Weird Chemist

DPP–3 : Stoichiometry

Chapter: Some Basic Concepts of Chemistry

Solution Sheet

TYPE I(i) : Basic Stoichiometry — Finding weight of species

Golden Rule of Stoichiometry:

$$\frac{\text{Given moles}}{\text{Stoichiometric coefficient}} = \text{Reaction scale factor}$$

Then: Required moles = Stoichiometric coefficient of desired species × Scale factor.

Step flow: Given mass $\xrightarrow{\div M_r}$ Moles $\xrightarrow{\times\text{ratio}}$ Moles of product $\xrightarrow{\times M_r}$ Mass

Q.1 The equation $2Al_{(s)} + \frac{3}{2}O_{2(g)} \longrightarrow Al_2O_{3(s)}$ shows that:

Explanation

Analysing each option from the balanced equation $2Al + \frac{3}{2}O_2 \rightarrow Al_2O_3$:

- Option (1): Says product = $\frac{7}{2}$ mol Al_2O_3 — **Wrong**. Coefficient of Al_2O_3 = 1, not $\frac{7}{2}$.
- Option (2): Says 2 g Al + $\frac{3}{2}$ g O_2 — **Wrong**. Coefficients represent moles, not grams.
- Option (3): Says $\frac{3}{2}$ L of O_2 — **Wrong**. Volume at STP needs separate calculation.
- Option (4): **Correct.** 2 mol Al reacts with $\frac{3}{2}$ mol O_2 to produce 1 mol Al_2O_3. The coefficients directly give molar ratios.

Approach / Analogy

A balanced chemical equation is a **mole recipe**, not a mass recipe or a volume recipe. The coefficients (2, $\frac{3}{2}$, 1) are mole ratios — like a recipe saying “2 cups flour + $\frac{3}{2}$ cups sugar makes 1 cake.” Those are counts (cups/moles), not weights.

Common Mistake

Students confuse stoichiometric coefficients with grams or litres. Coefficients ONLY represent moles directly. 2 mol Al $\neq$ 2 g Al (Al has molar mass 27 g/mol, so 2 mol = 54 g).

Answer

(4) 2 mol of Al reacts with $\frac{3}{2}$ mol of O_2 to produce 1 mol of Al_2O_3

Q.2 For the reaction $2P + Q \longrightarrow R$, 4 mol of P and excess of Q will produce:

Explanation

From the equation: 2 mol P produces 1 mol R (molar ratio P:R = 2:1).

$$n_R = \frac{4}{2} \times 1 = \mathbf{2} \text{ mol R}$$

Q is in excess so it is not the limiting reagent — P controls the yield.

Approach / Analogy

"Excess of Q" is the exam's way of saying: ignore Q, it will never run out. Only worry about P. 2 P makes 1 R, so 4 P makes 2 R. Like: if 2 workers build 1 house and you have 4 workers (and unlimited materials), they build 2 houses.

Common Mistake

Thinking Q also limits the reaction and trying to use Q's moles (which aren't even given). When the question says "excess of Q," Q is irrelevant to the yield calculation — P is the only limiting species.

Answer

(3) 2 mol of R

Q.3 If 1.5 moles of oxygen combine with Al to form Al_2O_3, the weight of Al used in the reaction is: $4Al + 3O_2 \longrightarrow 2Al_2O_3$

Explanation

From the equation: 3 mol O_2 reacts with 4 mol Al.

$$n(\text{Al}) = \frac{4}{3} \times 1.5 = 2 \text{ mol}$$

$$\text{Mass of Al} = 2 \times 27 = \mathbf{54} \text{ g}$$

Approach / Analogy

Set up the mole ratio: O_2:Al = 3:4. Given 1.5 mol O_2, scale: $1.5 \times \frac{4}{3} = 2$ mol Al. Multiply by molar mass (27) to get 54 g. Think of it as: if 3 parts O_2 need 4 parts Al, how much Al do 1.5 parts O_2 need? Cross-multiply.

Common Mistake

Using the ratio $\frac{4}{3}$ upside-down: writing $n(\text{Al}) = 1.5 \times \frac{3}{4} = 1.125$ mol instead of $1.5 \times \frac{4}{3} = 2$ mol. Always set up the ratio as: $\frac{n_{\text{Al}}}{4} = \frac{n_{O_2}}{3}$, then solve.

Answer

(3) 54 g

Q.4 In a given reaction, 9 g of Al will react with how much O_2? $2Al + \frac{3}{2}O_2 \longrightarrow Al_2O_3$

Explanation

$$\begin{aligned} n(\text{Al}) &= \frac{9}{27} = \frac{1}{3}\ \text{mol} \\ \text{Al} : \text{O}_2 &= 2 : \frac{3}{2} \\ n(\text{O}_2) &= \frac{1}{3} \times \frac{3/2}{2} = \frac{1}{3} \times \frac{3}{4} = \frac{1}{4}\ \text{mol} \\ \text{Mass of O}_2 &= \frac{1}{4} \times 32 = \mathbf{8}\ \text{g} \end{aligned}$$

Approach / Analogy

Scale factor from Al: $\frac{1/3}{2} = \frac{1}{6}$ (reaction runs at $\frac{1}{6}$ scale). O_2 needed = $\frac{3}{2} \times \frac{1}{6} = \frac{1}{4}$ mol = 8 g. Alternatively use direct ratio: Al:O_2 = 2:$\frac{3}{2}$ ⇒ $O_2 = \frac{1}{3} \times \frac{3/2}{2} = \frac{1}{4}$ mol.

Common Mistake

Forgetting fractional coefficients are still mole ratios. $\frac{3}{2}$ mol O_2 reacts with 2 mol Al — this is the ratio. Don't multiply 9 g by $\frac{3}{2}$ directly (that's mixing grams with coefficients).

Answer

(2) 8 g O_2

Q.5 1.5 mol of O_2 combine with Mg to form oxide MgO. The mass of Mg (At. mass 24) that has combined is:

Explanation

Balanced equation: $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$

Ratio: Mg : O_2 = 2 : 1

$$n(\text{Mg}) = 2 \times 1.5 = 3\ \text{mol}$$

$$\text{Mass} = 3 \times 24 = \mathbf{72}\ \text{g}$$

Approach / Analogy

The equation is not given — you need to write it yourself! 2 Mg atoms combine with 1 O_2 molecule (Mg forms MgO, a 1:1 ionic compound). So ratio Mg:O_2 = 2:1. Given 1.5 mol O_2, Mg needed = $2 \times 1.5 = 3$ mol = 72 g.

Common Mistake

Writing the ratio as 1:1 (thinking Mg:O_2 = 1:1 because "MgO has one of each"). The formula MgO tells you atom ratio, but the balanced *equation* gives molecule ratio: 2Mg + O_2 → 2MgO. O_2 is a diatomic molecule — 1 mol O_2 reacts with 2 mol Mg.

Answer

(1) 72 g

Q.6 What quantity of limestone on heating will give 56 kg of CaO?

Explanation

Reaction: $CaCO_3 \rightarrow CaO + CO_2$

Molar ratio: $CaCO_3$: CaO = 1 : 1

$$\begin{aligned} M_r(\text{CaO}) &= 56 \text{ g/mol}, \quad M_r(\text{CaCO}_3) = 100 \text{ g/mol} \\ n(\text{CaO}) &= \frac{56000}{56} = 1000 \text{ mol} \\ n(\text{CaCO}_3) &= 1000 \text{ mol} \\ \text{Mass} &= 1000 \times 100 = 100{,}000 \text{ g} = \mathbf{100} \text{ kg} \end{aligned}$$

Approach / Analogy

CaO has a molar mass of exactly 56 g/mol — so 56 kg CaO = 1000 mol CaO. Since 1 mol $CaCO_3$ gives 1 mol CaO, you need 1000 mol $CaCO_3$ = 100 kg. Notice: $CaCO_3$ (100 g/mol) is heavier than CaO (56 g/mol) — the extra mass is CO_2 that escapes.

Common Mistake

Writing the answer as 56 kg (same mass as CaO). The question involves a chemical change, not a physical one. $CaCO_3$ loses CO_2 (44 kg per 100 kg $CaCO_3$) when heated, so you always need **more** limestone than the CaO produced.

Answer

(4) 100 kg

Q.7 How much iron can be theoretically obtained in the reduction of 1 kg of Fe_2O_3? $Fe_2O_3 + 3H_2 \longrightarrow 2Fe + 3$

Explanation

$$\begin{aligned} M_r(\text{Fe}_2\text{O}_3) &= 2(56) + 3(16) = 160 \text{ g/mol} \\ n(\text{Fe}_2\text{O}_3) &= \frac{1000}{160} = 6.25 \text{ mol} \\ n(\text{Fe}) &= 2 \times 6.25 = 12.5 \text{ mol} \\ \text{Mass of Fe} &= 12.5 \times 56 = \mathbf{700} \text{ g} \end{aligned}$$

Approach / Analogy

From the equation: 1 mol Fe_2O_3 gives 2 mol Fe (ratio 1:2). 1 kg Fe_2O_3 / 160 g/mol = 6.25 mol Fe_2O_3 → 12.5 mol Fe → 12.5 × 56 = 700 g Fe. Notice Fe_2O_3 has 2 iron atoms per formula unit.

Common Mistake

Using a 1:1 ratio for Fe_2O_3 : Fe instead of 1:2. Each formula unit of Fe_2O_3 contains **two** iron atoms. Writing $n(\text{Fe}) = 6.25$ (same as Fe_2O_3) gives 350 g instead of 700 g — exactly half the correct answer.

Answer

(1) 700 g

Q.8 What amount of silver chloride is formed by the action of 5.850 g of NaCl on an excess of $AgNO_3$? $NaCl + AgNO_3 \longrightarrow AgCl + NaNO_3$

Explanation

$$\begin{aligned} M_r(\text{NaCl}) &= 58.5 \text{ g/mol}, \quad M_r(\text{AgCl}) = 108 + 35.5 = 143.5 \text{ g/mol} \\ n(\text{NaCl}) &= \frac{5.850}{58.5} = 0.1 \text{ mol} \\ n(\text{AgCl}) &= 0.1 \text{ mol (1:1 ratio)} \\ \text{Mass of AgCl} &= 0.1 \times 143.5 = \mathbf{14.35} \text{ g} \end{aligned}$$

Approach / Analogy

NaCl and AgCl are in a 1:1 ratio. 0.1 mol NaCl gives 0.1 mol AgCl. Note the big jump in mass: 5.85 g NaCl gives 14.35 g AgCl! This is because Ag (108 g/mol) is much heavier than Na (23 g/mol). Mass is not conserved per species — only moles follow the stoichiometric ratio.

Common Mistake

Expecting the mass of AgCl to equal the mass of NaCl (5.85 g). Mass changes because you're swapping Na (23) for Ag (108) — the AgCl product is much heavier. Always convert to moles first, then use the molar mass of the product.

Answer

(1) 14.35 g

Q.9 For the following reaction, the mass of water produced from 445 g of $C_{57}H_{110}O_6$ is: $2C_{57}H_{110}O_6 + 163O_2$
[JEE(Main)-2019(Jan)]

Explanation

$$M_r(\mathrm{C_{57}H_{110}O_6}) = 57(12) + 110(1) + 6(16) = 684 + 110 + 96 = 890 \text{ g/mol}$$
$$n(\mathrm{C_{57}H_{110}O_6}) = \frac{445}{890} = 0.5 \text{ mol}$$
$$\text{Ratio: } \mathrm{C_{57}H_{110}O_6} : \mathrm{H_2O} = 2 : 110$$
$$n(\mathrm{H_2O}) = 0.5 \times \frac{110}{2} = 27.5 \text{ mol}$$
$$\text{Mass of } \mathrm{H_2O} = 27.5 \times 18 = \mathbf{495} \text{ g}$$

Approach / Analogy

Large molecule — don't panic. Calculate molar mass systematically (57C + 110H + 6O). 445/890 = 0.5 mol of the fat molecule. Ratio of fat to water = 2:110, so 0.5 mol fat gives $0.5 \times 55 = 27.5$ mol water = 495 g. This molecule is a fat (triglyceride)!

Common Mistake

Making an arithmetic error in the molar mass: 57 × 12 = 684 (not 570 or 628). Use 12 × 57 = 12 × 50 + 12 × 7 = 600 + 84 = 684. Also, the ratio is 2:110, not 1:110 — dividing by 2 is essential.

Answer

(4) 495 g

Q.10 The mass of carbon anode consumed (giving only carbon dioxide) in the production of 270 kg of aluminium metal from bauxite by the Hall's process is: $2\mathrm{Al_2O_3} + 3\mathrm{C} \longrightarrow 4\mathrm{Al} + 3\mathrm{CO_2}$

Explanation

$$n(\mathrm{Al}) = \frac{270000}{27} = 10000 \text{ mol}$$
$$\text{Ratio Al : C} = 4 : 3$$
$$n(\mathrm{C}) = 10000 \times \frac{3}{4} = 7500 \text{ mol}$$
$$\text{Mass of C} = 7500 \times 12 = 90{,}000 \text{ g} = \mathbf{90} \text{ kg}$$

Approach / Analogy

270 kg Al = 270,000 g / 27 g/mol = 10,000 mol Al. From equation, Al:C = 4:3, so C needed = $10000 \times 3/4 = 7500$ mol = 90 kg. The carbon anode gets consumed in the process — it reacts with oxygen to form $\mathrm{CO_2}$.

Common Mistake

Using 4:3 ratio upside down: writing $n(\mathrm{C}) = 10000 \times 4/3$ instead of $10000 \times 3/4$. Set up the ratio carefully: $\frac{n_C}{3} = \frac{n_{Al}}{4}$, so $n_C = \frac{3}{4} \times n_{Al}$.

Answer

(4) 90 kg

TYPE I(ii) : Basic Stoichiometry — Volume of gas questions

Key Tool: Molar Volume at STP/NTP
At STP (0°C, 1 atm): 1 mol of any ideal gas = **22.4 L**

Step flow for volume problems:
Given mass $\xrightarrow{\div M_r}$ Moles $\xrightarrow{\times\text{stoich. ratio}}$ Moles of gas $\xrightarrow{\times 22.4}$ Volume at STP

Q.11 In the following reaction, if 10 g of H_2 reacts with N_2, the volume of NH_3 at STP will be: $N_2 + 3H_2 \longrightarrow 2NH_3$

Explanation

$$n(\mathrm{H_2}) = \frac{10}{2} = 5 \text{ mol}$$

$$\text{Ratio } \mathrm{H_2} : \mathrm{NH_3} = 3 : 2$$

$$n(\mathrm{NH_3}) = 5 \times \frac{2}{3} = \frac{10}{3} \text{ mol}$$

$$V(\mathrm{NH_3}) = \frac{10}{3} \times 22.4 = \mathbf{74.67} \text{ L}$$

Approach / Analogy

N_2 is in excess (not given, so irrelevant). H_2 controls the yield. H_2:NH_3 = 3:2, so 5 mol H_2 gives $5 \times \frac{2}{3} = \frac{10}{3}$ mol NH_3. At STP: $\frac{10}{3} \times 22.4 \approx 74.67$ L. The $\frac{2}{3}$ ratio comes from reading the equation: for every 3 mol H_2, 2 mol NH_3 is produced.

Common Mistake

Using ratio $\frac{3}{2}$ instead of $\frac{2}{3}$ for $H_2 \rightarrow NH_3$ conversion. The H_2:NH_3 molar ratio is 3:2, so to get NH_3 from H_2 you multiply by $\frac{2}{3}$ (you get *less* NH_3 than H_2 consumed because 3 parts H_2 make only 2 parts NH_3).

Answer

(1) 74.67 L

Q.12 Assuming 100% yield, how many moles of $NaHCO_3$ will produce 448 mL of CO_2 gas at STP? $NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$ (unbalanced)

Explanation

Step 1: Balance the equation.

$$2\text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O}$$

Ratio: $\text{NaHCO}_3 : \text{CO}_2 = 2 : 1$

Step 2: Moles of CO_2.

$$n(\text{CO}_2) = \frac{448 \text{ mL}}{22400 \text{ mL/mol}} = 0.02 \text{ mol}$$

Step 3: Moles of NaHCO_3.

$$n(\text{NaHCO}_3) = 2 \times 0.02 = \mathbf{0.04} \text{ mol}$$

Approach / Analogy

Key step: **balance the equation first!** The unbalanced equation gives a wrong ratio. 2 NaHCO_3 makes 1 Na_2CO_3 + 1 CO_2 + 1 H_2O. So for every 0.02 mol CO_2 produced, you need double that = 0.04 mol NaHCO_3.

Common Mistake

Using a 1:1 ratio (assuming 1 mol NaHCO_3 gives 1 mol CO_2) without balancing first. The balanced equation shows a 2:1 ratio — ignoring the balancing step gives 0.02 mol instead of 0.04 mol (half the correct answer).

Answer

(1) 0.04

Q.13 In the reaction $2\text{KClO}_3 \longrightarrow 2\text{KCl} + 3\text{O}_2$, when 36.75 g of KClO_3 is heated, the volume of oxygen evolved at NTP will be: [NSEC-2012]

Explanation

$$M_r(\text{KClO}_3) = 39 + 35.5 + 48 = 122.5 \text{ g/mol}$$

$$n(\text{KClO}_3) = \frac{36.75}{122.5} = 0.3 \text{ mol}$$

$$\text{Ratio: KClO}_3 : \text{O}_2 = 2 : 3$$

$$n(\text{O}_2) = 0.3 \times \frac{3}{2} = 0.45 \text{ mol}$$

$$V(\text{O}_2) = 0.45 \times 22.4 = \mathbf{10.08} \text{ dm}^3$$

Approach / Analogy

KClO_3:O_2 = 2:3 — so O_2 formed is 1.5 times the moles of KClO_3. $0.3 \times 1.5 = 0.45$ mol O_2. At NTP = 22.4 L/mol, so volume = $0.45 \times 22.4 = 10.08$ dm^3.

Common Mistake

Using the ratio 2:3 upside down: writing $n(\mathrm{O_2}) = 0.3 \times \frac{2}{3} = 0.2$ mol. Per the equation, $KClO_3$:O_2 = 2:3, so to convert $KClO_3$ moles to O_2 moles, multiply by $\frac{3}{2}$ (you get *more* O_2 than $KClO_3$).

Answer

(3) 10.08 dm^3

Q.14 Number of moles of $KClO_3$ that have to be heated to produce 1.0 L of O_2 at STP: [NSEC-2018]

Explanation

From Q.13: $KClO_3$: O_2 = 2 : 3

$$n(\mathrm{O_2}) = \frac{1.0\ \mathrm{L}}{22.4\ \mathrm{L/mol}} = \frac{1}{22.4}\ \mathrm{mol}$$
$$n(\mathrm{KClO_3}) = \frac{2}{3} \times n(\mathrm{O_2}) = \frac{2}{3} \times \frac{1}{22.4} = \boldsymbol{\frac{2}{3}\left(\frac{1}{22.4}\right)}$$

Approach / Analogy

This is a "formula answer" question — it asks for the expression, not a number. Convert 1 L to moles using 22.4, then apply the $KClO_3$:O_2 = 2:3 ratio backwards: moles of $KClO_3$ = $\frac{2}{3}$ times moles of O_2.

Common Mistake

Inverting the mole-to-volume step: writing $n = 1.0 \times 22.4$ instead of $1.0/22.4$. Dividing by 22.4 converts litres to moles; multiplying by 22.4 converts moles to litres. The question gives a volume and asks for moles, so divide.

Answer

(3) $\frac{2}{3}\left(\frac{1}{22.4}\right)$

Q.15 0.01 mol of iodoform (CHI_3) reacts with Ag to produce a gas whose volume at NTP is: $2CHI_3 + 6Ag \longrightarrow$

Explanation

The gas produced is C_2H_2 (acetylene/ethyne).

$$\text{Ratio: } \mathrm{CHI_3} : \mathrm{C_2H_2} = 2 : 1$$
$$n(\mathrm{C_2H_2}) = \frac{0.01}{2} = 0.005\ \mathrm{mol}$$
$$V = 0.005 \times 22400\ \mathrm{mL} = \mathbf{112}\ \mathrm{mL}$$

Approach / Analogy

2 mol CHI_3 gives 1 mol C_2H_2 (ratio 2:1). So 0.01 mol CHI_3 gives half that = 0.005 mol C_2H_2. Volume = $0.005 \times 22400 = 112$ mL. Note the unit — answer is in mL, not L.

Common Mistake

Using 1:1 ratio and getting $0.01 \times 22400 = 224$ mL (option 1). The equation clearly shows 2 CHI_3 produces 1 C_2H_2 — the ratio is 2:1, not 1:1.

Answer

(2) 112 mL

Q.16 How many litres of CO_2 at STP will be formed when 0.01 mol of H_2SO_4 reacts with excess of Na_2CO_3? $Na_2CO_3 + H_2SO_4 \longrightarrow Na_2SO_4 + CO_2 + H_2O$

Explanation

Ratio: H_2SO_4 : CO_2 = 1 : 1

$$n(\text{CO}_2) = 0.01 \text{ mol}$$
$$V(\text{CO}_2) = 0.01 \times 22.4 = \mathbf{0.224} \text{ L}$$

Approach / Analogy

1:1 ratio makes this straightforward: 0.01 mol H_2SO_4 gives 0.01 mol CO_2. Volume at STP = $0.01 \times 22.4 = 0.224$ L. Na_2CO_3 is in excess — it doesn't limit the yield.

Common Mistake

Multiplying by 22400 instead of 22.4 to get volume in litres. Remember: 22.4 L/mol for litres. If you want mL, use 22400 mL/mol. The answer 0.224 L = 224 mL — both are correct but match units to the option choices.

Answer

(3) 0.227 L

Note

Exact calculation: $0.01 \times 22.4 = 0.224$ L. Option (3) says 0.227 L — likely uses 22.7 L/mol (updated STP). Both are acceptable depending on the STP molar volume used. Verify with your source.

Q.17 The volume of gas at NTP produced by 100 g of CaC_2 with water is: $CaC_2 + 2H_2O \longrightarrow Ca(OH)_2 + C_2H$

Explanation

$$M_r(\text{CaC}_2) = 40 + 24 = 64 \text{ g/mol}$$
$$n(\text{CaC}_2) = \frac{100}{64} = 1.5625 \text{ mol}$$
$$\text{Ratio: CaC}_2 : \text{C}_2\text{H}_2 = 1 : 1$$
$$n(\text{C}_2\text{H}_2) = 1.5625 \text{ mol}$$
$$V = 1.5625 \times 22.4 = \mathbf{35} \text{ L}$$

Approach / Analogy

CaC_2 (calcium carbide) reacts with water 1:1 to give C_2H_2 (acetylene). Molar mass = 64 g/mol. 100/64 = 1.5625 mol → 1.5625 × 22.4 = 35 L. Used in carbide lamps and welding torches!

Common Mistake

Calculating $M_r(\text{CaC}_2) = 40 + 12 = 52$ (using just one carbon). CaC_2 has **two** carbon atoms: Ca(40) + 2C(24) = 64 g/mol. Using 52 g/mol gives the wrong number of moles and hence wrong volume.

Answer

(2) 35 L

Q.18 Aluminum carbide (Al_4C_3) liberates methane on treatment with water. The grams of aluminum carbide required to produce 11.2 L of methane under STP conditions is: [NSEC-2014]

Explanation

Reaction: $Al_4C_3 + 12H_2O \rightarrow 4Al(OH)_3 + 3CH_4$

Ratio: Al_4C_3 : CH_4 = 1 : 3

$$n(\text{CH}_4) = \frac{11.2}{22.4} = 0.5 \text{ mol}$$
$$n(\text{Al}_4\text{C}_3) = \frac{0.5}{3} = \frac{1}{6} \text{ mol}$$
$$M_r(\text{Al}_4\text{C}_3) = 4(27) + 3(12) = 108 + 36 = 144 \text{ g/mol}$$
$$\text{Mass} = \frac{1}{6} \times 144 = \mathbf{24} \text{ g}$$

Approach / Analogy

The equation is not given — write it from the information. Al_4C_3 has 3 carbon atoms, each giving 1 CH_4, so 1 mol Al_4C_3 gives 3 mol CH_4 (ratio 1:3). Working backwards from 0.5 mol CH_4: Al_4C_3 needed = 0.5/3 mol = 1/6 mol = 24 g.

Common Mistake

Using a 1:1 ratio for $Al_4C_3 : CH_4$. The subscript 3 in C_3 of Al_4C_3 tells you there are **3 carbon** atoms per formula unit, so 3 mol CH_4 are produced per mole of Al_4C_3. Always count atoms in the formula.

Answer

(4) 24 g

Q.19 The minimum quantity in grams of H_2S needed to precipitate 63.5 g of Cu^{2+} will be nearly: $Cu^{2+} + H_2S \longrightarrow CuS + H_2^+$

Explanation

$$\begin{aligned} M_r(\text{Cu}) &= 63.5\ \text{g/mol}, \quad M_r(\text{H}_2\text{S}) = 34\ \text{g/mol} \\ n(\text{Cu}^{2+}) &= \frac{63.5}{63.5} = 1\ \text{mol} \\ \text{Ratio: Cu}^{2+} : \text{H}_2\text{S} &= 1 : 1 \\ n(\text{H}_2\text{S}) &= 1\ \text{mol} \\ \text{Mass} &= 1 \times 34 = \mathbf{34}\ \text{g} \end{aligned}$$

Approach / Analogy

1:1 ratio from the equation. 63.5 g Cu^{2+} = 1 mol (convenient!). So 1 mol H_2S is needed. Mass of H_2S = 2 + 32 = 34 g. The Cu^{2+} precipitates as CuS (black precipitate).

Common Mistake

Confusing the mass of Cu^{2+} (63.5 g) with the required mass of H_2S. The question asks for mass of H_2S, not Cu. Since $M_r(\text{H}_2\text{S}) = 34 \neq 63.5$, the answer is 34 g, not 63.5 g.

Answer

(3) 34 g

Q.20 2.76 g of silver carbonate on being strongly heated yields a residue weighing: $Ag_2CO_3 \longrightarrow 2Ag + CO_2 +$

Explanation

$$\begin{aligned} M_r(\text{Ag}_2\text{CO}_3) &= 2(108) + 12 + 48 = 216 + 60 = 276\ \text{g/mol} \\ n(\text{Ag}_2\text{CO}_3) &= \frac{2.76}{276} = 0.01\ \text{mol} \\ n(\text{Ag}) &= 2 \times 0.01 = 0.02\ \text{mol} \\ \text{Mass of Ag} &= 0.02 \times 108 = \mathbf{2.16}\ \text{g} \end{aligned}$$

The solid residue is Ag (silver metal).

Approach / Analogy

When Ag_2CO_3 decomposes, the gases (CO_2 and O_2) escape and the solid residue is pure Ag metal. So mass of residue = mass of Ag produced. 2.76 g Ag_2CO_3 / 276 g/mol = 0.01 mol → 0.02 mol Ag → 2.16 g.

Common Mistake

Trying to subtract the mass of CO_2 and O_2 lost from 2.76 g instead of computing the product mass from moles. Both methods work, but the mole approach is cleaner. Also: don't forget the coefficient — 1 mol Ag_2CO_3 gives **2 mol Ag** (not 1 mol).

Answer

(1) 2.16 g

TYPE II(i) : Limiting Reagent — Moles or Weight given

How to Find the Limiting Reagent:
Divide moles of each reactant by its stoichiometric coefficient. The reactant with the **smallest** value is the limiting reagent.

$$\text{Divide each: } \frac{n_A}{\text{coeff}_A}, \frac{n_B}{\text{coeff}_B} \quad \Rightarrow \quad \text{Smallest} = \text{Limiting Reagent}$$

Q.21 For reaction $A + 5B \longrightarrow C + 3D$, given 10 mol of A and 10 mol of B. Which is the limiting reagent?

Explanation

Divide moles by coefficient:

$$\frac{n_A}{\text{coeff}_A} = \frac{10}{1} = 10$$
$$\frac{n_B}{\text{coeff}_B} = \frac{10}{5} = 2$$

B gives the smaller value ($2 < 10$), so **B is the limiting reagent.**

Approach / Analogy

Think of it like a sandwich shop: making 1 sandwich needs 1 bread + 5 fillings. You have 10 breads and 10 fillings. With 10 fillings you can only make $10/5 = 2$ sandwiches. With 10 breads you could make 10. Fillings (B) run out first — B is the limiting reagent.

Common Mistake

Thinking the species with fewer total moles is always the limiting reagent. Here both have 10 mol, but B has a coefficient of 5 — it gets consumed 5 times faster than A. Always divide by the stoichiometric coefficient before comparing.

Answer

(2) B

Q.22 For reaction A + 2B $\longrightarrow$ C. The amount of product formed by starting with 5 mol of A and 8 mol of B is:

Explanation

$$\frac{n_A}{1} = 5, \quad \frac{n_B}{2} = 4$$

B is limiting ($4 < 5$). Product C is formed in 1:1 ratio with A and 1:2 ratio with B.

From the limiting reagent (B):

$$n_C = \frac{8}{2} = \mathbf{4}\ \text{mol}$$

Approach / Analogy

B is limiting because $8/2 = 4 < 5$. Every 2 mol B gives 1 mol C, so 8 mol B gives 4 mol C. Alternatively: reaction scale = 4 (the smaller divided value), so C produced = $1 \times 4 = 4$ mol.

Common Mistake

Adding moles of A and B to get product: $5 + 8 = 13$ or averaging them. Products are determined by stoichiometry and the limiting reagent — never by adding reactant moles together.

Answer

(4) 4 mol

Q.23 If 0.5 mol of $BaCl_2$ is mixed with 0.2 mol of Na_3PO_4, the maximum number of moles of $Ba_3(PO_4)_2$ that can be formed is: $3BaCl_2 + 2Na_3PO_4 \longrightarrow Ba_3(PO_4)_2 + 6NaCl$

Explanation

$$\frac{n_{\text{BaCl}_2}}{3} = \frac{0.5}{3} = 0.1\overline{6}$$

$$\frac{n_{\text{Na}_3\text{PO}_4}}{2} = \frac{0.2}{2} = 0.1$$

Na_3PO_4 is limiting ($0.1 < 0.167$). Scale factor = 0.1.

$$n_{\text{Ba}_3(\text{PO}_4)_2} = 1 \times 0.1 = \mathbf{0.1}\ \text{mol}$$

Approach / Analogy

Scale factor for this reaction = 0.1 (from the limiting reagent Na_3PO_4). Since coefficient of $Ba_3(PO_4)_2$ = 1, product moles = $1 \times 0.1 = 0.1$ mol.

Common Mistake

Adding the moles: $0.5 + 0.2 = 0.7$ mol (option 1) — completely wrong approach. Or taking the smaller raw mole value (0.2) without dividing by the coefficient first. Always divide by stoichiometric coefficients before identifying the limiting reagent.

Answer

(4) 0.1

Q.24 Maximum number of moles of barium phosphate formed when 0.9 mol of barium chloride is mixed with 0.4 mol of sodium phosphate is: [NSEC-2010]

Explanation

Same reaction: $3BaCl_2 + 2Na_3PO_4 \rightarrow Ba_3(PO_4)_2 + 6NaCl$

$$\frac{0.9}{3} = 0.3, \quad \frac{0.4}{2} = 0.2$$

Na_3PO_4 is limiting $(0.2 < 0.3)$. Scale = 0.2.

$$n_{\text{product}} = 1 \times 0.2 = \mathbf{0.2} \text{ mol}$$

Approach / Analogy

Same reaction structure as Q.23. Quick check: 0.9/3 = 0.3 and 0.4/2 = 0.2. The phosphate (Na_3PO_4) limits again. Product = 0.2 mol.

Answer

(1) 0.2

Q.25 0.5 mole of H_2SO_4 is mixed with 0.2 mole of $Ca(OH)_2$. The maximum number of moles of $CaSO_4$ formed is:

Explanation

Reaction: $H_2SO_4 + Ca(OH)_2 \rightarrow CaSO_4 + 2H_2O$

Coefficients are 1:1:1.

$$\frac{0.5}{1} = 0.5, \quad \frac{0.2}{1} = 0.2$$

$Ca(OH)_2$ is limiting. $n(CaSO_4) = \mathbf{0.2}$ mol.

Approach / Analogy

Both coefficients = 1, so just compare raw moles: 0.5 vs 0.2. $Ca(OH)_2$ (0.2 mol) limits. $CaSO_4$ produced = 0.2 mol (1:1 ratio with $Ca(OH)_2$).

Answer

(1) 0.2

Q.26 4 g of hydrogen are ignited with 4 g of oxygen. The weight of water formed is:

Explanation

Reaction: $2H_2 + O_2 \rightarrow 2H_2O$

$$n(H_2) = \frac{4}{2} = 2\text{ mol}, \quad n(O_2) = \frac{4}{32} = 0.125\text{ mol}$$

$$\frac{n_{H_2}}{2} = 1, \quad \frac{n_{O_2}}{1} = 0.125$$

O_2 is limiting ($0.125 < 1$). Scale = 0.125.

$$n(H_2O) = 2 \times 0.125 = 0.25\text{ mol}$$

$$\text{Mass} = 0.25 \times 18 = \mathbf{4.5}\text{ g}$$

Approach / Analogy

Despite equal masses, 4 g H_2 = 2 mol but 4 g O_2 = only 0.125 mol. O_2 is massively outnumbered in moles. It runs out first. Scale factor = 0.125 → water = $2 \times 0.125 = 0.25$ mol = 4.5 g. Notice: mass isn't conserved here because excess H_2 escapes unreacted.

Common Mistake

Thinking that equal masses (4 g each) means equal moles. Masses are NOT moles! H_2 has $M_r = 2$, O_2 has $M_r = 32$. The same mass gives wildly different mole counts. Always convert to moles first.

Answer

(3) 4.5 g

Q.27 10 g of hydrogen and 64 g of oxygen were filled in a steel vessel and exploded. Amount of water produced in this reaction will be:

Explanation

$2H_2 + O_2 \rightarrow 2H_2O$

$$n(H_2) = \frac{10}{2} = 5\text{ mol}, \quad n(O_2) = \frac{64}{32} = 2\text{ mol}$$

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

O_2 is limiting ($2 < 2.5$). Scale = 2.

$$n(H_2O) = 2 \times 2 = \mathbf{4}\text{ mol}$$

Approach / Analogy

5 mol H_2 and 2 mol O_2. Scale factor from O_2 = 2/1 = 2. Water = $2 \times 2 = 4$ mol. Steel vessel means the gas is trapped — all reacts until limiting reagent runs out.

Answer

(4) 4 mol

Q.28 **$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(g)$. Given 4 g of H_2 and 32 g of O_2, the volume of water vapour at STP is:**

Explanation

$$n(\mathrm{H_2}) = \frac{4}{2} = 2\text{ mol}, \quad n(\mathrm{O_2}) = \frac{32}{32} = 1\text{ mol}$$

$$\frac{n_{\mathrm{H_2}}}{1} = 2, \quad \frac{n_{\mathrm{O_2}}}{1/2} = \frac{1}{0.5} = 2$$

Both equal — **neither is limiting** (exact stoichiometric ratio!).

$$n(\mathrm{H_2O}) = 2\text{ mol (from either reactant)}$$
$$V(\mathrm{H_2O}) = 2 \times 22.4 = \mathbf{44.8}\text{ L}$$

Approach / Analogy

The division gives 2 for both reactants — they are in exact stoichiometric proportion (neither is in excess). Both get fully consumed to give 2 mol H_2O. Volume of steam at STP = $2 \times 22.4 = 44.8$ L.

Common Mistake

Automatically assuming one must be limiting without checking. When both divide-by-coefficient values are equal, the reactants are in perfect stoichiometric ratio and both run out simultaneously. Neither limits — both are completely consumed.

Answer

(1) 44.8 L

Q.29 **28 g Lithium is mixed with 48 g O_2 to react according to $4Li + O_2 \longrightarrow 2Li_2O$. The mass of Li_2O formed is:**

Explanation

$$n(\text{Li}) = \frac{28}{7} = 4 \text{ mol}, \quad n(\text{O}_2) = \frac{48}{32} = 1.5 \text{ mol}$$

$$\frac{4}{4} = 1, \quad \frac{1.5}{1} = 1.5$$

Li is limiting ($1 < 1.5$). Scale = 1.

$$n(\text{Li}_2\text{O}) = 2 \times 1 = 2 \text{ mol}$$

$$\text{Mass} = 2 \times 30 = \mathbf{60} \text{ g}$$

($M_r(\text{Li}_2\text{O}) = 14 + 16 = 30$ g/mol)

Approach / Analogy

Li:O_2 = 4:1. Li runs out first (scale factor 1 vs 1.5 for O_2). Scale = 1 $\rightarrow$ Li_2O = $2 \times 1 = 2$ mol $\rightarrow$ 60 g. Li has $M_r = 7$ g/mol (lightest metal).

Common Mistake

Using $M_r(\text{Li}_2\text{O}) = 2 + 16 = 18$ (confusing Li = 7 with H = 1). Lithium's atomic mass is **7**, not 1. So Li_2O = 2(7) + 16 = 30 g/mol, not 18 g/mol.

Answer

(1) 60 g

Q.30 The mass of Mg_3N_2 produced if 48 g of Mg metal is reacted with 34 g NH_3 gas is: $3\text{Mg} + 2\text{NH}_3 \longrightarrow \text{Mg}_3\text{N}$

Explanation

$$n(\text{Mg}) = \frac{48}{24} = 2 \text{ mol}, \quad n(\text{NH}_3) = \frac{34}{17} = 2 \text{ mol}$$

$$\frac{2}{3} = 0.6\overline{6}, \quad \frac{2}{2} = 1$$

Mg is limiting ($0.67 < 1$). Scale = $\frac{2}{3}$.

$$n(\text{Mg}_3\text{N}_2) = 1 \times \frac{2}{3} = \frac{2}{3} \text{ mol}$$

$$M_r(\text{Mg}_3\text{N}_2) = 3(24) + 2(14) = 72 + 28 = 100 \text{ g/mol}$$

$$\text{Mass} = \frac{2}{3} \times 100 = \frac{200}{3} \text{ g}$$

Approach / Analogy

Both reactants = 2 mol, but Mg has coefficient 3 and NH_3 has coefficient 2. Dividing: Mg gives 2/3 = 0.67 and NH_3 gives 2/2 = 1. Mg limits. Scale factor = 2/3 → Mg_3N_2 = 2/3 mol = 200/3 g.

Common Mistake

Thinking equal moles (2 mol each) means neither limits. Coefficients differ (3 vs 2), so with equal moles Mg runs out first — Mg needs 3 mol per cycle, NH_3 needs only 2 mol per cycle. Equal moles $\neq$ equal supply.

Answer

(1) $\frac{200}{3}$ g

Q.31 If 10 g of Ag reacts with 1 g of sulphur, the amount of Ag_2S formed will be:

Explanation

Reaction: $2\text{Ag} + \text{S} \rightarrow \text{Ag}_2\text{S}$

$$n(\text{Ag}) = \frac{10}{108} = 0.0926 \text{ mol}, \quad n(\text{S}) = \frac{1}{32} = 0.03125 \text{ mol}$$

$$\frac{0.0926}{2} = 0.0463, \quad \frac{0.03125}{1} = 0.03125$$

S is limiting ($0.03125 < 0.0463$). Scale = 0.03125.

$$M_r(\text{Ag}_2\text{S}) = 216 + 32 = 248 \text{ g/mol}$$
$$n(\text{Ag}_2\text{S}) = 1 \times 0.03125 = 0.03125 \text{ mol}$$
$$\text{Mass} = 0.03125 \times 248 = \mathbf{7.75} \text{ g}$$

Approach / Analogy

S is limiting despite having only 1 g — because sulphur (32 g/mol) is heavier per mole and only 1 in the coefficient. Scale = 0.03125 → Ag_2S = 7.75 g.

Common Mistake

Not writing the balanced equation first. The equation 2Ag + S → Ag_2S has coefficients 2:1:1 — this must be known or derived. Using a 1:1 ratio for Ag:S gives a wrong limiting reagent identification.

Answer

(1) 7.75 g

Q.32 How many moles of lead(II) chloride will be formed from a reaction between 6.5 g of PbO and 3.2 g of HCl? (Atomic wt. of Pb = 207)

Explanation

Reaction: $PbO + 2HCl \to PbCl_2 + H_2O$

$$M_r(\text{PbO}) = 207 + 16 = 223 \text{ g/mol}, \quad M_r(\text{HCl}) = 36.5 \text{ g/mol}$$

$$n(\text{PbO}) = \frac{6.5}{223} = 0.02915 \text{ mol}$$

$$n(\text{HCl}) = \frac{3.2}{36.5} = 0.08767 \text{ mol}$$

$$\frac{0.02915}{1} = 0.02915, \quad \frac{0.08767}{2} = 0.04384$$

PbO is limiting. $n(\text{PbCl}_2) = 1 \times 0.02915 \approx \mathbf{0.029}$ mol

Approach / Analogy

PbO:HCl = 1:2 in the balanced equation. PbO has the smaller scale factor (0.029 vs 0.044), so PbO limits. Product = same moles as limiting PbO = 0.029 mol.

Common Mistake

Not knowing the balanced equation for metal oxide + HCl. The standard reaction is: metal oxide + HCl → metal chloride + water. For PbO: PbO + 2HCl → $PbCl_2$ + H_2O. The coefficient 2 for HCl is crucial.

Answer

(2) 0.029

Q.33 The maximum amount of CH_3Cl that can be prepared from 20 g of CH_4 and 10 g of Cl_2 by: $CH_4 + Cl_2 \longrightarrow CH_3Cl + HCl$ [NSEC-2012]

Explanation

$$n(\text{CH}_4) = \frac{20}{16} = 1.25 \text{ mol}, \quad n(\text{Cl}_2) = \frac{10}{71} = 0.1408 \text{ mol}$$

Coefficients both = 1:

$$\frac{1.25}{1} = 1.25, \quad \frac{0.1408}{1} = 0.1408$$

Cl_2 is limiting. $n(\text{CH}_3\text{Cl}) = 0.1408 \approx \mathbf{0.141}$ mol

Approach / Analogy

1:1 ratio — just compare moles directly. Cl_2 (0.141 mol) $<$ CH_4 (1.25 mol), so Cl_2 limits. Product CH_3Cl = 0.141 mol. Note: $M_r(\text{Cl}_2) = 71$ g/mol (two Cl atoms, each 35.5).

Common Mistake

Using $M_r(\text{Cl}_2) = 35.5$ (atomic mass of Cl) instead of 71 (molecular mass of Cl_2). Chlorine gas is **diatomic** — Cl_2 has molar mass = 2 × 35.5 = 71 g/mol, not 35.5.

Answer

(2) 0.141 mol

Q.34 For a reaction $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$; identify dihydrogen ($H_2$) as a limiting reagent in the following reaction mixtures: [JEE(Main)-2019(April)]

Explanation

H_2 is limiting when $\frac{n_{H_2}}{3} < \frac{n_{N_2}}{1}$, i.e., when $n_{H_2} < 3 \times n_{N_2}$.

Check each option:

- (1) 28 g N_2 + 6 g H_2: $n_{N_2} = 1$ mol, $n_{H_2} = 3$ mol. $3/3 = 1 = 1/1 = 1$. **Exact ratio — neither limits.**
- (2) 56 g N_2 + 10 g H_2: $n_{N_2} = 2$ mol, $n_{H_2} = 5$ mol. $5/3 = 1.67 < 2/1 = 2$. **H_2 limits.**
- (3) 14 g N_2 + 4 g H_2: $n_{N_2} = 0.5$ mol, $n_{H_2} = 2$ mol. $2/3 = 0.67 > 0.5/1 = 0.5$. N_2 limits.
- (4) 35 g N_2 + 8 g H_2: $n_{N_2} = 1.25$ mol, $n_{H_2} = 4$ mol. $4/3 = 1.33 > 1.25$. N_2 limits.

Approach / Analogy

For H_2 to be limiting: H_2 must have the smaller "divide by coefficient" value. Check option (2): $10/2 = 5$ mol H_2, $56/28 = 2$ mol N_2. Ratios: H_2: 5/3 = 1.67; N_2: 2/1 = 2. Since $1.67 < 2$, H_2 is limiting. This is the only option where H_2 limits.

Common Mistake

Checking only whether H_2 moles are less than N_2 moles, without dividing by coefficients. In option (2): 5 mol H_2 $>$ 2 mol N_2 in raw moles, yet H_2 is still limiting because it has a coefficient of 3 (gets consumed 3× faster).

Answer

(2) 56 g of N_2 + 10 g of H_2

Q.35 1.0 g of magnesium is burnt with 0.56 g O_2 in a closed vessel. Which reactant is left in excess and by how much? (At. wt. Mg = 24; O = 16)

Explanation

Reaction: $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$

$$n(\text{Mg}) = \frac{1.0}{24} = 0.04167 \text{ mol}, \quad n(\text{O}_2) = \frac{0.56}{32} = 0.0175 \text{ mol}$$

$$\frac{0.04167}{2} = 0.02083, \quad \frac{0.0175}{1} = 0.0175$$

O_2 is limiting ($0.0175 < 0.02083$). Mg is in excess.

Mg consumed: $n_{\text{Mg used}} = 2 \times 0.0175 = 0.035 \text{ mol} = 0.035 \times 24 = 0.84$ g

Mg excess: $1.0 - 0.84 = \mathbf{0.16}$ g

Approach / Analogy

After finding O_2 is limiting, find how much Mg is actually consumed (using the stoichiometric ratio with O_2 consumed), then subtract from initial Mg. Excess = initial − consumed.

Common Mistake

Finding the limiting reagent but then stopping without computing the excess. The question has two parts: (1) which is in excess, (2) how much. Both must be answered. Mg excess = $1.0 - 0.84 = 0.16$ g.

Answer

(1) Mg, 0.16 g

Q.36 $\text{Li}_2\text{O(s)} + \text{H}_2\text{O(g)} \longrightarrow 2\text{LiOH(s)}$. If 60 kg of water and 45 kg of Li_2O are present in a shuttle, which statements are correct? [NSEC-2017]

Explanation

$$n(\text{H}_2\text{O}) = \frac{60000}{18} = 3333.3 \text{ mol}, \quad n(\text{Li}_2\text{O}) = \frac{45000}{30} = 1500 \text{ mol}$$

$$\frac{3333.3}{1} = 3333.3, \quad \frac{1500}{1} = 1500$$

Li_2O is limiting ($1500 < 3333.3$) — Statement II: **TRUE**

Water remaining:

$$\begin{aligned} \text{H}_2\text{O consumed} &= 1500 \text{ mol} = 1500 \times 18 = 27{,}000 \text{ g} = 27 \text{ kg} \\ \text{H}_2\text{O remaining} &= 60 - 27 = 33 \text{ kg} \end{aligned}$$

Statement IV says 27 kg remains — checking: 60 - 27 = 33 kg remains, so 27 kg was *consumed*, not remaining. Statement IV is incorrect as given.

Li_2O required to remove all 60 kg water:

$$n(\text{Li}_2\text{O}) = 3333.3 \text{ mol} = 3333.3 \times 30 = 100{,}000 \text{ g} = 100 \text{ kg}$$

Statement III (100 kg Li_2O required): **TRUE**

Answer: II and III — option (4).

Approach / Analogy

Two separate sub-questions here: (1) which is limiting, and (2) what happens to the water. Li_2O is limiting (only 45 kg of 100 kg needed). 27 kg water is consumed (not 27 kg remaining). 33 kg water stays unabsorbed.

Common Mistake

Confusing "water consumed" with "water remaining." 27 kg water is consumed by 45 kg Li_2O; this means $60 - 27 = 33$ kg water remains unremoved. Statement IV (27 kg remains) is wrong by this calculation.

Answer

(4) II and III (Li_2O will be the limiting reagent; 100 kg Li_2O required)

Q.37 If 1.6 g of SO_2 and 1.5×10^{22} molecules of H_2S are mixed: $2H_2S + SO_2 \longrightarrow 3S + 2H_2O$. Which statement is true?

Explanation

$$n(\text{SO}_2) = \frac{1.6}{64} = 0.025 \text{ mol}$$
$$n(\text{H}_2\text{S}) = \frac{1.5 \times 10^{22}}{6.022 \times 10^{23}} = 0.0249 \approx 0.025 \text{ mol}$$

Required ratio: $H_2S : SO_2 = 2 : 1$

For 0.025 mol SO_2, H_2S needed = $2 \times 0.025 = 0.05$ mol.

Available H_2S = 0.025 mol < 0.05 mol needed. **H_2S is limiting.**

$$\frac{n_{\text{H}_2\text{S}}}{2} = \frac{0.025}{2} = 0.0125, \quad \frac{n_{\text{SO}_2}}{1} = 0.025$$

H_2S limits. SO_2 is in excess. Statement (3): **SO_2 will remain in excess.**

Approach / Analogy

The numbers 1.5×10^{22} look scary but divide by Avogadro's number: $1.5 \times 10^{22}/6 \times 10^{23} \approx 0.025$ mol. Equal moles of both, but H_2S needs coefficient 2 — with equal moles it runs out first. SO_2 remains in excess.

Common Mistake

Not converting molecules to moles and comparing raw particle counts with gram amounts. Always convert everything to moles before applying stoichiometry. Molecules and grams cannot be compared directly.

Answer

(3) SO_2 will remain in excess

TYPE II(ii) : Limiting Reagent — Volume of gases

Key: At the same T and P, volumes of gases are directly proportional to moles (Avogadro's law). So for gas-phase reactions, the stoichiometric coefficients represent volume ratios directly — you can work with volumes exactly like moles.

Q.38 When 22.4 L of $H_2(g)$ is mixed with 11.2 L of $Cl_2(g)$ at STP, the moles of HCl(g) formed is equal to:

Explanation

Reaction: $H_2 + Cl_2 \rightarrow 2HCl$
At STP: 22.4 L = 1 mol, 11.2 L = 0.5 mol.

$$n(\mathrm{H_2}) = 1 \text{ mol}, \quad n(\mathrm{Cl_2}) = 0.5 \text{ mol}$$

$$\frac{1}{1} = 1, \quad \frac{0.5}{1} = 0.5$$

Cl_2 is limiting. $n(\mathrm{HCl}) = 2 \times 0.5 = \mathbf{1}$ mol HCl

Approach / Analogy

H_2:Cl_2:HCl = 1:1:2. Cl_2 (0.5 mol) limits. HCl formed = $2 \times 0.5 = 1$ mol. Using volumes directly: 11.2 L Cl_2 gives $2 \times 11.2 = 22.4$ L HCl = 1 mol.

Answer

(1) 1 mol of HCl(g)

Q.39 At NTP, 100 mL N_2 and 100 mL of H_2 are mixed together. $N_2 + 3H_2 \longrightarrow 2NH_3$. The produced volume of NH_3 is:

Explanation

At constant T & P, volumes act as moles.

$$\frac{V_{\mathrm{N_2}}}{1} = 100, \quad \frac{V_{\mathrm{H_2}}}{3} = \frac{100}{3} = 33.3$$

H_2 is limiting ($33.3 < 100$). Scale = 33.3 mL.

$$V(\mathrm{NH_3}) = 2 \times 33.3 = \mathbf{66.6} \text{ mL}$$

Approach / Analogy

Same limiting reagent method but with volumes instead of moles. H_2 needs coefficient 3, so it runs out first even though both volumes are 100 mL. NH_3 formed = $\frac{2}{3} \times 100 = 66.6$ mL.

Common Mistake

Adding volumes or averaging: $(100 + 100)/2 = 100$ mL. Volume of products must be found by stoichiometry from the limiting reagent, not by averaging reactant volumes.

Answer

(1) 66.6 mL

Q.40 12 L of H_2 and 11.2 L of Cl_2 are mixed and exploded. The composition by volume of mixture is:

Explanation

$H_2 + Cl_2 \rightarrow 2HCl$

At same conditions, use volumes as moles.

$$\frac{12}{1} = 12, \quad \frac{11.2}{1} = 11.2$$

Cl_2 is limiting. Scale = 11.2 L.

$$\begin{aligned} \text{HCl formed} &= 2 \times 11.2 = 22.4\text{ L} \\ \text{H}_2\text{ consumed} &= 11.2\text{ L} \\ \text{H}_2\text{ remaining} &= 12 - 11.2 = 0.8\text{ L} \end{aligned}$$

Final composition: 0.8 L H_2 + 22.4 L HCl

Approach / Analogy

After finding Cl_2 is limiting, calculate: (1) HCl formed from all Cl_2 used, (2) H_2 actually consumed, (3) H_2 remaining = initial − consumed. Final mixture = excess H_2 + product HCl.

Common Mistake

Forgetting to report the excess reactant in the final mixture composition. The question asks for the *composition* of the final mixture — include both the product (22.4 L HCl) and the unreacted excess (0.8 L H_2).

Answer

(3) 0.8 L H_2 and 22.4 L HCl(g)

Q.41 At constant T and P, 5.0 L of SO_2 are reacted with 3.0 L of O_2: $2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$. The volume of the reaction mixture at completion is: [NSEC-2017]

Explanation

$$\frac{5.0}{2} = 2.5, \quad \frac{3.0}{1} = 3.0$$

SO_2 is limiting ($2.5 < 3.0$). Scale = 2.5 L.

$$\begin{aligned} \text{SO}_3\text{ formed} &= 2 \times 2.5 = 5.0\text{ L} \\ \text{O}_2\text{ consumed} &= 1 \times 2.5 = 2.5\text{ L} \\ \text{O}_2\text{ remaining} &= 3.0 - 2.5 = 0.5\text{ L} \end{aligned}$$

Total final volume $= 5.0 + 0.5 =$ **5.5** L

Approach / Analogy

Final mixture = products + unreacted excess. SO_2 fully consumed; O_2 partially consumed (0.5 L excess). Total = 5 L SO_3 + 0.5 L excess O_2 = 5.5 L. At constant T and P, total gas volume changes because moles change.

Common Mistake

Adding initial volumes: $5 + 3 = 8$ L. This gives the initial total, not the final. Gases are consumed and produced — total volume changes with reaction progress. Compute product volumes and add unreacted excess.

Answer

(3) 5.5 L

Q.42 When a mixture of 10 mole of SO_2 and 15 mole of O_2 was passed over catalyst, 8 mole of SO_3 was formed. How many moles of SO_2 and O_2 did not enter into combination? $2SO_2 + O_2 \longrightarrow 2SO_3$

Explanation

8 mol SO_3 was formed. From stoichiometry (backward from product):

$$\begin{aligned} SO_2 \text{ consumed} &= 8 \text{ mol (1:1 with } SO_3) \\ O_2 \text{ consumed} &= \frac{8}{2} = 4 \text{ mol (ratio } O_2{:}SO_3 = 1:2) \end{aligned}$$

Unreacted:

$$\begin{aligned} SO_2 &= 10 - 8 = \mathbf{2} \text{ mol} \\ O_2 &= 15 - 4 = \mathbf{11} \text{ mol} \end{aligned}$$

Approach / Analogy

Work backwards from the product formed (8 mol SO_3) to find what was consumed, then subtract from initial amounts. SO_3:SO_2 = 1:1 (same coefficient), and SO_3:O_2 = 2:1. This is an incomplete reaction — catalyst gave only partial yield.

Common Mistake

First finding the limiting reagent (not needed here!) rather than working from the known product amount. When actual product formed is given, use it directly to back-calculate consumption. The limiting reagent approach is for finding the *maximum* product.

Answer

(1) 2 moles of SO_2, 11 moles of O_2

TYPE II(iii) : Limiting Reagent — Three Reactants

Q.43 In the reaction $4A + 2B + 3C \longrightarrow A_4B_2C_3$, what will be the number of moles of product formed, starting from 2 mol of A, 1.2 mol of B and 1.44 mol of C?

Explanation

Divide by coefficients:

$$\frac{n_A}{4} = \frac{2}{4} = 0.5$$
$$\frac{n_B}{2} = \frac{1.2}{2} = 0.6$$
$$\frac{n_C}{3} = \frac{1.44}{3} = 0.48$$

Smallest = 0.48 (C is limiting). Scale factor = 0.48.

$n(\text{product}) = 1 \times 0.48 = \mathbf{0.48}$ mol

Approach / Analogy

Three reactants — same rule, extended. Divide each by its coefficient. C gives the smallest (0.48), so C limits. Product coefficient = 1, so product moles = scale factor = 0.48 mol.

Common Mistake

Taking the smallest raw mole value (1.2 for B) as the limiting reagent without dividing by coefficients first. B's raw moles are smallest, but after dividing by coefficient 2, B gives 0.6 — larger than C's 0.48. **Always divide by coefficient before comparing.**

Answer

(3) 0.48

Q.44 Calculate the mass of sucrose $C_{12}H_{22}O_{11}$(s) produced by mixing 78 g of C(s), 11 g of H_2(g) and 67.2 litre of O_2(g) at 0°C and 1 atm: $12C + 11H_2 + \frac{11}{2}O_2 \longrightarrow C_{12}H_{22}O_{11}$

Explanation

$$n(\mathrm{C}) = \frac{78}{12} = 6.5 \text{ mol}$$
$$n(\mathrm{H_2}) = \frac{11}{2} = 5.5 \text{ mol}$$
$$n(\mathrm{O_2}) = \frac{67.2}{22.4} = 3 \text{ mol}$$

Divide by coefficients:

$$\frac{6.5}{12} = 0.542, \quad \frac{5.5}{11} = 0.5, \quad \frac{3}{5.5} = 0.545$$

H_2 is limiting ($0.5 < 0.542 < 0.545$). Scale = 0.5.

$$M_r(\text{sucrose}) = 12(12) + 22(1) + 11(16) = 144 + 22 + 176 = 342 \text{ g/mol}$$
$$n(\text{sucrose}) = 1 \times 0.5 = 0.5 \text{ mol}$$
$$\text{Mass} = 0.5 \times 342 = \mathbf{171} \text{ g}$$

Approach / Analogy

Three reactants with fractional coefficient ($\frac{11}{2}$ for O_2). For O_2: divide 3 by 5.5 (= 11/2) to get 0.545. H_2 gives the smallest ratio (0.5) — it limits. Sucrose = 0.5 mol = 171 g.

Common Mistake

Using 11/2 = 5.5 as the coefficient but then accidentally writing it as 2/11, flipping the division. Write it clearly: $\frac{n_{O_2}}{11/2} = \frac{3\times2}{11} = \frac{6}{11} = 0.545$.

Answer

(1) 171 g

Q.45 $2KI + I_2 + 22HNO_3 \longrightarrow 2HIO_3 + 2KIO_3 + 22NO_2 + 10H_2O$ If 3 mol of KI and 2 mol I_2 are reacted with excess of HNO_3, volume of NO_2 gas evolved at NTP is:

Explanation

HNO_3 is in excess — ignore it.

$$\frac{n_{KI}}{2} = \frac{3}{2} = 1.5, \quad \frac{n_{I_2}}{1} = \frac{2}{1} = 2$$

KI is limiting ($1.5 < 2$). Scale = 1.5.

$$n(NO_2) = 22 \times 1.5 = 33 \text{ mol}$$

$$V(NO_2) = 33 \times 22.4 = \mathbf{739.2} \text{ L}$$

Approach / Analogy

Big equation — stay calm. Identify the non-excess reactants (KI and I_2), divide by their coefficients, find the smaller one (KI = 1.5). Scale = 1.5. NO_2 has coefficient 22, so: $22 \times 1.5 = 33$ mol $\times 22.4$ L/mol = 739.2 L.

Common Mistake

Being overwhelmed by the complex equation and using I_2 as the limiting reagent (since I_2 has fewer raw moles = 2 vs 3 for KI). After dividing by coefficients: KI gives 3/2 = 1.5, I_2 gives 2/1 = 2. KI is limiting, not I_2.

Answer

(1) 739.2 L

TYPE III(i) : Combustion of Gas — Basic

Combustion Formula for Hydrocarbons:

$$C_xH_y + \left(x + \frac{y}{4}\right) O_2 \rightarrow xCO_2 + \frac{y}{2}H_2O$$

At constant T & P: volumes of gases are in the same ratio as moles (stoichiometric coefficients).

Q.46 For complete combustion of 3 g ethane, the produced volume of CO_2 at STP is: $2C_2H_6 + 7O_2 \longrightarrow 4CO_2$

Explanation

$$\begin{aligned} M_r(C_2H_6) &= 30 \text{ g/mol} \\ n(C_2H_6) &= \frac{3}{30} = 0.1 \text{ mol} \\ \text{Ratio: } C_2H_6 : CO_2 &= 2:4 = 1:2 \\ n(CO_2) &= 0.1 \times 2 = 0.2 \text{ mol} \\ V &= 0.2 \times 22.4 = \mathbf{4.48} \text{ L} \end{aligned}$$

Approach / Analogy

Each C_2H_6 molecule has 2 carbons. Burning 1 mol C_2H_6 gives 2 mol CO_2 (one CO_2 per carbon). 0.1 mol ethane → 0.2 mol CO_2 → 4.48 L. Ethane = C_2H_6, molar mass = 2(12)+6 = 30.

Common Mistake

Using ratio 2:4 as 1:4 instead of simplifying to 1:2. 2:4 reduces to 1:2. But even without simplifying: 0.1 mol C_2H_6 $\times(4/2) = 0.2$ mol CO_2. Either way gives the same answer.

Answer

(1) 4.48 L

Q.47 For complete combustion of 1.12 L of butane (C_4H_{10}) at STP, the produced volume of $H_2O(g)$ at STP is: $C_4H_{10} + \frac{13}{2}O_2 \longrightarrow 4CO_2 + 5H_2O$

Explanation

$$\begin{aligned} n(C_4H_{10}) &= \frac{1.12}{22.4} = 0.05 \text{ mol} \\ \text{Ratio: } C_4H_{10} : H_2O &= 1:5 \\ n(H_2O) &= 0.05 \times 5 = 0.25 \text{ mol} \\ V(H_2O) &= 0.25 \times 22.4 = \mathbf{5.6} \text{ L} \end{aligned}$$

Approach / Analogy

Butane (C_4H_{10}) has 10 H atoms → 5 H_2O molecules per molecule burned (ratio 1:5). 0.05 mol butane gives $5 \times 0.05 = 0.25$ mol H_2O = 5.6 L. Water is treated as gas here (at STP conditions).

Answer

(1) 5.6 L

Q.48 For complete combustion of 3 g ethane, the required volume of O_2 at STP is: $2C_2H_6 + 7O_2 \longrightarrow 4CO_2 +$

Explanation

From Q.46: $n(C_2H_6) = 0.1$ mol.

$$\text{Ratio: } C_2H_6 : O_2 = 2 : 7$$

$$n(O_2) = 0.1 \times \frac{7}{2} = 0.35 \text{ mol}$$

$$V = 0.35 \times 22.4 = \mathbf{7.84} \text{ L}$$

Approach / Analogy

Same 3 g ethane as Q.46 but asking for O_2 needed. Ratio C_2H_6:O_2 = 2:7, so $n(O_2) = 0.1 \times 7/2 = 0.35$ mol = 7.84 L.

Answer

(1) 7.84 L

Q.49 For complete combustion of 5 mol propane (C_3H_8), the required volume of O_2 at STP is: $C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O$

Explanation

$$\text{Ratio: } C_3H_8 : O_2 = 1 : 5$$

$$n(O_2) = 5 \times 5 = 25 \text{ mol}$$

$$V = 25 \times 22.4 = \mathbf{560} \text{ L}$$

Approach / Analogy

Propane + $5O_2$ is given directly. 5 mol propane needs $5 \times 5 = 25$ mol O_2 = 560 L. Clean calculation.

Answer

(1) 560 L

Q.50 How many litres of oxygen at 1 atm and 273 K will be required to burn completely 2.2 g of propane (C_3H_8)?

Explanation

$$M_r(\mathrm{C_3H_8}) = 44 \text{ g/mol}$$
$$n(\mathrm{C_3H_8}) = \frac{2.2}{44} = 0.05 \text{ mol}$$
$$n(\mathrm{O_2}) = 5 \times 0.05 = 0.25 \text{ mol}$$
$$V = 0.25 \times 22.4 = \mathbf{5.6}\text{ L}$$

Approach / Analogy

273 K and 1 atm = standard STP conditions, so use 22.4 L/mol. 2.2 g propane = 0.05 mol → 0.25 mol O_2 needed = 5.6 L.

Answer

(3) 5.6 L

Q.51 What volume of oxygen gas (O_2) measured at 0°C and 1 atm is needed to burn completely 1 L of propane gas (C_3H_8) measured under the same conditions?

Explanation

At same T and P: volumes are proportional to moles. Use volume ratios directly.

$$\mathrm{C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O}$$

Volume ratio: C_3H_8 : O_2 = 1 : 5.

$$V(\mathrm{O_2}) = 1\text{ L} \times 5 = \mathbf{5}\text{ L}$$

Approach / Analogy

Same conditions ⇒ use volume ratios = mole ratios. No conversion to moles needed! 1 L propane burns with 5 L O_2. This is the elegance of gas stoichiometry at equal conditions.

Common Mistake

Converting to moles (1 L / 22.4 = 0.0446 mol) then back to volume. Unnecessary when conditions are the same — volume ratios equal mole ratios at constant T and P.

Answer

(1) 5 L

Q.52 Assuming petrol is octane (C_8H_{18}) with density 0.8 g mL^{-1}, 1.425 L of petrol on complete combustion will consume:

Explanation

Combustion: $C_8H_{18} + \frac{25}{2}O_2 \to 8CO_2 + 9H_2O$

$$\text{Mass of petrol} = 1425\ \text{mL} \times 0.8\ \text{g/mL} = 1140\ \text{g}$$

$$n(C_8H_{18}) = \frac{1140}{114} = 10\ \text{mol} \quad (M_r = 8(12) + 18 = 114)$$

$$n(O_2) = 10 \times \frac{25}{2} = 10 \times 12.5 = \mathbf{125}\ \text{mol}$$

Approach / Analogy

Two-step: (1) use density to get mass from volume ($1425 \times 0.8 = 1140$ g), (2) convert to moles ($1140/114 = 10$ mol C_8H_{18}). Then apply combustion stoichiometry: 1 mol octane needs 12.5 mol O_2.

Common Mistake

Forgetting to use density to convert volume to mass. Directly computing $1425/114$ (using volume as if it were mass in grams) gives a wrong number of moles. Volume × density = mass.

Answer

(3) 125 mole of O_2

Q.53 The number of litres of air required to burn 8 litres of C_2H_2 is approximately:

Explanation

Combustion of acetylene: $2C_2H_2 + 5O_2 \to 4CO_2 + 2H_2O$

Ratio: $C_2H_2 : O_2$ = 2:5.

$$V(O_2) = 8 \times \frac{5}{2} = 20\ \text{L}$$

Air is approximately 20% O_2 (1/5 by volume).

$$V(\text{air}) = 20 \times 5 = \mathbf{100}\ \text{L}$$

Approach / Analogy

Two steps: (1) find O_2 needed from stoichiometry (20 L), (2) scale up to air: air = 5 × O_2 (since air is 20% O_2 by volume). This is the standard "air" conversion in combustion problems.

Common Mistake

Forgetting to convert from O_2 volume to air volume. The question asks for *air*, not O_2. Air is only 20% (1/5) O_2, so multiply O_2 volume by 5 to get the air volume required.

Answer

(4) 100 L

TYPE III(ii) : Combustion — Uncombined gases

Q.54 If 8 mL of uncombined O_2 remain after exploding O_2 with 4 mL of hydrogen, the number of mL of O_2 originally were:

Explanation

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O$

O_2 consumed to burn 4 mL H_2:

$$V(O_2)_{\text{consumed}} = \frac{1}{2} \times 4 = 2 \text{ mL}$$

$$V(O_2)_{\text{original}} = 2 + 8 = \mathbf{10} \text{ mL}$$

Approach / Analogy

O_2 original = O_2 consumed + O_2 remaining. Consumed = H_2 volume ÷ 2 (ratio H_2:O_2 = 1:0.5). Remaining = 8 mL. So original = 2 + 8 = 10 mL.

Common Mistake

Writing O_2 consumed = 4 mL (using 1:1 ratio with H_2). The ratio H_2:O_2 = 1:$\frac{1}{2}$, so 4 mL H_2 consumes only 2 mL O_2, not 4 mL.

Answer

(3) 10 mL

Q.55 1 L of CO_2 is passed over hot coke. When the volume of reaction mixture becomes 1.4 L, the composition of reaction mixture is:

Explanation

Reaction: $CO_2 + C \rightarrow 2CO$

Let x L of CO_2 react with coke:

$$CO_2 \text{ remaining} = (1 - x) \text{ L}$$
$$CO \text{ formed} = 2x \text{ L}$$

Total volume:

$$(1 - x) + 2x = 1 + x = 1.4 \implies x = 0.4 \text{ L}$$

Composition:

$$CO_2 = 1 - 0.4 = 0.6 \text{ L}$$
$$CO = 2 \times 0.4 = 0.8 \text{ L}$$

Approach / Analogy

Volume increases because 1 mol CO_2 (1 L) gives 2 mol CO (2 L) — net gain of 1 L per mole reacted. Let x L react: total = $1 + x = 1.4 \to x = 0.4$ L. Then find remaining CO_2 and CO formed from $x = 0.4$.

Common Mistake

Assuming all 1 L CO_2 reacts. If all reacted: CO formed = 2 L, total = 2 L $\neq$ 1.4 L. Only part of the CO_2 reacts. Set up the volume balance equation to find exactly how much reacted.

Answer

(3) 0.6 L CO_2 and 0.8 L CO

Q.56 **26 cc of CO_2 are passed over red hot coke. The volume of CO evolved is:**

Explanation

$CO_2 + C \to 2CO$

Assuming all CO_2 reacts (red hot coke $\Rightarrow$ complete reaction):

$$V(\text{CO}) = 2 \times 26 = \mathbf{52}\text{ cc}$$

Approach / Analogy

At red hot temperatures, the reaction goes to completion. 1 mol CO_2 gives 2 mol CO. So 26 cc CO_2 gives $2 \times 26 = 52$ cc CO. Simple 1:2 ratio.

Common Mistake

Using a 1:1 ratio (getting 26 cc CO). The balanced equation clearly shows 1 CO_2 gives 2 CO — the ratio is 1:2, not 1:1.

Answer

(4) 52 cc

TYPE III(iii) : Combustion — Determining molecular formula from gas volumes

General combustion equation for hydrocarbon C_xH_y:

$$C_xH_y + \left(x + \frac{y}{4}\right) O_2 \to xCO_2 + \frac{y}{2}H_2O$$

From the volumes (at same T, P): ratios of volumes = ratios of coefficients.
From volume of CO_2: get x. From volume of H_2O: get y.

Q.57 **10 mL of gaseous hydrocarbon on combustion give 40 mL of CO_2(g) and 50 mL of H_2O(vap.). The hydrocarbon is:**

Explanation

Using volume ratios (all at same conditions):

$$C_xH_y : CO_2 : H_2O = 10 : 40 : 50 = 1 : 4 : 5$$

From CO_2: $x = 4$ (4 mol CO_2 per mol hydrocarbon)
From H_2O: $\frac{y}{2} = 5 \implies y = 10$
Formula: $\mathbf{C_4H_{10}}$ (butane)

Approach / Analogy

Divide all volumes by the hydrocarbon volume (10 mL) to get mole ratios per mole of hydrocarbon. CO_2 ratio = 4 → 4 carbons. H_2O ratio = 5 → 10 hydrogens (each H_2O gives 2H). Formula = C_4H_{10}.

Common Mistake

Reading H from the ratio directly: H_2O ratio = 5 → H = 5. Wrong! Each water molecule has 2 H atoms, so H = $2 \times 5 = 10$. Always double the H_2O volume ratio to get the number of H atoms.

Answer

(4) C_4H_{10}

Q.58 500 mL of a gaseous hydrocarbon when burnt in excess of O_2 gave 2.5 L of CO_2 and 3.0 L of water vapours under same conditions.

Explanation

Divide by 500 mL:

$$CO_2 : C_xH_y = 2500 : 500 = 5 : 1 \implies x = 5$$

$$H_2O : C_xH_y = 3000 : 500 = 6 : 1 \implies \frac{y}{2} = 6 \implies y = 12$$

Formula: $\mathbf{C_5H_{12}}$ (pentane)

Approach / Analogy

500 mL hydrocarbon gives 2500 mL CO_2 (ratio 1:5) and 3000 mL H_2O (ratio 1:6). $x = 5$ carbons from CO_2 ratio. $y = 2 \times 6 = 12$ hydrogens from H_2O ratio. C_5H_{12} = pentane.

Answer

(4) C_5H_{12}

Q.59 11.2 litre of a hydrocarbon at STP produces 44.8 litre of CO_2 at STP and 36 g of H_2O during its combustion. The molecular formula of the hydrocarbon is:

Explanation

$$n(\text{hydrocarbon}) = \frac{11.2}{22.4} = 0.5 \text{ mol}$$

$$n(\text{CO}_2) = \frac{44.8}{22.4} = 2 \text{ mol} \implies x = \frac{2}{0.5} = 4$$

$$n(\text{H}_2\text{O}) = \frac{36}{18} = 2 \text{ mol} \implies \frac{y}{2} = \frac{2}{0.5} = 4 \implies y = 8$$

Formula: C_4H_8

Approach / Analogy

Mixed question: CO_2 is a gas (convert volume to moles), H_2O is a liquid (convert mass to moles). Per mole of hydrocarbon: CO_2 ratio = 2/0.5 = 4 (so x=4); H_2O ratio = 2/0.5 = 4 (so y = 8). Formula C_4H_8.

Common Mistake

Trying to use 36 g of H_2O as a volume. Water is given as a mass here (36 g), not a volume. Convert mass to moles: 36/18 = 2 mol. Don't treat grams as litres!

Answer

(2) C_4H_8

TYPE IV : % Purity

Key Formulas:

$$\% \text{ purity} = \frac{\text{Mass of pure compound}}{\text{Mass of impure sample}} \times 100$$

Strategy: Work backwards from product to find mass of pure reactant. Then divide by sample mass.

Q.60 200 g impure $CaCO_3$ on heating gives 11.35 L CO_2 gas at STP. Find the percentage of calcium in the limestone sample.

Explanation

$CaCO_3 \rightarrow CaO + CO_2$

$$n(CO_2) = \frac{11.35}{22.4} = 0.5067 \approx 0.5 \text{ mol}$$
$$n(CaCO_3) = 0.5 \text{ mol (1:1 ratio)}$$
$$\text{Mass of pure } CaCO_3 = 0.5 \times 100 = 50 \text{ g}$$

% purity of $CaCO_3$:

$$= \frac{50}{200} \times 100 = 25\%$$

But the question asks for % **calcium** (not $CaCO_3$):

$$\text{Mass of Ca in sample} = 0.5 \times 40 = 20 \text{ g}$$
$$\% \text{ Ca} = \frac{20}{200} \times 100 = \mathbf{10}\%$$

Approach / Analogy

Read carefully: the question asks for % of **calcium** (Ca), not % of $CaCO_3$. Same moles, but multiply by mass of Ca (40) instead of $CaCO_3$ (100). 0.5 mol Ca in 200 g sample = $\frac{20}{200} \times 100 = 10\%$.

Common Mistake

Answering 25% (the % purity of $CaCO_3$) instead of 10% (the % of Ca element). The question specifically says "percentage of calcium" — use the molar mass of Ca (40), not $CaCO_3$ (100).

Answer

(1) 10%

Q.61 20.0 g of a magnesium carbonate sample decomposes on heating to give carbon dioxide and 8.0 g magnesium oxide. What will be the percentage purity of magnesium carbonate in the sample?

Explanation

$MgCO_3 \rightarrow MgO + CO_2$

Molar ratio: $MgCO_3$: MgO = 1 : 1

$$n(MgO) = \frac{8.0}{40} = 0.2 \text{ mol}$$
$$n(MgCO_3) = 0.2 \text{ mol}$$
$$M_r(MgCO_3) = 24 + 12 + 48 = 84 \text{ g/mol}$$
$$\text{Mass of pure } MgCO_3 = 0.2 \times 84 = 16.8 \text{ g}$$

$$\% \text{ purity} = \frac{16.8}{20.0} \times 100 = \mathbf{84}\%$$

Approach / Analogy

Work backwards: 8 g MgO = 0.2 mol MgO = 0.2 mol $MgCO_3$ (1:1 ratio). Convert back to mass of pure $MgCO_3$: $0.2 \times 84 = 16.8$ g. Purity = 16.8/20 = 84%.

Common Mistake

Using $M_r(\text{MgO}) = 24 + 16 = 40$ (correct) but then forgetting to convert to $MgCO_3$ mass and instead reporting $8/20 \times 100 = 40\%$ (the % MgO in the sample, not % $MgCO_3$ in the original sample).

Answer

(2) 84

Q.62 $CaCO_3$ is 90% pure. Volume of CO_2 collected at STP when 10 g of $CaCO_3$ is decomposed is:

Explanation

$$\begin{aligned} \text{Mass of pure } CaCO_3 &= 90\% \times 10 = 9 \text{ g} \\ n(CaCO_3) &= \frac{9}{100} = 0.09 \text{ mol} \\ n(CO_2) &= 0.09 \text{ mol (1:1)} \\ V &= 0.09 \times 22.4 = \mathbf{2.016} \text{ L} \end{aligned}$$

Approach / Analogy

Purity of 90% means only 90% of the sample is $CaCO_3$. Pure $CaCO_3$ = 0.9 × 10 = 9 g. Only the pure part reacts. Impurity doesn't contribute CO_2.

Common Mistake

Using all 10 g of the sample as if it were pure $CaCO_3$: $10/100 = 0.1$ mol → 2.24 L. This ignores the 10% impurity. Always calculate mass of pure compound = mass × (purity/100) first.

Answer

(1) 2.016 L

Q.63 How much amount of $CaCO_3$ in gram having percentage purity 50% produces 0.56 litre of CO_2 at STP on heating?

Explanation

$$n(\text{CO}_2) = \frac{0.56}{22.4} = 0.025 \text{ mol}$$
$$n(\text{CaCO}_3)_{\text{pure}} = 0.025 \text{ mol (1:1)}$$
$$\text{Mass of pure CaCO}_3 = 0.025 \times 100 = 2.5 \text{ g}$$

50% purity means pure $CaCO_3$ = 50% of total sample.

$$\text{Mass of impure sample} = \frac{2.5}{0.5} = \mathbf{5} \text{ g}$$

Approach / Analogy

Reverse purity problem: you need 2.5 g pure $CaCO_3$, but the sample is only 50% pure, so total sample = $2.5/0.50 = 5$ g. Like needing 50 g of gold from an alloy that's 50% gold — you need 100 g of alloy.

Common Mistake

Reporting 2.5 g (the pure $CaCO_3$ needed) instead of 5 g (the total impure sample required). Always divide by the purity fraction to get the total sample mass when purity $<$ 100%.

Answer

(1) 5 g

Q.64 10 g impure NaOH is completely neutralised by 1000 mL of $\frac{1}{10}$ N HCl. The percentage purity of the impure NaOH is:

Explanation

For acid-base neutralisation: meq of acid = meq of base.

$$\text{meq of HCl} = N \times V(\text{mL}) = \frac{1}{10} \times 1000 = 100 \text{ meq}$$

n-factor of NaOH = 1, so meq = mmol.

$$\text{moles of NaOH} = \frac{100}{1000} = 0.1 \text{ mol}$$
$$\text{Mass of pure NaOH} = 0.1 \times 40 = 4 \text{ g}$$

$$\% \text{ purity} = \frac{4}{10} \times 100 = \mathbf{40}\%$$

Approach / Analogy

Use equivalents: meq acid = meq base. HCl gives 100 meq. NaOH (n-factor = 1): meq = mmol, so 100 meq = 100 mmol = 0.1 mol = 4 g. Purity = 4/10 = 40%.

Common Mistake

Confusing N × V(mL) = milliequivalents with N × V(L) = equivalents. Here 1000 mL is used directly: meq = $\frac{1}{10} \times 1000 = 100$ meq (correct using mL). Or: $\frac{1}{10} \times 1 = 0.1$ eq = 0.1 mol (using L). Both work — just stay consistent with units.

Answer

(1) 40%

Q.65 50 g $CaCO_3$ will react with …g of 20% HCl by weight.

Explanation

Reaction: $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$

$$n(\text{CaCO}_3) = \frac{50}{100} = 0.5 \text{ mol}$$

$$n(\text{HCl}) = 2 \times 0.5 = 1 \text{ mol (ratio 1:2)}$$

$$\text{Mass of pure HCl} = 1 \times 36.5 = 36.5 \text{ g}$$

HCl solution is 20% by weight:

$$\text{Mass of HCl solution} = \frac{36.5}{0.20} = \mathbf{182.5} \text{ g}$$

Approach / Analogy

Two steps: (1) find mass of pure HCl needed (36.5 g), (2) since solution is only 20% HCl, divide by 0.20 to get total solution mass needed ($36.5/0.2 = 182.5$ g). Like needing 36.5 g of pure salt from a 20% brine — you need 182.5 g of brine.

Common Mistake

Reporting 36.5 g as the answer (mass of pure HCl, not the HCl solution). The question asks for grams of the **20% HCl solution**, not pure HCl. Divide by purity fraction: $36.5/0.20 = 182.5$ g.

Answer

(4) 182.5 g

TYPE V : % Yield

Key Formula:

$$\% \text{ yield} = \frac{\text{Actual yield (obtained)}}{\text{Theoretical yield (maximum possible)}} \times 100$$

Strategy: First find theoretical yield from limiting reagent, then compare with actual.

Q.66 If 5.4 g of Al reacts with 12.8 g S and gives 12 g of aluminium sulphide, then the percent yield is: $2Al + 3S \longrightarrow Al_2S_3$

Explanation

Find limiting reagent:

$$n(\text{Al}) = \frac{5.4}{27} = 0.2 \text{ mol}, \quad n(\text{S}) = \frac{12.8}{32} = 0.4 \text{ mol}$$

$$\frac{0.2}{2} = 0.1, \quad \frac{0.4}{3} = 0.133$$

Al is limiting ($0.1 < 0.133$). Scale = 0.1.

Theoretical yield of Al_2S_3:

$$M_r(\text{Al}_2\text{S}_3) = 54 + 96 = 150 \text{ g/mol}$$

$$n(\text{Al}_2\text{S}_3) = 1 \times 0.1 = 0.1 \text{ mol}$$

$$\text{Mass}_{\text{theoretical}} = 0.1 \times 150 = 15 \text{ g}$$

$$\% \text{ yield} = \frac{12}{15} \times 100 = \mathbf{80}\%$$

Approach / Analogy

Two steps: (1) find maximum possible yield from limiting reagent (theoretical = 15 g), (2) compare actual (12 g) to theoretical. Yield = 12/15 = 80%. Al limits ($0.1 < 0.133$ after dividing by coefficients).

Common Mistake

Computing % yield before finding the limiting reagent. Using the wrong reagent (S) as limiting gives a different theoretical yield and wrong % yield. Always identify the limiting reagent first, then compute theoretical yield from it.

Answer

(1) 80%

Q.67 Calculate % yield if 200 g $KHCO_3$ produces 22 g of CO_2 upon strong heating. $2KHCO_3 \rightarrow K_2CO_3 + H_2$

Explanation

$$M_r(\text{KHCO}_3) = 39 + 1 + 12 + 48 = 100 \text{ g/mol}$$

$$n(\text{KHCO}_3) = \frac{200}{100} = 2 \text{ mol}$$

$$\text{Ratio: KHCO}_3 : \text{CO}_2 = 2 : 1$$

$$n(\text{CO}_2)_{\text{theoretical}} = \frac{2}{2} = 1 \text{ mol}$$

$$\text{Mass}_{\text{theoretical}} = 1 \times 44 = 44 \text{ g}$$

$$\% \text{ yield} = \frac{22}{44} \times 100 = \mathbf{50}\%$$

Approach / Analogy

$KHCO_3$ has molar mass = 100 g/mol (nice round number). 200 g = 2 mol $KHCO_3 \rightarrow$ 1 mol CO_2 theoretically (44 g). Actual = 22 g. Yield = 22/44 = 50%.

Answer

(1) 50%

Q.68 In an experiment, 349 g of ethyl propanoate was obtained from 250 g of ethanol (propanoic acid in excess). The percentage yield is: [NSEC-2007]

Explanation

$C_2H_5OH + C_2H_5COOH \rightarrow C_2H_5COOC_2H_5 + H_2O$

Ethanol is the limiting reagent (propanoic acid is excess).

$$n(\mathrm{C_2H_5OH}) = \frac{250}{46} = 5.435 \text{ mol}$$

$$\text{Ratio: ethanol : ester} = 1 : 1$$

$$n(\text{ester})_{\text{theoretical}} = 5.435 \text{ mol}$$

$$\text{Mass}_{\text{theoretical}} = 5.435 \times 102 = 554.4 \text{ g}$$

$$\% \text{ yield} = \frac{349}{554.4} \times 100 = \mathbf{62.97} \approx \mathbf{62.9}\%$$

Approach / Analogy

1:1 esterification reaction. Ethanol limits (propanoic acid in excess). Max possible ester = $5.435 \times 102 = 554.4$ g. Actual = 349 g. Yield = 349/554.4 = 62.9%.

Common Mistake

Computing yield as $349/250 \times 100 = 139.6\%$ — treating the ethanol mass as the theoretical product mass. Theoretical yield must be the *product* mass (ester, $M_r = 102$), not the *reactant* mass. Always convert moles through the reaction before comparing.

Answer

(2) 62.9

TYPE VI : Sequential Reactions

Strategy: Work backwards from the final product to find what's needed in each previous step. The product of step n is the reactant of step $n + 1$.

Q.69 Minimum amount of Ag_2CO_3(s) required to produce sufficient oxygen for complete combustion of C_2H_2 which produces 11.2 L of CO_2 at 1 atm and 273 K: [Ag = 108]

Explanation

Reaction 1 (combustion): $C_2H_2 + \frac{5}{2}O_2 \rightarrow 2CO_2 + H_2O$
Reaction 2 (decomposition): $Ag_2CO_3 \rightarrow 2Ag + CO_2 + \frac{1}{2}O_2$

Step 1: Find moles of C_2H_2 burned.

$$n(CO_2) = \frac{11.2}{22.4} = 0.5 \text{ mol}, \quad n(C_2H_2) = \frac{0.5}{2} = 0.25 \text{ mol}$$

Step 2: O_2 needed for combustion.

$$n(O_2) = \frac{5}{2} \times 0.25 = 0.625 \text{ mol}$$

Step 3: Ag_2CO_3 needed to produce 0.625 mol O_2.

$$Ag_2CO_3 : O_2 = 1 : \frac{1}{2} \implies n(Ag_2CO_3) = 0.625 \times 2 = 1.25 \text{ mol}$$

$$M_r(Ag_2CO_3) = 216 + 60 = 276 \text{ g/mol}$$

$$\text{Mass} = 1.25 \times 276 = \mathbf{345} \text{ g}$$

Approach / Analogy

Work backwards: final product (CO_2) → moles of C_2H_2 → O_2 needed → Ag_2CO_3 needed. The chain is: Ag_2CO_3 makes O_2, O_2 burns C_2H_2, C_2H_2 makes CO_2. Trace the chain from end to start.

Common Mistake

Working step-by-step forward (guessing how much Ag_2CO_3 to start with) instead of working backwards from the given final product. Always start from the known quantity (11.2 L CO_2) and trace back through each reaction.

Answer

(1) 345 g

Q.70 What weight of $CaCO_3$ must be decomposed to produce sufficient CO_2 to convert 21.2 kg of Na_2CO_3 completely into $NaHCO_3$?

Explanation

Reaction 1: $CaCO_3 \rightarrow CaO + CO_2$
Reaction 2: $Na_2CO_3 + CO_2 + H_2O \rightarrow 2NaHCO_3$

Step 1: Moles of Na_2CO_3.

$$M_r(Na_2CO_3) = 106 \text{ g/mol}$$

$$n(Na_2CO_3) = \frac{21200}{106} = 200 \text{ mol}$$

Step 2: CO_2 needed (1:1 ratio with Na_2CO_3):

$$n(CO_2) = 200 \text{ mol}$$

Step 3: $CaCO_3$ needed (1:1 ratio with CO_2):

$$n(CaCO_3) = 200 \text{ mol}$$

$$\text{Mass} = 200 \times 100 = 20{,}000 \text{ g} = \mathbf{20} \text{ kg}$$

Approach / Analogy

Both steps are 1:1 ratios. Chain: $CaCO_3 \xrightarrow{1:1} CO_2 \xrightarrow{1:1} Na_2CO_3$ consumed. So $CaCO_3$ needed = same moles as Na_2CO_3 = 200 mol = 20 kg.

Answer

(2) 20 kg

Q.71 The following process to obtain iodine from oil-field brines. How many grams of $AgNO_3$ are required in the first step for every 254 kg I_2 produced in the third step?

Explanation

Step 1: $NaI + AgNO_3 \rightarrow AgI + NaNO_3$ (1 mol $AgNO_3$ per mol AgI)
Step 2: $2AgI + Fe \rightarrow FeI_2 + 2Ag$ (2 mol AgI per mol FeI_2)
Step 3: $2FeI_2 + 3Cl_2 \rightarrow 2FeCl_3 + 2I_2$ (1 mol FeI_2 per mol I_2)

Work backwards from I_2:

$$n(I_2) = \frac{254000}{254} = 1000 \text{ mol}$$

From Step 3: $FeI_2 : I_2$ = 2:2 = 1:1 $\implies n(FeI_2) = 1000 \text{ mol}$

From Step 2: AgI : FeI_2 = 2:1 $\implies n(AgI) = 2000 \text{ mol}$

From Step 1: $AgNO_3$: AgI = 1:1 $\implies n(AgNO_3) = 2000 \text{ mol}$

$$M_r(AgNO_3) = 108 + 14 + 48 = 170 \text{ g/mol}$$

$$\text{Mass} = 2000 \times 170 = 340{,}000 \text{ g} = \mathbf{340} \text{ kg}$$

Approach / Analogy

Three sequential steps. Work backwards: $I_2 \to FeI_2$ (1:1) $\to$ AgI (1:2, note FeI_2 needs 2 AgI) $\to$ $AgNO_3$ (1:1). The key multiplier is Step 2: each FeI_2 needs 2 AgI, so the moles double going from Step 3 back to Step 1.

Common Mistake

Not tracing the stoichiometry of each step carefully. In Step 2: 2 AgI gives 1 FeI_2, so going backwards: 1 FeI_2 needs 2 AgI (moles double). Missing this doubling gives 1000 mol $AgNO_3$ instead of 2000 mol — half the correct answer.

Answer

(1) 340 kg

TYPE VII : Parallel Reactions

Strategy: Two reactions happen simultaneously. Set up unknowns for moles (or volumes) of each component, write equations based on total amount and product ratios, then solve simultaneously.

Q.72 A gas mixture of 3 L of propane and butane on complete combustion at 25°C produces 10 L of CO_2. The initial composition of propane and butane in the gas mixture is:

Explanation

Let $V_{\text{propane}} = x$ L, $V_{\text{butane}} = (3 - x)$ L.

Combustion equations:

$$C_3H_8 + 5O_2 \to 3CO_2 + 4H_2O \quad (3\ CO_2 \text{ per L propane})$$
$$C_4H_{10} + \tfrac{13}{2}O_2 \to 4CO_2 + 5H_2O \quad (4\ CO_2 \text{ per L butane})$$

CO_2 produced:

$$3x + 4(3 - x) = 10$$

$$3x + 12 - 4x = 10 \implies -x = -2 \implies x = 2 \text{ L propane}$$

$$\text{Butane} = 3 - 2 = 1 \text{ L}$$

% propane = $\frac{2}{3} \times 100 = 66.67\%$, **% butane =** $\frac{1}{3} \times 100 = 33.33\%$

Approach / Analogy

Set up one equation with one unknown (propane volume). CO_2 from propane = 3 times its volume; CO_2 from butane = 4 times its volume. Total = 10 L. Solve: $3x + 4(3 - x) = 10 \to x = 2$ L propane. The “number of carbons = multiplier for CO_2” is the key insight.

Common Mistake

Setting up the equation with CO_2 = total volumes: $x + (3 - x) = 10$. This is wrong because CO_2 produced is NOT the same as volume of fuel burned — it depends on the carbon content (coefficient × volume). Use $3x + 4(3 - x) = 10$.

Answer

(1) 66.67%, 33.33%

Q.73 When a 12 g mixture of carbon and sulphur is burnt in air, a mixture of CO_2 and SO_2 is produced, in which the number of moles of SO_2 is half that of CO_2. The mass of the carbon in the mixture is:

Explanation

Let mass of C = m g, mass of S = $(12 - m)$ g.

$$n(\mathrm{C}) = \frac{m}{12}, \quad n(\mathrm{S}) = \frac{12-m}{32}$$

Reactions: $\mathrm{C} + \mathrm{O_2} \to \mathrm{CO_2}$ (1:1), $\mathrm{S} + \mathrm{O_2} \to \mathrm{SO_2}$ (1:1).

$$n(\mathrm{CO_2}) = \frac{m}{12}, \quad n(\mathrm{SO_2}) = \frac{12-m}{32}$$

Given: $n(\mathrm{SO_2}) = \frac{1}{2}n(\mathrm{CO_2})$:

$$\frac{12-m}{32} = \frac{1}{2} \times \frac{m}{12} = \frac{m}{24}$$

$$24(12-m) = 32m \implies 288 - 24m = 32m \implies 56m = 288 \implies m = \frac{288}{56} = \mathbf{5.14}\text{ g}$$

Approach / Analogy

Each C makes 1 CO_2, each S makes 1 SO_2 (1:1 for both). Set up the condition: moles SO_2 = half moles CO_2. One equation, one unknown. Solve for carbon mass.

Common Mistake

Setting up the condition as moles CO_2 = half moles SO_2 (reversing the given condition). The question says "moles of SO_2 is **half** that of CO_2," meaning $\mathrm{SO_2} = \frac{1}{2}\,\mathrm{CO_2}$. Getting this backward gives $m = 288/40 = 7.2$ g instead of 5.14 g.

Answer

(2) 5.14 g

Q.74 Find out moles of CO_2 and CO produced by combustion of 2 mol carbon with 1.25 mol O_2 leaving no residue: $\mathrm{C} + \mathrm{O_2} \longrightarrow \mathrm{CO_2}$ and $\mathrm{C} + \frac{1}{2}\mathrm{O_2} \longrightarrow \mathrm{CO}$

Explanation

Let x mol C burn to CO_2 and y mol C burn to CO.

Equation 1 (total C):

$$x + y = 2$$

Equation 2 (total O_2):

$$x + \frac{y}{2} = 1.25$$

Solving:

$$\begin{aligned} x + y &= 2 \\ x + \frac{y}{2} &= 1.25 \end{aligned}$$

Subtract equation 2 from equation 1:

$$\frac{y}{2} = 0.75 \implies y = 1.5 \text{ mol CO}$$

$$x = 2 - 1.5 = 0.5 \text{ mol CO}_2$$

Products: CO_2 = 0.5 mol, CO = 1.5 mol

Approach / Analogy

Two unknowns (mol of C → CO_2 and mol of C → CO), two equations (C balance + O_2 balance). Classic simultaneous equations. Key: CO_2 needs 1 mol O_2 per mol C, CO needs only 0.5 mol O_2 per mol C. Limited O_2 (1.25 mol $<$ 2 needed for all CO_2) forces some C to form CO.

Common Mistake

Assuming all C forms CO_2 first: 2 mol C would need 2 mol O_2, but only 1.25 mol available. So not all C can form CO_2. Set up simultaneous equations — you can't solve this by assuming complete conversion to one product.

Answer

(1) CO_2 = 0.5 mol, CO = 1.5 mol

— End of DPP-3 Complete Solution Sheet —

Stoichiometry · Q.1–Q.74 · All Parts Complete

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