Chapter 6

Chemical Proportions in Compounds

Solutions for Practice Problems

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1. Problem

A sample of a compound is analyzed and found to contain 0.90 g of calcium and 1.60 g of chlorine. The sample has a mass of 2.50 g. Find the percentage composition of the compound.

What Is Required?

You need to find the mass percents of calcium and chlorine in the compound.

What Is Given?

You know the mass of the compound, as well as the mass of each element in the compound.

Mass of compound = 2.50 g
Mass of Ca = 0.90 g
Mass of Cl = 1.60 g

Plan Your Strategy

To find the percentage composition of the compound, find the mass percent of each element. To do this, divide the mass of each element by the mass of the compound and multiply by 100%.

Act on Your Strategy

$$\text{Mass percent of Ca} = \frac{\text{Mass of Ca}}{\text{Mass of compound}} \times 100\%$$
$$= \frac{0.90\ \text{g}}{2.50\ \text{g}} \times 100\%$$
$$= 36\%$$
$$\text{Mass percent of Cl} = \frac{\text{Mass of Cl}}{\text{Mass of compound}} \times 100\%$$
$$= \frac{1.60\ \text{g}}{2.50\ \text{g}} \times 100\%$$
$$= 64\%$$

The percentage composition of the compound is 36% calcium and 64% chlorine.

Check Your Solution

The mass of calcium is 0.9 g per 2.50 g of the compound. This is roughly a little over one third of the mass of the compound, which is close to the calculated value of 36%.

2. Problem

Find the percentage composition of a pure substance that contains 7.22 g nickel, 2.53 g phosphorus, and 5.25 g oxygen only.

What Is Required?
You need to find the mass percents of nickel, phosphorus, and oxygen in the pure substance.

What Is Given?
You know the mass of each element in the compound.
Mass of Ni = 7.22 g
Mass of P = 2.53 g
Mass of O = 5.25 g

Plan Your Strategy
First calculate the mass of the pure substance. According to the Law of Conservation of Mass, the mass of any compound is the sum of the masses of its component elements. Then, to find the percentage composition of the compound, find the mass percent of each element. To do this, divide the mass of each element by the mass of the compound and multiply by 100%.

Act on Your Strategy
Mass of the pure substance = Mass of Ni + Mass of P + Mass of O
$$= 7.22\ \text{g} + 2.53\ \text{g} + 5.25\ \text{g}$$
$$= 15.0\ \text{g}$$
$$\text{Mass percent of Ni} = \frac{\text{Mass of Ni}}{\text{Mass of substance}} \times 100\%$$
$$= \frac{7.22\ \cancel{\text{g}}}{15.0\ \cancel{\text{g}}} \times 100\%$$
$$= 48.1\%$$
$$\text{Mass percent of P} = \frac{\text{Mass of P}}{\text{Mass of substance}} \times 100\%$$
$$= \frac{2.53\ \cancel{\text{g}}}{15.0\ \cancel{\text{g}}} \times 100\%$$
$$= 16.9\%$$
$$\text{Mass percent of O} = \frac{\text{Mass of O}}{\text{Mass of substance}} \times 100\%$$
$$= \frac{5.25\ \cancel{\text{g}}}{15.0\ \cancel{\text{g}}} \times 100\%$$
$$= 35.0\%$$
The percentage composition of the pure substance is 48.1% nickel, 16.9% phosphorus, and 35.0% oxygen.

Check Your Solution
The mass of nickel is 7.22 g per 15.0 g of the compound. This is roughly 50%, which is close to the calculated value of 48.1%.

3. Problem
A sample of a compound is analyzed and found to contain carbon, hydrogen, and oxygen. The mass of the sample is 650 mg, and the sample contains 257 mg of carbon and 50.4 mg of hydrogen. What is the percentage composition of the compound?

What Is Required?
You need to find the mass percents of carbon, hydrogen, and oxygen in the compound.

What Is Given?
You know the mass of the compound, as well as the mass of carbon and hydrogen in the compound.
Mass of compound = 650 mg
Mass of C = 257 mg

Mass of H = 50.4 mg

Plan Your Strategy

First calculate the mass of oxygen in the compound. According to the Law of Conservation of Mass, the mass of any compound is the sum of the masses of its component elements. To find the percentage composition of the compound, find the mass percent of each element. To do this, divide the mass of each element by the mass of the compound and multiply by 100%.

Act on Your Strategy

$$\text{Mass of O} = \text{Mass of compound} - (\text{Mass of C} + \text{Mass of H})$$
$$= 650 \text{ mg} - (257 \text{ mg} + 50.4 \text{ mg})$$
$$= 342.6 \text{ mg}$$

$$\text{Mass percent of C} = \frac{\text{Mass of C}}{\text{Mass of compound}} \times 100\%$$
$$= \frac{257 \cancel{\text{mg}}}{650 \cancel{\text{mg}}} \times 100\%$$
$$= 39.5\%$$

$$\text{Mass percent of H} = \frac{\text{Mass of H}}{\text{Mass of compound}} \times 100\%$$
$$= \frac{50.4 \cancel{\text{mg}}}{650 \cancel{\text{mg}}} \times 100\%$$
$$= 7.8\%$$

$$\text{Mass percent of O} = \frac{\text{Mass of O}}{\text{Mass of compound}} \times 100\%$$
$$= \frac{342.6 \cancel{\text{mg}}}{650 \cancel{\text{mg}}} \times 100\%$$
$$= 52.7\%$$

The percentage composition of the compound is 39.5% carbon, 7.8% hydrogen, and 52.7% oxygen.

Check Your Solution

The mass of oxygen is 342.6 mg per 650 mg of the compound. This is roughly 50%, which is close to the calculated value of 52.7%.

4. Problem

A scientist analyzes a 50.0 g sample and finds that it contains 13.3 g of potassium, 17.7 g of chromium, and another element. Later the scientist learns that the sample is potassium dichromate, $K_2Cr_2O_7$. Potassium dichromate is a bright orange compound that is used in the production of safety matches. What is the percentage composition of potassium dichromate?

What Is Required?

You need to find the mass percents of potassium, chromium, and oxygen in the potassium dichromate.

What Is Given?

You know the mass of potassium dichromate, as well as the mass of the potassium and chromium in the compound.

Mass of $K_2Cr_2O_7$ = 50.0 g
Mass of K = 13.3 g
Mass of Cr = 17.7 g

Plan Your Strategy

First calculate the mass of oxygen in the compound. According to the Law of Conservation of Mass, the mass of any compound is the sum of the masses of its component elements. To find the percentage composition of the compound, find the

mass percent of each element. To do this, divide the mass of each element by the mass of the compound and multiply by 100%.

Act on Your Strategy

Mass of O = Mass of $K_2Cr_2O_7$ − (Mass of K + Mass of Cr)

= 50.0 g − (13.3 g + 17.7 g)

= 19.0 g

$$\text{Mass percent of K} = \frac{\text{Mass of K}}{\text{Mass of compound}} \times 100\%$$

$$= \frac{13.3\ \text{g}}{50.0\ \text{g}} \times 100\%$$

$$= 26.6\%$$

$$\text{Mass percent of Cr} = \frac{\text{Mass of Cr}}{\text{Mass of compound}} \times 100\%$$

$$= \frac{17.7\ \text{g}}{50.0\ \text{g}} \times 100\%$$

$$= 35.4\%$$

$$\text{Mass percent of O} = \frac{\text{Mass of O}}{\text{Mass of compound}} \times 100\%$$

$$= \frac{19.0\ \text{g}}{50.0\ \text{g}} \times 100\%$$

$$= 38.0\%$$

The percentage composition of the compound is 26.6% potassium, 35.4% chromium, and 38.0% oxygen.

Check Your Solution

The mass of oxygen is 342.6 mg per 650 mg of the compound. This is roughly 50%, which is close to the calculated value of 52.7%.

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5. Problem

Calculate the mass percent of nitrogen in each compound.

(a) N_2O **(b)** $Sr(NO_3)_2$ **(c)** NH_4NO_3 **(d)** HNO_3

What Is Required?

You need to find the mass percent of nitrogen in the compounds listed.

What Is Given?

The molecular formula of each compound is given. This gives the total number of elements of each type in one mole of the compound.

(a) N_2O = 2 N and 1 O

(b) $Sr(NO_3)_2$ = 1 Sr, 2 N, and 6 O

(c) NH_4NO_3 = 2 N, 4 H, and 3 O

(d) HNO_3 = 1 H, 1 N, and 3 O

Plan Your Strategy

First calculate the molar mass of each compound. This is the sum of the molar masses of its component elements. To find the mass percent of nitrogen in the compound, divide the total molar mass of nitrogen in each compound by the molar mass of the compound, and multiply by 100%.

Act on Your Strategy

(a) Molar mass of N_2O = 2(Molar mass of N) + Molar mass of O

= 2(14.01 g/mol) + 16.00 g/mol

= 44.02 g/mol

Total molar mass of N in N_2O = 2(Molar mass of N)
= 2(14.01 g/mol)
= 28.02 g/mol

$$\text{Mass percent N in } N_2O = \frac{\text{Total molar mass of N}}{\text{Molar mass } N_2O} \times 100\%$$
$$= \frac{28.02\ \cancel{\text{g/mol}}}{44.02\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 63.65\%$$

(b) Molar mass of $Sr(NO_3)_2$
= Molar mass of Sr + 2(Molar mass of N) + 6(Molar mass of O)
= 87.62 g/mol + 2(14.01 g/mol) + 6(16.00 g/mol)
= 211.64 g/mol

Total molar mass of N in $Sr(NO_3)_2$ = 2(Molar mass of N)
= 2 (14.01 g/mol)
= 28.02 g/mol

$$\text{Mass percent N in } Sr(NO_3)_2 = \frac{\text{Total molar mass of N}}{\text{Molar mass } Sr(NO_3)_2} \times 100\%$$
$$= \frac{28.02\ \cancel{\text{g/mol}}}{211.64\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 13.24\%$$

(c) Molar mass of NH_4NO_3
= 2(Molar mass of N) + 4(Molar mass of H) + 3(Molar mass of O)
= 2(14.01 g/mol) + 4(1.01 g/mol) + 3(16.00 g/mol)
= 80.06 g/mol

Total molar mass of N in NH_4NO_3 = 2(Molar mass of N)
= 2 (14.01 g/mol)
= 28.02 g/mol

$$\text{Mass percent N in } NH_4NO_3 = \frac{\text{Total molar mass of N}}{\text{Molar mass } NH_4NO_3} \times 100\%$$
$$= \frac{28.02\ \cancel{\text{g/mol}}}{211.64\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 35.00\%$$

(d) Molar mass of HNO_3
= Molar mass of H + Molar mass of N + 3(Molar mass of O)
= 1.01 g/mol + 14.01 g/mol + 3(16.00 g/mol)
= 63.02 g/mol

Total molar mass of N in HNO_3 = 1(Molar mass of N)
= 14.01 g/mol

$$\text{Mass percent N in } HNO_3 = \frac{\text{Total molar mass of N}}{\text{Molar mass } HNO_3} \times 100\%$$
$$= \frac{14.01\ \cancel{\text{g/mol}}}{63.02\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 22.23\%$$

Check Your Solution

You can calculate the mass percent of the other elements in each compound in the same way. The total mass percents should equal 100%. For example, in N_2O, the mass percent of O is $\frac{16\ \cancel{\text{g/mol}}}{44.02\ \cancel{\text{g/mol}}} \times 100\% = 36.35\%$. The mass percent of the nitrogen was 63.65%. Both mass percents total 100%.

6. Problem

Sulfuric acid, H_2SO_4, is an important acid in laboratories and industries. Determine the percentage composition of sulfuric acid.

What Is required?

You need to find the mass percents of hydrogen, sulfur, and oxygen in the sulfuric acid.

What Is Given?

The molecular formula of sulfuric acid is given. This indicates the number of elements of each type in the compound.
H_2SO_4 = 4 H, 1 S, and 4 O

Plan Your Strategy

First determine the molar mass of the compound. To find the percentage composition of the compound, find the mass percent of each element. To do this, divide the total molar mass of each element by the molar mass of the compound and multiply by 100%.

Act on Your Strategy

Molar mass of H_2SO_4
= 2(Molar mass of H) + Molar mass of S + 4(Molar mass of O)
= 2(1.01 g/mol) + 32.07 g/mol + 4(16.00 g/mol)
= 98.09 g/mol

$$\text{Mass percent of H} = \frac{\text{Total molar mass of H}}{\text{Molar mass } H_2SO_4} \times 100\%$$
$$= \frac{2.02\ \cancel{\text{g/mol}}}{98.09\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 2.06\%$$

$$\text{Mass percent of S} = \frac{\text{Total molar mass of S}}{\text{Molar mass } H_2SO_4} \times 100\%$$
$$= \frac{32.07\ \cancel{\text{g/mol}}}{98.09\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 32.7\%$$

$$\text{Mass percent of O} = \frac{\text{Total molar mass of O}}{\text{Molar mass } H_2SO_4} \times 100\%$$
$$= \frac{64.00\ \cancel{\text{g/mol}}}{98.09\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 65.2\%$$

The percentage composition of sulfuric acid is 2.06% hydrogen, 32.7% sulfur, and 65.2% oxygen.

Check Your Solution

The mass percents add up to 100%.

7. Problem

Potassium nitrate, KNO_3, is used in fireworks. What is the mass percent of oxygen in potassium nitrate?

What Is required?

You need to find the mass percent of oxygen in the potassium nitrate.

What Is Given?

The molecular formula of KNO_3 is given. This indicates the number of elements of each type in the compound.
KNO_3 = 1 K, 1 N, and 3 O

Plan Your Strategy

First determine the molar mass of the compound. To find the mass percent of oxygen, divide the total molar mass of oxygen by the molar mass of the compound and multiply by 100%.

Act on Your Strategy

Molar mass of KNO_3 = Molar mass of K + Molar mass of N + 3(Molar mass of O)
= 39.10 g/mol + 14.01 g/mol + 3(16.00 g/mol)
= 101.11 g/mol

Total molar mass of O = 3(Molar mass of O)
= 3(16.00 g/mol)
= 48.00 g/mol

$$\text{Mass percent of O} = \frac{\text{Total molar mass of O}}{\text{Molar mass } KNO_3} \times 100\%$$
$$= \frac{48.00\ \cancel{\text{g/mol}}}{101.11\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 47.47\%$$

Check Your Solution

You can calculate the mass percent of the other elements in the compound in the same way. The total mass percents should add up to 100%.

8. Problem

A mining company wishes to extract manganese metal from pyrolusite ore, MnO_2.

(a) What is the percentage composition of pyrolusite ore?

(b) Use your answer from part (a) to calculate the mass of pure manganese that can be extracted from 250 kg pyrolusite ore.

What Is Required?

(a) You need to find the mass percents of manganese and oxygen in the pyrolusite ore.

(b) You need to find the mass of the manganese in 250 kg of pyrolusite ore.

What Is given?

(a) The molecular formula of pyrolusite ore is given. This indicates the number of elements of each type in the compound.
MnO_2 = 1 Mn and 2 O

(b) The mass of the pyrolusite is given.
MnO_2 = 250 kg

Plan Your Strategy

(a) First determine the molar mass of the compound. To find the mass percent of each element, divide the total molar mass of the element by the molar mass of the compound and multiply by 100%.

(b) To find the mass of manganese in the pyrolusite ore, multiply the mass percent of manganese by the mass of the ore.

First determine the molar mass of the compound. To find the mass percent of oxygen, divide the total molar mass of oxygen by the molar mass of the compound and multiply by 100%.

Act on Your Strategy

(a) Molar mass of MnO_2 = Molar mass of Mn + 2(Molar mass of O)
= 54.94 g/mol + 2(16.00 g/mol)
= 86.94 g/mol

$$\text{Mass percent of Mn} = \frac{\text{Total molar mass of Mn}}{\text{Molar mass of } MnO_2} \times 100\%$$
$$= \frac{54.94\ \cancel{\text{g/mol}}}{86.94\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 63.19\%$$

$$\text{Mass percent of O} = \frac{\text{Total molar mass of O}}{\text{Molar mass of } MnO_2} \times 100\%$$
$$= \frac{32.00\ \cancel{\text{g/mol}}}{86.94\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 36.81\%$$

The percentage composition of pyrolusite ore is 63.19% manganese and 36.81% oxygen.

(b) The mass of manganese in 250 kg of MnO_2 = Mass percent of Mn × 250 kg
= (63.19 / 100) × 250 kg
= 158 kg

Check Your Solution

(a) The mass percents add up to 100%.

(b) Calculated similarly, the mass of oxygen in the pyrolusite ore is 36.81% × 250 kg = 92 kg. The masses of manganese and oxygen (i.e. 158 kg + 92 kg) add up to 250 kg.

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9. Problem

A compound consists of 17.6% hydrogen and 82.4% nitrogen. Determine the empirical formula of the compound.

What Is Required?

You need to find the empirical formula of the compound.

What Is Given?

You know the percentage composition of the compound. You have access to a periodic table.

Plan Your Strategy

Assume you have 100 g of the compound. From the given mass percents of the elements in the compound, you can determine the mass of the element. This means you have 17.6 g of hydrogen and 82.4 g of nitrogen. Convert each mass to moles using the molar mass of the element. The number of moles can then be converted into a lower terms ratio of the element to get the empirical formula.

Act on Your Strategy

$$\text{Number of moles of H in 100 g sample} = \frac{\text{Mass of H}}{\text{Molar mass of H}}$$
$$= \frac{17.6\ \cancel{\text{g}}}{1.01\ \cancel{\text{g}}/\text{mol}}$$
$$= 17.425\ \text{mol}$$

$$\text{Number of moles of N in 100 g sample} = \frac{\text{Mass of N}}{\text{Molar mass of N}}$$
$$= \frac{82.4\ \cancel{\text{g}}}{14.01\ \cancel{\text{g}}/\text{mol}}$$
$$= 5.882\ \text{mol}$$

$$\text{The lowest whole number ratio} = \frac{\text{molar amount}}{\text{lowest molar amount}}$$

$$H_{\frac{17.425\ \cancel{\text{mol}}}{5.882\ \cancel{\text{mol}}}} N_{\frac{5.882\ \cancel{\text{mol}}}{5.882\ \cancel{\text{mol}}}} \rightarrow H_{2.96}N_{1.00} \rightarrow H_3N_1$$

Alternatively, you can set your solution as a table.

Element	Mass percent (%)	Grams per 100 g sample (g)	Molar mass (g/mol)	Number of moles (mol)	Molar amount + lowest molar amount
H	17.6	17.6	1.01	17.425	17.425/5.882 =2.96
N	82.4	82.4	14.01	5.882	5.882/5.882=1

The empirical formula of the compound is H_3N_1 or NH_3.

Check Your Solution

Work backward. Calculate the percentage composition of NH_3.

$$\text{Mass percent of N} = \frac{14.01\ \cancel{\text{g/mol}}}{17.04\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 82.2\%$$
$$\text{Mass percent of H} = \frac{3.03\ \cancel{\text{g/mol}}}{17.04\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 17.8\%$$

These calculated values are very close to the given percentage compositions (variations arise depending on the periodic table value used for the molar mass of N and H, which differs in different publications).

10. Problem

Find the empirical formula of a compound that is 46.3% lithium and 53.7% oxygen.

What Is Required?

You need to find the empirical formula of the compound.

What Is Given?

You know the percentage composition of the compound. You have access to a periodic table.

Plan Your Strategy

Assume you have 100 g of the compound. From the given mass percents of the elements in the compound, you can determine the mass of the element. This means you have 46.3 g of lithium and 53.7 g of oxygen. Convert each mass to moles using the molar mass of the element. The number of moles can then be converted into a lower terms ratio of the element to get the empirical formula.

Act on Your Strategy

$$\text{Number of moles of Li in 100 g sample} = \frac{\text{Mass of Li}}{\text{Molar mass of Li}}$$
$$= \frac{46.3\ \cancel{\text{g}}}{6.94\ \cancel{\text{g}}/\text{mol}}$$
$$= 6.67\ \text{mol}$$
$$\text{Number of moles of O in 100 g sample} = \frac{\text{Mass of O}}{\text{Molar mass of O}}$$
$$= \frac{53.7\ \cancel{\text{g}}}{16.00\ \cancel{\text{g}}/\text{mol}}$$
$$= 3.35\ \text{mol}$$
$$\text{The lowest whole number ratio} = \frac{\text{molar amount}}{\text{lowest molar amount}}$$
$$\text{Li}_{\frac{6.67}{3.35}}\text{O}_{\frac{3.35}{3.35}} \rightarrow \text{Li}_{1.99}\text{O}_{1.00} \rightarrow \text{Li}_2\text{O}_1$$

Alternatively, you can set your solution as a table.

Element	Mass percent (%)	Grams per 100 g sample (g)	Molar mass (g/mol)	Number of moles (mol)	Molar amount + lowest molar amount
Li	46.3	46.3	6.94	6.67	6.67/3.35=1.99
O	53.7	53.7	16.00	3.35	3.35/3.35=1

The empirical formula of the compound is Li_2O_1 or Li_2O.

Check Your Solution

Work backward. Calculate the percentage composition of Li_2O.

$$\text{Mass percent of Li} = \frac{13.88\ \cancel{\text{g/mol}}}{29.88\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 46.45\%$$
$$\text{Mass percent of O} = \frac{16.00\ \cancel{\text{g/mol}}}{29.88\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 53.55\%$$

These calculated values are very close to the given percentage compositions (variations arise depending on the periodic table value used for the molar mass of Li and O, which differs in different publications).

11. Problem

What is the empirical formula of a compound that is 15.9% boron and 84.1% fluorine.

What Is Required?

You need to find the empirical formula of the compound.

What Is Given?

You know the percentage composition of the compound. You have access to a periodic table.

Plan Your Strategy

Assume you have 100 g of the compound. From the given mass percents of the elements in the compound, you can determine the mass of the element. This means you have 15.9 g of boron and 84.1 g of fluorine. Convert each mass to moles using the molar mass of the element. The number of moles can then be converted into a lower terms ratio of the element to get the empirical formula.

Act on Your Strategy

$$\text{Number of moles of B in 100 g sample} = \frac{\text{Mass of B}}{\text{Molar mass of B}}$$
$$= \frac{15.9\ \cancel{\text{g}}}{10.81\ \cancel{\text{g}}/\text{mol}}$$
$$= 1.47\ \text{mol}$$
$$\text{Number of moles of F in 100 g sample} = \frac{\text{Mass of F}}{\text{Molar mass of F}}$$
$$= \frac{84.1\ \cancel{\text{g}}}{19.00\ \cancel{\text{g}}/\text{mol}}$$
$$= 4.43\ \text{mol}$$
$$\text{The lowest whole number ratio} = \frac{\text{molar amount}}{\text{lowest molar amount}}$$
$$B_{\frac{1.47}{1.47}}F_{\frac{4.43}{1.47}} \rightarrow B_{1.00}F_{3.01} \rightarrow BF_3$$

Alternatively, you can set your solution as a table.

Element	Mass percent (%)	Grams per 100 g sample (g)	Molar mass (g/mol)	Number of moles (mol)	Molar amount + lowest molar amount
B	15.9	15.9	10.81	1.47	1.47/1.47=1
F	84.1	84.1	19.00	4.43	4.43/1.47=3.01

The empirical formula of the compound is BF_3.

Check Your Solution

Work backward. Calculate the percentage composition of BF_3.

$$\text{Mass percent of B} = \frac{10.81\ \cancel{\text{g/mol}}}{67.81\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 15.9\%$$
$$\text{Mass percent of F} = \frac{57.00\ \cancel{\text{g/mol}}}{67.81\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 84.1\%$$

These calculated values are the same as the given percentage compositions.

12. Problem

Determine the empirical formula of a compound made up of 52.51% chlorine and 47.48% sulfur.

What Is Required?

You need to find the empirical formula of the compound.

What Is Given?

You know the percentage composition of the compound. You have access to a periodic table.

Plan Your Strategy

Assume you have 100 g of the compound. From the given mass percents of the elements in the compound, you can determine the mass of the element. This means you have 52.51 g of chlorine and 47.48 g of sulfur. Convert each mass to moles using the molar mass of the element. The number of moles can then be converted into a lower terms ratio of the element to get the empirical formula.

Act on Your Strategy

$$\text{Number of moles of Cl in 100 g sample} = \frac{\text{Mass of Cl}}{\text{Molar mass of Cl}}$$
$$= \frac{52.51\ \cancel{\text{g}}}{35.45\ \cancel{\text{g}}/\text{mol}}$$
$$= 1.48\ \text{mol}$$
$$\text{Number of moles of S in 100 g sample} = \frac{\text{Mass of S}}{\text{Molar mass of S}}$$
$$= \frac{47.48\ \cancel{\text{g}}}{32.07\ \cancel{\text{g}}/\text{mol}}$$
$$= 1.48\ \text{mol}$$
$$\text{The lowest whole number ratio} = \frac{\text{molar amount}}{\text{lowest molar amount}}$$
$$\text{Cl}_{\frac{1.48}{1.48}}\text{S}_{\frac{1.48}{1.48}} \rightarrow \text{Cl}_{1.00}\text{S}_{1.00} \rightarrow \text{ClS}$$

Alternatively, you can set your solution as a table.

Element	Mass percent (%)	Grams per 100 g sample (g)	Molar mass (g/mol)	Number of moles (mol)	Molar amount + lowest molar amount
Cl	52.51	52.51	35.45	1.48	1.48/1.48=1
S	47.48	47.48	32.07	1.48	1.48/1.48=1

The empirical formula of the compound is ClS.

Check Your Solution

Work backward. Calculate the percentage composition of ClS.

$$\text{Mass percent of Cl} = \frac{35.45\ \cancel{\text{g/mol}}}{67.52\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 52.5\%$$
$$\text{Mass percent of S} = \frac{32.07\ \cancel{\text{g/mol}}}{67.52\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 47.5\%$$

These calculated values are the same as the given percentage compositions.

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13. Problem

An oxide of chromium is made up of 68.4% chromium and 31.6% oxygen. What is the empirical formula of this oxide?

What Is Required?

You need to find the empirical formula of this oxide of chromium.

What Is Given?

You know the percentage composition of the compound. You have access to a periodic table.

Plan Your Strategy

Assume you have 100 g of the compound. Convert the mass percents of the elements to mass, then to the number of moles. Then find the lowest whole number ratio.

Act on Your Strategy

Element	Mass percent (%)	Grams per 100 g sample (g)	Molar mass (g/mol)	Number of moles (mol)	Molar amount + lowest molar amount
Cr	68.4	68.4	52.00	1.315	1.00
O	31.6	31.6	16.00	1.975	1.50

You now have the empirical formula $Cr_1O_{1.50}$. Convert the subscript 1.50 (3/2) to a whole number. $Cr_{1\times2}O_{1.5\times2} = Cr_2O_3$.

The empirical formula of the compound is Cr_2O_3.

Check Your Solution

Work backward. Calculate the percentage composition of Cr_2O_3.

$$\text{Mass percent of Cr} = \frac{104.00\ \cancel{\text{g/mol}}}{152.00\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 68.4\%$$
$$\text{Mass percent of O} = \frac{48.00\ \cancel{\text{g/mol}}}{152.00\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 31.6\%$$

These calculated values are the same as the given percentage compositions.

14. Problem

Phosphorus reacts with oxygen to give a compound that is 43.7% phosphorus and 56.4% oxygen. What is the empirical formula of the compound?

What Is Required?

You need to find the empirical formula of this compound.

What Is Given?
You know the percentage composition of the compound. You have access to a periodic table.

Plan Your Strategy
Assume you have 100 g of the compound. Convert the mass percents of the elements to mass, then to the number of moles. Then find the lowest whole number ratio.

Act on Your Strategy

Element	Mass percent (%)	Grams per 100 g sample (g)	Molar mass (g/mol)	Number of moles (mol)	Molar amount ÷ lowest molar amount
P	43.7	43.7	30.97	1.411	1.000
O	56.4	56.4	16.00	3.525	2.498

You now have the empirical formula $P_1O_{2.498}$. Convert the subscript 2.498 (5/2) to a whole number. $P_{1\times2}O_{2.498\times2} = P_2O_5$.
The empirical formula of the compound is P_2O_5.

Check Your Solution
Work backward. Calculate the percentage composition of P_2O_5.

$$\text{Mass percent of P} = \frac{61.94\ \cancel{\text{g/mol}}}{141.8\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 43.7\%$$
$$\text{Mass percent of O} = \frac{80.00\ \cancel{\text{g/mol}}}{141.8\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 56.4\%$$

These calculated values are the same as the given percentage compositions.

15. Problem
An inorganic salt is composed of 17.6% sodium, 39.7% chromium, and 42.8% oxygen. What is the empirical formula of this salt?

What Is Required?
You need to find the empirical formula of this inorganic salt.

What Is Given?
You know the percentage composition of the compound. You have access to a periodic table.

Plan Your Strategy
Assume you have 100 g of the compound. Convert the mass percents of the elements to mass, then to the number of moles. Then find the lowest whole number ratio.

Act on Your Strategy

Element	Mass percent (%)	Grams per 100 g sample (g)	Molar mass (g/mol)	Number of moles (mol)	Molar amount ÷ lowest molar amount
Na	17.6	17.6	22.99	0.765	1.002
Cr	39.7	39.7	52.00	0.763	1.000
O	42.8	42.8	16.00	2.675	3.505

You now have the empirical formula $Na_{1.002}Cr_{1.000}O_{3.505}$. Convert the subscript 3.505 (7/2) to a whole number. $Na_{1.002\times2}Cr_{1.000\times2}O_{3.505\times2} = Na_2Cr_2O_7$.
The empirical formula of the compound is $Na_2Cr_2O_7$.

Check Your Solution

Work backward. Calculate the percentage composition of $Na_2Cr_2O_7$.

$$\text{Mass percent of Na} = \frac{45.98\ \cancel{\text{g/mol}}}{261.98\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 17.6\%$$
$$\text{Mass percent of Cr} = \frac{104.00\ \cancel{\text{g/mol}}}{261.98\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 39.7\%$$
$$\text{Mass percent of O} = \frac{112.00\ \cancel{\text{g/mol}}}{261.98\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 42.8\%$$

These calculated values are the same as the given percentage compositions.

16. Problem

Compound X contains 69.9% carbon, 6.86% hydrogen, and 23.3% oxygen. Determine the empirical formula of compound X.

What Is Required?

You need to find the empirical formula of compound X.

What Is Given?

You know the percentage composition of the compound. You have access to a periodic table.

Plan Your Strategy

Assume you have 100 g of the compound. Convert the mass percents of the elements to mass, then to the number of moles. Then find the lowest whole number ratio.

Act on Your Strategy

Element	Mass percent (%)	Grams per 100 g sample (g)	Molar mass (g/mol)	Number of moles (mol)	Molar amount + lowest molar amount
C	69.9	69.9	12.01	5.820	3.99
H	6.86	6.86	1.01	6.792	4.665
O	23.3	23.3	16.00	1.456	1.000

You now have the empirical formula $C_{3.99}H_{4.664}O_{1.000}$. Convert the subscripts 3.99 (12/3) and 4.665 (14/3) to whole numbers. $C_{3.99\times3}H_{4.665\times3}O_{1.000\times3} = C_{12}H_{14}O_3$. The empirical formula of the compound is $C_{12}H_{14}O_3$.

Check Your Solution

Work backward. Calculate the percentage composition of $C_{12}H_{14}O_3$.

$$\text{Mass percent of C} = \frac{144.12\ \cancel{\text{g/mol}}}{206.26\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 69.9\%$$
$$\text{Mass percent of H} = \frac{14.4\ \cancel{\text{g/mol}}}{206.26\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 6.86\%$$
$$\text{Mass percent of O} = \frac{48.00\ \cancel{\text{g/mol}}}{206.26\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 23.3\%$$

These calculated values are the same as the given percentage compositions.

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17. Problem

The empirical formula of butane, the fuel used in disposable lighters, is C_2H_5. In an experiment, the molar mass of butane was determined to be 58 g/mol. What is the molecular formula of butane?

What Is Required?

You need to find the molecular formula of butane.

What Is Given?

You know the empirical formula and the molar mass of butane.

Plan Your Strategy

Divide the molar mass of butane by the "molar mass" of the empirical formula. The answer you get is the factor by which you multiply the empirical formula.

Act on Your Strategy

The "molar mass" of the empirical formula C_2H_5, determined using the periodic table, is

$$2(12.01\ \text{g/mol}) + 5(1.01\ \text{g/mol}) = 29.07\ \text{g/mol}$$

The molar mass of butane is = 58 g/mol.

$$\frac{58\ \cancel{\text{g/mol}}}{29.07\ \cancel{\text{g/mol}}} = 2$$

Molecular formula subscripts = 2 × Empirical formula subscripts

$$= C_{2\times2}H_{5\times2}$$

$$= C_4H_{10}$$

Therefore, the molecular formula of butane is C_4H_{10}.

Check Your Solution

Work backward by calculating the molar mass of C_4H_{10}.

$(4 \times 12.01\ \text{g/mol}) + (10 \times 1.01\text{g/mol}) = 58\ \text{g/mol}$

The calculated molar mass matches the molar mass that is given in the problem. The answer is reasonable.

18. Problem

Oxalic acid has the empirical formula CHO_2. Its molar mass is 90 g/mol. What is the molecular formula of oxalic acid?

What Is Required?

You need to find the molecular formula of oxalic acid.

What Is Given?

You know the empirical formula and the molar mass of oxalic acid.

Plan Your Strategy

Divide the molar mass of oxalic acid by the "molar mass" of the empirical formula. The answer you get is the factor by which you multiply the empirical formula.

Act on Your Strategy

The "molar mass" of the empirical formula CHO_2, determined using the periodic table, is

$$12.01\ \text{g/mol} + 1.01\ \text{g/mol} + 2(16.00\ \text{g/mol}) = 45.02\ \text{g/mol}$$

The molar mass of oxalic acid is = 90 g/mol.

$$\frac{90\ \cancel{\text{g/mol}}}{45.02\ \cancel{\text{g/mol}}} = 2$$

Molecular formula subscripts = 2 × Empirical formula subscripts
$= C_{1\times2}H_{1\times2}O_{2\times2}$
$= C_2H_2O_4$

Therefore, the molecular formula of butane is $C_2H_2O_4$.

Check Your Solution

Work backward by calculating the molar mass of $C_2H_2O_4$.
(2 × 12.01 g/mol) + (2 × 1.01 g/mol) + (4 × 16.00 g/mol) = 90 g/mol
The calculated molar mass matches the molar mass that is given in the problem. The answer is reasonable.

19. Problem

The empirical formula of codeine is $C_{18}H_{21}NO_3$. If the molar mass of codeine is 299 g/mol, what is its molecular formula?

What Is Required?

You need to find the molecular formula of codeine.

What Is Given?

You know the empirical formula and the molar mass of codeine.

Plan Your Strategy

Divide the molar mass of codeine by the "molar mass" of the empirical formula. The answer you get is the factor by which you multiply the empirical formula.

Act on Your Strategy

The "molar mass" of the empirical formula $C_{18}H_{21}NO_3$, determined using the periodic table, is
18(12.01 g/mol) + 21(1.01 g/mol) + 14.01 g/mol + 3(16.00 g/mol) = 299.4 g/mol.
The molar mass of the codeine is the same as that of its empirical formula. Therefore, the molecular formula is the same as its empirical formula: $C_{18}H_{21}NO_3$.

Check Your Solution

The calculated molar masses are the same, so the formula should be the same. The answer is reasonable.

20. Problem

A compound's molar mass is 240.28 g/mol. Its percentage composition is 75.0% carbon, 5.05% hydrogen, and 20.0% oxygen. What is the compound's molecular formula?

What Is Required?

You need to find the molecular formula of the unknown compound.

What Is Given?

You know the mass percent of the elements in the compound. You also know the molar mass of the compound.

Plan Your Strategy

(a) First determine the empirical formula. Assume you have 100 g of the compound. Convert the mass percents of the elements to mass, then to the number of moles. Then find the lowest whole number ratio.

(b) To determine the molecular formula, divide the molar mass of oxalic acid by the "molar mass" of the empirical formula. The answer you get is the factor by which you multiply the empirical formula.

Act on Your Strategy

(a) Consider a 100 g sample:

Element	Grams per 100 g sample (g)	Molar mass (g/mol)	Number of moles (mol)	Molar amount + lowest molar amount
C	75.0	12.01	75.0 / 12.01 = 6.25	5
H	5.05	1.01	5.05 / 1.01 = 5.00	4
O	20.0	16.00	20.0 / 16.01 = 1.25	1

The empirical formula of the compound is C_5H_4O.

(b) The "molar mass" of the empirical formula C_5H_4O, determined using the periodic table, is

$$5(12.01 \text{ g/mol}) + 4(1.01 \text{ g/mol}) + 16.00 \text{ g/mol} = 80.1 \text{ g/mol}$$

The molar mass of the compound is = 240.28 g/mol.

$$\frac{240.28 \cancel{\text{ g/mol}}}{80.1 \cancel{\text{ g/mol}}} = 3.0$$

Molecular formula subscripts = 3 × Empirical formula subscripts

$$= C_{5\times3}H_{4\times3}O_{1\times3}$$
$$= C_{15}H_{12}O_3$$

Therefore, the molecular formula of this unknown compound is $C_{15}H_{12}O_3$.

Check Your Solution

Work backward by calculating the molar mass of $C_{15}H_{12}O_3$.

$(15 \times 12.01 \text{ g/mol}) + (12 \times 1.01 \text{ g/mol}) + (3 \times 16.00 \text{ g/mol}) = 240.27 \text{ g/mol}$

The calculated molar mass matches the molar mass that is given in the problem. The answer is reasonable.

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21. Problem

A 0.539 g sample of a compound that contained only carbon and hydrogen was subjected to combustion analysis. The combustion produced 1.64 g of carbon dioxide and 0.807 g of water. Calculate the percentage composition and the empirical formula of the sample.

What Is Required?

(a) You need to find the mass percent of carbon and hydrogen in the sample.

(b) You need to find the empirical formula of the compound.

What Is Given?

You know the mass of the sample. You also know the masses of the water and the carbon dioxide in the combustion of the sample.

Plan Your Strategy

(a) All the carbon was converted into the carbon dioxide, and all the hydrogen was converted into water. Therefore, the mass percent of C in the sample substrate is its mass in CO_2 divided by the mass of the sample substrate, multiplied by 100%. Similarly, the mass percent of H in the sample is its mass in water divided by the mass of the sample substrate, multiplied by 100%. You obtain this mass by dividing the total molar mass of the element (H or C) in the products (H_2O or CO_2) by the molecular mass of the products, and multiplying by the given mass of the products.

(b) To determine the empirical formula of the substrate sample, convert the calculated masses of the C and H in (a) to the number of moles using the molar masses of C and H, respectively. Then find the lowest whole number ratio.

Act on Your Strategy

(a) Mass of carbon in the carbon dioxide produced

$$= \frac{12.01\ \cancel{\text{g/mol}}}{44.01\ \cancel{\text{g/mol}}} \times 1.64\ \text{g} = 0.44754\ \text{g}$$

Mass percent of C in the sample $= \frac{0.44754\ \cancel{\text{g}}}{0.538\ \cancel{\text{g}}} \times 100\% = 83.19\%$

Mass of hydrogen in the water produced $= \frac{2.02\ \cancel{\text{g/mol}}}{18.02\ \cancel{\text{g/mol}}} \times 0.807\ \text{g} = 0.09046\ \text{g}$

Mass percent of H in sample $= \frac{0.09046\ \cancel{\text{g}}}{0.538\ \cancel{\text{g}}} \times 100\% = 16.81\%$

Thus the percentage composition of the unknown sample is 83.19% C and 16.81% H.

(b) Moles of C in sample $= \frac{0.44754\ \cancel{\text{g}}}{12.01\ \cancel{\text{g}}/\text{mol}} = 0.03726\ \text{mol}$

Moles of H in sample $= \frac{0.09046\ \cancel{\text{g}}}{1.01\ \cancel{\text{g}}/\text{mol}} = 0.8956\ \text{mol}$

Empirical formula $= \text{C}_{\frac{0.03726}{0.03726}}\text{H}_{\frac{0.08956}{0.03726}} = \text{C}_1\text{H}_{2.404}$

Convert the subscript 2.404 ($\frac{12}{5}$) to a whole number. $\text{C}_{1\times5}\text{H}_{2.404\times5} = \text{C}_5\text{H}_{12}$.

Thus the empirical formula is C_5H_{12}.

Check Your Solution

(a) The sum of the masses of carbon and hydrogen is 0.44754 g + 0.09046 g = 0.538 g. This is close to the given mass of the sample.

(b) Work backward. Calculate the percentage composition of C_5H_{12}.

Mass percent of C $= \frac{60.05\ \cancel{\text{g/mol}}}{72.17\ \cancel{\text{g/mol}}} \times 100\%$

$= 83.21\%$

Mass percent of H $= \frac{12.12\ \cancel{\text{g/mol}}}{72.17\ \cancel{\text{g/mol}}} \times 100\%$

$= 16.79\%$

These mass percent values are close to the values calculated in part (a). Therefore your answers are reasonable.

22. Problem

An 874 mg sample of cortisol was subjected to carbon-hydrogen combustion analysis. 2.23 g of carbon dioxide and 0.652 g of water were produced. The molar mass of cortisol was found to be 362 g/mol using a mass spectrometer. If cortisol contains carbon, hydrogen, and oxygen, determine its molecular formula.

What Is Required?

You need to find the molecular formula of cortisol.

What Is Given?

The mass of the sample cortisol is given, as well as the masses of the carbon dioxide and water produced. The molar mass of cortisol is also given.

Plan Your Strategy

The empirical formula of cortisol must first be found in order to determine its molecular formula. Therefore, the number of moles of C, H, and O in cortisol has to be determined for the mole ratio. In a complete reaction, all the C would have been converted to CO_2 and all the H to H_2O. Therefore, the masses of C and H in the products are the same as that in the 874 mg sample of cortisol.

(a) You first obtain this mass by dividing the total molar mass of the element (H or C) in the products (H_2O or CO_2) by the molecular mass of the products, and multiplying by the given mass of the products.

(b) By the Law of Conservation of Mass, the mass of the O in the sample of cortisol as the given mass of cortisol minus the sum of the calculated masses of C and H.

(c) Finally, convert the masses of the C, H and O calculated in (a) and (b) to the number of moles using the molar masses of C, H, and O, respectively. Then find the lowest whole number ratio.

Act on Your Strategy

(a) Mass of carbon in the CO_2 produced $= \frac{12.01\ \cancel{\text{g/mol}}}{44.01\ \cancel{\text{g/mol}}} \times 2.23\ \text{g} = 0.60855\ \text{g}$

Mass of hydrogen in the H_2O produced $= \frac{2.02\ \cancel{\text{g/mol}}}{18.02\ \cancel{\text{g/mol}}} \times 0.652\ \text{g} = 0.07308\ \text{g}$

(b) Using the Law of Conservation of Mass, the O in the original compound

$$= 0.874\ \text{g} - (0.60855\ \text{g} + 0.07308\ \text{g}) = 0.19237\ \text{g}$$

(c) Moles of C in cortisol $= \frac{0.60855\ \cancel{\text{g}}}{12.01\ \cancel{\text{g}}/\text{mol}} = 0.05067\ \text{mol}$

Moles of H in cortisol $= \frac{0.07308\ \cancel{\text{g}}}{1.01\ \cancel{\text{g}}/\text{mol}} = 0.07235\ \text{mol}$

Moles of O in cortisol $= \frac{0.19237\ \cancel{\text{g}}}{16.00\ \cancel{\text{g}}/\text{mol}} = 0.01202\ \text{mol}$

Empirical formula $= C_{\frac{0.05067}{0.01202}} H_{\frac{0.07235}{0.01202}} O_{\frac{0.01202}{0.01202}} = C_{4.21}H_{6.02}O_{1.00}$

Convert the subscripts 4.21 ($\frac{21}{5}$) and 6.02 ($\frac{30}{5}$) to whole numbers.

$C_{4.21\times5}H_{6.02\times5}O_{1\times5} = C_{21}H_{30}O_5$

The "molar mass" of this empirical formula is

$= 21(12.01\ \text{g/mol}) + 30(1.01\ \text{g/mol}) + 5(16.00\ \text{g/mol}) = 362.51\ \text{g/mol}$

Since the given molar mass of cortisol was 362 g/mol, then the molecular formula is the same as the empirical formula, $C_{21}H_{30}O_5$.

Check Your Solution

The calculated molar masses are the same, so the formula should be the same. The answer is reasonable.

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23. Problem

What is the percent by mass of water in magnesium sulfite hexahydrate, $MgSO_3 \cdot 6H_2O$?

What Is Required?

You need to find the mass percent of water in the magnesium sulfite hexahydrate.

What Is Given?

The molecular formula of the sample is $MgSO_3 \cdot 6H_2O$.

Plan Your Strategy

Calculate the molar mass of the magnesium sulfite hexahydrate. Then calculate the total molar mass of the water portion of the substance. The mass percent is the molar mass of the water divided by the molar mass of the substance, multiplied by 100%.

Act on Your Strategy

Molar mass of $MgSO_3 \cdot 6H_2O$

= 24.31 g/mol + 32.07 g/mol + 3(16.00 g/mol) + 6[2(1.01 g/mol) + 16.00 g/mol]
= 212.50 g/mol
Molar mass of water portion = 6[2(1.01) + 16.00] = 108.12 g/mol

$$\text{Mass percent of water} = \frac{108.12\ \cancel{\text{g/mol}}}{212.50\ \cancel{\text{g/mol}}} \times 100\%$$
$$= 50.8\%$$

Check Your Solution
Similarly, the mass percent of the $MgSO_3$ portion is
$\frac{24.31\ \text{g/mol} + 32.07\ \text{g/mol} + 3(16.00\ \text{g/mol})}{212.50\ \text{g/mol}} \times 100\% = 49.2\%$
The two mass percents add up to 100%.

24. Problem
A 3.34 g sample of a hydrate has the formula $SrS_2O_3 \bullet xH_2O$, and contains 2.30 g of SrS_2O_3. Find the value of x.

What Is Required?
You need to find how many water molecules are bonded to each formula unit of SrS_2O_3.

What Is Given?
The mass of the sample is given, as well as the mass of the SrS_2O_3 portion of the hydrate. The formula of the hydrate is given.

Plan Your Strategy
To find the mass of water in the hydrate, subtract the given mass of SrS_2O_3 from the given mass of the hydrate. Then, find the number of moles of SrS_2O_3 and H_2O by dividing their mass by their respective molar mass. To find out how many water molecules bond to each formula unit of SrS_2O_3, divide each mole value obtained by the number of moles of SrS_2O_3.

Act on Your Strategy
Mass of water portion = Mass of hydrate − Mass of SrS_2O_3
= 3.34 g − 2.30 g
= 1.04 g

$$\text{Moles of } SrS_2O_3 = \frac{2.30\ \cancel{\text{g}}}{199.76\ \cancel{\text{g}}/\text{mol}} = 0.0115\ \text{mol}$$

$$\text{Moles of } H_2O = \frac{1.04\ \cancel{\text{g}}}{18.02\ \cancel{\text{g}}/\text{mol}} = 0.0577\ \text{mol}$$

The mole ratio is
$\frac{0.0115}{0.0115}$ mol SrS_2O_3 : $\frac{0.0577}{0.0115}$ mol H_2O = 1 mol SrS_2O_3 : 5.02 mol H_2O ,which is rounded off to 1 mol SrS_2O_3 : 5 mol H_2O
The value of x in $SrS_2O_3 \bullet xH_2O$ is 5. Therefore the molecular formula is $SrS_2O_3 \bullet 5H_2O$.

Check Your Solution
According to the formula, the percent by mass of water in $SrS_2O_3 \bullet 5H_2O$ is $\frac{5(18.02\ \cancel{\text{g/mol}})}{289.86\ \cancel{\text{g/mol}}} \times 100\% = 31.08\%$. The percent by mass of the SrS_2O_3 is $\frac{199.76\ \cancel{\text{g/mol}}}{289.86\ \cancel{\text{g/mol}}} \times 100\% = 68.92\%$. The two mass percents add up to 100%.

25. Problem
A hydrate of zinc chlorate, $Zn(ClO_3)_2 \bullet xH_2O$, contains 21.5% zinc by mass. Find the value of x.

What Is Required?

You need to find out how many water molecules are bonded to each formula unit of $Zn(ClO_3)_2$.

What Is Given?

The mass percent of zinc is given, as well as the formula of the zinc chlorate.

Plan Your Strategy

Find the molar mass of the zinc hydrate in terms of x by adding up the molar masses of all the elements in the given formula. Equate the given mass percent of Zn as being the molar mass of Zn divided by the calculated molar mass of the zinc chlorate in terms of x, multiplied by 100%. In this way, x can be solved.

Act on Your Strategy

Molar mass of the $Zn(ClO_3)_2 \bullet xH_2O$ in terms of x

$= 65.39 \text{ g/mol} + 2[35.45 \text{ g/mol} + 3(16.00 \text{ g/mol})] + x[2(1.01 \text{ g/mol}) + 16.00 \text{ g/mol}]$

$= 232.29 \text{ g/mol} + 18.02x \text{ g/mol}$

$$\text{Mass percent of Zn} = 21.5\% = \frac{65.39 \text{ g/mol}}{232.29 \text{ g/mol} + 18.02x \text{ g/mol}} \times 100\%$$

Therefore,

$$21.5 \times (232.29 \text{ g/mol} + 18.02x \text{ g/mol}) = 65.39 \text{ g/mol} \times 100$$

$$4996.385 + 387.43x = 6539$$

$$387.43x = 1542.615$$

$$x = 3.987, \text{ which can be rounded off to 4.}$$

The formula of the hydrate is $Zn(ClO_3)_2 \bullet 4H_2O$.

Check Your Solution

According to the formula, the percent by mass of water in $Zn(ClO_3)_2 \bullet 4H_2O$ is $\frac{4(18.02 \cancel{\text{g/mol}})}{304.37 \cancel{\text{g/mol}}} \times 100\% = 23.68\%$. The percent by mass of the $Zn(ClO_3)_2$ is $\frac{232.29 \cancel{\text{g/mol}}}{304.37 \cancel{\text{g/mol}}} \times 100\% = 76.32\%$. The two mass percents add up to 100%.