Calculating Empirical Formulas

An **empirical formula** provides the lowest whole-number ratio of the elements in a compound.

- Talking in mole form, an empirical formula is the lowest whole-number ratio of moles of atoms in a compound.

- sometimes empirical & molecular formulas for compounds are the same.

ex) CO_2 is an empirical formula & molecular formula for carbon dioxide whereas NH_2 is the empirical formula for N_2H_4

ex) What is empirical formula of compound made up of 60.3% Mg & 39.7% O?

This means in a mass of 100g this compound, 60.3g are Mg & 39.7g are O.

molar mass

Change ratio of masses to ratio of moles by GFM

$$\frac{60.3\text{ g Mg}}{} \times \frac{1\text{ mol Mg}}{24.3\text{ g}} = 2.48\text{ mol Mg}$$

$$\frac{39.7\text{ g O}}{} \times \frac{1\text{ mol O}}{16.0\text{ g}} = 2.48\text{ mol O}$$

since the ratio of moles of Mg to O is 2.48 : 2.48 we can simplify by dividing by common factors, which reduces our ratio to 1:1

→ this means our empirical formula is 1 Mg : 1 O or $\boxed{MgO}$ (which is the molecular formula too)

ex) 222.6 g Na combines w/ 77.4 g O

$$\frac{222.6\text{ g Na}}{} \times \frac{1\text{ mol Na}}{23.0\text{ g}} = 9.68\text{ mol Na}$$

$$\frac{77.4\text{ g O}}{} \times \frac{1\text{ mol O}}{16.0\text{ g}} = 4.84\text{ mol O}$$

Our ratio is 9.68 Na : 4.84 O. Simply by dividing by 4.84 and we get 2:1

so our empirical formula is $\boxed{Na_2O}$

74.05% C, 7.46% H, 9.86% O, 8.63% N

from P.T.

$$\frac{74.05\ g\ C}{} \left| \frac{1\ mol}{12.0\ g} \right. = \frac{6.2\ mol}{.62} \quad 10$$

$$\frac{7.46\ g\ H}{} \left| \frac{1\ mol}{1.0\ g} \right. = \frac{7.46\ mol}{.62} \quad 12$$

$$\frac{9.86\ g\ O}{} \left| \frac{1\ mol}{16.0\ g} \right. = \frac{0.62\ mol}{.62} \quad 1$$

$$\frac{8.63\ g\ N}{} \left| \frac{1\ mol}{14.0\ g} \right. = \frac{0.62\ mol}{.62} \quad 1$$

$$10C : 12H : 10 : 1N = \boxed{C_{10}H_{12}ON}$$

Empirical and Molecular Formulas Cheat Sheet

Express percent by mass in grams.

Percent composition

Mass of component elements

Find the number of moles of each element.

$$\frac{\text{Mass of each element}}{\text{Molar mass}}$$

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.

$$\frac{\text{Experimental molar mass}}{\text{Mass of empirical formula}} = n$$

Multiply the subscripts by n.

(Empirical formula)n

Write the molecular formula.

Molecular formula