Section 5: Empirical and Molecular Formulas

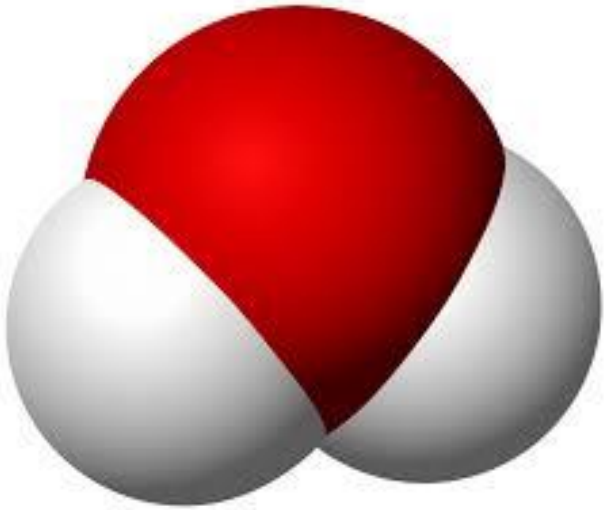
Chapter 7: Chemical Composition

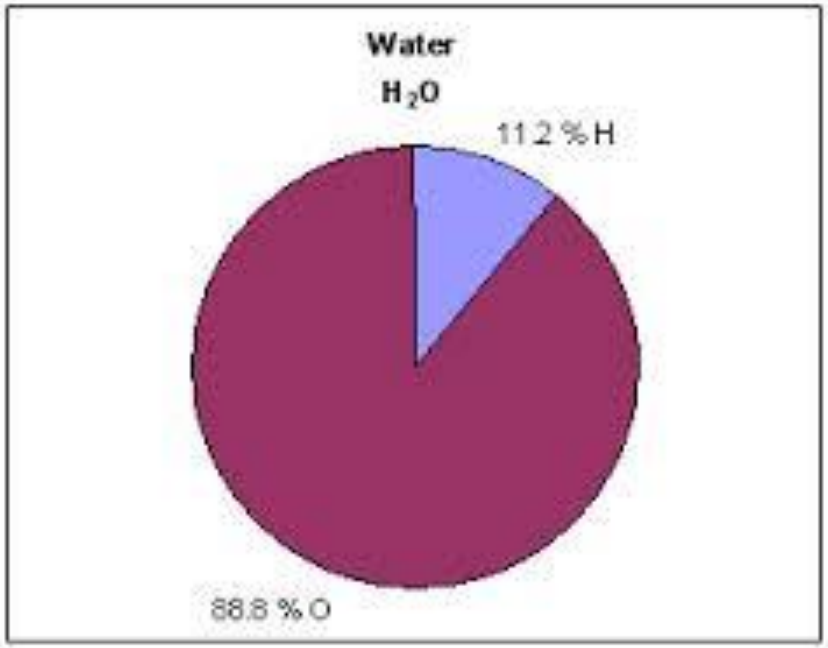
Learning Goals

- Use mass percent composition as a conversion factor.
- Determine mass percent composition from a chemical formula.
- Determine an empirical formula from experimental data.
- Calculate a molecular formula from an empirical formula and molar mass.

Percent Composition

- The **mass percent** of an element is the elements percentage of the total mass of the compound.





Percent Composition

- The **mass percent** of any element in a compound can be found by dividing the mass of the element by the mass of the compound and multiplying by 100.

$$\text{percent by mass (element)} = \frac{\text{mass of element}}{\text{mass of compound}} \times 100$$

Percent Composition

- A 0.358-g sample of chromium reacts with oxygen to form 0.523 g of the metal oxide.

$$\text{Mass percent Cr} = \frac{\text{Mass Cr}}{\text{Mass metal oxide}} \times 100\%$$

$$= \frac{0.358\ \text{g}}{0.523\ \text{g}} \times 100\% = 68.5\%$$

Percent Composition

- The mass percent of each element in a compound is the percent composition of a compound.
- Percent composition of a compound can also be determined from its chemical formula.

Percent Composition

$$\text{percent by mass} = \frac{\text{mass of element in 1 mol of compound}}{\text{molar mass of compound}} \times 100$$

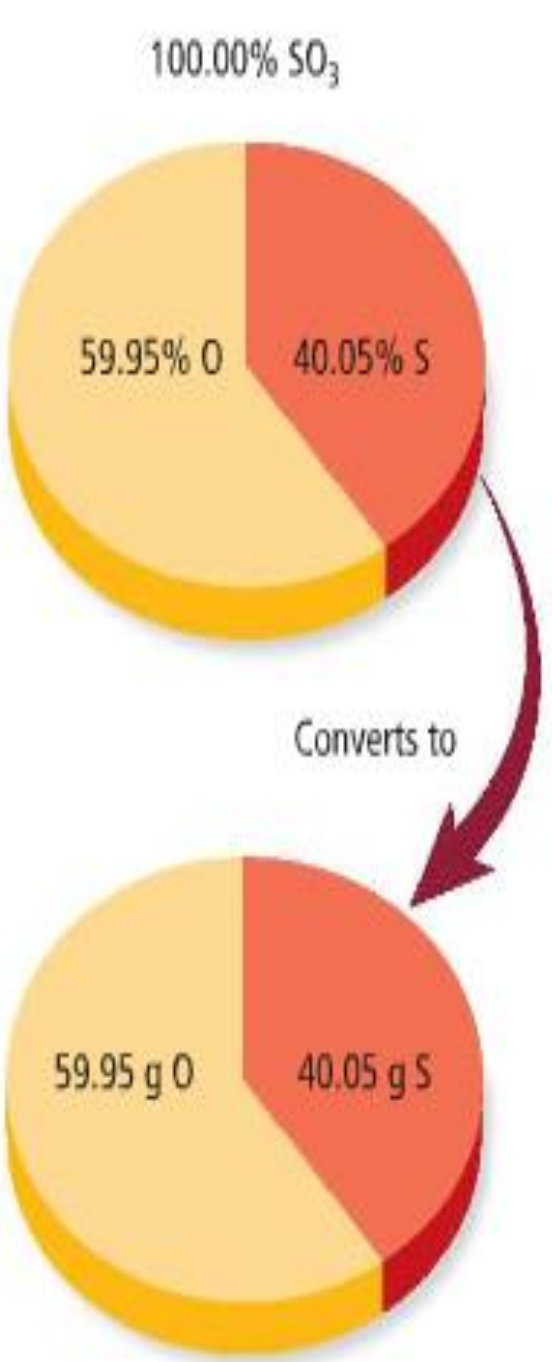
Percent Composition

- What is the percent composition of phosphoric acid?
- Which has the larger percent by mass of sulfur? H_2SO_3 or $H_2S_2O_8$

Empirical Formula

- The empirical formula for a compound is the smallest whole-number mole ratio of the elements.
- You can calculate the empirical formula from percent by mass by assuming you have 100.00 g of the compound. Then, convert the mass of each element to moles.

Figure 13 Keep this figure in mind when doing problems using percent composition. You can always assume that you have a 100-g sample of the compound and use the percents of the elements as masses of the elements.



Empirical Formula

- An empirical formula gives only the smallest whole-number *ratio* of each type of atom in a compound, not the specific number of each type of atom in a molecule.
  - The molecular formula is always a whole-number multiple of the empirical formula.

Empirical Formula

- The empirical formula may or may not be the same as the molecular formula.
  - Molecular formula of hydrogen peroxide = H_2O_2
  - Empirical formula of hydrogen peroxide = HO

Empirical Formula

1. Write down (or calculate) as given the masses of each element present in a sample of the compound. If you are given mass percent composition, assume a 100-g sample and calculate the masses of each element from the given percentages.
   - 63.16% O
   - 36.84% N

Empirical Formula

2. Convert each of the masses in Step 1 to moles by using the appropriate molar mass for each element as a conversion factor.
   - 63.16 g O
   - 36.84 g N

Empirical Formula

3. Write down a pseudo-formula for the compound, using the moles of each element (from Step 2) as subscripts.

Empirical Formula

4. Divide all the subscripts in the formula by the smallest subscript.

Empirical Formula

5. If the subscripts are not whole numbers, multiply all the subscripts by a small whole number (see the following table) to arrive at whole-number subscripts.

Practice

- Determine the empirical formula for a compound that contains 35.98% aluminum and 64.02% sulfur.

Practice

- The chemical analysis of aspirin indicates that the molecule is 60.00% carbon, 4.44% hydrogen, and 35.56% oxygen. Determine the empirical formula for aspirin.

Molecular Formula

- The molecular formula specifies the actual number of atoms of each element in one molecule or formula unit of the substance.
- Molecular formula is always a whole-number multiple of the empirical formula.

Molecular Formula

- Use the molar mass and the empirical formula molar mass to determine *n*
  - n = the integer by which you must multiply the empirical formula to get the molecular formula

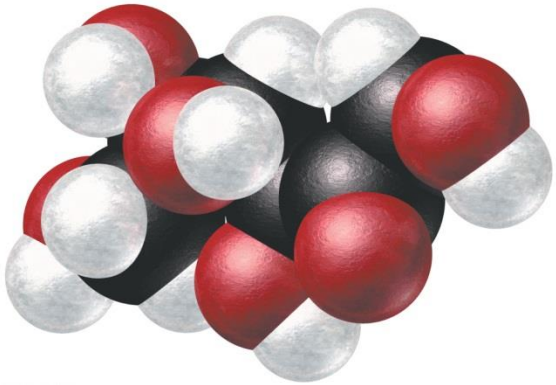
$$n = \frac{\text{Molar mass}}{\text{Empirical formula molar mass}}$$

Molecular Formula

- Multiply the subscripts in the empirical formula by *n* to arrive at the molecular formula.

Molecular Formula

- Find the molecular formula for fructose (a sugar found in fruit) from its empirical formula, CH_2O, and its molar mass, 180.2 g/mol.



Molecular Formula

- For fructose, the empirical formula molar mass is as follows:
  - Empirical formula molar mass = 1(12.01) + 2(1.01) + 16.00 = 30.03 g/mol

$$n = \frac{180.2\ \text{g/mol}}{30.03\ \text{g/mol}} = 6$$

Molecular Formula

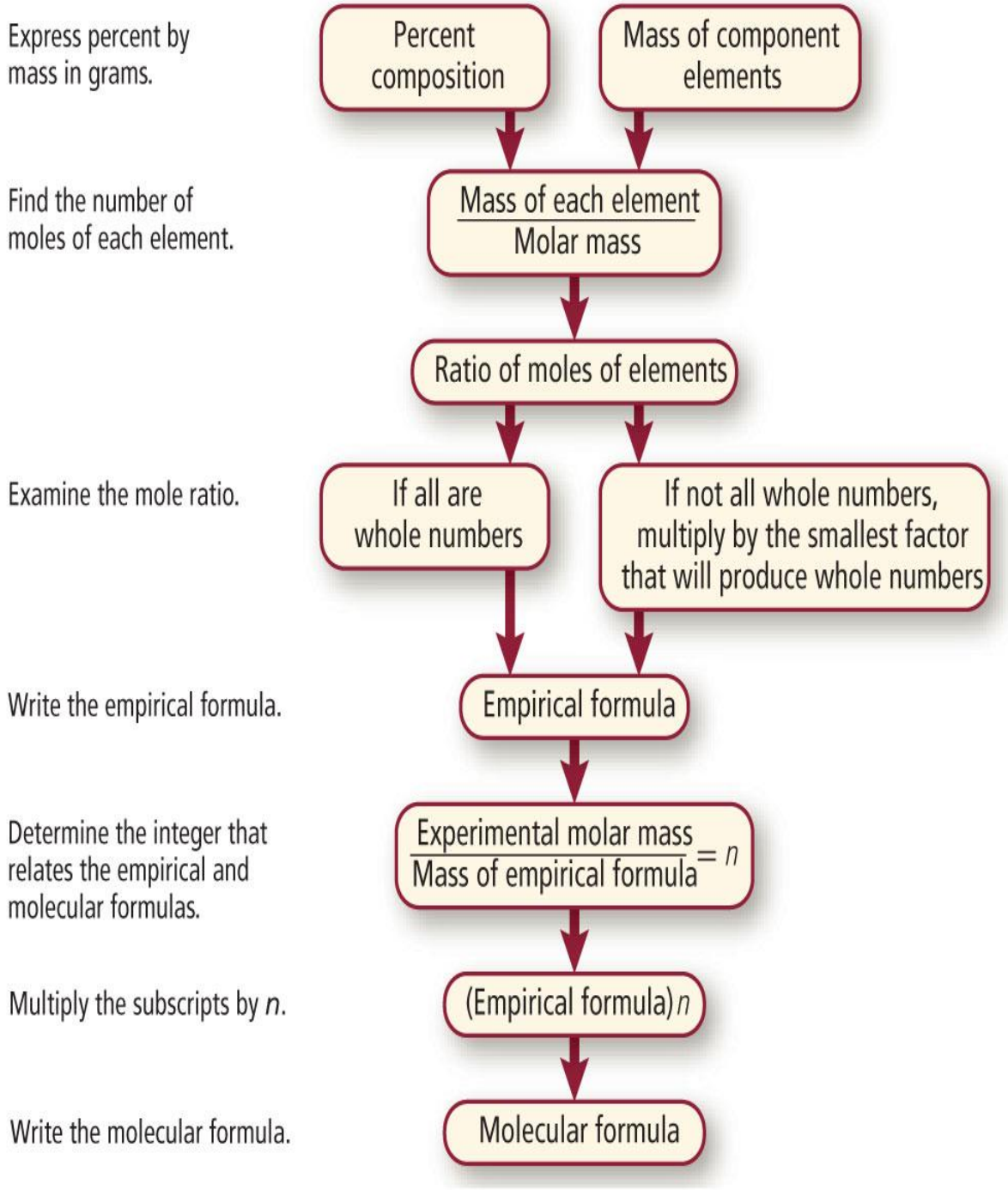
- We can then use this value of *n* to find the molecular formula.
- Molecular formula =
  - $CH_2O \times 6 =$

Practice

- Succinic acid is composed of 40.68% carbon, 5.08% hydrogen, and 54.24% oxygen and has a molar mass of 118.1 g/mol. Determine the empirical and molecular formulas for succinic acid.

Practice

- A compound was found to contain 49.98 g of carbon and 10.47 g of hydrogen. The molar mass of the compound is 58.12 g/mol. Determine the molecular formula.


Express percent by mass in grams.
Percent composition
Mass of component elements
Find the number of moles of each element.
Mass of each element / Molar mass
Ratio of moles of elements
Examine the mole ratio.
If all are whole numbers
If not all whole numbers, multiply by the smallest factor that will produce whole numbers
Write the empirical formula.
Empirical formula
Determine the integer that relates the empirical and molecular formulas.
Experimental molar mass / Mass of empirical formula = n
Multiply the subscripts by n.
(Empirical formula)n
Write the molecular formula.
Molecular formula