

Topic 1.3: Elemental Composition of Pure Substances

Revision Notes · AP Chemistry

UNIT 1 · ATOMIC STRUCTURE AND PROPERTIES

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What Percent Composition Tells You

The percent composition of a compound is the percentage of its total mass that comes from each element.

It is a property of the compound itself. A pure sample of water is always 11.2% hydrogen and 88.8% oxygen by mass, whether the sample is a drop or a tank of it.

Percent composition connects the formula of a compound to weighable quantities. It lets you predict how much of an element sits inside any mass of the compound, and, run in reverse, it lets you work out the formula from measured masses.

The formula. For each element: percent by mass = (number of atoms of the element × its molar mass) ÷ (molar mass of the compound) × 100.

The atom count comes from the subscripts in the formula, including any subscripts hidden inside brackets or after a dot in a hydrate.

Worked Example: Percent Composition of Glucose

Glucose has the formula $\text{C}_6\text{H}_{12}\text{O}_6$. Its molar mass is $6(12.01) + 12(1.008) + 6(16.00) = 72.06 + 12.10 + 96.00 = 180.16 \text{ g/mol}$.

Carbon. $\%C = 72.06 \div 180.16 \times 100 = 40.00\%$.

Hydrogen. $\%H = 12.10 \div 180.16 \times 100 = 6.71\%$.

Oxygen. $\%O = 96.00 \div 180.16 \times 100 = 53.29\%$.

The check. $40.00 + 6.71 + 53.29 = 100.00$. The percentages of all elements in a compound must always sum to 100, so add them up every time. A sum that misses 100 means an arithmetic slip somewhere above.

Empirical and Molecular Formulas

The **empirical formula** is the lowest whole number ratio of atoms in a compound. The **molecular formula** is the actual count of atoms in one molecule.

The two can be identical, as in CO_2 and H_2O , or different. Glucose has molecular formula $\text{C}_6\text{H}_{12}\text{O}_6$ and empirical formula CH_2O , because 6:12:6 reduces to 1:2:1.

Several different compounds can share one empirical formula. CH_2O is the empirical formula of glucose, of fructose, and also of ethanoic acid, $\text{C}_2\text{H}_4\text{O}_2$. An empirical formula identifies a ratio, never a unique compound.

Finding the Empirical Formula: The Method

Step 1: get masses. If you are given percentages, assume a 100 g sample. The percentages then become grams directly: 40.0% carbon means 40.0 g of carbon.

Step 2: convert each mass to moles. Divide each element's mass by its molar mass.

Step 3: divide every value by the smallest. This gives the mole ratio relative to the least abundant element.

Step 4: turn the ratio into whole numbers. If every value is within rounding of a whole number, you are done. If not, multiply the whole ratio by 2, 3 or 4 until it is.

Worked example. A compound is 40.0% C, 6.7% H and 53.3% O by mass. In a 100 g sample: $\text{C} = 40.0 \div 12.01 = 3.33 \text{ mol}$; $\text{H} = 6.7 \div 1.008 = 6.65 \text{ mol}$; $\text{O} = 53.3 \div 16.00 = 3.33 \text{ mol}$.

Dividing by the smallest, 3.33, gives 1 : 2 : 1. The empirical formula is CH_2O .

From Empirical to Molecular

An empirical formula alone can never give the molecular formula. You also need the compound's molar mass.

The multiplier. $n = \text{molar mass of the compound} \div \text{mass of the empirical formula unit}$. Then multiply every subscript in the empirical formula by n .

Worked example. The empirical formula CH_2O has a unit mass of $12.01 + 2(1.008) + 16.00 = 30.03$ g/mol. A compound with this empirical formula and a molar mass of 180.16 g/mol has $n = 180.16 \div 30.03 = 6.00$.

Its molecular formula is $\text{C}_6\text{H}_{12}\text{O}_6$: glucose again. Steps 1.3's two directions meet: the percent composition of glucose produced CH_2O , and its molar mass rebuilds the full formula.

Combustion Analysis

Combustion analysis finds a formula experimentally. A weighed sample of a compound containing carbon, hydrogen and possibly oxygen is burned completely in excess oxygen.

Where each element goes. All the carbon ends up in the CO_2 produced, and all the hydrogen ends up in the H_2O . The masses of CO_2 and H_2O are measured, and the moles of each product give the moles of carbon and hydrogen in the original sample.

Each mole of CO_2 holds one mole of carbon atoms. Each mole of H_2O holds two moles of hydrogen atoms. That factor of two is the step this method is built around.

Oxygen by difference. If the compound contains only C, H and O, the mass of oxygen is what is left: sample mass minus the mass of carbon minus the mass of hydrogen. Oxygen cannot be measured this way when other elements are present.

Worked example. A 0.200 g sample of a C, H and O compound produces 0.293 g of CO_2 and 0.120 g of H_2O on complete combustion.

Moles of C = $0.293 \div 44.01 = 0.00666$ mol, which is $0.00666 \times 12.01 = 0.0800$ g of carbon. Moles of H = $2 \times (0.120 \div 18.02) = 0.0133$ mol, which is 0.0134 g of hydrogen.

Mass of O = $0.200 - 0.0800 - 0.0134 = 0.1066$ g, so moles of O = $0.1066 \div 16.00 = 0.00666$ mol. The ratio $0.00666 : 0.0133 : 0.00666$ reduces to 1 : 2 : 1, and the empirical formula is CH_2O .

Traps That Cost Marks in Topic 1.3

Dividing by the largest value instead of the smallest. Step 3 normalizes against the least abundant element. Dividing by the largest produces fractions below 1 and a ratio that will not clean up.

Rounding 1.33 or 1.5 to 1 too early. A ratio of 1 : 1.5 is not 1 : 1 or 1 : 2. Multiply everything by 2 to get 2 : 3. Ratios ending in .5 need $\times 2$, in .33 or .67 need $\times 3$, in .25 or .75 need $\times 4$.

Forgetting the factor of two for hydrogen in combustion. The moles of H_2O are not the moles of hydrogen. Each water molecule carries two hydrogen atoms out of the sample.

Finding oxygen from the CO₂ and H₂O masses. The oxygen in the products mostly came from the oxygen gas used to burn the sample, not from the compound. Compound oxygen comes only from the mass difference, and only for C, H and O compounds.

Giving a molecular formula without a molar mass. Percent or mass data alone fix only the ratio. If a question asks for the molecular formula, it must give a molar mass somewhere; if it does not, the answer expected is the empirical formula.

Treating percent as moles. 40.0% carbon is a mass statement. It becomes moles only after dividing by the molar mass; skipping that division compares masses as if they were atom counts.

Key Relationships to Remember

- Percent by mass of an element = $(\text{atoms of the element} \times \text{molar mass of the element}) \div (\text{molar mass of the compound}) \times 100$; all percentages sum to 100
- Empirical formula method: mass of each element → moles → divide by the smallest → whole number ratio
- Percent data: assume a 100 g sample so percentages become grams
- Molecular formula = empirical formula $\times n$, where $n = (\text{molar mass of compound}) \div (\text{mass of empirical unit})$; n is always a whole number
- Combustion: moles of C = moles of CO₂; moles of H = $2 \times$ moles of H₂O; oxygen by difference for C, H and O compounds only

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