

Answers

A Simple Rhyme for a Simple Formula, by Joel S. Thompson

- Percent to mass
- Mass to mole
- Divide by small
- Multiply 'til whole

Calculate the percent composition of each element in each compound listed.

1) KNO_3 Molar mass = $(1 \times 39.10) + (1 \times 14.01) + (3 \times 16.00) = 101.11 \text{ g/mol}$
 $\%K = \frac{39.10}{101.11} \times 100 = 38.67\% K$ $\%N = \frac{14.01}{101.11} \times 100 = 13.86\% N$ $\%O = \frac{3 \times 16.00}{101.11} \times 100 = 47.47\% O$

2) H_2SO_4 Molar mass = $(2 \times 1.008) + (1 \times 32.07) + (4 \times 16.00) = 98.09 \text{ g/mol}$
 $\%H = \frac{2 \times 1.008}{98.09} \times 100 = 2.05\% H$ $\%S = \frac{32.07}{98.09} \times 100 = 32.69\% S$ $\%O = \frac{4 \times 16.00}{98.09} \times 100 = 65.25\% O$

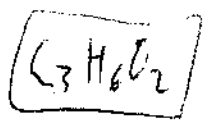
3) C_2H_5OH Molar mass = $(2 \times 12.01) + (6 \times 1.008) + (1 \times 16.00) = 46.07 \text{ g/mol}$
 $\%C = \frac{2 \times 12.01}{46.07} \times 100 = 52.14\% C$ $\%H = \frac{6 \times 1.008}{46.07} \times 100 = 13.13\% H$ $\%O = \frac{16.00}{46.07} \times 100 = 34.73\% O$

4) $C_6H_5NH_2$ Molar mass = $(6 \times 12.01) + (7 \times 1.008) + (1 \times 14.01) = 93.13 \text{ g/mol}$
 $\%C = \frac{6 \times 12.01}{93.13} \times 100 = 77.38\% C$ $\%H = \frac{7 \times 1.008}{93.13} \times 100 = 7.57\% H$ $\%N = \frac{14.01}{93.13} \times 100 = 15.04\% N$

Empirical Formula Examples

1) A compound is found to have (by mass) 48.38% carbon, 8.12% hydrogen and the rest oxygen. What is its empirical formula? Assume 100g. Mass of oxygen = $100 - 48.38 - 8.12 = 43.50 \text{ g}$

$48.38 \text{ g C} \times \left(\frac{1 \text{ mol C}}{12.01 \text{ g}}\right) = 4.0283 \text{ mol C} / 2.7188 = 1.48 = \frac{1}{2} \times 2 = 3$
 $8.12 \text{ g H} \times \left(\frac{1 \text{ mol H}}{1.008 \text{ g}}\right) = 8.0556 \text{ mol H} / 2.7188 = 2.963 = 3 \times 2 = 6$
 $43.50 \text{ g O} \times \left(\frac{1 \text{ mol O}}{16.00 \text{ g}}\right) = 2.7188 \text{ mol O} / 2.7188 = 1 \times 2 = 2$

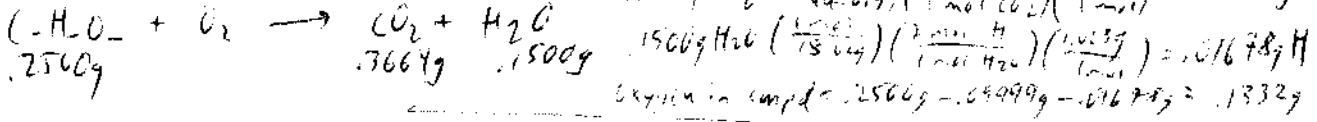


2) A compound is found to have 46.67% nitrogen, 6.70% hydrogen, 19.98% carbon and 26.65% oxygen. What is its empirical formula? Assume 100g.

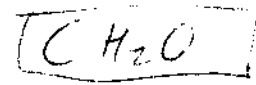
$46.67 \text{ g N} \left(\frac{1 \text{ mol N}}{14.01 \text{ g}}\right) = 3.3312 \text{ mol N} / 1.6636 = 2.002 = 2$
 $6.70 \text{ g H} \left(\frac{1 \text{ mol H}}{1.008 \text{ g}}\right) = 6.6468 \text{ mol H} / 1.6636 = 3.995 = 4$
 $19.98 \text{ g C} \left(\frac{1 \text{ mol C}}{12.01 \text{ g}}\right) = 1.6636 \text{ mol C} / 1.6636 = 1$
 $26.65 \text{ g O} \left(\frac{1 \text{ mol O}}{16.00 \text{ g}}\right) = 1.6656 \text{ mol O} / 1.6636 = 1.001 = 1$



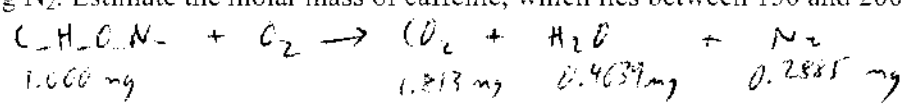
- 6) A 0.2500 g sample of a compound known to contain carbon, hydrogen and oxygen undergoes complete combustion to produce 0.3664 g of CO₂ and 0.1500 g of H₂O. What is the empirical formula of this compound?



$0.9999\text{g C} \left(\frac{1\text{ mol C}}{12.01\text{g}}\right) = 0.08326\text{ mol C} / 0.08325 = 1.001 = 1$
 $0.1678\text{g H} \left(\frac{1\text{ mol H}}{1.008\text{g}}\right) = 0.1665\text{ mol H} / 0.08325 = 2$
 $0.0823\text{g O} \times \left(\frac{1\text{ mol O}}{16.00\text{g}}\right) = 0.00514\text{ mol O} / 0.00514 = 1$



- 7) Caffeine, a stimulant found in coffee, tea, and certain soft drinks, contains C, H, O, and N. Combustion of 1.000 mg of caffeine produces 1.813 mg CO₂, 0.4639 mg H₂O, and 0.2885 mg N₂. Estimate the molar mass of caffeine, which lies between 150 and 200 g/mol.



Ratio is same, grams vs milligrams

$1.813\text{mg CO}_2 \left(\frac{1\text{ mol}}{44.01\text{g}}\right) \left(\frac{1\text{ mol C}}{1\text{ mol CO}_2}\right) \left(\frac{12.01\text{g}}{1\text{ mol}}\right) = 0.4948\text{mg}$
 $0.4639\text{mg H}_2\text{O} \left(\frac{1\text{ mol}}{18.02\text{g}}\right) \left(\frac{2\text{ mol H}}{1\text{ mol H}_2\text{O}}\right) \left(\frac{1.008\text{g}}{1\text{ mol}}\right) = 0.0519\text{mg}$
 $0.2885\text{mg N}_2 \left(\frac{1\text{ mol}}{2 \times 14.01\text{g}}\right) \left(\frac{2\text{ mol N}}{1\text{ mol N}_2}\right) \left(\frac{14.01\text{g}}{1\text{ mol}}\right) = 0.2885\text{mg N}$
 Oxygen in compd = $1.000\text{mg} - 0.4948\text{mg} - 0.0519\text{mg} - 0.2885\text{mg} = 0.1648\text{mg O}$

$0.4948\text{g C} \left(\frac{1\text{ mol C}}{12.01\text{g}}\right) = 0.04120\text{ mol C} / 0.0103 = 4$
 $0.0519\text{g H} \left(\frac{1\text{ mol H}}{1.008\text{g}}\right) = 0.05149\text{ mol H} / 0.0103 = 5$
 $0.2885\text{g N} \left(\frac{1\text{ mol N}}{14.01\text{g}}\right) = 0.02059\text{ mol N} / 0.0103 = 2$
 $0.1648\text{g O} \left(\frac{1\text{ mol O}}{16.00\text{g}}\right) = 0.0103\text{ mol O} / 0.0103 = 1$

Empirical Formula **C₄H₅N₂O**

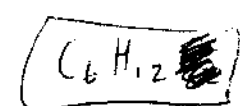
molar mass = $4 \times 12.01 + 5 \times 1.008 + 2 \times 14.01 + 1 \times 16.00 = 97.10\text{g}$

molar mass estimate **194.20 g/mol**

- 8) A compound is known to have an empirical formula of CH and a molar mass of 78.11 g/mol. What is its molecular formula?

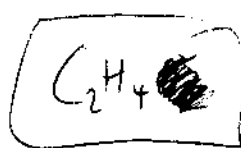
molar mass of FF = $12.01 + 1.008 = 13.02$

$\frac{78.11}{13.02} = 6$ ← multiplier
FF = EF → MF



- 9) Another compound, also with an empirical formula of CH is found to have a molar mass of 26.04 g/mol. What is its molecular formula?

$\frac{26.04}{13.02} = 2$



10) What are the empirical and molecular formulas for a compound with 86.88% carbon and 13.12% hydrogen and a molecular weight of about 345?

Assume 100g

$$86.88g C \left(\frac{1 \text{ mol } C}{12.01g} \right) = 7.234 \text{ mol } C / 7.234 = 1 \quad \times 5 = 5$$

$$13.12g H \left(\frac{1 \text{ mol } H}{1.008g} \right) = 13.016 \text{ mol } H / 7.234 = 1.799 = 1.8 = 1 \frac{4}{5} \times 5 = 9$$

Empirical Formula = C_5H_9 ← molar mass = $(5 \times 12.01) + (9 \times 1.008) = 69.12g/mol$
 $345/69 = 5 \rightarrow$ Molecular Formula = $C_{25}H_{45}$

11) What are the empirical and molecular formulas for a compound with 83.625% carbon and 16.375% hydrogen and a molecular weight of 388.78?

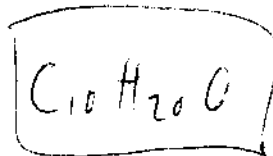
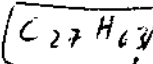
Assume 100g

$$83.625g C \left(\frac{1 \text{ mol } C}{12.01g} \right) = 6.96244 \text{ mol } C / 6.96244 = 1 \quad \times 3 = 3$$

$$16.375g H \left(\frac{1 \text{ mol } H}{1.008g} \right) = 16.245 \text{ mol } H / 6.96244 = 2.333 \times 3 = 7$$

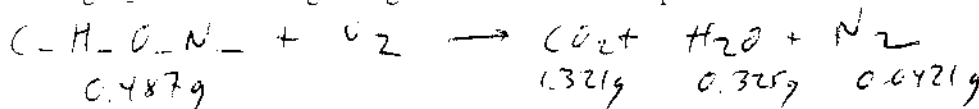
Empirical Formula = C_3H_7 ← molar mass = $(3 \times 12.01) + (7 \times 1.008) = 42.09g/mol$
 $388.78/42.09 = 9 \rightarrow$ molecular formula =

12) 0.1005 g of menthol is combusted, producing 0.2829 g of CO_2 and 0.1159 g of H_2O .
 What is menthol's empirical formula?



— see problem #5

13) 0.487 grams of quinine (molar mass = 324 g/mol) is combusted and found to produce 1.321 g CO_2 , 0.325 g H_2O and 0.0421 g nitrogen. Determine the empirical and molecular formulas.



$$0.0421g N \left(\frac{1 \text{ mol } N}{14.01g} \right) = 0.00301 \text{ mol } N / 0.00301 = 1$$

$$0.366g C \left(\frac{1 \text{ mol } C}{12.01g} \right) = 0.03052 \text{ mol } C / 0.00301 = 10$$

$$0.03636g H \left(\frac{1 \text{ mol } H}{1.008g} \right) = 0.03607 \text{ mol } H / 0.00301 = 12$$

$$0.04804g O \left(\frac{1 \text{ mol } O}{16.00g} \right) = 0.003003 \text{ mol } O / 0.00301 = 1$$

Empirical Formula = $C_{10}H_{12}ON$

molar mass of $C_{10}H_{12}ON$
 $(10 \times 12.01) + (12 \times 1.008) + (16.00) + (14.01) = 162.21g/mol$

$\frac{324}{162} = 2 \rightarrow$ molecular formula = $C_{20}H_{24}O_2N_2$

Oxygen in 0.487g sample =
 $- 0.0421g N$
 $- 0.366g C$
 $- 0.03636g H$
 $\hline 0.04804g O$

