CHAPTER 3
CHEMICAL COMPOUNDS
PRACTICE EXAMPLES

1A First we convert the number of chloride ions to the mass of MgCl₂.

\[
\text{mass}_{\text{MgCl}_2} = 5.0 \times 10^{23} \text{Cl}^- \times \frac{1 \text{MgCl}_2}{2 \text{Cl}^-} \times \frac{1 \text{mol MgCl}_2}{6.022 \times 10^{23} \text{MgCl}_2} \times \frac{95.211 \text{g MgCl}_2}{1 \text{mol MgCl}_2} = 4.0 \times 10^4 \text{ g MgCl}_2
\]

1B First we convert mass Mg(NO₃)₂ to moles Mg(NO₃)₂ and formula units Mg(NO₃)₂ then finally to NO₃⁻ ions.

\[
\begin{align*}
1.00 \mu \text{g Mg(NO}_3\text{)}_2 & \times \frac{1 \text{ g Mg(NO}_3\text{)}_2}{1,000,000 \mu \text{g Mg(NO}_3\text{)}_2} \times \frac{1 \text{ mol Mg(NO}_3\text{)}_2}{148.3148 \text{ g Mg(NO}_3\text{)}_2} \times \frac{6.022 \times 10^{22} \text{ formula units Mg(NO}_3\text{)}_2}{1 \text{ mol Mg(NO}_3\text{)}_2} \\
& = 4.06 \times 10^{15} \text{ formula units Mg(NO}_3\text{)}_2
\end{align*}
\]

Next, determine the number of oxygen atoms by multiplying by the appropriate ratio.

\[
\# \text{ atoms O} = 4.06 \times 10^{15} \text{ formula units Mg(NO}_3\text{)}_2 \times \frac{6 \text{ atoms O}}{1 \text{ formula unit Mg(NO}_3\text{)}_2} = 2.44 \times 10^{16} \text{ O}
\]

2A The volume of gold is converted to its mass and then to the amount in moles.

\[
\# \text{ Au atoms} = (2.50 \text{ cm})^2 \times \left( \frac{0.100 \text{ mm}}{10 \text{ mm}} \right) \times \frac{1 \text{ cm}}{1 \text{ cm}^3} \times \frac{1 \text{ mol Au}}{196.97 \text{ g Au}} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol Au}}
\]

\[
= 3.69 \times 10^{21} \text{ Au atoms}
\]

2B We need the molar mass of ethyl mercaptan for one conversion factor.

\[
M = (2 \times 12.011 \text{ g C}) + (6 \times 1.008 \text{ g H}) + (1 \times 32.066 \text{ g S}) = 62.136 \text{ g/mol C}_2\text{H}_6\text{S}
\]

Volume of room: 62 ft × 35 ft × 14 ft = 3.04 × 10⁴ ft³. We also need to convert ft³ to m³.

\[
3.04 \times 10^4 \text{ ft}^3 \times \left( \frac{12 \text{ in}}{1 \text{ ft}} \right)^3 \times \left( \frac{2.54 \text{ cm}}{1 \text{ in}} \right)^3 \times \left( \frac{1 \text{ m}}{100 \text{ cm}} \right)^3 = 8.6 \times 10^3 \text{ m}^3
\]

\[
[C_2\text{H}_6\text{S}] = \frac{1.0 \mu \text L \text{C}_2\text{H}_6\text{S}}{8.6 \times 10^2 \text{ m}^3} \times \frac{1 \text{L}}{1 \times 10^6 \mu \text{L}} \times \frac{1000 \text{ mL}}{1 \text{L}} \times \frac{0.84 \text{ g}}{1 \text{mol C}_2\text{H}_6\text{S}} \times \frac{1 \text{ mol C}_2\text{H}_6\text{S}}{10^6 \mu \text{mol}} \times \frac{62.136 \text{ g}}{1 \text{ mol}}
\]

\[
= 0.016 \mu \text{mol/m}^3 > 9.0 \times 10^{-4} \mu \text{mol/m}^3 = \text{ the detectable limit}
\]

Thus, the vapor will be detectable.
Chapter 3: Chemical Compounds

3A  The molar mass of halothane is given in Example 3-3 in the textbook as 197.4 g/mol. The rest of the solution uses conversion factors to change units.

\[
\text{mass Br} = 25.0 \text{ mL } C_2HBrClF_3 \times \frac{1.871 \text{ g } C_2HBrClF_3}{1 \text{ mL } C_2HBrClF_3} \times \frac{1 \text{ mol } C_2HBrClF_3}{197.4 \text{ g } C_2HBrClF_3} \times \frac{1 \text{ mol Br}}{1 \text{ mol } C_2HBrClF_3}
\]

\[
\times \frac{79.904 \text{ g Br}}{1 \text{ mol Br}} = 18.9 \text{ g Br}
\]

3B  Again, the molar mass of halothane is given in Example 3-3 in the textbook as 197.4 g/mol.

\[
V_{\text{halothane}} = 1.00 \times 10^{24} \text{ Br} \times \frac{1 \text{ mol Br}}{6.022 \times 10^{23} \text{ Br}} \times \frac{1 \text{ mol } C_2HBrClF_3}{1 \text{ mol Br}} \times \frac{197.4 \text{ g } C_2HBrClF_3}{1 \text{ mol } C_2HBrClF_3} \times \frac{1 \text{ mL}}{1.871 \text{ g}}
\]

\[
= 175 \text{ mL } C_2HBrClF_3
\]

4A  We use the same technique as before: determine the mass of each element in a mole of the compound. Their sum is the molar mass of the compound. The percent composition is determined by comparing the mass of each element with the molar mass of the compound.

\[
M = (10 \times 12.011 \text{ g C}) + (16 \times 1.008 \text{ g H}) + (5 \times 14.01 \text{ g N}) + (3 \times 30.97 \text{ g P}) + (13 \times 15.999 \text{ g O})
\]

\[
= 120.11 \text{ g C} + 16.13 \text{ g H} + 70.05 \text{ g N} + 92.91 \text{ g P} + 207.99 \text{ g O} = 507.19 \text{ g ATP/mol}
\]

\[
\%C = \frac{120.11 \text{ g C}}{507.19 \text{ g ATP}} \times 100\% = 23.68\% C
\]

\[
\%H = \frac{16.13 \text{ g H}}{507.19 \text{ g ATP}} \times 100\% = 3.18\% H
\]

\[
\%N = \frac{70.05 \text{ g N}}{507.19 \text{ g ATP}} \times 100\% = 13.81\% N
\]

\[
\%P = \frac{92.91 \text{ g P}}{507.19 \text{ g ATP}} \times 100\% = 18.32\% P
\]

\[
\%O = \frac{207.99 \text{ g O}}{507.19 \text{ g ATP}} \times 100\% = 41.008\% O
\]

(Note: the mass percents sum to 99.999%)

4B  Both (b) and (e) have the same empirical formula, that is, CH₂O. These two molecules have the same percent oxygen by mass.

5A  Once again, we begin with a 100.00 g sample of the compound. In this way, each elemental mass in grams is numerically equal to its percent. We convert each mass to an amount in moles, and then determine the simplest integer set of molar amounts. This determination begins by dividing all three molar amounts by the smallest.

\[
39.56 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 3.294 \text{ mol C} \quad \rightarrow \quad 1.000 \text{ mol C} \times 3.000 = 3.000 \text{ mol C}
\]

\[
7.74 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 7.68 \text{ mol H} \quad \rightarrow \quad 2.33 \text{ mol H} \times 3.000 = 6.99 \text{ mol H}
\]

\[
52.70 \text{ g O} \times \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 3.294 \text{ mol O} \quad \rightarrow \quad 1.000 \text{ mol O} \times 3.000 = 3.000 \text{ mol O}
\]
Thus, the empirical formula of the compound is \( \text{C}_3\text{H}_7\text{O}_3 \). The empirical molar mass of this compound is:
\[
(3\times12.01\text{ g C})+(7\times1.008\text{ g H})+(3\times16.00\text{ g O}) = 36.03\text{ g} + 7.056\text{ g} + 48.00\text{ g} = 91.09\text{ g/mol}
\]
The empirical mass is almost precisely one half the reported molar mass, leading to the conclusion that the molecular formula must be twice the empirical formula in order to double the molar mass. Thus, the molecular formula is \( \text{C}_6\text{H}_{14}\text{O}_6 \).

5B To answer this question, we start with a 100.00 g sample of the compound. In this way, each elemental mass in grams is numerically equal to its percent. We convert each mass to an amount in moles, and then determine the simplest integer set of molar amounts. This determination begins by dividing all molar amounts by the smallest number of moles in the group of four, viz. 1.1025 moles. Multiplication of the resulting quotients by eight produces the smallest possible set of whole numbers.

\[
\begin{align*}
21.51\text{ g C} \times \frac{1\text{ mol C}}{12.011\text{ g C}} &= 1.791\text{ mol C} + 1.1025 \rightarrow 1.624\text{ mol C} \times 8 = 12.99\text{ mol C} \\
2.22\text{ g H} \times \frac{1\text{ mol H}}{1.00794\text{ g H}} &= 2.20\text{ mol H} + 1.1025 \rightarrow 2.00\text{ mol H} \times 8 = 16.0\text{ mol H} \\
17.64\text{ g O} \times \frac{1\text{ mol O}}{15.9994\text{ g O}} &= 1.1025\text{ mol O} + 1.1025 \rightarrow 1.000\text{ mol O} \times 8 = 8.000\text{ mol O} \\
58.63\text{ g Cl} \times \frac{1\text{ mol Cl}}{35.453\text{ g Cl}} &= 1.654\text{ mol Cl} + 1.1025 \rightarrow 1.500\text{ mol Cl} \times 8 = 12.00\text{ mol Cl}
\end{align*}
\]
Thus, the empirical formula of the compound is \( \text{C}_{13}\text{H}_{16}\text{O}_8\text{Cl}_{12} \). The empirical molar mass of this compound is 725.7 g/mol.
The empirical mass is almost precisely the same as the reported molar mass, leading to the conclusion that the molecular formula must be the same as the empirical formula. Thus, the molecular formula is \( \text{C}_{13}\text{H}_{16}\text{O}_8\text{Cl}_{12} \).

6A We calculate the amount in moles of each element in the sample (determining the mass of oxygen by difference) and transform these molar amounts to the simplest integral amounts, by first dividing all three by the smallest.

\[
\begin{align*}
2.726\text{ g CO}_2 \times \frac{1\text{ mol CO}_2}{44.010\text{ g CO}_2} \times \frac{1\text{ mol C}}{1\text{ mol CO}_2} &= 0.06194\text{ mol C} \times \frac{12.011\text{ g C}}{1\text{ mol C}} = 0.7440\text{ g C} \\
1.116\text{ g H}_2\text{O} \times \frac{1\text{ mol H}_2\text{O}}{18.015\text{ g H}_2\text{O}} \times \frac{2\text{ mol H}}{1\text{ mol H}_2\text{O}} &= 0.1239\text{ mol H} \times \frac{1.008\text{ g H}}{1\text{ mol H}} = 0.1249\text{ g H} \\
(1.152\text{ g cmpd} - 0.7440\text{ g C} - 0.1249\text{ g H}) &= 0.283\text{ g O} \times \frac{1\text{ mol O}}{16.00\text{ g O}} = 0.0177\text{ mol O}
\end{align*}
\]
Chapter 3: Chemical Compounds

<table>
<thead>
<tr>
<th>Element</th>
<th>Mole</th>
<th>Mass (g)</th>
<th>Molar Mass (g/mol)</th>
<th>Molar Amount (mol)</th>
<th>O.S.</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>0.06194</td>
<td>0.0177</td>
<td>12.011</td>
<td>0.01789 mol C</td>
<td>3.999 mol C</td>
</tr>
<tr>
<td>H</td>
<td>0.1239</td>
<td>0.0177</td>
<td>1.008</td>
<td>0.01789 mol H</td>
<td>4.00 mol H</td>
</tr>
<tr>
<td>O</td>
<td>0.0177</td>
<td>0.0177</td>
<td>15.999</td>
<td>0.01789 mol O</td>
<td>1.000 mol O</td>
</tr>
</tbody>
</table>

All of these amounts in moles are multiplied by 2 to make them integral. Thus, the empirical formula of isobutyl propionate is $C_7H_{14}O_2$.

Notice that we do not have to obtain the mass of any element in this compound by difference; there is no oxygen present in the compound. We calculate the amount in one mole of each element in the sample and transform these molar amounts to the simplest integral amounts, by first dividing all three by the smallest.

Thus, the empirical formula of thiophene is $C_4H_4S$.

For an atom of a free element, the oxidation state is 0 (rule 1).

The sum of all the oxidation numbers in the ion is −2 (rule 2). The O.S. of each oxygen is −2 (rule 6). Thus, the total for all seven oxygens is −14. The total for bothchromiums must be +12. Thus, each Cr has an O.S. = +6.

The sum of all oxidation numbers in the compound is 0 (rule 2). The O.S. of oxygen is −2 (rule 6). The total for the two chlorines must be +2. Thus, each chlorine must have O.S. = +1.

The sum for all the oxidation numbers in the compound is 0 (rule 2). The O.S. of potassium is +1 (rule 3). The sum of the oxidation numbers of the two oxygens must be −1. Thus, each oxygen must have O.S. = −1/2.

The sum of all the oxidation numbers in the ion is −2 (rule 2). The O.S. of oxygen is −2 (rule 6). Thus, the total for three oxygens must be −6. The total for both sulfurs must be +4. Thus, each S has an O.S. = +2.

The O.S. of each Cl is +1 (rule 7). The sum of all O.S. is 0 (rule 2). Thus, the total for two Hg is +2 and each Hg has O.S. = +1.

The O.S. of each O is −2 (rule 6). Thus, the total for 4 oxygens must be −8. The K has O.S. = +1 (rule 3). The total of all O.S. is 0 (rule 2). Thus, the O.S. of Mn is +7.

The O.S. of each H is +1 (rule 5), producing a total for both hydrogens of +2. The O.S. of O is −2 (rule 6). Thus, the O.S. of C is 0, because the total of all O.Ss. is 0 (rule 2).
In each case, we determine the formula with its accompanying charge of each ion in the compound. We then produce a formula for the compound in which the total positive charge equals the total negative charge.

- lithium oxide: $\text{Li}^+ \text{ and } \text{O}^{2-}$
  - two $\text{Li}^+$ and one $\text{O}^{2-}$
  - $\text{Li}_2\text{O}$

- tin(II) fluoride: $\text{Sn}^{2+} \text{ and } \text{F}^-$
  - one $\text{Sn}^{2+}$ and two $\text{F}^-$
  - $\text{SnF}_2$

- lithium nitride: $\text{Li}^+ \text{ and } \text{N}^{3-}$
  - three $\text{Li}^+$ and one $\text{N}^{3-}$
  - $\text{Li}_3\text{N}$

Using a similar procedure as that provided in 8A:

- aluminum sulfide: $\text{Al}^{3+} \text{ and } \text{S}^{2-}$
  - two $\text{Al}^{3+}$ and three $\text{S}^{2-}$
  - $\text{Al}_2\text{S}_3$

- magnesium nitride: $\text{Mg}^{2+} \text{ and } \text{N}^{3-}$
  - three $\text{Mg}^{2+}$ and two $\text{N}^{3-}$
  - $\text{Mg}_3\text{N}_2$

- vanadium(III) oxide: $\text{V}^{3+} \text{ and } \text{O}^{2-}$
  - two $\text{V}^{3+}$ and three $\text{O}^{2-}$
  - $\text{V}_2\text{O}_3$

The name of each of these ionic compounds is the name of the cation followed by that of the anion. Each anion name is a modified (with the ending “ide”) version of the name of the element. Each cation name is the name of the metal, with the oxidation state appended in Roman numerals in parentheses if there is more than one type of cation for that metal.

- CsI cesium iodide
- CaF$_2$ calcium fluoride
- FeO The O.S. of O = $-2$ (rule 6). Thus, the O.S. of Fe = $+2$ (rule 2).
  - The name of the compound is iron(II) oxide.
- CrCl$_3$ The O.S. of Cl = $-1$ (rule 7). Thus, the O.S. of Cr = $+3$ (rule 2).
  - The cation is chromium (III). The compound is chromium (III) chloride.

The name of each of these ionic compounds is the name of the cation followed by that of the anion. Each anion name is a modified (with the ending “ide”) version of the name of the element. Each cation name is the name of the metal, with the oxidation state appended in Roman numerals in parentheses if there is more than one type of cation for that metal.

The oxidation state of Ca is $+2$ (rule 3). Hydrogen would therefore have an oxidation number of $-1$ (which is an exception to rule 5), based on rule 2.

- CaH$_2$ calcium hydride

The oxidation number of sulfur is $-2$ (rule 7), and therefore silver would be $+1$ for each silver atom based on rule 2.

- Ag$_2$S silver(I) sulfide

In the next two compounds, the oxidation state of chlorine is $-1$ (rule 7) and thus the oxidation state of the metal in each cation must be $+1$ (rule 2).

- CuCl copper(I) chloride
- Hg$_2$Cl$_2$ mercury(I) chloride
Chapter 3: Chemical Compounds

10A

SF₆  Both S and F are nonmetals. This is a binary molecular compound: sulfur hexafluoride.

HNO₂  The NO₂⁻ ion is the nitrite ion. Its acid is nitrous acid.

Ca(HCO₃)₂  HCO₃⁻ is the bicarbonate ion or the hydrogen carbonate ion.
This compound is calcium bicarbonate or calcium hydrogen carbonate.

FeSO₄  The SO₄²⁻ ion is the sulfate ion. The cation is Fe²⁺, iron(II).
This compound is iron(II) sulfate.

10B

NH₄NO₃  The cation is NH₄⁺, ammonium ion. The anion is NO₃⁻, nitrate ion.
This compound is ammonium nitrate.

PCl₃  Both P and Cl are nonmetals. This is a binary molecular compound: phosphorus trichloride.

HBrO  BrO⁻ is hypobromite, this is hypobromous acid.

AgClO₄  The anion is perchlorate ion, ClO₄⁻. The compound is silver(I) perchlorate.

Fe₂(SO₄)₃  The SO₄²⁻ ion is the sulfate ion. The cation is Fe³⁺, iron(III).
This compound is iron(III) sulfate.

11A

boron trifluoride  Both elements are nonmetals.
This is a binary molecular compound: BF₃.

potassium dichromate  Potassium ion is K⁺, and dichromate ion is Cr₂O₇²⁻.
This is K₂Cr₂O₇.

sulfuric acid  The anion is sulfate, SO₄²⁻. There must be two H⁺s.
This is H₂SO₄.

calcium chloride  The ions are Ca²⁺ and Cl⁻. There must be one Ca²⁺ and two Cl⁻s: CaCl₂.

11B

aluminum nitrate  Aluminum is Al³⁺; the nitrate ion is NO₃⁻. This is Al(NO₃)₃.

tetraphosphorus decoxide  Both elements are nonmetals.
This is a binary molecular compound, P₄O₁₀.

chromium(III) hydroxide  Chromium(III) ion is Cr³⁺; the hydroxide ion is OH⁻.
This is Cr(OH)₃.

iodic acid  The halogen “ic” acid has the halogen in a +5 oxidation state.
This is HIO₅.
12A
(a) Not isomers: molecular formulas are different (C₈H₁₈ vs C₉H₂₀).
(b) Molecules are isomers (same formula C₇H₁₆).

12B
(a) Molecules are isomers (same formula C₇H₁₄)
(b) Not isomers: molecular formulas are different (C₄H₈ vs C₅H₁₀).

13A
(a) The carbon to carbon bonds are all single bonds in this hydrocarbon. This compound is an alkane.
(b) In this compound, there are only single bonds, and a Cl atom has replaced one H atom. This compound is a chloroalkane.
(c) The presence of the carboxyl group (—CO₂H) in this molecule means that the compound is a carboxylic acid.
(d) There is a carbon to carbon double bond in this hydrocarbon. This is an alkene.

13B
(a) The presence of the hydroxyl group (—OH) in this molecule means that this compound is an alcohol.
(b) The presence of the carboxyl group (—CO₂H) in this molecule means that the compound is a carboxylic acid. This molecule also contains the hydroxyl group (—OH).
(c) The presence of the carboxyl group (—CO₂H) in this molecule means that the compound is a carboxylic acid. As well, a Cl atom has replaced one H atom. This compound is a chloroalkane. The compound is a chloro carboxylic acid.
(d) There is a carbon to carbon double bond in this compound; hence, it is an alkene. There is also one H atom that has been replaced by a Br atom. This compound is also a bromoalkane.

14A
(a) The structure is that of an alcohol with the hydroxyl group on the second carbon atom of a three carbon chain. The compound is 2-propanol (commonly isopropyl alcohol).
(b) The structure is that of an iodoalkane molecule with the I atom on the first carbon of a three-carbon chain. The compound is called 1-iodopropane.
(c) The carbon chain in this structure is four carbon atoms long with the end C atom in a carboxyl group. There is also a methyl group on the third carbon in the chain. The compound is 3-methylbutanoic acid.
(d) The structure is that of a three carbon chain that contains a carbon to carbon double bond. This compound is propene.

14B
(a) 2-chloropropane  (b) 1,4-dichlorobutane  (c) 2-methyl propanoic acid
15A
(a) pentane: \( \text{CH}_3(\text{CH}_2)_3\text{CH}_3 \)  
(b) ethanoic acid: \( \text{CH}_3\text{CO}_2\text{H} \)  
(c) 1-iodooctane: \( \text{ICH}_2(\text{CH}_2)_6\text{CH}_3 \)  
(d) 1-pentanol: \( \text{CH}_2(\text{OH})(\text{CH}_2)_3\text{CH}_3 \)

15B
(a) propene

(b) 1-heptanol

(c) chloroacetic acid

(d) hexanoic acid

INTEGRATIVE EXAMPLES

A.
First, determine the mole ratios of the dehydrated compound:
27.74 g Mg \( \times (1 \text{ mol Mg} / 24.305 \text{ g Mg}) = 1.14 \text{ mol Mg} \)
23.57 g P \( \times (1 \text{ mol P} / 30.97 \text{ g P}) = 0.76 \text{ mol P} \)
48.69 g O \( \times (1 \text{ mol O} / 16.0 \text{ g O}) = 3.04 \text{ mol O} \)

Mole ratios are determined by dividing by the smallest number:
1.14 mol Mg / 0.76 mol P = 1.5
0.76 mol P / 0.76 mol P = 1.0
3.04 mol O / 0.76 mol P = 4.0

Multiplying by 2 to get whole numbers, the empirical formula becomes \( \text{Mg}_3\text{P}_2\text{O}_8 \). The compound is magnesium phosphate, \( \text{Mg}_3(\text{PO}_4)_2 \).

To determine the number of waters of hydration, determine the mass of water driven off.
mass of \( \text{H}_2\text{O} \) = 2.4917 g – 1.8558 g = 0.6359 g
mol \( \text{H}_2\text{O} \) = 0.6359 g \( \times (1 \text{ mol H}_2\text{O} / 18.0 \text{ g H}_2\text{O}) = 0.0353 \text{ mol.} \)

Then, calculate the number of moles of dehydrated \( \text{Mg}_3(\text{PO}_4)_2 \) in the same manner above. The number of moles (using 262.86 g/mol for molecular weight) is 0.00706. Dividing the number of moles of \( \text{H}_2\text{O} \) by \( \text{Mg}_3(\text{PO}_4)_2 \) gives a ratio of 5. Therefore, the compound is \( \text{Mg}_3(\text{PO}_4)_2\cdot5 \text{ H}_2\text{O} \)
B. First, determine the mole ratio of the elements in this compound:

\[
\begin{align*}
17.15 \text{ g Cu} & \times \left( \frac{1 \text{ mol Cu}}{63.546 \text{ g Cu}} \right) = 0.27 \text{ mol Cu} \\
19.14 \text{ g Cl} & \times \left( \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} \right) = 0.54 \text{ mol Cl} \\
60.45 \text{ g O} & \times \left( \frac{1 \text{ mol O}}{16.0 \text{ g O}} \right) = 3.78 \text{ mol O}
\end{align*}
\]

Mass of H: 100 – (17.15 + 19.14 + 60.45) = 3.26 g H

\[
3.26 \text{ g H} \times \left( \frac{1 \text{ mol H}}{1.01 \text{ g H}} \right) = 3.23 \text{ mol H}
\]

Mole ratios are determined by dividing by the smallest number:

\[
\begin{align*}
0.27 \text{ mol Cu} / 0.27 \text{ mol Cu} &= 1.0 \\
0.54 \text{ mol Cl} / 0.27 \text{ mol Cu} &= 2.0 \\
3.78 \text{ mol O} / 0.27 \text{ mol Cu} &= 14.0 \\
3.23 \text{ mol H} / 0.27 \text{ mol Cu} &= 12.0
\end{align*}
\]

Now we know that since all the hydrogen atoms are taken up as water, half as many moles of O are also taken up as water. Therefore, if there are 12 moles of H, 6 moles of O are needed, 6 moles of \( \text{H}_2\text{O} \) are generated, and 8 moles of O are left behind.

To determine the oxidation state of Cu and Cl, we note that there are 4 times as many moles of O as there is Cl. If the Cl and O are associated, we have the perchlorate ion (\( \text{ClO}_4^- \)) and the formula of the compound is \( \text{Cu(ClO}_4)\text{)}_2\cdot6\text{H}_2\text{O} \). The oxidation state of Cu is +2 and Cl is +7.

**EXERCISES**

**Representing Molecules**

1. 
   (a) \( \text{H}_2\text{O}_2 \)  
   (b) \( \text{CH}_3\text{CH}_2\text{Cl} \)  
   (c) \( \text{P}_4\text{O}_{10} \)  
   (d) \( \text{CH}_3\text{CH(OH)}\text{CH}_3 \)  
   (e) \( \text{HCO}_2\text{H} \)

3. 
   (b) \( \text{CH}_3\text{CH}_2\text{Cl} \)  
   (d) \( \text{CH}_3\text{CH(OH)}\text{CH}_3 \)  
   (e) \( \text{HCO}_2\text{H} \)

\[
\begin{align*}
\text{(a) } & \text{H}_{\text{H}} \quad \text{H}_{\text{H}} \quad \text{Cl}_{\text{H}} \\
\text{(b) } & \text{H}_{\text{H}} \quad \text{H}_{\text{H}} \quad \text{OH}_{\text{H}} \\
\text{(d) } & \text{H}_{\text{H}} \quad \text{H}_{\text{H}} \quad \text{H}_{\text{H}} \\
\text{(e) } & \text{OH} \\
\end{align*}
\]
Chapter 3: Chemical Compounds

The Avogadro Constant and the Mole

5. (a) A trinitrotoluene molecule, \( CH_3C_6H_2(NO_2)_3 \), contains 7 C atoms, 5 H atoms, 3 N atoms, and \( 3 \times 2 \) O atoms = 6 O atoms, for a total of \( 7 + 5 + 3 + 6 = 21 \) atoms.

(b) \( CH_3(CH_2)_4CH_2OH \) contains 6 C atoms, 14 H atoms, and 1 O atom, for a total of 21 atoms.

Conversion pathway approach:

\[
\text{Number of atoms} = 0.00102 \text{ mol } CH_3(CH_2)_4CH_2OH \times \frac{6.022 \times 10^{23} \text{ molecules } CH_3(CH_2)_4CH_2OH}{1 \text{ mol } CH_3(CH_2)_4CH_2OH} \times \frac{21 \text{ atoms}}{1 \text{ molecule } CH_3(CH_2)_4CH_2OH} = 1.29 \times 10^{22} \text{ atoms}
\]

Stepwise approach:

\[
0.00102 \text{ mol } CH_3(CH_2)_4CH_2OH \times \frac{6.022 \times 10^{23} \text{ molecules } CH_3(CH_2)_4CH_2OH}{1 \text{ mol } CH_3(CH_2)_4CH_2OH} = 6.14 \times 10^{20} \text{ molecules}
\]

\[
6.14 \times 10^{20} \text{ mol } \times \frac{21 \text{ atoms}}{1 \text{ C}_6\text{H}_{14} \text{O molecule}} = 1.29 \times 10^{22} \text{ atoms}
\]

(c) Conversion pathway approach:

\[
\text{Number of F atoms} = 12.15 \text{ mol } C_2HBrClF_3 \times \frac{3 \text{ mol F}}{1 \text{ mol } C_2HBrClF_3} \times \frac{6.022 \times 10^{23} \text{ F atoms}}{1 \text{ mol F atoms}} = 2.195 \times 10^{25} \text{ F atoms}
\]

Stepwise approach:

\[
12.15 \text{ mol } C_2HBrClF_3 \times \frac{3 \text{ mol F}}{1 \text{ mol } C_2HBrClF_3} = 36.45 \text{ F}
\]

\[
36.45 \text{ F} \times \frac{6.022 \times 10^{23} \text{ F atoms}}{1 \text{ mol F atoms}} = 2.195 \times 10^{25} \text{ F atoms}
\]

7. (a) molecular mass (mass of one molecule) of \( C_3H_{11}NO_2S \) is:

\[
(5 \times 12.011 \text{ u C}) + (11 \times 1.0079 \text{ u H}) + 14.0067 \text{ u N} + (2 \times 15.9994 \text{ u O}) + 32.066 \text{ u S} = 149.213 \text{ u/C}_3\text{H}_{11}\text{NO}_2\text{S molecule}
\]

(b) Since there are 11 H atoms in each \( C_3H_{11}NO_2S \) molecule, there are 11 moles of H atoms in each mole of \( C_3H_{11}NO_2S \) molecules.
Chapter 3: Chemical Compounds

(c) mass $C = 1 \text{mol C}_5\text{H}_{11}\text{NO}_2\text{S} \times \frac{5 \text{ mol C}}{1 \text{mol C}_5\text{H}_{11}\text{NO}_2\text{S}} \times \frac{12.011 \text{ g C}}{1 \text{mol C}} = 60.055 \text{ g C}$

(d) # C atoms $= 9.07 \text{ mol C}_5\text{H}_{11}\text{NO}_2\text{S} \times \frac{5 \text{ mol C}}{1 \text{mol C}_5\text{H}_{11}\text{NO}_2\text{S}} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{mol C}}$

$= 2.73 \times 10^{23} \text{ C atoms}$

9. The greatest number of N atoms is found in the compound with the greatest number of moles of N.

The molar mass of N$_2$O $= (2 \text{ mol N} \times 14.0 \text{ g N}) + (1 \text{ mol O} \times 16.0 \text{ g O}) = 44.0 \text{ g/mol N}_2\text{O}$. Thus, 50.0 g N$_2$O is slightly more than 1 mole of N$_2$O, and contains slightly more than 2 moles of N. Each mole of N$_2$ contains 2 moles of N. The molar mass of NH$_3$ is 17.0 g. Thus, there is 1 mole of NH$_3$ present, which contains 1 mole of N.

The molar mass of pyridine is $\left(5 \text{ mol C} \times 12.0 \text{ g C}\right) + \left(5 \text{ mol H} \times 1.01 \text{ g H}\right) + 14.0 \text{ g N} = 79.1 \text{ g/mol}$. Because each mole of pyridine contains 1 mole of N, we need slightly more than 2 moles of pyridine to have more N than is present in the N$_2$O. But that would be a mass of about 158 g pyridine, and 150 mL has a mass of less than 150 g. Thus, the greatest number of N atoms is present in 50.0 g N$_2$O.

11.

(a) moles N$_2$O$_4 = 115 \text{ g N}_2\text{O}_4 \times \frac{1 \text{ mol N}_2\text{O}_4}{92.02 \text{ g N}_2\text{O}_4} = 1.25 \text{ mol N}_2\text{O}_4$

(b) moles N $= 43.5 \text{ g Mg(NO}_3)_2 \times \frac{1 \text{ mol Mg(NO}_3)_2}{148.33 \text{ g}} \times \frac{2 \text{ mol N}}{1 \text{ mol Mg(NO}_3)_2} = 0.587 \text{ mol N atoms}$

(c) moles N $= 12.4 \text{ g C}_6\text{H}_{12}\text{O}_6 \times \frac{1 \text{ mol C}_6\text{H}_{12}\text{O}_6}{180.16 \text{ g}} \times \frac{6 \text{ mol O}}{1 \text{ mol C}_6\text{H}_{12}\text{O}_6} \times \frac{1 \text{ mol C}_7\text{H}_5(\text{NO}_2)_3}{6 \text{ mol O}}$

$= \frac{3 \text{ mol N}}{1 \text{ mol C}_7\text{H}_5(\text{NO}_2)_3} = 0.206 \text{ mol N}$

13. The number of Fe atoms in 6 L of blood can be found using dimensional analysis.

$= 6 \text{ L blood} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{15.5 \text{ g Hb}}{100 \text{ mL blood}} \times \frac{1 \text{ mol Hb}}{64,500 \text{ g Hb}} \times \frac{4 \text{ mol Fe}}{1 \text{ mol Hb}} \times \frac{6.022 \times 10^{23} \text{ atoms Fe}}{1 \text{ mol Fe}}$

$= 3 \times 10^{22} \text{ Fe atoms}$
Chemical Formulas

15. For glucose (blood sugar), $C_6H_{12}O_6$,
   
   (a) FALSE  The percentages by mass of C and O are different than in CO. For one thing, CO contains no hydrogen.
   
   (b) TRUE  In dihydroxyacetone, $\left(\text{CH}_2\text{OH}\right)_2\text{CO}$ or $C_3H_6O_3$, the ratio of C:H:O = 3:6:3 or 1:2:1. In glucose, this ratio is C:H:O = 6:12:6 = 1:2:1. Thus, the ratios are the same.
   
   (c) FALSE  The proportions, by number of atoms, of C and O are the same in glucose. Since, however, C and O have different molar masses, their proportions by mass must be different.
   
   (d) FALSE  Each mole of glucose contains $(12\times1.01 = 12.1 \text{ g H. But each mole also contains 72.0 g C and 96.0 g O. Thus, the highest percentage, by mass, is that of O. The highest percentage, by number of atoms, is that of H.}

17.
   
   (a)  $\text{Cu(UO}_2\text{)}_2(\text{PO}_4)_2\cdot8\text{H}_2\text{O}$ has 1 Cu, 2 U, 2 P, 20 O, and 16 H, or a total of 41 atoms.
   
   (b)  By number, $\text{Cu(UO}_2\text{)}_2(\text{PO}_4)_2\cdot8\text{H}_2\text{O}$ has a H to O ratio of 16:20 or 4:5 or 0.800 H atoms/O atom.
   
   (c)  By number, $\text{Cu(UO}_2\text{)}_2(\text{PO}_4)_2\cdot8\text{H}_2\text{O}$ has a Cu to P ratio of 1:2.
   
   (d)  With a mass percent slightly greater than 50%, U has the largest mass percent, with oxygen coming in at ~34%.
   
   \[
   \text{mass % U} = \frac{\text{mass U in Cu(UO}_2\text{)}_2(\text{PO}_4)_2\cdot8\text{H}_2\text{O}}{\text{total mass of Cu(UO}_2\text{)}_2(\text{PO}_4)_2\cdot8\text{H}_2\text{O}} \times 100 \%
   \]
   
   \[
   = \frac{2 \times 238.029 \text{ g/mol}}{937.680 \text{ g/mol}} \times 100 \%
   \]
   
   \[
   = 50.77 \%
   \]
   
   \[
   \text{mass % O} = \frac{\text{mass O in Cu(UO}_2\text{)}_2(\text{PO}_4)_2\cdot8\text{H}_2\text{O}}{\text{total mass of Cu(UO}_2\text{)}_2(\text{PO}_4)_2\cdot8\text{H}_2\text{O}} \times 100 \%
   \]
   
   \[
   = \frac{20 \times 15.9994 \text{ g/mol}}{937.680 \text{ g/mol}} \times 100 \%
   \]
   
   \[
   = 34.13 \%
   \]
   
   (e)  $1.00 \text{ g P} \times \frac{1 \text{ mol P}}{30.9738 \text{ g P}} \times \frac{1 \text{ mol Cu(UO}_2\text{)}_2(\text{PO}_4)_2\cdot8\text{H}_2\text{O}}{2 \text{ mol P}} \times \frac{937.666 \text{ g Cu(UO}_2\text{)}_2(\text{PO}_4)_2\cdot8\text{H}_2\text{O}}{1 \text{ mol Cu(UO}_2\text{)}_2(\text{PO}_4)_2\cdot8\text{H}_2\text{O}} = 15.1 \text{ g of Cu(UO}_2\text{)}_2(\text{PO}_4)_2\cdot8\text{H}_2\text{O}}
Percent Composition of Compounds

19. The information obtained in the course of calculating the molar mass is used to determine the mass percent of H in decane.

\[
molar \ mass \ C_{10}H_{22} = \left( \frac{10 \ mol \ C}{1 \ mol \ C_{10}H_{22}} \times 12.011 \ g \ C \right) + \left( \frac{22 \ mol \ H}{1 \ mol \ C_{10}H_{22}} \times 1.00794 \ g \ H \right)
\]

\[
= \frac{120.11 \ g \ C}{1 \ mol \ C_{10}H_{22}} + \frac{22.1747 \ g \ H}{1 \ mol \ C_{10}H_{22}} = \frac{142.28 \ g}{1 \ mol \ C_{10}H_{22}}
\]

\[
\% H = \frac{22.1747 \ g \ H}{142.28 \ g \ C_{10}H_{22}} \times 100\% = 15.585\% H
\]

21. C(CH₃)₃CH₂CH(CH₃)₂ has a molar mass of 114.231 g/mol and one mole contains 18.143 g of H.

\[
\text{percent hydrogen in sample} = \frac{18.143 \ g}{114.231 \ g} \times 100\% = 15.88\% H
\]

23. \[
molar \ mass = (20 \ mol \ C \times 12.011 \ g \ C) + (24 \ mol \ H \times 1.00794 \ g \ H) + (2 \ mol \ N \times 14.0067 \ g \ N)
\]

\[
+ (2 \ mol \ O \times 15.9994 \ g \ O) = 324.42 \ g/mol
\]

\[
\% C = \frac{240.22}{324.42} \times 100\% = 74.046\% C \quad \% H = \frac{24.1906}{324.42} \times 100\% = 7.4566\% H
\]

\[
\% N = \frac{28.0134}{324.42} \times 100\% = 8.6349\% N \quad \% O = \frac{31.9988}{324.42} \times 100\% = 9.8634\% O
\]

25. In each case, we first determine the molar mass of the compound, and then the mass of the indicated element in one mole of the compound. Finally, we determine the percent by mass of the indicated element to four significant figures.

(a) \[
molar \ mass \ Pb(C₂H₅)₄ = 207.2 \ g \ Pb + (8 \times 12.011 \ g \ C) + (20 \times 1.00794 \ g \ H)
\]

\[
= 323.447 \ g/mol \ Pb(C₂H₅)₄
\]

\[
\text{mass} \ Pb/mol \ Pb(C₂H₅)₄ = \frac{1 \ mol \ Pb}{1 \ mol \ Pb(C₂H₅)₄} \times \frac{207.2 \ g \ Pb}{1 \ mol \ Pb} = 207.2 \ g/mol \ Pb(C₂H₅)₄
\]

\[
\% Pb = \frac{207.2 \ g \ Pb}{323.447 \ g \ Pb(C₂H₅)₄} \times 100\% = 64.06\% Pb
\]

(b) \[
molar \ mass \ Fe₄[Fe(CN)₆]₃ = (7 \times 55.847 \ g \ Fe) + (18 \times 12.011 \ g \ C) + (18 \times 14.0067 \ g \ N)
\]

\[
= 859.248 \ g/mol \ Fe₄[Fe(CN)₆]₃
\]

\[
\frac{\text{mass} \ Fe}{mol \ Fe₄[Fe(CN)₆]₃} = \frac{7 \ mol \ Fe}{1 \ mol \ Fe₄[Fe(CN)₆]₃} \times \frac{55.847 \ g \ Fe}{1 \ mol \ Fe} = 390.929 \ g/mol \ Fe₄[Fe(CN)₆]₃
\]
%Fe = \( \frac{390.929 \text{ g Fe}}{859.248 \text{ g Fe}_4[\text{Fe(CN)}_6]_3} \times 100\% = 45.497 \% \text{ Fe} \)

(c) Molar mass \( C_{55}H_{72}MgN_4O_{5} \)

\[
\text{mass Mg} = \frac{1 \text{ mol Mg}}{1 \text{ mol } C_{55}H_{72}MgN_4O_{5}} \times \frac{24.305 \text{ g Mg}}{1 \text{ mol Mg}} = \frac{24.305 \text{ g Mg}}{1 \text{ mol } C_{55}H_{72}MgN_4O_{5}}
\]

\[
\%Mg = \left( \frac{24.305 \text{ g Mg}}{893.505 \text{ g } C_{55}H_{72}MgN_4O_{5}} \right) \times 100\% = 2.7202 \% \text{ Mg}
\]

27. Oxide with the largest %Cr will have the largest number of moles of Cr per mole of oxygen.

\[
\text{CrO:} \quad \frac{1 \text{ mol Cr}}{1 \text{ mol } O} = 1 \text{ mol Cr/mol O} \quad \text{Cr_2O_3:} \quad \frac{2 \text{ mol Cr}}{3 \text{ mol O}} = 0.667 \text{ mol Cr/mol O}
\]

\[
\text{CrO}_2: \quad \frac{1 \text{ mol Cr}}{2 \text{ mol O}} = 0.500 \text{ mol Cr/mol O} \quad \text{CrO_3:} \quad \frac{1 \text{ mol Cr}}{3 \text{ mol O}} = 0.333 \text{ mol Cr/mol O}
\]

Arranged in order of increasing %Cr: \( \text{CrO}_3 < \text{CrO}_2 < \text{Cr}_2\text{O}_3 < \text{CrO} \)

**Chemical Formulas from Percent Composition**

29. \( \text{SO}_3 \) (40.05% S) and \( \text{S}_2\text{O} \) (80.0% S) (2 O atoms \( \approx \) 1 S atom in terms of atomic masses)

Note the molar masses are quite close (within 0.05 g/mol).

31. Determine the % oxygen by difference. \( \% \text{O} = 100.00\% - 45.27\% \text{C} - 9.50\% \text{H} = 45.23\% \text{O} \)

\[
\text{mol O} = 45.23 \text{ g} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.827 \text{ mol O} \quad + \quad 2.827 \rightarrow 1.000 \text{ mol O}
\]

\[
\text{mol C} = 45.27 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 3.769 \text{ mol C} \quad + \quad 2.827 \rightarrow 1.333 \text{ mol C}
\]

\[
\text{mol H} = 9.50 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 9.42 \text{ mol H} \quad + \quad 2.827 \rightarrow 3.33 \text{ mol H}
\]

Multiply all amounts by 3 to obtain integers. Empirical formula is \( C_4H_{10}O_3 \).
33.  
(a)  
\[ 74.01 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 6.162 \text{ mol C} \times 1.298 \rightarrow 4.747 \text{ mol C} \]
\[ 5.23 \text{ g H} \times \frac{1 \text{ mol H}}{1.01 \text{ g H}} = 5.18 \text{ mol H} \times 1.298 \rightarrow 3.99 \text{ mol H} \]
\[ 20.76 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.298 \text{ mol O} \times 1.298 \rightarrow 1.00 \text{ mol O} \]
Multiply each of the mole numbers by 4 to obtain an empirical formula of \( C_{19}H_{16}O_4 \).

(b)  
\[ 39.98 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 3.3286 \text{ mol C} \times 0.7397 \rightarrow 4.500 \text{ mol C} \]
\[ 3.73 \text{ g H} \times \frac{1 \text{ mol H}}{1.00794 \text{ g H}} = 3.70 \text{ mol H} \times 0.7397 \rightarrow 5.00 \text{ mol H} \]
\[ 20.73 \text{ g N} \times \frac{1 \text{ mol N}}{14.0067 \text{ g N}} = 1.480 \text{ mol N} \times 0.7397 \rightarrow 2.001 \text{ mol N} \]
\[ 11.84 \text{ g O} \times \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 0.7400 \text{ mol O} \times 0.7397 \rightarrow 1.000 \text{ mol O} \]
\[ 23.72 \text{ g S} \times \frac{1 \text{ mol S}}{32.066 \text{ g S}} = 0.7397 \text{ mol S} \times 0.7397 \rightarrow 1.000 \text{ mol S} \]
Multiply by 2 to obtain the empirical formula \( C_9H_{10}N_4O_2S_2 \).

35. Convert each percentage into the mass in 100.00 g, and then to the moles of that element.

\[ 94.34 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 7.854 \text{ mol C} \div 5.615 = 1.40 \text{ mol C} \times 5 = 7.00 \]
\[ 5.66 \text{ g H} \times \frac{1 \text{ mol H}}{1.00794 \text{ g H}} = 5.615 \text{ mol H} \div 5.615 = 1.00 \text{ mol H} \times 5 = 5.00 \]
Multiply by 5 to achieve whole number ratios. The empirical formula is \( C_9H_7 \), and the formula mass \( [(7 \times 12.011 \text{ g C})+(5 \times 1.00794 \text{ g H})] = 89.117 \text{ u} \). Since this empirical molar mass is one-half of the 178 u, the correct molecular mass, the molecular formula must be twice the empirical formula. Molecular formula: \( C_{14}H_{10} \)

37. Determine the mass of oxygen by difference. Then convert all masses to amounts in moles.

\[ \text{oxygen mass} = 100.00 \text{ g} - 73.27 \text{ g C} - 3.84 \text{ g H} - 10.68 \text{ g N} = 12.21 \text{ g O} \]
The empirical formula is $C_8H_5NO$, which has an empirical mass of 131 u. This is almost exactly half the molecular mass of 262.3 u. Thus, the molecular formula is twice the empirical formula and is $C_{16}H_{10}N_2O_2$.

**39.** The molar mass of element X has the units of grams per mole. We can determine the amount, in moles of Cl, and convert that to the amount of X, equivalent to 25.0 g of X.

$$\text{molar mass} = \frac{25.0 \text{ g X}}{75.0 \text{ g Cl}} \times \frac{35.453 \text{ g Cl}}{1 \text{ mol Cl}} \times \frac{4 \text{ mol Cl}}{1 \text{ mol X}} = 47.3 \text{ g X}$$

The atomic mass is 47.3 u. This atomic mass is close to that of the element titanium, which therefore is identified as element X.

**41.** Consider 100 g of chlorophyll, which contains 2.72 g of Mg. To answer this problem, we must take note of the fact that 1 mole of Mg contains 1 mole of chlorophyll.

$$\frac{100 \text{ g chlorophyll}}{2.72 \text{ g Mg}} \times \frac{24.305 \text{ g Mg}}{1 \text{ mol Mg}} \times \frac{1 \text{ mol chlorophyll}}{1 \text{ mol Mg}} = 894 \text{ g mol}^{-1}$$

Therefore, the molecular mass of chlorophyll is 894 u.

**Combustion Analysis**

**43. (a)** First we determine the mass of carbon and of hydrogen present in the sample. Remember that a hydrocarbon contains only hydrogen and carbon.

$$0.6260 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.01422 \text{ mol C} \times \frac{12.011 \text{ g C}}{1 \text{ mol C}} = 0.17085 \text{ g C}$$

$$0.1602 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0153 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.017785 \text{ mol H} \times \frac{1.00794 \text{ g H}}{1 \text{ mol H}} = 0.017926 \text{ g H}$$

Then the % C and % H are found.

$$\% C = \frac{0.17085}{0.1888 \text{ g cmpd}} \times 100\% = 90.49\% \text{ C} \quad \% H = \frac{0.017926 \text{ g H}}{0.1888 \text{ g cmpd}} \times 100\% = 9.495\% \text{ H}$$
(b) Use the moles of C and H from part (a), and divide both by the smallest value, namely 0.01422 mol. Thus 0.01422 mol C ÷ 0.01422 mol = 1 mol H; 0.017785 mol H ÷ 0.01422 mol = 1.251 mol H. The empirical formula is obtained by multiplying these mole numbers by 4. It is \( \text{C}_4\text{H}_5 \).

(c) The molar mass of the empirical formula \( \text{C}_4\text{H}_5 \) \( \left( \frac{4 \times 12.0 \text{ g C}}{1 \text{ mol C}} + \frac{5 \times 1.0 \text{ g H}}{1 \text{ mol H}} \right) \) = 53.0 g/mol. This value is 1/2 of the actual molar mass. The molecular formula is twice the empirical formula. \( \implies \) Molecular formula: \( \text{C}_8\text{H}_{10} \).

45. First, determine the mass of carbon and hydrogen present in the sample.

\[
0.458 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.0104 \text{ mol C} \times \frac{12.011 \text{ g C}}{1 \text{ mol C}} = 0.125 \text{ g C}
\]

\[
0.374 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0153 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.0415 \text{ mol H} \times \frac{1.00794 \text{ g H}}{1 \text{ mol H}} = 0.04185 \text{ g H}
\]

Then, the mass of N that this sample would have produced is determined. (Note that this is also the mass of \( \text{N}_2 \) produced in the reaction.)

\[
0.226 \text{ g N}_2 \times \frac{0.312 \text{ g 1st sample}}{0.486 \text{ g 2nd sample}} = 0.145 \text{ g N}_2
\]

From which we can calculate the mass of N in the sample.

\[
0.145 \text{ g N}_2 \times \frac{1 \text{ mol N}_2}{28.0134 \text{ g N}_2} \times \frac{2 \text{ mol N}}{1 \text{ mol N}_2} \times \frac{14.0067 \text{ g N}}{1 \text{ mol N}} = 0.145 \text{ g N}
\]

We may alternatively determine the mass of N by difference:

\[
0.312 \text{ g} - 0.125 \text{ g C} - 0.04185 \text{ g H} = 0.145 \text{ g N}
\]

Then, we can calculate the relative number of moles of each element.

\[
0.145 \text{ g N} \times \frac{1 \text{ mol N}}{14.0067 \text{ g N}} = \begin{cases} 0.01035 \text{ mol N} & \rightarrow 1.00 \text{ mol N} \\ 0.0104 \text{ mol C} & \rightarrow 1.00 \text{ mol C} \\ 0.0415 \text{ mol H} & \rightarrow 4.01 \text{ mol H} \end{cases}
\]

Thus, the empirical formula is \( \text{CH}_4\text{N} \).

47. Each mole of \( \text{CO}_2 \) is produced from a mole of C. Therefore, the compound with the largest number of moles of C per mole of the compound will produce the largest amount of \( \text{CO}_2 \) and, thus, also the largest mass of \( \text{CO}_2 \). Of the compounds listed, namely \( \text{CH}_4 \), \( \text{C}_2\text{H}_5\text{OH} \), \( \text{C}_{10}\text{H}_8 \),
and C₆H₅OH, C₁₀H₈ has the largest number of moles of C per mole of the compound and will produce the greatest mass of CO₂ per mole on complete combustion.

49. The molecular formula for CH₃CHOHCH₂CH₃ is C₄H₁₀O. Here we will use the fact that C₄H₁₀O has a molar mass of 74.1228 g/mol to calculate the masses of CO₂ and H₂O:

Mass of CO₂:

**Conversion pathway approach:**

\[
\text{Mass of CO}_2 = 1.562 \text{ g C}_4\text{H}_{10}\text{O} \times \frac{1 \text{ mol C}_4\text{H}_{10}\text{O}}{74.1228 \text{ g C}_4\text{H}_{10}\text{O}} \times \frac{4 \text{ mol C}}{1 \text{ mol C}_4\text{H}_{10}\text{O}} \times \frac{44.010 \text{ g CO}_2}{1 \text{ mol CO}_2} = 3.710 \text{ g CO}_2
\]

**Stepwise approach:**

\[
1.562 \text{ g C}_4\text{H}_{10}\text{O} \times \frac{1 \text{ mol C}_4\text{H}_{10}\text{O}}{74.1228 \text{ g C}_4\text{H}_{10}\text{O}} = 0.02107 \text{ mol C}_4\text{H}_{10}\text{O} \times \frac{4 \text{ mol C}}{1 \text{ mol C}_4\text{H}_{10}\text{O}} = 0.08429 \text{ mol C}
\]

\[
0.08429 \text{ mol C} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol C}} = 0.08429 \text{ mol CO}_2 \times \frac{44.010 \text{ g CO}_2}{1 \text{ mol CO}_2} = 3.710 \text{ g CO}_2
\]

Mass of H₂O:

**Conversion pathway approach:**

\[
\text{Mass of H}_2\text{O} = 1.562 \text{ g C}_4\text{H}_{10}\text{O} \times \frac{1 \text{ mol C}_4\text{H}_{10}\text{O}}{74.1228 \text{ g C}_4\text{H}_{10}\text{O}} \times \frac{10 \text{ mol H}}{1 \text{ mol C}_4\text{H}_{10}\text{O}} \times \frac{1 \text{ mol H}_2\text{O}}{2 \text{ mol H}} \times \frac{18.0153 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 1.898 \text{ g H}_2\text{O}
\]

**Stepwise approach:**

\[
1.562 \text{ g C}_4\text{H}_{10}\text{O} \times \frac{1 \text{ mol C}_4\text{H}_{10}\text{O}}{74.1228 \text{ g C}_4\text{H}_{10}\text{O}} = 0.02107 \text{ mol C}_4\text{H}_{10}\text{O}
\]

\[
0.02107 \text{ mol C}_4\text{H}_{10}\text{O} \times \frac{10 \text{ mol H}}{1 \text{ mol C}_4\text{H}_{10}\text{O}} = 0.2107 \text{ mol H}
\]

\[
0.2107 \text{ mol H} \times \frac{1 \text{ mol H}_2\text{O}}{2 \text{ mol H}} = 0.1054 \text{ mol H}_2\text{O} \times \frac{18.0153 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 1.898 \text{ g H}_2\text{O}
\]

**Oxidation States**

51. The oxidation state (O.S.) is given first, followed by the explanation for its assignment.

(a) C = −4 in CH₄ H has an oxidation state of +1 in its non-metal compounds. (Remember that the sum of the oxidation states in a neutral compound equals 0.)

(b) S = +4 in SF₄ F has O.S. = −1 in its compounds.

(c) O = −1 in Na₂O₂ Na has O.S. = +1 in its compounds.

(d) C = 0 in C₂H₅O₂⁻ H has O.S. = +1 in its non-metal compounds; that of O = −2 (usually). (Remember that the sum of the oxidation states in a
polyatomic ion equals the charge on that ion.)

(e) \( \text{Fe} = +6 \) in \( \text{FeO}_4^{2-} \) O has O.S. = \(-2\) in most of its compounds (especially metal containing compounds).

53. Remember that the oxidation state of oxygen is usually \(-2\) in its compounds. \( \text{Cr}^{3+} \) and \( \text{O}^{2-} \) form \( \text{Cr}_2\text{O}_3 \), chromium(III) oxide. \( \text{Cr}^{4+} \) and \( \text{O}^{2-} \) form \( \text{CrO}_2 \), chromium (IV) oxide. \( \text{Cr}^{6+} \) and \( \text{O}^{2-} \) form \( \text{CrO}_3 \), chromium(VI) oxide.

55. (a) \( \text{O} = +2 \) in \( \text{OF}_2 \) F has an oxidation state of -1 in its compounds.
(b) \( \text{O} = +1 \) in \( \text{O}_2\text{F}_2 \) F has O.S. = \(-1\) in its compounds.
(c) \( \text{O} = \frac{\text{O}}{2} \) in \( \text{CsO}_2 \) Cs has O.S. = +1 in its compounds.
(d) \( \text{O} = -1 \) in \( \text{BaO}_2 \) Ba has O.S. = +2 in its compounds.

Nomenclature

57. (a) \( \text{SrO} \) strontium oxide (b) \( \text{ZnS} \) zinc sulfide
(c) \( \text{K}_2\text{CrO}_4 \) potassium chromate (d) \( \text{Cs}_2\text{SO}_4 \) cesium sulfate
(e) \( \text{Cr}_2\text{O}_3 \) chromium(III) oxide (f) \( \text{Fe}_2(\text{SO}_4)_3 \) iron(III) sulfate
(g) \( \text{Mg}\left(\text{HCO}_3\right)_2 \) magnesium hydrogen carbonate or magnesium bicarbonate
(h) \( \left(\text{NH}_4\right)_2\text{HPO}_4 \) ammonium hydrogen phosphate
(i) \( \text{Ca}\left(\text{HSO}_3\right)_2 \) calcium hydrogen sulfite (j) \( \text{Cu(OH)}_2 \) copper(II) hydroxide
(k) \( \text{HNO}_3 \) nitric acid (l) \( \text{KClO}_4 \) potassium perchlorate
(m) \( \text{HBrO}_3 \) bromic acid (n) \( \text{H}_3\text{PO}_4 \) phosphorous acid

59. (a) \( \text{CS}_2 \) carbon disulfide (b) \( \text{SiF}_4 \) silicon tetrafluoride
(c) \( \text{ClF}_5 \) chlorine pentafluoride (d) \( \text{N}_2\text{O}_5 \) dinitrogen pentoxide
(e) \( \text{SF}_6 \) sulfur hexafluoride (f) \( \text{I}_2\text{Cl}_6 \) diiodine hexachloride
61. 
(a) $\text{Al}_2(\text{SO}_4)_3$ aluminum sulfate  
(b) $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$ ammonium dichromate  
(c) $\text{SiF}_4$ silicon tetrafluoride  
(d) $\text{Fe}_2\text{O}_3$ iron(III) oxide  
(e) $\text{C}_3\text{S}_2$ tricarbon disulfide  
(f) $\text{Co}(\text{NO}_3)_2$ cobalt(II) nitrate  
(g) $\text{Sr}(\text{NO}_3)_2$ strontium nitrite  
(h) $\text{HBr(aq)}$ hydrobromic acid  
(i) $\text{HIO}_3$ iodic acid  
(j) $\text{PCl}_3\text{F}_3$ phosphorus dichloride trifluoride

63. 
(a) $\text{Ti}^{4+}$ and $\text{Cl}^-$ produce $\text{TiCl}_4$  
(b) $\text{Fe}^{3+}$ and $\text{SO}_4^{2-}$ produce $\text{Fe}_2(\text{SO}_4)_3$  
(c) $\text{Cl}^{2+}$ and $\text{O}^{2-}$ produce $\text{Cl}_2\text{O}_7$  
(d) $\text{S}^{2+}$ and $\text{O}^{2-}$ produce $\text{S}_2\text{O}_8^{2-}$

65. 
(a) $\text{HClO}_2$ chlorous acid  
(b) $\text{H}_2\text{SO}_3$ sulfurous acid  
(c) $\text{H}_2\text{Se}$ hydroselenic acid  
(d) $\text{HNO}_2$ nitrous acid

67. 
(a) $\text{OF}_2$ oxygen difluoride  
(b) $\text{XeF}_2$ xenon difluoride  
(c) $\text{CuSO}_3$ copper (II) sulfite  
(d) $(\text{NH}_4)_2\text{HPO}_4$ ammonium hydrogen phosphate

Both (c) and (d) are ionic compounds.

Hydrates

69. The hydrate with the greatest mass percent $\text{H}_2\text{O}$ is the one that gives the largest result for the number of moles of water in the hydrate’s empirical formula, divided by the mass of one mole of the anhydrous salt for the hydrate.

$$\frac{5\text{H}_2\text{O}}{\text{CuSO}_4} = \frac{5\text{mol H}_2\text{O}}{159.6\text{g}} = 0.03133$$
$$\frac{6\text{H}_2\text{O}}{\text{MgCl}_2} = \frac{6\text{mol H}_2\text{O}}{95.2\text{g}} = 0.0630$$
$$\frac{18\text{H}_2\text{O}}{\text{Cr}_2(\text{SO}_4)_3} = \frac{18\text{mol H}_2\text{O}}{392.3\text{g}} = 0.04588$$
$$\frac{2\text{H}_2\text{O}}{\text{LiC}_2\text{H}_3\text{O}_2} = \frac{2\text{mol H}_2\text{O}}{66.0\text{g}} = 0.0303$$

The hydrate with the greatest % $\text{H}_2\text{O}$ therefore is $\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$

71. molar mass $\text{CuSO}_4 = 63.546\text{ g Cu} + 32.066\text{ g S} + (4 \times 15.9994\text{ g O}) = 159.61\text{ g CuSO}_4/\text{mol}$. Note that each $\text{CuSO}_4$ will pick up 5 equivalents of $\text{H}_2\text{O}$ to give $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$. 

54
Conversion pathway approach:

\[
\text{mass of required CuSO}_4 = 12.6 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0153 \text{ g H}_2\text{O}} \times \frac{1 \text{ mol CuSO}_4}{5 \text{ mol H}_2\text{O}} \times \frac{159.61 \text{ g CuSO}_4}{1 \text{ mol CuSO}_4}
\]

\[
= 22.3 \text{ g CuSO}_4 \quad \text{is the minimum amount required to remove all the water}
\]

Stepwise approach:

\[
12.6 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0153 \text{ g H}_2\text{O}} = 0.699 \text{ mol H}_2\text{O} \times \frac{1 \text{ mol CuSO}_4}{5 \text{ mol H}_2\text{O}} = 0.140 \text{ mol CuSO}_4
\]

\[
0.140 \text{ mol CuSO}_4 \times \frac{159.61 \text{ g CuSO}_4}{1 \text{ mol CuSO}_4} = 22.3 \text{ g CuSO}_4
\]

\[
= \text{is the minimum amount required to remove all the water}
\]

73. We start by converting to molar amounts for each element based on 100.0 g:

\[
20.3 \text{ g Cu} \times \frac{1 \text{ mol Cu}}{63.546 \text{ g Cu}} = 0.319 \text{ mol Cu} \quad \div \quad 0.319 \rightarrow 1.00 \text{ mol Cu}
\]

\[
8.95 \text{ g Si} \times \frac{1 \text{ mol Si}}{28.0855 \text{ g Si}} = 0.319 \text{ mol Si} \quad \div \quad 0.319 \rightarrow 1.00 \text{ mol Si}
\]

\[
36.3 \text{ g F} \times \frac{1 \text{ mol F}}{18.9984 \text{ g F}} = 1.91 \text{ mol F} \quad \div \quad 0.319 \rightarrow 5.99 \text{ mol F}
\]

\[
34.5 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0153 \text{ g H}_2\text{O}} = 1.915 \text{ mol H}_2\text{O} \quad \div \quad 0.319 \rightarrow 6.00 \text{ mol H}_2\text{O}
\]

Thus the empirical formula for the hydrate is CuSiF$_6$·6H$_2$O.

Organic Compounds and Organic Nomenclature

75. Answer is (b), 2-butanol is the most appropriate name for this molecule. It has a four carbon atom chain with a hydroxyl group on the carbon second from the end.

77. Molecules (a), (b), (c), and (d) are structural isomers. They share a common formula, namely C$_3$H$_{12}$O, but have different molecular structures. Molecule (e) has a different chemical formula (C$_6$H$_{14}$O) and hence cannot be classified as an isomer. It should be pointed out that molecules (a) and (c) are identical as well as being isomers of (b).

79.

(a) CH$_3$(CH$_2$)$_3$CH$_3$  
(b) CH$_3$CH$_2$CO$_2$H  
(c) CH$_3$CH$_2$CH$_2$CH(CH$_3$)CH$_2$OH  
(d) CH$_3$CH$_2$F

81.

(a) methanol; CH$_3$OH;  
(b) 2-chlorohexane; CH$_3$(CH$_2$)$_2$CHClCH$_3$  
Molecular mass = 120.6 u
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(c) pentanoic acid; \( \text{CH}_3(\text{CH}_2)_3\text{CO}_2\text{H} \)  Molecular mass = 102.1 u

(d) 2-methyl-1-propanol; \( \text{CH}_3\text{CH}(\text{CH}_3)\text{CH}_2\text{OH} \)  Molecular mass = 74.12 u

---

INTEGRATIVE AND ADVANCED EXERCISES

83. molar mass = \((1 \times 6.941 \text{ g Li}) + (1 \times 26.9815 \text{ g Al}) + (2 \times 28.0855 \text{ g Si}) + (6 \times 15.9994 \text{ g O})\) = 186.09 g/mol

*Conversion pathway approach:*

\[
\text{number of Li - 6 atoms} = \frac{518 \text{ g spodumene}}{186.09 \text{ g spodumene}} \times \frac{1 \text{ mol spodumene}}{186.09 \text{ g spodumene}} \times \frac{1 \text{ mol Li}}{1 \text{ mol spodumene}} \times \frac{7.40 \text{ mol Li - 6}}{100.00 \text{ mol total Li}} \times \frac{6.022 \times 10^{23} \text{ Li - 6 atoms}}{1 \text{ mol Li - 6}} = 1.24 \times 10^{23} \text{ Li - 6 atoms}
\]

*Stepwise approach:*

\[
518 \text{ g spodumene} \times \frac{1 \text{ mol spodumene}}{186.09 \text{ g spodumene}} = 2.78 \text{ mol spodumene}
\]

\[
2.78 \text{ mol spodumene} \times \frac{1 \text{ mol Li}}{1 \text{ mol spodumene}} = 2.78 \text{ mol lithium}
\]

\[
2.78 \text{ mol lithium} \times \frac{7.40 \text{ mol lithium - 6}}{100.00 \text{ mol total Li}} = 0.206 \text{ mol lithium-6}
\]

\[
0.206 \text{ mol lithium-6} \times \frac{6.022 \times 10^{23} \text{ lithium - 6 atoms}}{1 \text{ mol lithium - 6}} = 1.24 \times 10^{23} \text{ Li - 6 atoms}
\]

86. First, we determine the amount of each mineral necessary to obtain 1 kg or 1000 g of boron.

\[
1000 \text{ g B} \times \frac{1 \text{ mol B}}{10.811 \text{ g B}} \times \frac{1 \text{ mol } \text{Na}_2\text{B}_4\text{O}_7 \cdot 4\text{H}_2\text{O}}{4 \text{ mol B}} \times \frac{273.28 \text{ g } \text{Na}_2\text{B}_4\text{O}_7 \cdot 4\text{H}_2\text{O}}{1 \text{ mol } \text{Na}_2\text{B}_4\text{O}_7 \cdot 4\text{H}_2\text{O}} = 6319.5 \text{ g } \text{Na}_2\text{B}_4\text{O}_7 \cdot 4\text{H}_2\text{O}
\]

\[
1000 \text{ g B} \times \frac{1 \text{ mol B}}{10.811 \text{ g B}} \times \frac{1 \text{ mol } \text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}}{4 \text{ mol B}} \times \frac{381.372 \text{ g } \text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}}{1 \text{ mol } \text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}} = 8819.1 \text{ g } \text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}
\]

The difference between these two masses is the required additional mass. Hence, 8819.1 g - 6319.5 g = 2499.6 g. Thus, an additional 2.500 kg mass is required.

89. It is not possible to have less than 1 molecule of \( \text{S}_8 \). In order to determine whether it is possible to have \( 1.00 \times 10^{-23} \) g of \( \text{S}_8 \), determine how many molecules that number is equivalent to.
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\[ 1.00 \times 10^{-23} \text{ g S}_8 \times \frac{1 \text{ mol S}_8}{256.528 \text{ g S}_8} \times \frac{6.022 \times 10^{23} \text{ molecules S}_8}{1 \text{ mol S}_8} = 0.023 \text{ molecules S}_8 \]

Therefore it is **not** possible to have \( 1.00 \times 10^{-23} \text{ g of S}_8 \).

\[ 1 \text{ molec S}_8 \times \frac{1 \text{ mol S}_8}{6.022 \times 10^{23} \text{ molec S}_8} \times \frac{256.53 \text{ g S}}{1 \text{ mol S}_8} = 4.26 \times 10^{-22} \text{ g S} \]

\[ 4.26 \times 10^{-22} \text{ g S} \times \frac{1 \text{ yg S}}{10^{-24} \text{ g S}} = 426 \text{ yoctograms} \]

**91.** We determine the masses of CO\(_2\) and H\(_2\)O produced by burning the C\(_3\)H\(_8\).

\[
\text{mass}_{\text{CO}_2} = 6.00 \text{ g C}_3\text{H}_8 \times \frac{1 \text{ mol C}_3\text{H}_8}{44.0965 \text{ g C}_3\text{H}_8} \times \frac{3 \text{ mol C}}{1 \text{ mol C}_3\text{H}_8} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol C}} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 17.96 \text{ g CO}_2
\]

\[
\text{mass}_{\text{H}_2\text{O}} = 6.00 \text{ g C}_3\text{H}_8 \times \frac{1 \text{ mol C}_3\text{H}_8}{44.0965 \text{ g C}_3\text{H}_8} \times \frac{8 \text{ mol H}}{1 \text{ mol C}_3\text{H}_8} \times \frac{1 \text{ mol H}_2\text{O}}{2 \text{ mol H}} \times \frac{18.0153 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 9.805 \text{ g H}_2\text{O}
\]

Then, from the masses of CO\(_2\) and H\(_2\)O in the unknown compound, we determine the amounts of C and H in that compound and finally its empirical formula.

\[
\text{amount C} = \frac{(29.0 - 17.96) \text{ g CO}_2}{44.01 \text{ g CO}_2} = 0.251 \text{ mol C}
\]

\[
\text{amount H} = \frac{(18.8 - 9.805) \text{ g H}_2\text{O}}{18.0153 \text{ g H}_2\text{O}} = 0.9986 \text{ mol H}
\]

The empirical formula of the unknown compound is CH\(_4\). The C:H ratio is 0.9986/0.251 = 3.98.

The molecular formula can be calculated by knowing that we have 0.251 moles, which accounts for the 4.00 g of hydrocarbon (40 % of 10.0 g). This gives a molar mass of 4.00/0.251 = 15.9 g/mol. This is nearly the same as the molar mass of the empirical formula CH\(_4\) (16.04 g/mol).

**93.** Since the compound is composed of H\(_2\)SO\(_4\) and H\(_2\)O, we will need to determine the percent composition of both H\(_2\)SO\(_4\) and water.

\[
\% \text{ H}_2\text{SO}_4 = \frac{\# \text{ grams H}_2\text{SO}_4}{\text{total mass in grams}} \times 100\%
\]

\[
\# \text{ g H}_2\text{SO}_4 = 65.2 \text{ g (NH}_4\text{)}_2\text{SO}_4 \times \frac{1 \text{ mol (NH}_4\text{)}_2\text{SO}_4}{132.15 \text{ g (NH}_4\text{)}_2\text{SO}_4} \times \frac{1 \text{ mol H}_2\text{SO}_4}{1 \text{ mol (NH}_4\text{)}_2\text{SO}_4} \times \frac{98.08 \text{ g H}_2\text{SO}_4}{1 \text{ mol H}_2\text{SO}_4} = 48.4 \text{ g}
\]

\[
\text{total mass} = 32.0 \text{ mL mixture} \times \frac{1.78 \text{ g mixture}}{1 \text{ mL mixture}} = 57.0 \text{ g mixture}
\]

\[
\% \text{ H}_2\text{SO}_4 = \frac{48.4 \text{ g}}{57.0 \text{ g}} \times 100\% = 85.0 \%
\]
% H₂O = 100.0 – 85.0 = 15.0 %

95.

\[
9.0 \times 10^{-11} \text{ mol C}_2\text{H}_6\text{S per m}^3 \text{ air} \times \frac{1 \times 10^{-6} \text{ mol C}_2\text{H}_6\text{S}}{1 \text{ mol C}_2\text{H}_6\text{S}} \times \frac{62.13 \text{ g C}_2\text{H}_6\text{S}}{1 \text{ mol C}_2\text{H}_6\text{S}} \times \frac{1 \text{ m}^3}{(100)^3 \text{ cm}^3} \times \frac{1 \text{ cm}^3}{1 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1 \text{ L air}}{1.2 \text{ g air}} = 4.7 \times 10^{-11} \text{ g C}_2\text{H}_6\text{S per g air}
\]

\[
4.7 \times 10^{-11} \frac{\text{g C}_2\text{H}_6\text{S}}{\text{g air}} \times \frac{1 \times 10^9 \text{ g}}{1 \text{ billion grams}} = 0.0466 \text{ ppb} = 0.05 \text{ ppb}
\]

96. a) If we have one mole of entities, then we must have 0.7808 mol N₂, 0.2095 mol O₂, 0.0093 mol Ar, and 0.0004 mol CO₂.

\[
0.7808 \text{ mol N}_2 \times \frac{28.02 \text{ g N}_2}{1 \text{ mol N}_2} = 21.88 \text{ g N}_2
\]

\[
0.2095 \text{ mol O}_2 \times \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} = 6.704 \text{ g O}_2
\]

\[
0.0004 \text{ mol CO}_2 \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 0.0176 \text{ g CO}_2
\]

\[
0.0093 \text{ mol Ar} \times \frac{39.948 \text{ g Ar}}{1 \text{ mol Ar}} = 0.3715 \text{ g Ar}
\]

mass of air sample = 21.88 g N₂ + 6.704 g O₂ + 0.0176 g CO₂ + 0.3715 g Ar = 28.97 g

\[
1 \text{ m}^3 \times \frac{(100)^3 \text{ cm}^3}{1 \text{ m}^3} \times \frac{1 \text{ mL}}{1 \text{ cm}^3} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1.2 \text{ g}}{1 \text{ L}} = 1200 \text{ g dry air}
\]

b) 1200 g \times \frac{1 \text{ mol entities}}{28.97 \text{ g}} \times \frac{1.14 \times 10^{-4} \text{ mol Kr}}{100} = 4.72 \times 10^{-5} \text{ mol Kr}

\[
4.72 \times 10^{-5} \text{ mol Kr} \times \frac{83.80 \text{ g Kr}}{1 \text{ mol}} = 3.96 \times 10^{-3} \text{ g Kr} = 4.0 \text{ mg Kr}
\]
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99. 

\[ \text{mass } SO_4^{2-} = \frac{1.511 \text{ g } BaSO_4}{233.39 \text{ g } BaSO_4} \times \frac{1 \text{ mol } BaSO_4}{1 \text{ mol } BaSO_4} \times \frac{1 \text{ mol } SO_4^{2-}}{1 \text{ mol } SO_4^{2-}} = 0.006474 \text{ mol } SO_4^{2-} \]

\[ 0.006474 \text{ mol } SO_4^{2-} \times \frac{96.064 \text{ g } SO_4^{2-}}{1 \text{ mol } SO_4^{2-}} = 0.6219 \text{ g } SO_4^{2-} \]

\[ \text{amount } M = \frac{2 \text{ mol } M^{3+}}{3 \text{ mol } SO_4^{2-}} = 0.004316 \text{ mol } M^{3+} \]

\[ \text{mass } M = 0.738 \text{ g } M_2(SO_4)_2 - 0.6219 \text{ g } SO_4^{2-} = 0.116 \text{ g } M \]

\[ \text{atomic mass of } M = \frac{0.116 \text{ g } M}{0.004316 \text{ mol } M} = 26.9 \text{ g } M/\text{mol} \]

M is the element aluminum.

103. If we determine the mass of anhydrous ZnSO₄ in the hydrate, we then can determine the mass of water, and the formula of the hydrate.

\[ \text{mass } ZnSO_4 = \frac{1 \text{ mol } BaSO_4}{233.391 \text{ g } BaSO_4} \times \frac{1 \text{ mol } ZnSO_4}{1 \text{ mol } BaSO_4} \times \frac{161.454 \text{ g } ZnSO_4}{1 \text{ mol } ZnSO_4} = 0.5688 \text{ g } ZnSO_4 \]

The water present in the hydrate is obtained by difference.

\[ \text{mass } H_2O = 1.013 \text{ g hydrate} - 0.5688 \text{ g } ZnSO_4 = 0.444 \text{ g } H_2O \]

The hydrate’s formula is determined by a method similar to that for obtaining an empirical formula.

\[ \text{amt. } ZnSO_4 = \frac{0.5688 \text{ g } \times \frac{1 \text{ mol } ZnSO_4}{161.454 \text{ mol } ZnSO_4}}{1.00 \text{ mol } ZnSO_4} = 0.003523 \text{ mol } ZnSO_4 \]

\[ \text{amt. } H_2O = \frac{0.444 \text{ g } \times \frac{1 \text{ mol } H_2O}{18.0153 \text{ g } H_2O}}{7.00 \text{ mol } H_2O} = 0.02465 \text{ mol } H_2O \]

Thus, the formula of the hydrate is ZnSO₄·7H₂O.

106. First find the mass of carbon, hydrogen, chlorine, and oxygen. From the molar ratios, we determine the molecular formula.

\[ 2.094 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.001 \text{ g } CO_2} \times \frac{1 \text{ mol } C}{1 \text{ mol } CO_2} = 0.04759 \text{ mol } C \times \frac{12.011 \text{ g } C}{1 \text{ mol } C} = 0.5716 \text{ g } C \]

\[ 0.286 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.0153 \text{ g } H_2O} \times \frac{2 \text{ mol } H}{1 \text{ mol } H_2O} = 0.03175 \text{ mol } H \times \frac{1.00794 \text{ g } H}{1 \text{ mol } H} = 0.0320 \text{ g } H \]

\[ \text{moles of chlorine } = \frac{\text{mol } C}{2} \times \frac{0.04759}{2} = 0.02380 \text{ mol } Cl \]

\[ \text{mass of } Cl = 0.02380 \text{ mol } Cl \times \frac{35.4527 \text{ g } Cl}{1 \text{ mol } Cl} = 0.8436 \text{ g } Cl \]

\[ \text{mass of oxygen obtained by difference: } 1.510 \text{ g } - 0.8436 \text{ g } - 0.5716 \text{ g } - 0.0320 \text{ g } = 0.063 \text{ g } O \]

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moles of oxygen = 0.063 g O × \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 0.00394 \text{ mol O}

Divide the number of moles of each element by 0.00394 to give an empirical formula of C_{12.1} H_{8.06} Cl_{6.04} O_{1.00} owing to the fact that the oxygen mass is obtained by difference, and it has only two significant digits and thus a higher degree of uncertainty.

The empirical formula is C_{12} H_{8} Cl_{6} O, which with a molecular mass of 381 u has the same molecular mass as the molecular formula. Hence, this empirical formula is also the molecular formula.

110. Calculate the mass of chlorine: 0.244 L × 2.898 g/L = 0.707 g chlorine
    Calculate the mass of iodine: 1.553 g - 0.707 g = 0.846 g iodine
    Calculate the moles of chlorine: 0.707 g/35.45 g/mol = 0.0199 mol chlorine
    Calculate the moles of iodine: 0.846 g/126.90 g/mol = 0.00667 mol iodine
    Calculate the mole ratio: 0.0199:0.00667 = 1:2.98 ≈ 1:3
    Calculate the empirical molar mass: (126.90 + 3 × 35.45) g/mol = 233.25 g/mol
    Because 467/233.25 ≈ 2, the molecular formula is I_{2}Cl_{6}.

FEATURE PROBLEMS

111. (a) “5-10-5” fertilizer contains 5.00 g N (that is, 5.00% N), 10.00 g P_{2}O_{5}, and 5.00 g K_{2}O in 100.00 g fertilizer. We convert the last two numbers into masses of the two elements.

(1) \% P = \frac{10.00 \text{ mol P}_{2}O_{5} \times 2 \text{ mol P}}{141.9 \text{ g P}_{2}O_{5} \times 1 \text{ mol P}_{2}O_{5} \times 30.97 \text{ g P}} = 100\% \times \text{ Ca (H}_{2}\text{PO}_{4})_{2} \times \text{ 30.97 g P} \times \frac{1 \text{ mol P}}{1 \text{ mol P}} = 4.37\% P

(2) \% K = \frac{5.00 \text{ mol K}_{2}O \times 2 \text{ mol K}}{94.20 \text{ g K}_{2}O \times 1 \text{ mol K}_{2}O \times 39.10 \text{ g K}} = 100\% \times \text{ NH}_{4}\text{HPO}_{4} \times \text{ 39.10 g K} \times \frac{1 \text{ mol K}}{1 \text{ mol K}} = 4.15\% K

(b) First, we determine \% P and then convert it to \% P_{2}O_{5}, given that 10.0% P_{2}O_{5} is equivalent to 4.37% P.

(1) \% P_{2}O_{5} = \frac{2 \text{ mol P}}{1 \text{ mol \text{ Ca (H}_{2}\text{PO}_{4})_{2}} \times 1 \text{ mol P}} \times \frac{30.97 \text{ g P}}{234.05 \text{ g \text{ Ca (H}_{2}\text{PO}_{4})_{2}}} \times \frac{1 \text{ mol \text{ Ca (H}_{2}\text{PO}_{4})_{2}} {1 \text{ mol \text{ Ca (H}_{2}\text{PO}_{4})_{2}}} \times 100\% \times \frac{10.0\% P_{2}O_{5}}{4.37\% P} = 60.6\% P_{2}O_{5}

(2) \% P_{2}O_{5} = \frac{1 \text{ mol P}}{1 \text{ mol \text{ NH}_{4}\text{HPO}_{4}} \times 1 \text{ mol P}} \times \frac{30.97 \text{ g P}}{132.06 \text{ g \text{ NH}_{4}\text{HPO}_{4}}} \times \frac{1 \text{ mol \text{ NH}_{4}\text{HPO}_{4}} {1 \text{ mol \text{ NH}_{4}\text{HPO}_{4}}} \times 100\%
(c) If the mass ratio of (NH₄)₂HPO₄ to KCl is set at 5.00:1.00, then for every 5.00 g of (NH₄)₂HPO₄ in the mixture there must be 1.00 g of KCl. Let’s start by finding the %N, %P, and %K for the fertilizer mixture.

\[
\%N(\text{by mass}) = \frac{2 \text{ mol N}}{1 \text{ mol (NH}_4\text{)}_2\text{HPO}_4} \times \frac{1 \text{ mol (NH}_4\text{)}_2\text{HPO}_4}{132.06 \text{ g (NH}_4\text{)}_2\text{HPO}_4} \times \frac{14.0067 \text{ g N}}{1 \text{ mol N}} \times \frac{5.00 \text{ g (NH}_4\text{)}_2\text{HPO}_4}{6.00 \text{ g mixture}} \times 100\%
\]

\[= 17.7\% \text{ N}\]

\[
\%P(\text{by mass}) = \frac{1 \text{ mol P}}{1 \text{ mol (NH}_4\text{)}_2\text{HPO}_4} \times \frac{1 \text{ mol (NH}_4\text{)}_2\text{HPO}_4}{132.06 \text{ g (NH}_4\text{)}_2\text{HPO}_4} \times \frac{30.9738 \text{ g P}}{1 \text{ mol P}} \times \frac{5.00 \text{ g (NH}_4\text{)}_2\text{HPO}_4}{6.00 \text{ g mixture}} \times 100\%
\]

\[= 19.5\% \text{ P}\]

\[
\%K(\text{by mass}) = \frac{1 \text{ mol K}}{1 \text{ mol KCl}} \times \frac{1 \text{ mol KCl}}{74.55 \text{ g KCl}} \times \frac{39.0983 \text{ g K}}{1 \text{ mol K}} \times \frac{1.00 \text{ g KCl}}{6.00 \text{ g mixture}} \times 100\%
\]

\[= 8.74\% \text{ K}\]

Next, we convert %P to P₂O₅ and %K to K₂O.

\[
\%P_2O_5 = 19.5\% \text{ P} \times \frac{10.0\% P_2O_5}{4.37\% P} = 44.6\% P_2O_5
\]

\[
\%K_2O = 8.74\% K \times \frac{500\% K_2O}{41.5\% K} = 10.5\% K_2O
\]

Thus, the combination of 5.00 g (NH₄)₂HPO₄ with 1.00 g KCl affords a “17.7-44.6-10.5” fertilizer, that is, 17.7% N, a percentage of phosphorus expressed as 44.6% P₂O₅, and a percentage of potassium expressed as 10.5% K₂O.

(d) A “5-10-5” fertilizer must possess the mass ratio 5.00 g N: 4.37 g P: 4.15 g K per 100 g of fertilizer. Thus a “5-10-5” fertilizer requires an N:P relative mass ratio of 5.00 g N:4.37 g P = 1.00 g N:0.874 g P. Note specifically that the fertilizer has a somewhat greater mass of N than of P.

If all of the N and P in the fertilizer comes solely from (NH₄)₂HPO₄, then the atom ratio of N relative to P will remain fixed at 2 N:1 P. Whether or not an inert non-fertilizing filler is present in the mix is immaterial. The relative N:P mass ratio is (2 × 14.01) g N:30.97 g P, that is, 0.905 g N:1.00 g P. Note specifically that (NH₄)₂HPO₄ has a somewhat lesser mass of N than of P. Clearly, it is impossible to make a “5-10-5” fertilizer if the only fertilizing components are (NH₄)₂HPO₄ and KCl.
113.

(a) The formula for stearic acid, obtained from the molecular model, is \( \text{CH}_3(\text{CH}_2)_{16}\text{CO}_2\text{H} \). The number of moles of stearic acid in 10.0 grams is

\[
10.0 \text{ g stearic acid} \times \frac{1 \text{ mol stearic acid}}{284.48 \text{ g stearic acid}} = 3.515 \times 10^{-2} \text{ mol of stearic acid.}
\]

The layer of stearic acid is one molecule thick. According to the figure provided with the question, each stearic acid molecule has a cross-sectional area of \(-0.22 \text{ nm}^2\).

In order to find the stearic acid coverage in square meters, we must multiply the total number of stearic acid molecules by the cross-sectional area for an individual stearic acid molecule. The number of stearic acid molecules is:

\[
= 3.515 \times 10^{-2} \text{ mol of stearic acid} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol of stearic acid}} = 2.117 \times 10^{22} \text{ molecules of stearic acid}
\]

area in m² = \( \frac{2.117 \times 10^{22} \text{ molecules of stearic acid} \times 0.22 \text{ nm}^2}{\text{molecule}} \times \frac{(1 \text{ m})^2}{(1 \times 10^9 \text{ nm})^2} \)

The area in m² = 4657 m² or \( 4.7 \times 10^3 \text{ m}^2 \) (with correct number of sig. fig.)

(b) The density for stearic acid is 0.85 g cm⁻³. Thus, 0.85 grams of stearic acid occupies 1 cm³. Find the number of moles of stearic acid in 0.85 g of stearic acid

\[
= 0.85 \text{ grams of stearic acid} \times \frac{1 \text{ mol stearic acid}}{284.48 \text{ g stearic acid}} = 3.0 \times 10^{-3} \text{ mol of stearic acid.}
\]

This number of moles of acid occupies 1 cm³ of space. So, the number of stearic acid molecules in 1 cm³

\[
= 3.0 \times 10^{-3} \text{ mol of stearic acid} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol of stearic acid}} = 1.8 \times 10^{21} \text{ stearic acid molecules.}
\]

Thus, the volume for a single stearic acid molecule in nm³

\[
= 1 \text{ cm}^3 \times \frac{1 \text{ mole of stearic acid}}{1.8 \times 10^{21} \text{ molecules stearic acid}} \times \frac{(1.0 \times 10^7 \text{ nm})^3}{(1 \text{ cm})^3} = 0.556 \text{ nm}^3
\]

The volume of a rectangular column is simply the area of its base multiplied by its height (i.e., \( V = \text{area of base (in nm}^2\) \times \text{height (in nm)})

So, the average height of a stearic acid molecule = \( \frac{0.556 \text{ nm}^3}{0.22 \text{ nm}^2} = 2.5 \text{ nm} \)
(c) The density for oleic acid = 0.895 g mL\(^{-1}\). So, the concentration for oleic acid is
\[
\frac{0.895 \text{ g acid}}{10.00 \text{ mL}} = 0.0895 \text{ g mL}^{-1} \text{ (solution 1)}
\]
This solution is then divided by 10, three more times, to give a final concentration of \(8.95 \times 10^{-5}\) g mL\(^{-1}\). A 0.10 mL sample of this solution contains:
\[
\frac{8.95 \times 10^{-5} \text{ g acid}}{1.00 \text{ mL}} \times 0.10 \text{ mL} = 8.95 \times 10^{-6} \text{ g of acid.}
\]
The number of acid molecules = \(85 \text{ cm}^2 \times \frac{1}{4.6 \times 10^{-15} \text{ cm}^2 \text{ per molecule}} = 1.85 \times 10^{16}\) oleic acid molecules.
So, \(8.95 \times 10^{-6}\) g of oleic acid corresponds to \(1.85 \times 10^{16}\) oleic acid molecules.
The molar mass for oleic acid, \(\text{C}_{18}\text{H}_{34}\text{O}_2\), is 282.47 g mol\(^{-1}\).
The number of moles of oleic acid is
\[
\frac{8.95 \times 10^{-6} \text{ g of acid}}{282.47 \text{ g}} = 3.17 \times 10^{-8} \text{ mol}
\]
So, Avogadro’s number here would be equal to:
\[
\frac{1.85 \times 10^{16} \text{ oleic acid molecules}}{3.17 \times 10^{-8} \text{ oleic acid moles}} = 5.8 \times 10^{23}\) molecules per mole of oleic acid.

**SELF ASSESSMENT EXERCISES**

118. The answer is (c), because 12.01 g of \(\text{H}_2\text{O}\) = 0.667 mol \(\text{H}_2\text{O}\), which equates to 0.667\(\times\)3 = 2.00 moles of atoms. One mole of \(\text{Br}_2\) also has 2.00 moles of atoms.

119. The answer is (b). \(\text{N}_2\text{H}_4\) can be reduced further to an empirical formula of \(\text{NH}_2\).

120. The answer is (d), because total atomic mass is 14 for N and 7 for H.

121. Answer is (a).
(a) 50.0 g \(\text{N}_2\text{O}\) \(\times\) (1 mol \(\text{N}_2\text{O}\)/44.0 g \(\text{N}_2\text{O}\)) \(\times\) (2 mol N/1 mol \(\text{N}_2\text{O}\)) = 2.27 mol
(b) 17.0 g \(\text{NH}_3\) \(\times\) (1 mol \(\text{NH}_3\)/17.0 g \(\text{NH}_3\)) \(\times\) (1 mol N/1 mol \(\text{NH}_3\)) = 1.00 mol
**Chapter 3: Chemical Compounds**

(c) \[150 \text{ mL } \text{C}_5\text{H}_5\text{N} \times \left(0.983 \text{ g/mL}\right) \times \left(1 \text{ mol Pyr/79.0 g Pyr}\right) \times \left(1 \text{ mol N/1 mol Pyr}\right) = 1.87 \text{ mol}\]

(d) \[1 \text{ mol N}_2 \times \left(2 \text{ mol N/1 mol N}_2\right) = 2.0\]

**122.** Answer is (c).
Mass \% of F = \[\frac{19 \times 3}{X + 19 \times 3} = 0.65\]
Solving for X, we get X = 30.7 or 31 u

**123.** Answer is (c). Total formal charge on H: +4. Total charge on O: -12, and the ion has a negative charge. Therefore, oxidation state of I = -12 + 4 + 1 = 7.

**124.** The answer is (b). Ca is a +2 ion and ClO\(^{-}\) is -1 anion.

**125.** The answer is (d). Multiplying O atomic mass by 4 (64 u) is nearly the same as the atomic mass of Cu (63.55).

**126.** The answer is (d). Answer (a) isn’t correct. While having the correct number of atoms, it is not an isomer because it is only a molecular formula and gives no information on atom bonding. Answer (b) isn’t correct, because it’s the exact same molecule as stated in the question. Answer (c) isn’t correct because it doesn’t have enough atoms. Therefore, the answer is (d), because it has the correct number of atoms in a different configuration.

**127.** First, find out the mass of Na\(_2\)SO\(_3\), which is 126.0 g/mol. Then:
Mass H\(_2\)O (x) = 0.5 (x + Mass Na\(_2\)SO\(_3\)).
x = 0.5x + 63.
Solving for x, we obtain x = 126 g (mass of H\(_2\)O)
Since we have 126 g of water, the number of moles of H\(_2\)O is 126 g / 18.0 g mol\(^{-1}\) = 7
Therefore, the formula is Na\(_2\)SO\(_3\)·7 H\(_2\)O.

**128.**
(a) Based on this composition, molar mass of malachite is calculated to be 221.18 g/mol. Since there are two moles of Cu per mole of malachite, the %mass of Cu is:
\[\frac{1000 \text{ g Malachite} \times \left(1 \text{ mol Mal.}/221.18 \text{ g Mal.}\right) \times \left(2 \text{ mol Cu}/1 \text{ mol Mal.}\right) \times \frac{63.546 \text{ g Cu}}{1 \text{ mol Cu}}}{1000 \text{ g}} = 574.61 \text{ g Cu}\]
% Cu = \[\frac{574.61 \text{ g}}{1000 \text{ g}} \times 100 = 57.46\%\]

(b) The formula for copper (II) oxide is CuO. Therefore, for one mole of malachite, there are two moles of CuO. Therefore,
\[\frac{1000 \text{ g Malachite} \times \left(1 \text{ mol Mal.}/221.18 \text{ g Mal.}\right) \times \left(2 \text{ mol CuO}/1 \text{ mol Mal.}\right)}{1 \text{ mol CuO}} = 719.5 \text{ g}\]

**129.** Molar mass of acetaminophen is 151.2 u, or 151.2 g/mol. To determine the molecular formula, calculate the moles of various constituting elements, as shown below:
mol C = 63.56 g C × (1 mol C/12.01 g C) = 5.3
mol H = 6.00 g H × (1 mol H/1.01 g H) = 5.94
mol N = 9.27 g N × (1 mol N/14.0 g N) = 0.66
mol O = 21.17 g O × (1 mol O/16.0 g O) = 1.32

Then, divide all values by the smallest to determine mole ratios:
5.3 mol C / 0.66 mol N = 8.0
6.00 mol H / 0.66 mol N = 9.0
0.66 mol N / 0.66 mol N = 1
1.32 mol O / 0.66 mol N = 2

The C:H:N:O ratio is 8:9:1:2. The empirical formula is therefore C₈H₉NO₂. The molar mass of this formula unit is 151.1, which is the same as the molar mass of acetaminophen. Therefore, the empirical formula obtained is also the same as the molecular formula.

130. The first step is to determine the mass of C, H, and O.

mol C = 6.029 g CO₂ × \( \frac{1 \text{ mol CO}_2}{44.0 \text{ g CO}_2} \) × \( \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \) = 0.137 mol

mass of C = 0.137 mol C × (12.01 g C/1 mol C) = 1.645 g C

mol H = 1.709 g H₂O × \( \frac{1 \text{ mol H}_2O}{18.0 \text{ g H}_2O} \) × \( \frac{2 \text{ mol H}}{1 \text{ mol H}_2O} \) = 0.190 mol

mass of H = 0.190 mol H × (1.01 g H/1 mol H) = 0.190 g H

Mass of oxygen is obtained by difference: mass of O = 2.174 g – (1.645 + 0.190) = 0.339 g

mol O = 0.339 g O₂ × \( \frac{1 \text{ mol O}_2}{16.0 \text{ g O}_2} \) = 0.0212 mol

(a) % Composition:
1.645 g C / 2.174 g Ibo = 75.67% C
0.190 g H / 2.174 g Ibo = 8.795% H
0.339 g O / 2.174 g Ibo = 15.45% O

(b) To determine the empirical formula, divide all mole values by the lowest one:
0.137 mol C / 0.0212 mol O = 6.5
0.190 mol H / 0.0212 mol O = 9.0
0.0212 mol O / 0.0212 mol O = 1.0

The empirical formula is obtained by multiplying the above ratios by 2. The formula is C₁₃H₁₈O₂.