$$\text{mass percent of S} = \frac{1 \times 32.07\ \cancel{\text{g/mol}}}{159.59\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 20.095\%$$
$$= 20.1\%$$

$$\text{mass percent of O} = \frac{4 \times 16.00\ \cancel{\text{g/mol}}}{159.59\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 40.1027\%$$
$$= 40.1\%$$

The sample is 39.8% copper, 20.1% sulfur, and 40.1% oxygen. This closely matches the given data. The empirical formula is reasonable.

34. Practice Problem (page 273)

Determine the empirical formula for the compound with the percentage composition of 26.61% K, 35.38% Cr, and 38.01% O.

What Is Required?

You need to determine the empirical formula for the sample that contains K (potassium), Cr (chromium), and O (oxygen).

What Is Given?

You know the composition of the compound: 26.61% K, 35.38% Cr, and 38.01% O.

Plan Your Strategy

Assume that the mass of the sample is 100.00 g.
Determine the atomic molar masses of potassium, chromium, and oxygen using the periodic table.
Convert each mass to amount in moles.
Determine the ratio of amounts of each element in whole numbers by dividing each mole amount by the lowest mole amount.

Act on Your Strategy

In a 100.00 g sample of the compound, there will be 26.61 g K, 35.38 g Cr, and 38.01 g oxygen.

From the periodic table:
The molar mass of potassium is 39.10 g/mol.
The molar mass of chromium is 52.00 g/mol.
The molar mass of oxygen is 16.00 g/mol.

Amount in moles, n, of each element:

$$n_K = \frac{26.61\,\cancel{g}}{39.10\,\cancel{g}/\text{mol}}$$
$$= 0.680562 \text{ mol}$$

$$n_{Cr} = \frac{35.38\,\cancel{g}}{52.00\,\cancel{g}/\text{mol}}$$
$$= 0.680384 \text{ mol}$$

$$n_O = \frac{38.01\,\cancel{g}}{16.00\,\cancel{g}/\text{mol}}$$
$$= 2.375625 \text{ mol}$$

Whole-number ratio:

$$\frac{0.680562}{0.680384} : \frac{0.680384}{0.680384} : \frac{2.375625}{0.680384}$$

Ratio: 1.0002:1.0000:3.4915 = 1:1:3.5

The least common multiple that will make this a whole-number ratio is 2. The smallest whole-number ratio of the elements is 2:2:7.

The empirical formula is $K_2Cr_2O_7$.

Check Your Solution
Work backward. Determine the percentage composition of $K_2Cr_2O_7$.
Molar mass, M, of $K_2Cr_2O_7$:

$$M_{K_2Cr_2O_7} = 2M_K + 2M_{Cr} + 7M_O$$
$$= 2(39.10 \text{ g/mol}) + 2(52.00 \text{ g/mol}) + 7(16.00 \text{ g/mol})$$
$$= 294.2 \text{ g/mol}$$

$$\text{mass percent of K} = \frac{2 \times 39.10\ \cancel{\text{g/mol}}}{294.2\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 26.5805\%$$
$$= 26.58\%$$

$$\text{mass percent of Cr} = \frac{2 \times 52.00\ \cancel{\text{g/mol}}}{294.2\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 35.3501\%$$
$$= 35.35\%$$

$$\text{mass percent of O} = \frac{7 \times 16.00\ \cancel{\text{g/mol}}}{294.2\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 38.0693\%$$
$$= 38.07\%$$

The sample is 26.58% potassium, 35.35% chromium, and 38.07% oxygen. This closely matches the given data. The empirical formula is reasonable.

35. Practice Problem (page 273)

Determine the empirical formula for the compound with the composition by mass of 17.6 g of hydrogen and 82.4 g of nitrogen.

What Is Required?

You need to determine the empirical formula from mass data for a sample that contains hydrogen and nitrogen.

What Is Given?

You know the mass data for the compound: 17.6 g hydrogen, 82.4 g nitrogen

Plan Your Strategy

Determine the atomic molar masses of hydrogen and nitrogen using the periodic table.
Convert each mass to amount in moles.
Determine the ratio of amounts of each element in whole numbers by dividing each mole amount by the lower mole amount.

Act on Your Strategy
From the periodic table:
The molar mass of nitrogen is 14.01 g/mol.
The molar mass of hydrogen is 1.01 g/mol.

Amount in moles, n, of each element:

$$n_N = \frac{82.4\ \cancel{g}}{14.01\ \cancel{g}/\text{mol}}$$
$$= 5.881513\ \text{mol}$$

$$n_H = \frac{17.6\ \cancel{g}}{1.01\ \cancel{g}/\text{mol}}$$
$$= 17.42574\ \text{mol}$$

Whole-number ratio:

$$\frac{5.881513}{5.881513} : \frac{17.42574}{5.881513}$$

Ratio: 1:2.962 = 1:3

The empirical formula is NH_3.

Check Your Solution
Work backward. Assume there is 100.0 g of the compound. Determine the percentage composition of NH_3.

Molar mass, M, of NH_3:

$$M_{NH_3} = 1M_N + 3M_H$$
$$= 1(14.01\ \text{g/mol}) + 3(1.01\ \text{g/mol})$$
$$= 17.04\ \text{g/mol}$$

$$\text{mass percent of N} = \frac{1 \times 14.01\ \cancel{\text{g/mol}}}{17.04\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 82.2183\%$$
$$= 82.2\%$$

$$\text{mass percent of H} = \frac{3 \times 1.01\ \cancel{\text{g/mol}}}{17.04\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 17.7816\%$$
$$= 17.8\%$$

A 100.0 g sample would contain 82.2 g of nitrogen and 17.8 g of hydrogen. This closely matches the given data. The empirical formula is reasonable.

36. Practice Problem (page 273)

Determine the empirical formula for the compound with the composition by mass of 46.3 g of lithium and 53.7 g of oxygen.

What Is Required?

You need to determine the empirical formula from mass data for a sample that contains lithium and oxygen.

What Is Given?

You know the mass data for the compound: 46.3 g lithium, 53.7 g oxygen

Plan Your Strategy

Determine the atomic molar masses of lithium and oxygen using the periodic table.
Convert each mass to amount in moles.
Determine the ratio of amounts of each element in whole numbers by dividing each mole amount by the lower mole amount.

Act on Your Strategy

From the periodic table:
The molar mass of lithium is 6.94 g/mol.
The molar mass of oxygen is 16.00 g/mol.

Amount in moles, n, of each element:

$$n_{\text{Li}} = \frac{43.6\ \cancel{\text{g}}}{6.94\ \cancel{\text{g}}/\text{mol}}$$
$$= 6.671469\ \text{mol}$$

$$n_{\text{O}} = \frac{53.7\ \cancel{\text{g}}}{16.00\ \cancel{\text{g}}/\text{mol}}$$
$$= 3.35625\ \text{mol}$$

Whole-number ratio:

$$\frac{6.671469}{3.35625} : \frac{3.35625}{3.35625}$$

Ratio: 1.9877:1 = 2:1

The empirical formula is Li_2O.

Check Your Solution

Work backward. Assume there is 100.0 g of the compound. Determine the percentage composition of Li_2O.

Molar mass, M, of Li_2O:

$$\begin{aligned} M_{Li_2O} &= 2M_{Li} + 1M_O \\ &= 2(6.94 \text{ g/mol}) + 1(16.00 \text{ g/mol}) \\ &= 29.88 \text{ g/mol} \end{aligned}$$

$$\begin{aligned} \text{mass percent of Li} &= \frac{2 \times 6.94 \cancel{\text{g/mol}}}{29.88 \cancel{\text{g/mol}}} \times 100\% \\ &= 46.452\% \\ &= 46.4\% \end{aligned}$$

$$\begin{aligned} \text{mass percent of O} &= \frac{1 \times 16.00 \cancel{\text{g/mol}}}{29.88 \cancel{\text{g/mol}}} \times 100\% \\ &= 53.5475\% \\ &= 53.5\% \end{aligned}$$

A 100.0 g sample would contain 46.4 g of lithium and 53.5 g of oxygen. This closely matches the given data. The masses do not add to 100 g because of rounding. The empirical formula is reasonable.

37. Practice Problem (page 273)

Determine the empirical formula for the compound with the percentage composition of 15.9% boron and the rest being fluorine.

What Is Required?

You need to determine the empirical formula for the sample that contains boron and fluorine.

What Is Given?
You know the composition of the compound: 15.9% boron and the remainder is fluorine.

Plan Your Strategy
Assume that the mass of the sample is 100.0 g.
Determine the mass of the fluorine by subtracting the mass of boron from 100.0 g.
Determine the atomic molar masses of boron and fluorine using the periodic table.
Convert each mass to amount in moles.
Determine the ratio of amounts of each element in whole numbers by dividing each mole amount by the lower mole amount.

Act on Your Strategy
Mass of fluorine:

$$\begin{aligned}\text{mass of fluorine} &= \text{mass of compound} - \text{mass of boron}\\ &= 100.0\text{ g} - 15.9\text{ g}\\ &= 84.1\text{ g}\end{aligned}$$

In a 100.0 g sample of the compound, there will be 15.9 g of boron and 84.1 g of fluorine.

From the periodic table:
The molar mass of boron is 10.81 g/mol.
The molar mass of fluorine is 19.00 g/mol.

Amount in moles, n, of each element:

$$\begin{aligned}n_B &= \frac{15.9\ \cancel{g}}{10.81\ \cancel{g}/\text{mol}}\\ &= 1.47086\text{ mol}\end{aligned}$$

$$\begin{aligned}n_F &= \frac{84.1\ \cancel{g}}{19.00\ \cancel{g}/\text{mol}}\\ &= 4.426315\text{ mol}\end{aligned}$$

Whole-number ratio:

$$\frac{1.47086}{1.47086} : \frac{4.426315}{1.47086}$$

Ratio: 1:3.00933 = 1:3

The empirical formula is BF_3.

Check Your Solution
Work backward. Determine the percentage composition of BF_3.
Molar mass, M, of BF_3:

$$\begin{aligned} M_{BF_3} &= 1M_B + 3M_F \\ &= 1(10.81\text{ g/mol}) + 3(19.00\text{ g/mol}) \\ &= 67.81\text{ g/mol} \end{aligned}$$

$$\begin{aligned} \text{mass percent of B} &= \frac{1 \times 10.81\ \cancel{\text{g/mol}}}{67.81\ \cancel{\text{g/mol}}} \times 100\% \\ &= 15.9416\% \\ &= 15.9\% \end{aligned}$$

$$\begin{aligned} \text{mass percent of F} &= \frac{3 \times 19.00\ \cancel{\text{g/mol}}}{67.81\ \cancel{\text{g/mol}}} \times 100\% \\ &= 84.0583\% \\ &= 84.0\% \end{aligned}$$

The sample is 15.9% boron and 84.0% fluorine. This closely matches the given data. The percentages do not add to 100% because of rounding. The empirical formula is reasonable.

38. Practice Problem (page 273)
Determine the empirical formula for the compound with the percentage composition of 60.11% sulfur and the rest being chlorine.

What Is Required?
You need to determine the empirical formula for the sample that contains sulfur and chlorine.

What Is Given?
You know the composition of the compound: 60.11% sulfur and the remainder is chlorine

Plan Your Strategy
Assume that the mass of the sample is 100.00 g.
Determine the mass of chlorine by subtracting the mass of sulfur from 100.00 g.
Determine the atomic molar masses of sulfur and chlorine using the periodic table.
Convert each mass to amount in moles.
Determine the ratio of amounts of each element in whole numbers by dividing each mole amount by the lower mole amount.

Act on Your Strategy
Mass of chlorine:

$$\begin{aligned}\text{mass of chlorine} &= \text{mass of compound} - \text{mass of sulfur}\\ &= 100.0\text{ g} - 60.11\text{ g}\\ &= 39.89\text{ g}\end{aligned}$$

In a 100.00 g sample of the compound, there will be 60.11 g of sulfur and 39.89 g of chlorine.

From the periodic table:
The molar mass of sulfur is 32.07 g/mol.
The molar mass of chlorine is 35.45 g/mol.

Amount in moles, n, of each element:

$$\begin{aligned}n_S &= \frac{60.11\ \cancel{g}}{32.07\ \cancel{g}/\text{mol}}\\ &= 1.87433\text{ mol}\end{aligned}$$

$$\begin{aligned}n_{Cl} &= \frac{39.89\ \cancel{g}}{35.45\ \cancel{g}/\text{mol}}\\ &= 1.12524\text{ mol}\end{aligned}$$

Whole-number ratio:

$$\frac{1.87433}{1.12524} : \frac{1.12524}{1.12524}$$

$$\begin{aligned}\text{Ratio: } 1.6657 : 1 &= 1\frac{2}{3} : 1\\ &= \frac{5}{3} : 1\end{aligned}$$

The least common multiple that will make this a whole-number ratio is 3. The smallest whole-number ratio of the elements is 5:3.

The empirical formula is Cl_3S_5.

Check Your Solution
Work backward. Determine the percentage composition of Cl_3S_5.
Molar mass, M, of Cl_3S_5:

$$\begin{aligned} M_{Cl_3S_5} &= 3M_{Cl} + 5M_S \\ &= 3(35.45\text{ g/mol}) + 5(32.07\text{ g/mol}) \\ &= 266.7\text{ g/mol} \end{aligned}$$

$$\begin{aligned} \text{mass percent of Cl} &= \frac{3 \times 35.45\ \cancel{\text{g/mol}}}{266.7\ \cancel{\text{g/mol}}} \times 100\% \\ &= 39.876\% \\ &= 39.88\% \end{aligned}$$

$$\begin{aligned} \text{mass percent of S} &= \frac{5 \times 32.07\ \cancel{\text{g/mol}}}{266.7\ \cancel{\text{g/mol}}} \times 100\% \\ &= 60.1237\% \\ &= 60.12\% \end{aligned}$$

The sample is 39.88% chlorine and 60.12% sulfur. This closely matches the given data. The empirical formula is correct.

39. Practice Problem (page 273)
Determine the empirical formula for the compound with the percentage composition of 11.33% carbon, 45.29% oxygen, and the rest being sodium.

What Is Required?
You need to determine the empirical formula for the sample that contains carbon, oxygen, and sodium.

What Is Given?
You know the composition of the compound: 11.33% carbon, 45.29% oxygen, and the rest is sodium.

Plan Your Strategy
Assume that the mass of the sample is 100.00 g.
Determine the mass of sodium by subtracting the sum of the percentage carbon and percentage oxygen from 100.00 g
Determine the atomic molar masses of sodium, carbon, and oxygen using the periodic table.
Convert each mass to amount in moles.
Determine the ratio of amounts of each element in whole numbers by dividing each mole amount by the lowest mole amount.

Act on Your Strategy
Mass of sodium:

$$\begin{aligned}\text{mass of sodium} &= \text{mass of compound} - (\text{mass of carbon} + \text{mass of oxygen})\\ &= 100.0\text{ g} - (11.33\text{ g} + 45.29\text{ g})\\ &= 43.38\text{ g}\end{aligned}$$

In a 100.00 g sample of the compound, there will be 43.38 g of sodium, 11.33 g of carbon, and 45.29 g of oxygen.

From the periodic table:
The molar mass of sodium is 22.99 g/mol.
The molar mass of carbon is 12.01 g/mol.
The molar mass of oxygen is 16.00 g/mol.

Amount in moles, n, of each element:

$$\begin{aligned}n_{\text{Na}} &= \frac{43.38\ \cancel{\text{g}}}{22.99\ \cancel{\text{g}}/\text{mol}}\\ &= 1.88690\text{ mol}\end{aligned}$$

$$\begin{aligned}n_{\text{C}} &= \frac{11.33\ \cancel{\text{g}}}{12.01\ \cancel{\text{g}}/\text{mol}}\\ &= 0.94338\text{ mol}\end{aligned}$$

$$\begin{aligned}n_{\text{O}} &= \frac{45.29\ \cancel{\text{g}}}{16.00\ \cancel{\text{g}}/\text{mol}}\\ &= 2.83062\text{ mol}\end{aligned}$$

Whole-number ratio:

$$\frac{1.88690}{0.94338} : \frac{0.94338}{0.94338} : \frac{2.83062}{0.94338}$$

Ratio: 2.0001:1.0000:3.00642 = 2:1:3

The empirical formula is Na_2CO_3.

Check Your Solution
Work backward. Determine the percentage composition of Na_2CO_3.

Molar mass, M, of Na_2CO_3:

$$\begin{aligned} M_{Na_2CO_3} &= 2M_{Na} + 1M_C + 3M_O \\ &= 2(22.99\text{ g/mol}) + 1(12.01\text{ g/mol}) + 3(16.00\text{ g/mol}) \\ &= 105.99\text{ g/mol} \end{aligned}$$

$$\begin{aligned} \text{mass percent of Na} &= \frac{2 \times 22.99\ \cancel{\text{g/mol}}}{105.99\ \cancel{\text{g/mol}}} \times 100\% \\ &= 43.38145\% \\ &= 43.38\% \end{aligned}$$

$$\begin{aligned} \text{mass percent of C} &= \frac{1 \times 12.01\ \cancel{\text{g/mol}}}{105.99\ \cancel{\text{g/mol}}} \times 100\% \\ &= 11.33125\% \\ &= 11.33\% \end{aligned}$$

$$\begin{aligned} \text{mass percent of O} &= \frac{3 \times 16.00\ \cancel{\text{g/mol}}}{105.99\ \cancel{\text{g/mol}}} \times 100\% \\ &= 45.2872\% \\ &= 43.29\% \end{aligned}$$

The sample is 43.38% sodium, 11.33% carbon, and 45.29% oxygen. This matches the given data. The empirical formula is correct.

40. Practice Problem (page 273)
Determine the empirical formula for the compound with the percentage composition of 56.36% oxygen, and the rest being phosphorus.

What Is Required?
You need to determine the empirical formula for the sample that contains oxygen and phosphorus.

What Is Given?
You know the composition of the compound: 56.36% oxygen and the remainder is phosphorus

Plan Your Strategy
Assume that the mass of the sample is 100.00 g.
Determine the mass of the phosphorus by subtracting the mass of oxygen from 100.00 g.
Determine the atomic molar masses of phosphorus and oxygen using the periodic table.
Convert each mass to amount in moles.
Determine the ratio of amounts of each element in whole numbers by dividing each mole amount by the lower mole amount.

Act on Your Strategy
Mass of phosphorus:

$$\begin{aligned}\text{mass of phosphorus} &= \text{mass of compound} - \text{mass of oxygen}\\ &= 100.0\text{ g} - 56.36\text{ g}\\ &= 43.64\text{ g}\end{aligned}$$

In a 100.00 g sample of the compound, there will be 43.64 g of phosphorus and 56.36 g of oxygen.

From the periodic table:
The molar mass of phosphorus is 30.97 g/mol.
The molar mass of oxygen is 16.00 g/mol.

Amount in moles, n, of each element:

$$\begin{aligned}n_P &= \frac{43.64\ \cancel{g}}{30.97\ \cancel{g}/\text{mol}}\\ &= 1.40910\text{ mol}\end{aligned}$$

$$\begin{aligned}n_O &= \frac{56.36\ \cancel{g}}{16.00\ \cancel{g}/\text{mol}}\\ &= 3.52250\text{ mol}\end{aligned}$$

Whole-number ratio:

$$\frac{1.40910}{1.40910} : \frac{3.52250}{1.40910}$$

Ratio: 1:2.4998 = 1:2.5

The least common multiple that will make this a whole-number ratio is 2. The smallest whole-number ratio of the elements is 2:5.

The empirical formula is P_2O_5.

Check Your Solution
Work backward. Determine the percentage composition of P_2O_5.
Molar mass, M, of P_2O_5:

$$\begin{aligned} M_{P_2O_5} &= 2M_P + 5M_O \\ &= 2(30.97\text{ g/mol}) + 5(16.00\text{ g/mol}) \\ &= 141.94\text{ g/mol} \end{aligned}$$

$$\begin{aligned} \text{mass percent of P} &= \frac{2 \times 30.97\ \cancel{\text{g/mol}}}{141.94\ \cancel{\text{g/mol}}} \times 100\% \\ &= 43.36381\% \\ &= 43.36\% \end{aligned}$$

$$\begin{aligned} \text{mass percent of O} &= \frac{5 \times 16.00\ \cancel{\text{g/mol}}}{141.94\ \cancel{\text{g/mol}}} \times 100\% \\ &= 56.3618\% \\ &= 56.36\% \end{aligned}$$

The sample is 43.36% phosphorus and 56.36% oxygen. This closely matches the given data. The percentages do not add to 100% because of rounding. The empirical formula is reasonable.

Section 6.2 Empirical and Molecular Formulas

Solutions for Practice Problems

Student Edition pages 275-276

41. Practice Problem (page 275)

The empirical formula for glucose is $CH_2O(s)$. The molar mass of glucose is 180.18 g/mol. Determine the molecular formula for glucose.

What Is Required

You need to determine the molecular formula for glucose.