JEE MAIN · PAPER 1 · CHEMISTRY

01 Some Basic Concepts of Chemistry

Formula & concept sheet

The language and arithmetic of chemistry: matter, units and significant figures, the laws of chemical combination, the mole, empirical and molecular formulae, balanced equations with limiting reagent and yield, and every way of stating the concentration of a solution. Each idea is explained first, then the formulas, shortcuts and traps.

INSIDE THIS SHEET

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READ ALONGSIDE NCERT XI · Some Basic Concepts of Chemistry

HOW TO USE THIS SHEET

1 **Skim**
Read the Key ideas, then

2 **Memorise**
Learn the formulas and

3 **Test**
Do the worked examples,

4 **Revise**
Use the last-minute

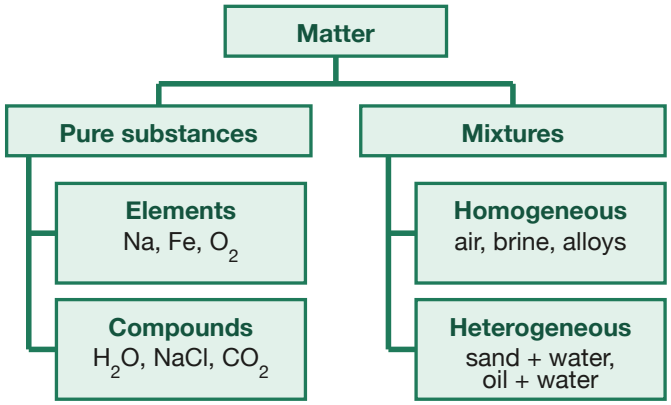
CHAPTER 01

Some Basic Concepts of Chemistry

NCERT XI · Some Basic Concepts of Chemistry

Covers: matter and its nature · Dalton's atomic theory · atoms, molecules, elements, compounds · laws of chemical combination · atomic and molecular masses · the mole and molar mass · percentage composition, empirical and molecular formulae · chemical equations and stoichiometry · limiting reagent and yield · molarity, molality, mole fraction, normality, ppm, volume strength.

1 Matter, units and measurement



Only a compound can be split into elements by chemical means; a mixture is separated physically

MATTER Anything that has mass and occupies space. A **physical change** (melting, dissolving, boiling) leaves the substance the same; a **chemical change** makes new substances, and only that is written as an equation.

ATOMS, MOLECULES, COMPOUNDS

- An **atom** is the smallest particle of an element that takes part in a reaction; a **molecule** is the smallest particle that can exist on its own (H_2, O_3, P_4, S_8 are molecules of elements, with atomicity 2, 3, 4, 8).
- A **compound** has two or more elements combined in a fixed ratio by mass; its properties differ from those of its elements. Ionic solids such as NaCl have **formula units**, not molecules.
- A **mixture** keeps the properties of its parts and is separated by physical methods.

SI BASE UNITS

QUANTITY	UNIT	SYMBOL
length	metre	m
mass	kilogram	kg
time	second	s
amount of substance	mole	mol
temperature	kelvin	K
electric current	ampere	A
luminous intensity	candela	cd

PREFIXES AND CONVERSIONS

PREFIX	FACTOR	PREFIX	FACTOR
tera T	10^{12}	centi c	10^{-2}
giga G	10^{9}	milli m	10^{-3}
mega M	10^{6}	micro μ	10^{-6}
kilo k	10^{3}	nano n	10^{-9}
deci d	10^{-1}	pico p	10^{-12}

$1\ \text{L} = 1\ \text{dm}^3 = 10^3\ \text{cm}^3$; $1\ \text{atm} = 101325\ \text{Pa} = 760\ \text{mm Hg} = 1.013\ \text{bar}$; $K = {}^\circ\text{C} + 273.15$; $1\ \text{cal} = 4.184\ \text{J}$; $1\ \text{Å} = 10^{-10}\ \text{m}$

DENSITY

$$d = \frac{m}{V}$$

g mL^{-1} (= g cm^{-3}) or kg m^{-3}; water is 1 g mL^{-1}; specific gravity is the same number without a unit

MASS AND WEIGHT

$$W = mg$$

mass is the amount of matter and never changes; weight is the pull of gravity and varies with place. Temperatures below 0 K do not exist

SIGNIFICANT FIGURES

1. Every non-zero digit counts; zeros between digits count; **leading** zeros never count (0.0035 has 2); trailing zeros count only after a decimal point (2.50 has 3, 250 has 2 or 3).
2. Exact numbers (a count of 2 mol, a definition like 100 cm $= 1$ m) have unlimited figures.
3. Multiply or divide: keep as many figures as the least precise factor. Add or subtract: keep as many **decimal places** as the least precise term.
4. Round only at the end of a calculation; round 5 upwards.

FACTOR-LABEL (DIMENSIONAL ANALYSIS)

1. Write what is given with its unit.
2. Multiply by ratios equal to 1, such as $\frac{1000\ \text{mL}}{1\ \text{L}}$ or $\frac{M\ \text{g}}{1\ \text{mol}}$, each chosen so the unwanted unit cancels.
3. Check the unit that is left is the one asked for. This catches $\times$ versus $\div$ mistakes before any arithmetic.

2 Atoms and the laws of chemical combination

THE LAWS

LAW	STATEMENT	EXAMPLE
Conservation of mass (Lavoisier)	total mass of reactants $=$ total mass of products in a closed system	2 g $\text{H}_2 + 16$ g O_2 gives 18 g H_2O
Definite proportions (Proust)	a pure compound has the same elements in the same mass ratio, whatever its source	H_2O: H : O $= 1 : 8$ by mass, always
Multiple proportions (Dalton)	if two elements form more than one compound, the masses of one that combine with a fixed mass of the other are in small whole-number ratio	CO and CO_2: 16 g and 32 g of O per 12 g of C, ratio $1 : 2$
Gaseous volumes (Gay-Lussac)	gases react in simple whole-number volume ratios, at the same T and p	2 vol $\text{H}_2 + 1$ vol $\text{O}_2 \rightarrow 2$ vol steam
Avogadro	equal volumes of gases at the same T and p hold equal numbers of molecules, so volume ratio $=$ mole ratio	explains Gay-Lussac; $V \propto n$

DALTON'S POSTULATES THAT STILL HOLD

- Matter is built of atoms; atoms of one element are alike in kind and differ from other elements.
- Compounds form when atoms of different elements combine in small whole-number ratios.
- Atoms are neither created nor destroyed in a chemical reaction; they only rearrange.

WHAT IS NOW KNOWN TO BE WRONG

- Atoms are indivisible: they hold electrons, protons and neutrons.
- All atoms of an element have the same mass: **isotopes** differ (Cl-35, Cl-37).
- Atoms of different elements always differ in mass: **isobars** (Ar-40, Ca-40) can match.

ATOMIC MASS UNIT

$$1\ \text{u} = \frac{1}{12}\,m(^{12}\text{C})$$

$= 1.66 \times 10^{-24}$ g $= \frac{1}{N_A}$ g; relative atomic mass has no unit

AVERAGE ATOMIC MASS

$$\bar{M} = \sum x_i M_i$$

x_i = fractional abundance of isotope i; Cl: $0.7577 \times 35 + 0.2423 \times 37 = 35.48$

AVOGADRO'S LAW

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

at fixed T and p; so $V \propto n$ for ideal gases

CHARGES TO KNOW FOR WRITING FORMULAE (CRISS-CROSS THE CHARGES)

CHARGE	CATIONS	ANIONS
$+1$ / -1	H^+, Na^+, K^+, Ag^+, NH_4^+, Cu^+	Cl^-, Br^-, I^-, OH^-, NO_3^-, HCO_3^-, MnO_4^-
$+2$ / -2	Mg^{2+}, Ca^{2+}, Ba^{2+}, Zn^{2+}, Fe^{2+}, Cu^{2+}	O^{2-}, S^{2-}, SO_4^{2-}, CO_3^{2-}, $Cr_2O_7^{2-}$
$+3$ / -3	Al^{3+}, Fe^{3+}, Cr^{3+}	PO_4^{3-}, N^{3-}

Criss-cross: Al^{3+} with SO_4^{2-} gives $Al_2(SO_4)_3$; Ca^{2+} with PO_4^{3-} gives $Ca_3(PO_4)_2$. Total positive charge = total negative charge.

KEY IDEA

A statement question usually pins one law to one example: "the ratio of masses of oxygen in CO and CO_2 is $1:2$" is multiple proportions; "water from any source is $1:8$" is definite proportions; "2 L of H_2 gives 2 L of steam" is Gay-Lussac. Dalton's postulate that atoms are conserved is still the basis of balancing; isotopes break only the equal-mass claim.

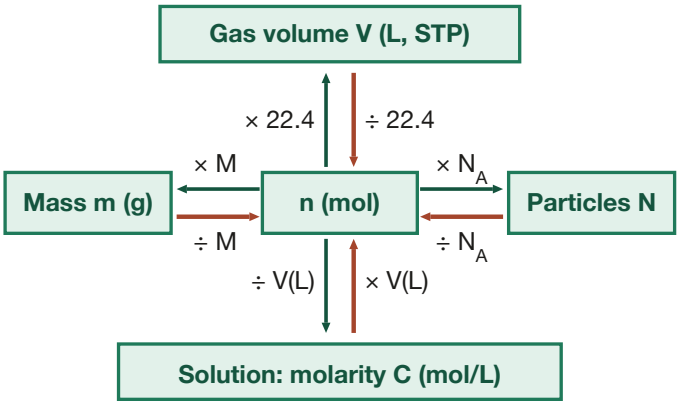
QUICK CHECK

Copper forms two oxides with 88.9% and 79.9% Cu. Do they obey multiple proportions?

Answer: Cu per g of O: $\frac{88.9}{11.1} = 8.01$ and $\frac{79.9}{20.1} = 3.98$; ratio $2:1$, so yes (Cu_2O and CuO).

3 The mole and molar mass

MOLE The amount of substance containing as many particles as there are atoms in 12 g of C-12: $N_A = 6.022 \times 10^{23}$ mol^{-1} (Avogadro's number). The molar mass in g mol^{-1} equals the relative atomic or molecular mass in u.



M = molar mass in g/mol; 22.4 L/mol at 1 atm and 273 K (22.7 L/mol at 1 bar)

ONE RELATION, FOUR FORMS

$$n = \frac{m}{M} = \frac{N}{N_A} = \frac{V_{gas}}{V_m} = C\,V_{soln}$$

m in g, M in g mol^{-1}, V_{gas} at STP, C in mol L^{-1} and V_{soln} in L; convert everything to moles first, then to what is asked

COUNTING ATOMS

- ✓ atoms $= n \times$ (atoms per formula) $\times N_A$: 18 g water is 1 mol, N_A molecules and $3N_A$ atoms
- ✓ compare samples in units of N_A to avoid large numbers

VAPOUR DENSITY

$$M = 2 \times \text{VD}$$

VD is relative to H_2 ($M = 2$); relative to air (29) use $M = 29\times$ density ratio

IDEAL GAS (AS A TOOL)

$$pV = nRT$$

$R = 0.0821$ L atm K^{-1} mol^{-1} $= 8.314$ J K^{-1} mol^{-1}; use when p and T are not STP

MEAN MOLAR MASS

$$\bar{M} = \sum x_i M_i$$

x_i = mole fraction (equal to volume fraction for gases at the same T, p)

MOLAR VOLUME OF AN IDEAL GAS: WHICH CONVENTION?

CONVENTION	CONDITIONS	V_m (L MOL^{-1})
old STP (most JEE papers)	273.15 K, 1 atm	22.4
IUPAC STP	273.15 K, 1 bar	22.7
SATP	298.15 K, 1 bar	24.8

Use the value the question states and keep it throughout one problem. The two STP values differ by about 1.3% (it changes the 11.2 volume-strength factor to 11.35).

COUNTING PARTICLES IN ONE MOLECULE

SPECIES	ATOMS	ELECTRONS
H_2O, NH_3, CH_4	3, 4, 5	10 each
CO_2	3	22
H_2SO_4	7	50
$C_6H_{12}O_6$ (glucose)	24	96
$C_{12}H_{22}O_{11}$ (sucrose)	45	182

electrons per molecule = sum of atomic numbers; per mole multiply by N_A

NUMBERS WORTH CARRYING

QUANTITY	VALUE
mass of one molecule of H_2O	$\frac{18}{N_A} = 2.99 \times 10^{-23}$ g
molecules in 1 mL of gas at STP	2.69×10^{19}
charge on 1 mol of electrons	96500 C (Faraday)
atoms in 1 g of H_2O	1.0×10^{23}
1 mol of O_2 at STP	32 g, 22.4 L, 6.022×10^{23} molecules

REMEMBER

One mole: 6.022×10^{23} particles, M grams, 22.4 L if it is a gas at STP. 18 g of water is one mole; 1 mol of any gas has the same volume, not the same mass.

QUICK CHECK

Which has more atoms: 4 g of He or 4 g of H_2?

Answer: He: 1 mol, N_A atoms; H_2: 2 mol, $4N_A$ atoms. H_2 has 4 times as many.

4 Composition and formulae

- The **empirical formula** is the simplest whole-number ratio of atoms; the **molecular formula** is the actual number of atoms and is a whole-number multiple of it: $(\mathrm{EF})_n$ with $n = M/M_{EF}$.
- Percentage composition fixes the empirical formula; the molar mass (from vapour density, osmotic data, mass spectrum) then fixes the molecular formula.

MASS PERCENTAGE

$$\%X = \frac{(\text{atoms of } X)\, M_X}{M_{compound}} \times 100$$

for a hydrate use the full molar mass including the water

MOLECULAR FORMULA

$$n = \frac{M_{molecular}}{M_{empirical}}$$

n must come out whole: if it does not, check the rounding

WATER OF CRYSTALLISATION

$$\%\,H_2O = \frac{18x}{M_{hydrate}} \times 100$$

$CuSO_4 \cdot 5H_2O$: $M = 249.5$, water $\frac{90}{249.5} = 36.1\%$

EMPIRICAL FORMULA FROM PERCENTAGES

1. Take 100 g, so each percentage becomes grams.
2. Divide each mass by the atomic mass to get moles.
3. Divide all by the smallest to get a ratio.
4. If a value is not whole, multiply: $1.5 \to \times 2$, 1.33 or $1.67 \to \times 3$, 1.25 or $1.75 \to \times 4$.
5. Divide M by the empirical mass to get n, then scale.

WORKED LAYOUT: C 40%, H 6.7%, O 53.3%, M = 180

ELEMENT	MASS (G)	÷ ATOMIC MASS	÷ SMALLEST
C	40	3.33	1
H	6.7	6.7	2
O	53.3	3.33	1

EF CH_2O (mass 30); $n = 180/30 = 6$, so $C_6H_{12}O_6$ (glucose).

CARBON FROM CO_2

$$\%\,\mathrm{C} = \frac{12}{44} \times \frac{m_{\mathrm{CO_2}}}{m_{sample}} \times 100$$

all the carbon of the sample ends up in the CO_2 collected in KOH

HYDROGEN FROM H_2O

$$\%\,\mathrm{H} = \frac{2}{18} \times \frac{m_{\mathrm{H_2O}}}{m_{sample}} \times 100$$

water is absorbed in $CaCl_2$ or $Mg(ClO_4)_2$ first

OXYGEN BY DIFFERENCE

$$\%\,\mathrm{O} = 100 - (\%\mathrm{C} + \%\mathrm{H} + \%\mathrm{N} + \dots)$$

oxygen is not estimated directly; nitrogen by Dumas or Kjeldahl (see **Purification and Characterisation**)

KEY IDEA

If a compound's mass percentages are given for an element and the molar mass is known, you can skip the empirical step: atoms per molecule $= \dfrac{\% \times M}{100 \times A_{element}}$.

For N in urea, $\dfrac{46.7 \times 60}{100 \times 14} = 2$.

TRAP

Mass percent of an element in a hydrate includes the water. An empirical ratio of 1.5 means multiply the whole ratio by 2, never write a half atom. Do not round 1.33 to 1.

5 Chemical equations and stoichiometry

- A balanced equation has the same number of each kind of atom, and the same charge, on both sides. Its coefficients are **mole ratios** (and, for gases at the same T and p, volume ratios), never mass ratios.
- Every stoichiometry problem is the same chain: **grams → moles → (coefficient ratio) → moles → grams, litres or particles**.

BALANCE BY INSPECTION

1. Balance the element that appears in the fewest species first (usually a metal), then non-metals.
2. Leave H and O for last; treat a polyatomic ion that stays intact (SO_4^{2-}) as one unit.
3. If a fraction appears, clear it by doubling. Recount every element at the end.

LIMITING REAGENT AND YIELD

1. Convert every reactant to moles.
2. Divide each by its coefficient: the **smallest** n/ν is the limiting reagent; the excess left $= n - \nu \times (n/\nu)_{min}$.
3. Product moles $= (n/\nu)_{min} \times \nu_{product}$; this is the **theoretical** yield.
4. $\%$ yield $= \dfrac{\text{actual}}{\text{theoretical}} \times 100$. For an impure sample use $m \times \dfrac{\%\text{purity}}{100}$ first.

reactant	n	coeff	n/coeff
H_2	1	2	**0.5**
O_2	1	1	1

smallest = limiting

divide moles by the coefficient; the smallest ratio runs out first

PERCENTAGE YIELD AND PURITY

$$\%\,\text{yield} = \frac{\text{actual}}{\text{theoretical}} \times 100, \qquad n_{pure} = \frac{m\,(\%\text{purity})}{100\,M}$$

in a multi-step synthesis the overall yield is the **product** of the step yields: 80% then 50% gives 40%

For a limiting-reagent problem with a gas, convert volumes to moles first; if all gases are at the same T and p you may use volumes directly in place of moles.

REACTIONS THE CHAPTER USES (ALL BALANCED)

REACTION	MOLE FACT WORTH KNOWING
$CaCO_3 \xrightarrow{\Delta} CaO + CO_2$	100 g gives 56 g residue and 22.4 L CO_2 (STP)
$2\,KClO_3 \xrightarrow[MnO_2]{\Delta} 2\,KCl + 3\,O_2$	245 g gives 96 g O_2 (67.2 L)
$2\,H_2O_2 \longrightarrow 2\,H_2O + O_2$	68 g H_2O_2 gives 22.4 L O_2 (volume strength)
$Zn + 2\,HCl \longrightarrow ZnCl_2 + H_2$	65 g Zn gives 22.4 L H_2
$4\,Al + 3\,O_2 \longrightarrow 2\,Al_2O_3$	54 g Al needs 48 g O_2
$Fe_2O_3 + 3\,CO \longrightarrow 2\,Fe + 3\,CO_2$	160 g Fe_2O_3 gives 112 g Fe
$CH_4 + 2\,O_2 \longrightarrow CO_2 + 2\,H_2O$	1 vol CH_4 needs 2 vol O_2
$C_3H_8 + 5\,O_2 \longrightarrow 3\,CO_2 + 4\,H_2O$	1 vol gives 3 vol CO_2
$CuSO_4 \cdot 5\,H_2O \xrightarrow{\Delta} CuSO_4 + 5\,H_2O$	blue to white; 36.1% mass lost
$NaOH + HCl \longrightarrow NaCl + H_2O$	1 : 1; $Ca(OH)_2 + 2\,HCl \longrightarrow CaCl_2 + 2\,H_2O$ is 1 : 2

EUDIOMETRY: BURNING A GASEOUS HYDROCARBON

$$C_xH_y + \left(x + \frac{y}{4}\right) O_2 \rightarrow x\,CO_2 + \frac{y}{2}\,H_2O$$

balanced by atoms: x C, y H, then O $= 2x + \frac{y}{2}$

$$\text{contraction on cooling} = \left(1 + \frac{y}{4}\right) V$$

gas before: $1 + (x + \frac{y}{4})$; gas after (water now liquid): x

$$\text{decrease on KOH} = xV$$

KOH absorbs the CO_2; what remains is excess O_2

KEY IDEA

Mixture problems (heating carbonates, gas mixtures) are two equations in two unknowns: one for the total mass and one for the product (moles of gas, mass of residue). Write the unknowns as x and y in **grams** and divide by each molar mass.

TRAP

Compare moles, never grams, to find the limiting reagent; compare n/ν, not n. Heating a carbonate: mass lost is only CO_2, and the residue is the oxide, not the metal.

6 Concentration of solutions

A solution is a homogeneous mixture; the substance in smaller amount is the **solute** and the other is the **solvent**. Concentration terms differ in what they divide by: **volume** of solution (changes with temperature) or **mass** (does not).

MOLARITY

$$M = \frac{n_{solute}}{V_{soln}(\text{L})}$$

mol L^{-1}; depends on temperature

MOLALITY

$$m = \frac{n_{solute}}{w_{solvent}(\text{kg})}$$

mol kg^{-1}; independent of temperature

MOLE FRACTION

$$x_A = \frac{n_A}{\sum n}, \quad \sum x = 1$$

no unit; independent of temperature

MASS PERCENT

$$\%\, w/w = \frac{\text{g solute}}{\text{g solution}} \times 100$$

$\%\, w/v$ divides by mL of solution; $\%\, v/v$ by volumes

PPM

$$\frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6$$

for dilute aqueous solutions ppm $=$ mg L^{-1}

NORMALITY

$$N = M \times n_f$$

n_f is equivalents per mole;

$$\text{equivalent mass} = \frac{M}{n_f}$$

CONVERTING BETWEEN TERMS

FROM	TO	FORMULA
% w/w (w) and d	molarity	$M = \dfrac{10\,w\,d}{M_{solute}}$ (d in g mL^{-1})
molarity and d	molality	$m = \dfrac{1000\,M}{1000\,d - M\,M_{solute}}$
mole fraction x_{solute}	molality	$m = \dfrac{1000\,x_{solute}}{x_{solvent}\,M_{solvent}}$
molality and d	molarity	$M = \dfrac{1000\,m\,d}{1000 + m\,M_{solute}}$

Without a formula: assume 1 L of solution (if M is given) or 100 g of solution (if % is given) or 1000 g of solvent (if m is given), find solute, solvent and solution masses with d, then compute what is asked.

DEPEND ON TEMPERATURE (VOLUME IN THE DENOMINATOR)

- Molarity M and normality N: the volume of a solution expands on heating, so the value falls.
- % w/v and % v/v change for the same reason.

INDEPENDENT OF TEMPERATURE (MASSES OR MOLES ONLY)

- Molality m, mole fraction x, % w/w and ppm: they compare masses or numbers of moles, which do not change.
- So colligative-property formulas use m and x, not M (see **Solutions**).

DILUTION

$$M_1V_1 = M_2V_2$$

the moles of solute do not change when solvent is added

MIXING THE SAME SOLUTE

$$M_{mix} = \frac{M_1V_1 + M_2V_2}{V_1 + V_2}$$

volumes taken additive; different solutes that do not react keep their own moles

VOLUME STRENGTH OF HYDROGEN PEROXIDE

$$2\,H_2O_2 \longrightarrow 2\,H_2O + O_2$$

68 g of H_2O_2 gives 22.4 L of O_2 at STP

1 L of 1 M (34 g) $\rightarrow 11.2$ L O_2

half of 22.4 L

$$\text{vol. strength} = 11.2\,M = 5.6\,N$$

$n_f = 2$ for H_2O_2; $11.35\,M$ if 22.7 L is used

N-FACTOR (EQUIVALENTS PER MOLE)

SPECIES	n_f
acid (basicity), e.g. H_2SO_4, H_3PO_4 (as used)	H^+ replaced: 2, 3
base (acidity), e.g. $Ca(OH)_2$	2
salt	total positive charge
$KMnO_4$ acidic / neutral / basic	5 / 3 / 1
$K_2Cr_2O_7$ (acidic), $Fe^{2+} \rightarrow Fe^{3+}$, $C_2O_4^{2-} \rightarrow 2CO_2$	6, 1, 2

Law of equivalence: $N_1V_1 = N_2V_2$ at the end point;
$\text{gram equivalents} = \dfrac{m}{\text{equivalent mass}}$.

KEY IDEA

Molarity of a dilute aqueous solution is a little **smaller** than molality, since 1 L of solution weighs about 1000 g but contains less than 1000 g of water; for concentrated solutions the order can reverse.

TRAP

Molarity divides by the volume of **solution**; molality divides by the mass of **solvent** in kg: subtract the solute from the solution mass. Dilution does not change the moles of solute but changes every concentration.

7 Worked examples

COUNTING ATOMS

Number of oxygen atoms in 4.4 g of CO_2 ($M = 44$)?

$n = \dfrac{4.4}{44} = 0.1$ mol; O atoms $= 0.1 \times 2 \times N_A = 0.2N_A = 1.2 \times 10^{23}$

LIMITING REAGENT

2 g H_2 and 32 g O_2 react to form water. Mass of water and of the excess reagent?

$n_{H_2} = 1,\ n_{O_2} = 1$; $\dfrac{n}{\nu}$: 0.5 and 1, so H_2 limits. Water $= 1$ mol $= 18$ g; O_2 left $= 1 - 0.5 = 0.5$ mol $= 16$ g

EMPIRICAL TO MOLECULAR FORMULA

An organic compound has 40% C, 6.7% H, 53.3% O and molar mass 180 g mol^{-1}. Molecular formula?

ratio C : H : O $= 3.33 : 6.7 : 3.33 = 1 : 2 : 1$; EF CH_2O (30); $n = \dfrac{180}{30} = 6$; $C_6H_{12}O_6$

MOLARITY FROM PERCENT

Orthophosphoric acid is 70% H_3PO_4 by mass with density 1.54 g mL^{-1} ($M = 98$). Molarity?

$M = \dfrac{10 \times 70 \times 1.54}{98} = 11.0$ M

MOLALITY FROM MOLARITY

A 0.2 M $CuSO_4$ solution has density 1.25 g mL^{-1} ($M = 159.5$). Molality?

$m = \dfrac{1000 \times 0.2}{1000 \times 1.25 - 0.2 \times 159.5} = \dfrac{200}{1218.1} = 0.164$ mol kg^{-1}

MIXING

20 mL of 2 M NaOH is added to 400 mL of 0.5 M NaOH. Final molarity?

moles $= 0.040 + 0.200 = 0.240$; volume 0.420 L; $M = \dfrac{0.240}{0.420} = 0.571$ M

VOLUME STRENGTH

30% w/v H_2O_2 ($M = 34$): volume strength at the 22.4 L convention?

30 g per 100 mL is 300 g L^{-1}, $M = \dfrac{300}{34} = 8.82$; $11.2 \times 8.82 = 98.8 \approx 99$ volume

CONCENTRATED ACID

98% H_2SO_4 by mass, density 1.84 g mL^{-1} ($M = 98$): molarity and molality?

$M = \dfrac{10 \times 98 \times 1.84}{98} = 18.4$ M; 100 g has 1 mol acid in 2 g water, so $m = \dfrac{1}{0.002} = 500$ mol kg^{-1}

HEATING A MIXTURE

18.4 g of $CaCO_3$ and $MgCO_3$ lose 8.8 g on strong heating. Mass of each? ($M = 100$ and 84)

$n_{CO_2} = \dfrac{8.8}{44} = 0.2$; $\dfrac{x}{100} + \dfrac{18.4 - x}{84} = 0.2 \Rightarrow x = 10$ g $CaCO_3$, 8.4 g $MgCO_3$

GAS FROM IMPURE SAMPLE

Mass of 90% pure $KClO_3$ ($M = 122.5$) needed for 11.2 L O_2 at STP?

$n_{O_2} = 0.5$; $n_{KClO_3} = 0.5 \times \dfrac{2}{3} = 0.333$; pure 40.8 g; sample $\dfrac{40.8}{0.90} = 45.4$ g

EUDIOMETRY

10 mL of a gaseous hydrocarbon burnt in excess O_2 contracts by 30 mL on cooling and by a further 30 mL on KOH. Formula?

$xV = 30 \Rightarrow x = 3$; $\left(1 + \dfrac{y}{4}\right) 10 = 30 \Rightarrow y = 8$; C_3H_8

NORMALITY

25 mL of Fe^{2+} solution needs 20 mL of 0.02 M $KMnO_4$ (acidic). Molarity of Fe^{2+}?

equivalents of $KMnO_4 = 0.02 \times 5 \times 20 = 2$ meq $= N_{Fe} \times 25$; $N = 0.08$, and as $n_f = 1$, $M = 0.08$ M

WHERE NUMERICALS GO WRONG

Using grams where moles are needed; forgetting to subtract solute mass when finding solvent mass; mixing 22.4 and 22.7 in one problem; a missing n-factor or coefficient; rounding 1.33 to 1 in an empirical formula.

AVERAGE ATOMIC MASS

Boron is 19% ^{10}B and 81% ^{11}B. Atomic mass?

$\bar{M} = 0.19 \times 10 + 0.81 \times 11 = 10.81$ u

PERCENTAGE YIELD

6 g of carbon burns to give 17.6 g of CO_2. Percentage yield?

$n_C = 0.5$ gives $0.5 \times 44 = 22$ g (theoretical); $\frac{17.6}{22} \times 100 = 80\%$

MOLE FRACTION TO MOLALITY

An aqueous solution has solute mole fraction 0.1. Molality?

$$m = \frac{1000\, x_{solute}}{x_{water} \times 18} = \frac{1000 \times 0.1}{0.9 \times 18} = 6.17 \text{ mol kg}^{-1}$$

HYDRATE

Residue and water lost on heating 2.5 g of $CuSO_4 \cdot 5\,H_2O$ ($M = 249.5$; $CuSO_4 = 159.5$)?

residue $= 2.5 \times \frac{159.5}{249.5} = 1.60$ g; water lost $= 2.5 - 1.60 = 0.90$ g

LIMITING GAS

10 g of N_2 and 3 g of H_2 react to make NH_3. Mass of NH_3 ?

$N_2 + 3\,H_2 \longrightarrow 2\,NH_3$; $\frac{n}{\nu}$: $\frac{0.357}{1}$ and $\frac{1.5}{3} = 0.5$, so N_2 limits; $NH_3 = 0.714$ mol $= 12.1$ g

DILUTION

To what volume must 10 mL of 12 M HCl be diluted to get 0.6 M?

$V_2 = \frac{12 \times 10}{0.6} = 200$ mL (add 190 mL of water)

8 Last-minute checklist

BEFORE YOU ANSWER

Moles first. Balanced? Which reagent limits? Volume of solution or mass of solvent? Which molar volume (22.4 or 22.7) does the question use?

- $n = m/M = N/N_A = V/22.4$ (STP); $N_A = 6.022 \times 10^{23}$; 1 u $= 1.66 \times 10^{-24}$ g; $M = 2\times$ VD.
- Laws: Lavoisier (mass), Proust (fixed ratio), Dalton (multiple), Gay-Lussac (volumes), Avogadro ($V \propto n$).
- Empirical: $\%$ $\rightarrow$ grams $\rightarrow \div A \rightarrow \div$ smallest $\rightarrow$ whole numbers; $n = M/M_{EF}$.
- Limiting reagent: smallest n/ν. Yield $= \frac{\text{actual}}{\text{theoretical}} \times 100$; use purity before moles.
- $M = \frac{10wd}{M_{solute}}$; $m = \frac{1000M}{1000d - MM_{solute}}$; $M_1V_1 = M_2V_2$; $N = Mn_f$.
- H_2O_2: volume strength $= 11.2M = 5.6N$; eudiometry contraction $= (1 + y/4)V$, KOH removes xV.

Now test it. Practise this chapter with past-paper and bank questions on Prep Vault. **Practice this chapter →**