

Weird Chemist

DPP-6 : Empirical Formula

Chapter: Some Basic Concepts of Chemistry

Solution Sheet — Part 1 (TYPE 1, Q.1–Q.20)

5-Step Method to Find Empirical Formula

Step 1. Assume 100 g of compound. Mass of each element = its % value in grams.

Step 2. Divide each mass by its atomic weight to get **moles of atoms**.

Step 3. Divide all mole values by the **smallest** mole value.

Step 4. If ratios are not whole numbers, multiply all by a suitable integer ($\times 2$ for 1.5, $\times 3$ for 1.33).

Step 5. Write empirical formula using these whole-number ratios as subscripts.

TYPE 1 : Empirical Formula from % Composition

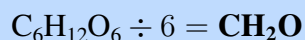
Q.1 Empirical formula of glucose is

Explanation

Glucose molecular formula: $C_6H_{12}O_6$.

To find empirical formula, divide all subscripts by their HCF (Highest Common Factor).

HCF of 6, 12, 6 = 6.



Empirical formula = CH_2O (molar mass = 30 g/mol).

Approach / Analogy

The empirical formula is the **simplest whole-number ratio** of atoms. Glucose has C:H:O = 6:12:6 = 1:2:1 (dividing by 6). So empirical formula = CH_2O . Think of it as the “reduced fraction” version of the molecular formula, like simplifying 6/12/6 to 1/2/1.

Common Mistake

Choosing $C_2H_4O_2$ (option 3) thinking HCF = 3. But HCF of 6, 12, 6 is **6**, not 3. $6/6 = 1$, $12/6 = 2$, $6/6 = 1 \rightarrow CH_2O$. Always find the *highest* common factor, not just any common factor.

Answer

(4) CH_2O

Q.2 Which of the following compounds has same empirical formula as that of glucose?

Explanation

Glucose empirical formula = CH_2O . Now check each option:

- $\text{CH}_3\text{CHO} = \text{C}_2\text{H}_4\text{O} \rightarrow \text{C:H:O} = 2:4:1$. HCF = 1. EF = $\text{C}_2\text{H}_4\text{O}$. ×
- **$\text{CH}_3\text{COOH} = \text{C}_2\text{H}_4\text{O}_2 \rightarrow \text{C:H:O} = 2:4:2 \div 2 = 1:2:1 = \text{CH}_2\text{O}$** . ✓
- $\text{CH}_3\text{OH} = \text{CH}_4\text{O} \rightarrow \text{C:H:O} = 1:4:1$. EF = CH_4O . ×
- $\text{C}_2\text{H}_6 \rightarrow \text{C:H} = 2:6 \div 2 = 1:3 = \text{CH}_3$. ×

CH_3COOH (acetic acid) has the same empirical formula CH_2O as glucose.

Approach / Analogy

CH_3COOH (acetic acid) has the molecular formula $\text{C}_2\text{H}_4\text{O}_2$. Dividing by 2: CH_2O . Same empirical formula as glucose ($\text{C}_6\text{H}_{12}\text{O}_6 \div 6 = \text{CH}_2\text{O}$). This is a beautiful chemistry fact: vinegar and blood sugar share the same empirical formula!

Common Mistake

Picking CH_3CHO (acetaldehyde) thinking it simplifies to CH_2O . $\text{CH}_3\text{CHO} = \text{C}_2\text{H}_4\text{O}$. $\text{C:H:O} = 2:4:1$. HCF = 1 (since O has subscript 1). EF = $\text{C}_2\text{H}_4\text{O} \neq \text{CH}_2\text{O}$. Always check **all three** subscripts for HCF, not just C and H.

Answer

(2) CH_3COOH

Q.3 In a compound x is 75.8% and y is 24.2% by weight present. If atomic weight of x and y are 24 and 16 respectively, then the empirical formula of the compound is

Explanation

In 100 g compound: 75.8 g of x, 24.2 g of y.

$$n_x = \frac{75.8}{24} = 3.158 \text{ mol atoms}$$

$$n_y = \frac{24.2}{16} = 1.5125 \text{ mol atoms}$$

Divide by smallest (1.5125):

$$\frac{n_x}{1.5125} = \frac{3.158}{1.5125} \approx 2.09 \approx 2, \quad \frac{n_y}{1.5125} = 1$$

Ratio x : y = 2 : 1. **Empirical formula = x_2y**

Approach / Analogy

Step-by-step: moles of x = $75.8/24 \approx 3.16$, moles of y = $24.2/16 \approx 1.51$. Divide both by the smaller (1.51): x gives ≈ 2 , y gives 1. Ratio 2:1. EF = x_2y . The element with atomic weight 24 is Mg and 16 is O, so this is Mg_2O — but we use x and y as given.

Common Mistake

Dividing the percentages directly by the atomic weights and getting $75.8/24 = 3.16$ and $24.2/16 = 1.51$ but then forgetting to divide by the **smallest** (1.51). Students sometimes report $3:1.5 = 2:1$ directly, which accidentally works here, but the proper step is always “divide by smallest.”

Answer

(1) X_2Y

Q.4 Two elements X (Atomic weight = 75) and Y (Atomic weight = 16) combine to give a compound having 75.8% of X. The empirical formula of compound is

Explanation

% Y = $100 - 75.8 = 24.2\%$

In 100 g: 75.8 g X, 24.2 g Y.

$$n_X = \frac{75.8}{75} = 1.0107$$

$$n_Y = \frac{24.2}{16} = 1.5125$$

Divide by smallest (1.0107):

$$X : \frac{1.0107}{1.0107} = 1, \quad Y : \frac{1.5125}{1.0107} \approx 1.497 \approx 1.5$$

Multiply by 2 to clear the fraction: X:Y = 2:3.

Empirical formula = X_2Y_3

Approach / Analogy

After dividing by the smallest, Y gives ≈ 1.5 (not a whole number). When you get a ratio ending in 0.5, multiply **all values** by 2 to get whole numbers: $1 \times 2 = 2$ for X, $1.5 \times 2 = 3$ for Y. So X:Y = 2:3. This is a classic “multiply to clear” step.

Common Mistake

Rounding 1.5 down to 1 and writing X_2Y instead of X_2Y_3 . **Never round 1.5 to 1.** If the ratio is 1.5, multiply everything by 2. Only ratios very close to whole numbers (within ± 0.05) can be rounded — 1.5 is exactly between 1 and 2 and must be handled by multiplying by 2.

Answer

(4) X_2Y_3

Q.5 In a compound, element A (Atomic weight = 12.5) is 25% and element B (Atomic weight = 37.5) is 75% by weight. The empirical formula of the compound is

Explanation

$$n_A = \frac{25}{12.5} = 2.0$$
$$n_B = \frac{75}{37.5} = 2.0$$

Divide by smallest (2.0): A:B = 1:1.

Empirical formula = AB

Approach / Analogy

Equal moles despite different percentages and different atomic weights — because both ratios simplify to 2. When two elements are both present in whole-number mole ratios equal to each other, the empirical formula is simply AB (1:1 ratio).

Common Mistake

Writing A_2B_2 (option 3) since both mole values = 2. But A_2B_2 simplifies to AB (divide by 2). Empirical formula must be in the **simplest** ratio. A_2B_2 has HCF = 2, so it reduces to AB. Always reduce the final ratio to its lowest terms.

Answer

(1) AB

Q.6 A compound of X and Y has equal mass of them. If their atomic weights are 30 and 20 respectively. Molecular formula of the compound is

Explanation

Equal mass \Rightarrow assume 50 g each (or any equal amount, say 1 g each).

$$n_X = \frac{50}{30} = \frac{5}{3} = 1.667$$
$$n_Y = \frac{50}{20} = 2.5$$

Divide by smallest (1.667):

$$X : \frac{1.667}{1.667} = 1, \quad Y : \frac{2.5}{1.667} = 1.5$$

Multiply by 2: X:Y = 2:3.

Empirical formula = X_2Y_3

The question says “Molecular formula” — since no molar mass information is given to find the multiplier n , the molecular formula = empirical formula here = X_2Y_3 .

Approach / Analogy

“Equal mass” means use the same number for both (say 1 g or 50 g). With 50 g each: X gives $50/30 = 5/3$ mol, Y gives $50/20 = 5/2$ mol. Ratio = $5/3 : 5/2 = 2:3$ (cross-multiply). EF = X_2Y_3 . Shortcut: for equal masses, ratio of moles = inverse ratio of atomic weights = $20:30 = 2:3$ (flip the atomic weights!).

Common Mistake

Assuming equal masses means equal moles, writing X:Y = 1:1. Equal mass \neq equal moles when atomic weights differ. Heavier atoms (X, at.wt. 30) give fewer moles than lighter atoms (Y, at.wt. 20) from the same mass.

Answer

(3) X₂Y₃

Q.7 The simplest formula of a compound containing 50% of element X (at. wt. = 10) and 50% of element Y (at. wt. = 20) is

Explanation

$$n_X = \frac{50}{10} = 5$$

$$n_Y = \frac{50}{20} = 2.5$$

Divide by smallest (2.5):

$$X : \frac{5}{2.5} = 2, \quad Y : \frac{2.5}{2.5} = 1$$

X:Y = 2:1. **Empirical formula = X₂Y**

Approach / Analogy

Equal % but different atomic weights \rightarrow different moles. Lighter X (at.wt. 10) gives 5 mol vs Y (at.wt. 20) gives 2.5 mol. Ratio 5:2.5 = 2:1. EF = X₂Y. Quick shortcut: for equal masses, mole ratio = inverse of atomic weight ratio = 20:10 = 2:1.

Common Mistake

Picking XY (1:1 ratio) thinking equal percentages must mean equal atoms. Equal mass fractions don't mean equal mole fractions unless atomic weights are equal. Since X is lighter (10 vs 20), the same mass contains twice as many X atoms.

Answer

(2) X₂Y

Q.8 A hydrocarbon contains 80% C. The vapour density of compound is 30. Empirical formula of compound is

Explanation

$\% \text{ H} = 100 - 80 = 20\%$.

Empirical formula (from %):

$$n_{\text{C}} = \frac{80}{12} = 6.67, \quad n_{\text{H}} = \frac{20}{1} = 20$$

Divide by smallest (6.67):

$$\text{C} : 1, \quad \text{H} : \frac{20}{6.67} = 3$$

EF = CH₃, EF mass = 12 + 3 = 15.

Molecular formula (from VD):

$$M_r = 2 \times \text{VD} = 2 \times 30 = 60$$

$$n = \frac{M_r}{\text{EF mass}} = \frac{60}{15} = 4$$

MF = (CH₃)₄ = C₄H₁₂

But the question asks for **empirical formula = CH₃**.

Approach / Analogy

Two-part question: (1) find EF from percentages (CH₃), (2) find MF from VD (C₄H₁₂). The question asks for EF, which is CH₃. The VD info is extra/confirmatory here. CH₃ itself is not a stable molecule (radical), but it is the empirical formula.

Common Mistake

Answering C₄H₁₂ (the molecular formula) when the question asks for the **empirical formula**. Read carefully. EF = CH₃ (simplest ratio), MF = C₄H₁₂ (actual formula). These are different answers to different questions.

Answer

(1) CH₃

Q.9 A hydrocarbon contains 80% of carbon, then the hydrocarbon is

Explanation

Same % as Q.8: EF = CH₃, EF mass = 15.

Check each option's carbon percentage:

- CH₄: $M_r = 16$. $\% \text{C} = 12/16 \times 100 = 75\%$. ×
- C₂H₄: $M_r = 28$. $\% \text{C} = 24/28 \times 100 = 85.7\%$. ×
- C₂H₆: $M_r = 30$. $\% \text{C} = 24/30 \times 100 = 80\%$. ✓
- C₂H₂: $M_r = 26$. $\% \text{C} = 24/26 \times 100 = 92.3\%$. ×

C₂H₆ has 80% carbon. Also verify: C₂H₆ has EF = CH₃ (divide by 2), consistent with EF from Q.8.

Approach / Analogy

Quick method: compute %C for each option and match to 80%. Or use the EF = CH₃ from Q.8 — which MF gives CH₃ as EF? C₂H₆ ÷ 2 = CH₃. Check: C₂H₆ has %C = 24/30 = 80%. Both approaches confirm C₂H₆.

Common Mistake

Picking CH₄ (methane) by instinct since it's the simplest hydrocarbon. But CH₄ has 75% C, not 80%. Always calculate %C for each candidate rather than guessing.

Answer

(3) C₂H₆

Q.10 An oxide of sulphur contains 50% of sulphur in it. Its empirical formula is

Explanation

% O = 100 – 50 = 50%.

$$n_S = \frac{50}{32} = 1.5625$$

$$n_O = \frac{50}{16} = 3.125$$

Divide by smallest (1.5625):

$$S : 1, \quad O : \frac{3.125}{1.5625} = 2$$

Empirical formula = SO₂

Approach / Analogy

50% S and 50% O by mass. Note: atomic weight of S (32) is double that of O (16). Equal masses of S and O → moles of O = 2 × moles of S. So S:O = 1:2. EF = SO₂. Real compound: sulphur dioxide, a gas responsible for acid rain.

Common Mistake

Writing S:O = 1:1 (SO) because both are 50% by mass. 50/50 by mass is NOT 1:1 by moles when atomic weights differ. S weighs 32 and O weighs 16 per atom, so equal masses give twice as many O atoms as S atoms.

Answer

(1) SO₂

Q.11 An oxide of metal M has 40% by mass of oxygen. Metal M has atomic mass of 24. The empirical formula of the oxide is

Explanation

$$\% M = 100 - 40 = 60\%.$$

$$n_M = \frac{60}{24} = 2.5$$

$$n_O = \frac{40}{16} = 2.5$$

Divide by smallest (2.5): M:O = 1:1.

Empirical formula = MO

The metal M with at. mass 24 is Magnesium (Mg), and MgO is the well-known magnesium oxide. Consistent!

Approach / Analogy

Equal moles of M and O \rightarrow 1:1 ratio \rightarrow MO. This is the empirical formula of MgO (magnesium oxide), a white powder used in antacids and refractory materials. The equal moles arise because 60% Mg (at.wt. 24) and 40% O (at.wt. 16) give $60/24 = 2.5$ each.

Common Mistake

Writing M_2O_3 guessing from common metal oxides. Don't assume the formula based on common compounds — always calculate. Here both divide to 2.5 \rightarrow ratio is 1:1 \rightarrow MO, not M_2O_3 .

Answer

(3) MO

Q.12 A compound contains 38.8% C, 16.0% H and 45.2% N. The formula of the compound would be

Explanation

$$n_C = \frac{38.8}{12} = 3.233$$

$$n_H = \frac{16.0}{1} = 16.0$$

$$n_N = \frac{45.2}{14} = 3.229$$

Divide by smallest (≈ 3.23):

$$C : 1, \quad H : \frac{16.0}{3.23} \approx 4.95 \approx 5, \quad N : 1$$

Empirical formula = CH_5N

EF mass = $12 + 5 + 14 = 31$. Check options: CH_5N has $M_r = 31$, consistent with methylamine (CH_3NH_2 , also written CH_5N).

Approach / Analogy

C:H:N \approx 3.23:16:3.23. Divide by 3.23 \rightarrow 1:4.95:1 \approx 1:5:1. Formula = CH₅N = methylamine (CH₃NH₂). Note that CH₃NH₂ rearranged = CH₅N. The H ratio is 4.95 \approx 5, so round to 5 (within ± 0.05 of a whole number).

Common Mistake

Rounding H = 4.95 down to 4 (to give CH₄N) instead of up to 5. Values within ± 0.05 of an integer should be rounded to that integer. 4.95 is clearly close to 5, not 4 — round up to 5. CH₄N is not even an option.

Answer

(2) CH₃NH₂ (written as CH₅N)

Note

CH₃NH₂ (methylamine) and CH₅N are the same compound written differently. Both represent C:H:N = 1:5:1.

Q.13 An organic compound contains C, H and O. Its elemental analysis gives 38.71% of C and 9.67% of H. The empirical formula of the compound would be

Explanation

% O = 100 – 38.71 – 9.67 = 51.62%

$$n_C = \frac{38.71}{12} = 3.226$$

$$n_H = \frac{9.67}{1} = 9.67$$

$$n_O = \frac{51.62}{16} = 3.226$$

Divide by smallest (3.226):

$$C : 1, \quad H : \frac{9.67}{3.226} \approx 3, \quad O : 1$$

Empirical formula = CH₃O

Approach / Analogy

Always compute %O by subtraction: 100 – %C – %H = %O. Here: 100 – 38.71 – 9.67 = 51.62% O. Three equal mole values for C and O (3.226 each), and H is exactly 3 times that. EF = CH₃O.

Common Mistake

Forgetting to calculate %O by subtraction. In organic compounds with C, H, and O, if only C and H percentages are given, %O = 100 – %C – %H. Students sometimes try to use %O = %C + %H or leave O out entirely.

Answer

(3) CH₃O

Q.14 An organic compound containing C, H and N gave the following analysis C = 40%, H = 13.33%, N = 46.67%. Its empirical formula would be

Explanation

$$n_C = \frac{40}{12} = 3.33$$

$$n_H = \frac{13.33}{1} = 13.33$$

$$n_N = \frac{46.67}{14} = 3.33$$

Divide by smallest (3.33):

$$C : 1, \quad H : \frac{13.33}{3.33} = 4, \quad N : 1$$

Empirical formula = CH₄N

Approach / Analogy

C:H:N = 3.33:13.33:3.33. Divide by 3.33 → 1:4:1. EF = CH₄N. C and N are equal, H is exactly 4 times them. Check: %C = $12 / (12 + 4 + 14) \times 100 = 12/30 \times 100 = 40\%$. ✓

Common Mistake

Confusing CH₄N (EF mass 30) with CH₅N from Q.12 (EF mass 31). These have different compositions. Here H:N = 4:1, not 5:1. Carefully re-read the analysis percentages for each new problem.

Answer

(1) CH₄N

Q.15 An organic compound is composed of 4.58% H, 40.92% C and 54.50% oxygen. What would be the empirical formula of the compound?

Explanation

$$n_C = \frac{40.92}{12} = 3.41$$

$$n_H = \frac{4.58}{1} = 4.58$$

$$n_O = \frac{54.50}{16} = 3.406$$

Divide by smallest (3.406):

$$C : \frac{3.41}{3.406} \approx 1, \quad H : \frac{4.58}{3.406} \approx 1.344 \approx \frac{4}{3}, \quad O : 1$$

Since H gives $\approx 1.33 = 4/3$, multiply all by 3:

$$C : 3, \quad H : 4, \quad O : 3$$

Empirical formula = C₃H₄O₃

Approach / Analogy

H ratio = 1.344 $\approx 4/3$. When you get fractions like 1.33 (= 4/3), multiply all ratios by 3 to clear the fraction: C: $1 \times 3 = 3$, H: $1.33 \times 3 = 4$, O: $1 \times 3 = 3$. EF = C₃H₄O₃.

Common Mistake

Rounding 1.344 to 1 and writing CHO instead of C₃H₄O₃. A ratio of 1.33 cannot be rounded to 1 (the error is 25%). When the ratio is ≈ 1.33 or ≈ 1.67 , you must multiply by 3. Use: 1.25 $\rightarrow \times 4$, 1.33 $\rightarrow \times 3$, 1.5 $\rightarrow \times 2$, 1.67 $\rightarrow \times 3$.

Answer

(1) C₃H₄O₃

Q.16 An organic compound containing C and H gave the following analysis C = 40%, H = 6.7%. Its empirical formula would be

Explanation

$$\% O = 100 - 40 - 6.7 = 53.3\%$$

$$n_C = \frac{40}{12} = 3.33$$

$$n_H = \frac{6.7}{1} = 6.7$$

$$n_O = \frac{53.3}{16} = 3.33$$

Divide by smallest (3.33):

$$C : 1, \quad H : \frac{6.7}{3.33} \approx 2, \quad O : 1$$

Empirical formula = CH₂O

Approach / Analogy

The question says “C and H” but the percentages only add to 46.7% — there must be a third element. $100 - 40 - 6.7 = 53.3\%$ which must be oxygen. Always check if percentages sum to 100; if not, the balance is the unstated element (usually O in organic compounds).

Common Mistake

Treating the compound as containing only C and H (since the question says “containing C and H”), and computing only n_C and n_H : 3.33 and 6.7 \rightarrow ratio 1:2 \rightarrow CH₂. But the percentages only sum to 46.7% \neq 100%, so oxygen must be present. Always check the sum of given percentages first.

Answer

(2) CH₂O

Q.17 A gas is found to contain 2.34 g of Nitrogen and 5.34 g of oxygen. Simplest formula of the compound is

Explanation

$$n_N = \frac{2.34}{14} = 0.1671$$

$$n_O = \frac{5.34}{16} = 0.33375$$

Divide by smallest (0.1671):

$$N : 1, \quad O : \frac{0.33375}{0.1671} \approx 2$$

Empirical formula = NO₂

Approach / Analogy

When masses (not percentages) are given, the method is the same — just divide each mass by its atomic weight to get moles, then find the simplest ratio. Here: moles of O = twice moles of N \rightarrow NO₂. This is nitrogen dioxide (the brownish gas in polluted air).

Common Mistake

Using the given masses as percentages (writing 2.34% N and 5.34% O). The question gives masses in grams, not percentages. But the method is identical: divide mass by atomic weight to get moles, then find the ratio. The total doesn't need to be 100 g.

Answer

(4) NO₂

Q.18 2.2 g of a compound of phosphorous and sulphur has 1.24 g of P in it. Its empirical formula is

Explanation

Mass of S = 2.2 – 1.24 = 0.96 g.

$$n_P = \frac{1.24}{31} = 0.04$$

$$n_S = \frac{0.96}{32} = 0.03$$

Divide by smallest (0.03):

$$P : \frac{0.04}{0.03} = 1.33 = \frac{4}{3}, \quad S : 1$$

Multiply by 3: P:S = 4:3.

Empirical formula = P₄S₃

Approach / Analogy

Find mass of S by subtraction: 2.2 – 1.24 = 0.96 g. Then $n_P = 1.24/31 = 0.04$ and $n_S = 0.96/32 = 0.03$. Ratio P:S = 0.04:0.03 = 4:3 (multiply both by 100/1 then simplify, or divide by 0.03 to get 1.33:1 then multiply by 3). P₄S₃ is a real compound (white phosphorus + sulphur).

Common Mistake

Forgetting to find mass of S by subtraction first. The total compound mass (2.2 g) minus P mass (1.24 g) = S mass (0.96 g). Without this step, there's no way to get n_S .

Answer

(4) P₄S₃

Q.19 On analysis, a certain compound was found to contain iodine and oxygen in the ratio of 254:80. The formula of the compound is: (At. mass I = 127, O = 16) [NSEC]

Explanation

Mass ratio I:O = 254:80.

$$n_I = \frac{254}{127} = 2$$

$$n_O = \frac{80}{16} = 5$$

Ratio I:O = 2:5.

Empirical formula = I₂O₅

I₂O₅ (iodine pentoxide) is a real compound, used as an oxidising agent.

Approach / Analogy

Mass ratio given directly (not percentages) — same method. 254/127 = 2 mol I, 80/16 = 5 mol O. Ratio 2:5. EF = I₂O₅. Notice 254 = 2 × 127 (atomic mass of I) — the numbers are designed to give clean mole values!

Common Mistake

Dividing the mass ratio numbers by each other directly: $254/80 \approx 3.175$ and trying to interpret this as an atom ratio. Always divide **each mass by its own atomic weight** before comparing.

Answer

(4) I_2O_5

Q.20 The number of atoms of Cr and O are 4.8×10^{10} and 9.6×10^{10} respectively. Its empirical formula is

Explanation

Number of atoms is directly proportional to moles (both divided by N_A).

Ratio of atoms = ratio of moles:

$$\text{Cr} : \text{O} = 4.8 \times 10^{10} : 9.6 \times 10^{10} = 4.8 : 9.6 = 1 : 2$$

Empirical formula = CrO_2

Approach / Analogy

When the number of *atoms* (not masses) is given, the ratio of atoms IS the ratio of moles directly (no need to divide by atomic weight). Simply simplify the atom count ratio: $4.8 : 9.6 = 1 : 2$. EF = CrO_2 . This is chromium(IV) oxide.

Common Mistake

Dividing the atom counts by atomic weights of Cr (52) and O (16). This would be wrong! Atom count ratios directly give mole ratios without any division by atomic weight. Dividing by atomic weights is only needed when going from **masses** to moles, not from atom counts.

Answer

(2) CrO_2

End of Part 1 — Q.1 to Q.20 (TYPE 1)

Part 2 covers TYPE 2: Molecular Formula from Empirical Formula (Q.21–Q.32)

TYPE 2 : Molecular Formula from Empirical Formula

Key Relationship

$$\text{Molecular Formula} = n \times \text{Empirical Formula}, \quad \text{where } n = \frac{M_r(\text{compound})}{\text{EF mass}}$$

n must be a positive integer. Find EF first, calculate EF mass, then divide M_r by EF mass to get n .

When VD is given: $M_r = 2 \times \text{VD}$

When 1 L of gas has same mass as 1 L of another gas (same T, P): both gases have the same M_r .

Q.21 Empirical formula of a compound is CH₂O. The molecular formula of the compound cannot be

Explanation

EF = CH₂O (EF mass = 30). MF must be (CH₂O)_n for some integer $n \geq 1$.

Check each option:

- CH₂O: $n = 1$. ✓ (formaldehyde)
- C₆H₁₂O₆: $n = 6$. $\div 6 = \text{CH}_2\text{O}$. ✓ (glucose)
- C₁₂H₂₂O₁₁: $M_r = 342$. $342/30 = 11.4$ — not an integer! ×
- CH₃COOH = C₂H₄O₂: $\div 2 = \text{CH}_2\text{O}$. ✓ (acetic acid)

C₁₂H₂₂O₁₁ (sucrose) cannot have the EF CH₂O.

Approach / Analogy

The MF must be an exact multiple of the EF. Divide M_r of each option by EF mass (30) — if the result is a whole number, it's possible. C₁₂H₂₂O₁₁ has $M_r = 342$: $342/30 = 11.4$, not whole. Alternatively, check the C:H:O ratio in sucrose: 12:22:11 — this cannot be reduced to 1:2:1 since HCF(12,22,11) = 1.

Common Mistake

Picking CH₂O itself (option 1) thinking n must be > 1 . $n = 1$ is perfectly valid — the molecular formula can equal the empirical formula. The EF is just the simplest form; if the compound happens to have that exact simple formula, MF = EF.

Answer

(3) C₁₂H₂₂O₁₁

Q.22 What is the simplest formula of the compound which has Carbon 80%, Hydrogen 20%? If the molecular mass is 30, the molecular formula is

Explanation

Empirical formula:

$$n_C = \frac{80}{12} = 6.67, \quad n_H = \frac{20}{1} = 20$$

Divide by smallest (6.67): C:H = 1:3. **EF = CH₃** (EF mass = 15).

Molecular formula:

$$n = \frac{30}{15} = 2$$

MF = (CH₃)₂ = C₂H₆

Approach / Analogy

EF = CH₃, $M_r = 30 = 2 \times \text{EF mass (15)}$. So MF = C₂H₆ (ethane). The question asks for both: simplest formula (CH₃) and molecular formula (C₂H₆). Both given in option (1).

Common Mistake

Computing EF mass as 16 instead of 15 (using H = 2 instead of H = 1). Hydrogen's atomic mass = 1 g/mol, not 2. EF mass of CH₃ = 12 + 3(1) = 15. Using H = 2 gives EF mass = 18, then $n = 30/18$ which is not an integer.

Answer

(1) CH₃ and C₂H₆

Q.23 In a compound Carbon is 52.2%, Hydrogen is 13%, Oxygen is 34.8% and molecular mass of the compound is 92. The molecular formula of the compound is

Explanation

Empirical formula:

$$n_C = \frac{52.2}{12} = 4.35, \quad n_H = \frac{13}{1} = 13, \quad n_O = \frac{34.8}{16} = 2.175$$

Divide by smallest (2.175):

$$C : \frac{4.35}{2.175} = 2, \quad H : \frac{13}{2.175} \approx 5.98 \approx 6, \quad O : 1$$

EF = C₂H₆O (EF mass = 24+6+16 = 46).

Molecular formula:

$$n = \frac{92}{46} = 2$$

MF = (C₂H₆O)₂ = C₄H₁₂O₂

Approach / Analogy

EF = C₂H₆O (EF mass 46). With $M_r = 92$, $n = 2$. MF = C₄H₁₂O₂. This could represent diethyl ether or butanediol. The H ratio $5.98 \approx 6$ — round to 6 (within 0.05 of a whole number).

Common Mistake

Stopping at the EF (C₂H₆O) and choosing option (1) C₂H₆O without using the given molecular mass. The question provides $M_r = 92$ specifically to find the MF, not just the EF. Always use M_r to calculate n and find the actual molecular formula.

Answer

(2) C₄H₁₂O₂

Q.24 An organic compound on analysis was found to contain 71.7% of chlorine, 4.04% of hydrogen and rest is carbon. If its molecular weight is 99, calculate molecular formula

Explanation

$$\% \text{ C} = 100 - 71.7 - 4.04 = 24.26\%$$

Empirical formula:

$$n_{\text{C}} = \frac{24.26}{12} = 2.022 \approx 2$$

$$n_{\text{H}} = \frac{4.04}{1} = 4.04 \approx 4$$

$$n_{\text{Cl}} = \frac{71.7}{35.5} = 2.019 \approx 2$$

Divide by smallest (2): C:H:Cl = 1:2:1.

EF = CH₂Cl (EF mass = 12 + 2 + 35.5 = 49.5).

Molecular formula:

$$n = \frac{99}{49.5} = 2$$

MF = (CH₂Cl)₂ = C₂H₄Cl₂

Approach / Analogy

Three steps: (1) find %C by subtraction, (2) compute EF (CH₂Cl, EF mass 49.5), (3) $n = 99/49.5 = 2$ → MF = C₂H₄Cl₂. Note Cl atomic mass = 35.5. C₂H₄Cl₂ is dichloroethane, an industrial solvent.

Common Mistake

Using Cl atomic mass as 35 instead of 35.5. Chlorine's standard atomic mass is 35.5 (average of isotopes ³⁵Cl and ³⁷Cl). Using 35 gives $n_{\text{Cl}} = 71.7/35 = 2.05$ which still rounds to 2, so the error is small here — but always use 35.5 for Cl.

Answer

(2) C₂H₄Cl₂

Q.25 A compound on analysis gave: C = 54.54%, H = 9.09% and vapour density of the compound = 88. The molecular formula of the compound is

Explanation

$$\% \text{ O} = 100 - 54.54 - 9.09 = 36.37\%$$

Empirical formula:

$$n_{\text{C}} = \frac{54.54}{12} = 4.545$$

$$n_{\text{H}} = \frac{9.09}{1} = 9.09$$

$$n_{\text{O}} = \frac{36.37}{16} = 2.273$$

Divide by smallest (2.273):

$$\text{C} : 2, \quad \text{H} : \frac{9.09}{2.273} \approx 4, \quad \text{O} : 1$$

EF = C₂H₄O (EF mass = 24+4+16 = 44).

Molecular formula:

$$M_r = 2 \times \text{VD} = 2 \times 88 = 176$$

$$n = \frac{176}{44} = 4$$

MF = (C₂H₄O)₄ = C₈H₁₆O₄

Approach / Analogy

EF = C₂H₄O (EF mass 44). VD = 88 → $M_r = 176$. $n = 176/44 = 4$. MF = C₈H₁₆O₄. Notice EF mass 44 = exactly $M_r(\text{CO}_2)$ — a coincidence, but always verify by checking percentages.

Common Mistake

Not converting VD to M_r and directly using 88 as molar mass: $n = 88/44 = 2 \rightarrow \text{C}_4\text{H}_8\text{O}_2$. Remember: $\text{VD} = M_r/2$, so $M_r = 2 \times 88 = 176$, not 88. The answer C₄H₈O₂ (option 2) results from this exact mistake.

Answer

(3) C₈H₁₆O₄

Q.26 Percentage composition of an organic compound is: C = 10.06, H = 0.84, Cl = 89.10. Which of the following corresponds to its molecular formula if the vapour density is 60.0?

Explanation

Empirical formula:

$$n_C = \frac{10.06}{12} = 0.8383$$

$$n_H = \frac{0.84}{1} = 0.84$$

$$n_{Cl} = \frac{89.10}{35.5} = 2.51$$

Divide by smallest (0.8383):

$$C : 1, \quad H : \frac{0.84}{0.8383} \approx 1, \quad Cl : \frac{2.51}{0.8383} \approx 2.99 \approx 3$$

EF = CHCl₃ (EF mass = 12 + 1 + 3 × 35.5 = 13 + 106.5 = 119.5).

Molecular formula:

$$M_r = 2 \times 60 = 120$$

$$n = \frac{120}{119.5} \approx 1$$

MF = CHCl₃ (chloroform)

Approach / Analogy

EF = CHCl₃ (EF mass ≈ 119.5). $M_r = 120$. $n \approx 1$. MF = CHCl₃. This is chloroform (trichloromethane), historically used as an anaesthetic! The close agreement of EF mass (119.5) and M_r (120) confirms $n = 1$.

Common Mistake

Miscalculating EF mass of CHCl₃ as 12 + 1 + 35.5 = 48.5 (counting only one Cl instead of three). CHCl₃ has 3 Cl atoms: 12 + 1 + 3(35.5) = 12 + 1 + 106.5 = 119.5. Always multiply atomic mass by subscript.

Answer

(2) CHCl₃

Q.27 1 L of a hydrocarbon weighs as much as 1 L of CO₂ under similar conditions. The molecular formula of the hydrocarbon is

Explanation

Same volume, same conditions \Rightarrow same moles \Rightarrow same molar mass.

$$M_r(\text{hydrocarbon}) = M_r(\text{CO}_2) = 44 \text{ g/mol}$$

Now find a hydrocarbon with $M_r = 44$:

- C_3H_8 : $3(12) + 8 = 44$. ✓ (propane)
- C_2H_6 : $24 + 6 = 30$. ✗
- C_2H_4 : $24 + 4 = 28$. ✗
- C_3H_6 : $36 + 6 = 42$. ✗

Molecular formula = C_3H_8

Approach / Analogy

Same T and P: equal volumes = equal moles (Avogadro's law). Equal mass means same molar mass. Hydrocarbon $M_r = M_r(\text{CO}_2) = 44$. Only C_3H_8 has $M_r = 44$ among the options. CO_2 is a useful mass reference because its $M_r = 44$ is exactly that of propane!

Common Mistake

Thinking "same volume" means same mass and then comparing volumes instead of molar masses. The logic is: same V + same T&P + same mass \Rightarrow same molar mass (from ideal gas law: $PV = nRT$, so $n = PV/RT$ is fixed for equal V, T, P; fixed mass per fixed moles = same M_r).

Answer

(1) C_3H_8

Q.28 The empirical formula of an organic compound containing carbon and hydrogen is CH_2 . The mass of one litre of this organic gas is exactly equal to that of one litre of N_2 at same temperature and pressure. Therefore, the molecular formula of the organic gas is

Explanation

Same T, P, V \Rightarrow same moles \Rightarrow same mass per mole \Rightarrow same M_r .

$$M_r(\text{hydrocarbon}) = M_r(\text{N}_2) = 28 \text{ g/mol}$$

EF = CH_2 , EF mass = 14.

$$n = \frac{28}{14} = 2$$

MF = $(\text{CH}_2)_2 = \text{C}_2\text{H}_4$ (ethylene)

Approach / Analogy

EF = CH_2 (EF mass 14). Hydrocarbon $M_r = 28$ (same as N_2). $n = 28/14 = 2$. MF = C_2H_4 . Ethylene (C_2H_4) is the most important industrial organic chemical — used to make polyethylene plastic!

Common Mistake

Using N's atomic mass (14) instead of N₂'s molar mass (28). The gas is N₂ (diatomic nitrogen), $M_r = 2 \times 14 = 28$. Using 14 gives $n = 14/14 = 1$ and MF = CH₂ — but CH₂ (carbene) is a reactive intermediate, not a stable gas. Always use the molecular mass of the gas (N₂ = 28), not the atomic mass of N (14).

Answer

(1) C₂H₄

Q.29 A metal chloride contains 55.0% of chlorine by weight. 100 mL vapours of the metal chloride at STP weigh 0.57 g. The molecular formula of the metal chloride is [JEE 2023]

Explanation

Step 1: Molar mass of metal chloride.

$$n = \frac{100 \text{ mL}}{22400 \text{ mL/mol}} = \frac{1}{224} \text{ mol}$$

$$M_r = \frac{0.57 \text{ g}}{1/224} = 0.57 \times 224 = 127.68 \approx 128 \text{ g/mol}$$

Step 2: Empirical formula.

% M = 100 - 55 = 45%. Let the chloride be MCl_x.

$$\text{Mass of Cl in formula} = 0.55 \times 128 = 70.4 \text{ g} \approx 70 \text{ g}$$

$$x = \frac{70.4}{35.5} \approx 1.98 \approx 2$$

So 2 Cl atoms per formula unit.

$$\text{Mass of M} = 128 - 2(35.5) = 128 - 71 = 57 \text{ g/mol}$$

Metal M has atomic mass 57. This corresponds to Fe (atomic mass ≈ 56) or similar. Formula = MCl₂.

Approach / Analogy

Two approaches needed: (1) find M_r from gas density data (volume + mass at STP), (2) find number of Cl atoms using %Cl and M_r . 100 mL at STP = 100/22400 mol. Mass = 0.57 g → $M_r = 0.57 \times 224 \approx 128$. %Cl = 55% → Cl mass = 70.4 g → 2 Cl atoms (71/35.5 = 2). Formula = MCl₂.

Common Mistake

Using 22.4 L/mol (litres) but dividing 100 mL into it without unit conversion: writing $n = 100/22.4$ instead of 100/22400. 100 mL = 0.1 L, so $n = 0.1/22.4$ OR use mL directly: $n = 100/22400$. Both give the same answer but the wrong approach gives $n = 4.46 \text{ mol}$ — 1000 times too large.

Answer

(3) MCl₂

Q.30 Number of hydrogen atoms per molecule of a hydrocarbon A having 85.8% carbon is. (Given:

Molar mass of A = 84 g mol⁻¹)

Explanation

% H = 100 – 85.8 = 14.2%

Empirical formula:

$$n_C = \frac{85.8}{12} = 7.15$$

$$n_H = \frac{14.2}{1} = 14.2$$

Divide by smallest (7.15): C:H = 1: $\frac{14.2}{7.15}$ = 1:1.986 \approx 1:2.

EF = CH₂ (EF mass = 14).

Molecular formula:

$$n = \frac{84}{14} = 6 \implies \text{MF} = \text{C}_6\text{H}_{12}$$

Number of H atoms = **12**.

Approach / Analogy

EF = CH₂ (EF mass 14). $M_r = 84$. $n = 84/14 = 6$. MF = C₆H₁₂ (cyclohexane or hexene). H atoms per molecule = 12. The question asks for H atoms, not H moles — the answer is 12 atoms per molecule.

Common Mistake

Answering 6 (the value of n) instead of 12 (the number of H atoms). $n = 6$ is the multiplier used to go from EF to MF — it tells you the formula is CH₂ \times 6 = C₆H₁₂. The actual number of H atoms per molecule = subscript of H in MF = 12.

Answer

(1) 12

Q.31 A gaseous hydrocarbon on combustion gives 0.72 g of water and 3.08 g of CO₂. The empirical formula of the hydrocarbon is

Explanation

Find moles of C and H from combustion products:

$$n(\text{CO}_2) = \frac{3.08}{44} = 0.07 \text{ mol} \implies n_C = 0.07 \text{ mol}$$

$$n(\text{H}_2\text{O}) = \frac{0.72}{18} = 0.04 \text{ mol} \implies n_H = 2 \times 0.04 = 0.08 \text{ mol}$$

Ratio C:H:

$$\frac{0.07}{0.07} : \frac{0.08}{0.07} = 1 : 1.143 \approx 1 : \frac{8}{7}$$

Multiply by 7: C:H = 7:8.

Empirical formula = C₇H₈ (toluene!)

Approach / Analogy

Combustion analysis method: every C in the hydrocarbon becomes 1 CO₂; every 2 H become 1 H₂O. So moles of C = moles of CO₂, and moles of H = 2 × moles of H₂O. Then find the C:H ratio and simplify. Here: C = 0.07, H = 0.08 → ratio = 7:8 → C₇H₈ (EF for toluene C₇H₈, which is already simplest).

Common Mistake

Writing $n_H = n(\text{H}_2\text{O}) = 0.04$ instead of $2 \times 0.04 = 0.08$. Each water molecule contains **2 H atoms**. The number of H atoms = 2 × moles of water. Forgetting this factor of 2 halves the H count and gives a wrong formula.

Answer

(4) C₇H₈

Q.32 An organic compound having molecular mass 60 is found to contain 20% of C, 6.67% of H and 46.67% of O, while rest is nitrogen. On heating it gives NH₃ along with a solid residue. The solid residue gives violet colour with alkaline copper sulphate solution. The compound is

Explanation

$$\% \text{ N} = 100 - 20 - 6.67 - 46.67 = 26.66\%$$

Empirical formula:

$$n_{\text{C}} = \frac{20}{12} = 1.667$$

$$n_{\text{H}} = \frac{6.67}{1} = 6.67$$

$$n_{\text{N}} = \frac{26.66}{14} = 1.904 \approx 1.9$$

$$n_{\text{O}} = \frac{46.67}{16} = 2.917$$

Divide by smallest (1.667):

$$\text{C} : 1, \quad \text{H} : 4, \quad \text{N} : \frac{1.9}{1.667} \approx 1.14 \approx 1, \quad \text{O} : \frac{2.917}{1.667} \approx 1.75 \approx \frac{7}{4}$$

The ratios are imprecise due to rounding in the given percentages. Try a direct approach using $M_r = 60$:

$$\text{Mass of C} = 0.20 \times 60 = 12 \text{ g} \implies n_{\text{C}} = 1$$

$$\text{Mass of H} = 0.0667 \times 60 = 4 \text{ g} \implies n_{\text{H}} = 4$$

$$\text{Mass of N} = 0.2666 \times 60 = 16 \text{ g} \implies n_{\text{N}} = \frac{16}{14} \approx 1.14$$

$$\text{Mass of O} = 0.4667 \times 60 = 28 \text{ g} \implies n_{\text{O}} = \frac{28}{16} = 1.75$$

These fractions suggest rounding errors. Check the option $(\text{NH}_2)_2\text{CO}$ (urea): $M_r = 60$, formula = $\text{CH}_4\text{N}_2\text{O}$.

$$\% \text{C} = 12/60 = 20\%, \quad \% \text{H} = 4/60 = 6.67\%, \quad \% \text{N} = 28/60 = 46.67\%, \quad \% \text{O} = 16/60 = 26.67\%.$$

Wait — %N and %O are swapped in the question. The compound is **urea** $(\text{NH}_2)_2\text{CO}$:

- Formula: $\text{CH}_4\text{N}_2\text{O}$, $M_r = 60$. ✓
- Gives NH_3 on heating. ✓
- Biuret test (violet with alkaline CuSO_4) confirms urea. ✓

Approach / Analogy

The clues identify urea uniquely: (1) $M_r = 60$, (2) gives NH_3 on heating, (3) biuret reaction gives violet colour with alkaline CuSO_4 . Even if the percentage calculation is tricky due to rounding, the chemical clues point directly to urea = $(\text{NH}_2)_2\text{CO}$.

Common Mistake

Getting confused by the apparently swapped %N and %O values in the printed question. Use the chemical identification clues (NH_3 on heating + biuret test + $M_r = 60$) as the primary identification method. Urea $(\text{NH}_2)_2\text{CO}$ is the only compound at $M_r = 60$ that fits all clues.

Answer

(1) $(\text{NH}_2)_2\text{CO}$ (Urea)

Note

There appears to be a swap in the given percentages for N and O in the original question (46.67% printed as O but actually matches N in urea, and vice versa). The compound is unambiguously urea based on: $M_r = 60$, production of NH_3 on heating, and the biuret test (violet colour with alkaline CuSO_4).

— End of DPP-6 Complete Solution Sheet —

Empirical Formula · Q.1–Q.32 · All Parts Complete

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