Answers

Miramar College, Chemistry 152 Practice Empirical Formulas

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.

Molar mass = (1×39.10)+(1×14.01)+(3×16.00) = 101.11 g/mol

1) KNO_3 %K = 39.10/101.11 ×100 = 38.67% K | %N = 14.01/101.11 ×100 = 13.86% N | %O = 3×16.00/101.11 ×100 = 47.47% O

2) H_2SO_4 Molar mass = (2×1.008)+(1×32.07)+(4×16.00) = 98.09 g/mol

%H = 2×1.008/98.09 ×100 = 2.055% H | S: 32.07/98.09 ×100 = 32.69% S | O: 4×16.00/98.09 ×100 = 65.25% O

3) C_2H_5OH molar mass = (2×12.01)+(6×1.008)+(1×16.00) = 46.07 g/mol

%C = 2×12.01/46.07 ×100 = 52.14% C | H: 6×1.008/46.07 ×100 = 13.13% H | O: 16.00/46.07 ×100 = 34.73% O

4) $C_6H_5NH_2$ Molar mass = (6×12.01)+(7×1.008)+(1×14.01) = 93.13 g/mol

C: 6×12.01/93.13 ×100 = 77.38% C | H: 7×1.008/93.13 ×100 = 7.577% H | N: 14.01/93.13 ×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 100 g

Mass of oxygen = 100 − 8.12 − 48.38 = 43.50 g

48.38 g C × (1 mol C / 12.01 g) = 4.0283 mol C / 2.7188 = 1.48 = 1½ × 2 = 3
8.12 g H × (1 mol H / 1.008 g) = 8.0556 mol H / 2.7188 = 2.963 = 3 × 2 = 6
43.50 g O × (1 mol O / 16.00 g) = 2.7188 mol O / 2.7188 = 1 × 2 = 2

$C_3H_6O_2$

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 100 g.

46.67 g N (1 mol N / 14.01 g) = 3.3312 mol N / 1.6656 = 2.002 = 2
6.70 g H (1 mol H / 1.008 g) = 6.6468 mol H / 1.6656 = 3.995 = 4
19.98 g C (1 mol C / 12.01 g) = 1.1636 mol C / 1.6656 = 1
26.65 g O (1 mol O / 16.00 g) = 1.6656 mol O / 1.6656 = 1.001 = 1

N_2H_4CO

Answers

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 CO_2 and 0.1500 g of H_2O. What is the empirical formula of this compound?

$C\_H\_O\_ + O_2 \rightarrow CO_2 + H_2O$
.2500g → .3664g, .1500g

.3664g CO_2 × (1 mol / 44.01g)(1 mol C / 1 mol CO_2)(12.01g / 1 mol) = .09999 g C
.1500g H_2O (1 mol / 18.02g)(2 mol H / 1 mol H_2O)(1.008g / 1 mol) = .01678g H
Oxygen in compd = .2500g − .09999g − .01678g = .1332g

.09999g C (1 mol C / 12.01g) = .008326 mol C / .008325 = 1.0001 = 1
.01678g H (1 mol H / 1.008g) = .01665 mol H / .008325 = 2
.1332g O × (1 mol O / 16.00g) = .008325 mol O / .008325 = 1

CH_2O

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 CO_2, 0.4639 mg H_2O, and 0.2885 mg N_2. Estimate the molar mass of caffeine, which lies between 150 and 200 g/mol.

$C\_H\_O\_N\_ + O_2 \rightarrow CO_2 + H_2O + N_2$
1.000 mg → 1.813 mg, 0.4639 mg, 0.2885 mg ← Ratio is same, grams or milligrams

1.813mg CO_2 (1 mol / 44.01g)(1 mol C / 1 mol CO_2)(12.01g / 1 mol) = 0.4948 mg
0.4639mg H_2O (1 mol / 18.02g)(2 mol H / 1 mol H_2O)(1.008g / 1 mol) = 0.05190 mg
0.2885mg N_2 (1 mol N_2 / 2×14.01g)(2 mol N / 1 mol N_2)(14.01g / 1 mol) = .2885 mg N ← (silly calc.) could be omitted
Oxygen in compd = 1.000 mg − .4948 − .0519 − .2885 = .1648 mg O

.4948g C (1 mol C / 12.01g) = .04120 mol C / .0103 = 4
.0519g H (1 mol H / 1.008g) = 0.05149 mol H / .0103 = 5
.2885g N (1 mol N / 14.01g) = .02059 mol N / .0103 = 2
.1648g O (1 mol O / 16.00g) = .0103 mol O / .0103 = 1

Empirical formula $C_4H_5N_2O$

molar mass = 4×12.01 + 5×1.008 + 2×14.01 + 1×16.00 = 97.10g

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 EF = 12.01 + 1.008 = 13.02

78.11 / 13.02 = 6 ← multiplier for EF → MF

C_6H_{12}

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

26.04 / 13.02 = 2

C_2H_4

Answers

3) A compound is found to have 1.121 g nitrogen, 0.161 g hydrogen, 0.480 g carbon and 0.640 g oxygen. What is its empirical formula? (Note that masses are given, NOT percentages.)

$1.121\,g\,N\left(\frac{1\,mol\,N}{14.01\,g}\right) = .080014\,mol\,N / .039967 = 2.002 = 2$

$0.161\,g\,H\left(\frac{1\,mol\,H}{1.008\,g}\right) = .15972\,mol\,H / .039967 = 3.996 = 4$

$0.480\,g\,C\left(\frac{1\,mol\,C}{12.01\,g}\right) = 0.039967\,mol\,C / .039967 = 1$

$0.640\,g\,O\left(\frac{1\,mol\,O}{16.00\,g}\right) = 0.040000\,mol\,O / .039967 = 1.0008 = 1$

N_2H_4CO

4) A compound containing only carbon, hydrogen and oxygen is found to be 48.38% carbon and 8.12% hydrogen by mass. What is its empirical formula?

% O = 100 − 48.38 − 8.12 = 43.50% O

Assume 100 g

48.38 g C

8.12 g H

43.50 g O

Same as #1

$C_3H_6O_2$

5) 95.6 mg of menthol (molar mass = 156 g/mol) are burned in oxygen gass to give 269 mg CO_2 and 110 mg H_2O. What is menthol's empirical formula?

molar masses

CO_2: (1×12.01)+(2×16.00) = 44.01 g/mol

H_2O: (2×1.008)+(1×16.00) = 18.02 g/mol

$C\_H\_O\_ + O_2 \rightarrow CO_2 + H_2O$

95.6 mg → 269 mg, 110 mg

$269\,mg\,CO_2\left(\frac{1\,g}{10^3\,mg}\right)\left(\frac{1\,mol\,CO_2}{44.01\,g}\right)\left(\frac{1\,mol\,C}{1\,mol\,CO_2}\right)\left(\frac{12.01\,g\,C}{1\,mol\,C}\right)\left(\frac{10^3\,mg}{1\,g}\right) = 73.4\,mg\,C$

$110\,mg\,H_2O\left(\frac{1\,g}{10^3\,mg}\right)\left(\frac{1\,mol\,H_2O}{18.02\,g}\right)\left(\frac{2\,mol\,H}{1\,mol\,H_2O}\right)\left(\frac{1.008\,g\,H}{1\,mol}\right)\left(\frac{10^3\,mg}{1\,g}\right) = 12.3\,mg\,H$

85.7 mg C and H

Remainder is oxygen

95.6 − 85.7 mg = 9.9 mg O

(Ratio would be same if mg or g.)

$73.4\,g\,C \times \left(\frac{1\,mol\,C}{12.01\,g}\right) = 6.112\,mol\,C / .6188 = 9.88$

$12.3\,g\,H\left(\frac{1\,mol\,H}{1.008\,g}\right) = 12.20\,mol\,H / .6188 = 19.7$

$9.9\,g\,O\left(\frac{1\,mol\,O}{16.00\,g}\right) = 0.6188\,mol\,O / .6188 = 1$

$C_{10}H_{20}O$

This is also the molecular formula.

↓

molar mass of empirical formula.

Need to "break" our rule of Thumb here, due to molar mass of 156 g/mol

= (10×12.01) + (20×1.008) + (16.00) = 156.26 g/mol

Answers

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 100 g

86.88 g C × (1 mol C / 12.01 g) = 7.234 mol C / 7.234 = 1 × 5 = 5

13.12 g H (1 mol H / 1.008 g) = 13.016 mol H / 7.234 = 1.799 = 1.8 = 1 4/5 × 5 = 9

Empirical Formula = C_5H_9 ← molar mass = (5 × 12.01) + (9 × 1.008) = 69.12 g/mol

345/69 = 5 → 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 100 g

83.625 g C (1 mol C / 12.01 g) = 6.96244 mol C / 6.96294 = 1 × 3 = 3

16.375 g H (1 mol H / 1.008 g) 16.245 mol H / 6.96294 = 2.333 × 3 = 7

Empirical Formula = C_3H_7 ← molar mass = (3 × 12.01) + (7 × 1.008) = 43.09 g/mol

388.78/43.09 = 9 → molecular formula = $C_{27}H_{63}$

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?

$C_{10}H_{20}O$ — 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.

C_H_O_N_ + O_2 → CO_2 + H_2O + N_2

0.487 g → 1.321 g, 0.325 g, 0.0421 g

.0421 g N (1 mol / 14.01 g) = .003011 mol N / .003003 = 1

.3605 g C (1 mol / 12.01 g) = .03002 mol C / .003003 = 10

.03636 g H (1 mol / 1.008 g) = .03607 mol H / .003003 = 12

.04804 g O (1 mol / 16.00 g) = .003003 mol O / .003003 = 1

Oxygen in 0.487 g sample = −.0421 g N − .3605 g C − .03636 g H = .04804 g O

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

molar mass of EF = (10 × 12.01) + (12 × 1.008) + (16.00) + (14.01) = 162.21 g/mol

324/162 = 2 → Molecular Formula = $C_{20}H_{24}O_2N_2$