

Topic 1.3: Elemental Composition

Worksheet · AP Chemistry

UNIT 1 · ATOMIC STRUCTURE AND PROPERTIES

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How to Use This Worksheet

Work through the sections in order. Section A checks the ideas, Section B builds the formula method, and Section C mixes everything the way an exam would.

Use a periodic table and a calculator. Keep full precision through each calculation and round only at the end, because empirical formula ratios are easily spoiled by early rounding.

Do not look at the answer key until you have attempted every question. The key gives full working, so compare your method, not just your final formula.

Section A: Recall and Concepts

A1. State what is meant by the percent composition of a compound, and state what the percentages of all the elements in a compound must add up to.

A2. Explain the difference between an empirical formula and a molecular formula. Give one example where the two are the same and one where they are different.

A3. A compound has the molecular formula C_8H_{18} . Write its empirical formula.

A4. In a combustion analysis, the masses of CO_2 and H_2O produced are measured. Explain how the moles of hydrogen in the original sample are obtained from the mass of H_2O , and why the mass of oxygen in the compound cannot be obtained from the CO_2 in the same direct way.

A5. A student finds the empirical formula of a compound and reports it as the molecular formula, but the question gave no molar mass. Explain the error.

Section B: Core Calculations

B1. Calculate the percent composition by mass of nitrogen in ammonium nitrate, NH_4NO_3 .

B2. A hydrocarbon is 85.7% carbon and 14.3% hydrogen by mass. Determine its empirical formula.

B3. A compound is 52.2% carbon, 13.0% hydrogen and 34.8% oxygen by mass. Determine its empirical formula.

B4. A compound has empirical formula CH_2O and a molar mass of 150.13 g/mol. Determine its molecular formula.

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B5. Calculate the mass of oxygen present in a 25.0 g sample of calcium carbonate, CaCO_3 .

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B6. A sample of a compound contains 0.300 mol of carbon, 0.600 mol of hydrogen and 0.150 mol of oxygen. Write its empirical formula, showing how you obtained the ratio.

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Section C: Exam Style

C1. A 0.300 g sample of a compound containing only carbon, hydrogen and oxygen is burned completely. It produces 0.440 g of CO_2 and 0.180 g of H_2O . (a) Determine the empirical formula of the compound. (b) The molar mass of the compound is 60.05 g/mol. Determine its molecular formula.

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C2. An oxide of iron is 69.9% iron and 30.1% oxygen by mass. Determine the empirical formula of the oxide, and explain why the first ratio you obtain cannot be rounded directly to whole numbers.

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C3. An organic compound contains 74.02% carbon, 8.71% hydrogen and 17.27% nitrogen by mass. Its molar mass is 162.23 g/mol. Determine its molecular formula.

Answer Key

A1. Percent composition is the percentage of the compound's total mass contributed by each element: the mass of that element in one mole of the compound divided by the molar mass of the compound, multiplied by 100. The percentages of all elements present must add up to exactly 100%.

A2. The empirical formula is the lowest whole number ratio of atoms in the compound; the molecular formula is the actual number of atoms of each element in one molecule. Same: H_2O is both the empirical and the molecular formula of water. Different: glucose has molecular formula $\text{C}_6\text{H}_{12}\text{O}_6$ but empirical formula CH_2O .

A3. The ratio 8:18 reduces by dividing both subscripts by 2, giving C_4H_9 .

A4. Every carbon and hydrogen atom in the products came from the sample. Each mole of H_2O contains 2 mol of hydrogen atoms, so moles of H = $2 \times (\text{mass of } \text{H}_2\text{O} \div 18.02 \text{ g/mol})$. The oxygen in the CO_2 cannot be traced to the compound in this way, because most of it came from the oxygen gas the sample was burned in. Oxygen in the compound is found by difference: sample mass minus the masses of carbon and hydrogen.

A5. Mass or percent data fix only the ratio of atoms, which is the empirical formula. The molecular formula is a whole number multiple of the empirical formula, and the multiple can only be found from the molar mass. Without it, several molecular formulas (CH_2O , $\text{C}_2\text{H}_4\text{O}_2$, $\text{C}_6\text{H}_{12}\text{O}_6$ and so on) all fit the same data.

B1. $M(\text{NH}_4\text{NO}_3) = 2(14.01) + 4(1.008) + 3(16.00) = 28.02 + 4.03 + 48.00 = 80.05 \text{ g/mol}$. Both nitrogen atoms count: $\%N = 28.02 \div 80.05 \times 100 = 35.0\%$.

B2. In 100 g: $C = 85.7 \div 12.01 = 7.14 \text{ mol}$; $H = 14.3 \div 1.008 = 14.19 \text{ mol}$. Divide by the smallest (7.14): 1 : 1.99, which is 1 : 2 within rounding. Empirical formula: CH_2 .

B3. In 100 g: $C = 52.2 \div 12.01 = 4.35 \text{ mol}$; $H = 13.0 \div 1.008 = 12.90 \text{ mol}$; $O = 34.8 \div 16.00 = 2.18 \text{ mol}$. Divide by the smallest (2.18): 2.00 : 5.93 : 1, which rounds to 2 : 6 : 1. Empirical formula: $\text{C}_2\text{H}_6\text{O}$.

B4. Mass of the empirical unit $\text{CH}_2\text{O} = 12.01 + 2(1.008) + 16.00 = 30.03 \text{ g/mol}$. $n = 150.13 \div 30.03 = 5.00$. Molecular formula: $\text{C}_5\text{H}_{10}\text{O}_5$.

B5. $M(\text{CaCO}_3) = 100.09 \text{ g/mol}$, of which oxygen contributes $3(16.00) = 48.00 \text{ g}$. $\%O = 48.00 \div 100.09 \times 100 = 47.96\%$. Mass of oxygen = $0.4796 \times 25.0 \text{ g} = 12.0 \text{ g}$.

B6. Divide each amount by the smallest, 0.150: $C = 0.300 \div 0.150 = 2$; $H = 0.600 \div 0.150 = 4$; $O = 0.150 \div 0.150 = 1$. Empirical formula: $\text{C}_2\text{H}_4\text{O}$.

C1. (a) Moles of C = $0.440 \div 44.01 = 0.01000 \text{ mol}$, so mass of C = $0.01000 \times 12.01 = 0.120 \text{ g}$. Moles of H = $2 \times (0.180 \div 18.02) = 0.01998 \text{ mol}$, so mass of H = $0.01998 \times 1.008 = 0.0201 \text{ g}$. Mass of O = $0.300 - 0.120 - 0.0201 = 0.160 \text{ g}$, so moles of O = $0.160 \div 16.00 = 0.0100 \text{ mol}$. Ratio C : H : O = 0.0100 :

$0.0200 : 0.0100 = 1 : 2 : 1$. Empirical formula: CH_2O . (b) $n = 60.05 \div 30.03 = 2.00$. Molecular formula: $\text{C}_2\text{H}_4\text{O}_2$.

C2. In 100 g: Fe = $69.9 \div 55.85 = 1.252$ mol; O = $30.1 \div 16.00 = 1.881$ mol. Divide by the smallest (1.252): $1 : 1.50$. A ratio of 1.50 is not within rounding of a whole number, so it cannot simply be rounded to 1 or 2 without changing the compound. Multiply the whole ratio by 2: $2 : 3$. Empirical formula: Fe_2O_3 .

C3. In 100 g: C = $74.02 \div 12.01 = 6.163$ mol; H = $8.71 \div 1.008 = 8.641$ mol; N = $17.27 \div 14.01 = 1.233$ mol. Divide by the smallest (1.233): $5.00 : 7.01 : 1$, so the empirical formula is $\text{C}_5\text{H}_7\text{N}$, with unit mass $5(12.01) + 7(1.008) + 14.01 = 81.12$ g/mol. $n = 162.23 \div 81.12 = 2.00$. Molecular formula: $\text{C}_{10}\text{H}_{14}\text{N}_2$.

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