Note that the percentage composition of C_2H_6 is found as follows.

Molar mass = 30.0 g

$$\% \text{ of C in compound} = \frac{24.0\ g}{30.0\ g} \times 100\% = 80.0\%$$

$$\% \text{ of H in compound} = \frac{6.0\ g}{30.0\ g} \times 100\% = 20.0\%$$

Therefore, both CH_3 and C_2H_6 contain 80.0% C and 20.0% H, and both have three times as many hydrogen atoms as carbon atoms. The **SIMPLEST** ratio which describes the relative number of carbons and hydrogens is

C_1H_3 or simply $\mathbf{CH_3}$.

EXAMPLE: **A compound contains 58.5% C, 7.3% H and 34.1% N. What is the empirical formula of the compound?**

Assume 100.0 g of the compound is taken.

mass of C = 58.5 g , mass of H = 7.3 g , mass of N = 34.1 g

$$\begin{array}{ll|l} & & \div 2.44 \\ \text{moles C} = 58.5\ g \times \frac{1\ mol}{12.0\ g} & = 4.88\ mol & 2 \\ \text{moles H} = 7.3\ g \times \frac{1\ mol}{1.0\ g} & = 7.3\ mol & 2.99 \simeq 3 \\ \text{moles N} = 34.1\ g \times \frac{1\ mol}{14.0\ g} & = 2.44\ mol & 1 \end{array}$$

Hence the empirical formula is $\mathbf{C_2H_3N}$.

EXAMPLE: **What is the empirical formula of a compound containing 81.8% C and 18.2% H?**

Assume 100.0 g of the compound is taken.

mass of C = 81.8 g , mass of H = 18.2 g

$$\begin{array}{ll|l|l} & & \div 6.82 & \times 3 \\ \text{moles C} = 81.8\ g \times \frac{1\ mol}{12.0\ g} & = 6.82\ mol & 1 & 3 \\ \text{moles H} = 18.2\ g \times \frac{1\ mol}{1.0\ g} & = 18.2\ mol & 2.67 & 8 \end{array}$$

Note that a third column is added to the above calculation. Dividing the calculated number of moles by the smallest number of moles present (6.82) leaves a fraction (2.67) in the second column. This is a problem because the empirical formula is a WHOLE–NUMBER ratio. Since the ".67" ending of 2.67 indicates the presence of a fraction involving "thirds", clear the fraction by multiplying the values in the second column by 3, so as to produce the whole–number values in the third column.

Therefore the empirical formula is $\mathbf{C_3H_8}$.

IMPORTANT: You must be able to recognize the following fractions and their decimal equivalents.

$0.20 = {}^1/_5$	$0.40 = {}^2/_5$	$0.67 = {}^2/_3$
$0.25 = {}^1/_4$	$0.50 = {}^1/_2$	$0.75 = {}^3/_4$
$0.33 = {}^1/_3$	$0.60 = {}^3/_5$	$0.80 = {}^4/_5$

SNEAKY TRICK: You don't have to re–write fractions such as 2.67 in the form ${}^8/_3$. All you have to do is to recognize that numbers such as 2.67, 1.33, 5.67 and 3.33 involve **THIRDS** and simply multiply the fraction by **3** to clear the fraction. Similarly, numbers like 1.75, 2.25 and 3.75 involve **QUARTERS,** so that multiplying by **4** will clear such fractions.

INCREDIBLY, VITALLY IMPORTANT NOTE:

Always carry out calculations to 3 or 4 digits and NEVER round off intermediate values. The numbers 3.60, 3.67, 3.75 and 3.80 are very close to one another and improper round–off of calculations will cause you to multiply by the wrong number when trying to "clear fractions".

EXAMPLE: **What is the empirical formula of a compound containing 39.0% Si and 61.0% O?**

Assume 100.0 g of the compound is taken.

mass of Si = 39.0 g , mass of O = 61.0 g

	$\div 1.39$	x 4
moles Si $= 39.0\,\text{g} \times \dfrac{1\ \text{mol}}{28.1\ \text{g}} = 1.39$ mol	1	4
moles O $= 61.0\,\text{g} \times \dfrac{1\ \text{mol}}{16.0\ \text{g}} = 3.81$ mol	$2.74 \cong 2.75$	11

The empirical formula is Si_4O_{11}.

EXERCISE:

46. Find the empirical formula for the following compounds.
(a) 15.9% B, 84.1% F
(b) 87.5% Si, 12.5% H
(c) 43.7% P, 56.3% O
(d) 77.9% I, 22.1% O
(e) 77.7% Fe, 22.3% O
(f) 70.0% Fe, 30.0% O
(g) 72.4% Fe, 27.6% O
(h) 46.3% Li, 53.7% O
(i) 24.4% C, 3.39% H, 72.2% Cl
(j) 26.6% K, 35.4% Cr, 38.0% O
(k) 21.8% Mg, 27.9% P, 50.3% O
(l) 3.66% H, 37.8% P, 58.4% O
(m) 46.2% C, 7.69% H, 46.2% O
(n) 50.5% C, 5.26% H, 44.2% N

FINDING THE MOLECULAR FORMULA

If the empirical formula can be found, it is straightforward to calculate the molar mass of the empirical formula; that is, the **empirical mass.**

The first example in this section pointed out that all of CH_2, C_2H_4, C_3H_6, C_4H_8 and C_5H_{10} have identical empirical formulae. Since all of these compounds have formulae which are whole–number multiples of CH_2, then the molar mass of all of the compounds must be a whole–number multiple of the empirical mass of CH_2.

Let N = the WHOLE NUMBER multiple of the empirical mass.

$$\text{multiple} = N = \frac{\text{molar mass}}{\text{empirical mass}}$$