Section 6.9: Molecular Formulas

Mini Investigation: Comparing Molecules and Molecular Formulas, page 297

A. Given: molecular formula for methanal, \( \text{CH}_2\text{O} \)

Required: molecular mass of methanal, \( M_{\text{CH}_2\text{O}} \)

Solution:

Step 1. Look up the molecular mass of each element in the substance.
\[
M_C = 12.01 \text{ u}; \quad M_H = 1.01 \text{ u}; \quad M_O = 16.00 \text{ u}
\]

Step 2. Add the molecular masses of the elements in the substance.
\[
M_{\text{CH}_2\text{O}} = M_C + 2M_H + M_O
\]
\[
= 12.01 \text{ u} + 2(1.01 \text{ u}) + 16.00 \text{ u}
\]
\[
M_{\text{CH}_2\text{O}} = 30.03 \text{ u}
\]

Statement: The molecular mass of methanal is 30.03 u.

B. The molar mass and molecular mass of a compound have the same numerical value but have different units. The units of molar masses are g/mol while the molecular masses are expressed in atomic mass units (u).

C. Given: molecular formula for methanal, \( \text{C}_2\text{H}_4\text{O}_2 \)

Required: molecular mass of methanal, \( M_{\text{C}_2\text{H}_4\text{O}_2} \)

Solution:

Step 1. Look up the molecular mass of each element in the substance.
\[
M_C = 12.01 \text{ u}; \quad M_H = 1.01 \text{ u}; \quad M_O = 16.00 \text{ u}
\]

Step 2. Add the molecular masses of the elements in the substance.
\[
M_{\text{C}_2\text{H}_4\text{O}_2} = 2M_C + 4M_H + 2M_O
\]
\[
= 2(12.01 \text{ u}) + 4(1.01 \text{ u}) + 2(16.00 \text{ u})
\]
\[
M_{\text{C}_2\text{H}_4\text{O}_2} = 60.06 \text{ u}
\]

Statement: The molecular mass of each compound with molecular formula \( \text{C}_2\text{H}_4\text{O}_2 \) is 60.06 u.

Given: molecular formula for methanal, \( \text{C}_3\text{H}_6\text{O}_3 \)

Required: molecular mass of methanal, \( M_{\text{C}_3\text{H}_6\text{O}_3} \)

Solution:

Step 1. Look up the molecular mass of each element in the substance.
\[
M_C = 12.01 \text{ u}; \quad M_H = 1.01 \text{ u}; \quad M_O = 16.00 \text{ u}
\]

Step 2. Add the molecular masses of the elements in the substance.
\[
M_{\text{C}_3\text{H}_6\text{O}_3} = 3M_C + 6M_H + 3M_O
\]
\[
= 3(12.01 \text{ u}) + 6(1.01 \text{ u}) + 3(16.00 \text{ u})
\]
\[
M_{\text{C}_3\text{H}_6\text{O}_3} = 90.09 \text{ u}
\]

Statement: The molecular mass of each compound with molecular formula \( \text{C}_3\text{H}_6\text{O}_3 \) is 90.09 u.

The molar mass of a compound is determined by its molecular formula and not by its structure.
D. Since the molecular formulas of these compounds are multiples of the molecular formula of methanal, their molecular masses can be predicted by multiplying the molecular mass of methanal by the multiple.

E. There may be several compounds that have the same molecular formula. One way to identify the compound is by determining a unique physical property of the substance such as its melting point or fragment pattern in a mass spectrum.

**Tutorial 1 Practice, page 298**

1. (a) Given: empirical formula, HO; molar mass of compound = 34.02 g/mol

**Required:** molecular formula of compound, $H_2O_y$

**Solution:**

**Step 1.** Calculate the empirical molar mass of the compound by adding the molar masses of the elements.

\[
M_{\text{HO}} = \left(1.01 \dfrac{\text{g}}{\text{mol}}\right) + \left(16.00 \dfrac{\text{g}}{\text{mol}}\right) = 17.01 \dfrac{\text{g}}{\text{mol}}
\]

**Step 2.** Solve for \(x\), the mass multiple.

\[
x = \dfrac{34.02 \dfrac{\text{g}}{\text{mol}}}{17.01 \dfrac{\text{g}}{\text{mol}}} = 2.000
\]

**Step 3.** The molar mass of the compound is 2 times the molar mass of the empirical formula. Multiply each of the subscripts by 2 to give $H_2O_2$.

**Statement:** The molecular formula of the compound is $H_2O_2$.

(b) Given: empirical formula, $SO_2$; molar mass of compound = 64.06 g/mol

**Required:** molecular formula of compound, $S_xO_y$

**Solution:**

**Step 1.** Calculate the empirical molar mass of the compound by adding the molar masses of the elements.

\[
M_{\text{SO}_2} = \left(32.07 \dfrac{\text{g}}{\text{mol}}\right) + \left(2 \times 16.00 \dfrac{\text{g}}{\text{mol}}\right) = 64.07 \dfrac{\text{g}}{\text{mol}}
\]

**Step 2.** Solve for \(x\), the mass multiple.

\[
x = \dfrac{64.06 \dfrac{\text{g}}{\text{mol}}}{64.07 \dfrac{\text{g}}{\text{mol}}} = 0.9998
\]

**Statement:** The molecular formula of the compound is $SO_2$. 
(c) Given: empirical formula, KSO₄; molar mass of compound = 270.32 g/mol
Required: molecular formula of compound, KₓSₓOₓ
Solution:
Step 1. Calculate the empirical molar mass of the compound by adding the molar masses of the elements.

\[
M_{KSO_4} = \left(39.10 \text{ g mol}^{-1}\right) + \left(32.07 \text{ g mol}^{-1}\right) + \left(4 \times 16.00 \text{ g mol}^{-1}\right)
\]

\[
M_{KSO_4} = 135.17 \text{ g mol}^{-1}
\]

Step 2. Solve for \(x\), the mass multiple.

\[
x = \frac{270.32 \text{ g mol}^{-1}}{135.17 \text{ g mol}^{-1}}
\]

\[
x = 2.000
\]

Step 3. The molar mass of the compound is 2 times the molar mass of the empirical formula. Multiply each of the subscripts by 2 to give K₂S₂O₈.

Statement: The molecular formula of the compound is K₂S₂O₈.

(d) Given: empirical formula, C₃H₅O₃; molar mass of compound = 445.40 g/mol
Required: molecular formula of compound, CₓHᵧOₙ
Solution:
Step 1. Calculate the empirical molar mass of the compound by adding the molar masses of the elements.

\[
M_{C_3H_5O_3} = \left(3 \times 12.01 \text{ g mol}^{-1}\right) + \left(5 \times 1.01 \text{ g mol}^{-1}\right) + \left(3 \times 16.00 \text{ g mol}^{-1}\right)
\]

\[
M_{C_3H_5O_3} = 89.08 \text{ g mol}^{-1}
\]

Step 2. Solve for \(x\), the mass multiple.

\[
x = \frac{445.40 \text{ g mol}^{-1}}{89.08 \text{ g mol}^{-1}}
\]

\[
x = 5.000
\]

Step 3. The molar mass of the compound is 5 times the molar mass of the empirical formula. Multiply each of the subscripts by 5 to give C₁₅H₂₅O₁₅.

Statement: The molecular formula of the compound is C₁₅H₂₅O₁₅.
2. (a) Given: molar mass of hydrocarbon = 42.09 g/mol; % C = 85.6 %; remainder is H
Required: molecular formula of hydrocarbon
Solution:
A 100.0 g sample of this compound contains 85.6 g of carbon and 14.4 g of hydrogen.
Step 1. Calculate the amount of each element in the 100.0 g sample.

\[ n_C = \left( \frac{85.6 \text{ g}}{12.01 \text{ g/mol}} \right) \]
\[ n_C = 7.1274 \text{ mol} \quad \text{[2 extra digits carried]} \]

\[ n_H = \left( \frac{14.4 \text{ g}}{1.01 \text{ g/mol}} \right) \]
\[ n_H = 14.257 \text{ mol} \quad \text{[2 extra digits carried]} \]

Step 2. Divide the amount of each element by the smallest amount.

\[ \frac{n_C}{n_C} = \frac{7.1274 \text{ mol}}{7.1274 \text{ mol}} = 1.00 \]

\[ \frac{n_H}{n_C} = \frac{14.257 \text{ mol}}{7.1274 \text{ mol}} = 2.00 \]

The ratio of carbon to hydrogen is 1:2. This ratio suggests an empirical formula of \( \text{CH}_2 \).

Step 3. Calculate the empirical molar mass of \( \text{CH}_2 \).

\[ M_{\text{CH}_2} = \left( 12.01 \text{ \frac{g}{mol}} \right) + \left( 2 \times 1.01 \text{ \frac{g}{mol}} \right) \]
\[ M_{\text{CH}_2} = 14.03 \text{ \frac{g}{mol}} \]

Step 4. Solve for \( x \), the mass multiple.

\[ x = \frac{42.09 \text{ \frac{g}{mol}}}{14.03 \text{ \frac{g}{mol}}} \]
\[ x = 3.000 \]

Step 5. The molar mass of the compound is 3 times the molar mass of the empirical formula. Multiply each of the subscripts by 3 to give \( \text{C}_3\text{H}_6 \).

Statement: The molecular formula of the hydrocarbon is \( \text{C}_3\text{H}_6 \).
(b) **Given:** molar mass of compound = 283.88 g/mol; % P = 43.6%; remainder is O  
**Required:** molecular formula of compound  
**Solution:**  
A 100.0 g sample of this compound contains 43.6 g of phosphorus and 56.4 g of oxygen.  
**Step 1.** Calculate the amount of each element in the 100.0 g sample.  
\[ n_P = \left( \frac{43.6\ g}{30.97\ g/mol} \right) = 1.4078\ mol \quad [2\ extra\ digits\ carried] \]
\[ n_O = \left( \frac{56.4\ g}{16.00\ g/mol} \right) = 3.5250\ mol \quad [2\ extra\ digits\ carried] \]

**Step 2.** Divide the amount of each element by the smallest amount.  
\[ n_P = \frac{1.4078\ mol}{1.4078\ mol} = 1.00 \]
\[ n_O = \frac{3.5250\ mol}{1.4078\ mol} = 2.50 \]

The ratio of phosphorus to oxygen is 1:2.5. This ratio suggests an empirical formula of \( \text{P}_1\text{O}_{2.5} \). Multiplying the subscripts of this formula by 2 gives \( \text{P}_2\text{O}_5 \).  
**Step 3.** Calculate the empirical molar mass of \( \text{P}_2\text{O}_5 \).  
\[ M_{\text{P}_2\text{O}_5} = \left( 2 \times 30.97\ \frac{g}{mol} \right) + \left( 5 \times 16.00\ \frac{g}{mol} \right) \]
\[ M_{\text{P}_2\text{O}_5} = 141.94\ \frac{g}{mol} \]

**Step 4.** Solve for \( x \), the mass multiple.  
\[ x = \frac{283.88\ \frac{g}{mol}}{141.94\ \frac{g}{mol}} \]
\[ x = 2.000 \]

**Step 5.** The molar mass of the compound is 2 times the molar mass of the empirical formula. Multiply each of the subscripts by 2 to give \( \text{P}_4\text{O}_{10} \).  
**Statement:** The molecular formula of the compound is \( \text{P}_4\text{O}_{10} \).
(c) **Given:** molar mass of compound = 168.21 g/mol; % C = 64.3 %; % H = 7.2 %; remainder is O  
**Required:** molecular formula of compound  
**Solution:**  
% O = 100 % – 64.3 % – 7.2 % = 28.5 %  
A 100.0 g sample of this compound contains 64.3 g of carbon, 7.2 g of hydrogen, and 28.5 g of oxygen.  
**Step 1.** Calculate the amount of each element in the 100.0 g sample.  
\[ n_C = \left( \frac{64.3 \text{ g}}{12.01 \text{ g/mol}} \right) = 5.3539 \text{ mol} \]  
\[ n_H = \left( \frac{7.2 \text{ g}}{1.01 \text{ g/mol}} \right) = 7.129 \text{ mol} \]  
\[ n_O = \left( \frac{28.5 \text{ g}}{16.00 \text{ g/mol}} \right) = 1.7813 \text{ mol} \]  
**Step 2.** Divide the amount of each element by the smallest amount.  
\[ \frac{n_O}{n_O} = \frac{1.7813 \text{ mol}}{1.7813 \text{ mol}} = 1.00 \]  
\[ \frac{n_C}{n_O} = \frac{5.3539 \text{ mol}}{1.7813 \text{ mol}} = 3.01 \]  
\[ \frac{n_H}{n_O} = \frac{7.129 \text{ mol}}{1.7813 \text{ mol}} = 4.0 \]  
The ratio of carbon to hydrogen to oxygen is 3:4:1. This ratio suggests an empirical formula of C₃H₄O.  
**Step 3.** Calculate the empirical molar mass of C₃H₄O.  
\[ M_{C,H,O} = \left( 3 \times 12.01 \frac{\text{g}}{\text{mol}} \right) + \left( 4 \times 1.01 \frac{\text{g}}{\text{mol}} \right) + \left( 16.00 \frac{\text{g}}{\text{mol}} \right) \]  
\[ M_{C,H,O} = 56.07 \frac{\text{g}}{\text{mol}} \]
Step 4. Solve for \( x \), the mass multiple.

\[
\frac{168.21 \text{ g}}{\text{mol}} \times \frac{1}{56.07 \text{ g/mol}} = x
\]

\( x = 3.000 \)

Step 5. The molar mass of the compound is 3 times the molar mass of the empirical formula. Multiply each of the subscripts by 3 to give \( C_9H_{12}O_3 \).

Statement: The molecular formula of the compound is \( C_9H_{12}O_3 \).

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1. The statement is valid because the molar masses of ethyl ethanoate and 2,4,6-trimethyl-1,3,5 trioxane are whole number multiples of the molar mass of ethanal. For example, the molar mass of ethyl ethanoate is twice the molar mass of ethanal. Similarly, multiplying the subscripts of ethanal’s molecular formula by 2 gives the molecular formula of ethyl ethanoate.

2. (a) Given: empirical formula, \( \text{NO}_2 \); molar mass of compound = 92.02 g/mol

Required: molecular formula of compound

Solution:

Step 1. Calculate the empirical molar mass of the compound.

\[
M_{\text{NO}_2} = \left( 14.01 \text{ g/mol} \right) + \left( 2 \times 16.00 \text{ g/mol} \right)
\]

\( M_{\text{NO}_2} = 46.01 \text{ g/mol} \)

Step 2. Solve for \( x \), the mass multiple.

\[
\frac{92.02 \text{ g}}{\text{mol}} \times \frac{1}{46.01 \text{ g/mol}} = x
\]

\( x = 2.000 \)

Step 3. The molar mass of the compound is 2 times the molar mass of the empirical formula. Multiply each of the subscripts by 2 to give \( \text{N}_2\text{O}_4 \).

Statement: The molecular formula of the compound is \( \text{N}_2\text{O}_4 \).

(b) Given: empirical formula, \( \text{CH}_2 \); molar mass of compound = 84.18 g/mol

Required: molecular formula of compound

Solution:

Step 1. Calculate the empirical molar mass of the compound.

\[
M_{\text{CH}_2} = \left( 12.01 \text{ g/mol} \right) + \left( 2 \times 1.01 \text{ g/mol} \right)
\]

\( M_{\text{CH}_2} = 14.03 \text{ g/mol} \)
Step 2. Solve for \( x \), the mass multiple.

\[
\begin{align*}
x &= \frac{84.18 \text{ g}}{14.03 \text{ mol}} \\
x &= 6.00
\end{align*}
\]

Step 3. The molar mass of the compound is 6 times the molar mass of the empirical formula. Multiply each of the subscripts by 6 to give \( C_6H_{12} \).

Statement: The molecular formula of the compound is \( C_6H_{12} \).

(c) Given: empirical formula, \( C_2H_3O_3 \); molar mass of compound = 225.15 g/mol

Required: molecular formula of compound

Solution:

Step 1. Calculate the empirical molar mass of the compound.

\[
M_{C_2H_3O_3} = \left( 2 \times 12.01 \text{ g mol} \right) + \left( 3 \times 1.01 \text{ g mol} \right) + \left( 3 \times 16.00 \text{ g mol} \right)
\]

\[
M_{C_2H_3O_3} = 75.05 \text{ g mol}
\]

Step 2. Solve for \( x \), the mass multiple.

\[
\begin{align*}
x &= \frac{225.15 \text{ g}}{75.05 \text{ g mol}} \\
x &= 3.000
\end{align*}
\]

Step 3. The molar mass of the compound is 3 times the molar mass of the empirical formula. Multiply each of the subscripts by 3 to give \( C_6H_9O_9 \).

Statement: The molecular formula of the compound is \( C_6H_9O_9 \).

(d) Given: empirical formula, \( CFBrO \); molar mass of compound = 253.82 g/mol

Required: molecular formula of compound

Solution:

Step 1. Calculate the empirical molar mass of the compound.

\[
M_{\text{CFBrO}} = \left( 12.01 \text{ g mol} \right) + \left( 19.00 \text{ g mol} \right) + \left( 79.90 \text{ g mol} \right) + \left( 16.00 \text{ g mol} \right)
\]

\[
M_{\text{CFBrO}} = 126.91 \text{ g mol}
\]

Step 2. Solve for \( x \), the mass multiple.

\[
\begin{align*}
x &= \frac{253.82 \text{ g}}{126.91 \text{ g mol}} \\
x &= 2.000
\end{align*}
\]
Step 3. The molar mass of the compound is 2 times the molar mass of the empirical formula. Multiply each of the subscripts by 2 to give C$_2$F$_2$Br$_2$O$_2$.

Statement: The molecular formula of the compound is C$_2$F$_2$Br$_2$O$_2$.

3. For some compounds, the molecular formula and the empirical formula are the same because the elements in the formula are already in the simplest whole-number ratio.

4. The percentage composition gives the percentage, by mass, of the elements in the compound. Percentage composition is the same for the empirical formula of a compound or any multiple of the empirical formula. For example, CH and C$_2$H$_2$ have the same percentage composition because they contain the same proportions of carbon to hydrogen. They would have different molecule masses since they contain different numbers of atoms of each element.

5. In almost all cases, the chemical formula of an ionic compound is an empirical formula.

6. Given: molar mass of caffeine = 194.19 g/mol; % C = 49.48%; % H = 5.15%; % N = 28.87%;% O = 16.49%

Required: empirical formula; molecular formula of caffeine

Solution:
A 100.0 g sample of this compound contains 49.48 g of carbon, 5.15 g of hydrogen, 28.87 g of nitrogen, and 16.49 g of oxygen.

Step 1. Calculate the amount of each element in the 100.0 g sample.

\[ n_C = \left( \frac{49.48 \text{ g}}{12.01 \text{ g/mol}} \right) \]
\[ n_C = 4.119 \text{ mol} \quad [2 \text{ extra digits carried}] \]

\[ n_H = \left( \frac{5.15 \text{ g}}{1.01 \text{ g/mol}} \right) \]
\[ n_H = 5.0990 \text{ mol} \quad [2 \text{ extra digits carried}] \]

\[ n_N = \left( \frac{28.87 \text{ g}}{14.01 \text{ g/mol}} \right) \]
\[ n_N = 2.060 \text{ mol} \quad [2 \text{ extra digits carried}] \]

\[ n_O = \left( \frac{16.49 \text{ g}}{16.00 \text{ g/mol}} \right) \]
\[ n_O = 1.030 \text{ mol} \quad (2 \text{ extra digits carried}) \]

Step 2. Divide the amount of each element by the smallest amount.

\[ n_O = \frac{1.030 \text{ mol}}{1.000} = 1.000 \]

\[ n_C = \frac{4.119 \text{ mol}}{3.997} = 3.997 \]
The ratio of carbon to hydrogen to nitrogen to oxygen is 4:5:2:1. This ratio suggests an empirical formula of \( \text{C}_4\text{H}_5\text{N}_2\text{O} \).

**Step 3.** Calculate the empirical molar mass of \( \text{C}_4\text{H}_5\text{N}_2\text{O} \).

\[
M_{\text{C}_4\text{H}_5\text{N}_2\text{O}} = \left(4 \times 12.01 \frac{\text{g}}{\text{mol}}\right) + \left(5 \times 1.01 \frac{\text{g}}{\text{mol}}\right) + \left(2 \times 14.01 \frac{\text{g}}{\text{mol}}\right) + \left(16.00 \frac{\text{g}}{\text{mol}}\right)
\]

\[
M_{\text{C}_4\text{H}_5\text{N}_2\text{O}} = 97.11 \frac{\text{g}}{\text{mol}}
\]

**Step 4.** Solve for \( x \), the mass multiple.

\[
x = \frac{194.19 \frac{\text{g}}{\text{mol}}}{97.11 \frac{\text{g}}{\text{mol}}}
\]

\[
x = 2.000
\]

**Step 5.** The molar mass of the compound is 2 times the molar mass of the empirical formula. Multiply each of the subscripts by 3 to give \( \text{C}_8\text{H}_{10}\text{N}_4\text{O}_2 \).

**Statement:** The empirical formula of caffeine is \( \text{C}_4\text{H}_5\text{N}_2\text{O} \) and its molecular formula is \( \text{C}_8\text{H}_{10}\text{N}_4\text{O}_2 \).

**7.** **Given:** molar mass of compound = 300.48 g/mol; % C = 80.0 %; % H = 9.41 %; % O = 10.6 %

**Required:** molecular formula of compound

**Solution:**
A 100.0 g sample of this compound contains 80.0 g of carbon, 9.41 g of hydrogen, and 10.6 g of oxygen.

**Step 1.** Calculate the amount of each element in the 100.0 g sample.

\[
n_\text{C} = (80.0 \frac{\text{g}}{\text{mol}}) \left(\frac{1 \text{ mol}}{12.01 \frac{\text{g}}{\text{mol}}}\right)
\]

\[
n_\text{C} = 6.6611 \text{ mol} \quad [2 \text{ extra digits carried}]
\]

\[
n_\text{H} = (9.41 \frac{\text{g}}{\text{mol}}) \left(\frac{1 \text{ mol}}{1.01 \frac{\text{g}}{\text{mol}}}\right)
\]

\[
n_\text{H} = 9.3168 \text{ mol} \quad [2 \text{ extra digits carried}]
\]
\[ n_0 = (10.6 \, \text{g}) \left( \frac{1 \, \text{mol}}{16.00 \, \text{g}} \right) \]

\[ n_0 = 0.662 \, 50 \, \text{mol} \quad [2 \, \text{extra digits carried}] \]

**Step 2.** Divide the amount of each element by the smallest amount.

\[ \frac{n_0}{n_0} = \frac{0.66250 \, \text{mol}}{0.66250 \, \text{mol}} = 1.00 \]

\[ \frac{n_0}{n_0} = \frac{6.6611 \, \text{mol}}{0.66250 \, \text{mol}} = 10.1 \]

\[ \frac{n_0}{n_0} = \frac{9.3168 \, \text{mol}}{0.66250 \, \text{mol}} = 14.1 \]

The ratio of carbon to hydrogen to oxygen is 10:14:1. This ratio suggests an empirical formula of \( \text{C}_{10}\text{H}_{14}\text{O} \).

**Step 3.** Calculate the empirical molar mass of \( \text{C}_{10}\text{H}_{14}\text{O} \).

\[ M_{\text{C}_{10}\text{H}_{14}\text{O}} = \left( 10 \times 12.01 \, \frac{\text{g}}{\text{mol}} \right) + \left( 14 \times 1.01 \, \frac{\text{g}}{\text{mol}} \right) + \left( 16.00 \, \frac{\text{g}}{\text{mol}} \right) \]

\[ M_{\text{C}_{10}\text{H}_{14}\text{O}} = 150.24 \, \frac{\text{g}}{\text{mol}} \]

**Step 4.** Solve for \( x \), the mass multiple.

\[ x = \frac{300.48 \, \frac{\text{g}}{\text{mol}}}{150.24 \, \frac{\text{g}}{\text{mol}}} \]

\[ x = 2.0000 \]

**Step 5.** The molar mass of the compound is 2 times the molar mass of the empirical formula. Multiply each of the subscripts by 2 to give \( \text{C}_{20}\text{H}_{28}\text{O}_2 \).

**Statement:** The molecular formula of this compound is \( \text{C}_{20}\text{H}_{28}\text{O}_2 \). (b) This analysis only gives the molecular formula of the compound. However, the molecular formula is not unique to one substance. Therefore, this analysis is insufficient evidence to accuse the athlete of using dianabol.

**8. Given:** ratio of H to O = 2 : 1; \% C = 40.0 \%; molar mass of compound = 180 g/mol  
**Required:** empirical formula; molecular formula of carbohydrate  
**Solution:** Since the ratio of hydrogen to oxygen is 2 : 1, the hydrogen and oxygen group has the simplest formula \( \text{H}_2\text{O} \). A 100.0 g sample of this compound contains 40.0 g of carbon and 60.0 g of \( \text{H}_2\text{O} \).
**Step 1.** Calculate the amount of each element/group in the 100.0 g sample.

\[
n_C = (40.0 \text{ g}) \left( \frac{1 \text{ mol}}{12.01 \text{ g}} \right)
\]

\[
n_C = 3.3306 \text{ mol} \ [2 \text{ extra digits carried}]
\]

\[
n_{H_2O} = (60.0 \text{ g}) \left( \frac{1 \text{ mol}}{18.02 \text{ g}} \right)
\]

\[
n_{H_2O} = 3.3296 \text{ mol} \ [2 \text{ extra digits carried}]
\]

**Step 2.** Divide each amount by the smallest amount.

\[
\frac{n_{H_2O}}{n_{H_2O}} = \frac{3.3296 \text{ mol}}{3.3296 \text{ mol}} = 1.00
\]

\[
\frac{n_C}{n_{H_2O}} = \frac{3.3306 \text{ mol}}{3.3296 \text{ mol}} = 1.00
\]

The ratio of carbon to H\(_2\)O group is 1:1. This ratio suggests an empirical formula of CH\(_2\)O.

**Step 3.** Calculate the empirical molar mass of CH\(_2\)O.

\[
M_{\text{CH}_2\text{O}} = \left( 12.01 \text{ g mol}^{-1} \right) + \left( 2 \times 1.01 \text{ g mol}^{-1} \right) + \left( 16.00 \text{ g mol}^{-1} \right)
\]

\[
M_{\text{CH}_2\text{O}} = 30.03 \text{ g mol}^{-1}
\]

**Step 4.** Solve for \(x\), the mass multiple.

\[
x = \frac{180 \text{ g mol}^{-1}}{30.03 \text{ g mol}^{-1}}
\]

\[
x = 5.99
\]

**Step 5.** The molar mass of the compound is 6 times the molar mass of the empirical formula. Multiply each of the subscripts by 6 to give C\(_6\)H\(_{12}\)O\(_6\).

**Statement:** The empirical formula of the carbohydrate is CH\(_2\)O and its molecular formula is C\(_6\)H\(_{12}\)O\(_6\).

9. Answers may vary. Sample answer:

(a) Anabolic steroids are synthetic versions of human hormones like testosterone. Like naturally-occurring hormones, anabolic steroids increase the rate of protein synthesis in cells, particularly in muscle cells which promotes the growth of muscle tissue. Anabolic steroids also promote the development of masculine characteristics like a deeper voice and body hair.

(b) Some of the risks of using anabolic steroids include: fluid retention and high blood pressure; severe acne; shrinking of the testes and a reduction in sperm count in men; changes in the menstrual cycle; growth of facial hair in women; changes in mood and behavior; and depression.