4(g)]}$$

$$\frac{1}{8}\text{S}_8(s) \longrightarrow \text{S}(g) \qquad \Delta H_2^\circ = \Delta H^\circ_{\text{f S}(g)}$$

$$2\text{F}_2(g) \longrightarrow 4\text{F}(g) \qquad 4\Delta H_3^\circ = 4\Delta H^\circ_{\text{f F}(g)}$$

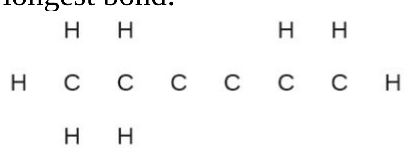
$$D_{\text{SF}_4} = \Delta H^\circ_{298} = -\Delta H_{\text{f[SF}_4(g)]} + \Delta H^\circ_{\text{f S}(g)} + 4\Delta H^\circ_{\text{f F}(g)}$$

$$= 728.43 + 278.81 + 4(79.4) = 1369.7 \text{ kJ}$$

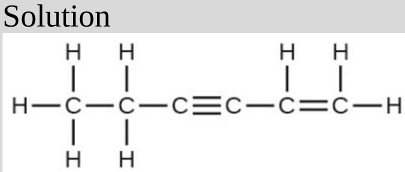
$$D_{\text{S–F}} = \frac{1324.84 \text{ kJ}}{4 \text{ S} - \text{F}} = 331.21 \text{ kJ}$$

Proceeding in the same manner, $-\Delta H_{\text{f[SF}_6(g)]} = 1220.5 \text{ kJ mol}^{-1}$. The 6F($g$) and S($g$) contribute 755.21 kJ; then $\text{DF}_{\text{SF}_6} = 1962.7 \text{ kJ}$ and $D_{\text{S–F}} = \frac{1975.71}{6} = 392.3 \text{ kJ mol}^{-1}$ per mole of bonds. The S–F bond in SF_4 is stronger.

73. Complete the following Lewis structure by adding bonds (not atoms), and then indicate the longest bond:



Solution



The C–C single bonds are longest.

75. Use principles of atomic structure to answer each of the following:[1]

1This question is taken from the Chemistry Advanced Placement Examination and is used with the permission of the Educational Testing Service.

(a) The radius of the Ca atom is 197 pm; the radius of the Ca^{2+} ion is 99 pm. Account for the difference.
(b) The lattice energy of CaO(s) is –3460 kJ/mol; the lattice energy of K_2O is –2240 kJ/mol. Account for the difference.
(c) Given these ionization values, explain the difference between Ca and K with regard to their first and second ionization energies.

Element	First Ionization Energy (kJ/mol)	Second Ionization Energy (kJ/mol)
K	419	3050
Ca	590	1140

(d) The first ionization energy of Mg is 738 kJ/mol and that of Al is 578 kJ/mol. Account for this difference.

Solution
(a) When two electrons are removed from the valence shell, the Ca radius loses the outermost energy level and reverts to the lower *n* = 3 level, which is much smaller in radius.
(b) The +2 charge on calcium pulls the oxygen much closer compared with K, thereby increasing the lattice energy relative to a less charged ion.
(c) Removal of the 4*s* electron in Ca requires more energy than removal of the 4*s* electron in K, because of the stronger attraction of the nucleus and the extra energy required to break the pairing of the electrons. The second ionization energy for K requires that an electron be removed from a lower energy level, where the attraction is much stronger from the nucleus for the electron. In addition, energy is required to unpair two electrons in a full orbital. For Ca, the second ionization potential requires removing only a lone electron in the exposed outer energy level.
(d) In Al, the removed electron is relatively unprotected and unpaired in a *p* orbital. The higher energy for Mg mainly reflects the unpairing of the 2*s* electron.

77. For which of the following substances is the least energy required to convert one mole of the solid into separate ions?
(a) MgO
(b) SrO
(c) KF
(d) CsF
(e) MgF_2

Solution
The lattice energy, *U*, is the energy required to convert the solid into separate ions. *U* may be calculated from the Born-Haber cycle.
The values in kJ/mol are approximately (a) 3791; (b) 3223; (c) 821; (d) 740; and (e) 2957.
The answer is (d), which requires about 740 kJ/mol.

79. The lattice energy of LiF is 1023 kJ/mol, and the Li–F distance is 201 pm. MgO crystallizes in the same structure as LiF but with a Mg–O distance of 205 pm. Which of the following values most closely approximates the lattice energy of MgO: 256 kJ/mol, 512 kJ/mol, 1023 kJ/mol, 2046 kJ/mol, or 4090 kJ/mol? Explain your choice.

Solution
4008 kJ/mol; both ions in MgO have twice the charge of the ions in LiF; the bond length is very similar and both have the same structure; a quadrupling of the energy is expected based on the equation for lattice energy

81. Which compound in each of the following pairs has the larger lattice energy? Note: Ba^{2+} and K^+ have similar radii; S^{2-} and Cl^- have similar radii. Explain your choices.
(a) K_2O or Na_2O
(b) K_2S or BaS
(c) KCl or BaS
(d) BaS or $BaCl_2$
Solution
(a) Na_2O; Na^+ has a smaller radius than K^+; (b) BaS; Ba has a larger charge than K; (c) BaS; Ba and S have larger charges; (d) BaS; S has a larger charge
83. Which of the following compounds requires the most energy to convert one mole of the solid into separate ions?
(a) K_2S
(b) K_2O
(c) CaS
(d) Cs_2S
(e) CaO
Solution
CaO

Chemistry 2e
7: Chemical Bonding and Molecular Geometry
7.6: Molecular Structure and Polarity

85. Explain why the HOH molecule is bent, whereas the HBeH molecule is linear.

Solution

The placement of the two sets of unpaired electrons in water forces the bonds to assume a tetrahedral arrangement, and the resulting HOH molecule is bent. The HBeH molecule (in which Be has only two electrons to bond with the two electrons from the hydrogens) must have the electron pairs as far from one another as possible and is therefore linear.

87. Explain the difference between electron-pair geometry and molecular structure.

Solution

Space must be provided for each pair of electrons whether they are in a bond or are present as lone pairs. Electron-pair geometry considers the placement of all electrons. Molecular structure considers only the bonding-pair geometry.

89. Explain how a molecule that contains polar bonds can be nonpolar.

Solution

As long as the polar bonds are compensated (for example. two identical atoms are found directly across the central atom from one another), the molecule can be nonpolar.

91. Predict the electron pair geometry and the molecular structure of each of the following molecules or ions:

(a) SF_6

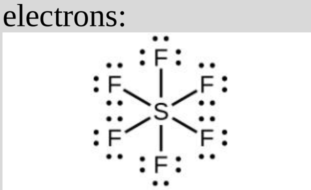
(b) PCl_5

(c) BeH_2

(d) CH_3^+

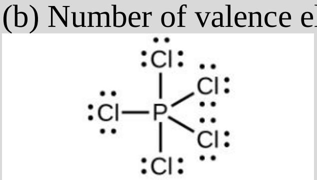
Solution

(a) Number of valence electrons: S = 6, F = 7 each, total 48. A single line bond represents two electrons:

.

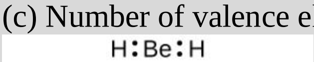
The total number of electrons used is 48; six bonds are formed and no nonbonded pairs exist. Therefore the molecule includes six regions of electron density and, from the table, the electron geometry is octahedral. Since no lone pairs exist, the electron geometry and molecular structure are the same.

(b) Number of valence electrons: P = 5, Cl = 7 each, total 40:

.

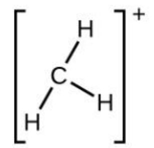
The total number of electrons is 40; there are five regions of electron density and, from the table, the geometry is trigonal bipyramid. Since no lone pairs exist on P, the electron geometry and molecular structure are the same.

(c) Number of valence electrons: Be = 2, H = 1 each, total 4:

.

There are only two regions of electron density and they must have a linear arrangement. These regions also correspond to the location of the bonds. Both the electron and molecular structures are linear.

(d) Number of valence electrons: C = 4, H = 1 each, less one electron because of the positive charge, for a total of six electrons:



.

There are three regions of electron density coincident with the three bonds. Therefore the shape is trigonal planar for both the electron geometry and molecular structure.

93. What are the electron-pair geometry and the molecular structure of each of the following molecules or ions?

(a) ClF_5
(b) ClO_2^-
(c) $TeCl_4^{2-}$
(d) PCl_3
(e) SeF_4
(f) PH_2^-

Solution

		Regions of High Electron Density		Structure	
Formula	Electrons	Total	Lone pairs	Electron	Molecular
(a) ClF_5	42	6	1	octahedral	square pyramidal
(b) ClO_2^-	20	4	2	tetrahedral	bent
(c) $TeCl_4^{2-}$	36	6	2	octahedral	square planar
(d) PCl_3	26	4	1	tetrahedral	trigonal pyramidal
(e) SeF_4	34	5	1	trigonal bipyramidal	seesaw
(f) PH_2^-	8	4	2	tetrahedral	bent (109°)
(g) SF_2	20	2	2	tetrahedral	bent (109°)

95. Identify the electron pair geometry and the molecular structure of each of the following molecules:

(a) ClNO (N is the central atom)
(b) CS_2
(c) Cl_2CO (C is the central atom)
(d) Cl_2SO (S is the central atom)
(e) SO_2F_2 (S is the central atom)
(f) XeO_2F_2 (Xe is the central atom)
(g) $ClOF_2^+$ (Cl is the central atom)

Solution

		Regions of High Electron Density		Structure	
Formula	Electrons	Total	Lone	Electron	Molecular

			pairs		
(a) ClNO	18	3	1	trigonal planar	bent (120°)
(b) CS_2	16	2	0	linear	linear
(c) Cl_2CO	24	3	0	trigonal planar	trigonal planar
(d) Cl_2SO	26	4	1	tetrahedral	trigonal pyramidal
(e) SO_2F_2	32	4	0	tetrahedral	tetrahedral
(f) XeO_2F_2	34	5	1	trigonal bipyramidal	seesaw
(g) $ClOF_2^+$	26	4	1	tetrahedral	trigonal pyramidal

97. Which of the following molecules and ions contain polar bonds? Which of these molecules and ions have dipole moments?
(a) ClF_5
(b) ClO_2^-
(c) $TeCl_4^{2-}$
(d) PCl_3
(e) SeF_4
(f) PH_2^-
(g) XeF_2

Solution
All of these molecules and ions contain polar bonds. Only ClF_5, ClO_2^-, PCl_3, SeF_4, and PH_2^- have dipole moments.

99. Which of the following molecules have dipole moments?
(a) CS_2
(b) SeS_2
(c) CCl_2F_2
(d) PCl_3 (P is the central atom)
(e) ClNO (N is the central atom)

Solution
(a) CS_2 is linear and has no dipole moment. (b) SeS_2 is bent. This leads to an overall dipole moment. (c) The C–Cl and C–F bonds are not balanced—that is, the dipoles do not completely cancel. Therefore, it has a dipole moment. (d) PCl3 is trigonal pyramidal. Due to this shape, the dipoles of the bonds do not cancel and there is an overall dipole moment. (e) The ClNO molecule is bent, leading to a dipole moment.

101. The molecule XF_3 has a dipole moment. Is X boron or phosphorus?

Solution
P

103. Is the Cl_2BBCl_2 molecule polar or nonpolar?

Solution
nonpolar

105. Describe the molecular structure around the indicated atom or atoms:
(a) the sulfur atom in sulfuric acid, H_2SO_4 [$(HO)_2SO_2$]
(b) the chlorine atom in chloric acid, $HClO_3$ [$HOClO_2$]
(c) the oxygen atom in hydrogen peroxide, HOOH
(d) the nitrogen atom in nitric acid, HNO_3 [$HONO_2$]

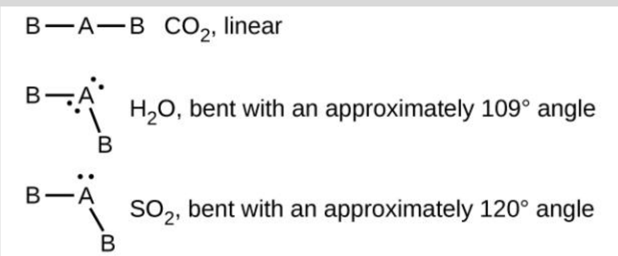
(e) the oxygen atom in the OH group in nitric acid, HNO_3 [$HONO_2$]
(f) the central oxygen atom in the ozone molecule, O_3
(g) each of the carbon atoms in propyne, CH_3CCH
(h) the carbon atom in Freon, CCl_2F_2
(i) each of the carbon atoms in allene, H_2CCCH_2

Solution

(a) tetrahedral; (b) trigonal pyramidal; (c) bent (109°); (d) trigonal planar; (e) bent (109°); (f) bent (109°); (g) $\underline{C}H_3CCH$ tetrahedral, $CH_3\underline{CC}H$ linear; (h) tetrahedral, (i) $H_2C\underline{C}CH_2$ linear; $H_2\underline{C}C\underline{C}H_2$ trigonal planar

107. A molecule with the formula AB_2, in which A and B represent different atoms, could have one of three different shapes. Sketch and name the three different shapes that this molecule might have. Give an example of a molecule or ion for each shape.

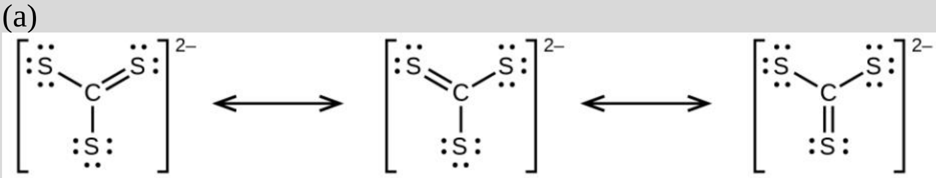
Solution



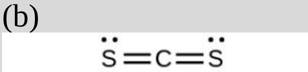
109. Draw the Lewis electron dot structures for these molecules, including resonance structures where appropriate:
(a) CS_3^{2-}
(b) CS_2
(c) CS
(d) predict the molecular shapes for CS_3^{2-} and CS_2 and explain how you arrived at your predictions

Solution

(a)


;

(b)


;

(c)

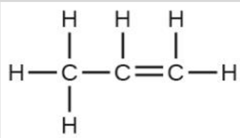

;

(d) CS_3^{2-} includes three regions of electron density (all are bonds with no lone pairs); the shape is trigonal planar; CS_2 has only two regions of electron density (all bonds with no lone pairs); the shape is linear

111. A compound with a molar mass of about 42 g/mol contains 85.7% carbon and 14.3% hydrogen. What is its molecular structure?

Solution

The empirical formula is CH_2 with a unit mass of14. $\frac{42}{14} = 3$. Therefore, the Lewis structure is made from three units, but the atoms must be rearranged:



113. Use this link: http://phet.colorado.edu/en/simulation/molecule-polarity to perform the following exercises for a real molecule. You may need to rotate the molecules in three dimensions to see certain dipoles.
(a) Sketch the bond dipoles and molecular dipole (if any) for O_3. Explain your observations.
(b) Look at the bond dipoles for NH_3. Use these dipoles to predict whether N or H is more electronegative.
(c) Predict whether there should be a molecular dipole for NH_3 and, if so, in which direction it will point. Check the molecular dipole box to test your hypothesis.

Solution
The molecular dipole points away from the hydrogen atoms.

115. Use this link: http://phet.colorado.edu/en/simulation/molecule-shapes to explore real molecules. On the Real Molecules tab, select H_2O. Switch between the “real” and “model” modes. Explain the difference observed.

Solution
The structures are very similar. In the model mode, each electron group occupies the same amount of space, so the bond angle is shown as 109.5°. In the “real” mode, the lone pairs are larger, causing the hydrogens to be compressed. This leads to the smaller angle of 104.5°.