

Weird Chemist

DPP-2 : Concentration Terms

Chapter: Some Basic Concepts of Chemistry

Solution Sheet

TYPE 1 : Solution Basics

Q.1 The example of a solid solution is

Explanation

A **solid solution** is formed when both the solute and solvent are in the solid state (or the solute dissolves into a solid matrix).

- Glucose in water: liquid solution
- **Copper in gold:** both are metals (solids); copper atoms dissolve into the gold lattice — this is a solid solution (alloy).
- Camphor in nitrogen: gas phase
- Oxygen in nitrogen: gas mixture

Approach / Analogy

Think of a solid solution like mixing two types of sand grains so thoroughly that you can't separate them by looking. Gold and copper atoms mix at the atomic level — no liquid involved. Alloys like brass (copper + zinc) and bronze are everyday examples.

Answer

(2) Copper in gold

Q.2 In a binary solution

Explanation

A binary solution has exactly **two components** (one solute + one solvent). There is no restriction on the physical states of either component:

- Solvent can be liquid (e.g., water), solid (e.g., alloys), or even gas.
- Solute can be solid, liquid, or gas (e.g., CO₂ dissolved in water).

So **any of these** statements can be true depending on the binary solution.

Approach / Analogy

Think of a binary solution as a two-person team. The team can be two runners, two swimmers, or one of each — as long as there are exactly two members. No rule says both must be the same type (state of matter).

Answer

(4) Any of these

Q.3 Which statement best explains the meaning of the phrase “like dissolves like”?

Explanation

“Like dissolves like” is the principle that:

- **Polar solvents** (e.g., water) dissolve **polar/ionic solutes** well.
- **Non-polar solvents** (e.g., benzene) dissolve **non-polar solutes** well.

The key reason is **similar intermolecular forces** — when solute and solvent have similar forces (both polar or both non-polar), the solute–solvent interactions are strong enough to break the solute–solute bonds and allow dissolution.

Approach / Analogy

Imagine two groups of friends at a party. People who like the same music (intermolecular forces) naturally mix together. A classical music fan (polar molecule) fits in with other classical fans (polar solvent), not with the heavy metal crowd (non-polar solvent). Same vibe = good mixing!

Answer

(2) A solvent and solute with similar intermolecular forces will readily form a solution

Q.4 An ionic compound that attracts atmospheric water so strongly that a hydrate is formed is said to be

Explanation

- **Hygroscopic**: absorbs moisture from air but may or may not form a hydrate.
- **Deliquescent**: absorbs so much moisture that it dissolves into a solution (extreme case of hygroscopic).
- The specific term for forming a **hydrate** by attracting atmospheric water is **hygroscopic** (in common usage at this level).
- Example: CaCl_2 absorbs water to form $\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$.

Approach / Analogy

Think of a hygroscopic compound as a sponge that grabs water molecules from the air and locks them into its crystal structure (hydrate). It’s like a magnet for water — the ionic compound is so attracted to water that it pulls it right out of the atmosphere.

Answer

(2) Hygroscopic

TYPE 2 : Percentage (%)

Q.5 A solution is prepared by adding 5 g of substance (x) to 18 g of water. The mass percent of the solute is

Explanation

Formula:

$$\text{Mass \%} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$

$$\text{Mass of solute} = 5 \text{ g}$$

$$\text{Mass of solvent} = 18 \text{ g}$$

$$\text{Mass of solution} = 5 + 18 = 23 \text{ g}$$

$$\text{Mass \%} = \frac{5}{23} \times 100 = \mathbf{21.74\%}$$

Approach / Analogy

Mass percent is simply: "Out of the total mixture, what fraction is the solute?" Like asking: if a 23 g trail mix has 5 g of cashews, what percent is cashews? $\frac{5}{23} \times 100 \approx 21.7\%$.

Common mistake: Don't divide by mass of solvent (18 g). Always use *total solution* mass in the denominator.

Answer

(1) 21.74%

Q.6 A solution is prepared by adding 4 g of substance 'A' and 6 g of substance 'B' to 30 g of water. What is the mass percentage of 'B' in the solution?

Explanation

$$\text{Mass of solution} = 4 + 6 + 30 = 40 \text{ g}$$

$$\text{Mass \% of B} = \frac{6}{40} \times 100 = \mathbf{15\%}$$

Approach / Analogy

With multiple solutes, total solution = all solutes + solvent combined. Just keep the target solute (B = 6 g) in the numerator and put the grand total in the denominator. The other solute (A) is part of the solution too.

Answer

(3) 15%

Q.7 When 400 g of a 20% solution by weight was cooled, 50 g of solute precipitated. The percentage concentration of the remaining solution is

Explanation

Step 1: Find initial solute mass.

$$\text{Solute}_{\text{initial}} = \frac{20}{100} \times 400 = 80 \text{ g}$$

Step 2: After precipitation.

$$\begin{aligned}\text{Solute remaining} &= 80 - 50 = 30 \text{ g} \\ \text{Solution remaining} &= 400 - 50 = 350 \text{ g}\end{aligned}$$

Step 3: New %

$$\text{New \%} = \frac{30}{350} \times 100 = \mathbf{8.57\%}$$

Approach / Analogy

When solute crystallises out (precipitates), it *leaves* the solution. So both the solute mass and total solution mass decrease by the same amount (50 g). Think of it like removing 50 g of dissolved sugar as solid crystals from a syrup — the remaining syrup is less sweet (lower %).

Answer

(1) 8.57%

Q.8 An aqueous solution of glucose is 10% in strength. The volume in which 2 mol of it is dissolved will be

Explanation

10% glucose means 10 g glucose per 100 g solution.

Step 1: Mass of 2 mol glucose:

$$2 \times 180 = 360 \text{ g of glucose}$$

Step 2: Using 10% by mass:

$$\frac{360}{M_{\text{solution}}} = \frac{10}{100} \implies M_{\text{solution}} = 3600 \text{ g}$$

Step 3: Assuming density $\approx 1 \text{ g/mL}$:

$$V = \frac{3600 \text{ g}}{1 \text{ g/mL}} = 3600 \text{ mL} = \mathbf{3.6 \text{ L}}$$

Approach / Analogy

If 10 g glucose lives in 100 g solution, then 360 g glucose lives in 3600 g solution. Assuming water-like density (1 g/mL), mass \approx volume in mL. Scaling up is like: if 1 small bottle of juice is 10% sugar, how many bottles hold 2 moles of sugar?

Answer

(2) 3.6 litre

Q.9 Ammonia gas is passed into water, yielding a solution of density 0.93 g/cm^3 and containing 18.6% NH_3 by weight. The mass of NH_3 per cc of the solution is

Explanation

Key idea: In 1 cc (1 mL) of solution, the mass of solution = density = 0.93 g .

$$\text{Mass of } \text{NH}_3 \text{ per cc} = \frac{18.6}{100} \times 0.93 = 0.186 \times 0.93 = \mathbf{0.17 \text{ g/cm}^3}$$

Approach / Analogy

Density tells you how heavy 1 cc of the solution is (0.93 g). The 18.6% tells you *what fraction* of that mass is NH_3 . Multiply the two: density \times % \div 100. It's like knowing a juice bottle weighs 0.93 g per mL and 18.6% is actual juice — how much actual juice is in 1 mL?

Answer

(1) 0.17 g/cm^3

TYPE 3 : Parts Per Million (ppm)

Q.10 A 500 g toothpaste sample has 0.02 g fluoride concentration. What is the concentration of fluorine in terms of ppm?

Explanation

Formula:

$$\text{ppm} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 10^6$$
$$\text{ppm} = \frac{0.02}{500} \times 10^6 = 4 \times 10^{-5} \times 10^6 = \mathbf{40 \text{ ppm}}$$

Approach / Analogy

ppm = “parts per million” = imagine scaling your 500 g sample up to $1,000,000 \text{ g}$. How many grams of fluoride would there be? If 500 g has 0.02 g , then $1,000,000 \text{ g}$ has $0.02 \times \frac{10^6}{500} = 40 \text{ g}$. So: 40 ppm .

Answer

(2) 40 ppm

Q.11 1000 g aqueous solution of CaCO_3 contains 10 g of calcium carbonate. The concentration of the solution is

Explanation

$$\text{ppm} = \frac{10}{1000} \times 10^6 = 0.01 \times 10^6 = \mathbf{10,000 \text{ ppm}}$$

Approach / Analogy

10 g out of 1000 g = 1% = 10,000 ppm. Remember: 1% = 10,000 ppm. This is a useful shortcut to memorise!

Answer

(4) 10,000 ppm

Q.12 A sample of toothpaste weighing 500 g, on analysis was found to contain 0.2 g of fluorine. The concentration of fluorine in ppm is

Explanation

$$\text{ppm} = \frac{0.2}{500} \times 10^6 = 4 \times 10^{-4} \times 10^6 = 4 \times 10^2 = \mathbf{400} \text{ ppm}$$

Approach / Analogy

Compare with Q.10: here 0.2 g fluorine (10 times more than Q.10's 0.02 g) in the same 500 g sample. So ppm is 10 times higher: $40 \times 10 = 400$ ppm.

Answer

(2) 4×10^2

Q.13 A given sample contains 30 mg of Hg in 500 mL of solution. The ppm of mercury in water is

Explanation

For aqueous solutions, density ≈ 1 g/mL, so 500 mL ≈ 500 g.

$$\text{Mass of Hg} = 30 \text{ mg} = 0.030 \text{ g}$$

$$\text{ppm} = \frac{0.030}{500} \times 10^6 = 6 \times 10^{-5} \times 10^6 = \mathbf{60} \text{ ppm}$$

Approach / Analogy

Always convert mg to g first! 30 mg = 0.030 g. Then apply the ppm formula. Think of it like: in 500 g of river water, how many "per million" parts are mercury? Scale 500 g to 1,000,000 g, scale the Hg proportionally.

Answer

(1) 60 ppm

Q.14 Persons are medically considered to have lead poisoning if they have a concentration greater than 10 micrograms of lead per decilitre of blood. Concentration in parts per billion is

Explanation

Given: 10 micrograms per decilitre of blood.

$$10 \mu\text{g} = 10 \times 10^{-6} \text{ g}$$

$$1 \text{ dL} = 100 \text{ mL} \approx 100 \text{ g (density} \approx 1)$$

$$\text{ppb} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^9 = \frac{10 \times 10^{-6}}{100} \times 10^9 = \frac{10^{-5}}{10^2} \times 10^9 = 10^{-7} \times 10^9 = \mathbf{100 \text{ ppb}}$$

Approach / Analogy

ppb = parts per *billion* = multiply ratio by 10^9 .

Key unit conversions: $1 \mu\text{g} = 10^{-6} \text{ g}$, $1 \text{ dL} = 100 \text{ mL} \approx 100 \text{ g}$. Once you have mass of Pb / mass of blood as a pure fraction, multiply by 10^9 for ppb.

Answer

(2) 100 ppb

TYPE 4 : Molarity

Key Formula:

$$M = \frac{\text{Moles of solute}}{\text{Volume of solution (in L)}} = \frac{n}{V} \quad \text{and} \quad n = \frac{\text{given mass}}{M_r}$$

TYPE 4(i): Molarity nikalni (Finding Molarity)

Q.15 A molar solution is one that contains one mole of solute in

Explanation

By definition, **1 Molar (1 M)** solution contains 1 mole of solute dissolved in **1 litre of solution** (not 1 litre of solvent — the solution's total volume is 1 L).

Approach / Analogy

Molarity is measured per litre of *solution* (solute + solvent together), not per litre of solvent alone. Imagine filling a 1 L flask: you add the solute first, then water up to the 1 L mark. The *flask* holds 1 L, not the water.

Answer

(2) One litre of the solution

Q.16 8 g NaOH is dissolved in one litre of solution. Its molarity is

Explanation

$$M_r(\text{NaOH}) = 23 + 16 + 1 = 40 \text{ g/mol}$$

$$n = \frac{8}{40} = 0.2 \text{ mol}$$

$$M = \frac{0.2}{1} = \mathbf{0.2 \text{ M}}$$

Approach / Analogy

Step 1: Find moles (mass \div molar mass). Step 2: Divide by volume in litres. Simple two-step recipe every time.

Answer

(3) 0.2 M

Q.17 A given solution of NaOH contains 4.00 g of NaOH per litre of solution. The molarity of this solution is

Explanation

$$n(\text{NaOH}) = \frac{4}{40} = 0.1 \text{ mol}$$

$$M = \frac{0.1}{1} = \mathbf{0.1 \text{ M}}$$

Approach / Analogy

Half the mass of Q.16 (4 g vs 8 g), same volume — so half the molarity (0.1 M vs 0.2 M). Cross-checking by proportion is a quick way to verify answers!

Answer

(1) 0.1 M

Q.18 Water is added to 10 g CaCO_3 to make 100 mL of solution. The molarity of the solution is

Explanation

$$M_r(\text{CaCO}_3) = 40 + 12 + 48 = 100 \text{ g/mol}$$

$$n = \frac{10}{100} = 0.1 \text{ mol}$$

$$V = 100 \text{ mL} = 0.1 \text{ L}$$

$$M = \frac{0.1}{0.1} = \mathbf{1.0 \text{ M}}$$

Approach / Analogy

CaCO_3 has a convenient molar mass of exactly 100 g/mol — easy to remember! 10 g = 0.1 mol. But volume is only 100 mL = 0.1 L. So $M = 0.1/0.1 = 1 \text{ M}$. Small volume, concentrated solution.

Answer

(1) 1.0 M

Q.19 If 18 g of glucose is present in 1000 g of solvent, the solution is said to be

Explanation

$$M_r(\text{glucose, C}_6\text{H}_{12}\text{O}_6) = 180 \text{ g/mol}$$

$$n = \frac{18}{180} = 0.1 \text{ mol}$$

The solute is in 1000 g = 1 kg of *solvent* — this is the definition of **molality**, not molarity!

$$\text{Molality} = \frac{0.1 \text{ mol}}{1 \text{ kg}} = \mathbf{0.1 \text{ molal}}$$

Approach / Analogy

This is a classic trap question! “1000 g of *solvent*” (not solution) is the clue. Whenever you see mass of *solvent*, think **molality** (m), not molarity (M). Molarity uses volume of *solution*; molality uses mass of *solvent*.

Answer

(4) 0.1 molal

Q.20 What is the molarity of NaOH solution if 250 mL of it contains 1 mg of NaOH?

Explanation

$$\text{Mass} = 1 \text{ mg} = 10^{-3} \text{ g}$$

$$n = \frac{10^{-3}}{40} = 2.5 \times 10^{-5} \text{ mol}$$

$$V = 250 \text{ mL} = 0.25 \text{ L}$$

$$M = \frac{2.5 \times 10^{-5}}{0.25} = 1 \times 10^{-4} = \mathbf{10^{-4} \text{ M}}$$

Approach / Analogy

Always convert mg to g first (divide by 1000). Very tiny mass = very low molarity. The answer 10^{-4} M makes sense — it’s an extremely dilute solution (1 mg is like a tiny pinch of NaOH in a glass of water).

Answer

(3) 10^{-4} M

Q.21 Concentration of glucose ($C_6H_{12}O_6$) in normal blood is approximately 90 mg per 100 mL. What is the molarity of the glucose solution in blood?

Explanation

$$\text{Mass} = 90 \text{ mg} = 0.09 \text{ g}$$

$$n = \frac{0.09}{180} = 5 \times 10^{-4} \text{ mol}$$

$$V = 100 \text{ mL} = 0.1 \text{ L}$$

$$M = \frac{5 \times 10^{-4}}{0.1} = 5 \times 10^{-3} = \mathbf{0.005 \text{ M}}$$

Approach / Analogy

This is a real-life value — your blood glucose is about 0.005 M (5 mM in medical terms). Tiny but critical. For JEE/NEET: always convert mg \rightarrow g and mL \rightarrow L before plugging in.

Answer

(2) 0.005 M

Q.22 The concentration of glucose in blood is 0.8 g L^{-1} . The molarity of glucose in the blood should be

Explanation

$$n = \frac{0.8}{180} = 4.44 \times 10^{-3} \text{ mol}$$

$$V = 1 \text{ L}$$

$$M = 4.44 \times 10^{-3} \approx \mathbf{4.4 \times 10^{-3} \text{ M}}$$

Approach / Analogy

0.8 g/L means 0.8 g dissolved in 1 L. Direct division by molar mass (180) gives moles per litre = molarity. Clean and quick.

Answer

(4) 4.4×10^{-3} M

Q.23 The molarity of a solution of NaCl (mol.wt. = 58.5) in water containing 5.85 g of NaCl in 500 mL of solution is

Explanation

$$n(\text{NaCl}) = \frac{5.85}{58.5} = 0.1 \text{ mol}$$
$$V = 500 \text{ mL} = 0.5 \text{ L}$$
$$M = \frac{0.1}{0.5} = \mathbf{0.2 \text{ M}}$$

Approach / Analogy

NaCl molar mass is 58.5 g/mol — a useful value to memorise. 5.85 g = exactly 0.1 mol. In 0.5 L, that's 0.2 M. Half-litre flask with 0.1 mol = 0.2 M.

Answer

(4) 0.2 M

Q.24 20.6 g NaBr is dissolved in 500 mL solution. What is the molarity of the resulting solution?

Explanation

$$M_r(\text{NaBr}) = 23 + 80 = 103 \text{ g/mol}$$
$$n = \frac{20.6}{103} = 0.2 \text{ mol}$$
$$V = 500 \text{ mL} = 0.5 \text{ L}$$
$$M = \frac{0.2}{0.5} = \mathbf{0.4 \text{ M}}$$

Approach / Analogy

NaBr: Na = 23, Br = 80 (Bromine has mass 80 — heavier than Cl which is 35.5). 20.6 g / 103 = 0.2 mol in 0.5 L = 0.4 M.

Answer

(2) 0.4 M

Q.25 6.025×10^{20} molecules of acetic acid are present in 500 mL of its solution. The concentration of solution is

Explanation

$$n = \frac{6.025 \times 10^{20}}{6.022 \times 10^{23}} \approx 10^{-3} \text{ mol} = 0.001 \text{ mol}$$
$$V = 500 \text{ mL} = 0.5 \text{ L}$$
$$M = \frac{0.001}{0.5} = \mathbf{0.002 \text{ M}}$$

Approach / Analogy

Molecules \rightarrow moles: always divide by Avogadro's number (6.022×10^{23}). Here, 6.025×10^{20} is very close to $\frac{N_A}{1000}$, so $n \approx 10^{-3}$ mol. Quick recognition of such fractions of N_A saves time.

Answer

(1) 0.002 M

Q.26 6.02×10^{21} molecules of urea are present in 100 mL of its solution. The concentration of urea solution is [AIEEE-2004]

Explanation

$$n = \frac{6.02 \times 10^{21}}{6.02 \times 10^{23}} = \frac{1}{100} = 0.01 \text{ mol}$$
$$V = 100 \text{ mL} = 0.1 \text{ L}$$
$$M = \frac{0.01}{0.1} = \mathbf{0.1 \text{ M}}$$

Approach / Analogy

$6.02 \times 10^{21} = \frac{N_A}{100}$, so $n = 0.01$ mol. In 0.1 L, that gives 0.1 M. Recognising these as fractions of N_A ($N_A/10$, $N_A/100$, $N_A/1000$) is a great speed trick for MCQs.

Answer

(4) 0.1 M

Q.27 The molarity of pure water is

Explanation

$$M_r(\text{H}_2\text{O}) = 18 \text{ g/mol}$$
$$\text{Density of water} = 1 \text{ g/mL}$$
$$\text{Mass of 1 L water} = 1000 \text{ g}$$
$$n = \frac{1000}{18} = 55.5 \text{ mol}$$
$$M = \frac{55.5}{1 \text{ L}} = \mathbf{55.5 \text{ M}}$$

Approach / Analogy

Pure water has a standard molarity of 55.5 M — memorise this! It seems paradoxical that water has “concentration” but it makes sense: 1 litre of pure water contains 55.5 moles of H_2O molecules. This value appears frequently in equilibrium problems too.

Answer

(2) 55.5 M

Q.28 Molarity of 720 g of pure water is

Explanation

720 g of water \neq 1 litre of water. Wait — do we know the volume of 720 g water?

$$V = \frac{720 \text{ g}}{1 \text{ g/mL}} = 720 \text{ mL} = 0.72 \text{ L}$$

$$n = \frac{720}{18} = 40 \text{ mol}$$

$$M = \frac{40}{0.72} = 55.5 \text{ M}$$

The molarity of pure water is **always 55.5 M** regardless of the quantity taken (it's an intensive property).

Approach / Analogy

Molarity is like density — it's an *intensive* property. Whether you take 720 g or 720,000 g of pure water, its molarity is always 55.5 M. .

Answer

(3) 55.5 M

Q.29 Equal weight of NaCl and KCl are dissolved separately in equal volumes of solutions. The molarity of the two solutions will be

Explanation

$$M_r(\text{NaCl}) = 58.5 \text{ g/mol}$$

$$M_r(\text{KCl}) = 74.5 \text{ g/mol}$$

For same mass w g in same volume V L:

$$M_{\text{NaCl}} = \frac{w/58.5}{V}$$

$$M_{\text{KCl}} = \frac{w/74.5}{V}$$

Since $58.5 < 74.5$: $\frac{w}{58.5} > \frac{w}{74.5}$

$$\therefore M_{\text{NaCl}} > M_{\text{KCl}}$$

Approach / Analogy

Lighter molecule = more moles from same mass. NaCl is lighter (58.5) than KCl (74.5). Same mass of a lighter substance gives MORE moles — like getting more pieces from lighter biscuits vs heavier cookies of the same total weight. More moles = higher molarity.

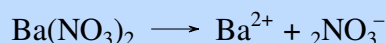
Answer

(3) That of NaCl will be more than that of KCl

Q.30 The total number of ions present in 1 mL of 0.1 M barium nitrate solution is

Explanation

Ba(NO₃)₂ dissociates as:



So 1 formula unit gives **3 ions** total.

$$n_{\text{Ba}(\text{NO}_3)_2} = M \times V = 0.1 \times 0.001 = 10^{-4} \text{ mol}$$

$$\begin{aligned} \text{Total ions} &= 3 \times 10^{-4} \times 6.02 \times 10^{23} \\ &= 3 \times 6.02 \times 10^{19} \\ &= \mathbf{3.0 \times 6.02 \times 10^{19}} \end{aligned}$$

(Note: 1 mL = 10⁻³ L)

Approach / Analogy

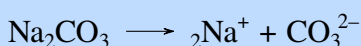
Ba(NO₃)₂ is like a family of 3: one Ba²⁺ dad and two NO₃⁻ kids. Every formula unit splits into 3 ions. So multiply moles by 3, then by Avogadro's number for total count.

Answer

(3) $3.0 \times 6.02 \times 10^{19}$

Q.31 25.3 g of Na₂CO₃ is dissolved in enough water to make 250 mL of solution. If Na₂CO₃ dissociates completely, molar concentration of Na⁺ and CO₃²⁻ are respectively. (Molar mass of Na₂CO₃ = 106 g mol⁻¹) [AIPMT (Prelims)-2010]

Explanation



$$n(\text{Na}_2\text{CO}_3) = \frac{25.3}{106} = 0.2388 \approx 0.239 \text{ mol}$$

$$V = 250 \text{ mL} = 0.25 \text{ L}$$

$$M(\text{Na}_2\text{CO}_3) = \frac{0.239}{0.25} = 0.955 \text{ M}$$

$$[\text{Na}^+] = 2 \times 0.955 = \mathbf{1.910 \text{ M}}$$

$$[\text{CO}_3^{2-}] = 1 \times 0.955 = \mathbf{0.955 \text{ M}}$$

Approach / Analogy

Na_2CO_3 has 2 sodium atoms per formula unit, so it gives *double* the Na^+ ions compared to the molarity of Na_2CO_3 . It's like a 2-for-1 deal on sodium: every mole of Na_2CO_3 gives 2 moles of Na^+ . Always write the dissociation equation first and count ions!

Answer

(2) 1.910 M and 0.955 M

Q.32 A solution of sodium sulfate contains 92 g of Na^+ ions per kilogram of water. The molality of Na^+ ions in that solution in mol kg^{-1} is [JEE-Main(Jan.)-2019]

Explanation

$$M_r(\text{Na}^+) = 23 \text{ g/mol}$$

$$n(\text{Na}^+) = \frac{92}{23} = 4 \text{ mol}$$

$$\text{Mass of solvent} = 1 \text{ kg}$$

$$\text{Molality} = \frac{4}{1} = 4 \text{ mol kg}^{-1}$$

Approach / Analogy

This directly gives moles of Na^+ per kg of solvent = molality definition. $92 \text{ g Na}^+ / 23 \text{ g mol}^{-1} = 4 \text{ mol}$. Done! No need to go back to Na_2SO_4 — the question directly asks for Na^+ molality.

Answer

(3) 4 mol kg^{-1}

Q.33 A 20.0 mL sample of CuSO_4 solution was evaporated to dryness, leaving 0.967 g of residue. What was the molarity of the original solution? ($\text{Cu} = 63.5$)

Explanation

The residue after evaporation is CuSO_4 (anhydrous, assuming no crystal water for this problem).

$$M_r(\text{CuSO}_4) = 63.5 + 32 + 64 = 159.5 \text{ g/mol}$$

$$n(\text{CuSO}_4) = \frac{0.967}{159.5} = 6.06 \times 10^{-3} \text{ mol}$$

$$V = 20 \text{ mL} = 0.02 \text{ L}$$

$$M = \frac{6.06 \times 10^{-3}}{0.02} = 0.303 \approx 0.0484 \text{ M}$$

Recalculating precisely: $n = 0.967/159.5 = 0.00606 \text{ mol}$; $M = 0.00606/0.020 = 0.303 \text{ M}$.

Hmm — checking option (4): 0.303 M seems correct. Let me verify: $0.303 \times 0.02 = 0.00606 \text{ mol}$, $\times 159.5 = 0.967 \text{ g}$. ✓

Approach / Analogy

Evaporation removes water but leaves the dissolved solid behind. So the residue mass = mass of solute. This gives us moles of CuSO_4 directly. Then standard molarity formula applies.

Answer

(4) 0.303 M

Q.34 Molarity of NaOH in 200 mL of an aqueous solution of it is 1 M. Find the change in molarity if 2 g of NaOH is added to it.

Explanation

Initial:

$$n_i(\text{NaOH}) = 1 \times 0.2 = 0.2 \text{ mol}$$

$$\text{Added NaOH} = \frac{2}{40} = 0.05 \text{ mol}$$

Final: (assuming volume stays ≈ 200 mL)

$$n_f = 0.2 + 0.05 = 0.25 \text{ mol}$$

$$M_f = \frac{0.25}{0.2} = 1.25 \text{ M}$$

$$\Delta M = 1.25 - 1.00 = \mathbf{0.25 \text{ M}}$$

Approach / Analogy

Adding solute to a solution increases molarity. The “change” is just new molarity minus old molarity. Key assumption: volume doesn't change significantly when a small amount of solid is added (standard exam assumption).

Answer

(2) 0.25 M

TYPE 4(ii): Solute weight find karna (Finding mass of solute)

Q.35 How many moles of NaCl are present in 250 mL of a 0.50 M NaCl solution?

Explanation

$$n = M \times V = 0.50 \times 0.250 = \mathbf{0.125 \text{ mol}}$$

Approach / Analogy

Rearranging $M = n/V$ gives $n = M \times V$. Always convert mL to L first. 250 mL = 0.25 L. Simple multiplication.

Answer

(1) 0.125 mol

Q.36 How many grams of H_2SO_4 are present in 500 mL of 0.2 M H_2SO_4 solution?

Explanation

$$\begin{aligned}n &= 0.2 \times 0.5 = 0.1 \text{ mol} \\M_r(\text{H}_2\text{SO}_4) &= 98 \text{ g/mol} \\ \text{Mass} &= 0.1 \times 98 = \mathbf{9.8 \text{ g}}\end{aligned}$$

Approach / Analogy

Three-step recipe: (1) $n = M \times V$, (2) $\text{mass} = n \times M_r$. Molar mass of $\text{H}_2\text{SO}_4 = 2(1) + 32 + 4(16) = 98 \text{ g/mol}$.

Answer

(1) 9.8 g

Q.37 How many grams of HNO_3 is required to prepare 400 mL solution of 0.2 M HNO_3 ?

Explanation

$$\begin{aligned}n &= 0.2 \times 0.4 = 0.08 \text{ mol} \\M_r(\text{HNO}_3) &= 1 + 14 + 48 = 63 \text{ g/mol} \\ \text{Mass} &= 0.08 \times 63 = \mathbf{5.04 \text{ g}}\end{aligned}$$

Approach / Analogy

Same three-step recipe. Remember molar mass of $\text{HNO}_3 = 63 \text{ g/mol}$ (H=1, N=14, 3O=48). 400 mL = 0.4 L.

Answer

(1) 5.04 g

Q.38 The amount of sugar ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) required to prepare 2 L of its 0.1 M aqueous solution is [JEE-Main(Jan.)-2019]

Explanation

$$M_r(\text{C}_{12}\text{H}_{22}\text{O}_{11}) = 12(12) + 22(1) + 11(16) = 144 + 22 + 176 = 342 \text{ g/mol}$$

$$n = 0.1 \times 2 = 0.2 \text{ mol}$$

$$\text{Mass} = 0.2 \times 342 = \mathbf{68.4 \text{ g}}$$

Approach / Analogy

Sucrose (table sugar) has molar mass 342 g/mol — worth memorising. 2 L of 0.1 M needs only 0.2 mol = 68.4 g. That's roughly 14 teaspoons of sugar — a pretty sweet drink!

Answer

(1) 68.4 g

Q.39 For preparing 0.1 M solution of H_2SO_4 in one litre, we need H_2SO_4

Explanation

$$n = 0.1 \times 1 = 0.1 \text{ mol}$$

$$\text{Mass} = 0.1 \times 98 = \mathbf{9.8 \text{ g}}$$

Approach / Analogy

1 L of 0.1 M H_2SO_4 needs 0.1 mol H_2SO_4 . At 98 g/mol, that's 9.8 g. Simple and direct.

Answer

(4) 9.8 g

Q.40 What is the mass of sodium acetate (CH_3COONa) required to make 200 mL of 0.245 molar aqueous solution? (Molar mass = 82 g mol^{-1}) [NCERT Pg. 23]

Explanation

$$n = 0.245 \times 0.200 = 0.049 \text{ mol}$$

$$\text{Mass} = 0.049 \times 82 = \mathbf{4.018 \approx 4 \text{ g}}$$

Approach / Analogy

$n = M \times V = 0.245 \times 0.2 = 0.049 \text{ mol}$. Mass = $0.049 \times 82 \approx 4 \text{ g}$. Classic NCERT numerical.

Answer

(2) 4 g

Q.41 How many grams of CH_3OH should be added to water to prepare 150 mL solution of 2 M

CH₃OH?

Explanation

$$M_r(\text{CH}_3\text{OH}) = 12 + 4 + 16 = 32 \text{ g/mol}$$

$$n = 2 \times 0.15 = 0.3 \text{ mol}$$

$$\text{Mass} = 0.3 \times 32 = \mathbf{9.6 \text{ g}}$$

Approach / Analogy

CH₃OH (methanol) molar mass = 32 g/mol. 150 mL = 0.15 L. $n = 2 \times 0.15 = 0.3 \text{ mol}$. Mass = $0.3 \times 32 = 9.6 \text{ g}$. Option (3) is 9.6 (not 9.6×10^3).

Answer

(3) 9.6

Q.42 If the density of methanol is 0.8 kg L^{-1} , what is the volume needed for making 2.5 L of its 0.4 M solution?

Explanation

$$n(\text{CH}_3\text{OH}) = 0.4 \times 2.5 = 1 \text{ mol}$$

$$\text{Mass needed} = 1 \times 32 = 32 \text{ g}$$

$$\text{Density} = 0.8 \text{ kg/L} = 800 \text{ g/L}$$

$$V = \frac{32}{800} = 0.04 \text{ L}$$

Approach / Analogy

First find how many grams of methanol you need, then use density to find the volume of pure liquid methanol to measure out. Like a recipe: you need 32 g of methanol, and the bottle has density $0.8 \text{ kg/L} = 800 \text{ g/L}$, so measure $32/800 = 0.04 \text{ L}$.

Answer

(3) 0.04 L

Q.43 5 g of an unknown solute is dissolved in 295 g solvent. If molarity and density of solution are 0.05 M and 1.5 g cc^{-1} respectively, the molecular weight of unknown solute is

Explanation

Step 1: Find volume of solution.

$$\text{Mass of solution} = 5 + 295 = 300 \text{ g}$$

$$V = \frac{300 \text{ g}}{1.5 \text{ g/cc}} = 200 \text{ cc} = 0.2 \text{ L}$$

Step 2: Find moles of solute.

$$n = M \times V = 0.05 \times 0.2 = 0.01 \text{ mol}$$

Step 3: Find molar mass.

$$M_r = \frac{\text{mass}}{n} = \frac{5}{0.01} = \mathbf{500 \text{ g/mol}}$$

Approach / Analogy

This is a “reverse” problem. You know molarity but need molar mass. Find volume from mass/density, find moles from $M \times V$, then $M_r = \text{mass}/\text{moles}$. Think of it as: you know the concentration and the quantity added, so you can “decode” the molecular identity.

Answer

(4) 500

Q.44 All of the water in a 0.20 M solution of NaCl was evaporated and 0.150 mol of NaCl was obtained. What was the original volume of the sample?

Explanation

$$V = \frac{n}{M} = \frac{0.150}{0.20} = 0.75 \text{ L} = \mathbf{750 \text{ mL}}$$

Approach / Analogy

Evaporation leaves behind all the solute (NaCl). So the 0.150 mol you collected *is* the amount of solute originally in the solution. Rearranging $M = n/V$ gives $V = n/M$. Straightforward reverse calculation.

Answer

(3) 750 mL

Q.45 Which of the following methods of expressing concentration varies with temperature?

Explanation

- **Molarity (M):** depends on volume. Volume changes with temperature (thermal expansion/contraction). **Temperature dependent.**
- **Molality (m):** depends on mass of solvent. Mass doesn't change with temperature. Temperature independent.
- **Mole fraction:** depends on moles only. Temperature independent.
- **Weight percent:** depends on mass. Temperature independent.

Approach / Analogy

Volume is sensitive to temperature (liquids expand when heated, contract when cooled). Molarity uses volume of solution, so it changes as temperature changes. Think of a balloon: heating makes it bigger, so the same amount of solute is now in a bigger volume — lower molarity. Molality, mole fraction, and % all use mass, which stays constant.

Answer

(3) Molarity

TYPE 5 : Molality

Key Formula:

$$m = \frac{\text{Moles of solute}}{\text{Mass of solvent (in kg)}}$$

Remember: Molality uses mass of **solvent** (not solution) in **kg** (not grams).

Q.46 A molal solution is one that contains one mole of a solute in

Explanation

By definition, **1 molal (1 m)** solution = 1 mole of solute dissolved in **1000 g (1 kg) of the solvent.**

Key distinction from molarity:

- Molarity (M): 1 mol solute per **litre of solution**
- Molality (m): 1 mol solute per **kg of solvent**

Approach / Analogy

Molality is the “weighing scale” version of concentration. Instead of measuring the total volume of the drink (solution), you weigh just the water (solvent) on a scale. That's why molality is temperature-independent — mass doesn't expand or contract with temperature.

Common Mistake

Students confuse “1 kg of solution” with “1 kg of solvent.” Molality specifically uses mass of **solvent alone** — not the total solution. The solute's mass is NOT included in the denominator.

Answer

(1) 1000 g of the solvent

Q.47 A solution contains 10 moles of sucrose in 1 kg of solvent. The molality of the solution is

Explanation

$$m = \frac{n_{\text{solute}}}{\text{mass of solvent (kg)}} = \frac{10}{1} = \mathbf{10 \text{ mol kg}^{-1}}$$

Directly given: 10 mol in 1 kg solvent.

Approach / Analogy

This is a definition-based question — values are directly plugged into the molality formula. No calculation needed beyond reading carefully.

Answer

(1) 10 mol kg⁻¹

Q.48 The molality of a solution containing 5.3 g of anhydrous Na₂CO₃ in 400 g of water is

Explanation

$$M_r(\text{Na}_2\text{CO}_3) = 2(23) + 12 + 3(16) = 46 + 12 + 48 = 106 \text{ g/mol}$$

$$n = \frac{5.3}{106} = 0.05 \text{ mol}$$

$$\text{Mass of solvent} = 400 \text{ g} = 0.4 \text{ kg}$$

$$m = \frac{0.05}{0.4} = \mathbf{0.125 \text{ m}}$$

Approach / Analogy

Two-step: (1) find moles from mass/molar mass, (2) divide by solvent mass in kg. Always convert solvent grams to kg (divide by 1000).

Common Mistake

Forgetting to convert 400 g to 0.4 kg. If you keep it in grams, your answer will be 1000 times too small (0.000125 instead of 0.125). **Always convert solvent mass to kg before dividing.**

Answer

(1) 0.125 m

Q.49 The molality of a solution containing 3 g CH₃COOH in 50 g benzene is

Explanation

$$M_r(\text{CH}_3\text{COOH}) = 12 + 4 + 12 + 16 + 16 = 60 \text{ g/mol}$$

$$n = \frac{3}{60} = 0.05 \text{ mol}$$

$$\text{Mass of solvent} = 50 \text{ g} = 0.05 \text{ kg}$$

$$m = \frac{0.05}{0.05} = \mathbf{1 \text{ m}}$$

Approach / Analogy

Solvent here is **benzene** — not water! Molality doesn't care about solvent identity, just its mass in kg. Same formula regardless of whether solvent is water, benzene, or anything else.

Common Mistake

Students sometimes panic seeing benzene as solvent and try to adjust the formula. No adjustment needed — molality formula is universal. Just use benzene's mass in kg.

Answer

(4) 1 m

Q.50 Calculate molality of the solution obtained by dissolving 11.7 g NaCl in 500 g water

Explanation

$$M_r(\text{NaCl}) = 23 + 35.5 = 58.5 \text{ g/mol}$$

$$n = \frac{11.7}{58.5} = 0.2 \text{ mol}$$

$$m = \frac{0.2}{0.5} = \mathbf{0.4 \text{ m}}$$

Approach / Analogy

11.7 g NaCl = exactly 0.2 mol (since $58.5 \times 0.2 = 11.7$). In 0.5 kg water: $m = 0.2/0.5 = 0.4 \text{ m}$.

Answer

(4) 0.4 m

Q.51 The molality of a urea solution in which 0.0100 g of urea $[(\text{NH}_2)_2\text{CO}]$ is added to 0.3000 dm³ of water at STP is

Explanation

$$M_r(\text{urea}) = 2(14) + 2(2) + 12 + 16 = 28 + 4 + 12 + 16 = 60 \text{ g/mol}$$

$$n = \frac{0.0100}{60} = 1.667 \times 10^{-4} \text{ mol}$$

At STP, water density $\approx 1 \text{ g/mL}$, so $0.3000 \text{ dm}^3 = 300 \text{ mL} \approx 300 \text{ g} = 0.3 \text{ kg}$.

$$m = \frac{1.667 \times 10^{-4}}{0.3} = 5.55 \times 10^{-4} \text{ m}$$

Approach / Analogy

$1 \text{ dm}^3 = 1 \text{ litre} = 1000 \text{ mL}$. So $0.3 \text{ dm}^3 = 300 \text{ mL}$ of water $\approx 300 \text{ g} = 0.3 \text{ kg}$. Very dilute solution — tiny amount of urea (0.01 g) in a large amount of water.

Common Mistake

Confusing dm^3 with mL. Remember: $1 \text{ dm}^3 = 1 \text{ L} = 1000 \text{ mL}$. Students sometimes treat 0.3 dm^3 as 0.3 mL and get a wildly wrong answer.

Answer

(2) $5.55 \times 10^{-4} \text{ m}$

Q.52 How many grams of KCl must be added to 75.0 grams of water to produce a solution that is 2.25 molal? (Molecular weight of KCl = 74.5)

Explanation

Rearranging molality formula:

$$n = m \times \text{mass of solvent (kg)} = 2.25 \times 0.075 = 0.16875 \text{ mol}$$

$$\text{Mass of KCl} = 0.16875 \times 74.5 = \mathbf{12.57 \text{ g}}$$

Approach / Analogy

Reverse molality: instead of finding m , find grams. Rearrange: $n = m \times kg_{\text{solvent}}$, then $\text{mass} = n \times M_r$. Like knowing the recipe concentration and batch size — calculate how much ingredient you need.

Common Mistake

Using 75 g as 75 kg instead of 0.075 kg. Always convert solvent grams \rightarrow kg. Also, don't confuse molality with molarity here — $m \times \text{kg solvent}$ gives moles, NOT $M \times \text{litres solution}$.

Answer

(4) 12.57 g

TYPE 6 : Mole Fraction

Key Formula:

$$\chi_A = \frac{n_A}{n_A + n_B + \dots}$$

Sum of all mole fractions = 1. Mole fraction has no units.

Q.53 23 g ethanol is dissolved in 36 g water. Find mole fraction of ethanol.

Explanation

$$M_r(\text{ethanol, C}_2\text{H}_5\text{OH}) = 46 \text{ g/mol}$$

$$M_r(\text{H}_2\text{O}) = 18 \text{ g/mol}$$

$$n_{\text{ethanol}} = \frac{23}{46} = 0.5 \text{ mol}$$

$$n_{\text{water}} = \frac{36}{18} = 2 \text{ mol}$$

$$\chi_{\text{ethanol}} = \frac{0.5}{0.5 + 2} = \frac{0.5}{2.5} = \mathbf{0.2}$$

Approach / Analogy

Mole fraction is the “share of moles.” If you have 0.5 mol ethanol and 2 mol water, total = 2.5 mol. Ethanol’s share = $0.5/2.5 = 0.2$ (or 20% of the molecules are ethanol). Like what fraction of people in a room are wearing red — count reds, divide by total.

Common Mistake

Using mass fraction instead of mole fraction — do NOT write $23/(23 + 36)$. You must first convert masses to **moles**, then compute the fraction. This is the most common error in mole fraction questions.

Answer

(3) 0.2

Q.54 Mole fraction of glycerine ($\text{C}_3\text{H}_8\text{O}_3$) in a solution of 36 g of water and 46 g of glycerine is

Explanation

$$M_r(\text{C}_3\text{H}_8\text{O}_3) = 3(12) + 8(1) + 3(16) = 36 + 8 + 48 = 92 \text{ g/mol}$$

$$n_{\text{glycerine}} = \frac{46}{92} = 0.5 \text{ mol}$$

$$n_{\text{water}} = \frac{36}{18} = 2 \text{ mol}$$

$$\chi_{\text{glycerine}} = \frac{0.5}{0.5 + 2} = \frac{0.5}{2.5} = \mathbf{0.2}$$

Approach / Analogy

Same structure as Q.53. Glycerine molar mass = 92 g/mol (3C + 8H + 3O). Despite having more mass (46 g vs 36 g of water), glycerine has fewer moles because its molecule is much heavier. Heavier molecules = fewer moles per gram.

Answer

(3) 0.20

Q.55 The mole fraction of oxygen in a mixture of 7 g of nitrogen and 8 g of oxygen is

Explanation

$$M_r(\text{N}_2) = 28 \text{ g/mol}, \quad M_r(\text{O}_2) = 32 \text{ g/mol}$$

$$n_{\text{N}_2} = \frac{7}{28} = 0.25 \text{ mol}$$

$$n_{\text{O}_2} = \frac{8}{32} = 0.25 \text{ mol}$$

$$\chi_{\text{O}_2} = \frac{0.25}{0.25 + 0.25} = \frac{0.25}{0.50} = \mathbf{0.5}$$

Approach / Analogy

Both gases give exactly 0.25 mol each (nice round numbers). Equal moles means mole fraction = 0.5 for each. This is a 50–50 mixture *in terms of moles*, even though masses are different (7 g N₂ vs 8 g O₂).

Common Mistake

Writing mole fraction as 8/15 (mass fraction) instead of converting to moles first. 7 g N₂ and 8 g O₂ are NOT equal moles just because masses are close. Always convert to moles first!

Answer

(2) 0.5

Q.56 In a solution of 7.8 g benzene (C₆H₆) and 46.0 g toluene (C₆H₅CH₃) the mole fraction of benzene is

Explanation

$$\begin{aligned}M_r(\text{C}_6\text{H}_6) &= 78 \text{ g/mol} \\M_r(\text{C}_6\text{H}_5\text{CH}_3) &= 92 \text{ g/mol} \\n_{\text{benzene}} &= \frac{7.8}{78} = 0.1 \text{ mol} \\n_{\text{toluene}} &= \frac{46.0}{92} = 0.5 \text{ mol} \\\chi_{\text{benzene}} &= \frac{0.1}{0.1 + 0.5} = \frac{0.1}{0.6} = \frac{1}{6}\end{aligned}$$

Approach / Analogy

Benzene ($M_r = 78$) and toluene ($M_r = 92$) are similar molecules — toluene is just benzene with a CH_3 group added. Nice clean numbers: $7.8/78 = 0.1$ and $46/92 = 0.5$. Fraction = $0.1/0.6 = 1/6$.

Answer

(1) $\frac{1}{6}$

Q.57 A solution is prepared by adding 360 g of glucose to 864 g of water. The mole fraction of glucose (molar mass = 180) is

Explanation

$$\begin{aligned}n_{\text{glucose}} &= \frac{360}{180} = 2 \text{ mol} \\n_{\text{water}} &= \frac{864}{18} = 48 \text{ mol} \\\chi_{\text{glucose}} &= \frac{2}{2 + 48} = \frac{2}{50} = \mathbf{0.04}\end{aligned}$$

Approach / Analogy

360 g of glucose = 2 mol (double the molar mass of 180). 864 g of water = 48 mol. Total = 50 mol. Glucose fraction = $2/50 = 0.04 = 4\%$. Small fraction makes sense — a lot of water, a little glucose.

Answer

(1) 0.04

Q.58 The mole fraction of carbon in diamond is

Explanation

Diamond is made of **pure carbon** only — 100% carbon atoms, nothing else.

$$\chi_C = \frac{n_C}{n_C} = \mathbf{1}$$

A pure substance has a mole fraction of 1 for itself.

Approach / Analogy

If there's only one component in the system, its mole fraction = 1 (it's the entire mixture). Diamond = pure C. Same logic applies to pure water (mole fraction of water = 1) or any pure substance.

Common Mistake

Trying to calculate using atomic mass (12 g/mol) and getting confused. No calculation needed — purity = mole fraction of 1. A pure substance has no other component to dilute it.

Answer

(4) 1

Q.59 Which of the following is correct?

Explanation

Evaluating each option:

- (1) **Sum of mole fractions = 1**: CORRECT. This is a fundamental property: $\sum \chi_i = 1$ always.
- (2) Mole fraction depends on temperature: FALSE. Mole fraction = moles/total moles. Moles don't change with temperature.
- (3) Mole fraction is always negative: FALSE. Mole fraction is always between 0 and 1.
- (4) Mole fraction is independent of solute content: FALSE. Adding more solute changes all mole fractions.

Approach / Analogy

Think of mole fractions as percentage shares of a company. All shareholders' percentages must add up to 100% (or 1 in fraction form). You can't have negative ownership, and adding a new shareholder changes everyone's fraction.

Answer

(1) The sum of mole fractions of all components in a solution is always unity

Q.60 8 g of NaOH is dissolved in 18 g of H₂O. Mole fraction of NaOH in solution and molality (in mol kg⁻¹) of the solutions respectively are [JEE-Main(Jan.)-2019]

Explanation

$$n_{\text{NaOH}} = \frac{8}{40} = 0.2 \text{ mol}$$

$$n_{\text{H}_2\text{O}} = \frac{18}{18} = 1 \text{ mol}$$

$$\chi_{\text{NaOH}} = \frac{0.2}{0.2 + 1} = \frac{0.2}{1.2} = \mathbf{0.167}$$

$$m = \frac{0.2 \text{ mol}}{0.018 \text{ kg}} = \mathbf{11.11 \text{ mol kg}^{-1}}$$

(18 g water = 0.018 kg)

Approach / Analogy

Two quantities, one setup. First find moles of both components (0.2 mol NaOH, 1 mol water). Use moles for mole fraction; use moles of solute and kg of solvent for molality. One set of mole calculations, two answers.

Common Mistake

For molality, students write 18 g water = 18 kg (massive error!) instead of 18 g = 0.018 kg. This gives 0.01111 m instead of 11.11 m — off by a factor of 1000! Always: divide grams by 1000 to get kg.

Answer

(1) 0.167, 11.11

TYPE 7 : Dilution

Key Formula (Dilution):

$$M_1V_1 = M_2V_2$$

Moles of solute before dilution = Moles of solute after dilution (only water is added, solute stays).

For Mixing:

$$M_1V_1 + M_2V_2 = M_f(V_1 + V_2)$$

Q.61 Which of the following is correct for a concentrated and diluted solution of the same substance?

Explanation

When a solution is diluted (water added), the **moles of solute do not change**:

$$n_{\text{before}} = n_{\text{after}} \implies M_1V_1 = M_2V_2$$

This is option (2). The other options are incorrect formulations.

Approach / Analogy

Dilution is like spreading the same amount of jam (solute) over a larger bread surface (solution volume). The jam doesn't increase — only the area does. So moles in = moles out: $M_1V_1 = M_2V_2$.

Common Mistake

Confusing $M_1V_1 = M_2V_2$ (dilution of same substance) with $M_1V_1 + M_2V_2 = M_f(V_1 + V_2)$ (mixing of two different solutions). Use the first formula **only** when you're adding pure solvent (water) to dilute.

Answer

$$(2) M_1V_1 = M_2V_2$$

Q.62 What volume of water is to be added to 100 cm³ of 0.5 M NaOH solution to make it 0.1 M solution?

Explanation

Using $M_1V_1 = M_2V_2$:

$$0.5 \times 100 = 0.1 \times V_2$$

$$V_2 = \frac{0.5 \times 100}{0.1} = 500 \text{ cm}^3$$

Volume of water to be added:

$$V_{\text{water}} = V_2 - V_1 = 500 - 100 = \mathbf{400 \text{ cm}^3}$$

Approach / Analogy

$M_1V_1 = M_2V_2$ gives the *final* volume (500 cm³), not the volume of water added. The water to add = final volume – initial volume = 500 – 100 = 400 cm³.

Common Mistake

The most common error: answering 500 cm³ (the final volume) instead of 400 cm³ (the water to add). The question asks “how much water to ADD,” not “what is the final volume.” Always subtract: water added = $V_2 - V_1$.

Answer

$$(2) 400 \text{ cm}^3$$

Q.63 How much water should be added to 2 M HCl solution to form 1 litre of 0.5 M HCl?

Explanation

Using $M_1V_1 = M_2V_2$:

$$2 \times V_1 = 0.5 \times 1$$

$$V_1 = \frac{0.5}{2} = 0.25 \text{ L}$$

Water to add:

$$V_{\text{water}} = 1 - 0.25 = \mathbf{0.75 \text{ L}}$$

Approach / Analogy

Final volume is given (1 L). Find the initial volume of concentrated solution needed (0.25 L), then the rest is water: $1 - 0.25 = 0.75 \text{ L}$. Think of it as: you need 1 L total, but only 250 mL is from the stock solution; the remaining 750 mL is water.

Common Mistake

Answering 0.25 L (the stock HCl volume) instead of 0.75 L (the water added). Re-read: the question asks for water added, not volume of stock solution.

Answer

(1) 0.75 L

Q.64 What volumes of 1 M and 2 M H_2SO_4 solution are required to produce 2 L of 1.75 M H_2SO_4 solution?

Explanation

Let volume of 1 M solution = V_1 and volume of 2 M solution = V_2 .

Condition 1 (total volume):

$$V_1 + V_2 = 2 \text{ L}$$

Condition 2 (moles conservation):

$$1 \times V_1 + 2 \times V_2 = 1.75 \times 2 = 3.5$$

Solving:

$$V_1 = 2 - V_2$$

$$(2 - V_2) + 2V_2 = 3.5$$

$$2 + V_2 = 3.5$$

$$V_2 = 1.5 \text{ L}, \quad V_1 = 0.5 \text{ L}$$

Approach / Analogy

Think of this as a mixing problem with two unknowns. Set up two equations: one for total volume, one for total moles. This is essentially the “alligation” method in disguise — mixing a weak (1 M) and a strong (2 M) solution to get an intermediate (1.75 M). Since 1.75 M is close to 2 M, you need more of the 2 M solution.

Common Mistake

Setting up only one equation and guessing. You **must** set up two simultaneous equations (volume + moles) to solve uniquely. Also, a quick sanity check: final molarity (1.75 M) is closer to 2 M than to 1 M, so you need more of the 2 M solution. Answer: 1.5 L of 2 M and 0.5 L of 1 M. This confirms option (1).

Answer

(1) 0.5 L of 1 M and 1.5 L of 2 M

TYPE 8 : Mixing of Solutions

Key Formula (Mixing):

$$M_f = \frac{M_1V_1 + M_2V_2}{V_1 + V_2}$$

Total moles \div total volume. For ions: find moles of the ion from each source, sum them, divide by total volume.

Q.65 50 mL of 0.2 M H_2SO_4 is mixed with 50 mL of 0.3 M H_2SO_4 . The molarity of the final solution is

Explanation

$$n_1 = 0.2 \times 0.050 = 0.010 \text{ mol}$$

$$n_2 = 0.3 \times 0.050 = 0.015 \text{ mol}$$

$$M_f = \frac{0.010 + 0.015}{0.050 + 0.050} = \frac{0.025}{0.100} = \mathbf{0.25 \text{ M}}$$

Equal volumes: shortcut = simple average = $(0.2 + 0.3)/2 = 0.25 \text{ M}$.

Approach / Analogy

When volumes are **equal**, final molarity = simple average of the two molarities. Like mixing equal cups of 0.2 M and 0.3 M coffee — you get the average strength 0.25 M. This shortcut only works for equal volumes!

Common Mistake

Using the average shortcut for **unequal** volumes. Always check volumes first. If $V_1 \neq V_2$, use the full weighted formula $M_f = (M_1V_1 + M_2V_2)/(V_1 + V_2)$.

Answer

(1) 0.25 M

Q.66 500 mL of 0.1 M HCl is mixed with 500 mL of 0.2 M HCl. The molarity of the final solution is

Explanation

Equal volumes: $M_f = \frac{0.1+0.2}{2} = \mathbf{0.15\ M}$

Approach / Analogy

Same equal-volume trick. Average of 0.1 and 0.2 = 0.15 M. Quick 5-second MCQ answer.

Answer

(1) 0.15 M

Q.67 2.5 litre of 1 M NaOH solution is mixed with another 3 litre of 0.5 M NaOH solution. Find out the molarity of the resultant solution.

Explanation

$$\begin{aligned}n_1 &= 1 \times 2.5 = 2.5 \text{ mol} \\n_2 &= 0.5 \times 3.0 = 1.5 \text{ mol} \\M_f &= \frac{2.5 + 1.5}{2.5 + 3.0} = \frac{4.0}{5.5} = 0.727 \approx \mathbf{0.73\ M}\end{aligned}$$

Approach / Analogy

Unequal volumes — must use full formula. Result (0.73 M) lies between 0.5 M and 1 M, pulled slightly toward 1 M since the 1 M solution contributes more moles (2.5 mol vs 1.5 mol).

Common Mistake

Averaging molarities: $(1+0.5)/2 = 0.75\ \text{M}$ — **wrong** for unequal volumes! This is the most common error in mixing questions. Weighted average is mandatory when volumes differ.

Answer

(3) 0.73 M

Q.68 10 mL of 2 M NaOH solution is added to 200 mL of 0.5 M NaOH solution. What is the final concentration? [JEE(Main-online)-2013]

Explanation

$$\begin{aligned}n_1 &= 2 \times 0.010 = 0.020 \text{ mol} \\n_2 &= 0.5 \times 0.200 = 0.100 \text{ mol} \\M_f &= \frac{0.020 + 0.100}{0.010 + 0.200} = \frac{0.120}{0.210} = \mathbf{0.571 \approx 0.57\ M}\end{aligned}$$

Approach / Analogy

10 mL is tiny compared to 200 mL — so result is close to 0.5 M but slightly above. When one volume dominates, the final molarity is pulled toward that solution's concentration. Sanity check before calculating!

Answer

(1) 0.57 M

Q.69 50 mL of 0.1 M HCl is mixed with 150 mL of 0.3 M HCl and 300 mL of water. The molarity of the final solution is

Explanation

$$\begin{aligned}n_1 &= 0.1 \times 0.050 = 0.005 \text{ mol} \\n_2 &= 0.3 \times 0.150 = 0.045 \text{ mol} \\V_{\text{total}} &= 50 + 150 + 300 = 500 \text{ mL} = 0.5 \text{ L} \\M_f &= \frac{0.005 + 0.045}{0.5} = \frac{0.050}{0.5} = \mathbf{0.10 \text{ M}}\end{aligned}$$

Approach / Analogy

Water contributes **volume but zero moles**. It just makes the container bigger, diluting the solution. Add water's 300 mL to denominator, add 0 to numerator.

Common Mistake

Forgetting to include 300 mL water in total volume. Ignoring it gives $0.050/0.200 = 0.25 \text{ M}$ instead of 0.10 M — more than double the correct answer. Water always adds to total volume.

Answer

(1) 0.10 M

Q.70 4.9 g H_2SO_4 is mixed with 250 mL H_2O and 250 mL of 0.1 M H_2SO_4 . The molarity of the final solution is

Explanation

$$\begin{aligned}n_{\text{solid}} &= \frac{4.9}{98} = 0.050 \text{ mol} \\n_{\text{soln}} &= 0.1 \times 0.250 = 0.025 \text{ mol} \\V_{\text{total}} &= 250 + 250 = 500 \text{ mL} = 0.5 \text{ L} \\M_f &= \frac{0.050 + 0.025}{0.5} = \frac{0.075}{0.5} = \mathbf{0.15 \text{ M}}\end{aligned}$$

Approach / Analogy

Three sources: solid acid (moles only, no volume counted), water (volume only, zero moles), solution (both moles and volume). Handle each separately then combine.

Common Mistake

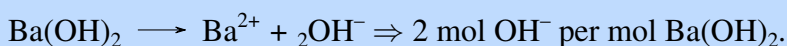
Treating water as a source of H_2SO_4 moles. Water adds 0 moles of solute. Also, the solid's volume is neglected — only the water and solution volumes count in the denominator.

Answer

(1) 0.15 M

Q.71 What is the $[\text{OH}^-]$ in the final solution prepared by mixing 20.0 mL of 0.050 M HCl with 30.0 mL of 0.10 M $\text{Ba}(\text{OH})_2$?

Explanation



$$n_{\text{H}^+} = 0.050 \times 0.020 = 0.001 \text{ mol}$$

$$n_{\text{OH}^-} = 2 \times 0.10 \times 0.030 = 0.006 \text{ mol}$$

$$n_{\text{OH}^- \text{ excess}} = 0.006 - 0.001 = 0.005 \text{ mol}$$

$$[\text{OH}^-] = \frac{0.005}{0.050} = \mathbf{0.10 \text{ M}}$$

Approach / Analogy

Acid and base neutralise first ($\text{H}^+ + \text{OH}^- \longrightarrow \text{H}_2\text{O}$). After neutralisation, find the excess ion and divide by total volume. OH^- is in excess here (0.005 mol in 50 mL total).

Common Mistake

Forgetting that $\text{Ba}(\text{OH})_2$ gives **2 moles of OH^-** per mole. Writing $n_{\text{OH}^-} = 0.10 \times 0.030 = 0.003$ (missing the $\times 2$ factor) gives wrong excess. Always write the dissociation equation and count OH^- ions per formula unit.

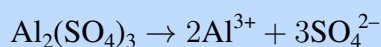
Answer

(2) 0.10 M

Q.72 20 mL of 0.5 M Na_2SO_4 is mixed with 50 mL of 0.2 M H_2SO_4 and 30 mL of 0.4 M $\text{Al}_2(\text{SO}_4)_3$ solution. Assuming 100% dissociation, the concentration of SO_4^{2-} in the final solution is

Explanation

Dissociations:



$$n(\text{SO}_4^{2-})_1 = 1 \times 0.5 \times 0.020 = 0.010 \text{ mol}$$

$$n(\text{SO}_4^{2-})_2 = 1 \times 0.2 \times 0.050 = 0.010 \text{ mol}$$

$$n(\text{SO}_4^{2-})_3 = 3 \times 0.4 \times 0.030 = 0.036 \text{ mol}$$

$$V_{\text{total}} = 20 + 50 + 30 = 100 \text{ mL} = 0.1 \text{ L}$$

$$[\text{SO}_4^{2-}] = \frac{0.010 + 0.010 + 0.036}{0.1} = \frac{0.056}{0.1} = \mathbf{0.56 \text{ M}}$$

Approach / Analogy

The key is the sulfate **multiplier** for each salt: Na_2SO_4 gives $1\times$, H_2SO_4 gives $1\times$, $\text{Al}_2(\text{SO}_4)_3$ gives $3\times$ sulfate. Write dissociation equations first, then count sulfate ions per formula unit.

Common Mistake

Using molarity of the salt directly as molarity of SO_4^{2-} without the multiplier. $\text{Al}_2(\text{SO}_4)_3$ gives **3** sulfate ions per formula unit — using multiplier 1 instead of 3 gives a drastically wrong answer. Always write the dissociation equation and count ions.

Answer

(1) 0.56 M

Q.73 A solution is made by mixing 300 mL 1.5 M $\text{Al}_2(\text{SO}_4)_3$ + 300 mL 2 M CaSO_4 + 400 mL 3.5 M CaCl_2 . Assuming complete dissociation, the final molarity of SO_4^{2-} is

Explanation

CaCl_2 gives no SO_4^{2-} . Sources only:

$$n(\text{SO}_4^{2-})_{\text{Al}_2(\text{SO}_4)_3} = 3 \times 1.5 \times 0.300 = 1.350 \text{ mol}$$

$$n(\text{SO}_4^{2-})_{\text{CaSO}_4} = 1 \times 2.0 \times 0.300 = 0.600 \text{ mol}$$

$$V_{\text{total}} = 300 + 300 + 400 = 1000 \text{ mL} = 1.0 \text{ L}$$

$$[\text{SO}_4^{2-}] = \frac{1.350 + 0.600}{1.0} = \mathbf{1.95 \text{ M}}$$

Approach / Analogy

CaCl_2 is a decoy — it contributes Ca^{2+} and Cl^- but **zero sulfate**. Total volume = 1 L makes division trivial. Identify which compounds contain the target ion before computing.

Common Mistake

Including CaCl_2 as a source of SO_4^{2-} (it has no sulfate) OR forgetting the $\times 3$ multiplier for $\text{Al}_2(\text{SO}_4)_3$. Either error gives a wrong answer. Check the formula for sulfate ions before computing.

Answer

(1) 1.95 M

Q.74 A solution is made by mixing 300 mL 1.5 M $\text{Al}_2(\text{SO}_4)_3$ + 300 mL 2 M CaSO_4 + 400 mL 3.5 M CaCl_2 . Assuming complete dissociation, the final molarity of Ca^{2+} is

Explanation

Only CaSO_4 and CaCl_2 give Ca^{2+} (1 mol each per formula unit):

$$n(\text{Ca}^{2+})_{\text{CaSO}_4} = 1 \times 2.0 \times 0.300 = 0.600 \text{ mol}$$

$$n(\text{Ca}^{2+})_{\text{CaCl}_2} = 1 \times 3.5 \times 0.400 = 1.400 \text{ mol}$$

$$[\text{Ca}^{2+}] = \frac{0.600 + 1.400}{1.0} = \mathbf{2.00 \text{ M}}$$

Approach / Analogy

$\text{Al}_2(\text{SO}_4)_3$ is the decoy this time — no calcium. Same total volume (1 L) from Q.73. Q.73–Q.75 all use the same mixture; set up the full ion table once and answer all three from it.

Common Mistake

Counting $\text{Al}_2(\text{SO}_4)_3$ as a calcium source. $\text{Al}_2(\text{SO}_4)_3$ contains Al^{3+} and SO_4^{2-} — no Ca^{2+} . Reading chemical formulas carefully before assigning ion contributions is essential.

Answer

(1) 2.00 M

Q.75 A solution is made by mixing 300 mL 1.5 M $\text{Al}_2(\text{SO}_4)_3$ + 300 mL 2 M CaSO_4 + 400 mL 3.5 M CaCl_2 . Assuming complete dissociation, the final molarity of Cl^- is

Explanation

Only CaCl_2 gives Cl^- with multiplier 2:

$$n(\text{Cl}^-) = 2 \times 3.5 \times 0.400 = 2.800 \text{ mol}$$

$$[\text{Cl}^-] = \frac{2.800}{1.0} = \mathbf{2.80 \text{ M}}$$

Approach / Analogy

CaCl_2 has 2 chloride ions per formula unit — multiplier = 2. Total volume = 1 L. Quick calculation once you identify CaCl_2 as the only chloride source.

Common Mistake

Using multiplier = 1 for Cl^- from CaCl_2 . The formula CaCl_2 has **2** chloride ions per formula unit. Writing $1 \times 3.5 \times 0.4 = 1.4 \text{ M}$ is wrong; correct is $2 \times 3.5 \times 0.4 = 2.8 \text{ M}$.

Answer

(1) 2.80 M

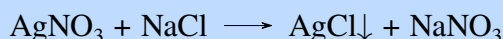
Q.76 What is the mass of the precipitate formed when 50 mL of 16.9% (w/v) solution of AgNO_3 is mixed with 50 mL of 5.8% (w/v) NaCl solution? ($A_g = 107.8$) [Re-AIPMT-2015]

Explanation

w/v: grams of solute per 100 mL solution.

$$\begin{aligned}\text{Mass of AgNO}_3 &= \frac{16.9}{100} \times 50 = 8.45 \text{ g} \\ n(\text{AgNO}_3) &= \frac{8.45}{170} = 0.0497 \approx 0.05 \text{ mol}\end{aligned}$$

$$\begin{aligned}\text{Mass of NaCl} &= \frac{5.8}{100} \times 50 = 2.9 \text{ g} \\ n(\text{NaCl}) &= \frac{2.9}{58.5} = 0.0496 \approx 0.05 \text{ mol}\end{aligned}$$



Both $\approx 0.05 \text{ mol}$: limiting reagent = either. Moles of $\text{AgCl} = 0.05 \text{ mol}$.

$$M_r(\text{AgCl}) = 107.8 + 35.5 = 143.3 \text{ g/mol}$$

$$\text{Mass of precipitate} = 0.05 \times 143.3 = 7.17 \approx 7 \text{ g}$$

Approach / Analogy

The precipitate is AgCl (white ppt). Steps: (1) find moles of both reactants, (2) identify limiting reagent, (3) moles of ppt = moles of limiting reagent, (4) mass = moles $\times M_r(\text{AgCl})$.

Common Mistake

Confusing w/v with w/w. w/v = g of solute per **100 mL** of solution. For 50 mL: mass = $(16.9/100) \times 50$. Students using density to convert volume to mass (as in w/w problems) get a wrong starting mass.

Answer

(1) 7 g

Q.77 Equal amounts of solutions of NaOH having concentration 10% (wt/wt) and 20% (wt/wt) are mixed. What will be the molality of NaOH in the resulting solution?

Explanation

Assume 100 g of each solution.

From 10%: NaOH = 10 g, H₂O = 90 g

From 20%: NaOH = 20 g, H₂O = 80 g

Total NaOH = 30 g, $n = 30/40 = 0.75$ mol

Total water = 170 g = 0.170 kg

$$m = \frac{0.75}{0.170} = 4.41 \approx 4.4 \text{ mol kg}^{-1}$$

Approach / Analogy

“Equal amounts” by mass \Rightarrow assume 100 g each (convenient base). Extract solute and solvent separately, combine, then apply molality formula. Note: mixing 10% and 20% gives 15% w/w but the molality is not the average of the two individual molalities.

Common Mistake

Averaging the two molalities directly. You must recompute molality from scratch after mixing — go back to solute moles and solvent kg separately. The average only works for simple cases, not molality mixing.

Answer

(2) 4.4 m

TYPE 9 : Cross-Concentration (Interconversion)

TYPE 9(i): Molarity \longleftrightarrow Molality

Standard Approach (take 1 L of solution):

Mass of solution = $\rho \times 1000$ g \Rightarrow Mass of solute = $n \times M_r$ \Rightarrow Mass of solvent = difference \Rightarrow molality.

Q.78 The molality of 1.2 M H₂SO₄ solution (density = 1.4 g/mL) is

Explanation

In 1 L of solution:

Mass of solution = $1000 \times 1.4 = 1400$ g

Mass of H₂SO₄ = $1.2 \times 98 = 117.6$ g

Mass of water = $1400 - 117.6 = 1282.4$ g = 1.2824 kg

$$m = \frac{1.2}{1.2824} = 0.936 \text{ mol kg}^{-1}$$

Approach / Analogy

Think of it as: “weigh the full bucket (1400 g), remove the salt (117.6 g), weigh just the water (1282.4 g), then compute molality.” This 4-step template works for every $M \rightarrow m$ conversion.

Common Mistake

Dividing by mass of **solution** (1400 g) instead of mass of **solvent** (1282 g). Molality denominator = solvent mass only. Forgetting to subtract solute mass is the most common error here.

Answer

(1) 0.936 m

Q.79 Density of a 2.05 M solution of acetic acid in water is 1.02 g mL^{-1} . The molality of the solution is

Explanation

$$\text{Mass of 1 L solution} = 1000 \times 1.02 = 1020 \text{ g}$$

$$\text{Mass of CH}_3\text{COOH} = 2.05 \times 60 = 123 \text{ g}$$

$$\text{Mass of water} = 1020 - 123 = 897 \text{ g} = 0.897 \text{ kg}$$

$$m = \frac{2.05}{0.897} = \mathbf{2.285} \approx \mathbf{2.28} \text{ mol kg}^{-1}$$

Approach / Analogy

Same 4-step template: density \times 1000 \rightarrow solution mass \rightarrow subtract solute mass \rightarrow solvent in kg \rightarrow molality. M_r of acetic acid $\text{CH}_3\text{COOH} = 60 \text{ g/mol}$.

Answer

(3) 2.28 mol kg^{-1}

Q.80 The density of 2 M aqueous solution of NaOH is 1.28 g/cm^3 . The molality of the solution is [NEET-2019 (Odisha)]

Explanation

$$\text{Mass of 1 L solution} = 1000 \times 1.28 = 1280 \text{ g}$$

$$\text{Mass of NaOH} = 2 \times 40 = 80 \text{ g}$$

$$\text{Mass of water} = 1280 - 80 = 1200 \text{ g} = 1.2 \text{ kg}$$

$$m = \frac{2}{1.2} = \mathbf{1.667} \approx \mathbf{1.67} \text{ mol kg}^{-1}$$

Approach / Analogy

Clean NEET numbers. $2 \text{ mol NaOH} \times 40 = 80 \text{ g}$. Water = $1200 \text{ g} = 1.2 \text{ kg}$. $m = 2/1.2 = 1.67 \text{ m}$.

Answer

(4) 1.67 m

Q.81 The density of 3 M solution of sodium chloride is 1.252 g mL^{-1} . The molality of the solution will be [JEE-Main-online-2013]

Explanation

$$\text{Mass of 1 L solution} = 1000 \times 1.252 = 1252 \text{ g}$$

$$\text{Mass of NaCl} = 3 \times 58.5 = 175.5 \text{ g}$$

$$\text{Mass of water} = 1252 - 175.5 = 1076.5 \text{ g} = 1.0765 \text{ kg}$$

$$m = \frac{3}{1.0765} = \mathbf{2.786} \approx \mathbf{2.79} \text{ mol kg}^{-1}$$

Approach / Analogy

JEE question, same template. No new concept — just careful arithmetic.

Answer

(4) 2.79 m

TYPE 9(ii): Mole Fraction \longleftrightarrow Molality**Key trick:**

Molality \rightarrow Mole Fraction: take 1 kg solvent as base $\Rightarrow n_{\text{solvent}} = 1000/M_{r,\text{solvent}}$

Mole Fraction \rightarrow Molality: take 1 mol total as base \Rightarrow convert solvent moles to kg.

Q.82 If we have 10 molal urea solution (MW = 60), the mole fraction of urea is

Explanation

10 molal = 10 mol urea per 1000 g water.

$$n_{\text{urea}} = 10 \text{ mol}$$

$$n_{\text{H}_2\text{O}} = \frac{1000}{18} = 55.56 \text{ mol}$$

$$\chi_{\text{urea}} = \frac{10}{10 + 55.56} = \frac{10}{65.56} = \mathbf{0.1526}$$

Approach / Analogy

Convert 1 kg of water to moles ($1000/18 = 55.56 \text{ mol}$). Add solute moles (10). Compute fraction. The 1 kg solvent base is the key starting point for molality \rightarrow mole fraction.

Common Mistake

Writing $n_{\text{H}_2\text{O}} = 1000$ instead of $1000/18 = 55.56$. Using mass (1000) instead of moles (55.56) for water gives a completely wrong denominator. **Always convert water mass to moles using $M_r = 18$.**

Answer

(1) 0.1526

Q.83 Calculate molality of aqueous urea solution which has $\chi_{\text{urea}} = 0.2$

Explanation

$\chi_{\text{urea}} = 0.2 \Rightarrow \chi_{\text{H}_2\text{O}} = 0.8$. Take 1 mol total: 0.2 mol urea + 0.8 mol water.

Mass of water = $0.8 \times 18 = 14.4 \text{ g} = 0.0144 \text{ kg}$

$$m = \frac{0.2}{0.0144} = \mathbf{13.88 \text{ mol kg}^{-1}}$$

Approach / Analogy

Take total = 1 mol as base. Convert water moles to kg ($0.8 \times 18 = 14.4 \text{ g} = 0.0144 \text{ kg}$). Molality = moles of solute / kg solvent. Clean and fast.

Answer

(3) 13.88 m

Q.84 An X molal solution of a compound in benzene has mole fraction of solute equal to 0.2. The value of X is

Explanation

Solvent = **benzene** ($M_r = 78 \text{ g/mol}$). Take 1 mol total: 0.2 mol solute + 0.8 mol benzene.

Mass of benzene = $0.8 \times 78 = 62.4 \text{ g} = 0.0624 \text{ kg}$

$$m = \frac{0.2}{0.0624} = \mathbf{3.205 \approx 3.2}$$

Approach / Analogy

Same method as Q.83 but solvent is benzene ($M_r = 78$) not water ($M_r = 18$). The molar mass of solvent matters — benzene is heavier, so the same moles of benzene weigh more than water.

Common Mistake

Using $M_r = 18$ (water) for solvent when solvent is benzene ($M_r = 78$). This gives $m = 0.2/(0.8 \times 0.018) = 13.88 \text{ m}$, which is the answer for aqueous solution — completely wrong for benzene.

Answer

(2) 3.2

Q.85 Mole fraction of the solute in a 1.00 molal aqueous solution is [AIPMT (Prelims)-2011]

Explanation

1 mol solute per 1000 g water.

$$n_{\text{H}_2\text{O}} = \frac{1000}{18} = 55.56 \text{ mol}$$

$$\chi_{\text{solute}} = \frac{1}{1 + 55.56} = \frac{1}{56.56} = \mathbf{0.01768} \approx \mathbf{0.0177}$$

Approach / Analogy

A 1 molal solution is quite dilute — 1 mol of solute vs 55.56 mol of water. So mole fraction is small: 0.0177. Sanity check: 10 molal gave 0.15 (Q.82), so 1 molal should give roughly 10 times less \approx 0.015. Close enough.

Answer

(3) 0.0177

Q.86 The mole fraction of a solvent in aqueous solution of a solute is 0.8. The molality (in mol kg⁻¹) of the aqueous solution is [JEE-Main(april)-2019]

Explanation

$\chi_{\text{solvent}} = 0.8 \Rightarrow \chi_{\text{solute}} = 0.2$. Take 1 mol total: 0.2 mol solute + 0.8 mol water.

$$\text{Mass of water} = 0.8 \times 18 = 14.4 \text{ g} = 0.0144 \text{ kg}$$

$$m = \frac{0.2}{0.0144} = \mathbf{13.88} \text{ mol kg}^{-1}$$

Same as Q.83 — identical calculation.

Approach / Analogy

The question gives mole fraction of **solvent** (0.8), not solute. Solute mole fraction = $1 - 0.8 = 0.2$. Then identical to Q.83.

Common Mistake

Using 0.8 as the **solute** mole fraction instead of the solvent fraction. The question says “mole fraction of **solvent** = 0.8.” Solute fraction = $1 - 0.8 = 0.2$. This mix-up completely flips the entire calculation.

Answer

(3) 13.88 mol kg⁻¹

Q.87 A 5.2 molal aqueous solution of methyl alcohol, CH₃OH, is supplied. What is the mole fraction of methyl alcohol in the solution? [AIEEE-2011]

Explanation

$$\begin{aligned}n_{\text{CH}_3\text{OH}} &= 5.2 \text{ mol} \\n_{\text{H}_2\text{O}} &= \frac{1000}{18} = 55.56 \text{ mol} \\ \chi_{\text{CH}_3\text{OH}} &= \frac{5.2}{5.2 + 55.56} = \frac{5.2}{60.76} = \mathbf{0.0856} \approx \mathbf{0.086}\end{aligned}$$

Approach / Analogy

Standard molality \rightarrow mole fraction: take 1 kg water, convert to moles (55.56), add solute moles (5.2), compute fraction.

Answer

(1) 0.086

Q.88 Mole fraction of solvent in aqueous solution of NaOH having molality of 3 is

Explanation

3 molal: 3 mol NaOH per 1000 g water.

$$\begin{aligned}n_{\text{H}_2\text{O}} &= \frac{1000}{18} = 55.56 \text{ mol} \\ \chi_{\text{H}_2\text{O}} &= \frac{55.56}{55.56 + 3} = \frac{55.56}{58.56} = \mathbf{0.949} \approx \mathbf{0.95}\end{aligned}$$

Approach / Analogy

The question asks for **solvent** (water) mole fraction. After computing: $\chi_{\text{NaOH}} \approx 0.051$, so $\chi_{\text{H}_2\text{O}} = 1 - 0.051 = 0.949 \approx 0.95$.

Common Mistake

Reporting the solute (NaOH) mole fraction instead of the solvent (water) mole fraction. The question explicitly says “mole fraction of **solvent**.” Read the question carefully before reporting the answer.

Answer

(4) 0.95

TYPE 9(iii): % \longleftrightarrow Molarity

Key Formula (w/w \rightarrow M):

$$M = \frac{\% \times \rho \times 10}{M_r} \quad \text{where } \% = \text{w/w}, \rho \text{ in g/mL}$$

For w/v: scale directly to per litre (multiply g/100 mL by 10), then divide by M_r .

Q.89 A sample of H_2SO_4 (density 1.8 g mL^{-1}) is labelled as 74.66% by weight. What is the molarity of the acid?

Explanation

$$M = \frac{74.66 \times 1.8 \times 10}{98} = \frac{1343.88}{98} = 13.71 \approx 14 \text{ M}$$

Approach / Analogy

One-line formula: $M = (\% \times \rho \times 10) / M_r$. The factor of 10 comes from converting % per 100 g to per 1000 mL (per litre).

Common Mistake

Forgetting the factor of 10 in the formula, writing $M = (\% \times \rho) / M_r$. This gives 1.37 instead of 13.7 — ten times too small. The factor of 10 is essential for unit consistency.

Answer

(1) 14 M

Q.90 The molarity of 93% (w/w) H_2SO_4 solution (density = 1.84 g/mL) is

Explanation

$$M = \frac{93 \times 1.84 \times 10}{98} = \frac{1711.2}{98} = 17.46 \text{ M}$$

Approach / Analogy

93% w/w concentrated H_2SO_4 (lab grade) has molarity $\approx 17.5 \text{ M}$. Worth memorising as a benchmark. Density 1.84 g/mL is also a standard value for conc. H_2SO_4 .

Answer

(1) 17.46 M

Q.91 Molarity of 29% (W/W) H_2SO_4 solution whose density is 1.22 g mL^{-1} is [NCERT Pg. 23]

Explanation

$$M = \frac{29 \times 1.22 \times 10}{98} = \frac{353.8}{98} = 3.61 \approx 3.6 \text{ M}$$

Approach / Analogy

NCERT standard. Direct formula plug-in. Notice density = 1.22 g/mL (less than 1.84 in Q.90) because this is a less concentrated solution.

Answer

(2) 3.6 M

Q.92 Concentrated aqueous sulphuric acid is 98% H_2SO_4 (w/v) and has a density of 1.80 gm L^{-1} . Molarity of solution is

Explanation

Note: w/v means grams per 100 mL — density is irrelevant here!

98 g H_2SO_4 per 100 mL = 980 g per 1000 mL (per litre).

$$M = \frac{980}{98} = 10 \text{ M}$$

Approach / Analogy

w/v percent directly gives g/100 mL. Scale to per litre: multiply by 10. Then divide by M_r . Do NOT use the density in the w/v formula — density is a red herring (or possibly a misprint in the question).

Common Mistake

Applying the w/w formula $M = \% \rho \times 10 / M_r$ when w/v percent is given. Using $98 \times 1.8 \times 10 / 98 = 18 \text{ M}$ is wrong because density is already embedded in w/v. For w/v: ignore density, just scale to per litre directly.

Answer

(3) 10 M

Q.93 The concentrated sulphuric acid that is peddled commercially is 95% H_2SO_4 by weight. If the density of this commercial acid is 1.834 g cm^{-3} , the molarity of this solution is [JEE-(Main)-2012]

Explanation

$$M = \frac{95 \times 1.834 \times 10}{98} = \frac{1742.3}{98} = 17.78 \approx 17.8 \text{ M}$$

Approach / Analogy

JEE standard numerical. 95% w/w conc. H_2SO_4 with density 1.834 g/cm^3 gives $\approx 17.8 \text{ M}$. Consistent with Q.90 (93% gives 17.46 M). Answer makes sense.

Answer

(1) 17.8 M

Q.94 How many grams of concentrated nitric acid solution should be used to prepare 250 mL of 2 M HNO_3 ? The concentrated acid is 70% (w/w) HNO_3 . [NEET-2013]

Explanation

$$\begin{aligned}n(\text{HNO}_3) &= 2 \times 0.250 = 0.5 \text{ mol} \\ \text{Mass of pure HNO}_3 &= 0.5 \times 63 = 31.5 \text{ g} \\ \text{Mass of conc. acid} &= \frac{31.5}{0.70} = \mathbf{45 \text{ g}}\end{aligned}$$

Approach / Analogy

Like needing 31.5 g of pure gold but the alloy is only 70% gold — you need $31.5/0.70 = 45$ g of alloy. The 70% purity means you need MORE of the concentrated solution to get the required pure acid.

Common Mistake

Reporting 31.5 g (mass of pure HNO_3) as the answer. The question asks for mass of **concentrated acid solution** (70% purity), not pure acid. Divide by purity fraction to get the total solution mass needed.

Answer

(4) 45.0 g

Q.95 Concentrated aqueous sulphuric acid is 98% H_2SO_4 by mass and has a density of 1.80 g mL^{-1} . Volume of acid required to make one litre of 0.1 M H_2SO_4 solution is

Explanation

Step 1: Molarity of concentrated acid.

$$M_{\text{conc}} = \frac{98 \times 1.80 \times 10}{98} = 18 \text{ M}$$

Step 2: Dilution.

$$18 \times V_1 = 0.1 \times 1000 \implies V_1 = \frac{100}{18} = \mathbf{5.55 \text{ mL}}$$

Approach / Analogy

Two steps: (1) find molarity of conc. acid using w/w formula, (2) apply $M_1V_1 = M_2V_2$ to find volume needed. Classic lab prep question.

Answer

(3) 5.55 mL

Q.96 What volume of 63% HNO_3 (wt/wt) having density 1.4 g mL^{-1} is required to prepare 200 mL of 0.7 M HNO_3 solution?

Explanation

Step 1: Molarity of conc. HNO_3 ($M_r = 63$):

$$M = \frac{63 \times 1.4 \times 10}{63} = 14 \text{ M}$$

Step 2: Dilution:

$$14 \times V_1 = 0.7 \times 200 \implies V_1 = \frac{140}{14} = \mathbf{10 \text{ mL}}$$

Approach / Analogy

Neat shortcut: when % value numerically equals M_r (both = 63 here for HNO_3), molarity = $\rho \times 10 = 1.4 \times 10 = 14 \text{ M}$. The percentage and molar mass cancel out!

Answer

(4) 10 mL

TYPE 9(iv): % \longleftrightarrow Other Concentration Terms

Q.97 If we have 10 molal urea solution ($MW = 60$), the % w/w of urea is

Explanation

10 molal = 10 mol urea per 1 kg water.

$$\text{Mass of urea} = 10 \times 60 = 600 \text{ g}$$

$$\text{Mass of solution} = 600 + 1000 = 1600 \text{ g}$$

$$\% \text{ w/w} = \frac{600}{1600} \times 100 = \mathbf{37.5\%}$$

Approach / Analogy

Start from molality definition (10 mol per 1 kg solvent), convert to masses, compute %. No formula to memorise — just unit logic.

Answer

(1) 37.5%

Q.98 Calculate the mole percentage of CH_3OH and H_2O respectively in 60% (by mass) aqueous solution of CH_3OH .

Explanation

In 100 g solution: 60 g CH₃OH + 40 g H₂O.

$$n_{\text{CH}_3\text{OH}} = \frac{60}{32} = 1.875 \text{ mol}$$

$$n_{\text{H}_2\text{O}} = \frac{40}{18} = 2.222 \text{ mol}$$

$$n_{\text{total}} = 4.097 \text{ mol}$$

$$\text{Mole \% of CH}_3\text{OH} = \frac{1.875}{4.097} \times 100 = \mathbf{45.8\%}$$

$$\text{Mole \% of H}_2\text{O} = \frac{2.222}{4.097} \times 100 = \mathbf{54.2\%}$$

Approach / Analogy

60% by mass but only 45.8% by moles — because methanol ($M_r = 32$) is heavier than water ($M_r = 18$), so the same mass of methanol contains fewer molecules. Mass % \neq mole % whenever molecular weights differ.

Common Mistake

Reporting 60% as the mole percent of CH₃OH. Mass percent and mole percent are completely different when M_r values differ. Convert to moles first, always.

Answer

(1) 45.8, 54.2

Q.99 What will be the density (in g mL⁻¹) of 3.60 molar sulphuric acid having 29% by mass? (Molar mass = 98 g mol⁻¹)

Explanation

Rearranging $M = \frac{\% \times \rho \times 10}{M_r}$ for density:

$$\rho = \frac{M \times M_r}{\% \times 10} = \frac{3.60 \times 98}{29 \times 10} = \frac{352.8}{290} = \mathbf{1.216 \approx 1.22 \text{ g/mL}}$$

Approach / Analogy

This is the standard formula rearranged to find density. Given M and %, solve for ρ . Write the formula, isolate the unknown, substitute. The formula triangle: any one of M, ρ , % can be the unknown.

Common Mistake

Not rearranging — trying to plug density as a known into the right side and getting confused. Write the formula, circle the unknown (density), rearrange algebraically before substituting numbers.

Answer

(2) 1.22 g/mL

Q.100 A 6.90 M solution of KOH in water has 30% by weight of KOH. The density of the solution is

Explanation

$M_r(\text{KOH}) = 39 + 16 + 1 = 56 \text{ g/mol}$.

$$\rho = \frac{M \times M_r}{\% \times 10} = \frac{6.90 \times 56}{30 \times 10} = \frac{386.4}{300} = \mathbf{1.288 \text{ g/mL}}$$

Approach / Analogy

Same reverse-density template as Q.99. Just plug in KOH values.

Answer

(2) 1.288 g/mL

Q.101 What would be the molality of 20% (mass/mass) aqueous solution of KI? (molar mass of KI = 166 g mol⁻¹) [JEE-Main(april)-2019]

Explanation

In 100 g solution: 20 g KI + 80 g water.

$$n_{\text{KI}} = \frac{20}{166} = 0.1205 \text{ mol}$$
$$m = \frac{0.1205}{0.080} = \mathbf{1.506 \approx 1.51 \text{ mol kg}^{-1}}$$

Approach / Analogy

% w/w → molality: assume 100 g solution, split into solute (20 g) and solvent (80 g), convert each appropriately and apply molality formula. $M_r(\text{KI}) = 166 \text{ g/mol}$.

Answer

(3) 1.51

Q.102 Mole fraction of ethanol in ethanol-water mixture is 0.25. Hence percentage concentration of ethanol by weight of mixture is

Explanation

Take 1 mol total: 0.25 mol ethanol + 0.75 mol water.

$$\text{Mass of ethanol} = 0.25 \times 46 = 11.5 \text{ g}$$

$$\text{Mass of water} = 0.75 \times 18 = 13.5 \text{ g}$$

$$\text{Total} = 25.0 \text{ g}$$

$$\% = \frac{11.5}{25.0} \times 100 = \mathbf{46\%}$$

Approach / Analogy

Mole fraction 0.25 gives mass percent 46% — not 25%! Because ethanol ($M_r = 46$) is heavier than water ($M_r = 18$), the same mole fraction gives a higher mass fraction.

Common Mistake

Reporting 25% (mole fraction $\times 100$). Mole fraction \neq mass fraction. Converting to masses is mandatory before computing mass percent.

Answer

(3) 46%

TYPE 9 Misc.: Miscellaneous Interconversions

Q.103 The molarity of $\text{CaCO}_3(\text{aq.})$ solution which has concentration of $\text{CaCO}_3 = 200$ ppm is

Explanation

200 ppm (mass/mass) = 200 g per 10^6 g solution. Assuming dilute (density ≈ 1 g/mL):

$$200 \text{ ppm} = \frac{200 \text{ g}}{10^6 \text{ mL}} = \frac{0.2 \text{ g}}{1000 \text{ mL}} = 0.2 \text{ g/L}$$

$$M = \frac{0.2}{100} = 2 \times 10^{-3} \text{ M}$$

($M_r(\text{CaCO}_3) = 100$ g/mol)

Approach / Analogy

ppm \rightarrow molarity: convert ppm to g/L (for dilute solution: 1 ppm = 0.001 g/L), then divide by molar mass. 200 ppm = 0.2 g/L. With $M_r = 100$: $M = 0.2/100 = 0.002 \text{ M} = 2 \times 10^{-3} \text{ M}$.

Answer

(1) $2 \times 10^{-3} \text{ M}$

Q.104 The density of a solution prepared by dissolving 120 g of urea (mol. mass = 60 u) in 1000 g of water is 1.15 g/mL. The molarity of this solution is [JEE-(Main)-2012]

Explanation

$$n_{\text{urea}} = \frac{120}{60} = 2 \text{ mol}$$

$$\text{Mass of solution} = 120 + 1000 = 1120 \text{ g}$$

$$V = \frac{1120}{1.15} = 973.9 \text{ mL} = 0.9739 \text{ L}$$

$$M = \frac{2}{0.9739} = 2.054 \approx 2.05 \text{ M}$$

Approach / Analogy

Given masses + density: total mass → divide by density to get volume → apply molarity formula. Volume of solution is NOT 1000 mL (that's the mass of water, not solution volume).

Common Mistake

Assuming volume = 1000 mL (= mass of water). The solution volume = total mass ÷ density = 1120/1.15 = 973.9 mL, which is **less** than 1000 mL. Solute dissolving compresses total volume slightly.

Answer

(1) 2.05 M

Q.105 The density of 3 M solution of $\text{Na}_2\text{S}_2\text{O}_3$ is 1.25 g mL^{-1} . Calculate the % by weight of $\text{Na}_2\text{S}_2\text{O}_3$ and mole fraction of $\text{Na}_2\text{S}_2\text{O}_3$ respectively.

Explanation

$$M_r(\text{Na}_2\text{S}_2\text{O}_3) = 2(23) + 2(32) + 3(16) = 46 + 64 + 48 = 158 \text{ g/mol.}$$

In 1 L solution:

$$\text{Mass of solution} = 1000 \times 1.25 = 1250 \text{ g}$$

$$\text{Mass of } \text{Na}_2\text{S}_2\text{O}_3 = 3 \times 158 = 474 \text{ g}$$

$$\text{Mass of water} = 1250 - 474 = 776 \text{ g}$$

$$\% = \frac{474}{1250} \times 100 = \mathbf{37.92\%}$$

$$n_{\text{Na}_2\text{S}_2\text{O}_3} = 3 \text{ mol}, \quad n_{\text{H}_2\text{O}} = \frac{776}{18} = 43.11 \text{ mol}$$

$$\chi = \frac{3}{3 + 43.11} = \frac{3}{46.11} = \mathbf{0.065}$$

Approach / Analogy

Two outputs from one 1 L basis. Find all masses first (solution, solute, solvent), then independently compute % and mole fraction. The “1 L basis” is the universal tool for cross-conversion problems involving molarity + density.

Answer

(1) 37.92% and 0.065

Q.106 10 mL of sulphuric acid solution (sp. gr. = 1.84) contains 98% by weight of pure acid. The volume of 2.5 M NaOH solution required to just neutralise the acid is

Explanation

Step 1: Molarity of H_2SO_4 :

$$M = \frac{98 \times 1.84 \times 10}{98} = 18.4 \text{ M}$$

Step 2: Moles of H_2SO_4 in 10 mL:

$$n = 18.4 \times 0.010 = 0.184 \text{ mol}$$

Step 3: $\text{H}_2\text{SO}_4 + 2\text{NaOH} \longrightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$

$$n(\text{NaOH}) = 2 \times 0.184 = 0.368 \text{ mol}$$

Step 4: Volume of 2.5 M NaOH:

$$V = \frac{0.368}{2.5} = 0.1472 \text{ L} = \mathbf{147.2 \text{ mL}}$$

Approach / Analogy

Four-step problem: (1) find molarity of conc. acid, (2) find moles in 10 mL, (3) use stoichiometry (1:2 ratio for H_2SO_4 :NaOH), (4) find volume of NaOH needed.

Common Mistake

Using 1:1 stoichiometry (H_2SO_4 : NaOH) instead of 1:2. H_2SO_4 is a **diprotic acid** — it needs 2 moles of NaOH per mole. Using 1:1 gives exactly half the correct answer: 73.6 mL instead of 147.2 mL.

Answer

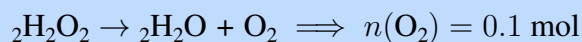
(1) 147.2 mL

Q.107 The concentration of a solution of H_2O_2 is 6.8% (w/V). Then the volume concentration of the solution is

Explanation

6.8 g H_2O_2 per 100 mL solution.

$$n(\text{H}_2\text{O}_2) = \frac{6.8}{34} = 0.2 \text{ mol per 100 mL}$$



$$V(\text{O}_2) = 0.1 \times 22400 \text{ mL} = 2240 \text{ mL (at STP)}$$

$$\text{Volume strength} = \frac{V(\text{O}_2)}{V(\text{solution})} = \frac{2240}{100} = \mathbf{22.4}$$

Approach / Analogy

Volume strength of H_2O_2 = mL of O_2 (at STP) released per mL of solution. Quick formula: Volume strength = $11.2 \times M(\text{H}_2\text{O}_2)$. Here: $M = 6.8/(34 \times 0.1) = 2 \text{ M}$, so volume strength = $11.2 \times 2 = 22.4$.

Common Mistake

Ignoring the 1:2 stoichiometry ($2\text{H}_2\text{O}_2$ gives only 1O_2) and writing $n(\text{O}_2) = n(\text{H}_2\text{O}_2) = 0.2$ mol. This gives volume strength 44.8 instead of 22.4 — exactly double the correct answer.

Answer

(1) 22.4

Note

Calculation gives volume concentration = 22.4 (option 1). If your answer key shows option (3) = 20, it likely uses a slightly different STP molar volume (22.0 L/mol or rounds differently). Standard STP gives 22.4. Please verify with your source.

TYPE 10 : Normality

Key Concepts:

$$N = M \times n\text{-factor} \quad \text{Gram equivalents} = N \times V(\text{L}) = \frac{\text{mass}}{\text{Eq. weight}}$$

n-factors: $\text{H}_2\text{SO}_4 = 2$, $\text{HCl}/\text{HNO}_3 = 1$, $\text{H}_3\text{PO}_4 = 3$, $\text{H}_3\text{PO}_3 = 2$ (diprotic!), $\text{H}_3\text{PO}_2 = 1$ (monoprotic!), $\text{NaOH}/\text{KOH} = 1$, $\text{Na}_2\text{CO}_3/\text{Ba}(\text{OH})_2 = 2$.

Equivalent weight = Molar mass \div n-factor

Q.108 The normality of 1.5 M H_3PO_4 is

Explanation

H_3PO_4 is triprotic — n-factor = 3.

$$N = M \times n = 1.5 \times 3 = \mathbf{4.5\ N}$$

Approach / Analogy

H_3PO_4 has 3 ionisable H atoms \Rightarrow n-factor = 3. Normality = 3 \times molarity. Think of n-factor as the multiplier based on reactive H^+ (or OH^-) provided per molecule.

Answer

(1) 4.5 N

Q.109 What is the normality of 1 M H_3PO_4 solution?

Explanation

n-factor of $\text{H}_3\text{PO}_4 = 3$.

$$N = 1 \times 3 = \mathbf{3\ N}$$

Approach / Analogy

Same concept, $M = 1$. $N = 3 \times 1 = 3 \text{ N}$.

Answer

(4) 3.0 N

Q.110 The molarity of 0.2 N Na_2CO_3 solution will be

Explanation

Na_2CO_3 n-factor = 2 (reacts with 2 H^+).

$$M = \frac{N}{n} = \frac{0.2}{2} = \mathbf{0.1 \text{ M}}$$

Approach / Analogy

Reverse: $N \rightarrow M$. Divide by n-factor. Na_2CO_3 reacts with 2 H^+ per formula unit, so n-factor = 2.

Common Mistake

Using n-factor = 1 for Na_2CO_3 . It reacts with 2 H^+ (diprotic base), giving n-factor = 2. $M = 0.2/1 = 0.2 \text{ M}$ is wrong; correct is $0.2/2 = 0.1 \text{ M}$.

Answer

(3) 0.1 M

Q.111 What is the molarity of 1 N H_2SO_4 solution?

Explanation

n-factor of $\text{H}_2\text{SO}_4 = 2$.

$$M = \frac{1}{2} = \mathbf{0.5 \text{ M}}$$

Approach / Analogy

1 N $\text{H}_2\text{SO}_4 = 0.5 \text{ M } \text{H}_2\text{SO}_4$. Important in titrations: 1 N acid always exactly neutralises 1 N base, regardless of identity.

Answer

(3) 0.5 M

Q.112 What is the normality of 2 M H_3PO_2 solution?

Explanation

H_3PO_2 (hypophosphorous acid): only **1 ionisable H** (the 2 H bonded directly to P are non-ionisable).
n-factor = 1.

$$N = 2 \times 1 = 2 \text{ N}$$

Approach / Analogy

H_3PO_2 looks triprotic but structurally has only 1 P–O–H bond (ionisable). The 2 H directly bonded to P cannot ionise. This is a classic “trap” question.

Common Mistake

Using n-factor = 3 for H_3PO_2 because it has 3 H atoms. Only H attached to O (as P–O–H) are ionisable. Structure: H_3PO_2 has 1 P–O–H \Rightarrow n-factor = 1. The phosphorus acids trio: H_3PO_4 (n=3), H_3PO_3 (n=2), H_3PO_2 (n=1).

Answer

(3) 2.0 N

Q.113 Normality of 0.3 M phosphorous acid is

Explanation

Phosphorous acid = H_3PO_3 : 2 ionisable H (2 P–O–H bonds, 1 P–H bond non-ionisable). n-factor = 2.

$$N = 0.3 \times 2 = 0.6 \text{ N}$$

Approach / Analogy

H_3PO_3 : 2 ionisable H, n-factor = 2. Remember the phosphorus acids: H_3PO_2 (n=1) < H_3PO_3 (n=2) < H_3PO_4 (n=3). Each step up adds one more ionisable H.

Common Mistake

Confusing H_3PO_3 (phosphorous acid, n=2) with H_3PO_4 (phosphoric acid, n=3). Their names differ by one letter (“ous” vs “ic”) but n-factors differ. Memorise the trio as a set.

Answer

(2) 0.6 N

Q.114 A 100 cm³ solution is prepared by dissolving 2 g of NaOH in water. The normality of the solution is

Explanation

n-factor of NaOH = 1. Eq. weight = 40 g/eq.

$$\text{Gram equivalents} = \frac{2}{40} = 0.05 \text{ eq}$$

$$N = \frac{0.05}{0.1} = 0.5 \text{ N} = \frac{N}{2}$$

Approach / Analogy

N/2 is the fractional notation for 0.5 N. For NaOH (n=1), normality = molarity. $M = 2/(40 \times 0.1) = 0.5 \text{ M} = 0.5 \text{ N}$.

Answer

(1) N/2

Q.115 Which of the following solutions has the highest normality?

Explanation

- 8 g KOH/L: $N = 8/56 = 0.143 \text{ N}$
- 1 N H₃PO₄: $N = 1.0 \text{ N}$
- 6 g NaOH/100 mL: $N = (6/40)/0.1 = 1.5 \text{ N}$
- 0.5 M H₂SO₄: $N = 0.5 \times 2 = 1.0 \text{ N}$

Highest = 1.5 N = option (3).

Approach / Analogy

Calculate all four and compare. Option (3) packs 6 g NaOH into only 100 mL = 1.5 N. Don't guess by inspection — compute each one systematically before choosing.

Common Mistake

Forgetting the n-factor multiplier for H₂SO₄ (×2) or miscalculating option (3) by using 1 L instead of 100 mL. Always re-check the volume unit in each option.

Answer

(3) 6 g of NaOH per 100 mL

Q.116 How many gram equivalents of H₂SO₄ are present in 200 mL of $\frac{N}{10}$ H₂SO₄ solution?

Explanation

$$\text{Gram equivalents} = N \times V(\text{L}) = \frac{1}{10} \times 0.200 = \mathbf{0.02 \text{ eq}}$$

Approach / Analogy

Like $n = M \times V$ for moles, here: gram eq = $N \times V$. $N/10 = 0.1 \text{ N}$. $0.1 \times 0.2 = 0.02 \text{ eq}$.

Answer

(1) 0.02

Q.117 The mass of H_2SO_4 in 150 mL of $\frac{N}{7}$ H_2SO_4 solution is**Explanation**Eq. weight of $\text{H}_2\text{SO}_4 = 98/2 = 49$ g/eq.

$$\text{Gram eq} = \frac{1}{7} \times 0.150 = \frac{0.150}{7} = 0.02143 \text{ eq}$$

$$\text{Mass} = 0.02143 \times 49 = \mathbf{1.05 \text{ g}}$$

Approach / AnalogyMass = gram eq \times eq. weight. Eq. weight of $\text{H}_2\text{SO}_4 = 49$ (half of 98 because $n=2$). Two-step: gram eq from $N \times V$, then multiply by eq. weight.**Answer**

(1) 1.05 g

Q.118 How many grams of dibasic acid (mol. weight 200) should be present in 100 mL of the aqueous solution to give strength of 0.1 N?**Explanation**Dibasic \Rightarrow n-factor = 2. Eq. weight = $200/2 = 100$ g/eq.

$$\text{Gram eq} = 0.1 \times 0.1 = 0.01 \text{ eq}$$

$$\text{Mass} = 0.01 \times 100 = \mathbf{1 \text{ g}}$$

Approach / AnalogyDibasic acid = 2 replaceable H = n-factor 2. Eq. weight = molar mass / $n = 200/2 = 100$ g/eq.**Answer**

(3) 1 g

Q.119 What volume of 0.1 N HNO_3 solution can be prepared from 6.3 g of HNO_3 ?**Explanation**n-factor of $\text{HNO}_3 = 1$. Eq. weight = 63 g/eq.

$$\text{Gram eq} = \frac{6.3}{63} = 0.1 \text{ eq}$$

$$V = \frac{0.1}{0.1} = \mathbf{1 \text{ L}}$$

Approach / Analogy

For monoprotic acids ($n=1$), gram equivalents = moles. Same as molarity calculation. $6.3 \text{ g} / 63 = 0.1 \text{ mol} = 0.1 \text{ gram eq}$. In 0.1 N solution: $0.1/0.1 = 1 \text{ L}$.

Answer

(1) 1 litre

Q.120 A 5 M solution of H_2SO_4 is diluted from 1 litre to a volume of 100 litres, the normality of the solution will be

Explanation

Initial normality: $N_1 = 5 \times 2 = 10 \text{ N}$.

$$N_1V_1 = N_2V_2 \implies 10 \times 1 = N_2 \times 100 \implies N_2 = \mathbf{0.1 \text{ N}}$$

Approach / Analogy

Dilution formula applies to normality too: $N_1V_1 = N_2V_2$. Convert M to N first ($5 \times 2 = 10 \text{ N}$), then dilute 100-fold to get 0.1 N.

Common Mistake

Using $N_1 = 5$ (the molarity value) without converting to normality first. For H_2SO_4 : $N = 2M$, so $N_1 = 10 \text{ N}$, not 5 N. Applying dilution with $N=5$ gives 0.05 N — half the correct answer.

Answer

(3) 0.1 N

Q.121 10 mL of $\frac{N}{2}$ HCl, 30 mL of $\frac{N}{10}$ HNO_3 and 75 mL of $\frac{N}{5}$ H_2SO_4 are mixed together. The normality of the resulting solution is

Explanation

$$\text{Gram eq from HCl: } 0.5 \times 0.010 = 0.005 \text{ eq}$$

$$\text{Gram eq from HNO}_3: 0.1 \times 0.030 = 0.003 \text{ eq}$$

$$\text{Gram eq from H}_2\text{SO}_4: 0.2 \times 0.075 = 0.015 \text{ eq}$$

$$V_{\text{total}} = 115 \text{ mL} = 0.115 \text{ L}$$

$$N_f = \frac{0.005 + 0.003 + 0.015}{0.115} = \frac{0.023}{0.115} = \mathbf{0.2 \text{ N}}$$

Approach / Analogy

Mixing acids of different types: normality mixing formula works identically to molarity mixing. Total gram equivalents divided by total volume. The power of normality — different acids handled uniformly.

Common Mistake

Misreading N/2 as 2 N (instead of 0.5 N). N/x notation: N/2 = 0.5 N, N/10 = 0.1 N, N/5 = 0.2 N. Reading these backwards inflates gram equivalents massively and gives a wrong answer.

Answer

(3) 0.2 N

Q.122 100 mL decinormal HCl is mixed with 100 mL seminormal H₂SO₄ solution. The normality of the resulting mixture is

Explanation

Decinormal = N/10 = 0.1 N. Seminormal = N/2 = 0.5 N.

Gram eq (HCl): $0.1 \times 0.1 = 0.01$ eq

Gram eq (H₂SO₄): $0.5 \times 0.1 = 0.05$ eq

$$N_f = \frac{0.01 + 0.05}{0.2} = \frac{0.06}{0.2} = \mathbf{0.30\ N}$$

Approach / Analogy

Lab vocabulary: decinormal = 0.1 N (deci = one-tenth), seminormal = 0.5 N (semi = half). Equal volumes here, but the molarities differ so we can't just average — use the full gram equivalents formula.

Common Mistake

Thinking “seminormal = 2 N” (double). Semi means **half**, not double. Seminormal = 0.5 N. Using 2 N gives a completely wrong answer.

Answer

(1) 0.30 N

Q.123 Calculate normality of 2.1% (w/V) H₂SO₄ solution.

Explanation

w/v: 2.1 g per 100 mL = 21 g per litre.

Eq. weight of H₂SO₄ = 49 g/eq.

$$N = \frac{21}{49} = \mathbf{0.4286 \approx 0.428\ N}$$

Approach / Analogy

Normality from w/v: scale to per litre (21 g/L), divide by **equivalent weight** (not molar mass). Eq. weight = $M_r/n = 98/2 = 49$ g/eq.

Answer

(3) 0.428 N

Q.124 For 1% H₂SO₄ solution find its normality (approx.)

Explanation

Dilute solution: density ≈ 1 g/mL, so 1% w/v $\approx 1\%$ w/w. Scale to 1 L: 10 g H₂SO₄/L.

$$N \approx \frac{10}{49} \approx 0.204 \approx \mathbf{0.2\ N}$$

Approach / Analogy

For dilute solutions, density ≈ 1 g/mL, so w/w \approx w/v. Quick estimate: 1% ≈ 10 g/L. Divide by eq. weight (49) to get normality ≈ 0.2 N.

Answer

(2) 0.2 N

Q.125 Na₂SO₄ is added in water to make 12% w/w solution having density 1.2 g/mL. Correct statement(s) is/are

Explanation

$M_r(\text{Na}_2\text{SO}_4) = 142$ g/mol. In 100 g solution: 12 g Na₂SO₄ + 88 g water.

Statement (a) — Molality:

$$m = \frac{(12/142)}{0.088} = \frac{0.08451}{0.088} = 0.96\ \text{m} \approx 0.96\ \text{mol/kg} \quad \checkmark$$

Statement (b) — Normality of Na⁺ vs SO₄²⁻:

By electroneutrality: total positive charge = total negative charge in any solution.

Normality of Na⁺ (charge +1, 2 per formula unit) = Normality of SO₄²⁻ (charge 2-, 1 per formula unit). Charge equivalents balance \Rightarrow **TRUE** under electroneutrality interpretation.

Statement (c) — Molarity vs Molality:

$$V = \frac{100}{1.2} = 83.33\ \text{mL} \implies M = \frac{12/142}{0.08333} = 1.01\ \text{M}$$

Molarity (1.01 M) > Molality (0.96 m). Statement (c) says M < m — **FALSE**.

Correct: (a) and (b).

Approach / Analogy

Multi-statement question — tackle independently. The key insight for (b): electroneutrality principle ensures that the total charge from cations = total charge from anions (expressed as normality, i.e., equivalents/L). For (c): since density > 1, molarity > molality (solution volume < water mass in mL).

Common Mistake

For statement (c): when density > 1 g/mL, volume of solution < mass in mL, so molarity (moles/L) > molality (moles/kg solvent). This counterintuitive result trips many students. Compute both M and m numerically to confirm rather than guessing.

Answer

(4) Both (a) and (b)

Q.126 One litre of N/2 HCl solution is heated in a beaker. It was observed that when the volume of the solution was reduced to 600 mL, 3.25 g of HCl was lost. Calculate the normality of the new solution.

Explanation

n-factor of HCl = 1, so N/2 = 0.5 N = 0.5 M.

$$\text{Initial moles HCl} = 0.5 \times 1 = 0.5 \text{ mol}$$

$$\text{Initial mass HCl} = 0.5 \times 36.5 = 18.25 \text{ g}$$

After heating:

$$\text{Mass HCl remaining} = 18.25 - 3.25 = 15.0 \text{ g}$$

$$\text{Moles HCl remaining} = \frac{15.0}{36.5} = 0.4110 \text{ mol}$$

$$N_f = M_f = \frac{0.4110}{0.6} = \mathbf{0.685 \text{ N}}$$

Approach / Analogy

HCl is a volatile acid — it escapes along with steam. Both moles (numerator) and volume (denominator) decrease when heated. Unlike evaporating a NaCl solution where only water leaves, here HCl itself is lost. Track both changes separately.

Common Mistake

Assuming only water evaporates and keeping moles of HCl at 0.5 mol. This gives $N = 0.5/0.6 = 0.833 \text{ N}$ — wrong. HCl is **volatile**: it escapes too. The 3.25 g HCl loss must be subtracted from initial solute mass before computing final normality.

Answer

(2) 0.685 N

— End of DPP-2 Complete Solution Sheet —

Concentration Terms · Q.1–Q.126 · All Parts Complete

“The difference between average and confident students is assignment completion.”