Empirical Formula - A formula that gives the simplest whole-number ratio of atoms in a compound.

Steps for Determining an Empirical Formula

1. Start with the number of grams of each element, given in the problem.
   - If percentages are given, assume that the total mass is 100 grams so that

     the mass of each element = the percent given.

2. Convert the mass of each element to moles using the molar mass from the periodic table.
3. Divide each mole value by the smallest number of moles calculated.
4. Round to the nearest whole number. This is the mole ratio of the elements and is

   represented by subscripts in the empirical formula.

   - If the number is too far to round (x.1 ~ x.9), then multiply each solution by the same

     factor to get the lowest whole number multiple.

     - e.g. If one solution is 1.5, then multiply each solution in the problem by 2 to get 3.
     - e.g. If one solution is 1.25, then multiply each solution in the problem by 4 to get 5.

Once the empirical formula is found, the molecular formula for a compound can be determined if the molar mass of the compound is known. Simply calculate the mass of the empirical formula and divide the molar mass of the compound by the mass of the empirical formula to find the ratio between the molecular formula and the empirical formula. Multiply all the atoms (subscripts) by this ratio to find the molecular formula. (See Example #2)

Example Problem #1

A compound was analyzed and found to contain 13.5 g Ca, 10.8 g O, and 0.675 g H. What is the empirical formula of the compound?

Start with the number of grams of each element, given in the problem.

Given ↓

13.5 g Ca

10.8 g O

0.675 g H

Convert the mass of each element to moles using the molar mass from the periodic table.

Given ↓ P.T. ↓

$$13.5\ \cancel{\text{g Ca}} \times \frac{1\ \text{mol Ca}}{40.1\ \cancel{\text{g Ca}}} = 0.337\ \text{mol Ca}$$

$$10.8\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.0\ \cancel{\text{g O}}} = 0.675\ \text{mol O}$$

$$0.675\ \cancel{\text{g H}} \times \frac{1\ \text{mol Ca}}{1.01\ \cancel{\text{g H}}} = 0.668\ \text{mol H}$$

Divide each mole value by the smallest number of moles calculated. Round to the nearest whole number.

Given ↓ P.T. ↓

$$13.5\ \cancel{\text{g Ca}} \times \frac{1\ \text{mol Ca}}{40.1\ \cancel{\text{g Ca}}} = \frac{0.337}{0.337}\ \text{mol Ca} \Rightarrow 1.00$$

$$10.8\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.0\ \cancel{\text{g O}}} = \frac{0.675}{0.337}\ \text{mol O} \Rightarrow 2.00$$

$$0.675\ \cancel{\text{g H}} \times \frac{1\ \text{mol Ca}}{1.01\ \cancel{\text{g H}}} = \frac{0.668}{0.337}\ \text{mol H} \Rightarrow 1.98 \approx 2.00$$

This is the mole ratio of the elements and is represented by subscripts in the empirical formula.

$$\boxed{CaO_2H_2 \Rightarrow Ca(OH)_2}$$

Example Problem #2

NutraSweet is 57.14% C, 6.16% H, 9.52% N, and 27.18% O. Calculate the empirical formula of NutraSweet and find the molecular formula. (The molar mass of NutraSweet is 294.30 g/mol)

Start with the number of grams of each element, given in the problem.

- If percentages are given, assume that the total mass is 100 grams so that the mass of each element = the percent given.

Given $\downarrow$

57.14 g C

6.16 g H

9.52 g N

27.18 g O

Convert the mass of each element to moles using the molar mass from the periodic table.

Given $\downarrow$ P.T. $\downarrow$

$$57.14\ \cancel{g\,C} \times \frac{1\ \text{mol C}}{12.0\ \cancel{g\,C}} = 4.76\ \text{mol C}$$

$$6.16\ \cancel{g\,H} \times \frac{1\ \text{mol H}}{1.01\ \cancel{g\,H}} = 6.10\ \text{mol H}$$

$$9.52\ \cancel{g\,N} \times \frac{1\ \text{mol N}}{14.0\ \cancel{g\,N}} = 0.68\ \text{mol N}$$

$$27.18\ \cancel{g\,O} \times \frac{1\ \text{mol O}}{16.0\ \cancel{g\,O}} = 1.70\ \text{mol O}$$

Divide each mole value by the smallest number of moles calculated. Round to the nearest whole number.

Given ↓ P.T. ↓

$$57.14\ \cancel{g\,C}\ x\ \frac{1\ \text{mol C}}{12.0\ \cancel{g\,C}} = \frac{4.76}{0.68}\ \text{mol C} \Rightarrow 7$$

$$6.16\ \cancel{g\,H}\ x\ \frac{1\ \text{mol H}}{1.01\ \cancel{g\,H}} = \frac{6.10}{0.68}\ \text{mol H} \Rightarrow 8.97 \approx 9$$

$$9.52\ \cancel{g\,N}\ x\ \frac{1\ \text{mol N}}{14.0\ \cancel{g\,N}} = \frac{0.68}{0.68}\ \text{mol N} \Rightarrow 1$$

$$27.18\ \cancel{g\,O}\ x\ \frac{1\ \text{mol O}}{16.0\ \cancel{g\,O}} = \frac{1.70}{0.68}\ \text{mol O} \Rightarrow 2.5$$

This is the mole ratio of the elements and is represented by subscripts in the empirical formula.

- If the number is too far to round (x.1 ~ x.9), then multiply each solution by the same
factor to get the lowest whole number multiple.

Given ↓ P.T. ↓

$$57.14\ \cancel{g\,C}\ x\ \frac{1\ \text{mol C}}{12.0\ \cancel{g\,C}} = \frac{4.76}{0.68}\ \text{mol C} \Rightarrow 7(2) = 14$$

$$6.16\ \cancel{g\,H}\ x\ \frac{1\ \text{mol H}}{1.01\ \cancel{g\,H}} = \frac{6.10}{0.68}\ \text{mol H} \Rightarrow 8.97 \approx 9(2) = 18$$

$$9.52\ \cancel{g\,N}\ x\ \frac{1\ \text{mol N}}{14.0\ \cancel{g\,N}} = \frac{0.68}{0.68}\ \text{mol N} \Rightarrow 1(2) = 2$$

$$27.18\ \cancel{g\,O}\ x\ \frac{1\ \text{mol O}}{16.0\ \cancel{g\,O}} = \frac{1.70}{0.68}\ \text{mol O} \Rightarrow 2.5(2) = 5$$

↑ too far to round, multiply
to get a whole number

$$\text{Empirical formula} = \boxed{C_{14}H_{18}N_2O_5}$$

Now, we can find the molecular formula by finding the mass of the empirical formula and setting up a ratio:

$$C_{14}H_{18}N_2O_5 = 14(12.0g)+18(1.01g)+2(14.0g)+5(16.0g)=294\ g/mol$$

from P.T. $\longrightarrow$ C H N O

$$\frac{\text{molar mass}}{\text{empirical formula}} = \frac{294.30\ g/mol}{294\ g/mol} \approx 1$$

So the empirical formula is the molecular formula

$$\boxed{C_{14}H_{18}N_2O_5}$$