ExamFitter

examfitter.com · Free AP Chemistry quizzes, revision notes and worksheets

Topic 3.7: Solutions and Mixtures

Revision Notes · AP Chemistry

UNIT 3 · PROPERTIES OF SUBSTANCES AND MIXTURES

On ExamFitter.com: Topic 3.7 Quiz · Topic 3.7 Worksheet · Unit 3: All Topics · AP Chemistry Home

Mixtures: Variable Composition, Uniform When Dissolved

A mixture holds two or more substances physically combined, each keeping its own identity. Unlike a compound, whose elements sit in a fixed mass ratio, its composition varies freely with how much of each component was added.

A homogeneous mixture has the same composition and properties throughout. A solution is a homogeneous mixture: the solute is dispersed uniformly through the solvent, so every sample drawn from it has the same concentration.

The solute is the substance that dissolves. The solvent does the dissolving and is usually the component present in the greatest amount: mix 20 mL of ethanol with 80 mL of water and water is the solvent, because there is more of it.

Solutions are not only liquids. Brass, copper and zinc uniformly mixed, is a solid solution; gases dissolve in liquids, and liquids in liquids. What makes a mixture a solution is uniform dispersion at the particle level, not its phase.

A heterogeneous mixture keeps distinguishable regions or phases: sand in water remains a separate solid phase, its composition differing from place to place.

The Particulate Picture: What Is Actually in the Beaker

An ionic solid dissolves by dissociating: $NaCl(s) \rightarrow Na^{+}(aq) + Cl^{-}(aq)$. The crystal's ions separate and move independently through the liquid. No intact NaCl units survive; the dissolved species are the ions themselves.

© 2026 ExamFitter.com · Free AP Chemistry revision notes

Each ion travels inside a hydration shell of water molecules. Water is polar, its oxygen end carrying a partial negative charge, so around a cation such as Na^+ the oxygen ends point inward. Around an anion such as Cl^- the hydrogen ends point inward instead.

A molecular solute dissolves differently. Sucrose molecules separate and disperse as intact, neutral molecules, hydrated but never split into ions. A sucrose solution holds molecules; a sodium chloride solution holds ions. That difference drives the particle counts and the conductivity of the two.

Solution particles are far too small to scatter light, so a solution shows no Tyndall effect, and they never settle out on standing. They pass straight through filter paper, so filtration cannot recover a dissolved solute; it removes only undissolved solids, such as a precipitate.

Molarity: Moles per Litre of Solution

Molarity is the amount of solute per litre of solution: M = n_solute ÷ V_solution, in mol/L. The volume is that of the finished solution, not the solvent started with, because dissolving changes the total volume.

The relationship runs three ways: n = M × V, and V = n ÷ M. Convert every volume to litres before substituting: 250 mL is 0.250 L.

Dissolve the solute in less solvent than the final volume, transfer it to a volumetric flask, then add solvent until the meniscus sits on the calibration mark. Fill past the mark and the molarity comes out lower than intended: the moles are unchanged, but the volume is too large.

Molarity is temperature dependent. Warm a solution and the liquid expands: the same moles fill more litres, so the molarity falls slightly. Mass based measures such as mass percent do not drift this way, because mass does not change with temperature.

Mass percent works in mass, molarity in volume. Converting between them requires the solution's density, the property linking a known mass of solution to the volume it occupies.

Counting Ions: Formula Units Are Not Particles

For an ionic solute, each ion's concentration is the molarity multiplied by that ion's count in the formula. In 0.20 M $Al(NO_3)_3$, aluminium ions sit at 0.20 M while nitrate ions sit at 3 × 0.20 = 0.60 M, because each formula unit releases three.

Total particle concentration adds every dissolved species. A 0.40 M $CaCl_2$ solution holds 0.40 M of Ca^{2+} and 0.80 M of Cl^-: 1.20 M of dissolved particles, three times the formula unit molarity.

© 2026 ExamFitter.com · Free AP Chemistry revision notes

Electrical conductivity follows the total concentration of mobile ions, and nothing else. Among 1.0 M solutions, $AlCl_3$ releases 4.0 M of ions and conducts most strongly; sucrose releases no ions at all, and its solution does not conduct however concentrated it is.

Dilution and Mixing: The Moles Do Not Change

Dilution adds solvent and nothing else. The moles of solute are untouched; only the spacing between particles grows as the volume rises and the molarity falls. At the particulate level it is simply the same particles, drawn further apart.

Because the moles match before and after, $M_1V_1 = M_2V_2$. The two volumes may use any unit, provided both sides use the same one.

Mixing two solutions of the same solute works the same way: add the moles from each portion, add the volumes, and divide. The result is the simple average of the two molarities only when the volumes are equal; otherwise the larger portion pulls the result toward its own concentration.

Macroscopic properties track concentration, so they dilute too. A blue $Cu(NO_3)_2$ solution pales as water is added, because its colour intensity follows the concentration of Cu^{2+} ions; the fading signals that concentration has fallen.

Worked Examples

Example 1: Molarity from a measured mass. 5.85 g of NaCl (molar mass 58.44 g/mol) is dissolved and made up to 250.0 mL. First the amount: n = 5.85 ÷ 58.44 ≈ 0.100 mol. Then M = 0.100 ÷ 0.2500 = 0.400 M.

Example 2: Ion concentrations, then a mole count. A 0.300 M $CaCl_2$ solution holds Ca^{2+} at 0.300 M and Cl^- at 2 × 0.300 = 0.600 M, 0.900 M of ions in total. In 150.0 mL of it, the moles of Cl^- are 0.600 × 0.1500 = 0.0900 mol: that one ion's concentration, not the compound's, multiplies the volume.

Example 3: A dilution. 25.0 mL of 6.00 M HCl is diluted to a final volume of 500.0 mL. The portion transferred carries 6.00 × 0.0250 = 0.150 mol, and dilution neither creates nor destroys solute. M_2 = (6.00 × 25.0) ÷ 500.0 = 0.300 M, one twentieth of the stock.

© 2026 ExamFitter.com · Free AP Chemistry revision notes

Example 4: Mixing. Mix 100.0 mL of 0.200 M KCl with 300.0 mL of 0.400 M KCl. The moles are 0.200 × 0.1000 = 0.0200 mol and 0.400 × 0.3000 = 0.120 mol, totalling 0.140 mol in 0.4000 L, so M = 0.140 ÷ 0.4000 = 0.350 M. Not the 0.300 M average: the larger portion supplies three quarters of the moles, so the mixture sits three quarters of the way from 0.200 M toward 0.400 M.

Traps That Cost Marks in Topic 3.7

Dividing by the solvent volume. Molarity divides by litres of finished solution. Dissolve a solute in exactly 1.0 L of water and the solution's volume is not 1.0 L; make the solution up to the mark instead.

Counting formula units as particles. A 0.40 M $CaCl_2$ solution is 1.20 M in dissolved particles. Conductivity and every particle count run on the ions actually present, never on the units in the formula.

Expecting sucrose to behave like salt. A molecular solute disperses as whole molecules: one particle per formula unit, no ions, and no conductivity at any concentration.

Averaging molarities when mixing. The simple average is correct only for equal volumes. For any other mix, add the moles, add the volumes, then divide.

Believing dilution removes solute. Adding water changes spacing and concentration only; the moles and mass of solute are exactly what they were.

Trusting one molarity at every temperature. A molarity is stamped at the temperature of preparation: warm the flask and the reading falls, because the volume grew while the moles stood still.

© 2026 ExamFitter.com · Free AP Chemistry revision notes

Key Relationships to Remember

- Solution: a homogeneous mixture; the solute is uniformly dispersed in the solvent, the component present in the greatest amount; every sample has the same concentration
- M = n_solute ÷ V_solution, V in litres of finished solution; n = M × V; prepare by dissolving first, then making up to the mark in a volumetric flask
- Ionic solutes dissociate into hydrated ions (water's oxygen ends toward cations, hydrogen ends toward anions); molecular solutes disperse as intact neutral molecules and do not conduct
- Ion concentration = molarity × the ion's count in the formula; total particle concentration adds all dissolved species; conductivity ranks by total mobile ion concentration
- Dilution: moles constant, $M_1V_1 = M_2V_2$; mixing: add moles, add volumes, divide (average only for equal volumes)
- Molarity depends on temperature through volume; mass percent does not; mass percent to molarity needs the density

Continue Revising

Topic 3.7 Quiz: Solutions and Mixtures

Topic 3.7 Worksheet

Unit 3: All Topics

© 2026 ExamFitter.com · Free AP Chemistry revision notes