Weird Chemist

DPP–8 : Laws of Chemical Combination

Chapter: Some Basic Concepts of Chemistry

Solution Sheet

Quick Reference: Five Laws of Chemical Combination

1. Conservation of Mass (Lavoisier): Total mass of reactants = Total mass of products.

2. Definite Proportions / Constant Composition (Proust): A pure compound always contains the same elements in the same mass ratio, regardless of source or method of preparation.

3. Multiple Proportions (Dalton): When two elements form more than one compound, the masses of one element that combine with a fixed mass of the other are in a simple whole-number ratio.

4. Gay-Lussac's Law of Gaseous Volumes: Gases react in simple whole-number ratios by volume (at same T and P). Volume ratios = mole ratios.

5. Avogadro's Law: Equal volumes of all gases at the same T and P contain equal numbers of molecules.

TYPE 1 : Law of Conservation of Mass

Q.1 $BaCl_2$ reacts with 24.4 g of sodium sulphate to produce 46.6 g of barium sulphate and 23.4 g of sodium chloride. The weight of $BaCl_2$ that reacted is:

Explanation

By the Law of Conservation of Mass:

$$\text{Total mass of reactants} = \text{Total mass of products}$$

$$m(\text{BaCl}_2) + 24.4 = 46.6 + 23.4$$

$$m(\text{BaCl}_2) = 70.0 - 24.4 = \mathbf{45.6}\ \text{g}$$

Approach / Analogy

Think of a balance scale: everything put in must equal everything that comes out. No atoms are created or destroyed. Reactants in = Products out. So: $BaCl_2$ + 24.4 g = 46.6 g + 23.4 g. Solve for $BaCl_2$.

Common Mistake

Directly reading one of the product masses (46.6 g or 23.4 g) as the answer for $BaCl_2$. The question gives you both products' masses — you must add them (70.0 g total) and subtract the known reactant mass (24.4 g) to find $BaCl_2$ = 45.6 g.

Answer

(1) 45.6 g

Q.2 10 g of $CaCO_3$ on heating gives 4.4 g of CO_2. The weight of CaO produced in quintal is:

Explanation

By conservation of mass: mass of CaO = $10.0 - 4.4 = 5.6$ g.

Convert to quintal: 1 quintal = 100 kg = 10^5 g.

$$5.6 \text{ g} = \frac{5.6}{10^5} \text{ quintal} = \mathbf{5.6 \times 10^{-5}} \text{ quintal}$$

Approach / Analogy

Two steps: (1) use conservation of mass to find CaO in grams (5.6 g), (2) convert units. 1 quintal = 100 kg = 100,000 g = 10^5 g. So 5.6 g = $5.6/10^5 = 5.6 \times 10^{-5}$ quintal.

Common Mistake

Confusing the unit conversion: 1 quintal = 100 kg, not 1 kg or 10 kg. Students sometimes use 1 quintal = 1000 g (thinking of 1 kg) and get 5.6×10^{-3} quintal. Remember: 1 quintal = 100 kg = 10^5 g.

Answer

(1) 5.6×10^{-5} quintal

Q.3 10.0 g of $CaCO_3$ on heating gave 4.4 g of CO_2 and x g of CaO. Applying law of conservation of mass, the mass of CaO is

Explanation

$$10.0 = 4.4 + x \implies x = 10.0 - 4.4 = \mathbf{5.6} \text{ g}$$

Approach / Analogy

Direct application of conservation of mass: mass in = mass out. $CaCO_3$ (10 g) = CaO (x g) + CO_2 (4.4 g). So $x = 10 - 4.4 = 5.6$ g. Verify: $5.6 + 4.4 = 10.0$ g. ✓

Common Mistake

Answering 4.4 g (same as CO_2 mass) or 10.0 g (reactant mass). The CaO mass is neither — it's the *difference* between reactant mass (10 g) and CO_2 product mass (4.4 g).

Answer

(1) 5.6 g

Q.4 The law of conservation of mass holds good for all of the following except

Explanation

The law of conservation of mass holds for **chemical reactions** (both endothermic and exothermic). However, in **nuclear reactions** (fission, fusion), a small fraction of mass converts to energy via $E = mc^2$. Mass is NOT conserved in nuclear reactions — mass+energy is conserved instead.

Approach / Analogy

Chemical reactions: bonds break and form but atoms are conserved → mass conserved. Nuclear reactions: protons and neutrons can combine or split, and mass converts to huge amounts of energy. Think of nuclear bombs and nuclear power plants — the energy comes from mass itself being "destroyed."

Common Mistake

Picking endothermic or exothermic reactions as exceptions. In both types of chemical reactions, mass is perfectly conserved — energy is released or absorbed from chemical bonds, not from mass. Only nuclear reactions violate conservation of mass.

Answer

(2) Nuclear reactions

Q.5 A chemical equation is balanced according to the law of

Explanation

When we balance a chemical equation, we ensure the number of atoms of each element on the LHS equals the RHS. This directly implements the **Law of Conservation of Mass** — same atoms, same mass on both sides.

Approach / Analogy

Balancing an equation = making sure nothing appears or disappears during the reaction. You add coefficients to make atom counts equal on both sides. This is the mathematical expression of "mass in = mass out" = Law of Conservation of Mass.

Common Mistake

Choosing "Constant composition" (option 2). Constant composition tells us the ratio within a compound, not how to balance an equation. Balancing is about equal atoms on both sides of the arrow = Conservation of Mass.

Answer

(4) Conservation of mass

Q.6 28 g of N_2 and 6 g of H_2 react to give 34 g of NH_3. This statement illustrates the law of [NCERT Pg. 14]

Explanation

Check conservation of mass: Reactants = $28 + 6 = 34$ g. Products = 34 g. Reactants = Products. ✓

This directly illustrates that mass is neither created nor destroyed in a chemical reaction = **Law of Conservation of Mass.**

Approach / Analogy

The statement gives specific masses of reactants and products and shows they add up: $28 + 6 = 34$. This is a direct demonstration of mass balance. It does not involve two different compounds of the same elements (multiple proportions) or fixed composition (definite proportions).

Common Mistake

Picking "Definite proportion" thinking the fixed ratio of N:H means definite proportions. While NH_3 does have a definite composition, *this specific statement* is demonstrating mass balance ($28+6 = 34$), which is conservation of mass. The key is: what does the *given statement* demonstrate?

Answer

(1) Conservation of mass

TYPE 2 : Law of Definite Proportions (Constant Composition)

Q.7 Cu forms two oxides, cuprous and cupric oxides. Which law can be proved by the weights of Cu and O?

Explanation

Cu forms **two different oxides**: Cu_2O (cuprous oxide) and CuO (cupric oxide). These contain different mass ratios of Cu:O. When two elements form **more than one compound**, the masses of one element combining with a fixed mass of the other are in a simple whole-number ratio. This is the **Law of Multiple Proportions**.

Cu_2O: Cu:O = 128:16 = 8:1.
CuO: Cu:O = 64:16 = 4:1.
Ratio of O combining with fixed Cu (8 g): $1 : 2$ (simple whole-number ratio).

Approach / Analogy

The key phrase is "two oxides." When two elements form **multiple** compounds, it's the Law of Multiple Proportions. If there were only one oxide, it would illustrate Definite Proportions. Two different compounds of the same pair of elements = Multiple Proportions.

Common Mistake

Choosing "constant composition" (option 1) or "definite proportions" (option 4). These are the same law (Proust's law) and apply to a *single* pure compound having a fixed ratio. Since there are *two* different Cu oxides with different ratios, this illustrates Multiple Proportions (Dalton's law), not Definite Proportions.

Answer

(2) Multiple proportions

Q.8 Weight of copper oxide obtained by treating 2.16 g of metallic copper: 2.70 g. In another experiment, 1.15 g of copper oxide on reduction yielded 0.92 g of copper. The ratio of Cu:O in experiment I and experiment II respectively is

Explanation

Experiment I:
Mass of Cu = 2.16 g, Mass of CuO = 2.70 g $\implies$ Mass of O = $2.70 - 2.16 = 0.54$ g.

$$\text{Cu:O} = 2.16 : 0.54 = \frac{2.16}{0.54} : 1 = 4 : 1$$

Experiment II:
Mass of Cu = 0.92 g, Mass of CuO = 1.15 g $\implies$ Mass of O = $1.15 - 0.92 = 0.23$ g.

$$\text{Cu:O} = 0.92 : 0.23 = \frac{0.92}{0.23} : 1 = 4 : 1$$

Both give Cu:O = **4:1** — illustrating the Law of Definite Proportions (same compound, same ratio).

Approach / Analogy

Both experiments produce/use the same compound (CuO). By the Law of Definite Proportions, the mass ratio Cu:O must be the same regardless of method. Mass of O = mass of oxide − mass of Cu. In both cases: ratio = 4:1. This is verification of Proust's Law.

Common Mistake

Computing the ratio as Cu:CuO (metal to oxide) instead of Cu:O (metal to oxygen alone). The oxygen mass = oxide mass − metal mass. In Exp I: O = 2.70 − 2.16 = 0.54 g. In Exp II: O = 1.15 − 0.92 = 0.23 g. Always find the oxygen mass by subtraction before computing the ratio.

Answer

(1) 4:1 in both

Q.9 In an experiment 2.4 g of FeO on reduction with hydrogen gives 1.68 g of Fe. In another experiment 2.9 g of FeO gives 2.03 g of Fe on reduction with hydrogen. The ratio of Fe:O in both experiments is

Explanation

Experiment I:
Fe = 1.68 g, FeO = 2.4 g $\implies$ O = $2.4 - 1.68 = 0.72$ g.

$$\text{Fe:O} = \frac{1.68}{0.72} = 2.333\ldots : 1 \approx \frac{7}{3} : 1 = 7 : 3$$

Experiment II:
Fe = 2.03 g, FeO = 2.9 g $\implies$ O = $2.9 - 2.03 = 0.87$ g.

$$\text{Fe:O} = \frac{2.03}{0.87} = 2.333\ldots : 1 \approx 7 : 3$$

Both give Fe:O ≈ **2.33:1** (i.e., $7:3$) — same ratio in both experiments. Law of Definite Proportions.

Approach / Analogy

Same compound (FeO) in both experiments → same Fe:O ratio. O mass = FeO mass − Fe mass. Both give $\approx 2.33 : 1$ ($= 7 : 3$). This confirms Proust's Law: a pure compound always has the same mass composition.

Common Mistake

Reporting different ratios for both experiments (option 4) due to arithmetic errors. If the calculations are done correctly, both give Fe:O = 2.33:1. A different ratio would suggest impurity or error — but the question designs both to give the same ratio.

Answer

(1) 2.33:1 in both

Q.10 Copper oxide was prepared by two different methods. In one case, 1.75 g of the metal gave 2.19 g of oxide. In the second case, 1.14 g of the metal gave 1.43 g of the oxide. The law illustrated is

Explanation

Method 1: Cu = 1.75 g, CuO = 2.19 g $\implies$ O = 0.44 g. Ratio Cu:O = $1.75 : 0.44 = 3.977 \approx 4 : 1$.

Method 2: Cu = 1.14 g, CuO = 1.43 g $\implies$ O = 0.29 g. Ratio Cu:O = $1.14 : 0.29 = 3.931 \approx 4 : 1$.

Both methods give the same Cu:O ratio (≈ 4:1), regardless of how the oxide was prepared. This illustrates the **Law of Constant Proportions (Definite Proportions)**.

Approach / Analogy

Key phrase: "prepared by two different methods." Same compound from different methods having the same ratio = Law of Definite Proportions. If two *different compounds* of the same elements were compared, it would be Multiple Proportions. Here it's the same compound (CuO) from two methods.

Common Mistake

Confusing this with Multiple Proportions. The Law of Multiple Proportions applies when two *different* compounds of the same elements show whole-number ratios (e.g., CuO vs Cu_2O). Here *the same compound* (CuO) prepared by two methods gives the same ratio — that's Definite Proportions.

Answer

(1) Law of constant proportions

TYPE 3 : Law of Multiple Proportions

Method for Multiple Proportions problems:
Fix the mass of one element (the "common" element) at 1 g or some convenient value. Find how much of the other element combines with that fixed amount in each compound. The ratio of those amounts must be a simple whole-number ratio.

Q.11 Hydrogen peroxide (H_2O_2) and water (H_2O) contain 5.93% and 11.2% of hydrogen respec-

tively. The ratio of masses of oxygen that combine with fixed mass (1 g) of hydrogen in H_2O_2 and H_2O is

Explanation

Fix hydrogen at 1 g in each compound.

In H_2O_2: H = 5.93%, O = 94.07%.

For 1 g H: mass of O $= \frac{94.07}{5.93} = 15.86 \text{ g} \approx 16 \text{ g}$.

In H_2O: H = 11.2%, O = 88.8%.

For 1 g H: mass of O $= \frac{88.8}{11.2} = 7.93 \text{ g} \approx 8 \text{ g}$.

Ratio of O (in H_2O_2) : O (in H_2O) = 16 : 8 = 2:1

Simple whole-number ratio confirms **Law of Multiple Proportions**.

Approach / Analogy

"Fix" hydrogen at 1 g as the reference. Then ask: how much O goes with that 1 g H in each compound? Answer: 16 g in H_2O_2 and 8 g in H_2O. Ratio = 2:1 (a simple integer ratio). This is exactly what Dalton's Law of Multiple Proportions says — when fixed H combines with oxygen in two compounds, the oxygen masses are in a simple ratio.

Common Mistake

Computing the ratio of H percentages ($11.2/5.93 \approx 2$) and concluding O ratio is 2:1 by inverse. This accidentally gives the right answer here but the logic is flawed. The correct method: fix H at 1 g, compute O in each compound separately, then find their ratio.

Answer

(1) 2:1

Q.12 Carbon combines with hydrogen in compounds P, Q and R. The % of hydrogen in P, Q and R are 25, 14.3 and 7.7 respectively. The ratio of carbon in P, Q and R is

Explanation

% of C: P = 75%, Q = 85.7%, R = 92.3%.

Fix hydrogen at 1 g. Find carbon combining with it.

$$\text{P: } \frac{\% C}{\% H} = \frac{75}{25} = 3 \text{ g C per g H}$$
$$\text{Q: } \frac{\% C}{\% H} = \frac{85.7}{14.3} = 5.993 \approx 6 \text{ g C per g H}$$
$$\text{R: } \frac{\% C}{\% H} = \frac{92.3}{7.7} = 11.987 \approx 12 \text{ g C per g H}$$

Ratio of C: $3 : 6 : 12 = \mathbf{1} : \mathbf{2} : \mathbf{4}$

(Simple whole-number ratio — Law of Multiple Proportions.)

Approach / Analogy

Fix H at 1 g as the reference element. For each compound, C:H ratio = %C/%H. Simplify: P gives 3, Q gives 6, R gives 12. Ratio 3:6:12 = 1:2:4. Compounds P, Q, R are likely CH_4, C_2H_4, C_4H_4 (or similar hydrocarbons). The 1:2:4 whole-number ratio confirms Dalton's Law.

Common Mistake

Using the %H values directly as the ratio ($25 : 14.3 : 7.7$). This is the ratio of H abundances, not the ratio of C combining with fixed H. Fix H, find C per 1g H in each compound, then compare those C values.

Answer

(1) 1:2:4

Q.13 Hydrogen and oxygen are known to form two or more compounds. The hydrogen content in one of these is 5.93% while in the other it is 11.2%. The law illustrated by this data is

Explanation

Same data as Q.11: H_2O_2 (5.93% H) and H_2O (11.2% H) are two compounds formed by the same two elements (H and O). The oxygen masses combining with a fixed hydrogen mass are in the ratio 2:1 (simple whole-number ratio) — this is the **Law of Multiple Proportions**.

Approach / Analogy

Clue: "two or more compounds" formed by the *same two elements*. Whenever you see two different compounds of the same pair of elements with different compositions, think Law of Multiple Proportions. This is Dalton's contribution to atomic theory.

Common Mistake

Choosing Law of Constant Proportions. That law applies to a *single* compound always having the same composition. Here, *two different* compounds (H_2O and H_2O_2) are being compared — that's Multiple Proportions, not Constant Proportions.

Answer

(1) Law of multiple proportions

Q.14 The percentage of hydrogen in water and hydrogen peroxide is 11.1 and 5.9 respectively. These figures illustrate

Explanation

Same concept as Q.13: H_2O (11.1% H) and H_2O_2 (5.9% H) are two different compounds of hydrogen and oxygen. The O masses combining with fixed H are in a simple ratio (2:1). This is the **Law of Multiple Proportions**.

Approach / Analogy

Two different compounds of H and O → Law of Multiple Proportions. Note: Q.11, Q.13, and Q.14 all give the same H_2O / H_2O_2 data in slightly different ways. Recognise the pattern: two compounds, same elements, different %H → always Multiple Proportions.

Answer

(1) Law of multiple proportions

Q.15 Different proportions of oxygen in the various oxides of nitrogen prove the law of

Explanation

Nitrogen forms multiple oxides: N_2O, NO, N_2O_3, NO_2, N_2O_5. In each, a fixed mass of N combines with different masses of O that are in simple whole-number ratios. This proves the **Law of Multiple Proportions**.

Approach / Analogy

"Various oxides" of the same element pair = multiple compounds of the same two elements. Multiple compounds + whole-number oxygen ratios = Law of Multiple Proportions. Think: if nitrogen formed only one oxide, it would illustrate Definite Proportions. Multiple oxides → Multiple Proportions.

Common Mistake

Choosing "Constant proportion." That would apply if nitrogen had only one oxide. Multiple different oxides with different O:N ratios illustrate Multiple Proportions.

Answer

(2) Multiple proportion

Q.16 Oxygen combines with two isotopes of carbon ^{12}C and ^{14}C to form two samples of carbon dioxide. The data illustrates

Explanation

Both $^{12}CO_2$ and $^{14}CO_2$ are formed by carbon and oxygen. However, isotopes are atoms of the **same element** (carbon) with different masses. The two "compounds" being compared here differ only in the isotope of carbon used — they are both carbon dioxide, just with different C isotopes.

This doesn't illustrate Multiple Proportions (which requires *different compound types*), nor Definite Proportions (which would apply to a single compound). It doesn't cleanly fit any of the standard laws. **Answer: None of these.**

Approach / Analogy

This is a tricky question. Multiple Proportions requires two or more *structurally different* compounds from the same two elements (like H_2O and H_2O_2). Here, $^{12}CO_2$ and $^{14}CO_2$ differ only in isotope — they are the same chemical compound (CO_2). Isotope variation is a nuclear/quantum phenomenon, not a chemical combination law. Hence: None of these.

Common Mistake

Choosing Law of Multiple Proportions by analogy with other questions. Multiple Proportions requires the compounds to have the same pair of *elements* but different *combining ratios*. Here both samples are CO_2 with the same C:O ratio (just different isotopes of C). The same formula = same law of definite proportions should apply, but isotopes complicate this. The answer "None of these" is correct.

Answer

(4) None of these

Q.17 Which one of the following pairs of compounds illustrates the law of multiple proportions?

Explanation

Law of Multiple Proportions requires the **same two elements** forming two **different compounds**:

- H_2O and Na_2O: different second elements (O, but first elements H vs Na — not same pair). ×
- MgO and Na_2O: Mg and O vs Na and O — different first elements. ×
- Na_2O and BaO: Na+O vs Ba+O — different first elements. ×
- **$SnCl_2$ and $SnCl_4$**: both contain Sn and Cl. In $SnCl_2$: Cl:Sn = 2:1. In $SnCl_4$: Cl:Sn = 4:1. Ratio of Cl masses per fixed Sn = 2:4 = **1:2 (simple whole number ratio).** ✓

Approach / Analogy

The pair must be: same two elements, two different compounds. $SnCl_2$ and $SnCl_4$ are both tin chlorides (Sn + Cl). Per mole of Sn: Cl is 2 in one and 4 in other. Ratio = 1:2. This is the textbook definition of Multiple Proportions.

Common Mistake

Choosing H_2O and Na_2O thinking both have oxygen. But Multiple Proportions requires the *same pair* of elements. H_2O has H and O; Na_2O has Na and O — these share only one element (O), not both.

Answer

(4) $SnCl_2$, $SnCl_4$

Q.18 Element X forms five stable oxides with oxygen of formula X_2O, XO, X_2O_3, X_2O_4, X_2O_5. The formation of these oxides explains

Explanation

Fix mass of X. The oxygen combining with fixed X in each oxide:

$$X_2O : XO : X_2O_3 : X_2O_4 : X_2O_5 \rightarrow O \text{ per 2 X} = 1 : 2 : 3 : 4 : 5$$

These are in simple whole-number ratios $1 : 2 : 3 : 4 : 5$ — the definition of the **Law of Multiple Proportions**.

Approach / Analogy

Five different oxides of the same element X with oxygen. Masses of O combining with fixed X are in ratio 1:2:3:4:5. Simple integer ratio = Multiple Proportions. This is analogous to nitrogen's oxides (N_2O, NO, N_2O_3, NO_2, N_2O_5).

Answer

(3) Law of multiple proportions

Q.19 Which of the following pairs of compound illustrate the law of multiple proportions?

Explanation

Check each pair for: same two elements, different compounds.

- KOH, CsOH: K+O+H vs Cs+O+H — different elements (K vs Cs). ×
- H_2O, D_2O: H+O vs D+O — D (deuterium) is isotope of H, not a different element chemically. Same pair essentially. ×
- **Ethane (C_2H_6) and benzene (C_6H_6)**: both contain only C and H. C:H ratios differ ($2:6 = 1:3$ vs $6:6 = 1:1$). The C masses per fixed H: ethane gives $12/1 = 12$, benzene gives $12 \times 6/6 = 12$... Actually both have the same per-H carbon mass! Let's check carefully. Per 1g H: in C_2H_6 = $24/6 = 4$ g C. In C_6H_6 = $72/6 = 12$ g C. Ratio = 4:12 = 1:3. Simple whole-number ratio. ✓
- KCl, KBr: K+Cl vs K+Br — different halogens (Cl vs Br), not same two elements. ×

Ethane and benzene both contain C and H with different C:H ratios.

Approach / Analogy

Both ethane (C_2H_6) and benzene (C_6H_6) contain only carbon and hydrogen — same two elements, different formulas. C per 1g H: ethane = 4g, benzene = 12g. Ratio = 1:3. Simple integer ratio = Multiple Proportions.

Common Mistake

Dismissing H_2O / D_2O as a valid pair. While D (deuterium) is technically a different isotope of H, chemically they are considered the same element. Standard chemistry treats isotopes as the same element, so H_2O and D_2O illustrate isotope effects, not Multiple Proportions in the classical sense.

Answer

(3) Ethane, benzene

Q.20 The law of multiple proportion was proposed by

Explanation

The Law of Multiple Proportions was proposed by **John Dalton** in 1803, as part of his atomic theory. It explained why elements can combine in different ratios to form different compounds.

For reference:

- **Lavoisier**: Law of Conservation of Mass
- **Proust**: Law of Definite Proportions (Constant Composition)
- **Dalton**: Law of Multiple Proportions + Atomic Theory
- **Gay-Lussac**: Law of Gaseous Volumes
- **Avogadro**: Equal volumes of gases contain equal molecules

Approach / Analogy

Memory trick for the laws and their proposers:
Lavoisier = **L**oss (nothing lost, Conservation of Mass)
Proust = **P**ure compound (same Proportion always)
Dalton = **D**ifferent compounds (Multiple ratios)
Gay-Lussac = **G**as volumes
Avogadro = **A**ll gases equal molecules

Common Mistake

Confusing Proust and Dalton. Proust proposed Definite Proportions (one compound, fixed ratio). Dalton proposed Multiple Proportions (two compounds, simple ratio between them). Both involve ratios but for different scenarios.

Answer

(2) Dalton

Q.21 Suppose the elements X and Y combine to form two compounds XY_2 and X_3Y_2. When 0.1 mole of XY_2 weighs 10 g and 0.05 mole of X_3Y_2 weighs 9 g, the atomic weights of X and Y are

Explanation

From XY_2:

$$M_r(\mathrm{XY_2}) = \frac{10}{0.1} = 100 \text{ g/mol}$$

$$A_X + 2A_Y = 100 \quad \cdots (1)$$

From X_3Y_2:

$$M_r(\mathrm{X_3Y_2}) = \frac{9}{0.05} = 180 \text{ g/mol}$$

$$3A_X + 2A_Y = 180 \quad \cdots (2)$$

Subtract (1) from (2):

$$2A_X = 80 \implies A_X = 40$$

From (1):

$$40 + 2A_Y = 100 \implies A_Y = 30$$

Atomic weights: X = 40, Y = 30.

Approach / Analogy

Two unknowns (A_X and A_Y), two equations from the two compounds. Use M_r = mass/moles to get M_r of each compound. Set up simultaneous equations. Subtract to eliminate A_Y and find $A_X = 40$, then $A_Y = 30$.

Common Mistake

Setting up $A_X + 2A_Y = 10/0.1$ but writing $A_X + A_Y = 10/0.1$ (forgetting the subscript 2 on Y). Always include the subscripts from the molecular formula in the equation: XY_2 has 1 X and 2 Y atoms per formula unit.

Answer

(3) 40, 30

End of Part 1 — Q.1 to Q.21 (TYPE 1–3)

Part 2 covers TYPE 4 (Gay-Lussac) and TYPE 5 (Avogadro) — Q.22–Q.31

TYPE 4 : Gay-Lussac's Law of Gaseous Volumes

Key: At the same T and P, gas volumes react in the same ratio as the stoichiometric coefficients (mole ratios). Volume ratios = mole ratios for gases at equal conditions.

Q.22 For the gaseous reaction: $H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$. If 40 mL of hydrogen completely reacts with chlorine, the required volume of Cl_2 and volume of produced HCl are

Explanation

From the equation: $H_2 : Cl_2 : HCl = 1 : 1 : 2$ (volume ratios at same T, P).

Given: 40 mL H_2 reacts completely.

$$V(\mathrm{Cl_2}) = 1 \times 40 = \mathbf{40}\ \mathrm{mL}$$
$$V(\mathrm{HCl}) = 2 \times 40 = \mathbf{80}\ \mathrm{mL}$$

Approach / Analogy

Gay-Lussac's law: volumes of gases react in the same simple ratios as the equation coefficients (1:1:2). 40 mL H_2 needs 40 mL Cl_2 and produces 80 mL HCl. Like a recipe: 1 cup A + 1 cup B makes 2 cups C. Use 40 cups A: need 40 cups B, make 80 cups C.

Common Mistake

Confusing Gay-Lussac's Law with Avogadro's Law. Both involve volumes and gases, but Gay-Lussac's deals with reacting volumes (stoichiometry), while Avogadro's deals with equal volumes having equal molecules. For stoichiometry: volume ratio = coefficient ratio.

Answer

(1) 40 mL Cl_2 and 80 mL HCl

Q.23 For the gaseous reaction: $H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$. Initially 20 mL of $H_2(g)$ and 30 mL of $Cl_2(g)$ are present. The volume of HCl(g) produced and unreacted Cl_2 are

Explanation

Coefficients: H_2 : Cl_2 : HCl = 1 : 1 : 2.

Divide by coefficients to find limiting gas:

$$\frac{V_{H_2}}{1} = 20, \qquad \frac{V_{Cl_2}}{1} = 30$$

H_2 is limiting $(20 < 30)$. Scale = 20 mL.

$$V(\text{HCl produced}) = 2 \times 20 = \mathbf{40}\text{ mL}$$
$$V(\text{Cl}_2\text{ consumed}) = 1 \times 20 = 20\text{ mL}$$
$$V(\text{Cl}_2\text{ unreacted}) = 30 - 20 = \mathbf{10}\text{ mL}$$

Approach / Analogy

Limiting reagent for gases: divide volume by coefficient, take the smaller one. H_2 (20/1 = 20) limits vs Cl_2 (30/1 = 30). All 20 mL H_2 reacts with 20 mL Cl_2 to give 40 mL HCl. Leftover Cl_2 = 30 − 20 = 10 mL.

Common Mistake

Not identifying the limiting reagent and assuming all of both gases react. With 20 mL H_2 and 30 mL Cl_2 in a 1:1 ratio, only 20 mL Cl_2 reacts (limited by H_2). The remaining 10 mL Cl_2 is unreacted and must be included in the answer.

Answer

(1) 40 mL HCl and 10 mL Cl_2

Q.24 In the reaction $N_2 + 3H_2 \longrightarrow 2NH_3$, the ratio by volume of N_2, H_2 and NH_3 is 1:3:2. This illustrates the law of

Explanation

The volumes of gases reacting and produced are in simple whole-number ratios (1:3:2), at the same temperature and pressure. This is the statement of **Gay-Lussac's Law of Gaseous Volumes**.

Approach / Analogy

Whenever a question mentions volumes of gases reacting in simple integer ratios — it's Gay-Lussac's Law. If it mentions equal volumes having equal molecules — it's Avogadro's Law. Different concepts, both about gases!

Common Mistake

Choosing Conservation of Mass. That law is about masses being equal on both sides. Here the statement is about the simple 1:3:2 ratio of *volumes* — which is Gay-Lussac's Law of Gaseous Volumes.

Answer

(4) Gaseous volumes

TYPE 5 : Avogadro's Law

Avogadro's Law: Equal volumes of all gases, at the same T and P, contain equal numbers of **molecules** (not atoms — molecules!).

Consequence: At same T and P: equal volumes $\Rightarrow$ equal moles $\Rightarrow$ same number of molecules.
Corollary: M_r of unknown gas / M_r of reference gas = (mass of unknown) / (mass of same volume of reference gas at same T, P).

Q.25 Which of the following represents Avogadro's hypothesis?

Explanation

Avogadro's hypothesis (1811) states: **"Equal volumes of all gases under the same conditions of temperature and pressure contain equal numbers of molecules."**

This is exactly option (2). The other options describe:
- (1): Gay-Lussac's Law of Gaseous Volumes
- (3): Wrong — it says "equal atoms" (incorrect; Avogadro said molecules, not atoms)
- (4): Graham's Law of Diffusion

Approach / Analogy

The key word is **molecules**, not atoms. Avogadro specifically corrected Dalton's atomic theory by distinguishing atoms from molecules. Equal volumes have equal *molecules*. A diatomic gas like H_2 has 2 atoms per molecule — so equal volumes of H_2 and O_2 have equal molecules but different total atom counts.

Common Mistake

Choosing option (3) which says "equal atoms." Avogadro's hypothesis is specifically about **molecules**, not atoms. One mole of H_2 and one mole of O_2 have the same number of molecules (N_A) but different atom counts ($2N_A$ for H_2 vs $2N_A$ for O_2 — actually same here, but for CH_4: $5N_A$ atoms vs $2N_A$ for O_2). Molecules = correct; atoms = wrong.

Answer

(2) Equal volumes of all gases under same conditions of T and P contain equal number of molecules

Q.26 Equal volume of different gases at any definite temperature and pressure have

Explanation

By Avogadro's Law: equal volumes of gases at same T and P $\Rightarrow$ equal number of molecules $\Rightarrow$ **equal moles**.

They do *not* have equal masses (different molar masses) or equal densities (density = mass/volume, different for different gases) or equal atoms (different gases have different atomicities).

Approach / Analogy

Equal molecules = equal moles (since moles = molecules/N_A, and N_A is the same for all). But masses differ: 1 mol H_2 = 2 g, 1 mol O_2 = 32 g. Same number of molecules, very different masses.

Common Mistake

Choosing "equal masses." Equal volumes of gases at the same T and P have equal *moles*/molecules, but their masses differ because molar masses differ. Equal molecules $\neq$ equal masses.

Answer

(4) Equal molecules

Q.27 Number of molecules in 100 mL of each of O_2, NH_3 and CO_2 at STP are

Explanation

By Avogadro's Law: equal volumes of gases at the same T and P (STP) contain equal numbers of molecules.

All three samples (100 mL each, all at STP) $\Rightarrow$ **same number of molecules.**

Approach / Analogy

All three containers have the same volume (100 mL), the same temperature (STP = 0°C), and the same pressure (1 atm). Avogadro's law directly says: same V, T, P $\rightarrow$ same number of molecules, regardless of the gas type.

Common Mistake

Thinking heavier gases (CO_2, $M_r = 44$) have more molecules per 100 mL than lighter gases (NH_3, $M_r = 17$). Molecule count depends on volume/T/P only, not on molar mass. All three have the same molecule count at the same V, T, P.

Answer

(3) The same

Q.28 Two flasks A and B of equal capacity of volume contain NH_3 and SO_2 gas respectively under similar conditions. Which flask has more number of moles?

Explanation

Equal volume, same temperature and pressure $\Rightarrow$ equal number of molecules (Avogadro's Law) $\Rightarrow$ equal number of moles.

Both flasks have the same number of moles.

Approach / Analogy

NH_3 ($M_r = 17$) and SO_2 ($M_r = 64$) are very different in molar mass, but at the same V, T, P they contain the same number of moles (and molecules). SO_2 just has more total mass (heavier molecules), but not more molecules.

Common Mistake

Choosing Flask B (SO_2) thinking heavier gas has more moles. More mass $\neq$ more moles when molecular masses are different. Moles depend on number of molecules (Avogadro), and equal V, T, P gives equal molecules regardless of molecular weight.

Answer

(3) Both have same moles

Q.29 Four one-litre flasks are separately filled with the gases hydrogen, helium, oxygen and ozone at same room temperature and pressure. The ratio of total number of atoms of these gases present in the different flasks would be

Explanation

By Avogadro's Law: same volume, T, P $\Rightarrow$ same number of molecules (n) in each flask.

Now count atoms per molecule:

Gas	Formula	Atoms/molecule	Total atoms
Hydrogen	H_2	2	$2n$
Helium	He	1	n
Oxygen	O_2	2	$2n$
Ozone	O_3	3	$3n$

Ratio $= 2n : n : 2n : 3n = \mathbf{2 : 1 : 2 : 3}$

Approach / Analogy

Equal molecules in each flask (Avogadro). But molecules are not atoms: H_2 has 2 atoms, He has 1 (monatomic), O_2 has 2, O_3 has 3. Multiply molecule count by atomicity to get atom count. Ratio of atoms = ratio of atomicities = 2:1:2:3.

Common Mistake

Answering 1:1:1:1 (equal atoms) by confusing Avogadro's equal molecules with equal atoms. Equal volumes give equal *molecules*, not equal atoms. Molecules differ in atomicity (number of atoms per molecule). He (monatomic) has 1 atom per molecule; O_3 (triatomic) has 3.

Answer

(3) 2:1:2:3

Q.30 A container of volume V contains 0.28 g of N_2 gas. If same volume of an unknown gas under similar conditions of temperature and pressure weighs 0.44 g, the molecular mass of the gas is

Explanation

Same volume, same T, P ⇒ same number of moles (Avogadro's Law).

$$n(\mathrm{N_2}) = \frac{0.28}{28} = 0.01 \text{ mol}$$

$$n(\text{unknown}) = 0.01 \text{ mol (same moles)}$$

$$M_r(\text{unknown}) = \frac{0.44}{0.01} = \mathbf{44} \text{ g/mol}$$

The unknown gas with $M_r = 44$ is CO_2.

Approach / Analogy

Equal volume at same T, P → equal moles. So moles of unknown = moles of N_2 = 0.01. Then $M_r = 0.44/0.01 = 44$. Alternatively: ratio of masses = ratio of molar masses (since moles are equal). $M_r(\text{unknown})/28 = 0.44/0.28 \Rightarrow M_r = 28 \times 44/28 = 44$.

Common Mistake

Setting up $M_r = 0.44/V$ without using Avogadro's Law to equate moles. The key insight: same volume at same T, P → same moles. Once you know moles of unknown = moles of N_2 = 0.28/28 = 0.01 mol, then $M_r = 0.44/0.01 = 44$ is straightforward.

Answer

(2) 44

Q.31 A and B are two identical vessels. A contains 15 g ethane at 1 atm and 298 K. The vessel B contains 75 g of a gas X_2 at same temperature and pressure. The vapour density of X_2 is

Explanation

Same volume, same T, P ⇒ same moles.

$$n(\text{ethane, } \mathrm{C_2H_6}) = \frac{15}{30} = 0.5 \text{ mol}$$

Therefore, moles of X_2 = 0.5 mol.

$$M_r(\mathrm{X_2}) = \frac{75}{0.5} = 150 \text{ g/mol}$$

$$\text{VD} = \frac{M_r}{2} = \frac{150}{2} = \mathbf{75}$$

Approach / Analogy

Avogadro's Law: same V, T, P → same moles. Ethane ($M_r = 30$): 15/30 = 0.5 mol. So X_2 also = 0.5 mol. $M_r(\mathrm{X_2}) = 75/0.5 = 150$. VD = $M_r/2 = 75$. Two-step: find M_r from Avogadro, then halve for VD.

Common Mistake

Reporting 150 (the molar mass) as the VD. VD = $M_r/2$, not M_r. VD is defined relative to hydrogen ($M_r = 2$), so VD = $M_r/2 = 150/2 = 75$. Always divide M_r by 2 to get VD.

Answer

(1) 75

— End of DPP-8 Complete Solution Sheet —

Laws of Chemical Combination · Q.1–Q.31 · All Parts Complete

"The difference between average and confident students is assignment completion."