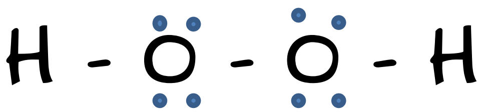
Chemistry Lecture #55: Empirical Formulas

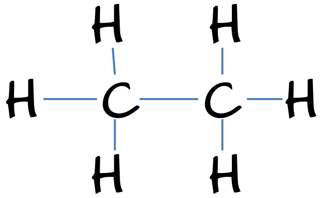
The empirical formula of a compound gives the smallest, whole number ratio of atoms in a compound.

A molecular formula gives the actual numbers of atoms in a molecule.

To illustrate the meaning and difference between an empirical and molecular formula, look at the Lewis structure of hydrogen peroxide below.



The structure above is made of two hydrogens and two oxygens, The ratio of hydrogen to oxygen in the compound is 2:2, which can be reduced to 1:1. Thus, the empirical formula would be H_1O_1 , or just HO. The molecular formula would be H_2O_2 .



The structure above is made of 2 carbons and 6 hydrogens, which gives us a ratio of 2:6 or 1:3. The empirical formula would be C_1H_3 or CH_3 . The molecular formula would be C_2H_6 .

We can calculate the empirical formula of a compound if we know its percent composition by following the steps below:

1. Assume you have 100 g of compound and that % is the same as grams. Convert percent or grams of the constituent elements into moles.

2. Find the smallest whole number ratio of moles by dividing by the smallest number of moles.

3. If you don't get small whole numbers, try multiplying your answer from step #2 by the numbers 2 through 8 to convert the values to something close to whole numbers. Use the if/then chart as a guide in deciding how to multiply the numbers.

If it ends with	then multiply by
.125	8
.6	5
.25	4
.33	3
.5	2
.66	3
.75	4

A compound is 25.9 % nitrogen and 74.1 % oxygen. Find the empirical formula.

$$\frac{25.9 \text{ g N}}{1} \times \frac{\text{mole N}}{14.0 \text{ g N}} = 1.85 \text{ moles N}$$

$$\frac{74.1 \text{ g O}}{1} \times \frac{\text{mole O}}{16.0 \text{ g O}} = 4.63 \text{ moles O}$$

The smallest number of moles is 1.85, so we divide both moles by 1.85.

$$\frac{1.85 \text{ moles N}}{1.85} = 1 \text{ mole N}$$

$$\frac{4.63 \text{ moles O}}{1.85} = 2.50 \text{ moles O}$$

We should have gotten whole numbers, but 2.5 is not a whole number. 2.5 ends with .5. The if/then chart tells us to multiply everything by 2 to get the smallest whole number ratios

1 mole N x 2 = 2 moles N
2.50 moles O x 2 = 5 moles O

The empirical formula is N_2O_5

Find the empirical formula of a compound that is 87.2 % C and 12.8 % H.

$$\frac{87.2\text{ g C}}{1} \times \frac{\text{mole C}}{12.0\text{ g C}} = 7.27\text{ moles C}$$

$$\frac{12.8\text{ g H}}{1} \times \frac{\text{mole H}}{1.01\text{ g H}} = 12.7\text{ moles H}$$

The smallest number of moles is 7.27 , so we divide by this number.

$$\frac{7.27\text{ moles C}}{7.27} = 1\text{ mole C}$$

$$\frac{12.7\text{ moles H}}{7.27} = 1.75\text{ moles H}$$

1.75 is not close to a whole number. Since it ends with .75, the if/then chart says to multiply everything by 4.

1 mole C x 4 = 4 moles C
1.75 moles H x 4 = 7 moles H

The empirical formula is C_4H_7 .

Find the empirical formula of a compound that is 52.1 % C, 13.1 % H, and 34.7 % O.

$$\frac{52.1 \text{ g C}}{1} \times \frac{\text{mole C}}{12.0 \text{ g C}} = 4.34 \text{ moles C}$$

$$\frac{13.1 \text{ g H}}{1} \times \frac{\text{mole H}}{1.01 \text{ g H}} = 13.0 \text{ moles H}$$

$$\frac{34.7 \text{ g O}}{1} \times \frac{\text{mole O}}{16.0 \text{ g O}} = 2.17 \text{ moles O}$$

The smallest number of moles is 2.17, so we divide by this number.

$$\frac{4.34 \text{ moles C}}{2.17} = 2.00 \text{ moles C}$$

$$\frac{13.0 \text{ moles H}}{2.17} = 5.99 \text{ moles H (which we'll round to 6)}$$

$$\frac{2.17 \text{ moles O}}{2.17} = 1 \text{ mole O}$$

We end up with values that are very close to whole numbers, so we don't need to use the if/then chart. If we round 5.99 to 6, then the empirical formula is C_2H_6O .