Chapter 3: Chemical Compounds

Teaching for Conceptual Understanding

As recommended in the introduction for Chapter 2 on elements, show students macroscopic samples of a variety of compounds together with a model, particulate level diagram, and symbol for each. Include both ionic and covalent compounds and be sure that all three states of matter are represented.

Students sometimes misinterpret the subscripts in chemical formulas. Examples of two common errors are 1) \( \text{K}_2\text{S} \) consists of 1 potassium and 2 sulfur instead of the reverse and 2) \( \text{Mg(OH)}_2 \) consists of 1 magnesium, 1 oxygen, and 2 hydrogen instead of 2 oxygen and 2 hydrogen. It is important to test for these errors because they will lead to more mistakes when calculating molar masses, writing balanced chemical equations, and doing stoichiometric calculations.

Although we want our students to think conceptually and not rely on algorithms for problem solving, some algorithms, such as the one for naming compounds, are worth teaching our students. Two simple questions (and hints) students should ask themselves are: 1) Is there a metal in the formula? (If not, prefixes will be used in the name.) and 2) Does the metal form more than one cation? (If yes, Roman numerals in parentheses will be used in the name.)

A common misconception students have about ions, is that a positive ion has gained electrons and a negative ion has lost electrons (see Questions for Review and Thought #118). Use a tally of the protons, neutrons and electrons (Figure 3.1) in an atom and the ion it forms to show students the basis of the charge of an ion.

The dissociation of an ionic compound into its respective ions when dissolved in water is another troublesome area. Research on student particulate level diagrams has revealed misconceptions such as these: 1) \( \text{NiCl}_2 \) dissociates into \( \text{Ni}^{2+} \) and \( \text{Cl}^- \). 2) \( \text{NaOH} \) dissociates into \( \text{Na}^+ \), \( \text{O}^{2-} \), and \( \text{H}^+ \). 3) \( \text{Ni(OH)}_2 \) dissociates into \( \text{Ni}^{2+} \) and \( (\text{OH})_2^{2-} \). Having students draw particulate level diagrams is an excellent way of testing their understanding of dissociation. Questions for Review and Thought #111, 112 and 113 address this issue.

Some students completely ignore the charges when they look at formulas of ions; hence they see no difference between molecules and the ions with the same number and type of atoms, e.g., \( \text{SO}_3 \) and \( \text{SO}_3^{2-} \) or \( \text{NO}_2 \) and \( \text{NO}_2^- \). It is important to point out that an ion’s formula is incomplete unless the proper charge is given, and that a substance is a compound if there is no charge specified.

Another means of assessing student understanding is to give them incorrect examples of a concept, term, or problem solution and have them explain why it is incorrect. For some examples see Questions for Review and Thought #114 - 116.

Suggestions for Effective Learning

Many instructors skip organic and biochemistry topics in an introductory chemistry course because they think students will take organic or biochemistry courses. The reality is that the majority of the students will not. Some of these students will be exposed to organic and biochemistry courses dealing with living systems (human, animal or plant). The rest will leave college with a very limited view of chemistry. Do not skip the organic and biochemistry topics; they will add breadth to your course and spark student interest.

Not all college students are abstract thinkers; many are still at the concrete level when it comes to learning chemistry. The confusion surrounding the writing and understanding ionic chemical formulas can be eliminated.
by the use of simple jig saw puzzle pieces. Below are templates for cation and anion cutouts. The physical manipulation or visualization of how these pieces fit together is enough for most students to grasp the concept.

Show them that the ions must fit together with all the notches paired. For example, magnesium chloride, is an example of a compound with a 2+ cation and two 1– anions.

The pieces fit together as shown here:

A tip for writing correct chemical formulas of ionic compounds is that the magnitude of charge on the cation is the subscript for the anion and vice versa. Consider the formula of the ionic compound made from \( \text{Al}^{3+} \) and \( \text{O}^{2–} \):

\[ \text{Al}^{3+} \quad \text{O}^{2–} \quad \text{results in the formula} \quad \text{Al}_2\text{O}_3. \]

Caution students that this method can lead to incorrect formulas as in the case of \( \text{Mg}^{2+} \) and \( \text{O}^{2–} \) resulting in \( \text{Mg}_2\text{O}_2 \) instead of the correct formula of \( \text{MgO} \).

In addition to showing representative samples of the compounds, demonstrate physical properties, e.g., conductivity of molten ionic compounds (Figure 3.5), electrolytes and nonelectrolytes (Figure 3.6), and chemical changes, e.g., vinegar and baking soda, electrolysis of water. Both will prepare students for chemical reactions in Chapter 4.

**Cooperative Learning Activities**
• Have students complete a matrix of names and/or formulas of compounds formed by specified cation and anion. This exercise can be used a drill-and-practice or an assessment of student knowledge prior to or after instruction. Use only those cations and anions most relevant to your course.

<table>
<thead>
<tr>
<th></th>
<th>Cl(^{-})</th>
<th>O(^{2-})</th>
<th>NO(_{3}^{-})</th>
<th>PO(_{4}^{3-})</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na(^{+})</td>
<td>NaCl</td>
<td>sodium chloride</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Fe(^{2+})</td>
<td></td>
<td></td>
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<tr>
<td>Fe(^{3+})</td>
<td></td>
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<td></td>
</tr>
<tr>
<td>Al(^{3+})</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

• Conceptual Challenge Problems: CP-3.A, B, and C.

• Concept map terms: anion, cation, crystal lattice, electrolytes, formula unit, ionic compounds, ions, molar mass, molecular compounds, molecular formula, nonelectrolytes, structural formula.
**Solutions to Chapter 3**

**Questions for Review and Thought**

**Review Questions**

1. The “parts” that make up a chemical compound are atoms. Three pure (or nearly pure) compounds people often encounter are: water (H₂O), table sugar (sucrose, C₁₂H₂₂O₁₁), and salt (NaCl). A compound is different from a mixture because it has specific properties; the elements are present in definite proportions and can only be separated by chemical means. Mixtures can have variable properties and proportions, and the components of a mixture can be separated by physical means.

2. (a) Molecular formula C₄H₈
   Condensed formula: CH₃CH=CH₂

   (b) Molecular formula C₃H₆O₂
   Condensed formula: CH₃COOH

   (c) Molecular formula C₃H₇NO₂
   Condensed formula: NH₂CH₂CH₂COOH

3. (a) Structural

   ![Structural formula 1](image1)

   (b) Structural

   ![Structural formula 2](image2)

   (c) Structural

   ![Structural formula 3](image3)

   Molecular: CH₄O
   C₂H₇N
   C₄H₁₀S

4. (a) Molecular formula of pyruvic acid is C₃H₄O₃.

   (b) Molecular formula of isocitric acid is C₆H₈O₇.

5. (a) Molecular formula of valine is C₅H₁₁NO₂.

   (b) Molecular formula of 4-methyl-2-hexanol is C₇H₁₆O.

6. (a) NF₃ is nitrogen trifluoride.

   (b) HI is hydrogen iodide.

   (c) BBr₃ is boron tribromide.

   (d) C₆H₁₄ is hexane.

7. (a) C₈H₁₈ is octane.

   (b) P₂S₃ is diphosphorus trisulfide.

   (c) OF₂ is oxygen difluoride.

   (d) XeF₄ is xenon tetrafluoride.

8. (a) Sulfur trioxide is SO₃.

   (b) Dinitrogen pentoxide N₂O₅.

   (c) Phosphorus pentachloride is PCl₅.

   (d) Silicon tetrachloride is SiCl₄.

   (e) Diborox trioxide is B₂O₃.

**Molecular and Structural Formulas**
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9. (a) Bromine trifluoride is BrF₃. 
   (b) Xenon difluoride XeF₂. 
   (c) Diphosphorus tetrafluoride is P₂F₄. 
   (d) Pentadecane is C₁₅H₃₂. 
   (e) Hydrazine is N₂H₄.

10. (a) Butane (4 carbon alkane)

(b) Nonane (9 carbon alkane)

(c) Hexane (6 carbon alkane)

(d) Octane (8 carbon alkane)

(e) Octadecane (18 carbon alkane)

11.

<table>
<thead>
<tr>
<th>Molecular Formula</th>
<th>Condensed Formula</th>
<th>Structural Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>butanol C₄H₁₀O</td>
<td>CH₃CH₂CH₂CH₂OH</td>
<td>H-C-C-C-OH</td>
</tr>
<tr>
<td>pentanol C₅H₁₂O</td>
<td>CH₃CH₂CH₂CH₂CH₂OH</td>
<td>H-C-C-C-OH</td>
</tr>
</tbody>
</table>

12. Define the problem: Given the volume of gasoline, the density of gasoline, and the relationship between
gallons and milliliters, determine the mass of gasoline in grams and in pounds.

**Develop a plan:** Convert gallons into cubic centimeters. Use the density as a conversion factor to convert cubic centimeters into grams. Convert grams into pounds.

**Execute the plan:**

\[
18 \text{ gal} \times \frac{3790 \text{ mL}}{1 \text{ gal}} \times \frac{1 \text{ cm}^3}{1 \text{ mL}} \times \frac{0.692 \text{ g}}{1 \text{ cm}^3} = 4.7 \times 10^4 \text{ g}
\]

\[
4.7 \times 10^4 \text{ g} \times \frac{1 \text{ lb}}{453.59 \text{ g}} = 1.0 \times 10^2 \text{ lb}
\]

**Check your answers:** Each conversion factor comes from a valid relationship. The units cancel except for the desired ones. A “gram” is smaller than a pound, so the number of grams should be more than the number of pounds. These answers look okay.

13. Sucrose, C\textsubscript{12}H\textsubscript{22}O\textsubscript{11}, has eleven oxygen atoms per molecule. Glutathione, C\textsubscript{10}H\textsubscript{17}N\textsubscript{3}O\textsubscript{6}S, has only six oxygen atoms per molecule. Therefore sucrose has more oxygen atoms. Sucrose has (12 + 22 + 11 =) 45 atoms, total. Glutathione has (10 + 17 + 3 + 6 + 1 =) 37 atoms, total. Therefore sucrose has more atoms of all kinds.

14. (a) Benzene Molecular Formula: C\textsubscript{6}H\textsubscript{6}
(b) Vitamin C Molecular Formula: C\textsubscript{6}H\textsubscript{8}O\textsubscript{6}

15. (a) Heptane Molecular Formula: C\textsubscript{7}H\textsubscript{16}
(b) Acrylonitrile Molecular Formula: C\textsubscript{3}H\textsubscript{3}N
(c) Fenclorac Molecular Formula: C\textsubscript{14}H\textsubscript{16}Cl\textsubscript{2}O\textsubscript{2}

16. **Note:** Atoms in a formula found inside parentheses that are followed by a subscript get multiplied by that subscript.

(a) CaC\textsubscript{2}O\textsubscript{4} contains one atom of calcium, two atoms of carbon, and four atoms of oxygen.
(b) C\textsubscript{6}H\textsubscript{5}CHCH\textsubscript{2} contains eight atoms of carbon and eight atoms of hydrogen.
(c) (NH\textsubscript{4})\textsubscript{2}SO\textsubscript{4} contains two (1 \times 2) atoms of nitrogen, eight (4 \times 2) atoms of hydrogen, one atom of sulfur, and four atoms of oxygen.
(d) Pt(NH\textsubscript{3})\textsubscript{2}Cl\textsubscript{2} contains one atom of platinum, two (1 \times 2) atoms of nitrogen, six (3 \times 2) atoms of hydrogen, and two atoms of chlorine.
(e) K\textsubscript{4}Fe(CN)\textsubscript{6} contains four atoms of potassium, one atom of iron, six (1 \times 6) atoms of carbon, and six (1 \times 6) atoms of nitrogen.

17. **Note:** Atoms in a formula found inside parentheses that are followed by a subscript get multiplied by that subscript.

(a) C\textsubscript{4}H\textsubscript{3}COOC\textsubscript{2}H\textsubscript{5} contains nine atoms of carbon, ten atoms of hydrogen, two atoms of oxygen.
(b) HOOCCH\textsubscript{2}CH\textsubscript{2}COOH contains four atoms of carbon, four atoms of oxygen, and six atoms of hydrogen.
(c) NH₂CH₂CH₃COOH contains one atom of nitrogen, seven \((2 + 2 + 2 + 1)\) atoms of hydrogen, three \((1 + 1 + 1)\) atoms of carbon, and two atoms of oxygen.

(d) C₁₀H₉NH₂Fe contains ten atoms of carbon, eleven \((9 + 2)\) atoms of hydrogen, one atom of nitrogen, and one atom of iron.

(e) C₆H₂CH₃(NO₂)₃ contains seven atoms of carbon, five atoms of hydrogen, three \((1 \times 3)\) atoms of nitrogen, six \((2 \times 3)\) atoms of oxygen.

**Constitutional Isomers**

18. (a) Two molecules that are constitutional isomers have the same formula (i.e., on the molecular level, these molecules have the same number of atoms of each kind).

(b) Two molecules that are constitutional isomers of each other have their atoms in different bonding arrangements.

19. Five constitutional hexane isomers:

   (1) straight six-carbon chain: CH₃CH₂CH₂CH₂CH₂CH₃

   (2) five-carbon chain with branch on second carbon: CH₃CH(CH₃)CH₂CH₂CH₃

   (3) five-carbon chain with branch on third carbon: CH₃CH₂CH(CH₃)CH₂CH₃

   (4) four-carbon chain with two branches on second carbon: CH₃C(CH₃)₂CH₂CH₃

   (5) four-carbon chain with a branches on second carbon and a branch of third carbon: CH₃CH(CH₃)CH(CH₃)CH₃

**Predicting Ion Charges**

20. A general rule for the charge on a metal cation: the group number represents the number of electrons lost. Hence, the group number will be the cation’s positive charge.

   (a) Lithium (Group 1A) Li⁺

   (b) Strontium (Group 2A) Sr²⁺

   (c) Aluminum (Group 3A) Al³⁺

   (d) Calcium (Group 2A) Ca²⁺

   (e) Zinc (Group 2B) Zn²⁺

21. For nonmetal elements in Groups 5A–7A, the electrons gained by an atom to form a stable anion are calculated using the formula: \(8 - \text{(group number)}\). That means the \((\text{group number}) - 8\) is the negative charge of the anion.

   (a) nitrogen (Group 5A) \(5 - 8 = -3\) \(\text{N}^{3-}\)

   (b) sulfur (Group 6A) \(6 - 8 = -2\) \(\text{S}^{2-}\)

   (c) chlorine (Group 7A) \(7 - 8 = -1\) \(\text{Cl}^-\)
(d) iodine (Group 7A) \( 7 - 8 = -1 \) \( I^- \)
(e) phosphorus (Group 5A) \( 5 - 8 = -3 \) \( P^{3-} \)

22. A general rule for the charge on a metal cation: the group number represents the number of electrons lost. Hence, the group number will be the cation’s positive charge. For nonmetal elements in Groups 5A-7A, the electrons gained by an atom to form a stable anion are calculated using the formula: \( 8 - \) (group number). That means the (group number) – 8 is the negative charge of the anion.

Barium (Group 2A) has a +2 charge. Bromine (Group 7A) has a –1 charge. So, the ions are \( \text{Ba}^{2+} \) and \( \text{Br}^- \).

23. A general rule for the charge on a metal cation: the group number represents the number of electrons lost. Hence, the group number will be the cation’s positive charge. Transition metals often have a +2 charge. Some have +3 and +1 charged ions, as well.

(a) magnesium (Group 2A) has a +2 charge. \( \text{Mg}^{2+} \)
(b) zinc (Group 8B) has a +2 charge. \( \text{Zn}^{2+} \)
(c) iron (a transition metal) has a +2 or +3 charge. \( \text{Fe}^{2+} \) or \( \text{Fe}^{3+} \)
(d) gallium (Group 3A) has a +3 charge. \( \text{Ga}^{3+} \)

24. For nonmetal elements in Groups 5A-7A, the electrons gained by an atom to form a stable anion are calculated using the formula: \( 8 - \) (group number). That means the (group number) – 8 is the negative charge of the anion. Transition metals often have a +2 charge. Some have +3 and +1 charged ions, as well.

(a) selenium (Group 6A) \( 6 - 8 = -2 \) \( \text{Se}^{2-} \)
(b) fluorine (Group 7A) \( 7 - 8 = -1 \) \( \text{F}^- \)
(c) nickel (a transition metal) \( \text{Ni}^{2+} \)
(d) nitrogen (Group 5A) \( 5 - 8 = -3 \) \( \text{N}^{3-} \)

25. Cobalt ions are \( \text{Co}^{2+} \) and \( \text{Co}^{3+} \). Oxide ion is \( \text{O}^{2-} \). The two compounds containing cobalt and oxide are made from the neutral combination of the charged ions:

one \( \text{Co}^{2+} \) and one \( \text{O}^{2-} \) [net charge = +2 + (-2) = 0 ] \( \text{CoO} \)
two \( \text{Co}^{3+} \) and three \( \text{O}^{2-} \) [net charge = 2 × (+3) + 3 × (-2) = 0 ] \( \text{Co}_2\text{O}_3 \)

26. The two compounds containing lead and chloride are made from the neutral combination of the charged ions:

one \( \text{Pb}^{2+} \) and two \( \text{Cl}^- \) [net charge = +2 + 2 × (-1) = 0 ] \( \text{PbCl}_2 \)
one \( \text{Pb}^{4+} \) and four \( \text{Cl}^- \) [net charge = +4 + 4 × (-1) = 0 ] \( \text{PbCl}_4 \)

27. (a) Aluminum ion (from Group 3A) is \( \text{Al}^{3+} \). Chloride ion (from Group 7A) is \( \text{Cl}^- \).

\( \text{AlCl} \) is not a neutral combination of these two ions. The proper formula would be \( \text{AlCl}_3 \). [net charge = +3 + 3 × (-1) = 0 ]

(b) Sodium ion (Group 1A) is \( \text{Na}^+ \). Fluoride ion (from Group 7A) is \( \text{F}^- \).

\( \text{NaF}_2 \) is not a neutral combination of these two ions. The proper formula would be \( \text{NaF} \). [net charge = +1 + (-1) = 0 ]
(c) Gallium ion (from Group 3A) is Ga$^{3+}$. Oxide ion (from Group 6A) is O$^{2-}$.

Ga$_2$O$_3$ is the proper neutral combination of these two ions.

[net charge = $2 \times (+3) + 3 \times (-2) = 0$]

(d) Magnesium ion (from Group 2A) is Mg$^{2+}$. Sulfide ion (from Group 6A) is S$^{2-}$.

MgS is the proper neutral combination of these two ions.

[net charge = $+2 + (-2) = 0$]

28. (a) Calcium ion (from Group 2A) is Ca$^{2+}$. Oxide ion (from Group 6A) is O$^{2-}$.

Ca$_2$O is not a neutral combination of these two ions. The proper formula would be CaO. [net charge = $+2 + (-2) = 0$]

(b) Strontium ion (Group 2A) is Sr$^{2+}$. Chloride ion (from Group 7A) is Cl$^-$.  

SrCl$_2$ is the proper neutral combination of these two ions.

[net charge = $+2 + 2 \times (-1) = 0$]

(c) Iron ion (from the transition elements) is Fe$^{3+}$ or Fe$^{2+}$. Oxide ion (from Group 6A) is O$^{2-}$. Fe$_2$O$_5$ is not a neutral combination of these ions. The proper possible formulas would be FeO [net charge = $+2 + (-2) = 0$] or Fe$_2$O$_3$ [net charge = $2 \times (+3) + 3 \times (-2) = 0$]

(d) Potassium ion (from Group 1A) is K$^+$. Oxide ion (from Group 6A) is O$^{2-}$.

K$_2$O is the proper neutral combination of these two ions.

[net charge = $2 \times (+1) + (-2) = 0$]

**Polyatomic Ions**

29. (a) Pb(NO$_3$)$_2$ has one ion of lead(II) (Pb$^{2+}$) and two ions of nitrate (NO$_3^-$).

(b) NiCO$_3$ has one ion of nickel(II) (Ni$^{2+}$) and one ion of carbonate (CO$_3^{2-}$).

(c) (NH$_4$)$_3$PO$_4$ has three ions of ammonium (NH$_4^+$) and one ion of phosphate (PO$_4^{3-}$).

(d) K$_2$SO$_4$ has two ions of potassium (K$^+$) and one ion of sulfate (SO$_4^{2-}$).

30. (a) Ca(CH$_3$CO$_2$)$_2$ has one ion of calcium (Ca$^{2+}$) and two ions of acetate (CH$_3$CO$_2^-$ also written: CH$_3$COO$^-$).

(b) CO$_2$(SO$_4$)$_3$ has two ions of cobalt(III) (Co$^{3+}$) and three ions of sulfate (SO$_4^{2-}$).

(c) Al(OH)$_3$ has one ion of aluminum (Al$^{3+}$) and three ions of hydroxide (OH$^-$).

(d) (NH$_4$)$_2$CO$_3$ has two ions of ammonium (NH$_4^+$) and one ion of carbonate (CO$_3^{2-}$).

31. Barium sulfate is BaSO$_4$. It contains barium ion (Ba$^{2+}$) and sulfate ion (SO$_4^{2-}$).

Magnesium nitrate is Mg(NO$_3$)$_2$. It contains magnesium ion (Mg$^{2+}$) and nitrate ions (NO$_3^-$).
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Sodium acetate is NaCH$_3$CO$_2$ or NaCH$_3$COO. It contains sodium ion (Na$^+$) and acetate ion (CH$_3$COO$^-$). (Note: Occasionally the Na$^+$ is written on the other end of the acetate formula like this CH$_3$COONa. It is done that way because the negative charge on acetate is on one of the oxygen atoms, so that’s where the Na$^+$ cation will be attracted.)

32. Calcium nitrate is Ca(NO$_3$)$_2$, potassium chloride is KCl, and barium phosphate is Ba$_3$(PO$_4$)$_2$. The ions in calcium nitrate are Ca$^{2+}$, called calcium ion with a 2+ charge, and NO$_3^-$, called nitrate ion with a 1– charge. The ions in potassium chloride are K$^+$, called potassium ion with a 1+ charge, and Cl$^-$, called chloride ion with a 1– charge. The ions in barium phosphate are Ba$^{2+}$, called barium ion with a 2+ charge, and PO$_4^{3–}$, called phosphate ion with a 3– charge.

33. (a) Nickel(II) nitrate Ni(NO$_3$)$_2$  (b) sodium bicarbonate NaHCO$_3$
    (c) Lithium hypochlorite LiClO  (d) magnesium chlorate Mg(ClO$_3$)$_2$
    (e) Calcium sulfite CaSO$_3$

**Ionic Compounds**

34. To tell if a compound is ionic or not, look for metals and nonmetals together, or common cations and anions. If a compound contains only nonmetals or metalloids and nonmetals, it is probably not ionic.

   (a) CF$_4$ contains only nonmetals. Not ionic.
   (b) SrBr$_2$ has a metal and nonmetal together. Ionic.
   (c) Co(NO$_3$)$_3$ has a metal and nonmetals together. Ionic.
   (d) SiO$_2$ contains a metalloid and a nonmetal. Not ionic.
   (e) KCN has a metal and a common diatomic ion (CN$^–$) together. Ionic.
   (f) SCl$_2$ contains only nonmetals. Not ionic.

35. To tell if a compound is ionic or not, look for metals and nonmetals together, or common cations and anions. If a compound contains only nonmetals or metalloids and nonmetals, it is probably not ionic.

   (a) Methane, CH$_4$, contains only nonmetals. Not ionic.
   (b) Dinitrogen pentoxide, N$_2$O$_5$, contains only nonmetals. Not ionic.
   (c) Ammonium sulfide, (NH$_4$)$_2$S, has a common cation and anion together. Ionic.
   (d) Hydrogen selenide, H$_2$Se, contains a metalloid and a nonmetal. Not ionic.
   (e) Sodium perchlorate, NaClO$_4$, has a metal and a common anion together. Ionic.

36. (a) Ammonium carbonate, (NH$_4$)$_2$CO$_3$
   (b) Calcium iodide, CaI$_2$
   (c) Copper(II) bromide, CuBr$_2$
37. (a) Calcium hydrogen carbonate, Ca(HCO₃)₂
(b) Potassium permanganate, KMnO₄
(c) Magnesium perchlorate, Mg(ClO₄)₂
(d) Ammonium hydrogen phosphate, (NH₄)₂HPO₄

38. (a) K₂S is potassium sulfide.
(b) NiSO₄ is nickel(II) sulfate.
(c) (NH₄)₃PO₄ is ammonium phosphate.

39. (a) Ca(CH₃CO₂)₂ is calcium acetate.
(b) Co₂(SO₄)₃ is cobalt(III) sulfate.
(c) Al(OH)₃ is aluminum hydroxide.

40. (a) KH₂PO₄ is potassium dihydrogen phosphate.
(b) CuSO₄ is copper(II) sulfate.
(c) CrCl₆ is chromium (VI) chloride.

41. Magnesium oxide is MgO, and it is composed of Mg²⁺ ions and O²⁻ ions. The relatively high melting temperature of MgO compared to NaCl (composed of Na⁺ ions and Cl⁻ ions) is probably due to the higher ionic charges and smaller sizes of the ions. The large opposite charges sitting close together have very strong attractive forces between the ions. Melting requires that these attractive forces be overcome.

42. A white crystalline powder in a bottle that melts at 310 °C is probably the ionic compound, NaNO₃. NO₂ and NH₃ are covalent compounds and are in the gaseous state at room temperature.

**Electrolytes**

43. An electrolyte is a compound that conducts electricity when dissolved in water because it dissociates into ions. Electricity is conducted when ions are present in the solution. We can differentiate a strong electrolyte by checking whether the solution conducts electricity. When a strong electrolyte (such as NaCl) dissolves in water, it will ionize completely (producing Na⁺ and Cl⁻). When a nonelectrolyte (such as table sugar, sucrose, C₁₂H₂₂O₁₁) dissolves in water it does not ionizes (In the solution is found only the molecular form, C₁₂H₂₂O₁₁).

44. Epsom salt contains two common ions, Mg²⁺ and SO₄²⁻. It is probably a strong electrolyte. Methanol, CH₃OH, which consists of only nonmetals is probably not ionic, and therefore non-electrolytic.

45. “Molecular compounds are generally non-electrolytes.” This general trend is sensible, since the molecular compounds are generally not ionic compounds, and therefore would not ionize in water.

46. “Ionic compounds are generally electrolytes.” This general trend is sensible, since the ionic compounds are composed of ions that could separate when the compound is dissolved in water. Therefore they would be
electrolytes.

47. (a) The ions present in a solution of KOH are $\text{K}^+$ and $\text{OH}^-$.
(b) The ions present in a solution of $\text{K}_2\text{SO}_4$ are $\text{K}^+$ and $\text{SO}_4^{2-}$.
(c) The ions present in a solution of NaNO$_3$ are $\text{Na}^+$ and $\text{NO}_3^-$.
(d) The ions present in a solution of NH$_4$Cl are $\text{NH}_4^+$ and Cl$^-$.

48. (a) The ions present in a solution of CaI$_2$ are $\text{Ca}^{2+}$ and I$^-$.
(b) The ions present in a solution of Mg$_3$(PO$_4$)$_2$ are $\text{Mg}^{2+}$ and PO$_4^{3-}$.
(c) The ions present in a solution of NiS are Ni$^{2+}$ and S$^{2-}$.
(d) The ions present in a solution of MgBr$_2$ are Mg$^{2+}$ and Br$^-$.

49. (a) NaCl is an ionic compound and will ionize to form Na$^+$ and Cl$^-$. When NaCl is dissolved in water, the resulting solution will conduct electricity.
(b) CH$_3$CH$_2$CH$_3$ is an organic hydrocarbon compound and will not ionize. When CH$_3$CH$_2$CH$_3$ is dissolved in water, the resulting solution will not conduct electricity.
(c) CH$_3$OH is an organic compound and will not ionize. When CH$_3$OH is dissolved in water, the resulting solution will not conduct electricity.
(d) Ca(NO$_3$)$_2$ is an ionic compound and will ionize to form Ca$^{2+}$ and NO$_3^-$.
When Ca(NO$_3$)$_2$ is dissolved in water, the resulting solution will conduct electricity.

Moles of Compounds

50. Consider a sample of 1 mol of methanol (M).

<table>
<thead>
<tr>
<th>CH$_3$OH</th>
<th>Carbon</th>
<th>Hydrogen</th>
<th>Oxygen</th>
</tr>
</thead>
<tbody>
<tr>
<td>No. of moles</td>
<td>1 mol</td>
<td>1 mol</td>
<td>4 mol</td>
</tr>
<tr>
<td>No. of molecules or atoms</td>
<td>$6.022 \times 10^{23}$ molecules</td>
<td>$6.022 \times 10^{23}$ atoms</td>
<td>$2.409 \times 10^{24}$ atoms</td>
</tr>
<tr>
<td>Molar mass</td>
<td>32.042 g/mol M</td>
<td>12.011 g/mol M</td>
<td>4.0316 g/mol M</td>
</tr>
</tbody>
</table>

51. Consider a sample of 1 mol of glucose (G).

<table>
<thead>
<tr>
<th>C$<em>6$H$</em>{12}$O$_6$</th>
<th>Carbon</th>
<th>Hydrogen</th>
<th>Oxygen</th>
</tr>
</thead>
<tbody>
<tr>
<td>No. of moles</td>
<td>1 mol</td>
<td>6 mol</td>
<td>12 mol</td>
</tr>
<tr>
<td>No. of molecules or atoms</td>
<td>$6.022 \times 10^{23}$ molecules</td>
<td>$3.613 \times 10^{24}$ atoms</td>
<td>$7.226 \times 10^{24}$ atoms</td>
</tr>
<tr>
<td>Molar mass</td>
<td>180.157 g/mol G</td>
<td>72.066 g/mol G</td>
<td>12.095 g/mol G</td>
</tr>
</tbody>
</table>

52. To calculate the molar mass of a compound, we perform a series of steps: First, look up the atomic mass of
each element in the compound and identify the atomic mass as the molar mass of each element. Second, determine the number of moles of atoms in one mole of the compound. Third, multiply the number of moles of the element by the molar mass of the element. Last, add up all the individual mass contributions.

(a) 1 mol of $\text{Fe}_2\text{O}_3$ contains 2 mol of Fe and 3 mol of O.

$$\left( \frac{2 \text{ mol Fe}}{1 \text{ mol } \text{Fe}_2\text{O}_3} \times \frac{55.845 \text{ g}}{1 \text{ mol Fe}} \right) + \left( \frac{3 \text{ mol O}}{1 \text{ mol } \text{Fe}_2\text{O}_3} \times \frac{15.9994 \text{ g}}{1 \text{ mol O}} \right) = 159.688 \frac{\text{g}}{\text{mol } \text{Fe}_2\text{O}_3}$$

(b) 1 mol of $\text{BF}_3$ contains 1 mol of B and 3 mol of F.

$$\left( \frac{1 \text{ mol B}}{1 \text{ mol } \text{BF}_3} \times \frac{10.811 \text{ g}}{1 \text{ mol B}} \right) + \left( 3 \text{ mol F} \times \frac{18.9984 \text{ g}}{1 \text{ mol F}} \right) = 67.806 \frac{\text{g}}{\text{mol } \text{BF}_3}$$

(c) 1 mol of $\text{N}_2\text{O}$ contains 2 mol of N and 1 mol of O.

$$\left( \frac{2 \text{ mol N}}{1 \text{ mol } \text{N}_2\text{O}} \times \frac{14.0067 \text{ g}}{1 \text{ mol N}} \right) + \left( \frac{1 \text{ mol O}}{1 \text{ mol } \text{N}_2\text{O}} \times \frac{15.9994 \text{ g}}{1 \text{ mol O}} \right) = 44.0128 \frac{\text{g}}{\text{mol } \text{N}_2\text{O}}$$

(d) 1 mol of $\text{MnCl}_2\cdot4\text{H}_2\text{O}$ compound has 1 mol of MnCl$_2$ with 4 mol of H$_2$O molecules. So, it contains 1 mol Mn, 2 mol Cl, (4 × 2 =) 8 mol H and (4 × 1 =) 4 mol O.

$$\left( \frac{1 \text{ mol Mn}}{1 \text{ mol comp}} \times \frac{54.938 \text{ g}}{1 \text{ mol Mn}} \right) + \left( \frac{2 \text{ mol Cl}}{1 \text{ mol } \text{MnCl}_2} \times \frac{35.453 \text{ g}}{1 \text{ mol Cl}} \right) + \left( 8 \text{ mol H} \times \frac{1.0079 \text{ g}}{1 \text{ mol H}} \right) + \left( 4 \text{ mol O} \times \frac{15.9994 \text{ g}}{1 \text{ mol O}} \right) = 197.905 \frac{\text{g}}{\text{mol comp}}$$

(e) 1 mol of $\text{C}_6\text{H}_8\text{O}_6$ compound contains 6 mol of C, 8 mol of H, and 6 mol of O.

$$\left( \frac{6 \text{ mol C}}{1 \text{ mol } \text{C}_6\text{H}_8\text{O}_6} \times \frac{12.0107 \text{ g}}{1 \text{ mol C}} \right) + \left( \frac{8 \text{ mol H}}{1 \text{ mol } \text{C}_6\text{H}_8\text{O}_6} \times \frac{1.0079 \text{ g}}{1 \text{ mol H}} \right) + \left( \frac{6 \text{ mol O}}{1 \text{ mol } \text{C}_6\text{H}_8\text{O}_6} \times \frac{15.9994 \text{ g}}{1 \text{ mol O}} \right) = 176.1238 \frac{\text{g}}{\text{mol comp}}$$

53. To calculate the molar mass of a compound, we perform a series of steps: First, look up the atomic mass of each element in the compound and identify the atomic mass as the molar mass of each element. Second, determine the number of moles of atoms in one mole of the compound. Third, multiply the number of moles of the element by the molar mass of the element. Last, add the individual masses.

(a) 1 mol of $\text{B}_10\text{H}_{14}$ contains 10 mol of B and 14 mol of H.

$$\left( \frac{10 \text{ mol B}}{1 \text{ mol } \text{B}_10\text{H}_{14}} \times \frac{10.811 \text{ g}}{1 \text{ mol B}} \right) + \left( \frac{14 \text{ moles H}}{1 \text{ mol } \text{B}_10\text{H}_{14}} \times \frac{1.0079 \text{ g}}{1 \text{ mol H}} \right) = 122.221 \frac{\text{g}}{\text{mol } \text{B}_10\text{H}_{14}}$$

(b) 1 mol of $\text{C}_6\text{H}_2(\text{CH}_3)(\text{NO}_2)_3$ compound contains 7 mol of C, 5 mol of H, 3 mol of N, and 6 mol of O.

$$\left( \frac{7 \text{ mol C}}{1 \text{ mol } \text{C}_6\text{H}_2(\text{CH}_3)(\text{NO}_2)_3} \times \frac{12.0107 \text{ g}}{1 \text{ mol C}} \right) + \left( \frac{5 \text{ mol H}}{1 \text{ mol } \text{C}_6\text{H}_2(\text{CH}_3)(\text{NO}_2)_3} \times \frac{1.0079 \text{ g}}{1 \text{ mol H}} \right)$$
Chapter 3: Chemical Compounds

54. Define the problem: Determine the number of moles in a given mass of a compound.

Develop a plan: Adapt the method described in the answer to Question 51 to calculate the molar mass for the compound, then use the molar mass as a conversion factor between grams and moles.

Execute the plan:

(a) Molar mass CH$_3$OH = (12.0107 g) + 4 × (1.0079 g) + (15.9994 g) = 32.0417 g/mol

\[ 1.00 \text{ g CH}_3\text{OH} \times \frac{1 \text{ mol CH}_3\text{OH}}{32.0417 \text{ g CH}_3\text{OH}} = 0.0312 \text{ mol CH}_3\text{OH} \]

(b) Molar mass Cl$_2$CO = 2 × (35.453 g) + (12.0107 g) + (15.9994 g) = 98.916 g/mol

\[ 1.00 \text{ g Cl}_2\text{CO} \times \frac{1 \text{ mol Cl}_2\text{CO}}{98.916 \text{ g Cl}_2\text{CO}} = 0.0101 \text{ mol Cl}_2\text{CO} \]

(c) Molar mass NH$_4$NO$_3$ = 2 × (14.0067 g) + 4 × (1.0079 g) + 3 × (15.9994 g) = 80.043 g/mol

\[ 1.00 \text{ g NH}_4\text{NO}_3 \times \frac{1 \text{ mol NH}_4\text{NO}_3}{80.043 \text{ g NH}_4\text{NO}_3} = 0.0125 \text{ mol NH}_4\text{NO}_3 \]

(d) Molar mass MgSO$_4$·7H$_2$O

\[ = (24.305 \text{ g}) + (32.065 \text{ g}) + 11 \times (15.9994 \text{ g}) + 14 \times (1.0079 \text{ g}) = 246.474 \text{ g/mol} \]

\[ 1.00 \text{ g MgSO}_4\cdot7\text{H}_2\text{O} \times \frac{1 \text{ mol MgSO}_4\cdot7\text{H}_2\text{O}}{246.474 \text{ g MgSO}_4\cdot7\text{H}_2\text{O}} = 0.00406 \text{ mol MgSO}_4\cdot7\text{H}_2\text{O} \]
(e) Molar mass AgCH$_3$CO$_2$

\[ (107.8682 \text{ g}) + 2 \times (12.0107 \text{ g}) + 3 \times (1.0079 \text{ g}) + 2 \times (15.9994 \text{ g}) = 166.9121 \text{ g/mol} \]

\[ 1.00 \text{ g} \text{ AgCH}_3\text{CO}_2 \times \frac{1 \text{ mol} \text{ AgCH}_3\text{CO}_2}{166.9121 \text{ g} \text{ AgCH}_3\text{CO}_2} = 0.00599 \text{ mol} \text{ AgCH}_3\text{CO}_2 \]

Check your answers: The quantity in moles is always going to be smaller than the mass in grams. These numbers look right.

55. Define the problem: Determine the number of moles in a given mass of a compound.

Develop a plan: Adapt the method described in the answer to Question 52 to calculate the molar mass for the compound, then use the molar mass as a conversion factor between grams and moles.

Execute the plan:

(a) Molar mass C$_7$H$_5$NO$_3$S = 7 \times (12.0107 \text{ g}) + 5 \times (1.0079 \text{ g}) + (14.0067 \text{ g})
+ 3 \times (15.9994 \text{ g}) + (32.065 \text{ g}) = 183.184 \text{ g/mol}

\[ 0.250 \text{ g} \text{ C}_7\text{H}_5\text{NO}_3\text{S} \times \frac{1 \text{ mol} \text{ C}_7\text{H}_5\text{NO}_3\text{S}}{183.184 \text{ g} \text{ C}_7\text{H}_5\text{NO}_3\text{S}} = 0.00136 \text{ mol} \text{ C}_7\text{H}_5\text{NO}_3\text{S} \]

(b) Molar mass C$_{13}$H$_{20}$N$_2$O$_2$ = 13 \times (12.0107 \text{ g}) + 20 \times (1.0079 \text{ g}) + 2 \times (14.0067 \text{ g})
+ 2 \times (15.9994 \text{ g}) = 236.3093 \text{ g/mol}

\[ 0.250 \text{ g} \text{ C}_{13}\text{H}_{20}\text{N}_2\text{O}_2 \times \frac{1 \text{ mol} \text{ C}_{13}\text{H}_{20}\text{N}_2\text{O}_2}{236.3093 \text{ g} \text{ C}_{13}\text{H}_{20}\text{N}_2\text{O}_2} = 0.00106 \text{ mol} \text{ C}_{13}\text{H}_{20}\text{N}_2\text{O}_2 \]

(c) Molar mass C$_{20}$H$_{14}$O$_4$ = 20 \times (12.0107 \text{ g}) + 14 \times (1.0079 \text{ g}) + 4 \times (15.9994 \text{ g}) = 318.3222 \text{ g/mol}

\[ 0.250 \text{ g} \text{ C}_{20}\text{H}_{14}\text{O}_4 \times \frac{1 \text{ mol} \text{ C}_{20}\text{H}_{14}\text{O}_4}{318.3222 \text{ g} \text{ C}_{20}\text{H}_{14}\text{O}_4} = 7.85 \times 10^{-4} \text{ mol} \text{ C}_{20}\text{H}_{14}\text{O}_4 \]

Check your answers: The quantity in moles is always going to be smaller than the mass in grams. These numbers look right.

56. Define the problem: Determine the molar mass of a compound and then determine the mass of a given number of moles and the number of moles in a given mass.

Develop a plan: Adapt the method described in the answer to Question 51 to calculate the molar mass for the compound, then use the molar mass as a conversion factor between grams and moles.

Execute the plan:

(a) Molar mass Fe(NO$_3$)$_2$ = (55.845 \text{ g}) + 2 \times (14.0067 \text{ g}) + 6 \times (15.9994 \text{ g}) = 179.855 \text{ g/mol}

\[ 0.200 \text{ mol} \text{ Fe(NO}_3)_2 \times \frac{179.857 \text{ g} \text{ Fe(NO}_3)_2}{1 \text{ mol} \text{ Fe(NO}_3)_2} = 36.0 \text{ g} \text{ Fe(NO}_3)_2 \]

(b) 4.66 \text{ g} \text{ Fe(NO}_3)_2 \times \frac{1 \text{ mol} \text{ Fe(NO}_3)_2}{179.855 \text{ g} \text{ Fe(NO}_3)_2} = 0.0259 \text{ mol} \text{ Fe(NO}_3)_2}

Check your answers: The quantity in moles is always going to be smaller than the mass in grams. These
numbers look right.

57. Define the problem: Determine the molar mass of a compound and then determine the mass of a given number of moles and the number of moles in a given mass.

Develop a plan: Adapt the method described in the answer to Question 52 to calculate the molar mass for the compound, then use the molar mass as a conversion factor between grams and moles.

Execute the plan:

(a) Molar mass \( \text{C}_8\text{H}_9\text{O}_2\text{N} = 8 \times (12.0107 \text{ g}) + 9 \times (1.0079 \text{ g}) + 2 \times (15.9994 \text{ g}) + (14.0067 \text{ g}) = 151.1622 \text{ g/mol} \)

(b) \( 5.32 \text{ g C}_8\text{H}_9\text{O}_2\text{N} \times \frac{1 \text{ mol C}_8\text{H}_9\text{O}_2\text{N}}{151.1622 \text{ g C}_8\text{H}_9\text{O}_2\text{N}} = 0.0352 \text{ mol C}_8\text{H}_9\text{O}_2\text{N} \)

(c) \( 0.166 \text{ mol C}_8\text{H}_9\text{O}_2\text{N} \times \frac{151.1622 \text{ g C}_8\text{H}_9\text{O}_2\text{N}}{1 \text{ mol C}_8\text{H}_9\text{O}_2\text{N}} = 25.1 \text{ g C}_8\text{H}_9\text{O}_2\text{N} \)

Check your answers: The quantity in moles is always going to be smaller than the mass in grams. These numbers look right.

58. Define the problem: Given the mass of a substance, determine the number of moles.

Develop a plan: Adapt the method described in the answer to Question 52 to calculate the molar mass for the compound, then use the molar mass as a conversion factor between grams and moles.

(a) Molar mass \( \text{H}_2\text{SO}_4 = 2 \times (1.0079 \text{ g}) + (32.065 \text{ g}) + 4 \times (15.9994 \text{ g}) = 98.078 \text{ g/mol} \)

\( 39.2 \text{ g H}_2\text{SO}_4 \times \frac{1 \text{ mol H}_2\text{SO}_4}{98.078 \text{ g H}_2\text{SO}_4} = 0.400 \text{ mol H}_2\text{SO}_4 \)

(b) Molar mass \( \text{O}_2 = 2 \times (15.9994 \text{ g}) = 31.9988 \text{ g/mol} \)

\( 8.00 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{31.9988 \text{ g O}_2} = 0.250 \text{ mol O}_2 \)

(b) Molar mass \( \text{NH}_3 = 14.0067 \text{ g} + 3 \times (1.0079 \text{ g}) = 17.0304 \text{ g/mol} \)

\( 10.7 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.0304 \text{ g NH}_3} = 0.628 \text{ mol NH}_3 \)

Check your answers: The quantity in moles is always going to be smaller than the mass in grams. These numbers look right.

59. Define the problem: Given the masses of three compounds in a mixture, determine the number of moles of each, then determine the number of molecules of one compound.

Develop a plan: Adapt the method described in the answer to Question 51 to calculate the molar mass for the compounds. Convert milligrams to grams, then use the molar mass as a conversion factor between grams and moles. Use Avogadro’s number to determine the actual number of molecules.

Execute the plan:

(a) Molar mass \( \text{C}_9\text{H}_6\text{O}_4 = 9 \times (12.0107 \text{ g}) + 8 \times (1.0079 \text{ g}) + 4 \times (15.9994 \text{ g}) = 180.1571 \text{ g/mol} \)
Molar mass \( \text{NaHCO}_3 = (22.98977 \text{ g}) + (1.0079 \text{ g}) + (12.0107 \text{ g}) + 3 \times (15.9994 \text{ g}) = 84.0066 \text{ g/mol} \)

Molar mass \( \text{C}_6\text{H}_8\text{O}_7 = 6 \times (12.0107 \text{ g}) + 8 \times (1.0079 \text{ g}) + 7 \times (15.9994 \text{ g}) = 192.1232 \text{ g/mol} \)

\[
324 \text{ mg} \ \text{C}_9\text{H}_8\text{O}_4 \times \frac{1 \text{ g} \ \text{C}_9\text{H}_8\text{O}_4}{1000 \text{ mg}} \times \frac{1 \text{ mol} \ \text{C}_9\text{H}_8\text{O}_4}{180.1571 \text{ g} \ \text{C}_9\text{H}_8\text{O}_4} = 0.00180 \text{ mol} \ \text{C}_9\text{H}_8\text{O}_4
\]

\[
1904 \text{ mg} \ \text{NaHCO}_3 \times \frac{1 \text{ g} \ \text{NaHCO}_3}{1000 \text{ mg}} \times \frac{1 \text{ mol} \ \text{NaHCO}_3}{84.0066 \text{ g} \ \text{NaHCO}_3} = 0.02266 \text{ mol} \ \text{NaHCO}_3
\]

\[
1000 \cdot \text{mg} \ \text{C}_6\text{H}_8\text{O}_7 \times \frac{1 \text{ g} \ \text{C}_6\text{H}_8\text{O}_7}{1000 \text{ mg} \ \text{C}_6\text{H}_8\text{O}_7} \times \frac{1 \text{ mol} \ \text{C}_6\text{H}_8\text{O}_7}{192.1232 \text{ g} \ \text{C}_6\text{H}_8\text{O}_7} = 0.005205 \text{ mol} \ \text{C}_6\text{H}_8\text{O}_7
\]

(b) \[
0.00180 \text{ mol} \ \text{C}_9\text{H}_8\text{O}_4 \times \frac{6.022 \times 10^{23} \ \text{C}_9\text{H}_8\text{O}_4 \ \text{molecules}}{1 \ \text{mol} \ \text{C}_9\text{H}_8\text{O}_4} = 1.08 \times 10^{21} \ \text{C}_9\text{H}_8\text{O}_4 \ \text{molecules}
\]

Check your answers: The quantity in moles is always going to be smaller than the mass in grams or milligrams. The number of molecules for a macroscopic sample will be huge. These numbers look right.

60. Define the problem: Given the mass of a compound, determine the number of molecules of that compound.

Develop a plan: Adapt the method described in the answer to Question 52 to calculate the molar mass for the compound. Use the molar mass as a conversion factor between grams and moles, then use Avogadro’s number to determine the actual number of molecules.

Execute the plan:

Molar mass \( \text{CCl}_2\text{F}_2 = (12.0107 \text{ g}) + 2 \times (35.453 \text{ g}) + 2 \times (18.9984 \text{ g}) = 120.914 \text{ g/mol} \)

\[
250 \ \text{g} \ \text{CCl}_2\text{F}_2 \times \frac{1 \ \text{mol} \ \text{CCl}_2\text{F}_2}{120.914 \ \text{g} \ \text{CCl}_2\text{F}_2} \times \frac{6.022 \times 10^{23} \ \text{molecules} \ \text{CCl}_2\text{F}_2}{1 \ \text{mol} \ \text{CCl}_2\text{F}_2} = 1.2 \times 10^{24} \ \text{molecules} \ \text{CCl}_2\text{F}_2
\]

Check your answer: The number of molecules for a macroscopic sample will be huge. These numbers look right.

61. Define the problem: Given the mass of a compound, determine the number of moles of that compound, the number of molecules of that compound, and the number of atoms of both elements.

Develop a plan: Adapt the method described in the answer to Question 51 to calculate the molar mass for the compound. Convert pounds to grams, then use the molar mass as a conversion factor between grams and moles, then use Avogadro’s number to determine the actual number of molecules, then use the formula stoichiometry to determine the number of atoms of each type.

Execute the plan:

Molar mass \( \text{SO}_3 = (32.065 \text{ g}) + 3 \times (15.9994 \text{ g}) = 80.063 \text{ g/mol} \)

(a) \[
1.00 \ \text{lb} \ \text{SO}_3 \times \frac{453.6 \ \text{g} \ \text{SO}_3}{1 \ \text{lb} \ \text{SO}_3} \times \frac{1 \ \text{mol} \ \text{SO}_3}{80.063 \ \text{g} \ \text{SO}_3} = 5.67 \ \text{mol} \ \text{SO}_3
\]

(b) \[
5.67 \ \text{mol} \ \text{SO}_3 \times \frac{6.022 \times 10^{23} \ \text{molecules} \ \text{SO}_3}{1 \ \text{mol} \ \text{SO}_3} = 3.41 \times 10^{24} \ \text{molecules} \ \text{SO}_3
\]

(c) Stoichiometry of the chemical formula: Each molecule of \( \text{SO}_3 \) contains one atom of S.
3.41 \times 10^{24} \text{ molecules } SO_3 \times \frac{1 \text{ S atom}}{1 \text{ SO}_3 \text{ molecule}} = 3.41 \times 10^{24} \text{ S atoms}

(d) Stoichiometry of the chemical formula: Each molecule of SO$_3$ contains three atoms of O.

3.41 \times 10^{24} \text{ molecules } SO_3 \times \frac{3 \text{ O atom}}{1 \text{ SO}_3 \text{ molecule}} = 1.02 \times 10^{25} \text{ O atoms}

Check your answers: The number of molecules for a macroscopic sample will be huge. The atom ratio of sulfur to oxygen in the compound SO$_3$ is 1:3, so the number of atoms of O will be three times greater than the number of atoms of S. These numbers look right.

62. Define the problem: Given the mass of a compound, determine the number of moles of that compound, and the number of atoms of one of the elements.

Develop a plan: Adapt the method described in the answer to Question 51 to calculate the molar mass for the compound. Use the molar mass as a conversion factor between grams and moles, then use Avogadro’s number to determine number of molecules, then use the formula stoichiometry to determine number of atoms of each type.

Execute the plan:

(a) Molar mass CF$_3$CH$_2$F

\[ = 2 \times (12.0107 \text{ g}) + 4 \times (18.9984 \text{ g}) + 2 \times (1.0079 \text{ g}) = 102.0308 \text{ g/mol} \]

25.5 g CF$_3$CH$_2$F \times \frac{1 \text{ mol CF}_3\text{CH}_2\text{F}}{102.0308 \text{ g CF}_3\text{CH}_2\text{F}} = 0.250 \text{ mol CF}_3\text{CH}_2\text{F}

(b) Stoichiometry of the chemical formula: Each mol of CF$_3$CH$_2$F contains 4 mol of F atoms.

\[0.250 \text{ mol CF}_3\text{CH}_2\text{F} \times \frac{4 \text{ mol F}}{1 \text{ mol CF}_3\text{CH}_2\text{F}} \times \frac{6.022 \times 10^{23} \text{ F atoms}}{1 \text{ mol F}} = 6.02 \times 10^{23} \text{ F atoms} \]

Check your answers: The number of atoms in a macroscopic sample will be huge. The mass of a substance will always be larger than the number of moles. These numbers look right.

63. Define the problem: Given the volume of a compound and its density, determine the number of molecules of the compound.

Develop a plan: Adapt the method described in the answer to Question 51 to calculate the molar mass for the compound. Use the density to convert the volume from milliliters to grams, then use the molar mass as a conversion factor between grams and moles, then use Avogadro’s number to determine the number of molecules.

Execute the plan:

Molar mass H$_2$O = 2 \times (1.0079 \text{ g}) + (15.9994 \text{ g}) = 18.0152 \text{ g/mol}

0.050 \text{ mL H}_2\text{O} \times \frac{1.0 \text{ g H}_2\text{O}}{1 \text{ mL H}_2\text{O}} \times \frac{1 \text{ mol H}_2\text{O}}{18.0152 \text{ g H}_2\text{O}} \times \frac{6.022 \times 10^{23} \text{ H}_2\text{O molecules}}{1 \text{ mol H}_2\text{O}} = 1.7 \times 10^{21} \text{ H}_2\text{O molecules}

Check your answer: The number of atoms in a macroscopic sample will be huge. These numbers look right.

64. Define the problem: Given the volume of water with the concentration of a contaminant and the density, determine the number of molecules of the contaminant.
Develop a plan: Adapt the method described in the answer to Question 52 to calculate the molar mass for the compound. Use the density to convert the volume from milliliters to grams. Then use the concentration of the contaminant as a conversion factor between grams of solution and grams of contaminant. Then use the molar mass of the contaminant to determine the moles of the contaminant. Then use Avogadro’s number to determine the number of molecules.

Execute the plan:

Molar mass CHCl₃ = (12.0107 g) + (1.0079 g) + 3 × (35.453 g) = 119.378 g/mol

We treat the unit “ppb” (parts per billion) the same way we treat “percent” (parts per hundred). Both of them relate the mass of part of the sample to the mass of the whole sample. In this case, there are 0.50 ppb CHCl₃ in the water. That means:

In every 1,000,000,000 grams of solution, we find 0.10 gram of CHCl₃.

% Pb = \frac{mass \ of \ Pb \ per \ mol \ PbS \times \ 100 \ %}{mass \ of \ PbS \ per \ mol \ PbS \times \ 100 \ %} \times \ \frac{207.2 \ g \ Pb}{239.3 \ g \ PbS} \times \ 100 \ % = 86.60 \ % \ Pb \ in \ PbS

% S = 100 \ % - 86.60 \ % \ Pb = 13.40 \ % \ S \ in \ PbS

Check your answer: The number of atoms of contaminant present at ppb levels is significantly smaller than the number of atoms in a macroscopic sample (see Question 61 in the Student Solution Manual). This number looks right, though it still seems quite large.

Percent Composition

Define the problem: Given the formula of a compound, determine the molar mass, and the mass percent of each element.

Develop a plan: Calculate the mass of each element in one mole of compound, while calculating the molar mass of the compound. Divide the calculated mass of the element by the molar mass of the compound and multiply by 100 % to get mass percent. To get the last element’s mass percent, subtract the other percentages from 100 

Execute the plan: (a) Mass of Pb per mole of PbS = 207.2 g Pb

Molar mass PbS = (207.2 g) + (32.065 g) = 239.3 g/mol PbS

% Pb = \frac{mass \ of \ Pb \ per \ mol \ PbS \times \ 100 \ %}{mass \ of \ PbS \ per \ mol \ PbS \times \ 100 \ %} \times \ \frac{207.2 \ g \ Pb}{239.3 \ g \ PbS} \times \ 100 \ % = 86.60 \ % \ Pb \ in \ PbS

% S = 100 \ % - 86.60 \ % \ Pb = 13.40 \ % \ S \ in \ PbS

(b) Mass of C per mole of C₂H₆ = 2 × (12.0107 g) = 24.0214 g C

Mass of H per mole of C₂H₆ = 6 × (1.0079 g) = 6.0474 g H

Molar mass C₂H₆ = (24.0214 g) + (6.0474 g) = 30.0688 g/mol C₂H₆
% C = \frac{\text{mass of C}}{\text{mol } C_2H_6} \times \frac{\text{mol } C_2H_6}{\text{mol } C_2H_6} \times 100 % = \frac{24.0214 \text{ g C}}{30.0688 \text{ g } C_2H_6} \times 100 % = 79.8881 \% \text{ C in } C_2H_6

\% H = 100 \% - 79.8881 \% \text{ C} = 20.1119 \% \text{ H in } C_2H_6

(c) Mass of C per mole of CH$_3$CO$_2$H = 2 \times (12.0107 \text{ g}) = 24.0214 \text{ g C}

Mass of H per mole of CH$_3$CO$_2$H = 4 \times (1.0079 \text{ g}) = 4.0316 \text{ g H}

Mass of O per mole of CH$_3$CO$_2$H = 2 \times (15.9994 \text{ g}) = 31.9988 \text{ g O}

Molar mass CH$_3$CO$_2$H = (24.0214 \text{ g}) + (4.0316 \text{ g}) + (31.9988 \text{ g}) = 60.0518 \text{ g/mol CH}_3\text{CO}_2\text{H}

% C = \frac{\text{mass of C}}{\text{mol } CH_3CO_2H} \times \frac{\text{mol } CH_3CO_2H}{\text{mol } CH_3CO_2H} \times 100 % = \frac{24.0214 \text{ g C}}{60.0518 \text{ g } CH_3CO_2H} \times 100 % = 40.0011 \% \text{ C in } CH_3CO_2H

% H = \frac{\text{mass of H}}{\text{mol } CH_3CO_2H} \times \frac{\text{mol } CH_3CO_2H}{\text{mol } CH_3CO_2H} \times 100 % = \frac{4.0316 \text{ g H}}{60.0518 \text{ g } CH_3CO_2H} \times 100 % = 6.7135 \% \text{ H in } CH_3CO_2H

% O = 100 \% - 40.0011 \% \text{ C} - 6.7135 \% \text{ H} = 53.2854 \% \text{ O in } CH_3CO_2H

(d) Mass of N per mole of NH$_4$NO$_3$ = 2 \times (14.0067 \text{ g}) = 28.0134 \text{ g N}

Mass of H per mole of NH$_4$NO$_3$ = 4 \times (1.0079 \text{ g}) = 4.0316 \text{ g H}

Mass of O per mole of NH$_4$NO$_3$ = 3 \times (15.9994 \text{ g}) = 47.9982 \text{ g O}

Molar mass NH$_4$NO$_3$ = (28.0134 \text{ g}) + (4.0316 \text{ g}) + (47.9982 \text{ g}) = 80.0432 \text{ g/mol NH}_4\text{NO}_3

% N = \frac{\text{mass of N}}{\text{mol } NH_4NO_3} \times \frac{\text{mol } NH_4NO_3}{\text{mol } NH_4NO_3} \times 100 % = \frac{28.0134 \text{ g N}}{80.0432 \text{ g } NH_4NO_3} \times 100 % = 34.9979 \% \text{ N in } NH_4NO_3

% H = \frac{\text{mass of H}}{\text{mol } NH_4NO_3} \times \frac{\text{mol } NH_4NO_3}{\text{mol } NH_4NO_3} \times 100 % = \frac{4.0316 \text{ g H}}{80.0432 \text{ g } NH_4NO_3} \times 100 % = 5.0368 \% \text{ H in } NH_4NO_3

% O = 100 \% - 34.9979 \% \text{ C} - 5.0368 \% \text{ H} = 59.9654 \% \text{ O in } NH_4NO_3

Note: When masses of different things are used in the same problem, make sure your units clearly specify what each mass refers to.

Check your answers: Calculating the last element’s mass percent using the formula gives the same answer as subtracting the other percentages from 100 %. These answers are consistent.
66. **Define the problem:** Given the formula of a compound, determine the molar mass, and the mass percent of each element

**Develop a plan:** Calculate the mass of each element in one mole of compound, while calculating the molar mass of the compound. Divide the calculated mass of the element by the molar mass of the compound and multiply by 100 % to get mass percent. To get the last element’s mass percent, subtract the other percentages from 100 %.

**Execute the plan:**

(a) Mass of Mg per mole of MgCO$_3$ = 24.3050 g Mg

\[
\text{Mass of Mg per mole of MgCO}_3 = 24.3050 \text{ g Mg}
\]

Mass of C per mole of MgCO$_3$ = 12.0107 g C

\[
\text{Mass of C per mole of MgCO}_3 = 12.0107 \text{ g C}
\]

Mass of O per mole of MgCO$_3$ = 3 × (15.9994 g) = 47.9982 g O

\[
\text{Mass of O per mole of MgCO}_3 = 3 \times (15.9994 \text{ g}) = 47.9982 \text{ g O}
\]

Molar mass MgCO$_3$ = (24.3050 g) + (12.0107 g) + (47.9982 g) = 84.3139 g/mol MgCO$_3$

\[
\text{Molar mass MgCO}_3 = (24.3050 \text{ g}) + (12.0107 \text{ g}) + (47.9982 \text{ g}) = 84.3139 \text{ g/mol MgCO}_3
\]

\[
\% \text{ Mg} = \left( \frac{\text{mass of Mg}}{\text{mass of MgCO}_3} \right) \times 100 \%
\]

\[
\% \text{ Mg} = \left( \frac{24.3050 \text{ g Mg}}{84.3139 \text{ g MgCO}_3} \right) \times 100 \%
\]

\[
\% \text{ Mg} = 28.8268 \% \text{ Mg in MgCO}_3
\]

\[
\% \text{ C} = \left( \frac{\text{mass of C}}{\text{mass of MgCO}_3} \right) \times 100 \%
\]

\[
\% \text{ C} = \left( \frac{12.0107 \text{ g C}}{84.3139 \text{ g MgCO}_3} \right) \times 100 \%
\]

\[
\% \text{ C} = 14.2452 \% \text{ C in MgCO}_3
\]

\[
100 \% - 28.8268 \% \text{ Mg} - 14.2452 \% \text{ C} = 56.9280 \% \text{ O in MgCO}_3
\]

(b) Mass of C per mole of C$_6$H$_5$OH = 6 × (12.0107 g) = 72.0642 g C

Mass of H per mole of C$_6$H$_5$OH = 6 × (1.0079 g) = 6.0474 g H

Mass of O per mole of C$_6$H$_5$OH = 15.9994 g O

Molar mass C$_6$H$_5$OH = (72.0642 g) + (6.0474 g) + (15.9994 g) = 94.1110 g/mol C$_6$H$_5$OH

\[
\text{Molar mass C}_6\text{H}_5\text{OH} = (72.0642 \text{ g}) + (6.0474 \text{ g}) + (15.9994 \text{ g}) = 94.1110 \text{ g/mol C}_6\text{H}_5\text{OH}
\]

\[
\% \text{ C} = \left( \frac{\text{mass of C}}{\text{mass of C}_6\text{H}_5\text{OH}} \right) \times 100 \%
\]

\[
\% \text{ C} = \left( \frac{72.0642 \text{ g C}}{94.1110 \text{ g C}_6\text{H}_5\text{OH}} \right) \times 100 \%
\]

\[
\% \text{ C} = 76.5736 \% \text{ C in C}_6\text{H}_5\text{OH}
\]

\[
\% \text{ H} = \left( \frac{\text{mass of H}}{\text{mass of C}_6\text{H}_5\text{OH}} \right) \times 100 \%
\]

\[
\% \text{ H} = \left( \frac{6.0474 \text{ g H}}{94.1110 \text{ g C}_6\text{H}_5\text{OH}} \right) \times 100 \%
\]

\[
\% \text{ H} = 6.4258 \% \text{ H in C}_6\text{H}_5\text{OH}
\]

\[
100 \% - 76.5736 \% \text{ C} - 6.4258 \% \text{ H} = 17.0006 \% \text{ O in C}_6\text{H}_5\text{OH}
\]

(c) Mass of C per mole of C$_2$H$_3$O$_3$N = 2 × (12.0107 g) = 24.0214 g C

Mass of H per mole of C$_2$H$_3$O$_3$N = 3 × (1.0079 g) = 3.0237 g H

Mass of O per mole of C$_2$H$_3$O$_3$N = 5 × (15.9994 g) = 79.9970 g O
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Mass of N per mole of C₂H₃O₅N = 14.0067 g N

Molar mass C₂H₃O₅N = (24.0214 g) + (3.0237 g) + (79.9970 g) + (14.0067 g) = 121.0488 g/mol C₂H₃O₅N

% C = \frac{\text{mass of } C}{\text{mol } C} \times 100 \%
= \frac{24.0214 \text{ g } C}{121.0488 \text{ g } C₂H₃O₅N} \times 100 \% = 19.8444 \% C in C₂H₃O₅N

% H = \frac{\text{mass of } H}{\text{mol } C} \times 100 \%
= \frac{3.0237 \text{ g } H}{121.0488 \text{ g } C₂H₃O₅N} \times 100 \% = 2.4979 \% H in C₂H₃O₅N

% O = \frac{\text{mass of } O}{\text{mol } C} \times 100 \%
= \frac{79.9970 \text{ g } O}{121.0488 \text{ g } C₂H₃O₅N} \times 100 \% = 66.0866 \% O in C₂H₃O₅N

% N = 100 \% - 19.8444 \% C - 2.4979 \% H - 66.0866 \% O = 11.5711 \% N in C₂H₃O₅N

(d) Mass of C per mole of C₄H₁₀O₃NPS = 4 \times (12.0107 g) = 48.0428 g C

Mass of H per mole of C₄H₁₀O₃NPS = 10 \times (1.0079 g) = 10.079 g H

Mass of O per mole of C₄H₁₀O₃NPS = 3 \times (15.9994 g) = 47.9982 g O

Mass of N per mole of C₄H₁₀O₃NPS = 14.0067 g N

Mass of P per mole of C₄H₁₀O₃NPS = 30.9738 g P

Mass of S per mole of C₄H₁₀O₃NPS = 32.065 g S

Molar mass C₄H₁₀O₃NPS = (48.0428 g) + (10.0790 g) + (47.9982 g) + (30.9738 g) + (32.065 g S) = 183.166 g/mol C₄H₁₀O₃NPS

% C = \frac{\text{mass of } C}{\text{mol } C₄H₁₀O₃NPS} \times 100 \%
= \frac{48.0428 \text{ g } C}{183.166 \text{ g } C₄H₁₀O₃NPS} \times 100 \% = 26.2292 \% C in C₄H₁₀O₃NPS

% H = \frac{\text{mass of } H}{\text{mol } C₄H₁₀O₃NPS} \times 100 \%
= \frac{10.079 \text{ g } H}{183.166 \text{ g } C₄H₁₀O₃NPS} \times 100 \%
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\[ \frac{10.0079 \text{ g H}}{183.166 \text{ g C}_4\text{H}_{10}\text{O}_3\text{NPS}} \times 100\% = 5.50267\% \text{ H in C}_4\text{H}_{10}\text{O}_3\text{NPS} \]

\[ \% \text{ O} = \frac{\text{mass of O}}{\text{mol C}_4\text{H}_{10}\text{O}_3\text{NPS}} \times 100\% = \frac{47.9982 \text{ g O}}{183.166 \text{ g C}_4\text{H}_{10}\text{O}_3\text{NPS}} \times 100\% = 26.2048\% \text{ O in C}_4\text{H}_{10}\text{O}_3\text{NPS} \]

\[ \% \text{ N} = \frac{\text{mass of N}}{\text{mol C}_4\text{H}_{10}\text{O}_3\text{NPS}} \times 100\% = \frac{14.0067 \text{ g N}}{183.166 \text{ g C}_4\text{H}_{10}\text{O}_3\text{NPS}} \times 100\% = 7.64702\% \text{ N in C}_4\text{H}_{10}\text{O}_3\text{NPS} \]

\[ \% \text{ P} = \frac{\text{mass of P}}{\text{mol C}_4\text{H}_{10}\text{O}_3\text{NPS}} \times 100\% = \frac{30.9738 \text{ g P}}{183.166 \text{ g C}_4\text{H}_{10}\text{O}_3\text{NPS}} \times 100\% = 16.9103\% \text{ P in C}_4\text{H}_{10}\text{O}_3\text{NPS} \]

\[ \% \text{S} = 100\% - 26.2292\% \text{ C} - 5.50267\% \text{ H} - 26.2048\% \text{ O} - 7.64702\% \text{ N} - 16.9103\% \text{ P} = 17.5020\% \text{ O in C}_4\text{H}_{10}\text{O}_3\text{NPS} \]

Note: When masses of different things are used in the same problem, make sure your units clearly specify what each mass refers to.

Check your answers: Calculating the last element’s mass percent using the formula gives the same answer as subtracting the other percentages from 100%. These answers are consistent.

67. Define the problem: Given the mass percent of one compound, M$_2$O, containing one known element, O, and one unknown element, M, calculate the percent by mass of another compound, MO.

Develop a plan: Choose a convenient sample mass of M$_2$O, such as 100.0 g. Find the mass of M and O in the sample, using the given mass percent. Using the molar mass of oxygen as a conversion factor, determine the number of moles of oxygen, then using the formula stoichiometry of M$_2$O as a conversion factor determine the number of moles of M. Find the molar mass of M by dividing the mass of M by the moles of M. Use the molar mass of M, and the formula stoichiometry of MO, to determine the mass percent of M in MO.

Execute the plan:

73.4 % M in M$_2$O means that 100.0 grams of M$_2$O contains 73.4 grams of M.

Mass of O = 100.0 g M$_2$O – 73.4 g M = 26.6 g O

Formula Stoichiometry: 1 mol of M$_2$O contains 2 mol M and 1 mol O.

\[ \frac{26.6 \text{ g O} \times 1 \text{ mol O}}{15.9994 \text{ g O} \times 1 \text{ mol O}} = 3.33 \text{ mol M} \]
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Molar Mass of \( M \) = \( \frac{\text{mass of } M \text{ in sample}}{\text{mol of } M \text{ in sample}} \) = \( \frac{73.4 \text{ g } M}{3.33 \text{ mol } M} = 22.1 \text{ g/mol} \)

Molar mass of \( MO \) = 22.1 g + 15.9994 g = 38.07 g/mol

\[
\% M = \frac{\text{mass of } M}{\text{mol } MO} \times 100\% = \frac{22.1 \text{ g } M}{38.07 \text{ g } MO} \times 100\% = 58.0\% \text{ } M \text{ in } MO
\]

Check your answer: It makes sense that the compound with more atoms of \( M \) has a higher mass percent of \( M \). The closest element to \( M \)'s atomic mass (22.1) is sodium (atomic mass = 22.99). If \( M \) is sodium, the two compounds would probably be sodium oxide (\( \text{Na}_2\text{O} \)) and sodium peroxide (\( \text{Na}_2\text{O}_2 \), a compound made up of two \( \text{Na}^+ \) ions and one \( \text{O}_2^{2-} \) ion. The simple ratio of Na and O atoms in this compound is 1:1). The results make sense.

68. Define the problem: Given the formula of a compound, determine the molar mass, and the mass percent of each element

Develop a plan: Calculate the mass of each element in one mole of compound, while calculating the molar mass of the compound (see full method on Question 52). Divide the calculated mass of the element by the molar mass of the compound and multiply by 100% to get mass percent.

Execute the plan:

Mass of Cu per mole of \( \text{Cu(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O} \) = 63.546 g Cu

Mass of N per mole of \( \text{Cu(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O} \) = \( 4 \times (14.0067) = 56.0268 \) g N

Mass of H per mole of \( \text{Cu(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O} \) = \( (3 \times 4 + 2) \times (1.0079) = 14.1106 \) g H

Mass of S per mole of \( \text{Cu(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O} \) = 32.065 g S

Mass of O per mole of \( \text{Cu(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O} \) = \( (4 + 1) \times (15.9994) = 79.9970 \) g O

Molar mass \( \text{Cu(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O} \) = \( 63.546 \text{ g} + 56.0268 \text{ g} + 14.1106 \text{ g} + 32.065 \text{ g} + 79.9970 \text{ g} \)

\[
= 245.745 \text{ g/mol C}_{4}\text{H}_{10}\text{O}_{3}\text{NPS}
\]

\[
\% \text{ Cu} = \frac{\text{mass of } \text{Cu} / \text{mol } \text{Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}}{\text{mass of } \text{Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O} / \text{mol } \text{Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}} \times 100\% = \frac{63.546 \text{ g Cu}}{245.745 \text{ g } \text{Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}} \times 100\% = 25.858\% \text{ Cu in } \text{Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}
\]

\[
\% \text{ N} = \frac{\text{mass of } \text{N} / \text{mol } \text{Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}}{\text{mass of } \text{Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O} / \text{mol } \text{Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}} \times 100\% = \frac{56.0268 \text{ g N}}{245.745 \text{ g } \text{Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}} \times 100\% = 22.7992\% \text{ N in } \text{Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}
\]

\[
\% \text{ H} = \frac{\text{mass of } \text{H} / \text{mol } \text{Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}}{\text{mass of } \text{Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O} / \text{mol } \text{Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}} \times 100\% = \frac{14.1106 \text{ g H}}{245.745 \text{ g } \text{Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}} \times 100\% = 5.74197\% \text{ H in } \text{Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}
\]
% S = \frac{\text{mass of S}}{\text{mol Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}} \times \frac{\text{mol Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}}{\text{mass of Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}} \times 100 \%
\begin{align*}
32.065 \text{ g S} & = \frac{245.745 \text{ g Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}}{100 \%} \times 13.048 \% \text{ S in Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O} \\
\% O = & \frac{\text{mass of O}}{\text{mol Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}} \times \frac{\text{mol Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}}{\text{mass of Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}} \times 100 \%
\end{align*}
\begin{align*}
79.9970 \text{ g O} & = \frac{245.745 \text{ g Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}}{100 \%} \times 32.5528 \% \text{ O in Co(NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}
\end{align*}

Check your answers: The sum of the percentages add up to 100 %. These answers are consistent.

69. Define the problem: Determine the molar mass of a compound and then determine the mass of a given

Develop a plan: Use mass conversions to calculate the grams of sucrose, then calculate and use the molar mass of sucrose to get the moles of sucrose. Use the formula stoichiometry as a conversion factor to get the moles of carbon. Use the molar mass of carbon to determine the mass of carbon.

Execute the plan:

Molar mass of C_{12}H_{22}O_{11} = 12 \times (12.0107 \text{ g}) + 22 \times (1.0079 \text{ g}) + 11 \times (15.9994 \text{ g}) = 342.2956 \text{ g/mol}

\begin{align*}
1 \text{ lb sucrose} \times \frac{453.59237 \text{ g}}{1 \text{ lb}} \times \frac{1 \text{ mole sucrose}}{342.2956 \text{ g sucrose}} \times \frac{12 \text{ mole C}}{1 \text{ mole sucrose}} \times \frac{12.0107 \text{ g C}}{1 \text{ mol C}} = 190.9915 \text{ g C}
\end{align*}

Check your answer: The percentage of carbon in sucrose is about 40%; therefore, about 40% of the pound needs to be carbon. This answer has the right size.

70. Define the problem: Given the formulas of three compounds, determine the percentage of iron in each of them.

Develop a plan: Calculate the mass of Fe in one mole of compound, while calculating the molar mass of the compound (see full method on Question 52). Divide the calculated mass of the element by the molar mass of the compound and multiply by 100 % to get mass percent.

Execute the plan: For FeCO_{3}: Mass of Fe per mole of FeCO_{3} = 55.845 \text{ g Fe}

\begin{align*}
\text{Molar mass FeCO}_{3} = (55.845 \text{ g}) + (12.0107 \text{ g}) + 3 \times (15.9994 \text{ g}) = 115.856 \text{ g/mol FeCO}_{3}
\end{align*}

\begin{align*}
\% \text{ Fe} = & \frac{\text{mass of Fe/mol FeCO}_{3}}{\text{mass of FeCO}_{3}/\text{mol FeCO}_{3}} \times 100 \%
\end{align*}
\begin{align*}
& \frac{55.845 \text{ g Fe}}{115.856 \text{ g FeCO}_{3}} \times 100 \% = 48.203 \% \text{ Fe in FeCO}_{3}
\end{align*}

For Fe_{2}O_{3}: Mass of Fe per mole of Fe_{2}O_{3} = 2 \times (55.845) \text{ g Fe} = 111.690 \text{ g Fe}

\begin{align*}
\text{Molar mass Fe}_{2}O_{3} = 111.690 \text{ g} + 3 \times (15.9994 \text{ g}) = 159.688 \text{ g/mol Fe}_{2}O_{3}
\end{align*}

\begin{align*}
\% \text{ Fe} = & \frac{\text{mass of Fe/mol Fe}_{2}O_{3}}{\text{mass of Fe}_{2}O_{3}/\text{mol Fe}_{2}O_{3}} \times 100 \%
\end{align*}
\begin{align*}
& \frac{111.694 \text{ g Fe}}{159.688 \text{ g Fe}_{2}O_{3}} \times 100 \% = 69.9426 \% \text{ Fe in Fe}_{2}O_{3}
\end{align*}

For Fe_{3}O_{4}: Mass of Fe per mole of Fe_{3}O_{4} = 3 \times (55.847) \text{ g Fe} = 167.535 \text{ g Fe}

\begin{align*}
\text{Molar mass Fe}_{3}O_{4} = 167.535 \text{ g} + 4 \times (15.9994 \text{ g}) = 231.533 \text{ g/mol Fe}_{3}O_{4}
\end{align*}
\[
\% \text{ Fe} = \frac{\text{mass of Fe}}{\text{mol Fe}_3\text{O}_4} \times \text{mass of Fe}_3\text{O}_4 \times \text{mol Fe}_3\text{O}_4 \times 100 \% = \frac{167.535 \text{ g Fe}}{231.533 \text{ g Fe}_3\text{O}_4} \times 100 \% = 72.3591 \% \text{ Fe in Fe}_3\text{O}_4
\]

Check your answer: The percentage of iron increases as the formula includes more iron and less of other elements.

71. Define the problem: Given the percent by mass of an element in an enzyme and the number of atoms of that element in one molecule of the enzyme, determine the molar mass of the enzyme.

Develop a plan: Choose a convenient sample of the enzyme, such as 100.0 g. Using the percent by mass, determine the number of grams of Mo in the sample. Use the molar mass of Mo as a conversion factor to get the moles of Mo. Use the formula stoichiometry as a conversion factor to get the moles of enzyme. Determine the molar mass by dividing the grams of enzyme in the sample, by the moles of enzyme in the sample.

Execute the plan: The enzyme contains 0.0872 % Mo by mass, this means that 100.0 g of enzyme contains 0.0872 grams Mo.

Formula stoichiometry: One molecule of enzyme contains 2 atoms of Mo, so 1 mol of enzyme molecules contains 2 mol of Mo atoms.

\[
\text{Molar Mass of enzyme} = \frac{\text{mass of enzyme in sample}}{\text{moles of enzyme in sample}} = \frac{100.0000 \text{ g enzyme}}{4.54 \times 10^{-4} \text{ mol enzyme}} = 2.20 \times 10^5 \text{ g/mol}
\]

Check your answer: Enzymes are large molecules with large molar masses. This answer makes sense.

72. (a) A group 6A element is likely to have a –2 charge, since:

\[
\text{anion charge} = (\text{group number}) - 8 = 6 - 8 = -2
\]

(b) Aluminum ion, Al\(^{3+}\) combines with X\(^{2-}\) to form Al\(_2\)X\(_3\).

(c) Define the problem: Given the percent by mass of an element in a compound and the compound’s formula including an unknown element, determine the identity of the unknown element.

Develop a plan: Choose a convenient sample of Al\(_2\)X\(_3\), such as 100.00 g. Using the percent by mass, determine the number of grams of Al and X in the sample. Use the molar mass of Al as a conversion factor to get the moles of Al. Use the formula stoichiometry as a conversion factor to get the moles of X. Determine the molar mass by dividing the grams of X in the sample, by the moles of X in the sample. Using the Periodic Table, determine which Group 6A element has a molar mass nearest this value.

Execute the plan: The compound is 18.55 % Al by mass. This means that 100.00 g of Al\(_2\)X\(_3\) contains 18.55 grams Al and the rest of the mass is from X.

Mass of X in sample = 100.00 g Al\(_2\)X\(_3\) – 18.55 g Al = 81.45 g X

Formula stoichiometry: 1 mol of Al\(_2\)X\(_3\) contains 2 mol of Al atoms.

\[
18.55 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.9815 \text{ g Al}} \times \frac{3 \text{ mol X}}{2 \text{ mol Al}} = 1.031 \text{ mol X}
\]
Molar Mass of X = \frac{\text{mass of X in sample}}{\text{moles of X in sample}} = \frac{81.45 \text{ g X}}{1.031 \text{ mol X}} = 78.98 \text{ g/mol}

Periodic Table indicates that X = ^{34}\text{Se} \text{ (with atomic weight = 78.96 g/mol)}

Check your answer: The molar mass calculated is quite close to that of the Group 6A element, Se. This answer makes sense.

73. Define the problem: Given the percent by mass of an element in a compound and the compound’s formula with an unknown subscript, determine the value of the unknown subscript.

Develop a plan: Choose a convenient sample of Si$_2$H$_x$, such as 100.00 g. Using the percent by mass, determine the number of grams of Si and H in the sample. Use the molar mass of Si as a conversion factor to get the moles of Si. Use the molar mass of H as a conversion factor to get the moles of H. Set up a mole ratio to determine the value of x.

Execute the plan: The compound is 90.28 % Si by mass. This means that 100.00 g of Si$_2$H$_x$ contains 90.28 grams Si and the rest of the mass is from H.

Mass of H in sample = 100.00 g Si$_2$H$_x$ – 90.28 g Si = 9.72 g H

\[
\frac{90.28 \text{ g Si} \times \frac{1 \text{ mol Si}}{28.0855 \text{ g Si}}}{3.214 \text{ mol Si}} = \frac{9.72 \text{ g H} \times \frac{1 \text{ mol H}}{1.0079 \text{ g H}}}{9.64 \text{ mol H}} = \frac{9.64 \text{ mol H}}{3 \text{ mol H}} = \frac{6 \text{ mol H}}{2 \text{ mol Si}}
\]

Molar Ratio = \frac{\text{moles of H in sample}}{\text{moles of Si in sample}} = \frac{9.64 \text{ mol H}}{3.214 \text{ mol Si}} = \frac{6 \text{ mol H}}{2 \text{ mol Si}}

Therefore, the formula is Si$_2$H$_6$ and x = 6.

Check your answer: The molar ratio is clearly a whole number relationship indicating a sensible number of hydrogen atoms in this molecule. The answer makes sense.

74. (a) Define the problem: Given the formula of a hydrate compound, and the formula of the hydrate produced after some of the water has been removed, determine the percentage of mass lost during dehydration.

Develop a plan: Find the molar mass of the original hydrate compound. Determine the number of moles of water it lost, and use the molar mass of water to determine mass of water lost per mole of original hydrate compound. Divide the calculated water mass by the molar mass of the compound and multiply by 100 % to get percent mass lost.

Execute the plan: The original hydrate compound is Na$_2$B$_4$O$_7$⋅10H$_2$O.

Molar Mass of Na$_2$B$_4$O$_7$⋅10H$_2$O = 2 × (22.9898 g) + 4 × (10.811 g) + (7+ 10) × (15.9994 g) + 20 × (1.0079 g) = 381.371 g/mol

Dehydrating 1 mol Na$_2$B$_4$O$_7$⋅10H$_2$O forms 1 mol Na$_2$B$_4$O$_7$⋅5H$_2$O.

Moles of water lost per mole of Na$_2$B$_4$O$_7$⋅10H$_2$O = 5 mol H$_2$O

Mass of water in 5 mol H$_2$O = 10 × (1.0079 g) + 5 × (15.9994 g) = 90.0760 g

\[
\% \text{ H}_2\text{O lost} = \frac{90.0760 \text{ g H}_2\text{O} / \text{mol hydrate}}{381.371 \text{ g hydrate / mol hydrate}} \times 100 \% = 23.6190 \% \text{ H}_2\text{O lost}
\]

Check your answer: 90 is about one quarter of 370. This answer makes sense.
(b) The percent boron by mass will not be the same in these two compounds. Clearly the number of atoms of boron are the same; however, the number of other atoms are significantly different (due to the different amounts of water in the hydrate). The product hydrate, \( \text{Na}_2\text{B}_4\text{O}_7\cdot 5\text{H}_2\text{O} \), will have a larger percent by mass of boron than original hydrate, \( \text{Na}_2\text{B}_4\text{O}_7\cdot 10\text{H}_2\text{O} \), since there are 5 water molecules fewer in the product than the original.

**Empirical and Molecular Formula**

75. An empirical formula shows the simplest whole number ratio of the elements in a compound. The molecular formula gives the actual number of atoms of each element in one formula unit of the compound. For ethane, \( \text{C}_2\text{H}_6 \) is the molecular formula. \( \text{CH}_3 \) is the empirical formula, since 6 is exactly divisible by 2 to give the smallest whole number ratio of 1:3.

76. **Define the problem:** Given the empirical formula of a compound and the molar mass, determine the molecular formula.

**Develop a plan:** Find the mass of 1 mol of the empirical formula. Divide the molar mass of the compound by the calculated empirical mass to get a whole number. Multiply all the subscripts in the empirical formula by this whole number.

**Execute the plan:** The empirical formula is \( \text{CHO} \), the molecular formula is \( (\text{CHO})_n \).

\[
\begin{align*}
\text{Mass of 1 mol of CH}_3 &= 12.0107 \text{ g} + 1.0079 \text{ g} + 15.9994 \text{ g} = 29.0180 \text{ g/mol CH}_3 \\
n &= \frac{\text{mass of 1 mol of molecule}}{\text{mass of 1 mol of CH}_3} = \frac{116.4 \text{ g}}{29.0180 \text{ g}} = 4.011 = 4 \\
\text{Molecular Formula is } (\text{CHO})_4 = \text{C}_4\text{H}_4\text{O}_4 \\
\end{align*}
\]

**Check your answer:** The molar mass is about 4 times larger than the mass of one mole of the empirical formula, so the molecular formula \( \text{C}_4\text{H}_4\text{O}_4 \) makes sense.

77. **Define the problem:** Given the empirical formula of a compound and the molar mass, determine the molecular formula.

**Develop a plan:** Find the mass of 1 mol of the empirical formula. Divide the molar mass of the compound by the calculated empirical mass to get a whole number. Multiply all the subscripts in the empirical formula by this whole number.

**Execute the plan:** The empirical formula is \( \text{C}_2\text{H}_4\text{NO} \), the molecular formula is \( (\text{C}_2\text{H}_4\text{NO})_n \).

\[
\begin{align*}
\text{Mass of 1 mol C}_2\text{H}_4\text{NO} &= 2 \times (12.0107 \text{ g}) + 4 \times (1.0079 \text{ g}) + 14.0067 \text{ g} + 15.9994 \text{ g} = 58.0591 \text{ g/mol} \\
n &= \frac{\text{mass of 1 mol of molecule}}{\text{mass of 1 mol of C}_2\text{H}_4\text{NO}} = \frac{116.1 \text{ g}}{58.0591 \text{ g}} = 2.000 = 2 \\
\text{Molecular Formula is } (\text{C}_2\text{H}_4\text{NO})_2 = \text{C}_4\text{H}_8\text{N}_2\text{O}_2 \\
\end{align*}
\]

**Check your answer:** The molar mass is about 2 times larger than the mass of one mole of the empirical formula, so the molecular formula \( \text{C}_4\text{H}_8\text{N}_2\text{O}_2 \) makes sense.

78. **Define the problem:** Given the molar mass of a compound and the percents by mass of elements in a compound, find the empirical formula and the molecular formula.

**Develop a plan:** Choose a convenient sample of \( \text{C}_x\text{H}_y\text{O}_z\text{F}_w \), such as 100.00 g. Using the percent by mass,
determine the number of grams of each element in the sample. Use the molar mass of each element to get the moles of that element. Set up a mole ratio to determine the values of \( x, y, z, \) and \( w \). Use those integers as subscripts in the empirical formula. Next, calculate the molar mass of the empirical formula, and divide the given molar mass of the compound into this value to determine the number of empirical formulas in the molecular formula and use that to write the formula of the molecule.

**Execute the plan:**

Exactly 100.0 g of \( C_xH_yO_zF_w \) contains 24.0 grams C, 3.0 grams H, 16.0 grams O, and 57.0 grams F.

\[
egin{align*}
24.0 \text{ g C} & \times \frac{1 \text{ mol C}}{12.0107 \text{ g C}} = 2.00 \text{ mol C} \\
3.0 \text{ g H} & \times \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 3.0 \text{ mol H} \\
16.0 \text{ g O} & \times \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 1.00 \text{ mol O} \\
57.0 \text{ g F} & \times \frac{1 \text{ mol F}}{18.9984 \text{ g F}} = 3.00 \text{ mol F}
\end{align*}
\]

Set up ratio and simplify: 

\[
2.00 \text{ mol C} : 3.0 \text{ mol H} : 1.00 \text{ mol O} : 3.00 \text{ mol F}
\]

simplify and write the empirical formula: 

\( 2 \text{ C} : 3 \text{ H} : 1 \text{ O} : 3 \text{ F} \)

The empirical formula is \( \text{C}_2\text{H}_3\text{OF}_3 \).

\[
\text{Mass of 1 mol of } \text{C}_2\text{H}_3\text{OF}_3 = 2 \times (12.0107 \text{ g}) + 3 \times (1.0079 \text{ g}) + 15.9994 \text{ g} + 3 \times (18.9984 \text{ g})
\]

\[
= 100.0397 \text{ g/mol C}_2\text{H}_3\text{OF}_3
\]

\[
n = \frac{\text{mass of 1 mol of molecule}}{\text{mass of 1 mol of } \text{C}_2\text{H}_3\text{OF}_3} = \frac{100.0 \text{ g}}{100.0397 \text{ g}} = 1.000 = 1
\]

Molecular Formula is \( \text{C}_2\text{H}_3\text{OF}_3 \).

**Check your answer:** The empirical mass is very close to the given molecular mass.

79. **Define the problem:** Given the percent mass of the elements in a compound and the molar mass, determine the empirical and molecular formulas.

**Develop a plan:** Choose a convenient mass sample of acetylene, such as 100.00 g. Using the given mass percents, determine the mass of C and H in the sample. Using molar masses of the elements, determine the moles of each element in the sample. Find the whole number mole ratio of the elements C and H to determine the subscripts in the empirical formula. Find the mass of 1 mol of the empirical formula. Divide the molar mass of the compound by the calculated empirical mass to get a whole number. Multiply all the subscripts in the empirical formula by this whole number.

**Execute the plan:**

A 100.00 gram sample will have 92.26 grams of C and 7.74 grams of H.

Find moles of C and H in the sample:

\[
92.26 \text{ g C} \times \frac{1 \text{ mol C}}{12.0107 \text{ g C}} = 7.681 \text{ mol C} \\
7.74 \text{ g H} \times \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 7.68 \text{ mol H}
\]

Mole ratio = \( \frac{7.681 \text{ mol C}}{7.68 \text{ mol H}} = 1.00 = \frac{1 \text{ mol C}}{1 \text{ mol H}} \)

The empirical formula is \( \text{CH} \), so the molecular formula is \( (\text{CH})_n \).

Mass of 1 mol CH = 12.0107 g + 1.0079 g = 13.0186 g / mol
\[
n = \frac{\text{mass of 1 mol of molecule}}{\text{mass of 1 mol of CH}} = \frac{26.02 \text{ g}}{13.0186 \text{ g}} = 1.999 \approx 2
\]

Molecular Formula is \((\text{CH})_2 = \text{C}_2\text{H}_2\)

*Check your answer:* The moles of C and H in the sample are nearly identical, so the empirical formula makes sense. The molar mass is about 2 times larger than the mass of one mole of the empirical formula, so the molecular formula \(\text{C}_2\text{H}_2\) makes sense.

80. **Define the problem:** Given the percent by mass of one element in a compound containing two elements with unknown subscripts and a list of possible empirical formulas, determine the empirical formula.

**Develop a plan:** Choose a convenient sample of \(\text{B}_x\text{H}_y\), such as 100.0 g. Using the percent by mass, determine the number of grams of B and H in the sample. Use the molar masses of B and H as conversion factors to get the moles of B and H. Set up a mole ratio to determine the \(y/x\) ratio. Compare to the list given to select the empirical formula that has the closest mole ratio.

**Execute the plan:** The compound is 88.5 % B by mass. This means that 100.0 g of \(\text{B}_x\text{H}_y\) contains 88.5 grams of B and the rest of the mass is from H.

Mass of H in sample = 100.00 g \(\text{B}_x\text{H}_y\) – 88.5 g B = 11.5 g H

\[
\begin{align*}
88.5 \text{ g B} \times \frac{1 \text{ mol B}}{10.811 \text{ g B}} &= 8.19 \text{ mol B} \\
11.5 \text{ g H} \times \frac{1 \text{ mol H}}{1.0079 \text{ g H}} &= 11.41 \text{ mol H}
\end{align*}
\]

Molar Ratio = \[
\frac{\text{moles of H in sample}}{\text{moles of B in sample}} = \frac{11.41 \text{ mol H}}{8.19 \text{ mol B}} = 1.39
\]

Therefore, the formula is \(\text{B}_5\text{H}_7\).

*Check your answer:* The other possible formulas’ mole ratios were 2.5, 2.2, and 2.0, so the 1.4 ratio (from \(7/5\)) was the closest to 1.39. The answer makes sense.

81. **Define the problem:** Given the percent by mass of one element in a compound containing two elements with unknown subscripts, determine the empirical formula.

**Develop a plan:** Choose a convenient sample of \(\text{N}_x\text{O}_y\), such as 100.00 g. Using the percent by mass, determine the number of grams of N and O in the sample. Use the molar masses of N and O as conversion factors to get the moles of N and O. Set up a mole ratio to determine the whole number \(y/x\) ratio for the empirical formula.

**Execute the plan:** The compound is 36.84 % N by mass. This means that 100.00 g of \(\text{N}_x\text{O}_y\) contains 36.84 grams of N and the rest of the mass is from O.

Mass of O in sample = 100.00 g \(\text{N}_x\text{O}_y\) – 36.84 g B = 63.16 g O

\[
\begin{align*}
36.84 \text{ g N} \times \frac{1 \text{ mol N}}{14.0067 \text{ g N}} &= 2.630 \text{ mol N} \\
63.16 \text{ g O} \times \frac{1 \text{ mol O}}{15.9994 \text{ g O}} &= 3.948 \text{ mol O}
\end{align*}
\]

Molar Ratio = \[
\frac{\text{moles of O in sample}}{\text{moles of N in sample}} = \frac{3.948 \text{ mol O}}{2.630 \text{ mol N}} = 1.501 \approx \frac{3}{2}
\]

Therefore, the formula is \(\text{N}_2\text{O}_3\).
Check your answer: The mole ratio is very close to \( \frac{3}{2} \). The answer makes sense.

82. Define the problem: Given the percent by mass of one element in a compound containing two elements with unknown subscripts and the molar mass, determine the empirical formula and the molecular formula.

Develop a plan: Choose a convenient sample of \( \text{C}_x\text{H}_y \), such as 100.00 g. Using the percent by mass, determine the number of grams of C and H in the sample. Use the molar masses of C and H as conversion factors to get the moles of C and H. Set up a mole ratio to determine the whole number \( y/x \) ratio for the empirical formula. Find the mass of 1 mol of the empirical formula. Divide the molar mass of the compound by the calculated empirical mass to get a whole number. Multiply all the subscripts in the empirical formula by this whole number.

Execute the plan: The compound is 89.94 % C by mass. This means that 100.00 g of \( \text{C}_x\text{H}_y \) contains 89.94 grams C and the rest of the mass is from H.

Mass of H in sample = 100.00 g \( \text{C}_x\text{H}_y \) – 89.94 g C = 10.06 g H

Find moles of C and H in the sample:

\[
\begin{align*}
89.94 \text{ g C} &\times \frac{1 \text{ mol C}}{12.0107 \text{ g C}} = 7.488 \text{ mol C} \\
10.06 \text{ g H} &\times \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 9.981 \text{ mol H}
\end{align*}
\]

Mole ratio = \( \frac{9.981 \text{ mol H}}{7.488 \text{ mol C}} = 1.333 \approx \frac{4}{3} \)

The empirical formula is \( \text{C}_3\text{H}_4 \), so the molecular formula is \( (\text{C}_3\text{H}_4)_n \).

Mass of 1 mol \( \text{C}_3\text{H}_4 \) = \( 3 \times (12.0107 \text{ g}) + 4 \times (1.0079 \text{ g}) = 40.0637 \text{ g/mol} \)

\[
n = \frac{\text{mass of 1 mol of molecule}}{\text{mass of 1 mol of } \text{C}_3\text{H}_4} = \frac{120.2 \text{ g}}{40.0637 \text{ g}} = 3.000 = 3
\]

Molecular Formula is \( (\text{C}_3\text{H}_4)_3 = \text{C}_9\text{H}_{12} \)

Check your answer: The mole ratio of C and H in the sample is very close to \( \frac{4}{3} \), so the empirical formula makes sense. The molar mass is 3 times larger than the mass of one mole of the empirical formula, so the molecular formula \( \text{C}_9\text{H}_{12} \) makes sense.

83. Define the problem: Given the percent by mass of all the elements in a compound and the molar mass, determine the empirical formula and the molecular formula.

Develop a plan: Choose a convenient sample of \( \text{C}_x\text{H}_y\text{O}_z \), such as 100.00 g. Using the percent by mass, determine the number of grams of C, H, and O in the sample. Use the molar masses of C, H, and O as conversion factors to get the moles of C, H, and O. Set up a mole ratio to determine the whole number \( x:y:z \) ratio for the empirical formula. Find the mass of 1 mol of the empirical formula. Divide the molar mass of the compound by the calculated empirical mass to get a whole number. Multiply all the subscripts in the empirical formula by this whole number.

Execute the plan:

(a) The compound is 40.0 % C, 6.71 % H, and 53.29 % O by mass. This means that 100.00 g of \( \text{C}_x\text{H}_y\text{O}_z \) contains 40.0 grams C, 6.71 grams H, and 53.29 grams O.
Chapter 3: Chemical Compounds

Find moles of C, H, and N in the sample:

\[
\begin{align*}
40.0 \text{ g C} & \times \frac{1 \text{ mol C}}{12.0107 \text{ g C}} = 3.33 \text{ mol C} \\
6.71 \text{ g H} & \times \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 6.66 \text{ mol H} \\
17.35 \text{ g O} & \times \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 3.331 \text{ mol O}
\end{align*}
\]

Mole ratio: 3.33 mol C : 6.66 mol H : 3.331 mol O

Divide each term by the smallest number of moles, 3.33 mol

Atom ratio: 1 C : 2 H : 1 O

The empirical formula is CH\(_2\)O

(b) The molecular formula is \((\text{CH}_2\text{O})_n\).

Mass of 1 mol CH\(_2\)O = 12.0107 g + 2 \times (1.0079 g) + 15.9994 g = 30.0259 g / mol

\[
n = \frac{\text{mass of 1 mol of molecule}}{\text{mass of 1 mol of CH}_2\text{O}} = \frac{60.0 \text{ g}}{30.0259 \text{ g}} = 2.00 = 2
\]

Molecular Formula is \((\text{CH}_2\text{O})_2 = \text{C}_2\text{H}_4\text{O}_2\)

Check your answers: The mole ratios of C, H, and O in the sample are very close to integer values, so the empirical formula makes sense. The molar mass is 2 times larger than the mass of one mole of the empirical formula, so the molecular formula C\(_2\)H\(_4\)O\(_2\) makes sense.

84. Define the problem: Given the percent by mass of all the elements in a compound and the molar mass, determine the empirical formula and the molecular formula.

Develop a plan: Choose a convenient sample of C\(_x\)H\(_y\)N\(_z\), such as 100.00 g. Using the percent by mass, determine the number of grams of C, H, and N in the sample. Use the molar masses of C, H, and N as conversion factors to get the moles of C, H and N. Set up a mole ratio to determine the whole number x:y:z ratio for the empirical formula. Find the mass of 1 mol of the empirical formula. Divide the molar mass of the compound by the calculated empirical mass to get a whole number. Multiply all the subscripts in the empirical formula by this whole number.

Execute the plan:

(a) The compound is 74.0 % C, 8.65 % H, and 17.35 % N by mass. This means that 100.00 g of C\(_x\)H\(_y\)N\(_z\) contains 74.0 grams C, 8.65 grams H, and 17.35 grams N.

Find moles of C, H, and N in the sample:

\[
\begin{align*}
74.0 \text{ g C} & \times \frac{1 \text{ mol C}}{12.0107 \text{ g C}} = 6.16 \text{ mol C} \\
8.65 \text{ g H} & \times \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 8.58 \text{ mol H} \\
17.35 \text{ g N} & \times \frac{1 \text{ mol N}}{14.0067 \text{ g N}} = 1.239 \text{ mol N}
\end{align*}
\]

Mole ratio: 6.16 mol C : 8.58 mol H : 1.239 mol N

Divide each term by the smallest number of moles, 1.239 mol

Atom ratio: 5 C : 7 H : 1 N
(b) The molecular formula is \((\text{C}_5\text{H}_7\text{N})_n\).

Mass of 1 mol \(\text{C}_5\text{H}_7\text{N} = 5 \times (12.0107 \text{ g}) + 7 \times (1.0079 \text{ g}) + 14.0067 \text{ g} = 81.1155 \text{ g / mol}\)

\[
n = \frac{\text{mass of 1 mol of molecule}}{\text{mass of 1 mol of C}_5\text{H}_7\text{N}} = \frac{162 \text{ g}}{81.1155 \text{ g}} = 2.00 = 2
\]

Molecular Formula is \((\text{C}_5\text{H}_7\text{N})_2 = \text{C}_{10}\text{H}_{14}\text{N}_2\)

_Check your answers:_ The mole ratios of C, H, and N in the sample are close to integer values, so the empirical formula makes sense. The molar mass is twice the mass of one mole of the empirical formula, so the molecular formula \(\text{C}_{10}\text{H}_{14}\text{N}_2\) makes sense.

85. _Define the problem:_ Given the percent by mass of all the elements in a compound and the molar mass, determine the empirical formula and the molecular formula.

_Develop a plan:_ Choose a convenient sample of \(\text{C}_x\text{H}_y\text{As}_z\), such as 100.00 g. Using the percent by mass, determine the number of grams of C, H, and As in the sample. Use the molar masses of C, H, and As as conversion factors to get the moles of C, H and As. Set up a mole ratio to determine the whole number \(x:y:z\) ratio for the empirical formula. Find the mass of 1 mol of the empirical formula. Divide the molar mass of the compound by the calculated empirical mass to get a whole number. Multiply all the subscripts in the empirical formula by this whole number.

_Execute the plan:_

(a) The compound is 22.88 % C, 5.76 % H, and 71.36 % As by mass. This means that 100.00 g of \(\text{C}_x\text{H}_y\text{As}_z\) contains 22.88 grams C, 5.76 grams H, and 71.36 grams As.

Find moles of C, H, and As in the sample:

\[
22.88 \text{ g C} \times \frac{1 \text{ mol C}}{12.0107 \text{ g C}} = 1.905 \text{ mol C}
\]

\[
5.76 \text{ g H} \times \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 5.72 \text{ mol H}
\]

\[
71.36 \text{ g As} \times \frac{1 \text{ mol As}}{74.9216 \text{ g As}} = 0.9525 \text{ mol As}
\]

Mole ratio: \(1.905 \text{ mol C} : 5.72 \text{ mol H} : 0.9525 \text{ mol As}\)

Divide each term by the smallest number of moles, 0.9525 mol

Atom ratio: \(2 \text{ C} : 6 \text{ H} : 1 \text{ As}\)

The empirical formula is \(\text{C}_2\text{H}_6\text{As}\)

(b) The molecular formula is \((\text{C}_2\text{H}_6\text{As})_n\).

Mass of 1 mol \(\text{C}_2\text{H}_6\text{As} = 2 \times (12.0107 \text{ g}) + 6 \times (1.0079 \text{ g}) + 74.9216 \text{ g} = 104.9904 \text{ g / mol}\)

\[
n = \frac{\text{mass of 1 mol of molecule}}{\text{mass of 1 mol of C}_2\text{H}_6\text{As}} = \frac{210 \text{ g}}{104.9904 \text{ g}} = 2.00 = 2
\]

Molecular Formula is \((\text{C}_2\text{H}_6\text{As})_2 = \text{C}_4\text{H}_{12}\text{As}_2\)

_Check your answers:_ The mole ratios of C, H, and As in the sample are close to integer values, so the empirical formula makes sense. The molar mass is 2 times larger than the mass of one mole of the empirical formula, so the molecular formula \(\text{C}_4\text{H}_{12}\text{As}_2\) makes sense.
Define the problem: Given the percent by mass of all the elements in a compound and the molar mass, determine the molecular formula.

Develop a plan: Choose a convenient sample of C$_x$H$_y$N$_z$, such as 100.00 g. Using the percent by mass, determine the number of grams of C, H, and N in the sample. Use the molar masses of C, H, and N as conversion factors to get the moles of C, H and N. Set up a mole ratio to determine the whole number $x:y:z$ ratio for the empirical formula. Find the mass of 1 mol of the empirical formula. Divide the molar mass of the compound by the calculated empirical mass to get a whole number. Multiply all the subscripts in the empirical formula by this whole number.

Execute the plan:

The compound is 58.77 % C, 13.81 % H, and 27.42 % N by mass. This means that 100.00 g of C$_x$H$_y$N$_z$ contains 58.77 grams C, 13.81 grams H, and 27.42 grams N.

Find moles of C, H, and N in the sample:

\[
\begin{align*}
58.77 \text{ g C} & \times \frac{1 \text{ mol C}}{12.0107 \text{ g C}} = 4.893 \text{ mol C} \\
13.81 \text{ g H} & \times \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 13.70 \text{ mol H} \\
27.42 \text{ g N} & \times \frac{1 \text{ mol N}}{14.0067 \text{ g N}} = 1.958 \text{ mol N}
\end{align*}
\]

Mole ratio: 4.893 mol C : 13.70 mol H : 1.958 mol N

Divide each term by the smallest number of moles, 1.958 mol, to get atom ratio:

Atom ratio: 2.5 C : 7 H : 1 N

Multiply each term by 2 to get whole number atom ratio:

Whole number atom ratio: 5 C : 14 H : 2 N

The empirical formula is C$_5$H$_{14}$N$_2$, so the molecular formula is (C$_5$H$_{14}$N$_2$)$_n$.

Check your answer: The mole ratios of C, H, and N in the sample are close to integer values, so the empirical formula makes sense. The molar mass is almost the same as the mass of one mole of the empirical formula, so the molecular formula C$_5$H$_{14}$N$_2$ makes sense.

Define the problem: Given the percent by mass of two of the three elements in a compound, determine the empirical formula.

Develop a plan: Choose a convenient sample of C$_x$H$_y$Cl$_z$, such as 100.00 g. Using the percent by mass, determine the number of grams of C, H, and Cl in the sample. Use the molar masses of C, H, and Cl as conversion factors to get the moles of C, H and Cl. Set up a mole ratio to determine the whole number $x:y:z$ ratio for the empirical formula.

Execute the plan:

The compound is 47.5 % C, 2.45 % H, and the rest is Cl. This means that 100.00 g of C$_x$H$_y$Cl$_z$ contains 47.5 grams C, 2.45 grams H, and the rest is Cl.
Mass of Cl = 100.00 g C\textsubscript{x}H\textsubscript{y}Cl\textsubscript{z} – 47.5 g C – 2.54 g H = 50.0 g Cl

Find moles of C, H, and Cl in the sample:

\[
\begin{align*}
47.5 \text{ g C} & \times \frac{1 \text{ mol C}}{12.0107 \text{ g C}} = 3.95 \text{ mol C} \\
2.54 \text{ g H} & \times \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 2.52 \text{ mol H} \\
50.0 \text{ g Cl} & \times \frac{1 \text{ mol Cl}}{35.4527 \text{ g Cl}} = 1.41 \text{ mol Cl}
\end{align*}
\]

\[
\text{Mole ratio: } 3.95 \text{ mol C} : 2.52 \text{ mol H} : 1.41 \text{ mol Cl}
\]

Divide each term by the smallest number of moles, 1.41 mol, to get atom ratio:

\[
\text{Atom ratio: } 2.80 \text{ C} : 1.79 \text{ H} : 1 \text{ Cl}
\]

Fractions that are multiples of 0.2 can be multiplied by 5 to make whole numbers, since \(5 \times 0.2 = 1\). Multiply each term by 5 to get whole number atom ratio:

\[
\text{Whole number atom ratio: } 14 \text{ C} : 9 \text{ H} : 5 \text{ Cl}
\]

The empirical formula is C\textsubscript{14}H\textsubscript{9}Cl\textsubscript{5}.

\textit{Check your answer:} The simple atom ratios of C, H, and N in the sample are not close to integer values, so a multiplier of 5 has to be applied. After this multiplication, the whole numbers are not so approximate. This empirical formula makes sense. If the mass percentages are based on data with possible large errors, the answer might also be C\textsubscript{11}H\textsubscript{7}Cl\textsubscript{4} (if 0.80 is rounded to 0.75) or C\textsubscript{3}H\textsubscript{2}Cl (if 0.80 is rounded to 1).

88. Define the problem: The mass of a sample of a hydrate compound is given. The formula of the hydrate is known except for the amount of water in it. All the water is dried out of the sample using high temperature, leaving a mass of completely dehydrated salt. Determine out the number of water molecules in the formula of the hydrate compound.

\textit{Develop a plan:} Use the molar mass of the dehydrated compound as a conversion factor to convert the mass of the dehydrated compound into moles. Since the only thing lost was water, a mole relationship can be established between the dehydrated and hydrated compound. In addition, the difference between the mass of the hydrated compound and the mass of the dehydrated compound gives the amount of water lost by the sample. Convert that water into moles, using the molar mass of water. Divide the moles of water by the moles of hydrate, to determine how many moles of water are in one mole of hydrate compound.

\textit{Execute the plan:}

1.687 g of hydrated compound, MgSO\textsubscript{4}·xH\textsubscript{2}O, is dehydrated to make 0.824 g of the dehydrated compound, MgSO\textsubscript{4}.

Molar Mass MgSO\textsubscript{4} = 24.305 g + 32.065 g + 4 \times (15.9994 g) = 120.368 g/mol

Molar Mass H\textsubscript{2}O = 2 \times (1.0079 g) + 2 \times (15.9994 g) = 18.0152 g/mol

Find moles of MgSO\textsubscript{4}·xH\textsubscript{2}O in the sample:

\[
0.824 \text{ g MgSO}_4 \times \frac{1 \text{ mol MgSO}_4}{120.368 \text{ g MgSO}_4} \times \frac{1 \text{ mol MgSO}_4 \cdot x\text{H}_2\text{O}}{1 \text{ mol MgSO}_4} = 6.85 \times 10^{-3} \text{ mol MgSO}_4 \cdot x\text{H}_2\text{O}
\]

Find mass of water lost by the sample: (1.687 g MgSO\textsubscript{4}·xH\textsubscript{2}O) – (0.824 g MgSO\textsubscript{4}) = 0.863 g H\textsubscript{2}O
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Find moles of H\(_2\)O lost from sample: \[0.863 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0152 \text{ g H}_2\text{O}} = 0.0479 \text{ mol H}_2\text{O}\]

\[\text{Mole ratio} = \frac{\text{mol H}_2\text{O from sample}}{\text{mol hydrate is sample}} = \frac{0.0479 \text{ mol H}_2\text{O}}{6.85 \times 10^{-3} \text{ mol hydrate}} = 7.00 = 7\]

The proper formula for the hydrate is MgSO\(_4\)\(_7\)H\(_2\)O and \(x = 7\).

Check your answer: The formula for epsom salt was given earlier in Chapter 3 (Question 44). It had 7 water molecules then, so it better have 7 water molecules, now!

89. Define the problem: The mass of a sample of a hydrate compound is given. The formula of the hydrate is known except for the amount of water in it. All the water is dried out of the sample using high temperature, losing a mass of water. Determine the number of moles of water in the formula of the hydrate compound.

Develop a plan: The difference between the mass of the hydrated compound and the mass of the water lost by the sample gives the mass of dehydrated compound. Use the molar mass of the dehydrated compound as a conversion factor to convert the mass of the dehydrated compound into moles. Since the only thing lost was water, a mole relationship can be established between the dehydrated and hydrated compound. Convert the mass of water into moles, using the molar mass of water. Divide the moles of water by the moles of hydrate, to determine how many moles of water came from one mole of hydrate compound.

Execute the plan:

A sample of 4.74 g of hydrated compound, KAl(SO\(_4\))\(_2\)-\(x\)H\(_2\)O, is dehydrated, losing 2.16 g H\(_2\)O, producing KAl(SO\(_4\))\(_2\).

Find mass of dehydrated compound produced from the sample:

\[= 4.74 \text{ g KAl(SO}_4)_2\cdot x\text{H}_2\text{O} - 2.16 \text{ g H}_2\text{O} = 2.58 \text{ g KAl(SO}_4)_2\]

Molar Mass KAl(SO\(_4\))\(_2\) = 39.0983 g + 26.9815 g + 2 \times (32.065 g) + 8 \times (15.9994 g) = 258.206 g/mol

Molar Mass H\(_2\)O = 2 \times (1.0079 g) + 2 \times (15.9994 g) = 18.0152 g/mol

Find moles of KAl(SO\(_4\))\(_2\)-xH\(_2\)O in sample:

\[2.58 \text{ g KAl(SO}_4)_2\cdot x\text{H}_2\text{O} \times \frac{1 \text{ mol KAl(SO}_4)_2\cdot x\text{H}_2\text{O}}{258.206 \text{ g KAl(SO}_4)_2\cdot x\text{H}_2\text{O}} \times \frac{1 \text{ mol KAl(SO}_4)_2\cdot x\text{H}_2\text{O}}{1 \text{ mol KAl(SO}_4)_2\cdot x\text{H}_2\text{O}} = 9.99 \times 10^{-3} \text{ mol KAl(SO}_4)_2\cdot x\text{H}_2\text{O}\]

Find moles of H\(_2\)O lost from sample:

\[2.16 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0152 \text{ g H}_2\text{O}} = 0.120 \text{ mol H}_2\text{O}\]

\[\text{Mole ratio} = \frac{\text{mol H}_2\text{O from sample}}{\text{mol hydrate is sample}} = \frac{0.120 \text{ mol H}_2\text{O}}{9.99 \times 10^{-3} \text{ mol hydrate}} = 12.0 = 12\]

The proper formula of the hydrated compound is: KAl(SO\(_4\))\(_2\)-12H\(_2\)O and \(x = 12\).

Check your answer: The mole ratio is very close to a whole number. This looks right.
Biological Periodic Table

90. According to Table 3.11, the ten elements most abundant in the human body are: H, O, C, N, Ca, P, Cl, S, Na, K.

91. One compound makes up about 60% of the human body and is nearly 90% oxygen. This compound is H₂O. Other compounds that have significant oxygen contribution are phosphate and carbonate compounds. Carbohydrates have more oxygen than fats and oils.

92. (a) Metals are found in the body as ions.

(b) Two uses for metals in the body are calcium (Ca²⁺) in bones and Fe²⁺ in hemoglobin. There are many others.

93. Macrominerals are also called “major minerals” which are more abundant in the human body (see Figure 3.7 in the textbook). Major minerals are present in greater than 0.01% of body weight (100 mg per kg). Microminerals are also called “trace minerals” which are less plentiful. Trace minerals are present in less than 0.01% of body weight (100 mg per kg), sometimes far less.

94. An essential mineral that is toxic at high levels is selenium. (See the “Chemistry in the News” box on page 107). This is the only one mentioned in the text. Most heavy metals such as iron and copper and nonmetals such as iodine are safe and necessary, but they are toxic at high concentrations.

General Questions

95. (a) A crystal of sodium chloride is a three dimensional array of alternating anions and cations:

(b) Cleavage at these locations would still keep ion pairs (Na⁺Cl⁻) together.

96. (a) Trinitrotoluene, TNT, has seven C atoms, six in the ring and one more in the –CH₃ at the top of the structure. It has five H atoms, two on ring carbons and three more in the –CH₃ at the top of the structure. It has three N atoms (one in each of the three –NO₂ groups attached to the ring carbons). It also has six O atoms (two in each of the three –NO₂ groups attached to the ring carbons). So, the molecular formula is C₇H₅N₃O₆.

(b) Serine has three carbon atoms. It has three hydrogen atoms attached to carbons, two H atoms in the two –OH groups, and two H atoms in the –NH₂ groups. It has one N atom. It has two O atoms in the –OH groups and one more in the =O. So, the molecular formula is C₃H₇NO₃.
97. **Define the problem:** Given one molecule of nitrogen, determine its mass in grams. Given one molecule of oxygen, determine its mass in grams. Find the ratio of the masses of these two atoms and compare that ratio to the ratio of atomic weights of nitrogen and oxygen.

**Develop a plan:** Use the periodic table to get the atomic weights of nitrogen and oxygen and equate those to the mass in grams of one mole of nitrogen molecules. Divide those numbers by Avogadro’s number to get the masses of one nitrogen and one oxygen molecule in grams. Take the ratio and compare to the direct ratio.
Chapter 3: Chemical Compounds

Execute the plan:

(a) \[
\frac{28.0134 \text{ g N}_2}{\text{1 mol N}_2 \text{ molecules}} \times \frac{1 \text{ mol N}_2 \text{ molecules}}{6.022 \times 10^{23} \text{ N}_2 \text{ molecules}} = 4.652 \times 10^{-23} \text{ g N}_2 \text{ molecules}
\]

(b) \[
\frac{31.9988 \text{ g O}_2}{\text{1 mol O}_2 \text{ molecules}} \times \frac{1 \text{ mol O}_2 \text{ molecules}}{6.022 \times 10^{23} \text{ O}_2 \text{ molecules}} = 5.314 \times 10^{-23} \text{ g O}_2 \text{ molecules}
\]

(c) molecule mass ratio = \[
\frac{5.314 \times 10^{-23} \text{ g O}_2}{4.652 \times 10^{-23} \text{ g N}_2} \approx 1.14227
\]

atomic weight ratio = \[
\frac{31.9994 \text{ amu O}_2}{28.0134 \text{ amu N}_2} \approx 1.14227
\]
The ratios are the identical.

Check your answer: The molecules are very small, so their mass should be very small. Avogadro’s number is a physical constant, so the ratio of the masses must be the same.

98. (a) Chlorine (Cl) and bromine (Br) are not likely to form an ionic compound, since they are both nonmetals in Group 7A.

(b) Lithium (Li) and tellurium (Te) might make an ionic compound. Lithium is a metal and tellurium is a metalloid. The likely compound contains ions Li\(^+\) (Group 1A cation; charge is +1) and Te\(^{2-}\) (Group 6A anion; charge is –2). The compound’s formula will be Li\(_2\)Te.

(c) Sodium (Na) and argon (Ar) are not likely to form an ionic compound, since argon is in Group 8A. Those elements are very unreactive and do not form ions at all.

(d) Magnesium (Mg) and fluorine (F) will make an ionic compound. Magnesium is a metal and fluorine is a nonmetal. The likely compound contains ions Mg\(^{2+}\) (Group 2A cation; charge is +2) and F\(^-\) (Group 7A anion; charge is –1). The compound’s formula will be MgF\(_2\).

(e) Nitrogen (N) and bromine (Br) are not likely to form an ionic compound, since they are both nonmetals in Groups 5A and 7A, respectively.

(f) Indium (In) and sulfur (S) will make an ionic compound. Indium is a metal and sulfur is a nonmetal. The likely compound contains ions In\(^{3+}\) (Group 3A cation; charge is +3) and S\(^{2-}\) (Group 6A anion; charge is –2). The compound’s formula will be In\(_2\)S\(_3\).

(g) Selenium (Se) and bromine (Br) are not likely to form an ionic compound, since they are both nonmetals in Groups 5A and 7A, respectively.

99. (a) binary molecular compound (both nonmetals; not ionic): chlorine tribromide

(b) binary molecular compound (both nonmetals; not ionic): nitrogen trichloride

(c) ionic compound (common cation, Ca\(^{2+}\), and anion, SO\(_4^{2-}\)): calcium sulfate

(d) organic compound (alkane): heptane

(e) binary molecular compound (both nonmetals; not ionic): xenon tetrafluoride

(f) binary molecular compound (both nonmetals; not ionic): oxygen difluoride

(g) ionic compound (common cation, Na\(^+\), and anion, I\(^-\)): sodium iodide
(h) ionic compound (common cation, Al$^{3+}$, and anion, S$^{2-}$): aluminum sulfide
(i) binary molecular compound (both nonmetals; not ionic): phosphorus pentachloride
(j) ionic compound (common cation, K$^{+}$, and anion, PO$_4^{3-}$): potassium phosphate

100. (a) sodium hypochlorite (common cation, Na$^+$, and anion, ClO$^-$; ionic): NaClO
(b) aluminum perchlorate (common cation, Al$^{3+}$, and anion, ClO$_4^-$; ionic): Al(ClO$_4$)$_3$
(c) potassium permanganate (common cation, K$^+$, and anion, MnO$_4^-$; ionic): KMnO$_4$
(d) potassium dihydrogen phosphate (common cation, K$^+$, and anion, H$_2$PO$_4^-$; ionic): KH$_2$PO$_4$
(e) chlorine trifluoride (binary molecular compound since both elements are nonmetals; not ionic). ClF$_3$
(f) boron tribromide (binary molecular compound since both elements are nonmetals; not ionic). BBr$_3$
(g) calcium acetate (common cation, Ca$^{2+}$, and anion, CH$_3$CO$_2^-$ or CH$_3$COO$^-$; ionic): Ca(CH$_3$COO)$_2$ or Ca(CH$_3$CO)$_2$ $\text{Either one is correct.}$
(h) sodium sulfite (common cation, Na$^+$, and anion, SO$_3^{2-}$; ionic): Na$_2$SO$_3$
(i) disulfur tetrachloride (binary molecular compound since both elements are nonmetals; not ionic). S$_2$Cl$_4$
(j) phosphorus trifluoride (binary molecular compound since both elements are nonmetals; not ionic). PF$_3$

101. Define the problem: Given the mass of a block of platinum in troy ounces and the density of platinum, determine the number of moles of metal and the volume of the block in cubic centimeters.

Develop a plan: Use the given relationship between troy ounces and grams as a conversion factor to determine the mass of the block in grams. (a) Use the periodic table to get the molar mass of platinum (Pt). Use the molar mass as a conversion factor to get the number of moles of Pt in the block. (b) Use the density as a conversion factor to get the block’s volume in cubic centimeters.

Execute the plan:

(a) $15.0 \text{ troy oz Pt} \times \frac{31.1 \text{ g Pt}}{1 \text{ troy oz Pt}} \times \frac{1 \text{ mol Pt}}{195.078 \text{ g Pt}} = 2.39 \text{ mol Pt}$

(b) $15.0 \text{ troy oz Pt} \times \frac{31.1 \text{ g Pt}}{1 \text{ troy oz Pt}} \times \frac{1 \text{ cm}^3 \text{ Pt}}{21.45 \text{ g Pt}} = 21.7 \text{ cm}^3 \text{ Pt}$

Check your answers: The macroscopic sample has a convenient number of moles. The volume should be approximately $\frac{3}{2}$ times the troy ounces. These numbers look right.

102. Define the problem: Given the moles of a cube of lithium and the density of lithium, determine the volume of the cube in cubic centimeters and the length of one edge of the cube.

Develop a plan: Start with the sample – the number of moles in the cube. Use the molar mass of lithium (Li) as a conversion factor to get the number of grams of Li. Then use the density as a conversion factor to get the volume in cubic centimeters. Use the relationship between the edges of a cube and the volume of a cube to find the length of one edge of the cubic block.
Chapter 3: Chemical Compounds

Execute the plan:

(a) \[ \text{Volume} = \frac{256 \text{ mol Li} \times 6.941 \text{ g Li}}{1 \text{ mol Li}} \times \frac{1 \text{ cm}^3 \text{ Li}}{0.534 \text{ g Li}} = 3.33 \times 10^3 \text{ cm}^3 \text{ Li} \]

(b) \[ V = (x \text{ cm})^3 \]
\[ 3.33 \times 10^3 \text{ cm}^3 = (x \text{ cm})^3 \]
\[ x \text{ cm} = \sqrt[3]{3.33 \times 10^3 \text{ cm}^3} = 14.9 \text{ cm is the length of one side} \]

Check your answer: The question made it sound like this cube might not fit on the starship. Even this large number of moles only occupies a cubic space that has 15 cm (about 6 inches) per side. That’s not very big. Rechecking the math, though, confirms that these numbers are calculated correctly; therefore, we don’t have to write the authors of Star Trek and tell them to redesign their ships.

103. Define the problem: Given the percent by mass of all the elements in a compound and the molar mass, determine the empirical formula and the molecular formula.

Develop a plan: Choose a convenient sample of \( \text{C}_x\text{O}_y\text{F}_z \), such as 100.0 g. Using the percent by mass, determine the number of grams of C, O, and F in the sample. Use the molar masses of C, O, and F as conversion factors to get the moles of C, O and F. Set up a mole ratio to determine the whole number \( x:y:z \) ratio for the empirical formula. Find the mass of 1 mol of the empirical formula. Divide the molar mass of the compound by the calculated empirical mass to get a whole number. Multiply all the subscripts in the empirical formula by this whole number.

Execute the plan:

(a) The compound is 14.6 % C, 39.0 % O, and 46.3 % F by mass. This means that 100.00 g of \( \text{C}_x\text{O}_y\text{F}_z \) contains 14.6 grams C, 39.0 grams O, and 46.3 grams F.

Find moles of C, O, and F in the sample:
\[ \frac{14.6 \text{ g C}}{12.0107 \text{ g C}} = 1.22 \text{ mol C} \]
\[ \frac{39.0 \text{ g O}}{15.9994 \text{ g O}} = 2.44 \text{ mol O} \]
\[ \frac{46.3 \text{ g F}}{18.9984 \text{ g F}} = 2.44 \text{ mol F} \]

Mole ratio: \( 1.22 \text{ mol C} : 2.44 \text{ mol O} : 2.44 \text{ mol F} \)

Divide each term by the smallest number of moles, 1.22 mol

Atom ratio: \( 1 \text{ C} : 2 \text{ O} : 2 \text{ F} \)

The empirical formula is \( \text{CO}_2\text{F}_2 \)

(b) The molecular formula is \( (\text{CO}_2\text{F}_2)_n \).

Mass of 1 mol \( \text{CO}_2\text{F}_2 = 12.0107 \text{ g} + 2 \times (15.9994 \text{ g}) + 2 \times (18.9984 \text{ g}) = 82.0063 \text{ g/mol} \)

\[ n = \frac{\text{mass of 1 mol of molecule}}{\text{mass of 1 mol of CO}_2\text{F}_2} = \frac{82 \text{ g}}{82.0063 \text{ g}} = 1.00 = 1 \]
Chapter 3: Chemical Compounds

The molecular formula is \((\text{CO}_2\text{F}_2)_1\) = \(\text{CO}_2\text{F}_2\)

*Check your answer:* The simplest ratios of C, O, and F in the sample end up very close to whole number values, so the empirical formula makes sense. The molar mass is the same as the mass of one mole of the empirical formula, so the molecular formula \(\text{CO}_2\text{F}_2\) makes sense.

104. Define the problem: Given the percent by mass of carbon in a hydrocarbon compound and the molar mass, determine the empirical formula and the molecular formula.

Develop a plan: A hydrocarbon contains just carbon and hydrogen. Choose a convenient sample of \(\text{C}_x\text{H}_y\), such as 100.00 g. Using the percent by mass of carbon, determine percent by mass of hydrogen, and the number of grams of C and H in the sample. Use the molar masses of C and H as conversion factors to get the moles of C and H. Set up a mole ratio to determine the whole number \(x:y\) ratio for the empirical formula. Find the mass of 1 mol of the empirical formula. Divide the molar mass of the compound by the calculated empirical mass to get a whole number. Multiply all the subscripts in the empirical formula by this whole number.

Execute the plan:

(a) The compound is 93.71 % C and the rest is hydrogen. This means that 100.00 g of \(\text{C}_x\text{H}_y\) contains 93.71 grams C and the mass of hydrogen is calculated like this:

\[
100.00 \text{ g \text{C}_x\text{H}_y} \times \frac{1 \text{ mol C}}{12.0107 \text{ g C}} = 7.802 \text{ mol C}
\]

\[
6.29 \text{ g \text{H}_y} \times \frac{1 \text{ mol H}}{1.0079 \text{ g H}} = 6.24 \text{ mol H}
\]

Mole ratio: \(7.802 \text{ mol C} : 6.24 \text{ mol H}\)

Divide each term by the smallest number of moles, 6.24 mol

Atom ratio: \(1.25 \text{ C} : 1 \text{ H}\)

Multiply by 4 (since \(4 \times 1.25 = 5\)) to get the whole number ratio:

Atom ratio: \(5 \text{ C} : 4 \text{ H}\)

The empirical formula is \(\text{C}_5\text{H}_4\)

(b) The molecular formula is \((\text{C}_5\text{H}_4)_n\).

Mass of 1 mol \(\text{C}_5\text{H}_4\) = \(5 \times (12.0107 \text{ g}) + 4 \times (1.0079 \text{ g}) = 64.0851 \text{ g/mol}\)

\[
n = \frac{\text{mass of 1 mol of molecule}}{\text{mass of 1 mol of \text{C}_5\text{H}_4}} = \frac{128.16 \text{ g}}{64.0851 \text{ g}} = 1.9998 = 2
\]

Molecular Formula is \((\text{C}_5\text{H}_4)_2 = \text{C}_{10}\text{H}_8\)

*Check your answers:* The simplest ratio of C and H in the sample ends up very close to whole number values, so the empirical formula makes sense. The molar mass is very close to double the mass of one mole of the empirical formula, so the molecular formula \(\text{C}_{10}\text{H}_8\) makes sense.

105. Define the problem: Given the percent by mass of all the elements in a compound, determine the empirical formula.
Develop a plan: Choose a convenient sample of $\text{Mn}_x\text{C}_y\text{H}_z\text{O}_v$, such as 100.0 g. Using the percent by mass, determine the number of grams of C, H, O, and Mn in the sample. Use the molar masses of C, H, O, and Mn as conversion factors to get the moles of C, H, O, and Mn. Set up a mole ratio to determine the whole number $x:y:z:v$ ratio for the empirical formula.

Execute the plan:

The compound is 49.5 % C, 3.2 % H, 22.0 % O, and 25.2 % Mn by mass. This means that 100.0 g of $\text{Mn}_x\text{C}_y\text{H}_z\text{O}_v$ contains 49.5 grams C, 3.2 grams H, 22.0 grams O, and 25.2 grams Mn.

Find moles of C, H, O, and Mn in the sample:

\[
\begin{align*}
49.5 \text{ g C} \times \frac{1 \text{ mol C}}{12.0107 \text{ g C}} &= 4.12 \text{ mol C} \\
3.2 \text{ g H} \times \frac{1 \text{ mol H}}{1.0079 \text{ g H}} &= 3.2 \text{ mol H} \\
22.0 \text{ g O} \times \frac{1 \text{ mol O}}{15.9994 \text{ g O}} &= 1.38 \text{ mol O} \\
25.2 \text{ g Mn} \times \frac{1 \text{ mol Mn}}{54.938 \text{ g Mn}} &= 0.459 \text{ mol Mn}
\end{align*}
\]

Mole ratio: $0.459 \text{ mol Mn} : 4.12 \text{ mol C} : 3.2 \text{ mol H} : 1.38 \text{ mol O}$

Divide each term by the smallest number of moles, 0.459 mol, and round to whole numbers:

\[1 \text{ Mn} : 9 \text{ C} : 7 \text{ H} : 3 \text{ O}\]

The empirical formula is $\text{MnC}_9\text{H}_7\text{O}_3$.

Check your answer: The simplest ratios of Mn, C, H, and O in the sample end up very close to whole number values, so the empirical formula makes sense.

106. Define the problem: Given the mass of an element used to make a mass of a compound and the molar mass of the compound, determine the empirical formula and the molecular formula of the compound.

Develop a plan: Use the mass of the compound, $I_x\text{Cl}_y$, and the mass of $I_2$ to determine the mass of Cl used. Use the molar masses of $I_2$ and Cl and the stoichiometry of their formulas to determine the moles of I and moles of Cl in the sample of the compound. Set up a mole ratio to determine the whole number $x:y$ ratio for the empirical formula. Find the mass of 1 mol of the empirical formula. Divide the molar mass of the compound by the calculated empirical mass to get a whole number. Multiply all the subscripts in the empirical formula by this whole number.

Execute the plan:

(a) The $I_x\text{Cl}_y$ sample has a mass of 1.246 grams. This sample was produced using 0.678 grams of iodine. The rest of the mass is from chlorine.

\[
1.246 \text{ g } I_x\text{Cl}_y - 0.678 \text{ g iodine} = 0.568 \text{ g chlorine}
\]

Molar mass of $I_2$ = $2 \times (126.90 \text{ g}) = 253.80 \text{ g/mol}$

Molar mass of Cl$_2$ = $2 \times (35.453 \text{ g}) = 70.906 \text{ g/mol}$

Find moles of I and Cl in the sample:

\[
0.678 \text{ g } I_2 \times \frac{1 \text{ mol } I_2}{253.80 \text{ g } I_2} \times \frac{2 \text{ mol I}}{1 \text{ mol } I_2} = 5.34 \times 10^{-3} \text{ mol I}
\]
0.568 g Cl₂ \times \frac{1 \text{ mol Cl}_2}{70.906 \text{ g Cl}_2} \times 2 \text{ mol Cl} = 1.60 \times 10^{-2} \text{ mol Cl}

Mole ratio: 5.34\times10^{-3} \text{ mol I} : 1.60\times10^{-2} \text{ mol Cl}

Divide each term by the smallest number of moles, 5.34\times10^{-3} \text{ mol}, and round to whole numbers:

1 I : 3 Cl

The empirical formula is ICl₃

(b) The molecular formula is (ICl₃)ₙ.

Mass of 1 mol ICl₃ = 126.90 g + 3 \times (35.453 g) = 233.26 g/mol

\[ n = \frac{\text{mass of 1 mol of molecule}}{\text{mass of 1 mol of ICl₃}} = \frac{467 \text{ g}}{233.26 \text{ g}} = 2.00 = 2 \]

Molecular Formula is (ICl₃)₂ = I₂Cl₆

Check your answer: The simplest ratio of I and Cl in the sample ends up very close to whole number values, so the empirical formula makes sense. The molar mass is twice the mass of one mole of the empirical formula, so the molecular formula I₂Cl₆ makes sense.

107. Define the problem: Given the number of tablets consumed, the mass of a compound in each tablet, and the formula of the compound, determine the moles of the compound consumed and the mass of one element consumed.

Develop a plan: Always start with the sample – in this case, the number of tablets consumed. We assume that C₇H₅BiO₄ is the “active ingredient”. Use the mass of C₇H₅BiO₄ per tablet to determine the mass of C₇H₅BiO₄ in the sample. Then use the molar mass of C₇H₅BiO₄ to determine the moles of C₇H₅BiO₄ in the sample. Then use the formula stoichiometry to get the moles of Bi in the sample. Then use the molar mass of Bi to get the grams of Bi in the sample.

Execute the plan: The sample is composed of two tablets of Pepto-Bismol. Find grams of C₇H₅BiO₄:

\[ 2 \text{ tablets} \times \frac{300. \text{ mg C₇H₅BiO₄}}{1 \text{ tablet}} \times \frac{1 \text{ g C₇H₅BiO₄}}{1000 \text{ mg C₇H₅BiO₄}} = 0.600 \text{ g C₇H₅BiO₄} \]

Molar mass of C₇H₅BiO₄

\[ = 7 \times (12.0107 \text{ g}) + 5 \times (1.0079 \text{ g}) + 208.9804 \text{ g} + 4 \times (15.9994 \text{ g}) = 362.0924 \text{ g/mol} \]

(a) Find moles of C₇H₅BiO₄ in the sample

\[ 0.600 \text{ g C₇H₅BiO₄} \times \frac{1 \text{ mol C₇H₅BiO₄}}{362.0924 \text{ g C₇H₅BiO₄}} = 1.66 \times 10^{-3} \text{ mol C₇H₅BiO₄} \]

(b) Find grams of Bi in the sample

\[ 1.66 \times 10^{-3} \text{ mol C₇H₅BiO₄} \times \frac{1 \text{ mol Bi}}{1 \text{ mol C₇H₅BiO₄}} \times \frac{208.9804 \text{ g Bi}}{1 \text{ mol Bi}} = 0.346 \text{ g Bi} \]

Check your answers: (a) The sample is somewhere between macroscale and microscale, so it makes sense that the number of moles is somewhat small. (b) The bismuth is almost 60 % of the C₇H₅BiO₄ compound mass so it makes sense that the number of grams of Bi in the sample is about 60 % of the mass.
of the compound in the sample.

108. **Define the problem:** Given the mass of an iron pyrite sample in kilograms and the formula of iron pyrite, determine the mass of one element in the sample.

**Develop a plan:** Always start with the sample – in this case, the kilograms of the FeS\(_2\). Use metric conversions to determine the sample mass in grams. Then use the molar mass of FeS\(_2\) to determine the moles of FeS\(_2\) in the sample. Then use the formula stoichiometry to get the moles of Fe in the sample. Then use the molar mass of Fe to get the grams of Fe in the sample. Use metric conversions to determine the Fe mass in kilograms.

**Execute the plan:** The sample is composed of 15.8 kg of FeS\(_2\).

Molar mass of FeS\(_2\) = 55.845 g + 2 × (32.065 g) = 119.975 g/mol

Find grams of Fe in the sample:

\[
15.8 \text{ kg } \text{FeS}_2 \times \frac{1000 \text{ g } \text{FeS}_2}{1 \text{ kg } \text{FeS}_2} \times \frac{1 \text{ mol } \text{FeS}_2}{119.975 \text{ g } \text{FeS}_2} \times \frac{1 \text{ mol } \text{Fe}}{1 \text{ mol } \text{FeS}_2} \times \frac{55.847 \text{ g } \text{Fe}}{1 \text{ mol } \text{Fe}} \times \frac{1 \text{ kg } \text{Fe}}{1000 \text{ g } \text{Fe}} = 7.35 \text{ kg } \text{Fe}
\]

Notice that the molar mass and stoichiometry can be interpreted in kilograms and kilomoles, removing the redundant metric conversions and making the problem shorter:

\[
15.8 \text{ kg } \text{FeS}_2 \times \frac{1 \text{ kmol } \text{FeS}_2}{119.975 \text{ kg } \text{FeS}_2} \times \frac{1 \text{ kmol } \text{Fe}}{1 \text{ kmol } \text{FeS}_2} \times \frac{55.847 \text{ kg } \text{Fe}}{1 \text{ kmol } \text{Fe}} = 7.35 \text{ kg } \text{Fe}
\]

**Check your answer:** The mass percentage of Fe in FeS\(_2\) is around 50%. So, it makes sense that the mass of Fe in the sample is about half of the sample mass. This number looks right.

109. **Define the problem:** Given the mass of a sample of an ore, the mass percent of titanium in the ore, the fact that the ore contains a compound ilmenite, and the formula of ilmenite (FeTiO\(_3\)), determine the mass of FeTiO\(_3\) in the sample of ore.

**Develop a plan:** Always start with the sample – in this case, 1.00 metric ton of ore. Use the given relationship and metric conversions to determine the sample mass in grams. Then use the mass percent of Ti as a conversion factor to determine the mass of Ti in the sample. Then use the molar mass of Ti to find the moles of Ti in the sample. Then use the formula stoichiometry of ilmenite to determine the moles of ilmenite, FeTiO\(_3\). Then use the molar mass of FeTiO\(_3\) to get the grams of FeTiO\(_3\) in the sample.

**Execute the plan:**

The ore is 6.75% Ti. That means, 100.00 grams of ore contains 6.75 grams Ti.

Formula Stoichiometry: 1 mol Ti is contained in 1 mol FeTiO\(_3\).

Molar mass of FeTiO\(_3\) = 55.845 g + 47.867 g + 3 × (15.9994 g) = 151.710 g/mol

\[
1.00 \text{ metric ton ore} \times \frac{1000 \text{ kg ore}}{1 \text{ metric ton ore}} \times \frac{1000 \text{ g ore}}{1 \text{ kg ore}} \times \frac{6.75 \text{ g Ti}}{100 \text{ g ore}} \times \frac{1 \text{ mol Ti}}{47.867 \text{ g Ti}} \times \frac{1 \text{ mol FeTiO}_3}{1 \text{ mol Ti}} \times \frac{151.710 \text{ g FeTiO}_3}{1 \text{ mol FeTiO}_3} = 2.14 \times 10^5 \text{ g FeTiO}_3
\]

**Check your answer:** Notice that the molar mass and stoichiometry can be interpreted with a larger scale.
100.00 metric ton of ore contains 6.75 metric ton Ti.

Formula Stoichiometry: 1 megamol Ti is contained in 1 megamol FeTiO$_3$.

Molar mass of FeTiO$_3$ = 55.845 g + 47.867 g + 3 x (15.9994 g) 
= 151.710 g/mol = \frac{151.710 \text{ g} \times (1,000,000)}{1 \text{ mol} \times (1,000,000)} = 151.710 \frac{\text{ metric ton}}{1 \text{ Mmol}}

\frac{1.00 \text{ metric ton ore} \times \frac{6.75 \text{ metric ton Ti}}{100 \text{ metric ton ore}} \times \frac{1 \text{ Mmol Ti}}{47.867 \text{ metric ton Ti}} \times \frac{1 \text{ Mmol FeTiO}_3}{1 \text{ Mmol Ti}}}{1 \text{ Mmol FeTiO}_3} = 0.214 \text{ metric ton FeTiO}_3

Ti represents about one third of the mass of FeTiO$_3$. This mass of FeTiO$_3$ in the ore is about three times more than the % Ti in the ore. This number looks right.

110. Define the problem: Given the mass of a sample of an ore, the mass percent of antimony in the ore, the fact that the ore contains the compound stibnite, and the formula of stibnite (Sb$_2$S$_3$), determine the mass in grams of Sb$_2$S$_3$ in the sample of ore.

Develop a plan: Always start with the sample – in this case, 1.00 pound of ore. Use English to metric conversion factor to determine the sample mass in grams. Then use the mass percent of Sb as a conversion factor to determine the mass of Sb in the sample. Then use the molar mass of Sb to find the moles of Sb in the sample. Then use the formula stoichiometry of stibnite to determine the moles of stibnite, Sb$_2$S$_3$. Then use the molar mass of Sb$_2$S$_3$ to get the grams of Sb$_2$S$_3$ in the sample.

Execute the plan: It is not clear how precisely the sample’s mass is measured. Mass measurements are pretty easy to do precisely, so we’ll use three significant figures, the limit of the other given information.

The ore is 10.6 % Sb. That means, 100.0 grams of ore contains 10.6 grams Sb.

Formula Stoichiometry: 2 mol Sb is contained in 1 mol Sb$_2$S$_3$.

Molar mass of Sb$_2$S$_3$ = 2 x (121.760 g) + 3 x (32.065 g) = 339.715 g/mol

\frac{1.00 \text{ lb ore} \times \frac{453.59 \text{ g ore}}{1 \text{ lb ore}} \times \frac{10.6 \text{ g Sb}}{100.0 \text{ g ore}} \times \frac{1 \text{ mol Sb}}{121.760 \text{ g Sb}} \times \frac{1 \text{ mol Sb$_2$S$_3$}}{2 \text{ mol Sb}}}{1 \text{ mol Sb$_2$S$_3$}} = 67.1 \text{ g Sb$_2$S$_3$}

Check your answer: Sb represents about \frac{2}{3} of the mass of Sb$_2$S$_3$. This calculated mass of Sb$_2$S$_3$ is about 0.15 pounds, or 15 % of the ore sample mass. Since the %Sb in the ore is 10 %, then approximately \frac{3}{2}(10 \%) of the ore will be Sb$_2$S$_3$. This number looks right.

111. Liquid bromine: Br$_2$ (ℓ)  
Solid lithium fluoride: LiF (s)
Chapter 3: Chemical Compounds
112. (a) A crystal of sodium chloride has an alternating lattice of Na\(^+\) and Cl\(^-\) ions. (b) The sodium chloride after it is melted has paired ions randomly distributed.

(c) Molten Al\(_2\)O\(_3\) has clusters of cations and anions randomly distributed.

113. (a) Solid lithium nitrate has an alternating lattice of Li\(^+\) and NO\(_3\)\(^-\) ions. (b) Molten lithium nitrate has ion pairs randomly distributed.

(c) Molten lithium nitrate when positive and negative electrodes are present will have the anions, NO\(_3\)\(^-\), (white circles) crowding around the positive electrode and the cations, Li\(^+\), (the black circles) crowding around the negative electrode:
114. There are three isomers given for C₃H₆O. Each of these isomers is represented by two of the structures. The following pairs are identical:

Isomer number one:
\[
\text{CH}_3-\text{CH}_2-\text{CH}_2-\text{OH} \\
\text{HO} - \text{CH}_2 - \text{CH}_2 \\
\text{CH}_3
\]

Isomer number two:
\[
\text{CH}_3-\text{CH} - \text{CH}_3 \\
\text{OH} \\
\text{HO} - \text{CH} - \text{CH}_3 \\
\text{CH}_3
\]

Isomer number three:
\[
\text{CH}_3-\text{O} - \text{CH}_2 - \text{CH}_3 \\
\text{CH}_3-\text{CH}_2-\text{O} - \text{CH}_3
\]

115. “Metals form positive ions by losing electrons.” Remember that a neutral atom has equal numbers of positive charges and negative charges. When some electrons are removed, we are reducing the number of negative charges. Now, there are more protons than there are electrons and the overall charge is positive as a result.

116. Thallium nitrate is TlNO₃. Since NO₃⁻ has a –1 charge, that means that thallium ion has a +1 charge, and is represented by Tl⁺. The carbonate compound containing thallium will be a combination of Tl⁺ and CO₃²⁻, and the compound’s formula will look like this: Tl₂CO₃. The sulfate compound containing thallium will be a combination of Tl⁺ and SO₄²⁻, and the compound’s formula will look like this: Tl₂SO₄.

117. (a) CaF₂ is calcium fluoride. Don’t use the “di-” prefix when naming ionic compounds.
(b) CuO is copper(II) oxide. The transition elements have several ions with different charges, so the specific valence is described by a Roman numeral indicating the cation’s charge. Since oxide ion has the formula O²⁻, the copper ion present is, Cu²⁺, the copper(II) ion.
(c) $\text{NaNO}_3$ is sodium nitrate. The incorrect name inappropriately used the naming system for binary molecular compounds, and this compound has more than two elements. It needs to be named using the ionic compound naming system with the name of the common cation ($\text{Na}^+$, sodium) and anion ($\text{NO}_3^-$, nitrate).

(d) $\text{NI}_3$ is a binary molecular compound. It contains only two elements. The proper name is nitrogen triiodide.

(e) $\text{FeCl}_3$ is iron(III) chloride. The Roman numeral indicates the charge of the iron cation. Since chloride ion is $\text{Cl}^-$, and three of them are in this neutral compound, that means the iron ion present is $\text{Fe}^{3+}$, which is called iron(III) ion, not iron(I).

(f) $\text{Li}_2\text{SO}_4$ is lithium sulfate. The incorrect name inappropriately used the naming system for binary molecular compounds, and this compound has more than two elements. It needs to be named using the ionic compound naming system with the name of the common cation ($\text{Li}^+$, lithium) and anion ($\text{SO}_4^{2-}$, sulfate).

118. (a) Based on the guidelines for naming oxyanions in a series, relate the chloro-oxyanions to the bromo-oxyanions, since chlorine and bromine are in the same group (Group 7A).

- $\text{BrO}_4^-$ is perbromate (like $\text{ClO}_4^-$ is perchlorate)
- $\text{BrO}_3^-$ is bromate (like $\text{ClO}_3^-$ is chlorate)
- $\text{BrO}_2^-$ is bromite (like $\text{ClO}_2^-$ is chlorite)
- $\text{BrO}^-$ is hypobromite (like $\text{ClO}^-$ is hypochlorite)

(b) Based on the guidelines for naming oxyanions in a series, relate the sulfur-oxyanions to the selenium-oxyanions.

- $\text{SeO}_4^{2-}$ is selenate (like $\text{SO}_4^{2-}$ is sulfate)
- $\text{SeO}_3^{2-}$ is selenite (like $\text{SO}_3^{2-}$ is sulfite)

119. Figure (f) represents the best representation of CaCl$_2$ dissolved in water. It shows the proper ions, Ca$^{2+}$ and Cl$^-$, floating around in the solution.

120. Define the problem: Given three samples, determine which has the largest quantity of NH$_3$.

Develop a plan: To compare these samples, put them in the same units. Convert all samples into moles.

Execute the plan:

(a) $6.022 \times 10^{24}$ molecules NH$_3 \times \frac{1 \text{ mol NH}_3}{6.022 \times 10^{23} \text{ molecules NH}_3} = 10.00 \text{ mol NH}_3$

(c) $17.03 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.030 \text{ g NH}_3} = 1.000 \text{ mol NH}_3$

Sample (a) has a larger amount of NH$_3$. (a) 10.00 mol > (c) 1.000 mol > (b) 0.10 mol.

Check your answer: The first sample is larger than Avogadro’s number. The third sample weighs the same as the molar mass. This order makes sense.
More Challenging Problems

121. Define the problem: Given the dimensions of a sample of nickel foil, the density of nickel metal, and the mass of a nickel fluoride compound \((\text{Ni}_x\text{F}_y)\) produced using that foil as a source of nickel, determine the moles of nickel, and the formula of the nickel fluoride compound.

Develop a plan: Always start with the sample – in this case, the dimensions of the Ni foil sample. Use metric conversion factors and the relationship between length, width, and height and volume to determine the sample’s volume. Then use the density to determine the mass of Ni in the sample. Subtract the mass of Ni from the mass of the compound produced to find the mass of F in the sample. Then use the molar masses of Ni and F to find the moles of Ni and moles of F in the sample. Then set up a mole ratio and find the simplest whole number relationship between the atoms.

Execute the plan: The foil is 0.550 mm thick, 1.25 cm long and 1.25 cm wide.

\[
V = \left( \frac{0.550 \text{ mm}}{1000 \text{ mm/m}} \right) \left( \frac{1 \text{ m}}{1 \text{ m}} \right) \times (1.25 \text{ cm}) \times (1.25 \text{ cm}) = 0.0859 \text{ cm}^3
\]

(a) \[
0.0859 \text{ cm}^3 \text{Ni} \times \frac{8.908 \text{ g Ni}}{1 \text{ cm}^3 \text{Ni}} \times \frac{1 \text{ mol Ni}}{58.69 \text{ g Ni}} = 0.0130 \text{ mol Ni}
\]

(b) \[
0.0859 \text{ cm}^3 \text{Ni} \times \frac{8.908 \text{ g Ni}}{1 \text{ cm}^3 \text{Ni}} = 0.765 \text{ g Ni}
\]

Mass of sulfur in the sample = \(\text{Ni}_x\text{S}_y\) compound mass – nickel mass

\[1.261 \text{ g Ni}_x\text{S}_y – 0.765 \text{ g Ni} = 0.496 \text{ g F}\]

Find moles of F in sample:

\[0.496 \text{ g F} \times \frac{1 \text{ mol F}}{18.9984 \text{ g F}} = 0.0261 \text{ mol F}\]

Set up mole ratio: \(0.0130 \text{ mol Ni} : 0.0261 \text{ mol F}\)

Divide all the numbers in the ratio by the smallest number of moles, 0.0130 mol, and round to whole numbers: \(1 \text{ Ni} : 2 \text{ F}\)

Empirical formula: \(\text{NiF}_2\)

(c) \(\text{NiF}_2\) is called nickel(II) fluoride. (It looks like it is composed of nickel(II) ion, \(\text{Ni}^{2+}\), and fluoride ion, \(\text{F}^-\)

Check your answers: The mole ratio was very close to a whole number ratio, so the empirical formula of \(\text{NiF}_2\) makes sense. One of the common ionic forms of nickel is nickel(II) ion, so this also confirms that the result is sensible.
118. (a) Define the problem: Given the mass a sample of a uranium oxide and the mass of uranium metal in the sample, determine the moles of uranium, the formula of the uranium oxide compound, and the moles of uranium oxide produced.

Develop a plan: The sample is the mass of $U_xO_y$. Subtract the mass of $U$ from the mass of $U_xO_y$ to find the moles of $O$ in the sample. Then use the molar masses of $U$ and $O$ to find the moles of $U$ and moles of $O$ in the sample. Then set up a mole ratio and find the simplest whole number relationship between the atoms.

Execute the plan: Mass of $O$ in the sample = $U_xO_y$ compound mass – $U$ mass

Notice that the number of sig. fig's in the oxygen mass has dropped to two after this subtraction.

Find moles of $U$ and $O$ in sample:

$$0.169 \text{ g } U \times \frac{1 \text{ mol } U}{238.03 \text{ g } U} = 7.10 \times 10^{-4} \text{ mol } U$$

$$0.030 \text{ g } O \times \frac{1 \text{ mol } O}{15.9994 \text{ g } O} = 1.9 \times 10^{-3} \text{ mol } O$$

Set up mole ratio: $7.10 \times 10^{-4} \text{ mol } U : 1.9 \times 10^{-3} \text{ mol } O$

Divide all the numbers in the ratio by the smallest number of moles, $7.10 \times 10^{-4}$ mol:

$1 \text{ U : 2.6 O}$

2.6 is close to both 2.66 and 2.5, and since this number has an uncertainty of $\pm 0.1$, either one is equally valid. It is unclear whether we should round down or up to find the whole number ratio. So, look at each case.

If the real ratio is $\frac{2}{3}$, we multiply by 3, and get $3 \text{ U : 8 O}$ and an empirical formula of $U_3O_8$.

If the real ratio is $\frac{1}{2}$, we multiply by 2, and get $2 \text{ U : 5 O}$ and an empirical formula of $U_2O_5$.

Check your answer: Of the two empirical formulas of $U_3O_8$ and $U_2O_5$, the second formula makes more sense. The common ion of oxy gen is oxide ion, $O^{2-}$. $U_2O_5$ looks like the simple combination of $U^{5+}$ and $O^{2-}$. The $U_3O_8$ formula is either wrong, or contains some mixture of uranium oxides with different formulas, such as a mixture of $UO_3$ and $U_2O_5$. With this insight, let’s presume that the $U_2O_5$ formula is the right formula. $U_2O_5$ would be called uranium (V) oxide.

Molar mass of $U_2O_5 = 2 \times (238.03 \text{ g}) + 5 \times (15.9994 \text{ g}) = 556.06 \text{ g/mol}$

$$0.199 \text{ g } U_2O_5 \times \frac{1 \text{ mol } U_2O_5}{556.06 \text{ g } U_2O_5} = 3.58 \times 10^{-4} \text{ mol } U_2O_5$$

(b) Define the problem: Given the mass of a sample of a uranium oxide hydrate and the mass of uranium oxide after it is dehydrated, determine the number of molecules of water associated with the hydrate.

Develop a plan: Subtract the mass of the dehydrated compound, $UO_2(NO_3)$, from the mass of the
hydrate, UO$_2$(NO$_3$)$_n$H$_2$O, to get the mass of water. Use the molar mass of water to get the moles of water. Use the molar mass of the UO$_2$(NO$_3$) and its relationship to the hydrated compound to find the number of moles of hydrate. Then set up a mole ratio between moles of water and moles of hydrate to find out how many molecules of water are associated with one hydrate formula.

**Execute the plan:**

Mass of H$_2$O in the sample
\[ \text{Mass of H}_2\text{O in the sample} = 0.865 \text{ g UO}_2(\text{NO}_3)\cdot n\text{H}_2\text{O} - 0.679 \text{ g UO}_2(\text{NO}_3) = 0.186 \text{ g H}_2\text{O} \]

Molar mass of H$_2$O = 2 \times (1.0079 \text{ g}) + 15.9994 \text{ g} = 18.0152 \text{ g/mol}

Molar mass of UO$_2$(NO$_3$) = 238.03 \text{ g} + 5 \times (15.9994 \text{ g}) + 14.0067 \text{ g} = 332.03 \text{ g/mol}

Moles of water in sample:
\[ 0.186 \text{ g H}_2\text{O}\times \frac{1 \text{ mol H}_2\text{O}}{18.0152 \text{ g H}_2\text{O}} = 0.0103 \text{ mol H}_2\text{O} \]

Moles of UO$_2$(NO$_3$)$_n$H$_2$O in sample:
\[ 0.679 \text{ g UO}_2(\text{NO}_3)\times \frac{1 \text{ mol UO}_2(\text{NO}_3) \cdot n\text{H}_2\text{O}}{332.03 \text{ g UO}_2(\text{NO}_3)}\times \frac{1 \text{ mol UO}_2(\text{NO}_3) \cdot n\text{H}_2\text{O}}{1 \text{ mol UO}_2(\text{NO}_3)} = 0.00204 \text{ mol UO}_2(\text{NO}_3)\cdot n\text{H}_2\text{O} \]

Mole ratio: 0.0103 mol H$_2$O : 0.00204 mol UO$_2$(NO$_3$)$_n$H$_2$O

Divide by the smallest number of moles, 0.00204 mol, and round to whole numbers:

5 H$_2$O : 1 UO$_2$(NO$_3$)$_n$H$_2$O

That means n = 5, the formula is UO$_2$(NO$_3$)$_5$H$_2$O, and there are 5 molecules of water of hydration in the original hydrated compound.

**Check your answer:** The mole ratio is very close to a whole number.

123. Define the problem: Given the mass of one molecule of a compound, the mass percentage of carbon, and the fact that the rest of the compound is oxygen, determine the identity of the compound.

**Develop a plan:** Pick a convenient sample size, such as 100.0 g of compound, C$_x$O$_y$. Use the percent by mass of carbon in the compound to determine the mass of carbon in the sample. Subtract the mass of carbon from the mass of the compound to find the mass of oxygen in the sample. Then use the molar masses of C and O to find the moles of C and moles of O in the sample. Then set up a mole ratio and find the simplest whole number relationship between the atoms. Use the mass of one molecule to determine the molar mass of the compound. Use the molar mass and the empirical formula to determine the molecular formula.

**Execute the plan:** The compound is 27.3 % C and the rest is O. That means that a 100.0-gram sample of the compound will contain 27.3 grams of carbon. The rest is oxygen:

\[ 100.0 \text{ g C}_x\text{O}_y - 27.3 \text{ g C} = 72.7 \text{ g O} \]

Determine the moles of C and O in the sample.
27.3 g C × \( \frac{1 \text{ mol C}}{12.0107 \text{ g C}} \) = 2.27 mol C

72.7 g O × \( \frac{1 \text{ mol O}}{15.9994 \text{ g O}} \) = 4.54 mol O

Set up mole ratio: 2.27 mol C : 4.54 mol O

Divide all the numbers in the ratio by the smallest number of moles, 2.27 mol, and round to whole numbers:

1 C : 2 O

Empirical formula: CO$_2$

Molecular formula: \((\text{CO}_2)_n\)

Mass of 1 mol CO$_2$ = 12.0107 g + 2 × (15.9994 g) = 44.010 g/mol

Molar mass of compound:

\[
\frac{7.308 \times 10^{-23} \text{ g}}{1 \text{ molecule}} \times \frac{6.022 \times 10^{23} \text{ molecule}}{1 \text{ mol}} = 44.0095 \text{ g mol}^{-1}
\]

Set up a ratio to find n:

\[
n = \frac{\text{mass of 1 mol compound}}{\text{mass of 1 mol empirical formula}} = \frac{44.01 \text{ g compound}}{44.0095 \text{ g CO}_2} = 1.000 \cong 1
\]

Molecular formula is \((\text{CO}_2)_1\) or \(\text{CO}_2\)

Check your answer: The mole ratio is very close to a whole number, and we recognize this compound as carbon dioxide, a common compound of carbon and oxygen. This makes sense.

124. Define the problem: Given the mass percentage of two elements in a compound, and the fact that the rest of the compound is a third element, determine the empirical formula of the compound.

Develop a plan: Pick a convenient sample size, such as 100.0 g of compound, \(\text{Co}_x\text{Mo}_y\text{Cl}_z\). Use the percent by mass of Co and Mo in the compound to determine the mass of chlorine in the sample. Subtract the mass of Co and Mo from the mass of the compound to find the mass of Cl in the sample. Then use the molar masses of Co, Mo, and Cl to find the moles of Co, Mo and Cl in the sample. Then set up a mole ratio and find the simplest whole number relationship between the atoms.

Execute the plan: Exactly 100.0 grams of the compound will contain 23.3 grams of Co and 25.3 grams of Mo. The rest is chlorine:

\[
100.0 \text{ g } \text{Co}_x\text{Mo}_y\text{Cl}_z - 23.3 \text{ g Co} - 25.3 \text{ g Mo} = 51.4 \text{ g Cl}
\]

Determine the moles of Co, Mo and Cl in the sample.

\[
23.3 \text{ g Co} \times \frac{1 \text{ mol Co}}{58.9332 \text{ g Co}} = 0.395 \text{ mol Co}
\]

\[
25.3 \text{ g Mo} \times \frac{1 \text{ mol Mo}}{95.94 \text{ g Mo}} = 0.264 \text{ mol Mo}
\]

\[
51.4 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.453 \text{ g Cl}} = 1.45 \text{ mol Cl}
\]

Set up mole ratio: 0.395 mol Co : 0.264 mol Mo : 1.45 mol Cl

Divide all the numbers in the ratio by the smallest number of moles, 0.264 mol

\[
1.50 \text{ Co} : 1.00 \text{ mol Mo} : 5.50 \text{ mol Cl}
\]

Multiply all the numbers in the ratio by 2, to resolve the fractions and get whole numbers:
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3 Co : 2 Mo : 11 Cl

Empirical formula: \( \text{Co}_3\text{Mo}_2\text{Cl}_{11} \)

*Check your answer:* The mole ratio is very close to a whole number, and we recognize this compound as carbon dioxide, a common compound of carbon and oxygen. This makes sense.

125. *Define the problem:* Given the mass percentage of two elements in a compound, and the fact that the rest of the compound is a third element, determine the empirical formula and name of the compound.

*Develop a plan:* Pick a convenient sample size, such as 100.0 g of compound, \( \text{Mg}_x\text{S}_y\text{O}_z \). Use the percent by mass of Mg and O in the compound to determine the mass of S in the sample. Subtract the mass of Mg and O from the mass of the compound to find the mass of S in the sample. Then use the molar masses of Mg, S, and O to find the moles of Mg, S and O in the sample. Then set up a mole ratio and find the simplest whole number relationship between the atoms.

*Execute the plan:* Exactly 100.0 grams of the compound will contain 23.3 grams of Mg and 46.0 grams of O. The rest is chlorine:

\[
100.0 \text{ g Mg}_x\text{S}_y\text{O}_z - 23.3 \text{ g Mg} - 46.0 \text{ g O} = 30.7 \text{ g S}
\]

Determine the moles of Mg, S and O in the sample.

\[
23.3 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.3050 \text{ g Mg}} = 0.959 \text{ mol Mg} \quad 30.7 \text{ g S} \times \frac{1 \text{ mol S}}{32.065 \text{ g S}} = 0.957 \text{ mol S}
\]

\[
46.0 \text{ g O} \times \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 2.88 \text{ mol O}
\]

Set up mole ratio: \( 0.959 \text{ mol Mg} : 0.957 \text{ mol S} : 2.88 \text{ mol O} \)

Divide all the numbers in the ratio by the smallest number of moles, 0.957 mol to get whole numbers

\( 1 \text{ Mg} : 1 \text{ S} : 3 \text{ O} \)

Empirical formula: \( \text{MgSO}_3 \) This is magnesium sulfite.

*Check your answer:* The mole ratio is very close to a whole number, and we recognize this compound as carbon dioxide, a common compound of carbon and oxygen. This makes sense.

126. *Define the problem:* For (a) and (b), given the name of a compound and the fact that it has an unknown amount of water associated as a hydrate and the mass percentage of water, determine the empirical formula of the compound. For (c), given the percents by mass of all the elements and water in a hydrate, determine the empirical formula of the compound.

*Develop a plan:* Set up a formula with the number of water molecules defined with the variable, \( n \).

Determine the mass of water in one mole of the hydrate, in terms of \( n \). Determine the molar mass of the compound, also in terms of \( n \). Use the definition of percent by mass of water in the compound to design an equation with one variable, \( n \). Solve for \( n \), algebraically.

*Execute the plan:* (a) The hydrate will have the formula \( \text{Fe(SCN)}_3\cdot n\text{H}_2\text{O} \); determine the value of \( n \).

Mass of water in one mole of \( \text{Fe(SCN)}_3\cdot n\text{H}_2\text{O} \) = \( n \times [1 \times (1.0079 \text{ g}) + (15.9994 \text{ g})] = 18.0152 \text{ n} \)

Molar mass \( \text{Fe(SCN)}_3\cdot n\text{H}_2\text{O} \)
\[= (55.845 \text{ g}) + 3 \times [(32.065 \text{ g}) + (12.0107 \text{ g}) + (14.0067 \text{ g})] + n \times [2 \times (1.0079 \text{ g}) + (15.9994 \text{ g})]\]
\[= 230.092 \text{ g} + 18.0152 n\]
%
\[\text{H}_2\text{O} = \frac{\text{mass of } \text{H}_2\text{O/ mol hydrate}}{\text{mass of hydrate/mol hydrate}} \times 100 \% = \frac{18.0152 n}{230.092 + 18.0152 n} \times 100 \% = 19.0 \%\]

Solve for \(n\):
\[18.0152 n = 0.190 \left(230.092 + 18.0152 n\right)\]
\[18.0152 n = 43.7175 + 3.4229 n\]
\[(18.0152 - 3.4229)n = 14.59 \quad n = 43.7175\]
\[n = 2.996 \approx 3 \quad \text{Empirical Formula: Fe(SCN)}_3\cdot3\text{H}_2\text{O}\]

(b) The hydrate will have the formula \(\text{ZnSO}_4\cdot n\text{H}_2\text{O}\); determine the value of \(n\).

Mass of water in one mole of \(\text{ZnSO}_4\cdot n\text{H}_2\text{O}\) = \(n \times [2 \times (1.0079 \text{ g}) + (15.9994 \text{ g})]\) = 18.0152 \(n\)

Molar mass \(\text{ZnSO}_4\cdot n\text{H}_2\text{O}\)
\[= (65.409 \text{ g}) + (32.065 \text{ g}) + 4 \times (15.9994 \text{ g}) + n \times [2 \times (1.0079 \text{ g}) + (15.9994 \text{ g})]\]
\[= 161.472 \text{ g} + 18.0152 n\]
%
\[\text{H}_2\text{O} = \frac{\text{mass of } \text{H}_2\text{O/ mol hydrate}}{\text{mass of hydrate/mol hydrate}} \times 100 \% = \frac{18.0152 n}{161.45 + 18.0152 n} \times 100 \% = 43.86 \%\]

Solve for \(n\):
\[18.0152 n = 0.4386 \left(161.472 + 18.0152 n\right)\]
\[18.0152 n = 70.8214 + 7.90147 n\]
\[(18.0152 - 7.90147)n = 70.8214\]
\[n = 7.0025 \approx 7 \quad \text{Empirical Formula: ZnSO}_4\cdot 7\text{H}_2\text{O}\]

(c) The hydrate compound is 12.10\% Na, 14.19\% Al, 22.14\% Si, 42.09\% O and 9.48\% \(\text{H}_2\text{O}\). That means that a 100.0-gram sample of the compound will contain 12.10 g Na, 14.19 g Al, 22.14 g Si, 42.09 g O and 9.48 g \(\text{H}_2\text{O}\).

Determine the moles of each in the sample.

downside up

Set up mole ratio: 0.5263 mol Na : 0.5259 mol Al : 0.7883 mol Si : 2.631 mol O : 0.526 mol \(\text{H}_2\text{O}\)

Divide all the numbers in the ratio by the smallest number of moles, 0.5263 moles:


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1 Na : 1 Al : 1.5 Si : 5 O : 1 H₂O

Multiply all the numbers in the ratio by 2, to resolve the fraction and get whole numbers:

2 Na : 2 Al : 3 Si : 10 O : 2 H₂O

Empirical formula: Na₂Al₂Si₃O₁₀·2H₂O

Check your answers: The mole ratios are all very close to whole numbers (or simple fractions, as in (c)).

127. Define the problem: Given the percent by mass of some elements and compounds in a fertilizer, determine the mass of three elements in a given mass of the fertilizer.

Develop a plan: Use the given mass and the definition of percent (parts per hundred) to determine the mass of each compound in the sample. Then determine the mass of each element in one mole of each of the compounds (as is done when determining mass percent of the element), and use that ratio to determine the mass of the elements in the sample.

Execute the plan: Assuming at least two significant figures precision, the fertilizer is 5.0% N, 10.0% P₂O₅, and 5.0% K₂O. That means, a sample of 100 lb of the fertilizer mixture, has 5.0 lb N, 10 lb P₂O₅, and 5.0 lb K₂O.

Mass of P in one mole of P₂O₅ = 2 × (30.9738 g) = 61.9476 g P

Molar mass P₂O₅ = 61.9476 g + 5 × (15.9994 g) = 141.9446 g P₂O₅

Mass of P in the fertilizer = 10 lb P₂O₅ × \frac{453.59237 g}{1 lb} × \frac{61.9476 g P}{141.9446 g P₂O₅} × \frac{1 lb}{453.59237 g} = 4.4 lb P

Mass of P in the fertilizer = 10 lb P₂O₅ × \frac{453.59237 g}{1 lb} × \frac{61.9476 g P}{141.9446 g P₂O₅} × \frac{1 lb}{453.59237 g} = 4.4 lb P

Mass of K in one mole of K₂O = 2 × (39.0983 g) = 78.1966 g K

Molar mass K₂O = 78.1966 g + 15.9994 g = 94.196 g K₂O

Mass of K in the fertilizer = 5.0 lb K₂O × \frac{453.59237 g}{1 lb} × \frac{78.1966 g K}{94.196 g K₂O} × \frac{1 lb}{453.59237 g} = 4.2 lb K

So, 100 lb of the fertilizer contains 5.0 lb N, 4.4 lb P, and 4.2 lb K

Check your answers: It makes sense that the percentage of P₂O₅ had to be higher to get approximately the same mass of P, since that compound has a significant mass of oxygen.

Conceptual Challenge Problems

CP-3.A The students should be asked to look at the numbers and contemplate what they can determine about them, before picking up a calculator. The three compounds have different formulas, as manifested most clearly by the %Ey differences. The second of these has a larger number of Ey atoms in the formula, due to its larger % mass; the third has the least Ey.

The best way to quantitatively compare these compounds is to scale the samples so that they have the same masses of one element. For example:

100.00 g of compound B has 40.002 g Ex, 6.7142 g Ey, and 53.284 g Ez. Scaling sample A by a factor of 1.0671 (calculated by dividing the %Exₐ/(%Eyₐ) shows that 106.71 grams of A has 40.002 g, Ex 13.427 g
Ey, and 53.284 g Ez. Scaling sample C by a factor of 0.983212 (calculated by dividing the \%Ex_B/\%Ey_C) shows that 98.3212 grams of C has 40.002 g Ex 5.0356 g Ey, and 53.283 g Ez.

A glance at these scaled masses shows that these three compounds have the same proportion of Ex to Ez. So the ratio of Ex atoms to Ez atoms in each of these formulas is the same. The ratio of Ey in each of them are: A/B = 2/1 and B/C = 1.333 = 4/3 Comparing this information, it is possible to determine that the Ey ratio in the three compounds is A:B:C = 8:4:3.

CP-3.B. Using what is known about the relationship between Ex and Ez in the given formula, it is possible to determine the coefficient for these two atoms in the other two formulas (since they must be the same). The 8:4:3 ratio found in the first problem allows for the relative determination of the Ey element in the other two formulas.

<table>
<thead>
<tr>
<th>Compound A</th>
<th>Compound B</th>
<th>Compound C</th>
</tr>
</thead>
<tbody>
<tr>
<td>ExEy$_4$Ez</td>
<td>ExEy$_2$Ez</td>
<td>ExEy$_3/2$Ez</td>
</tr>
<tr>
<td>Ex$_6$Ey$_8$Ez$_4$</td>
<td>Ex$_6$Ey$_4$Ez$_3$</td>
<td>Ex$_6$Ey$_3$Ez$_3$</td>
</tr>
<tr>
<td>Ex$_9$Ey$_4$Ez$_6$</td>
<td>Ex$_9$Ey$_2$Ez$_6$</td>
<td>Ex$_9$Ey$_3/2$Ez$_6$</td>
</tr>
<tr>
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<td>ExEy$_{8/3}$Ez$_3$</td>
<td>ExEy$_2$Ez$_3$</td>
</tr>
<tr>
<td>Ex$_5$Ey$_8$Ez$_3$</td>
<td>Ex$_5$Ey$_4$Ez$_3$</td>
<td>Ex$_5$Ey$_3$Ez$_3$</td>
</tr>
</tbody>
</table>

CP-3.B. (a) If the mass of Ez is 1.3320 times heavier than the mass of Ex, then the atoms must be present in equal proportion, since the mass ratio in the scaled samples is the same (54.284 g / 40.002 g). That means the formula must be Ex$_n$Ey$_8$Ez$_n$.

(b) If the mass of Ex is 11.916 times heavier than the mass of Ey, then there must be 1:2 atom ratio of Ex to Ey in compound B, where the mass ratio is 11.916/2.

(c) If the mass ratios are known, then the formulas can be determined.