STOICHIOMETRY

Detailed Chemistry Notes for JEE Advanced

Mole Concept • Chemical Equations • Limiting Reagent • Concentration • Titration

How to use these notes
Stoichiometry is the quantitative language of chemistry. Convert every given quantity into a suitable chemical amount (usually moles), use a balanced equation, identify the limiting reactant if needed, and convert the result into the requested unit.

Contents

- Atomic mass, molecular mass, formula mass, and the mole
- Molar mass, Avogadro constant, particles, and gas volume
- Chemical equations and stoichiometric coefficients
- Limiting reagent, excess reagent, and extent of reaction
- Percentage purity, yield, and composition analysis
- Empirical and molecular formula determination
- Concentration terms: molarity, molality, mole fraction, ppm, and normality
- Dilution, mixing, titration, and equivalent concept
- Gaseous stoichiometry, combustion analysis, and redox bookkeeping
- JEE Advanced solved illustrations and practice questions

Use consistent units and retain significant figures until the final step. Numerical values in examples are illustrative.

1. Atomic Mass, Molecular Mass, and the Mole

The unified atomic mass unit (u) is defined as one-twelfth of the mass of one carbon-12 atom. Relative atomic mass is a ratio and is dimensionless; molar mass is the mass of one mole of entities and is expressed in g mol^{-1}.

The mole

One mole contains exactly $6.02214076 \times 10^{23}$ specified elementary entities. The entities must be stated: atoms, molecules, ions, electrons, formula units, or other particles. This number is the Avogadro constant, N_a.

Quantity	Relation	Use
Moles (n)	n = mass / molar mass = m/M	Convert mass to amount.
Number of entities (N)	$N = nN_a$	Count atoms, molecules, ions, etc.
Mass	m = nM	Convert amount to mass.
Gas volume (ideal)	PV = nRT	Relate amount to P, V, T.

Molecular and formula mass

For a molecule, molecular mass is the sum of the atomic masses in its formula. For ionic solids, use formula mass rather than molecular mass because the solid is an extended ionic lattice. For example, formula mass of NaCl = $A_r(Na) + A_r(Cl)$.

Entity-counting trap
One mole of H_2O contains 1 mol of H_2O molecules, 2 mol of H atoms, and 1 mol of O atoms. It contains 3 mol of atoms in total. Always identify which entity the question asks you to count.

Average atomic mass

For an element with isotopes, average atomic mass is the abundance-weighted mean: $A_r = \Sigma(f_i \times m_i)$, where f_i is the fractional abundance and $\Sigma f_i = 1$. If percentages are supplied, divide each percentage by 100 before using the expression.

2. Molar Mass, Particles, and Gas Volume

Molar mass is the mass of one mole of a specified entity. Use the correct chemical formula and account for atoms per formula unit when converting between molecules and atoms.

Useful conversions

- Molecules in a sample = (sample mass / molar mass) × N_a.
- Total atoms in a molecular sample = molecules × number of atoms per molecule.
- For ionic compounds, formula units × stoichiometric subscript gives the count of each ion type.
- At a stated temperature and pressure, use the ideal gas equation unless non-ideality is explicitly relevant.

Ideal gas relation

PV = nRT. Use consistent units: for P in atm and V in L, R = 0.082057 L atm mol^{-1} K^{-1}; for SI units, R = 8.314 J mol^{-1} K^{-1}. Temperature must be in kelvin: T(K) = t(°C) + 273.15.

Gas density and molar mass

For an ideal gas, density d = PM/RT, so M = dRT/P. For two gases at the same T and P, density ratio equals molar-mass ratio. Relative vapour density with respect to H_2 is M/2 for an ideal gas under the same conditions.

Do not blindly use 22.4 L mol^{-1}
The value near 22.4 L mol^{-1} applies approximately to an ideal gas at 273.15 K and 1 atm. At other conditions, calculate molar volume from PV = nRT; standard-state conventions can differ.

3. Chemical Equations and Stoichiometric Coefficients

A balanced chemical equation obeys conservation of atoms and charge. Its coefficients give mole ratios, not mass ratios or direct particle masses.

General reaction

For aA + bB → cC + dD, the stoichiometric coefficients are $\nu A = -a$, $\nu B = -b$, $\nu C = +c$, and $\nu D = +d$ when reactants are assigned negative signs and products positive signs.

Mole-ratio method

If A is consumed, the reaction ratio gives n(B consumed) = n(A consumed) × (b/a), and n(C formed) = n(A consumed) × (c/a). Convert the given quantity to moles before applying the coefficient ratio.

Balancing methods

- Inspection: useful for simple equations; adjust coefficients while preserving atom counts.
- Algebraic method: assign variables to coefficients and solve atom-balance equations; choose the smallest whole-number solution.
- Redox half-reaction method: balance atoms other than O and H, then O using H_2O, H using H^+ (acidic medium), and charge using electrons. For basic medium, neutralise H^+ using OH^- and simplify water.
- In ionic equations, conserve both atoms and net charge; spectator ions may be cancelled only after writing complete ionic forms.

Coefficient meaning
For $2H_2 + O_2 \rightarrow 2H_2O$, 2 mol H_2 reacts with 1 mol O_2 to form 2 mol H_2O. The same ratio applies to molecules, but not to grams.

4. Limiting Reagent and Excess Reagent

The limiting reagent is consumed first and determines the maximum possible amount of product. Other reactants are present in excess.

Two reliable methods

Method A – Compare n/coefficient: For each reactant i, calculate n_i/a_i. The smallest value identifies the limiting reactant for a single reaction proceeding as written.

Method B – Product test: Calculate the amount of a chosen product that each reactant could produce independently. The smaller product amount is the theoretical maximum.

Step	Action
1	Write and balance the reaction.
2	Convert each reactant quantity to moles.
3	Divide each reactant's moles by its stoichiometric coefficient.
4	The smallest quotient identifies the limiting reagent.
5	Use the limiting reagent to calculate theoretical product.
6	Calculate leftover excess reagent by subtracting the amount consumed.

Extent of reaction

For a reaction with initial amount n_i, the amount at extent ξ is $n_i(\text{final}) = n_i(\text{initial}) + \nu_i\xi$. For a reactant, ν_i is negative. The maximum possible extent is limited by the first reactant to reach zero: $\xi_{max} = \min[n_i(\text{initial})/|\nu_i|]$ over reactants.

Solved illustration

Example: 4 mol H_2 reacts with 1 mol O_2 according to $2H_2 + O_2 \rightarrow 2H_2O$. Quotients: $H_2 = 4/2 = 2$; $O_2 = 1/1 = 1$. O_2 is limiting. It consumes 2 mol H_2 and forms 2 mol H_2O. Leftover H_2 = 2 mol.

JEE trap
The limiting reagent is not necessarily the reactant with the smallest mass or smallest number of moles. Compare moles divided by their balanced-equation coefficients.

5. Yield, Purity, and Composition

Theoretical yield is the maximum product calculated from the limiting reagent. Actual yield is the amount obtained experimentally.

Term	Expression
Percentage yield	(Actual yield / Theoretical yield) × 100
Percentage purity	(Mass of pure substance / Mass of impure sample) × 100
Mass of pure reactant	Sample mass × purity fraction
Percentage composition	(Mass of element in one mole of compound / Molar mass of compound) × 100

Purity in stoichiometric calculations

If a sample is p% pure, only p/100 of its stated mass is the desired reactant. Use this pure mass to calculate moles. The remaining mass is impurity and should not be included unless its identity and reaction are specified.

Solved illustration

A 10.0 g limestone sample is 80% $CaCO_3$ by mass. Pure $CaCO_3$ = 8.0 g. $M(CaCO_3) \approx 100$ g mol^{-1}, so n = 0.080 mol. On heating, $CaCO_3 \rightarrow CaO + CO_2$; theoretical CO_2 = 0.080 mol, approximately 3.52 g.

Yield vs purity
Purity describes the composition of the starting sample. Yield compares actual product with the theoretical product. Do not interchange the two percentages.

6. Empirical and Molecular Formula

The empirical formula gives the simplest whole-number ratio of atoms. The molecular formula gives the actual number of each atom in a molecule.

Empirical formula from percentage composition

- Assume 100 g of compound if percentages are given; each percentage then represents grams.
- Convert each element's mass to moles by dividing by its atomic mass.
- Divide all mole values by the smallest value.
- If ratios are fractional, multiply by the smallest integer that makes all subscripts whole numbers (commonly ×2, ×3, ×4).
- Check that the resulting formula is chemically plausible and that rounding has not distorted the ratio.

Molecular formula = $(\text{Empirical formula})_n$, where n = molar mass / empirical-formula mass. The value of n should be a positive integer within experimental uncertainty.

Combustion analysis

For a compound containing C and H, complete combustion gives CO_2 and H_2O. Moles of C = moles of CO_2; moles of H atoms = 2 × moles of H_2O. If the compound also contains O, oxygen mass can be found by difference after determining C and H, assuming no other elements are present.

For nitrogen or halogens, use the specific analytical method and product stoichiometry stated in the problem. Do not assume every element can be inferred from CO_2 and H_2O alone.

Solved illustration

A compound contains 40.0% C, 6.7% H, and 53.3% O. In 100 g: C = 40/12 ≈ 3.33 mol; H = 6.7/1 ≈ 6.7 mol; O = 53.3/16 ≈ 3.33 mol. Divide by 3.33: C:H:O ≈ 1:2:1. Empirical formula = CH_2O. If molar mass is 180 g mol^{-1}, empirical mass = 30, n = 6, molecular formula = $C_6H_{12}O_6$.

7. Concentration Terms

Term	Definition / formula	Unit / note
Molarity (M)	M = moles of solute / volume of solution in L	mol L^{-1}; temperature-dependent through volume.
Molality (m)	m = moles of solute / mass of solvent in kg	mol kg^{-1}; independent of temperature if mass unchanged.
Mole fraction (x_i)	$x_i = n_i / \Sigma n_j$	Dimensionless; sum of all mole fractions = 1.
Mass %	(mass solute / mass solution) × 100	Mass-based.
Volume %	(volume component / volume solution) × 100	Used for suitable liquid mixtures.
Mass/volume %	grams solute per 100 mL solution	Common in solution preparation.
ppm	(mass solute / mass solution) × 10^6	For dilute aqueous solutions, often approximated as mg L^{-1} only when density is about 1 kg L^{-1}.

Molarity and dilution

For dilution with no reaction or loss of solute: $M_1V_1 = M_2V_2$. Volumes must be in the same units. This relation is a solute mole balance, not a universal formula for mixing reacting solutions.

Mixing solutions of the same solute

If volumes are additive, final molarity = $(M_1V_1 + M_2V_2)/(V_1 + V_2)$. For different solutes, calculate each solute's final concentration separately. If volumes are not additive, use the actual final volume.

Molarity–molality relation

For a solution with density d in g mL^{-1}, solute molar mass M_r in g mol^{-1}, and molality m in mol kg^{-1}: Molarity = $(1000\ d\ m)/(1000 + mM_r)$. The denominator is the mass in grams of 1 kg solvent plus the solute mass corresponding to m moles.

Concentration trap
Molarity uses volume of the whole solution, not volume of solvent. Molality uses mass of solvent, not mass of solution.

8. Titration and Equivalent Concept

In a titration, the equivalence point is reached when reactants have combined in the exact stoichiometric ratio of the balanced reaction. The end point is the observed indicator change and may differ slightly from the equivalence point.

General stoichiometric relation

For $aA + bB \rightarrow$ products, at equivalence: $n(A)/a = n(B)/b$. If both are in solution, $M_A V_A/a = M_B V_B/b$, provided the balanced equation and reacting species are correctly identified.

Acid–base examples

Reaction	Mole relation at equivalence
$HCl + NaOH \rightarrow NaCl + H_2O$	$n(HCl) = n(NaOH)$
$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$	$n(NaOH) = 2n(H_2SO_4)$
$H_3PO_4 + 3NaOH \rightarrow Na_3PO_4 + 3H_2O$	For complete neutralisation: $n(NaOH) = 3n(H_3PO_4)$

Normality and n-factor

Equivalent amount = moles × n-factor for the specified reaction. Normality N = equivalents per litre = M × n-factor. The n-factor is reaction-dependent: for acids it is the number of H^+ actually furnished per formula unit in the stated reaction; for bases it is the number of H^+ consumed; for redox agents it is the number of electrons transferred per formula unit.

Do not memorise one fixed n-factor
For polyprotic acids and redox agents, n-factor depends on the reaction endpoint and medium. For example, H_3PO_4 can have n-factor 1, 2, or 3 depending on how many protons are neutralised.

Redox titration bookkeeping

Balance the redox equation in the stated medium, identify electrons lost and gained per formula unit, and equate total electron transfer. Never assign n-factor from the formula alone without specifying the reaction.

9. Gaseous Stoichiometry and Combustion

For ideal gases at the same temperature and pressure, volume ratios equal mole ratios. This is why balanced gaseous equations can be interpreted as volume ratios under the same T and P.

Gas reaction example

$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$. At the same T and P, 1 volume N_2 reacts with 3 volumes H_2 to form 2 volumes NH_3, if all gases are measured under the same conditions and the product remains gaseous.

Combustion calculations

For a hydrocarbon C_xH_y, complete combustion is: $C_xH_y + (x + y/4)O_2 \rightarrow xCO_2 + (y/2)H_2O$. Thus, moles of CO_2 reveal carbon; moles of H_2O reveal hydrogen. Oxygen required follows from the balanced equation.

Mixture problems

- For a mixture of non-reacting components, total moles are the sum of component moles.
- For an ideal-gas mixture, partial pressure $p_i = x_iP_{total}$ (Dalton's law).
- If a gas is collected over water, $P_{dry\ gas} = P_{total} - P_{water\ vapour}$ at the stated temperature.
- At the same T and P, gas volume is proportional to moles; at different conditions, use $PV = nRT$.

10. Advanced Problem-Solving Framework

A reliable workflow for multi-step JEE Advanced problems

- Translate every datum into a labelled quantity with units. Distinguish mass of sample, mass of pure substance, volume of solution, and gas volume.
- Write the chemical equation for the actual reaction and balance it, including ionic charge if relevant.
- Convert all reactants to moles. For solutions, use M×V with V in litres; for gases, use PV/RT.
- Identify the limiting reagent or use the extent-of-reaction variable ξ.
- Apply the mole ratio to calculate product, leftover reactant, or titration equivalence.
- Convert the result to the requested unit and check dimensional consistency.
- Check physical plausibility: yield cannot exceed 100% for a correctly defined theoretical yield; leftover limiting reagent should be zero; atom and charge balances must hold.

Solved JEE-style illustrations

Example 1 – Limiting reagent: 5.0 g H_2 reacts with 16.0 g O_2. Reaction: $2H_2 + O_2 \rightarrow 2H_2O$. $n(H_2)=2.5$ mol; $n(O_2)=0.5$ mol. Quotients are 1.25 and 0.5, so O_2 is limiting. Water formed = $2\times0.5 = 1.0$ mol = 18 g. H_2 consumed = 1.0 mol, so 1.5 mol H_2 remains.

Example 2 – Titration: 25.0 mL H_2SO_4 is neutralised by 20.0 mL of 0.50 M NaOH. $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$. Moles NaOH = $0.50\times0.0200 = 0.0100$ mol. Moles $H_2SO_4 = 0.0100/2 = 0.00500$ mol. Acid molarity = $0.00500/0.0250 = 0.200$ M.

Example 3 – Gas volume: At 300 K and 1.00 atm, volume of 0.20 mol ideal gas is $V=nRT/P = 0.20\times0.082057\times300 = 4.92$ L.

Example 4 – Yield: A reaction has a theoretical product mass of 12.0 g and gives 9.0 g experimentally. Percentage yield = $(9.0/12.0)\times100 = 75\%$.

11. JEE Advanced Practice Set

Try each problem before reading its answer. Assume ideal behaviour where required and use standard atomic masses unless a value is supplied.

Q1. How many oxygen atoms are present in 0.25 mol of $Al_2(SO_4)_3$ formula units?

Answer: Each formula unit contains 12 O atoms. Oxygen atoms = $0.25 \times 12 \times N_a = 3N_a \approx 1.807 \times 10^{24}$ atoms.

Q2. 10 g $CaCO_3$ is heated completely: $CaCO_3 \rightarrow CaO + CO_2$. Find CO_2 mass.

Answer: $M(CaCO_3) \approx 100$ g mol^{-1}; n=0.10 mol. Ratio is 1:1, so $CO_2 = 0.10 \text{ mol} \times 44 = 4.4$ g.

Q3. 3 mol N_2 reacts with 6 mol H_2: $N_2 + 3H_2 \rightarrow 2NH_3$. Identify limiting reagent and NH_3 formed.

Answer: H_2 quotient=6/3=2; N_2 quotient=3. H_2 is limiting. $NH_3 = 2 \times (6/3) = 4$ mol.

Q4. A 20 g impure sample is 75% Na_2CO_3. Calculate moles of pure Na_2CO_3. (M=106 g mol^{-1})

Answer: Pure mass=15 g. Moles=15/106≈0.1415 mol.

Q5. 100 mL of 2.0 M solution is diluted to 500 mL. Find final molarity.

Answer: $M_2 = M_1V_1/V_2 = 2.0 \times 100/500 = 0.40$ M.

Q6. A compound contains 52.2% C, 13.0% H, and 34.8% O. Find empirical formula.

Answer: For 100 g: C=4.35 mol, H=13.0 mol, O=2.175 mol. Divide by 2.175: 2:6:1. Empirical formula C_2H_6O.

Q7. 25 mL 0.10 M H_3PO_4 is completely neutralised by NaOH. Find NaOH moles at equivalence.

Answer: $H_3PO_4 + 3NaOH \rightarrow Na_3PO_4 + 3H_2O$. Acid moles=$0.025 \times 0.10 = 0.0025$; NaOH=$3 \times 0.0025 = 0.0075$ mol.

Q8. A gas has density 1.25 g L^{-1} at 1 atm and 300 K. Estimate molar mass.

Answer: $M = dRT/P = 1.25 \times 0.082057 \times 300/1 \approx 30.8$ g mol^{-1}.

Q9. 2 mol A and 3 mol B react by $A + 2B \rightarrow AB_2$. Find limiting reagent and leftover excess.

Answer: A quotient=2; B quotient=3/2=1.5. B is limiting; it consumes 1.5 mol A, leaving 0.5 mol A.

Q10. A reaction's theoretical yield is 8.0 g and actual yield is 6.4 g. Find percentage yield.

Answer: Yield=$(6.4/8.0) \times 100 = 80\%$.

12. Formula Sheet and Final Checklist

Concept	Formula / relation
Moles	$n = m/M$
Particles	$N = nN_a$
Ideal gas	$PV = nRT$
Density of ideal gas	$d = PM/RT$
Limiting reagent	Smallest n_i / coefficient$_i$
Extent of reaction	$n_i(\text{final}) = n_i(\text{initial}) + \nu_i\xi$
Percentage yield	Actual / theoretical × 100
Percentage purity	Pure mass / sample mass × 100
Molarity	M = nsolute / Vsolution(L)
Molality	m = nsolute / mass solvent(kg)
Mole fraction	$x_i = n_i / \Sigma n_j$
Dilution	$M_1V_1 = M_2V_2$
Titration	nA/a = nB/b at equivalence
Normality	N = M × n-factor

Final checklist before submitting an answer

- Is the chemical equation balanced for atoms and charge?
- Did I convert °C to K and mL to L where required?
- Did I use the mass of pure reactant rather than total impure sample?
- Did I identify the limiting reagent using coefficients?
- Did I distinguish solvent mass from solution volume?
- Did I use the correct n-factor for the specified reaction?
- Does the answer have the correct unit and a reasonable magnitude?

Core principle
Stoichiometry is a conservation problem. Atoms, charge, and the amount of each species must be accounted for through a balanced reaction. Most difficult problems become manageable when the data are converted to moles and organised systematically.