

- A. After combustion with excess oxygen, a 12.501 g sample of a petroleum compound produced 38.196 g of carbon dioxide and 18.752 of water. A previous analysis determined that the compound does not contain oxygen. Establish the empirical formula of the compound.

$$38.196 \text{ g CO}_2 \times \frac{1 \text{ mol C}}{44.011 \text{ g CO}_2} = 0.86787 \text{ mol C} \div 0.86787 = 1 \text{ mol C} \times 5 = 5 \text{ mol C}$$

$$18.752 \text{ g H}_2\text{O} \times \frac{2 \text{ mol H}}{18.016 \text{ g H}_2\text{O}} = 1.0817 \text{ mol H} \div 0.86787 = 2.3996 \text{ mol H} \times 5 = 11.998 \approx 12 \text{ mol H}$$



- 3.98 A sample of a compound of Cl and O reacts with an excess of H<sub>2</sub> to give 0.233 g of HCl and 0.403 g of H<sub>2</sub>O. Determine the empirical formula of the compound.

We assume that all the Cl in the compound ends up as HCl and all the O ends up as H<sub>2</sub>O.

$$0.233 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} \times \frac{1 \text{ mol Cl}}{1 \text{ mol HCl}} = 0.00639 \text{ mol Cl} \div 0.00639 = 1 \times 2 = 2 \text{ Cl}$$

$$0.403 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{1 \text{ mol O}}{1 \text{ mol H}_2\text{O}} = 0.0224 \text{ mol O} \div 0.00639 = 3.5 \times 2 = 7 \text{ O}$$



- 3.137 When 12.10 g of *tert*-butyl ether (a compound of C, H, and O) are burned in an apparatus like the one shown in Fig 3.6, 30.1 g of CO<sub>2</sub> and 14.8 g of H<sub>2</sub>O are formed. What is the empirical formula of the compound?

$$? \text{ g C} = 30.2 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 8.24 \text{ g C}$$

$$? \text{ g H} = 14.8 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 1.66 \text{ g H}$$

$$\text{mass O} = 12.10 \text{ g compound} - (8.24 \text{ g C} + 1.66 \text{ g H}) = 2.20 \text{ g O}$$

$$? \text{ mol C} = 8.24 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.686 \text{ mol C} \div 0.138 = 5$$

$$? \text{ mol H} = 1.66 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 1.65 \text{ mol H} \div 0.138 = 12$$

$$? \text{ mol O} = 2.20 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.138 \text{ mol O} \div 0.138 = 1$$

The empirical formula is  
**C<sub>5</sub>H<sub>12</sub>O.**

- B. After combustion with excess oxygen, a 28.39 g sample of propanethiol produced 49.22 g of carbon dioxide, 26.87 of water and 29.84 g SO<sub>3</sub>. The compound produced no other oxides. Determine the empirical formula of the compound.

$$49.22 \text{ g CO}_2 \times \frac{1 \text{ mol C}}{44.01 \text{ g CO}_2} = 1.118 \text{ mol C} \div 0.3728 = 3 \text{ mol C}$$

$$26.87 \text{ g H}_2\text{O} \times \frac{2 \text{ mol H}}{18.02 \text{ g H}_2\text{O}} = 2.982 \text{ mol H} \div 0.3728 = 8 \text{ mol H}$$

$$29.848 \text{ g SO}_3 \times \frac{1 \text{ mol S}}{80.06 \text{ g SO}_3} = 0.3728 \text{ mol S} \div 0.3728 \text{ mol S} = 1$$



- 3.52 The empirical formula of a compound is CH. If the molar mass of this compound is 78 g, what is its molecular formula?

$$\frac{\text{molar mass}}{\text{empirical molar mass}} = \frac{78 \text{ g}}{13.02 \text{ g}} \approx 6, \text{ thus the molecular formula is } (\text{CH})_6, \text{ or } \mathbf{C_6H_6}.$$

- 3.54 Monosodium glutamate (MSG) has the following composition by mass: 35.51% C, 4.77% H, 37.85% O, 8.29% N, and 13.60% Na. What is its molecular formula if its molar mass is 169 g?

$$n_{\text{C}} = 35.51 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 2.957 \text{ mol C} \div 0.5916 = 4.998 \approx 5$$

$$n_{\text{H}} = 4.77 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 4.73 \text{ mol H} \div 0.5916 = 8$$

$$n_{\text{O}} = 37.85 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.366 \text{ mol O} \div 0.5916 = 3.999 \approx 4$$

$$n_{\text{N}} = 8.29 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.592 \text{ mol N} \div 0.5916 = 1$$

$$n_{\text{Na}} = 13.60 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} = 0.5916 \text{ mol Na} \div 0.5916 = 1$$

Thus, the empirical formula is  $\mathbf{C_5H_8O_4NNa}$ .

Divide to find the factor for the molecular formula:  $\frac{\text{molar mass}}{\text{empirical molar mass}} = \frac{169 \text{ g}}{169.11 \text{ g}} \approx 1$

Hence, the molecular formula and the empirical formula are the same,  $\mathbf{C_5H_8O_4NNa}$ .

- 3.119 Lysine contains C, H, O, and N. In one experiment, the complete combustion of 2.175 g of lysine gave 3.94 g  $\text{CO}_2$  and 1.89 g  $\text{H}_2\text{O}$ . When reduced in hydrogen, a second 2.175 g sample of lysine produced 0.506 g  $\text{NH}_3$ . First determine the empirical formula for lysine, then given the molecular mass of lysine is 150 g/mol, determine its molecular formula.

$$? \text{ g C} = 3.94 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 1.075 \text{ g C}$$

$$? \text{ g H} = 1.89 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.2114 \text{ g H}$$

$$? \text{ g N} = 0.506 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3} \times \frac{1 \text{ mol N}}{1 \text{ mol NH}_3} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = 0.4165 \text{ g N}$$

$$? \text{ g O} = 2.175 \text{ g} - (1.075 \text{ g} + 0.2114 \text{ g} + 0.4165 \text{ g}) = 0.4721 \text{ g O}$$

$$? \text{ mol C} = 1.075 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.08951 \text{ mol C} \div 0.02951 = 3$$

$$? \text{ mol H} = 0.2114 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 0.2097 \text{ mol H} \div 0.02951 = 7$$

$$? \text{ mol N} = 0.4165 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.02973 \text{ mol N} \div 0.02951 = 1$$

$$? \text{ mol O} = 0.4721 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.02951 \text{ mol O} \div 0.02951 = 1$$

Thus, the empirical formula is  $\mathbf{C_3H_7NO}$ .

$$\frac{150 \text{ g}}{73.10 \text{ g}} \approx 2 \Rightarrow \text{molecular formula is } (\text{C}_3\text{H}_7\text{NO})_2 \text{ or } \mathbf{C_6H_{14}N_2O_2}.$$