CHEMISTRY NERD

MOLE CONCEPT

Detailed JEE Advanced-Level Notes

What you will master
Mole and particle counting • Atomic and molecular mass • Stoichiometric calculations • Limiting reagent • Mixtures and purity • Gas stoichiometry • Concentration terms • Equivalent concept • Advanced multi-step problems

Study approach: Build the meaning of each quantity first, write a balanced chemical equation before using mole ratios, and keep units visible in every calculation.

Prepared for JEE Main & JEE Advanced aspirants

1. The Mole: Meaning and Foundation

Chemistry deals with an enormous number of atoms, ions and molecules. Counting them one by one is impractical, so chemists use the **mole** as a counting unit, just as a dozen represents 12 objects.

One mole contains exactly **$6.02214076 \times 10^{23}$** specified elementary entities. This fixed number is the Avogadro constant, N_A.

The entity must always be specified: atoms, molecules, ions, electrons, formula units, or another defined particle. One mole of NaCl means one mole of NaCl formula units, not one mole of Na atoms plus one mole of Cl atoms unless the composition is being counted separately.

Core relationships

Number of entities, $N = n \times N_A$
Amount in moles, $n = N / N_A$
$N_A = 6.02214076 \times 10^{23}\ mol^{-1}$

For an ionic solid, the smallest representative entity is a formula unit. Thus 1 mol of $CaCl_2$ contains 1 mol Ca^{2+} ions and 2 mol Cl^- ions, i.e. 3 mol of ions in total.

Examples

- 0.5 mol H_2 contains $0.5N_A$ molecules and $1.0N_A$ hydrogen atoms.
- 2 mol CO_2 contains 2 mol C atoms and 4 mol O atoms.
- 0.2 mol $Al_2(SO_4)_3$ contains 0.4 mol Al^{3+}, 0.6 mol S atoms and 2.4 mol O atoms.

2. Atomic Mass, Molecular Mass and Molar Mass

Unified atomic mass unit

One atomic mass unit (u) is defined as one-twelfth of the mass of one carbon-12 atom. Numerically, $1\ u \approx 1.66054 \times 10^{-27}$ kg.

Relative atomic mass is a dimensionless weighted average based on isotopic abundances. For an element with isotopes of masses m_i and fractional abundances f_i:

Average atomic mass = $\Sigma(f_i m_i)$, where $\Sigma f_i = 1$

For example, if an element has isotopes of mass 10 u (20%) and 11 u (80%), average atomic mass = 0.20(10) + 0.80(11) = 10.8 u.

Molecular and formula mass

Molecular mass is the sum of the atomic masses of all atoms in one molecule. Formula mass is used for ionic compounds and other substances represented by formula units.

- Molecular mass of H_2SO_4 = 2(1) + 32 + 4(16) = 98 u.
- Formula mass of $Ca(OH)_2$ = 40 + 2(16 + 1) = 74 u.

Molar mass

Molar mass (M) = mass of one mole; unit: g mol^{-1}
$n = m/M$; $m = nM$

The numerical value of molar mass in g mol^{-1} is equal to the relative molecular/formula mass, but the quantities and units are different.

Mass of one particle

Mass of one entity = molar mass (in g mol^{-1}) / N_A (in mol^{-1})

For a molecule of mass 18 u, mass per molecule ≈ $18 \times 1.66054 \times 10^{-24}$ g = 2.989×10^{-23} g.

3. Mole Calculations: The Conversion Map

Mass (m) ■ Moles (n) ■ Number of entities (N)
$n = m/M$ $N = nN_A$

For gases, another useful bridge is volume. At a specified temperature and pressure, ideal-gas volume is connected to moles by PV = nRT.

$PV = nRT$; $n = PV/RT$
At 273.15 K and 1 atm, ideal molar volume ≈ 22.414 L mol^{-1} (often rounded to 22.4 L mol^{-1}).

Do not use 22.4 L mol^{-1} at arbitrary conditions. At 1 bar and 273.15 K, the ideal molar volume is about 22.71 L mol^{-1}.

Density and molar mass of a gas

For an ideal gas: $d = PM/RT$; $M = dRT/P$
At the same T and P, density ratio = molar-mass ratio.

For a gas mixture, use total moles and the mixture's average molar mass, not the molar mass of an arbitrary component.

4. Stoichiometry and Balanced Chemical Equations

A balanced chemical equation gives the relative numbers of particles and hence the relative numbers of moles that react or form. Its coefficients are mole ratios, not mass ratios.

For $aA + bB \rightarrow cC + dD$:

n(A) consumed : n(B) consumed : n(C) formed : n(D) formed = a : b : c : d

Reliable workflow:

• Write correct formulas and balance the equation.

• Convert every given mass, volume or particle count into moles.

• Use the coefficient ratio to calculate the required moles.

• Convert the result into the requested unit and check significant figures.

Worked example 1: mass-to-mass

$CaCO_3 \rightarrow CaO + CO_2$. What mass of CO_2 is formed from 25.0 g pure $CaCO_3$?

$M(CaCO_3)$ = 100 g mol^{-1}; n = 25.0/100 = 0.250 mol. The ratio $CaCO_3$:CO_2 is 1:1, so $n(CO_2)$ = 0.250 mol. Mass = 0.250 × 44 = **11.0 g**.

Worked example 2: gas volume

$2KClO_3 \rightarrow 2KCl + 3O_2$. On complete decomposition of 24.5 g $KClO_3$ (M = 122.5), n = 0.200 mol. Oxygen formed = 0.200 × 3/2 = 0.300 mol. At 273.15 K and 1 atm, V ≈ 0.300 × 22.414 = **6.72 L**.

5. Limiting Reagent and Excess Reagent

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

Systematic method

• Convert each reactant quantity to moles.

• For each reactant, calculate n/coefficient.

• The smallest value identifies the limiting reagent.

• Calculate products using only the limiting reagent; calculate leftover excess reagent separately.

For aA + bB → products, compare n(A)/a and n(B)/b. The smaller ratio is limiting.

Worked example

$N_2 + 3H_2 \rightarrow 2NH_3$. If 2 mol N_2 and 3 mol H_2 are mixed, n/a values are 2/1 = 2 and 3/3 = 1. Hence H_2 is limiting. NH_3 formed = 3 × 2/3 = 2 mol. N_2 consumed = 1 mol, so 1 mol N_2 remains.

Advanced insight: extent of reaction

For a reaction with stoichiometric numbers ν_i (negative for reactants, positive for products), $n_i = n_{i,0} + \nu_i\xi$. The maximum allowed extent ξ is constrained by every reactant: $\xi \leq n_{i,0}/|\nu_i|$. The smallest bound controls completion.

6. Percentage Yield, Purity and Impure Samples

% yield = (actual yield / theoretical yield) × 100
% purity = (mass of pure substance / mass of sample) × 100

Theoretical yield is calculated from the limiting reagent assuming complete reaction and no loss. Actual yield is measured experimentally.

Worked example: purity

A 10.0 g limestone sample contains $CaCO_3$ and inert impurities. On heating, it gives 3.52 g CO_2. Since $CaCO_3 \rightarrow CaO + CO_2$, $n(CO_2) = 3.52/44 = 0.0800$ mol. Pure $CaCO_3 = 0.0800 \times 100 = 8.00$ g. Purity = 8.00/10.0 × 100 = **80.0%**.

If a sample is x% pure, pure-reactant mass = (x/100) × sample mass. Apply stoichiometry to the pure mass, not the total sample mass.

7. Composition: Percentage, Empirical and Molecular Formula

Mass percentage

% element = (mass of that element in 1 mol compound / molar mass of compound) × 100

For H_2O, mass percentages are H = 2/18 × 100 = 11.11% and O = 16/18 × 100 = 88.89%.

Empirical formula

The empirical formula gives the simplest whole-number mole ratio of atoms.

• Assume 100 g sample if percentages are given; each percentage becomes grams.

• Convert grams of each element to moles by dividing by atomic mass.

• Divide all mole values by the smallest.

• If needed, multiply all ratios by 2, 3, 4, etc. to obtain whole numbers.

Molecular formula

Molecular formula = $(\text{Empirical formula})_k$
k = molar mass / empirical-formula mass (must be a whole number for a valid molecular formula).

Example: empirical formula CH_2O has empirical-formula mass 30. If molar mass is 180 g mol^{-1}, k = 6 and molecular formula is $C_6H_{12}O_6$.

Combustion analysis

For a compound containing C, H and possibly O: all carbon becomes CO_2, and all hydrogen becomes H_2O under complete combustion.

$n(C) = n(CO_2) = m(CO_2)/44$
$n(H) = 2n(H_2O) = 2m(H_2O)/18$
If only C, H, O are present: $m(O) = \text{sample mass} - m(C) - m(H)$.

Oxygen is often found by mass difference; do not infer oxygen moles directly from the water and carbon dioxide masses.

8. Concentration Terms

Molarity (M)

Molarity = moles of solute / volume of solution in litres

Molarity depends on temperature because solution volume can change with temperature.

Molality (m)

Molality = moles of solute / mass of solvent in kg

Molality is independent of temperature as it uses mass, not volume.

Mole fraction

$x_i = n_i / \Sigma n_j$; $\Sigma x_i = 1$

For a binary mixture, $x_A + x_B = 1$. Mole fraction is dimensionless.

Mass percentage, volume percentage and ppm

Mass % = (mass solute / mass solution) × 100
Volume % = (volume component / volume solution) × 100
ppm by mass ≈ (mass solute / mass solution) × 10^6

For dilute aqueous solutions with density close to 1 g mL^{-1}, 1 ppm by mass is approximately 1 mg L^{-1}; this approximation is not universal.

Dilution and mixing

Dilution of same solute: $M_1V_1 = M_2V_2$
For mixing solutions without reaction: $n_{total} = \Sigma(M_iV_i)$, and $M_{final} = n_{total}/V_{final}$.

Use volumes in consistent units. If volumes are not additive, the final volume must be measured or provided.

9. Equivalent Concept and Normality

An equivalent is reaction-dependent. The n-factor is the number of equivalents per mole for the specified reaction.

Equivalent mass = molar mass / n-factor
Number of equivalents = mass / equivalent mass = moles × n-factor
Normality (N) = equivalents / litre of solution = M × n-factor

Acid-base examples: HCl has n-factor 1; H_2SO_4 has n-factor 2 for complete neutralization by a strong base; H_3PO_4 can have n-factor 1, 2 or 3 depending on the extent of neutralization.

Redox n-factor is the total electrons exchanged per formula unit in the specified reaction. It must be obtained from the balanced ionic/electron change, not memorized as a universal property.

At equivalence for a reaction: $N_1V_1 = N_2V_2$ (when normalities refer to the reaction's equivalent basis).

A safer general method is to equate reacting moles using the balanced equation. This avoids mistakes when n-factors differ by reaction.

10. Gas Mixtures and Partial Pressures

Ideal gas mixture: $P_{total} = \Sigma P_i$; $P_i = x_i P_{total}$
$x_i = n_i/n_{total}$

For gases collected over water, the dry gas pressure is $P_{dry} = P_{total} - P_{water\ vapour}$.

At the same temperature and pressure, gas-volume ratios equal mole ratios for ideal gases. For reacting gases, use the balanced equation and ensure all volumes refer to the same T and P.

Average molar mass of a mixture

$M_{avg} = \Sigma x_i M_i$ = (total mass)/(total moles)

If mass fractions w_i are known, $1/M_{avg} = \Sigma(w_i/M_i)$.

11. High-Value JEE Advanced Problem Strategies

A. Conservation approach

Track atoms, total mass, and charge. In a closed system, total mass is conserved. In ionic/redox equations, charge must also balance.

B. Sequential reactions

For A → B → C, calculate the amount of B actually formed in the first step before using it as the input to the second. Do not assume 100% conversion unless stated.

C. Mixture analysis

Introduce variables for unknown component amounts. Use independent experimental observations as equations. Check whether the number of independent equations is sufficient.

D. Hydrates

For salt·xH_2O: x = (moles of water lost)/(moles of anhydrous salt).

Obtain water lost from the mass difference after heating to constant mass, and calculate moles of anhydrous salt from its final mass.

E. Titration

Write the balanced reaction first. Convert titrant volume to litres, calculate moles from molarity, then use the reaction ratio to obtain analyte moles.

F. Significant figures and units

• Keep extra digits during intermediate steps; round only the final result.

• Use Kelvin in gas-law calculations and litres with R = 0.082057 L atm mol■¹ K■¹.

• Use R = 8.314 J mol■¹ K■¹ with SI pressure-volume-energy units.

• Check dimensional consistency: mass ÷ molar mass gives moles; molarity × volume in litres gives moles.

12. Solved JEE-Style Illustrations

Illustration 1: molecules and atoms

Find total atoms in 0.25 mol of $Al_2(SO_4)_3$. Each formula unit has 2 + 3 + 12 = 17 atoms. Total atoms = 0.25 × 17 × N_A = **$4.25N_A$** ≈ 2.56 × 10^{24} atoms.

Illustration 2: limiting reagent with masses

$4Al + 3O_2 \rightarrow 2Al_2O_3$. Mix 10.8 g Al and 9.6 g O_2. n(Al) = 10.8/27 = 0.40 mol; n(O_2) = 9.6/32 = 0.30 mol. Ratios are 0.40/4 = 0.10 and 0.30/3 = 0.10: exact stoichiometric mixture. Al_2O_3 formed = 0.20 mol = **20.4 g**.

Illustration 3: empirical formula

A compound contains 40.0% C, 6.67% H and 53.33% O. In 100 g: C = 40/12 = 3.333 mol; H = 6.67/1 = 6.67 mol; O = 53.33/16 = 3.333 mol. Divide by 3.333: 1:2:1. Empirical formula = **CH_2O**.

Illustration 4: gas mixture

A mixture has 2 mol He and 3 mol Ne at total pressure 10 atm. x(He) = 2/5, so P(He) = 4 atm; x(Ne) = 3/5, so P(Ne) = 6 atm.

Illustration 5: hydrate

A 2.50 g sample of $MgSO_4 \cdot xH_2O$ is heated and leaves 1.20 g $MgSO_4$. Water lost = 1.30 g, n(H_2O) = 1.30/18 = 0.0722 mol. n($MgSO_4$) = 1.20/120 = 0.0100 mol. x = 7.22, so within experimental rounding the hydrate is approximately **$MgSO_4 \cdot 7H_2O$**.

13. Practice Set: JEE Advanced Level

1. A sample contains 3.01×10^{23} molecules of H_2SO_4. Calculate (i) moles of acid, (ii) moles of oxygen atoms, and (iii) total number of atoms.

2. A 5.00 g mixture of Na_2CO_3 and $NaHCO_3$ is heated strongly: $2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$. The residue weighs 4.38 g. Find the initial mass percentage of $NaHCO_3$.

3. For $2A + B \rightarrow 3C$, initially 5 mol A and 2 mol B are present. Find the limiting reagent, maximum moles of C, and moles of excess reactant left.

4. A gaseous hydrocarbon gives 8.8 g CO_2 and 5.4 g H_2O on complete combustion of 2.9 g sample. Determine its empirical formula.

5. A 10.0 mL sample of H_2SO_4 is neutralized by 25.0 mL of 0.200 M NaOH. Find the molarity of the acid.

6. A gas mixture contains 16 g O_2 and 28 g N_2. Calculate mole fraction of each gas and average molar mass.

7. A 4.00 g impure sample of $CaCO_3$ gives 0.672 L CO_2 at STP (273.15 K, 1 atm). Find its percentage purity, assuming ideal gas behavior.

8. A 1.00 g compound containing C, H and O produces 1.466 g CO_2 and 0.600 g H_2O on combustion. Determine its empirical formula.

9. A solution is prepared by dissolving 9.80 g H_2SO_4 to make 500 mL solution. Calculate molarity and molality if the solution density is 1.05 g mL^{-1}.

10. For $MnO_4^- \rightarrow Mn^{2+}$ in acidic medium, determine the n-factor of $KMnO_4$ and its equivalent mass. Use M($KMnO_4$) = 158 g mol^{-1}.

11. A mixture of CH_4 and C_2H_6 contains 0.50 mol total and requires 1.30 mol O_2 for complete combustion. Find the mole fraction of each gas.

12. A 10.0 g sample of a metal oxide MO is reduced completely by hydrogen to give 7.00 g metal M and water. Determine the atomic mass of M if oxygen atomic mass is 16.

14. Answer Key with Reasoning

1. n = $3.01\times10^{23}/6.022\times10^{23}$ = 0.500 mol. O atoms = 4 × 0.500 = 2.00 mol. Atoms per molecule = 7; total atoms = 3.5N■ ≈ 2.11×10^{2}■.

2. Mass loss = 5.00 – 4.38 = 0.62 g = CO■ + H■O released per 2 mol NaHCO■. For x g NaHCO■, gas loss = x × 62/168. Thus x = 0.62×168/62 = 1.68 g; percentage = 33.9%.

3. n(A)/2 = 2.5; n(B)/1 = 2, so B is limiting. C = 3×2 = 6 mol. A consumed = 4 mol; A left = 1 mol.

4. n(C) = 8.8/44 = 0.20; n(H) = 2×(5.4/18) = 0.60. Ratio C:H = 1:3. Empirical formula CH■. (A stable neutral hydrocarbon molecular formula must be consistent with its valence; data may represent an empirical ratio.)

5. n(NaOH) = 0.0250×0.200 = 0.00500 mol. H■SO■ + 2NaOH → ...; acid moles = 0.00250. M = 0.00250/0.0100 = 0.250 M.

6. n(O■)=0.500, n(N■)=1.00; total 1.50. x(O■)=1/3; x(N■)=2/3. Mavg = (16+28)/1.50 = 29.33 g mol■¹.

7. n(CO■)=0.672/22.414=0.02998 mol. CaCO■ moles same; pure mass ≈ 3.00 g. Purity = 3.00/4.00×100 = 75.0%.

8. n(C)=1.466/44=0.03332; n(H)=2×0.600/18=0.06667. m(C)=0.400 g; m(H)=0.0667 g; m(O)=0.5333 g, n(O)=0.03333. Ratio ≈ 1:2:1; CH■O.

9. Moles acid = 9.80/98 = 0.100; M = 0.100/0.500 = 0.200 M. Solution mass = 1.05×500 = 525 g; solvent mass = 525–9.80 = 515.2 g = 0.5152 kg. Molality = 0.100/0.5152 = 0.194 mol kg■¹.

10. Mn changes from +7 to +2, gain of 5 e■; n-factor = 5. Equivalent mass = 158/5 = 31.6 g eq■¹.

11. Let CH■ = x and C■H■ = 0.50–x. O■ needed = 2x + 3.5(0.50–x) = 1.30. Solving gives x = 0.30 mol CH■ and 0.20 mol C■H■. Mole fractions 0.60 and 0.40.

12. O mass = 10–7 = 3.00 g; n(O)=3/16=0.1875 mol. MO has 1:1 ratio, so n(M)=0.1875. Atomic mass M = 7/0.1875 = 37.33 g mol■¹.

15. Formula Sheet and Final Checklist

Fundamental: n = m/M = N/N_A ; N = nN_A ; PV = nRT
Stoichiometry: For aA+bB→..., compare n/a and n/b; smallest is limiting.
Yield and purity: % yield = actual/theoretical ×100 ; % purity = pure/sample ×100.
Composition: % element = mass of element per mole compound / molar mass ×100.
Concentration: M = n/V(L) ; m = n/kg solvent ; x■ = n■/Σn ; N = M×n-factor.
Gas mixture: P■=x■Ptotal ; Mavg=Σx■M■ ; Pdry=Ptotal–Pwater.
Dilution: M■V■=M■V■ for same solute with no reaction.

- Did I specify what entity one mole counts?
- Did I balance the equation before applying mole ratios?
- Did I convert all given quantities to moles?
- Did I identify the limiting reagent rather than assume one?
- Did I account for purity, yield, and excess reagent separately?
- Did I use the correct temperature, pressure and gas constant?
- Did I define the reaction before assigning an equivalent n-factor?
- Are units and significant figures consistent?

Important constants: N_A = 6.02214076×10²³ mol■¹; R = 0.082057 L atm mol■¹ K■¹ = 8.314 J mol■¹ K■¹; ideal molar volume at 273.15 K and 1 atm ≈ 22.414 L mol■¹.

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