1. The empirical formula of a compound is also called the simplest formula. It represents the relative or smallest whole-number ratio of atoms in a cmpd.

2. The molecular formula represents the actual number of atoms of each element in a molecule of the compound.

3. The empirical formula and the molecular formula are mathematically related as follows:
   
   Molecular formula = n x empirical formula.

4. Can the molecular formula be the same as the empirical formula? Explain.  
   Yes. The molecular and empirical formula may be the same when the molecular formula does not have subscripts that can be reduced (ex: formaldehyde, CH₂O). They are different in any case where the molecular formula may be reduced to a smaller whole-number ratio of elements (ex: acetic acid, C₂H₄O₂ has the same empirical formula as formaldehyde, CH₂O).

5. The molecular formula for glucose is C₆H₁₂O₆.
   a. What is its empirical formula? CH₂O
   b. The molecular formula of glucose, C₆H₁₂O₆ = 6 x CH₂O
   c. The molecular weight of glucose is 180 g/mol. It is equal to 6 x 30 g/mol.

6. The molecular formula of benzene is C₆H₆.
   a. What is its empirical formula? CH
   b. The molecular formula of benzene, C₆H₆ = 6 x CH

7. The empirical formula of a compound is NH, and its molecular weight is 30.0 amu.
   What is its molecular formula? N₂H₂
   
   NH ≈ 15 g/mol.
   n = 30.0 g/mol ÷ 15.0 g/mol = 2  ∴  2 x NH = N₂H₂

8. The empirical formula of a compound is CH₃. Its molecular weight is 30.0 amu.
   What is its molecular formula? C₂H₆
   
   CH₃ ≈ 15 g/mol.
   n = 30.0 g/mol ÷ 15.0 g/mol = 2  ∴  2 x CH₃ = C₂H₆
9. A compound is 81.7% carbon and 18.3% hydrogen.
   a. What is its empirical formula? $\text{C}_3\text{H}_8$
      Assume 100g of the compound to simplify the problem.
      81.7g C x (1 mol C / 12.01g C) = 6.803 mol C … divide both by 6.803 to reduce the ratio = 1 mol C
      18.3g H x (1mol H / 1.01g H) = 18.12 mol H … to 2.663 mol H
      Multiplying by 3 will give the smallest whole number ratio: 3 (C$_{1.6}$H$_{2.6}$) = C$_3$H$_8$
   b. The formula weight of this compound is 44.0 amu. Is the molecular formula different than the empirical formula? **No.** The mass of the empirical formula above is 44.0 g/mol, so the empirical and molecular formulas are the same.

10. Butyric acid is 54.5% carbon, 9.09% hydrogen and 36.4% oxygen.
   a. What is its empirical formula? $\text{C}_2\text{H}_4\text{O}$
      54.5g C x (1 mol C / 12.01g C) = 4.538 mol C … divide all by 2.275 to reduce the ratio = 1.995 mol C
      9.09g H x (1mol H / 1.01g H) = 9.000 mol H … to 3.956 mol H
      36.4g O x (1mol H / 16.00g H) = 2.275 mol O … to 1 mol O
      This gives the smallest whole number ratio:  C$_2$H$_4$O
   b. Its molar mass is 88.0 g/mol. What is the molecular formula of butyric acid? $\text{C}_4\text{H}_8\text{O}_2$
      Empirical molar mass = 44.0 g/mol
      $n = \frac{88.0 \text{ g/mol}}{44.0 \text{ g/mol}} = 2.000$.
      $2 \times \text{C}_2\text{H}_4\text{O} = \text{C}_4\text{H}_8\text{O}_2$

11. Isopropyl alcohol contains C, H, and O. When we burn 11.63 g of this compound, the products are 25.5 g CO$_2$ and 14.0 g H$_2$O.
   a. What is the empirical formula? $\text{C}_3\text{H}_8\text{O}$
      You can assume all the moles of C and all the moles of H came from the isopropanol.
      $0.5795 \text{ mol C} \times (12.01\text{g C} / 1 \text{ mol C}) = 6.96\text{g carbon}$
      $0.7778 \text{ mol H}_2\text{O} \times (2 \text{ mol H} / 1 \text{ mol H}_2\text{O}) = 1.5556 \text{mol H} \times (1.01 \text{ g H} / 1 \text{ mol H}) = 1.57 \text{ g hydrogen}$
      $11.63\text{g total} - (6.96\text{g C} + 1.57\text{g H}) = 3.10\text{g oxygen} \times (1 \text{ mol O} / 16.00\text{g O}) = 0.1937 \text{ mol O}$
      Dividing all moles by 0.1937 to reduce the ratio gives: C$_3$H$_8$O
   b. The molar mass of the alcohol is 60.0 g/mol. What is its molecular formula? $\text{C}_3\text{H}_8\text{O}$
      The molar mass of the empirical formula C$_3$H$_8$O is 60.0 g/mol, so the molecular formula is the same.
12. The complete combustion of a 0.5728 g sample of a compound that contains only C, H, and O produced 0.840 g of carbon dioxide and 0.254 g of water. The molar mass of the compound was determined to be about 60.0 g/mol.

What is the molecular formula of this compound? **C_2H_3O_2**

\[
\text{C}_x\text{H}_y\text{O}_z + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}
\]

Convert to moles:

<table>
<thead>
<tr>
<th>Substance</th>
<th>Mass</th>
<th>Moles</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.5728 g total</td>
<td>-</td>
<td>0.01909 mol</td>
</tr>
<tr>
<td>0.840 g CO₂</td>
<td>-</td>
<td>0.01410 mol</td>
</tr>
<tr>
<td>0.254 g H₂O</td>
<td>-</td>
<td></td>
</tr>
</tbody>
</table>

\[0.01909 \text{ mol C} \times (12.01 \text{ g C} / 1 \text{ mol C}) = 0.2293 \text{ g carbon}\]

\[0.01410 \text{ mol H}_2\text{O} \times (2 \text{ mol H} / 1 \text{ mol H}_2\text{O}) = 0.02819 \text{ mol H}\]

\[0.5728 \text{ g total} - (0.2293 \text{ g C} + 0.02847 \text{ g H}) = 0.3151 \text{ g oxygen}\]

\[0.3151 \text{ g oxygen} \times (1 \text{ mol O} / 16.00 \text{ g O}) = 0.01969 \text{ mol O}\]

Dividing all moles by 0.01909 to reduce the ratio gives: \(C_1H_{1.48}O_{1.06}\)

Multiplying by 2 will give the smallest whole number ratio: \(2 (C_1H_{1.48}O_{1.06}) = C_2H_3O_2\)

Empirical molar mass = 59 g/mol, which is approximately 60 g/mol. ∴ the molecular formula is the same.