CHAPTER 7. Lewis Formulas

7-1. See the text. Because silicon is below carbon in the periodic table, we can determine the Lewis formula of SiCl₄ in the same manner as for CCl₄.

7-2. Because phosphorus is the unique atom in this molecule, we shall assume that it is central and that the four chlorine atoms are attached to it.

\[
\begin{array}{c}
\text{Cl} \\
\text{Cl} \\
\text{P} \\
\text{Cl} \\
\text{Cl}
\end{array}
\]

The total number of valence electrons is \((1 \times 5) + (4 \times 7) - 1 = 32\). We use eight of the valence electrons to form phosphorus–chlorine bonds. We now place valence electrons as lone pairs on the four chlorine atoms, accounting for the remaining 24 valence electrons. The completed Lewis formula is

\[
\begin{array}{c}
\vdots \\
\vdots \\
\vdots \\
\vdots
\end{array}
\]

7-3. We first arrange the atoms as

\[
\begin{array}{c}
\text{H} \\
\text{P} \\
\text{H} \\
\text{H} \\
\text{H}
\end{array}
\]

The total number of valence electrons is \((1 \times 5) + (3 \times 1) = 8\). We use six valence electrons to form the P–H bonds. We place the remaining two valence electrons as a lone pair on the P atom. The Lewis formula is

\[
\begin{array}{c}
\vdots \\
\vdots \\H
\end{array}
\]

7-4. We first arrange the atoms as

\[
\begin{array}{c}
\text{H} \\
\text{N} \\
\text{H}
\end{array}
\]
The total number of valence electrons is $(1 \times 5) + (2 \times 1) + 1 = 8$. We use four valence electrons to form the N–H bonds and place the remaining four valence electrons as lone pairs on the N atom. The Lewis formula is

\[
\begin{array}{c}
\text{H} \quad \text{N} \quad \text{H} \\
\end{array}
\]

7-5. See the text.

7-6. The Lewis formula for the tetrafluoroborate ion is

\[
\begin{array}{c}
\text{F} \quad \text{B} \quad \text{F} \quad \text{F} \\
\end{array}
\]

The formal charges on the boron atom and fluorine atoms are calculated using Equation 7.1.

\[
\begin{align*}
\text{formal charge on the B atom} &= 3 - 0 - \frac{1}{2}(4) = -1 \\
\text{formal charge on each F atom} &= 7 - 6 - \frac{1}{2}(2) = 0
\end{align*}
\]

Thus, the Lewis formula for a BF\(_4^-\) ion with the formal charges indicated is

\[
\begin{array}{c}
\text{F} \quad \text{B} \quad \text{F} \quad \text{F} \\
\end{array}
\]

7-7. The formal charges on the oxygen and hydrogen atoms in Structure I are assigned according to Equation 7.1.

\[
\begin{align*}
\text{formal charge on end O atom} &= 6 - 6 - \frac{1}{2}(2) = -1 \\
\text{formal charge on bonded O atom} &= 6 - 2 - \frac{1}{2}(6) = +1 \\
\text{formal charge on each H atom} &= 1 - 0 - \frac{1}{2}(2) = 0
\end{align*}
\]

The formal charges in Structure II are

\[
\begin{align*}
\text{formal charge on each O atom} &= 6 - 6 - \frac{1}{2}(4) = 0 \\
\text{formal charge on each H atom} &= 1 - 0 - \frac{1}{2}(2) = 0
\end{align*}
\]

Structure II has zero formal charges everywhere and is the preferred Lewis formula for hydrogen peroxide.

7-8. (a) We first arrange the atoms as

\[
\begin{array}{c}
\text{H} \\
\text{H} \quad \text{C} \quad \text{O}
\end{array}
\]
The total number of valence electrons is \((1 \times 4) + (1 \times 6) + (2 \times 1) = 12\). As in Example 7-8, we cannot satisfy the octet rule with just single bonds. Therefore, we form a double bond between the carbon atom and the oxygen atom, form two H–C bonds, and place the remaining four valence electrons as lone pairs on the oxygen atom to write

\[
\begin{align*}
H & \quad \text{H} \\
\text{H} & \quad \text{C} \quad \text{O}_2 \\
\end{align*}
\]

(b) The Lewis formula for the methyl group is

\[
\begin{align*}
H & \quad \text{H} \\
\text{H} & \quad \text{C} \\
\text{H} & \quad \text{H}
\end{align*}
\]

Thus, the Lewis formula of ethanal is

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{C} \quad \text{O}_2 \\
\text{H} & \quad \text{H}
\end{align*}
\]

7-9. As in Example 7-9, we cannot satisfy the octet rule with all single bonds or a double bond. Therefore, we form a triple bond between the two carbon atoms and two C–H bonds to write

\[
\begin{align*}
\text{H} & \quad \text{C} \equiv \text{C} \quad \text{H}
\end{align*}
\]

7-10. We first arrange the atoms as

\[
\begin{align*}
\text{O} & \quad \text{C} \\
\text{O} & \quad \text{O}
\end{align*}
\]

The total number of valence electrons is \((1 \times 4) + (3 \times 6) + (2 \times 1) = 24\). The resonance formulas are

\[
\begin{align*}
\text{O} & \quad \text{O} \\
\quad \quad \text{O} & \quad \quad \text{O} \\
\quad \quad \text{O} & \quad \quad \text{O}
\end{align*}
\]

The formal charges indicated are calculated by using Equation 7.1.

\[
\begin{align*}
\text{formal charge on C atom} & = 4 - 0 - \frac{1}{2}(4) = 0 \\
\text{formal charge on each single bonded O atom} & = 6 - 6 - \frac{1}{2}(2) = -1 \\
\text{formal charge on double bonded O atom} & = 6 - 4 - \frac{1}{2}(4) = 0
\end{align*}
\]
The superposition of these resonance forms gives the resonance hybrid

\[
\text{O} = \text{C} = \text{O}
\]

The three carbon–oxygen bonds in a \( \text{CO}_2^- \) ion are equivalent; they have the same bond length and the same bond energy.

7-11. A nitrogen dioxide molecule has \((1 \times 5) + (2 \times 6) = 17\) valence electrons and so is a free radical. The two resonance forms of a \( \text{NO}_2 \) molecule are

7-12. We first arrange the atoms as

\[
\text{O} \quad \text{Cl} \quad \text{P} \quad \text{Cl} \\
\text{Cl}
\]

The total number of valence electrons is \((1 \times 5) + (1 \times 6) + (3 \times 7) = 32\). We use eight valence electrons to form three \( \text{P–Cl} \) bonds and one \( \text{P–O} \) bond. The remaining 24 valence electrons are lone pairs on the oxygen atom and three chlorine atoms. The Lewis formula is

Both the oxygen atom and phosphorus atom have a formal charge. We can eliminate the formal charges by expanding the the octet on the phosphorus atom to form a double bond between the phosphorus and oxygen atoms to obtain the Lewis formula

Notice that all the atoms now have a zero formal charge.

7-13. A \( \text{PCl}_6^- \) ion has \((1 \times 5) + (6 \times 7) + 1 = 48\) valence electrons. Twelve are used to form the six \( \text{P–Cl} \) bonds. The remaining 36 valence electrons are placed as lone electron pairs on the six chlorine atoms. The Lewis formula is
7-14. A $\text{PO}_4^{3-}$ ion has $(1 \times 5) + (4 \times 6) + 3 = 32$ valence electrons. The various resonance forms are

The superpositions of these resonance forms give the resonance hybrid

The four phosphorus–oxygen bonds in the $\text{PO}_4^{3-}$ ion are equivalent; they have the same bond length and bond energy.

7-15. An oxygen atom is more electronegative than a hydrogen atom, and so we have

7-16. The difference in the electronegativities of a hydrogen atom and a silicon atom is $2.1 - 1.90 = 0.2$. Thus, the bonding in a $\text{SiH}_4$ molecule is covalent with a small degree of ionic character. The difference in the electronegativities of a beryllium atom and a chlorine atom is $3.16 - 1.57 = 1.59$. Thus, the bonding in a $\text{SiH}_4$ molecule is polar covalent with a fair degree of ionic character.
CHAPTER 8. Prediction of Molecular Geometries

8-1. The Lewis formula of a CHCl₃ molecule is

\[
\text{H} - \text{C} = \text{Cl} - \text{Cl} - \text{Cl}
\]

There are four valence-shell electron pairs about the central carbon atom, so it is tetrahedral. The bond angles are 109.5°.

8-2. The Lewis formula of a BF₄⁻ molecule is

\[
\begin{array}{c}
\text{F} \\
\text{B} \\
\text{F} \\
\text{F} \\
\end{array}
\]

There are four valence-shell electron pairs about the central boron atom, and so the ion is tetrahedral.

8-3. The Lewis formula of an AlF₆⁻ ion is

\[
\begin{array}{c}
\text{F} \\
\text{Al} \\
\text{F} \\
\text{F} \\
\text{F} \\
\text{F} \\
\end{array}
\]

An AlF₆⁻ ion contains six covalent bonds, and so it is an octahedral ion with 90° bond angles.

8-4. The Lewis formula of a SF₂ molecule is

\[
\text{F} - \text{S} - \text{F}
\]

It is an AX₂E₂ molecule, and so it is bent with bond angles less than the tetrahedral bond angle of 109°.

8-5. The Lewis formula of a CS₂ molecule is

\[
\text{S} = \text{C} = \text{S}
\]

It is an AX₂ molecule, and so it is linear.
8-6. The Lewis formula of a ClO$_2^-$ ion is \( \text{O} \text{Cl O} \). It is an AX$_2$E$_2$ ion, and so it is bent.

The resonance formulas of a ClO$_3^-$ ion are

\[
\begin{align*}
\text{O} & \equiv \text{Cl} \equiv \text{O} \\
\text{O} & \equiv \text{Cl} \equiv \text{O} \\
\text{O} & \equiv \text{Cl} \equiv \text{O} \\
\text{O} & \equiv \text{Cl} \equiv \text{O} \\
\end{align*}
\]

It is an AX$_3$E ion, and so it is trigonal pyramidal.

8-7. The Lewis formulas of a SO$_2$ molecule are

\[
\text{O} \equiv \text{S} \equiv \text{O}
\]

It is an AX$_2$E molecule, and so is bent with a bond angle of less than 120° because of the two lone pairs.

8-8. The Lewis formula of a XeF$_2$ molecule is \( \text{F} \equiv \text{Xe} \equiv \text{F} \). It is an AX$_2$E$_3$ molecule, and so it is linear.

8-9. The Lewis formulas of a XeO$_3$ molecule are

\[
\begin{align*}
\text{O} & \equiv \text{Xe} \equiv \text{O} \\
\text{O} & \equiv \text{Xe} \equiv \text{O} \\
\text{O} & \equiv \text{Xe} \equiv \text{O} \\
\text{O} & \equiv \text{Xe} \equiv \text{O} \\
\end{align*}
\]

Thus, it is an AX$_3$E molecule, and so is trigonal pyramidal.

8-10. (a) tetrahedral with one chlorine atom; polar

(b) T-shaped; polar

(c) square pyramidal; polar

(d) (symmetric) trigonal bipyramidal; nonpolar

8-11. See the text.
CHAPTER 9. Covalent Bonding

9-1. See the text.

9-2. Bond order \( \frac{9 - 9}{2} = 0 \). See the text.

9-3. Using Figure 9.13, we have

- CN\(^+\): 12 electrons \((\sigma_1)^2(\sigma_{1s}^*)^2(\sigma_{2s})^2(\pi_{2p})^4\) bond order \(\frac{8 - 4}{2} = 2\)
- CN: 13 electrons \((\sigma_1)^2(\sigma_{1s}^*)^2(\sigma_{2s})^2(\pi_{2p})^4(\sigma_{2p})^1\) bond order \(\frac{9 - 4}{2} = 2\frac{1}{2}\)
- CN\(^-\): 14 electrons \((\sigma_1)^2(\sigma_{1s}^*)^2(\sigma_{2s})^2(\pi_{2p})^4(\sigma_{2p})^2\) bond order \(\frac{10 - 4}{2} = 3\)

The species with the largest bond order and thus the largest bond energy is CN\(^-\).

9-4. See the text.

9-5. See the text.

9-6. See the text.

9-7. See the text.

9-8. See the text.

9-9. See the text.
10-1. See the text.
10-2. See the text.
10-3. See the text.
10-4. See the text.
10-5. The oxalate ion has a charge of $-2$ (Table 10.1), and so we predict that it is a diprotic acid.
10-6. Using Tables 10.2 and 10.4 as a guide, we get the answers in the text.
10-7. The unreactive gas is $N_2(g)$ and the metallic mirrorlike deposit is $Na(s)$. Thus, we predict that the decomposition reaction is described by

$$2 NaN_3(s) \rightarrow 2 Na(s) + 3 N_2(g)$$

10-8. See the text.
10-9. Aluminum replaces the hydrogen atoms in $H_2SO_4(aq)$, giving us

$$2 Al(s) + 3 H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3 H_2(g)$$

10-10. See the text.
10-11. $MgO(s)$ combines with $SO_2(g)$ to yield $MgSO_3(s)$ as described in Section 10-1.
10-12. Chlorine is more reactive than bromine and so liberates $Br_2(l)$ from solutions of bromide salts, and so

$$2 Br^- (aq) + Cl_2(g) \rightarrow 2 Cl^- (aq) + Br_2(l)$$

10-13. (a) Soluble by Rule 1
(b) Insoluble by Rule 6
(c) Soluble by Rule 2
(d) Soluble by Rule 1
10-14. (a) $Hg_2Br_2$ is insoluble by Rule 3.

$$Hg_2(NO_3)_2(aq) + 2 NaBr(aq) \rightarrow Hg_2Br_2(s) + 2 NaNO_3(aq)$$

(b) No reaction. All possible products are soluble by Rule 1.
(c) CdS is insoluble by Rule 5.

\[ \text{Na}_2\text{S}(aq) + \text{Cd(NO}_3)_2(aq) \rightarrow 2\text{NaNO}_3(aq) + \text{CdS(s)} \]

10-15. Using the solubility rules, we find that the resulting double-replacement reaction produces a silver chromate precipitate. The complete equation for the reaction is

\[ 2\text{AgClO}_4(aq) + \text{Na}_2\text{CrO}_4(aq) \rightarrow 2\text{NaClO}_4(aq) + \text{Ag}_2\text{CrO}_4(s) \]

The net ionic equation is

\[ 2\text{Ag}^+(aq) + \text{CrO}_4^{2-}(aq) \rightarrow \text{Ag}_2\text{CrO}_4(s) \]

10-16. This is essentially an acid-base reaction, where Li\textsubscript{2}O(s) dissolves in water to yield a base (see Section 10.3), according to

\[ \text{Li}_2\text{O}(s) + \text{H}_2\text{O}(l) \rightarrow 2\text{LiOH}(aq) \]

Thus,

\[ \text{H}_2\text{SO}_4(aq) + 2\text{LiOH}(aq) \rightarrow \text{Li}_2\text{SO}_4(aq) + 2\text{H}_2\text{O}(l) \]

or

\[ \text{H}_2\text{SO}_4(aq) + \text{Li}_2\text{O}(s) \rightarrow \text{Li}_2\text{SO}_4(aq) + \text{H}_2\text{O}(l) \]

The net ionic equation is

\[ 2\text{H}^+(aq) + \text{Li}_2\text{O}(s) \rightarrow 2\text{Li}^+(aq) + \text{H}_2\text{O}(l) \]

10-17. The chemical formula for sodium hydrogen carbonate is NaHCO\textsubscript{3} (Table 10.1) and that of acetic acid is CH\textsubscript{3}COOH (Table 10.5). Thus, the equation for the reaction is

\[ \text{NaHCO}_3(s) + \text{CH}_3\text{COOH}(aq) \rightarrow \text{NaCH}_3\text{COO}(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(l) \]

10-18. (a) The ionic charges on iron in iron metal, Fe(s), and chlorine in chlorine gas, Cl\textsubscript{2}(g), are zero. However, the ionic charge on a chloride ion, Cl\textsuperscript{−}, is −1, and so that on the iron atom in a FeCl\textsubscript{3} formula unit must be +3. Because the ionic charges of the iron and chlorine atoms are changing, this is an oxidation-reduction reaction.

(b) Because the nitrate ion, NO\textsuperscript{−}_3, has an ionic charge of −1, the silver atom in a AgNO\textsubscript{3} formula unit must have an ionic charge of +1. The ionic charges of the Na and S atoms in a Na\textsubscript{2}S formula unit are +1 and −2, respectively. Similarly, we find for the products that the ionic charges of the Ag and S atoms in a Ag\textsubscript{2}S formula unit are +1 and −2, respectively, and the ionic charge of the Na atom in a NaNO\textsubscript{3} formula unit is +1. Thus, we conclude that this is not an oxidation-reduction reaction because the ionic charges of each atom in the reactants are identical to their ionic charges in the products.

(c) The ionic charge on zinc in zinc metal, Zn(s), is zero, and that of the zinc atom in a ZnCl\textsubscript{2} formula unit is +2. Similarly, the ionic charge on the mercury atom in a HgCl\textsubscript{2} formula unit is +2, and that in mercury metal is 0. Thus, we conclude that this is an oxidation-reduction reaction.

10-19. The ionic charge of the aluminum atom changes from 0 in aluminum metal to +3 in aluminum oxide. Similarly, the ionic charge of the manganese atom changes from +3 in manganese oxide to 0 in manganese metal. Thus, in the chemical equation, two aluminum atoms are oxidized (0 → +3) and two manganese atoms are reduced (+3 → 0), and so Al(s) is the reducing agent and Mn\textsubscript{2}O\textsubscript{3}(s) is the oxidizing agent.
CHAPTER 11. Chemical Calculations

11-1. \( \text{H}_2\text{SO}_4(\ell) \)

\[
\text{number of moles} = (37 \, 515 \times 10^3 \text{ metric ton}) \left(\frac{1000 \text{ kg}}{1 \text{ metric ton}}\right) \left(\frac{1000 \text{ g}}{1 \text{ kg}}\right) \left(\frac{1 \text{ mol H}_2\text{SO}_4}{98.078 \text{ g H}_2\text{SO}_4}\right)
\]

\[
= 3.83 \times 10^{11} \text{ mol H}_2\text{SO}_4
\]

\( \text{N}_2(g) \)

\[
\text{number of moles} = (26 \, 675 \times 10^3 \text{ metric ton}) \left(\frac{1000 \text{ kg}}{1 \text{ metric ton}}\right) \left(\frac{1000 \text{ g}}{1 \text{ kg}}\right) \left(\frac{1 \text{ mol N}_2}{28.013 \text{ g N}_2}\right)
\]

\[
= 9.52 \times 10^{11} \text{ mol N}_2
\]

\( \text{C}_2\text{H}_4(g) \)

\[
\text{number of moles} = (25 \, 682 \times 10^3 \text{ metric ton}) \left(\frac{1000 \text{ kg}}{1 \text{ metric ton}}\right) \left(\frac{1000 \text{ g}}{1 \text{ kg}}\right) \left(\frac{1 \text{ mol C}_2\text{H}_4}{28.053 \text{ g C}_2\text{H}_4}\right)
\]

\[
= 9.15 \times 10^{11} \text{ mol C}_2\text{H}_4
\]

\( \text{O}_2(g) \)

\[
\text{number of moles} = (19 \, 539 \times 10^3 \text{ metric ton}) \left(\frac{1000 \text{ kg}}{1 \text{ metric ton}}\right) \left(\frac{1000 \text{ g}}{1 \text{ kg}}\right) \left(\frac{1 \text{ mol O}_2}{31.9988 \text{ g O}_2}\right)
\]

\[
= 6.11 \times 10^{11} \text{ mol O}_2
\]

\( \text{H}_2(g) \)

\[
\text{number of moles} = (17 \, 698 \times 10^3 \text{ metric ton}) \left(\frac{1000 \text{ kg}}{1 \text{ metric ton}}\right) \left(\frac{1000 \text{ g}}{1 \text{ kg}}\right) \left(\frac{1 \text{ mol H}_2}{2.0158 \text{ g H}_2}\right)
\]

\[
= 8.78 \times 10^{12} \text{ mol H}_2
\]

On a mole basis, \( \text{H}_2(g) \) is produced in the greatest quantity annually in the United States.
11-2. The mass of a CO₂ molecule is

\[
mass \text{ of one molecule} = \left( \frac{44.0095 \text{ g CO}_2}{1 \text{ mol CO}_2} \right) \left( \frac{1 \text{ kg}}{1000 \text{ g}} \right) \left( \frac{1 \text{ mol}}{6.022 \times 10^{23}} \right) = 7.308 \times 10^{-26} \text{ kg}
\]

The mass of a SF₆ molecule is

\[
mass \text{ of one molecule} = \left( \frac{146.055 \text{ g SF}_6}{1 \text{ mol SF}_6} \right) \left( \frac{1 \text{ kg}}{1000 \text{ g}} \right) \left( \frac{1 \text{ mol}}{6.022 \times 10^{23}} \right) = 2.425 \times 10^{-25} \text{ kg}
\]

11-3. The number of each is given by

\[
number = (1.00 \times 10^{-12} \text{ L}) \left( \frac{1000 \text{ mL}}{1 \text{ L}} \right) \left( \frac{1.00 \text{ g}}{1 \text{ mL}} \right) \left( \frac{1 \text{ mol H}_2\text{O}}{18.015 \text{ g H}_2\text{O}} \right) \\
\times \left( \frac{6.022 \times 10^{23} \text{ molecule}}{1 \text{ mol}} \right)
= 3.34 \times 10^{13} \text{ H}_2\text{O molecules} \\
= 3.34 \times 10^{13} \text{ O atoms} \\
= 6.68 \times 10^{13} \text{ H atoms}
\]

11-4. As usual, take a 100-gram sample and write

\[
52.1 \text{ g C} \cong 13.2 \text{ g H} \cong 34.7 \text{ g O}
\]

Divide each value by the corresponding atomic mass

\[
\frac{52.1 \text{ C}}{12.01} \cong \frac{13.2 \text{ H}}{1.0079} \cong \frac{34.7 \text{ O}}{16.00}
\]

to get

\[
4.34 \text{ mol C} \cong 13.10 \text{ mol H} \cong 2.17 \text{ mol O}
\]

Dividing by 2.17, the smallest number, we get

\[
2.00 \text{ mol C} \cong 6.04 \text{ mol H} \cong 1.00 \text{ mol O}
\]

or C₂H₆O.

11-5. The amount of oxygen is given by

\[
mass \text{ of oxygen} = 3.37 \text{ g Sc O} - 2.18 \text{ g Sc} = 1.16 \text{ g}
\]

Thus, we have

\[
2.18 \text{ g Sc} \cong 1.16 \text{ g O}
\]

Dividing each value by the corresponding atomic mass gives

\[
0.0485 \text{ mol Sc} \cong 0.0725 \text{ mol O}
\]
Dividing by 0.0485, we get

\[ 1 \text{ mol Sc} \cong 1.49 \text{ mol O} \]

or, multiplying by 2, \( \text{Sc}_2\text{O}_3 \).

11-6. The amount of oxygen is given by \((4.111 - 3.058)\) grams = 1.053 grams. Thus, we have

\[ 3.058 \text{ g M} \cong 1.053 \text{ g O} \]

The number of moles of oxygen is

\[ \text{moles of O} = \left( \frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) (1.053 \text{ g O}) = 0.06581 \text{ mol} \]

The empirical formula of the oxide is \( \text{M}_2\text{O}_3 \), so we have

\[ \text{moles of M} = (0.06581 \text{ mol O}) \left( \frac{2 \text{ mol M}}{3 \text{ mol O}} \right) = 0.04388 \text{ mol} \]

and so

\[ 3.058 \text{ g M} \cong 0.04388 \text{ mol M} \]

or

\[ 69.69 \text{ g M} \cong 1 \text{ mol M} \]

The metal is gallium.

11-7. As usual, take a 100-gram sample and write

\[ 24.47 \text{ g C} \cong 4.075 \text{ g H} \cong 71.65 \text{ g Cl} \]

\[ 2.037 \text{ mol C} \cong 4.043 \text{ mol H} \cong 2.021 \text{ mol Cl} \]

\[ 1 \text{ mol C} \cong 2 \text{ mol H} \cong 1 \text{ mol Cl} \]

or \( \text{CH}_2\text{Cl} \) is the empirical formula. Given that its molecular mass is 98.95 and the empirical formula mass is 49.48, the molecular formula must be \( \frac{98.95}{49.48} = 2 \) times the empirical formula. Thus, the molecular formula is \( \text{C}_2\text{H}_4\text{Cl}_2 \).

11-8. We have that

\[ \text{mass of C} = \left( 3.710 \text{ g CO}_2 \right) \left( \frac{12.01 \text{ g C}}{44.01 \text{ g CO}_2} \right) = 1.012 \text{ g C} \]

\[ \text{mass of H} = \left( 1.013 \text{ g H}_2\text{O} \right) \left( \frac{2 \times 1.008 \text{ g H}}{18.02 \text{ g H}_2\text{O}} \right) = 0.1133 \text{ g H} \]

\[ \text{mass \% C} = \frac{1.012 \text{ g C}}{2.475 \text{ g sample}} \times 100 = 40.89\% \]
mass % H = \frac{0.1133 \text{ g H}}{2.475 \text{ g sample}} \times 100 = 4.578% \\

mass % O = (100.00 - 40.89 - 4.578)\% = 54.53% \\

As usual, take a 100-gram sample and write

40.89 g C \cong 4.578 g H \cong 54.53 g O \\
3.405 \text{ mol C} \cong 4.542 \text{ mol H} \cong 3.408 \text{ mol O} \\
1.00 \text{ mol C} \cong 1.33 \text{ mol H} = \frac{4}{3} \text{ mol H} \cong 1.00 \text{ mol O} \\
3 \text{ mol C} \cong 4 \text{ mol H} \cong 3 \text{ mol O} \\

The empirical formula is C₃H₄O₃. Given that the empirical formula is 88.06 and that its molecular mass is 176, the molecular formula of vitamin C is C₆H₈O₆.

11-9. (a) 2 \text{ KClO}_3(s) \rightarrow 2 \text{ KCl}(s) + 3 \text{ O}_2(g) \\
(b) \text{ moles of O}_2 = (0.50 \text{ mol KClO}_3) \left( \frac{3 \text{ mol O}_2}{2 \text{ mol KClO}_3} \right) = 0.75 \text{ mol} \\
(c) \text{ mass of O}_2 = (30.6 \text{ g KClO}_3) \left( \frac{1 \text{ mol KClO}_3}{122.55 \text{ g KClO}_3} \right) \left( \frac{3 \text{ mol O}_2}{2 \text{ mol KClO}_3} \right) \left( \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) = 12.0 \text{ g} \\

11-10. We have that

\text{ moles of ND}_3 = (7.15 \text{ mg ND}_3) \left( \frac{1 \text{ g}}{1000 \text{ mg}} \right) \left( \frac{1 \text{ mol ND}_3}{20.049 \text{ g ND}_3} \right) \\
= 3.566 \times 10^{-4} \text{ mol} \\

\text{ mass of D}_2\text{O} = (3.566 \times 10^{-4} \text{ mol D}_2\text{O}) \left( \frac{3 \text{ mol D}_2\text{O}}{1 \text{ mol ND}_3} \right) \left( \frac{20.027 \text{ g D}_2\text{O}}{1 \text{ mol D}_2\text{O}} \right) \left( \frac{1000 \text{ mg}}{1 \text{ g}} \right) \\
= 21.4 \text{ mg D}_2\text{O} \\

The number of milliliters of D₂O(l) required is given by

\text{ volume} = \frac{21.4 \text{ mg}}{1 \text{ mL}} \left( \frac{1 \times 10^{-3} \text{ g}}{1 \text{ mg}} \right) \left( \frac{1 \text{ mL}}{1.106 \text{ g}} \right) = 0.0194 \text{ mL} \\

11-11. We have that

\text{ mass of SF}_6 = (5.00 \text{ g S}) \left( \frac{1 \text{ mol S}}{32.065 \text{ g S}} \right) \left( \frac{1 \text{ mol SF}_6}{1 \text{ mol S}} \right) \left( \frac{146.05 \text{ g SF}_6}{1 \text{ mol SF}_6} \right) = 22.8 \text{ g} \\

and

\text{ mass of F}_2 = (5.00 \text{ g S}) \left( \frac{1 \text{ mol S}}{32.065 \text{ g S}} \right) \left( \frac{3 \text{ mol F}_2}{1 \text{ mol S}} \right) \left( \frac{37.997 \text{ g F}_2}{1 \text{ mol F}_2} \right) = 17.8 \text{ g} \\

11-12. (a) The number of grams of PI₃(s) prepared is given by

\text{ mass of PI}_3 = (1.25 \text{ g P}) \left( \frac{1 \text{ mol P}}{30.974 \text{ g P}} \right) \left( \frac{2 \text{ mol PI}_3}{2 \text{ mol P}} \right) \left( \frac{411.69 \text{ g PI}_3}{1 \text{ mol PI}_3} \right) = 16.6 \text{ g}
(b) The number of grams of $I_2(s)$ required is given by

$$\text{mass of } I_2 = (1.25 \text{ g } P) \left( \frac{1 \text{ mol } P}{30.974 \text{ g } P} \right) \left( \frac{3 \text{ mol } I_2}{2 \text{ mol } P} \right) \left( \frac{253.81 \text{ g } I_2}{1 \text{ mol } I_2} \right) = 15.4 \text{ g}$$

11-13. All the copper in $\text{CuFeS}_2(s)$ ends up as copper metal. Thus,

$$\text{mass of } Cu = (1000 \text{ kg } \text{ore}) \left( \frac{0.30 \text{ CuFeS}_2}{100\% \text{ ore}} \right) \left( \frac{1000 \text{ g}}{1 \text{ kg}} \right) \left( \frac{1 \text{ mol } CuFeS_2}{183.52 \text{ g } CuFeS_2} \right) \times \left( \frac{1 \text{ mol } Cu}{1 \text{ mol } CuFeS_2} \right) \left( \frac{63.546 \text{ g } Cu}{1 \text{ mol } Cu} \right) = 1.04 \times 10^5 \text{ g} = 0.104 \text{ metric ton}$$

11-14. Let $x$ be the mass percentage of $\text{NaCl}(s)$ in the mixture. The mass of $\text{NaCl}(s)$ in the mixture is given by

$$\text{mass of } \text{NaCl} = (2.86 \text{ g}) \left( \frac{x}{100} \right)$$

and the mass of $\text{BaCl}_2$ is given by

$$\text{mass of } \text{BaCl}_2 = (2.86 \text{ g}) \left( 1 - \frac{x}{100} \right)$$

The number of moles of each compound is

$$\text{moles } \text{NaCl} = (2.86 \text{ g } \text{NaCl}) \left( \frac{x}{100} \right) \left( \frac{1 \text{ mol } \text{NaCl}}{58.44 \text{ g } \text{NaCl}} \right) = 4.89 \times 10^{-4}x \text{ mol}$$

$$\text{moles } \text{BaCl}_2 = \left[ (2.86 \text{ g}) \left( 1 - \frac{x}{100} \right) \right] \left( \frac{1 \text{ mol } \text{BaCl}_2}{208.23 \text{ g } \text{BaCl}_2} \right) = (0.0137 - 1.37 \times 10^{-4}x) \text{ mol}$$

The number of moles of $\text{AgCl}(s)$ is given by

$$\text{moles } \text{AgCl} = (4.81 \text{ g } \text{AgCl}) \left( \frac{1 \text{ mol } \text{AgCl}}{143.32 \text{ g } \text{AgCl}} \right) = 0.0336 \text{ mol } \text{AgCl}$$

Because all the chloride in the $\text{AgCl}(s)$ comes from the $\text{NaCl}(s)$ and $\text{BaCl}_2(s)$, from the reaction stoichiometry, we have

$$\text{moles } \text{AgCl} = \text{moles } \text{NaCl} + 2 \text{moles } \text{BaCl}_2$$

Therefore, we have

$$0.0336 \text{ mol} = (4.89 \times 10^{-4}x) \text{ mol} + (2)(0.0137 - 1.37 \times 10^{-4}x) \text{ mol}$$

$$6.2 \times 10^{-3} = 2.15 \times 10^{-4}x$$

or $x = 29\%$. Therefore, the mixture contains 29% $\text{NaCl}(s)$ and 71% $\text{BaCl}_2(s)$.

11-15. The balanced chemical equation is

$$\text{CaSO}_4(s) + 4 \text{C}(s) \rightarrow \text{CaS}(s) + 4 \text{CO}(g)$$
We must check to see if there is a limiting reactant.

moles of CaSO\(_4\) = \((125 \text{ g CaSO}_4) \left( \frac{1 \text{ mol CaSO}_4}{136.14 \text{ g CaSO}_4} \right)\) = 0.918 mol

moles of C = \((125 \text{ g C}) \left( \frac{1 \text{ mol C}}{12.0107 \text{ g C}} \right)\) = 10.4 mol

Dividing each by its respective stoichiometric coefficient in the balanced equation, we find that the carbon is in great excess. The number of grams of CaS(s) produced is given by

mass of CaS = \((0.918 \text{ mol CaSO}_4) \left( \frac{1 \text{ mol CaS}}{1 \text{ mol CaSO}_4} \right) \left( \frac{72.14 \text{ g CaS}}{1 \text{ mol CaS}} \right)\) = 66.2 g

11-16. For a theoretical yield, the number of grams of tin required is given by

mass of Sn = \((0.106 \text{ g SnCl}_4) \left( \frac{1 \text{ mol SnCl}_4}{260.52 \text{ g SnCl}_4} \right) \left( \frac{1 \text{ mol Sn}}{1 \text{ mol SnCl}_4} \right) \left( \frac{118.71 \text{ g Sn}}{1 \text{ mol Sn}} \right)\) = 0.0483 g

For a percentage yield of 64.3%, the mass of tin required is

mass of Sn = \((0.0483 \text{ g Sn}) \left( \frac{100\%}{64.3\%} \right)\) = 0.0751 g