AP Chemistry • Unit 1: Atomic Structure and Properties

Topic 1.3

Elemental Composition of Pure Substances

Percent Composition, Empirical & Molecular Formulas

What You'll Master in This Guide

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Law of Definite Proportions

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⚖️ Concept 1: Law of Definite Proportions

📖 What is the Law of Definite Proportions?

The **Law of Definite Proportions** (also called the **Law of Constant Composition**) was established by French chemist Joseph Proust in 1799. It states:

> "A pure chemical compound always contains the same elements in the same proportions by mass, regardless of the source or amount."

💧 Example: Water (H_2O) Composition

Every sample of pure water, no matter where it comes from, has the exact same composition:

11.2%	88.8%
Hydrogen by mass	Oxygen by mass

Rain ☔ | Ocean 🌊 | Ice ❄️ | Lab 🧪 — ALL have identical composition!

🧮 Calculating Water's Composition

Molar masses: H = 1.008 g/mol, O = 16.00 g/mol

H_2O composition:

- H: 2 × 1.008 = 2.016 g/mol
- O: 1 × 16.00 = 16.00 g/mol
- Total: 18.016 g/mol

Percent H: (2.016 ÷ 18.016) × 100% = **11.2%**

Percent O: (16.00 ÷ 18.016) × 100% = **88.8%**

💡 Why This Law Matters

This law allows chemists to:
- Verify the purity of compounds
- Detect counterfeit products (like fake medicines)
- Ensure quality control in manufacturing
- Calculate expected masses in reactions

⚠️ Pure Compounds vs. Mixtures

Pure compounds: Fixed composition (H_2O, NaCl, CO_2) — follow this law ✓
Mixtures: Variable composition (air, seawater, soil) — do NOT follow this law ✗

⚖️ Concept 1: Law of Definite Proportions (Continued)

🔬 More Examples of Fixed Composition

Compound	Formula	Element 1	Element 2	Element 3
Water	H_2O	H: 11.2%	O: 88.8%	—
Carbon dioxide	CO_2	C: 27.3%	O: 72.7%	—
Table salt	NaCl	Na: 39.3%	Cl: 60.7%	—
Ammonia	NH_3	N: 82.3%	H: 17.7%	—
Glucose	$C_6H_{12}O_6$	C: 40.0%	H: 6.7%	O: 53.3%

💊 Real-World Application: Pharmaceutical Quality Control

Pharmaceutical companies use this law to verify drug purity. If a medication sample doesn't have the expected elemental composition, it indicates a problem:

❌ Counterfeit Product

Wrong or missing active ingredients

❌ Contamination

Impurities affecting composition

❌ Chemical Degradation

Breakdown over time or improper storage

❌ Manufacturing Error

Incorrect synthesis or formulation

📊 Example: Aspirin ($C_9H_8O_4$)

Pure aspirin always contains:

- Carbon: 60.0% (calculated: 9×12.01/180.16 × 100%)
- Hydrogen: 4.5% (calculated: 8×1.008/180.16 × 100%)
- Oxygen: 35.5% (calculated: 4×16.00/180.16 × 100%)

Analysis shows different values? → The sample is NOT pure aspirin!

⚡ Key Points to Remember

✓ Pure compounds have FIXED composition by mass — always the same ratio

✓ The composition is independent of source, location, or sample size

✓ Mixtures have VARIABLE composition — ratios can change

✓ This law is the foundation for quality control in industry

✓ Deviation from expected composition indicates impurity or counterfeit

📊 Concept 2: Percent Composition by Mass

📖 What is Percent Composition?

Percent composition by mass tells us what percentage of a compound's total mass comes from each element. It quantifies the Law of Definite Proportions numerically.

$$\% \text{ Element} = (\text{Mass of element in 1 mol compound} / \text{Molar mass of compound}) \times 100\%$$

📐 Step-by-Step Process

1	→	2	→	3	→	4
Write Formula Identify all elements		**Calculate Mass** Each element × subscript		**Find Total** Sum all element masses		**Calculate %** (part/total) × 100

🧮 Example 1: Carbon Dioxide (CO_2)

Step 1: Formula = CO_2 (1 carbon, 2 oxygen atoms)

Step 2: Calculate mass of each element:

C: 1 atom × 12.01 g/mol = 12.01 g/mol

O: 2 atoms × 16.00 g/mol = 32.00 g/mol

Step 3: Total molar mass = 12.01 + 32.00 = 44.01 g/mol

Step 4: Calculate percent composition:

%C = (12.01 / 44.01) × 100% = **27.3%**

%O = (32.00 / 44.01) × 100% = **72.7%**

Verification: 27.3% + 72.7% = 100.0% ✓

🧮 Example 2: Glucose ($C_6H_{12}O_6$)

Step 1: Formula = $C_6H_{12}O_6$ (6 C, 12 H, 6 O)

Step 2: Calculate element masses:

C: 6 × 12.01 = 72.06 g/mol

H: 12 × 1.008 = 12.10 g/mol

O: 6 × 16.00 = 96.00 g/mol

Step 3: Total = 72.06 + 12.10 + 96.00 = 180.16 g/mol

Step 4: Percent composition:

%C = (72.06/180.16) × 100% = **40.0%**

%H = (12.10/180.16) × 100% = **6.7%**

%O = (96.00/180.16) × 100% = **53.3%**

Example 3: Calcium Nitrate $Ca(NO_3)_2$ — Handling Parentheses!

Step 1: Expand the formula carefully:

$Ca(NO_3)_2$ means: 1 Ca, 2 N, and 6 O (the 2 multiplies EVERYTHING inside)

Step 2: Calculate element masses:

Ca: 1 × 40.08 = 40.08 g/mol

N: 2 × 14.01 = 28.02 g/mol

O: 6 × 16.00 = 96.00 g/mol

Step 3: Total = 40.08 + 28.02 + 96.00 = 164.10 g/mol

Step 4: Percent composition:

%Ca = (40.08/164.10) × 100% = **24.4%**

%N = (28.02/164.10) × 100% = **17.1%**

%O = (96.00/164.10) × 100% = **58.5%**

Verification: 24.4 + 17.1 + 58.5 = 100.0% ✓

Common Mistakes to Avoid

❌ **Mistake 1: Forgetting to multiply by subscripts**

In H_2O, hydrogen mass = 2 × 1.008 = 2.016, NOT just 1.008

❌ **Mistake 2: Using atomic NUMBER instead of atomic MASS**

Carbon = 12.01 g/mol (mass), NOT 6 (atomic number)

❌ **Mistake 3: Ignoring parentheses**

In $Ca(NO_3)_2$: oxygen = 3 × 2 = 6 atoms, NOT 3

❌ **Mistake 4: Not verifying your answer**

All percentages MUST sum to 100% (within rounding error)

Reverse Calculation: Finding Mass from Percent

You can use percent composition to find the mass of an element in any sample:

$$\text{Mass of element} = (\%\ \text{composition} / 100) \times \text{Total mass of sample}$$

Example: How much oxygen is in 50.0 g of water?

Given: Water is 88.8% oxygen by mass

Calculate: Mass of O = (88.8/100) × 50.0 g = **44.4 g oxygen**

Check: Mass of H = (11.2/100) × 50.0 g = 5.6 g hydrogen

Verify: 44.4 g + 5.6 g = 50.0 g ✓ (Total matches original sample)

Percent Composition Checklist

✓ Count ALL atoms of each element (watch for subscripts and parentheses!)

✓ Multiply atomic mass by number of atoms for EACH element

✓ Sum all element masses to get total molar mass

✓ Divide each element's mass by total, then multiply by 100%

✓ ALWAYS VERIFY: All percentages must sum to 100%

🔬 Concept 3: Empirical Formulas

📖 What is an Empirical Formula?

An **empirical formula** shows the **simplest whole-number ratio** of atoms in a compound. It represents the smallest set of subscripts that maintain the correct proportions, but may NOT show the actual number of atoms in a molecule.

Compound Name	Molecular Formula	Empirical Formula	Simplification
Glucose	$C_6H_{12}O_6$	CH_2O	÷ 6
Hydrogen peroxide	H_2O_2	HO	÷ 2
Benzene	C_6H_6	CH	÷ 6
Acetic acid	$C_2H_4O_2$	CH_2O	÷ 2
Water	H_2O	H_2O	already simplest

📐 How to Calculate Empirical Formula

1. **Mass → Moles** — Divide by molar mass →
2. **÷ Smallest** — Find mole ratio →
3. **× to Whole #** — If needed →
4. **Write Formula** — Use subscripts

🧮 Example 1: From Mass Data

Problem: A compound contains 40.0 g carbon, 6.7 g hydrogen, and 53.3 g oxygen. Find the empirical formula.

Step 1: Convert mass to moles

C: 40.0 g ÷ 12.01 g/mol = 3.33 mol

H: 6.7 g ÷ 1.008 g/mol = 6.65 mol

O: 53.3 g ÷ 16.00 g/mol = 3.33 mol

Step 2: Divide all by smallest value (3.33 mol)

C: 3.33 ÷ 3.33 = 1.00

H: 6.65 ÷ 3.33 = 2.00

O: 3.33 ÷ 3.33 = 1.00

Step 3: All are already whole numbers! No multiplication needed.

Step 4: Write empirical formula → CH_2O

💡 Working with Percentages

When given percent composition instead of mass, **assume a 100 g sample**. This clever trick converts percentages directly to grams!

Example: 40.0% C, 6.7% H, 53.3% O → In 100 g: 40.0 g C, 6.7 g H, 53.3 g O

🧮 Example 2: Handling Non-Whole Numbers

Problem: A compound contains 2.90 g nitrogen and 8.30 g oxygen. Find the empirical formula.

Step 1: Convert to moles

N: 2.90 g ÷ 14.01 g/mol = 0.207 mol

O: 8.30 g ÷ 16.00 g/mol = 0.519 mol

Step 2: Divide by smallest (0.207 mol)

N: 0.207 ÷ 0.207 = 1.00

O: 0.519 ÷ 0.207 = 2.51 ≈ **2.5**

Step 3: Multiply to get whole numbers

2.5 is the same as 5/2, so multiply ALL values by 2:

N: 1.00 × 2 = 2

O: 2.5 × 2 = 5

Step 4: Write empirical formula → N_2O_5

🔢 Common Decimal-to-Whole Number Conversions

When you get a decimal after dividing, recognize the fraction and multiply accordingly:

If you get approximately...	It represents fraction...	Multiply ALL by...	Example result
1.25	5/4	**4**	1.25 × 4 = 5
1.33	4/3	**3**	1.33 × 3 = 4
1.50	3/2	**2**	1.50 × 2 = 3
1.67	5/3	**3**	1.67 × 3 = 5
2.50	5/2	**2**	2.50 × 2 = 5
2.67	8/3	**3**	2.67 × 3 = 8

⚠️ Common Mistakes to Avoid

❌ **Mistake 1: Rounding too aggressively**

1.49 should stay as 1.5 (then multiply by 2), NOT rounded to 1

❌ **Mistake 2: Not multiplying ALL elements**

If you multiply one element by 2, you MUST multiply ALL elements by 2

❌ **Mistake 3: Leaving decimals in the final answer**

Empirical formulas must have whole number subscripts only — never $CH_{2.5}O$

⚡ Empirical Formula Checklist

✓ For percentages, assume 100 g sample (% becomes grams directly)

✓ Convert ALL masses to moles using molar mass from periodic table

✓ Divide ALL mole values by the SMALLEST mole value

✓ If ratios aren't whole numbers, multiply ALL by 2, 3, or 4 as needed

✓ Final formula must have only whole number subscripts

Concept 4: Molecular Formulas

Empirical vs. Molecular Formula

The molecular formula shows the **actual number** of atoms in one molecule, while the empirical formula shows only the **simplest ratio**.

Empirical Formula

- Simplest whole-number ratio
- May NOT represent actual molecule
- Example: CH_2O (simplest ratio)

Molecular Formula

- Actual number of atoms
- True composition of molecule
- Example: $C_6H_{12}O_6$ (glucose)

$$\text{Molecular Formula} = (\text{Empirical Formula}) \times n$$

where n = Molar mass of compound ÷ Empirical formula mass
n must be a whole number (1, 2, 3, 4, ...)

How to Find Molecular Formula

Step 1: Determine the empirical formula (if not already given)

Step 2: Calculate the empirical formula mass (add up all atomic masses)

Step 3: Find n = (Given molar mass) ÷ (Empirical formula mass)

Step 4: Multiply ALL subscripts in empirical formula by n

Step 5: Verify by calculating the molar mass of your answer

Example 1: Finding Molecular Formula

Problem: A compound has empirical formula CH_2O and molar mass 180 g/mol. Find the molecular formula.

Step 1: Empirical formula = CH_2O (given)

Step 2: Calculate empirical formula mass:

C: 1 × 12.01 = 12.01 g/mol
H: 2 × 1.008 = 2.016 g/mol
O: 1 × 16.00 = 16.00 g/mol
Total empirical mass = 30.03 g/mol

Step 3: Find n:

n = 180 g/mol ÷ 30.03 g/mol = **6**

Step 4: Multiply all subscripts by 6:

$(CH_2O) \times 6 = \mathbf{C_6H_{12}O_6}$

Step 5: Verify: 6(12.01) + 12(1.008) + 6(16.00) = 180.16 g/mol ✓

Example 2: Complete Problem (% → Empirical → Molecular)

Problem: A compound is 85.7% C and 14.3% H. Molar mass = 56 g/mol. Find molecular formula.

Part A - Find Empirical Formula:

Assume 100 g sample: 85.7 g C, 14.3 g H

Moles of C: 85.7 g ÷ 12.01 g/mol = 7.14 mol

Moles of H: 14.3 g ÷ 1.008 g/mol = 14.2 mol

Divide by smallest: C = 7.14/7.14 = 1; H = 14.2/7.14 = 2

Empirical formula: CH_2

Part B - Find Molecular Formula:

Empirical mass = 12.01 + 2(1.008) = 14.03 g/mol

n = 56 g/mol ÷ 14.03 g/mol = 4

Molecular formula: $(CH_2) \times 4 = C_4H_8$

Same Empirical, Different Molecular Formulas

Multiple compounds can share the same empirical formula but have different molecular formulas. The molar mass determines which one it is!

Empirical	Compound	Molecular	n	Molar Mass
CH	Acetylene	C_2H_2	2	26 g/mol
	Benzene	C_6H_6	6	78 g/mol
	Styrene	C_8H_8	8	104 g/mol
CH_2O	Formaldehyde	CH_2O	1	30 g/mol
	Acetic acid	$C_2H_4O_2$	2	60 g/mol
	Glucose	$C_6H_{12}O_6$	6	180 g/mol

When n = 1

When n = 1, the empirical and molecular formulas are **identical**. The compound's molar mass equals its empirical formula mass. Examples: H_2O, NaCl, CO_2, NH_3, CH_4 — all already in simplest form!

Common Mistakes to Avoid

❌ **Skipping empirical formula:** You MUST find empirical before molecular
❌ **Non-integer n:** If n isn't close to a whole number, recheck your calculations
❌ **Partial multiplication:** Multiply ALL subscripts by n, not just some

Molecular Formula Checklist

✓ Find empirical formula first (from mass or percent data)

✓ Calculate empirical formula mass accurately

✓ n = Molar mass ÷ Empirical mass (must be a whole number!)

✓ Multiply ALL subscripts in empirical formula by n

✓ Verify: Molar mass of molecular formula should match given value

📋 Quick Reference Summary

⚖️ Concept 1: Law of Definite Proportions

Pure compounds have **fixed composition by mass**, regardless of source, location, or sample size. Mixtures do NOT follow this law — their composition can vary.

$$\% \text{ Element} = (\text{Mass of element} / \text{Molar mass of compound}) \times 100\%$$

All percentages must sum to 100% — use this to verify your answer!

🔬 Concept 3: Empirical Formula Process

1. **% → grams** — Assume 100g sample →
2. **g → mol** — ÷ molar mass →
3. **÷ smallest** — Get ratio →
4. **× if needed** — Whole numbers

$$\text{Molecular Formula} = \text{Empirical Formula} \times n$$

n = Molar mass ÷ Empirical formula mass (must be a whole number)

🔢 Common Multipliers for Non-Whole Numbers

Decimal	≈ Fraction	× by	Decimal	≈ Fraction	× by
1.25	5/4	4	1.67	5/3	3
1.33	4/3	3	2.50	5/2	2
1.50	3/2	2	2.67	8/3	3

🚫 Top 5 Mistakes to Avoid

1. Using atomic NUMBER (e.g., 6 for C) instead of atomic MASS (12.01 for C)
2. Forgetting subscripts outside parentheses: $Ca(NO_3)_2$ has 6 oxygen atoms, not 3
3. Not dividing ALL mole values by the smallest when finding empirical formula
4. Leaving decimals in final formula — subscripts must be whole numbers
5. Not verifying that percent compositions sum to 100%

🎯 AP Exam Success Tips

- **Show ALL work** — partial credit is available for correct steps
- **Include units** in every calculation (g, mol, g/mol)
- **Box or circle** your final answer
- **Verify your answer** — does it make sense? Do percentages = 100%?

Question 1

The Law of Definite Proportions states that:

A) Mixtures have fixed composition

B) Pure compounds have fixed composition by mass

C) Elements combine in any ratio

D) Air has constant composition

✅ Answer: B

Explanation: The Law of Definite Proportions (Proust, 1799) states that pure compounds always contain the same elements in the same proportions by mass, regardless of source. Water is always 11.2% H and 88.8% O by mass, whether from rain, ocean, or a lab. Mixtures (A) have variable composition. Elements can combine in specific ratios, not any ratio (C). Air is a mixture with variable composition (D).

Question 2

What is the molar mass of H_2O? (H = 1.008 g/mol, O = 16.00 g/mol)

A) 17.008 g/mol

B) 18.016 g/mol

C) 18.00 g/mol

D) 19.008 g/mol

✅ Answer: B

Explanation: Molar mass = sum of all atomic masses × their subscripts.
H_2O = 2(1.008) + 1(16.00) = 2.016 + 16.00 = **18.016 g/mol**
Common error (A): Using only 1 hydrogen instead of 2. Always multiply by the subscript!

Question 3

Which of the following is an empirical formula?

A) $C_6H_{12}O_6$

B) $C_2H_4O_2$

C) CH_2O

D) $C_3H_6O_3$

✅ Answer: C

Explanation: An empirical formula shows the **simplest whole-number ratio**. Let's check each:
- $C_6H_{12}O_6$ → divide by 6 → CH_2O (not simplest)
- $C_2H_4O_2$ → divide by 2 → CH_2O (not simplest)
- CH_2O → already 1:2:1 ratio (SIMPLEST!) ✓
- $C_3H_6O_3$ → divide by 3 → CH_2O (not simplest)

All four compounds share the same empirical formula: CH_2O

Question 4

To find percent composition, you divide element mass by:

A) Atomic number

B) Total molar mass of compound

C) Number of atoms

D) Avogadro's number

✅ Answer: B

Explanation: The percent composition formula is:
% Element = (Mass of element / Total molar mass) × 100%
You compare the mass contribution of one element to the total mass of the compound. Atomic number (A) is the number of proton related to percent composition.

Question 5

The molecular formula is always:

A) Smaller than the empirical formula

B) A whole-number multiple of the empirical formula

C) Equal to the empirical formula

D) Unrelated to the empirical formula

✅ Answer: B

Explanation: Molecular formula = Empirical formula × n, where n ≥ 1.
- If n = 1: Molecular = Empirical (like H_2O)
- If n = 2: Molecular = 2 × Empirical (like H_2O_2 from HO)
- If n = 6: Molecular = 6 × Empirical (like $C_6H_{12}O_6$ from CH_2O)

It's never smaller (A) and always related (D). Not always equal (C) — only when n = 1.

Question 6

When given percent composition data, what sample size should you assume?

A) 1 g sample

B) 10 g sample

C) 100 g sample

D) 1 mol sample

✅ Answer: C

Explanation: Assuming a **100 g sample** is the standard technique because it converts percentages directly to grams!
Example: If a compound is 40.0% C, 6.7% H, 53.3% O:
In 100 g sample → 40.0 g C, 6.7 g H, 53.3 g O
This makes calculations much simpler than using other sample sizes.

Question 7

Which of the following follows the Law of Definite Proportions?

A) Air

B) Seawater

C) Pure NaCl

D) Soil

✅ Answer: C

Explanation: Only **pure compounds** follow the Law of Definite Proportions.
- Air (A): Mixture of gases — composition varies by location and altitude
- Seawater (B): Mixture — salt content varies by location
- **Pure NaCl (C): Compound — always 39.3% Na, 60.7% Cl ✓**
- Soil (D): Mixture — composition varies greatly by location

Question 8

In $Ca(NO_3)_2$, how many oxygen atoms are present?

A) 2

B) 3

C) 5

D) 6

✅ Answer: D

Explanation: The subscript 2 outside the parentheses multiplies EVERYTHING inside:
$Ca(NO_3)_2$ = Ca + 2 × (NO_3) = Ca + 2N + 2×3O = Ca + 2N + **6O**
Common error: Thinking there are only 3 oxygens (ignoring the ×2 multiplier).

Question 11

The empirical formula of C_4H_{10} is:

A) C_4H_{10}

B) C_2H_5

C) $CH_{2.5}$

D) CH_3

✅ Answer: B

Explanation: To find empirical formula, divide all subscripts by their greatest common factor (GCF).
C_4H_{10}: GCF of 4 and 10 is 2
C: 4 ÷ 2 = 2
H: 10 ÷ 2 = 5
Empirical formula: C_2H_5
Note: C is wrong because you can't have decimal subscripts ($CH_{2.5}$). Empirical formulas must have whole numbers.

Question 12

Percent composition values for all elements in a compound should always:

A) Equal 50% each

B) Be greater than 100%

C) Sum to 100%

D) Be equal for all elements

✅ Answer: C

Explanation: The percent composition represents the fraction of total mass from each element. Since all mass must be accounted for:
$\%Element_1 + \%Element_2 + \%Element_3 + ... = 100\%$
This is a crucial check for your calculations! If your percentages don't sum to approximately 100%, you've made an error.

Question 13

To find n (the multiplier for molecular formula), you calculate:

A) Empirical mass ÷ molar mass

B) Molar mass ÷ empirical mass

C) Atomic mass ÷ molar mass

D) Molar mass ÷ atomic mass

✅ Answer: B

Explanation: The formula is:
n = Molar mass of compound ÷ Empirical formula mass
Example: If empirical formula CH_2O has mass 30 g/mol and the compound has molar mass 180 g/mol:
n = 180 ÷ 30 = 6
Molecular formula = $(CH_2O) \times 6 = C_6H_{12}O_6$

Question 14

H_2O_2 has the empirical formula:

A) H_2O_2

B) HO

C) H_2O

D) HO_2

✅ Answer: B

Explanation: H_2O_2 has 2 hydrogen and 2 oxygen atoms. To find the empirical formula, divide by the GCF:
GCF of 2 and 2 is 2
H: 2 ÷ 2 = 1

Question 16

What is the percent composition of oxygen in H_2SO_4? (H=1.008, S=32.07, O=16.00 g/mol)

A) 16.3%

B) 32.7%

C) 65.3%

D) 81.6%

✅ Answer: C

Explanation:
Step 1: Calculate molar mass of H_2SO_4
H: 2 × 1.008 = 2.016 g/mol
S: 1 × 32.07 = 32.07 g/mol
O: 4 × 16.00 = 64.00 g/mol
Total = 2.016 + 32.07 + 64.00 = 98.09 g/mol

Step 2: Calculate %O
%O = (64.00 / 98.09) × 100% = **65.3%**

Question 17

A compound contains 3.0 mol C and 6.0 mol H. The empirical formula is:

A) C_3H_6

B) CH_2

C) C_2H_4

D) CH

✅ Answer: B

Explanation:
Step 1: Already have moles — find smallest value
Smallest = 3.0 mol

Step 2: Divide both by smallest
C: 3.0 ÷ 3.0 = 1
H: 6.0 ÷ 3.0 = 2

Empirical formula: CH_2
Note: C_3H_6 (A) and C_2H_4 (C) are not simplified — they can both be reduced to CH_2.

Question 18

After dividing moles by the smallest value, you get $N_1O_{2.5}$. You should:

A) Round to NO_3

B) Round to NO_2

C) Multiply both by 2 to get N_2O_5

D) Leave as $NO_{2.5}$

✅ Answer: C

Explanation: Empirical formulas cannot have decimal subscripts!
When you get 2.5 (which = 5/2), multiply ALL subscripts by 2:
N: 1 × 2 = 2
O: 2.5 × 2 = 5
Empirical formula: N_2O_5

Never round decimals like 2.5 (would give wrong formula). Never leave decimals in final answer.

FRQ 1: Percent Composition of Water

8 Points

Water (H_2O) is essential for all known forms of life. Use the following atomic masses: H = 1.008 g/mol, O = 16.00 g/mol.

(a) Calculate the molar mass of water. Show your work clearly. *(2 points)*

(b) Calculate the percent composition of hydrogen and oxygen in water. *(2 points)*

(c) If you have a 45.0 g sample of pure water, calculate the mass of oxygen it contains. *(2 points)*

(d) A student analyzes water from a mountain stream and finds it is 11.0% hydrogen. Is this consistent with the Law of Definite Proportions? Explain. *(2 points)*

📝 Complete Solution — FRQ 1

(a) Molar mass of water (2 points)

H_2O contains 2 hydrogen atoms and 1 oxygen atom.
Mass from H: 2 × 1.008 g/mol = 2.016 g/mol **[1 point]**
Mass from O: 1 × 16.00 g/mol = 16.00 g/mol
Total molar mass = 2.016 + 16.00 = 18.016 g/mol [1 point]

(b) Percent composition (2 points)

%H = (mass of H / total mass) × 100%
%H = (2.016 / 18.016) × 100% = **11.19% ≈ 11.2% [1 point]**

%O = (16.00 / 18.016) × 100% = **88.81% ≈ 88.8% [1 point]**

Verification: 11.2% + 88.8% = 100% ✓

(c) Mass of oxygen in 45.0 g sample (2 points)

Mass of O = (% oxygen / 100) × sample mass **[1 point for setup]**
Mass of O = (88.8 / 100) × 45.0 g
Mass of O = 0.888 × 45.0 g = **40.0 g oxygen [1 point for answer]**

Alternative verification: Mass of H = 0.112 × 45.0 g = 5.0 g; Total = 40.0 + 5.0 = 45.0 g ✓

(d) Consistency with Law of Definite Proportions (2 points)

Yes, this is approximately consistent. [1 point]

The calculated value is 11.2% hydrogen, and the measured value is 11.0%. This small difference (0.2%) is within experimental error for most analytical methods. According to the Law of Definite Proportions, pure water from ANY source should have the same composition. The slight variation is due to measurement uncertainty, not a violation of the law. **[1 point for explanation]**

Note: If the difference were large (e.g., 15% vs 11.2%), it would suggest the sample is not pure water or contains dissolved substances.

FRQ 2: Basic Empirical Formula Determination

7 Points

A hydrocarbon (compound containing only carbon and hydrogen) is analyzed. The analysis shows the compound contains 3.0 mol of carbon and 8.0 mol of hydrogen.

(a) Determine the empirical formula of this compound. Show all steps in your calculation. *(3 points)*

(b) Calculate the percent composition of carbon and hydrogen in this compound. *(2 points)*

(c) If the molar mass of the compound is 44 g/mol, determine the molecular formula. *(2 points)*

Complete Solution — FRQ 2

(a) Empirical formula (3 points)

Step 1: Identify the smallest number of moles **[1 point]**
C: 3.0 mol, H: 8.0 mol
Smallest = 3.0 mol

Step 2: Divide all mole values by the smallest **[1 point]**
C: 3.0 ÷ 3.0 = 1.00
H: 8.0 ÷ 3.0 = 2.67

Step 3: Convert to whole numbers **[1 point]**
2.67 ≈ 8/3, so multiply all by 3:
C: 1.00 × 3 = 3
H: 2.67 × 3 = 8

Empirical formula: C_3H_8

(b) Percent composition (2 points)

First, calculate empirical formula mass:
C_3H_8 = 3(12.01) + 8(1.008) = 36.03 + 8.064 = 44.09 g/mol

%C = (36.03 / 44.09) × 100% = **81.7% [1 point]**
%H = (8.064 / 44.09) × 100% = **18.3% [1 point]**

Verification: 81.7% + 18.3% = 100% ✓