Collegedunia NCERT Formula Sheet

The Ultimate Formula Reference for Class 12 (12th) Chemistry — NCERT 2026-27

Chapter 1: Solutions

Concentration · Solubility · Henry's & Raoult's Laws · Ideal/Non-ideal · Colligative Properties · van't Hoff Factor

Quick Reference — Symbols, Units & Key Constants

Symbol	**Meaning**	**SI unit**	**Notes**
x_A, x_B	Mole fraction of A, B	dimensionless	$x_A + x_B = 1$ for binary solution
M	Molarity	mol L^{-1}	Depends on T (volume changes)
m	Molality	mol kg^{-1}	T-independent (mass invariant)
K_H	Henry's law constant	bar or Pa	Larger $K_H \Rightarrow$ lower solubility
p_1°, p_2°	Pure-component vapour pressure	bar, kPa, mm Hg	Increases with T
ΔT_b, ΔT_f	BP elevation, FP depression	K	Always positive
K_b, K_f	Ebullioscopic, cryoscopic constant	K kg mol^{-1}	Solvent-specific (water: 0.52, 1.86)
Π	Osmotic pressure	bar, atm, Pa	$\Pi = CRT$ for dilute solutions
i	van't Hoff factor	dimensionless	$i > 1$ dissociation, $i < 1$ association
R	Gas constant	0.0821 L atm K^{-1} mol^{-1}	= 0.0831 L bar K^{-1} mol^{-1} = 8.314 J K^{-1} mol^{-1}

1 Types of Solutions (NCERT 1.1)

NCERT Section 1.1 defines a **solution** as a homogeneous mixture of two or more chemically non-reacting substances. A binary solution has two components — the one in larger amount is the **solvent**, the other is the **solute**. The physical state of the solvent decides the state of the solution.

Nine types of solutions (NCERT Table 1.1)

Gaseous solvent: gas-in-gas (air), liquid-in-gas (water vapour in air), solid-in-gas (camphor in N_2).
Liquid solvent: gas-in-liquid (O_2 in water, soda), liquid-in-liquid (ethanol–water), solid-in-liquid (NaCl, glucose in water).
Solid solvent: gas-in-solid (H_2 in palladium), liquid-in-solid (amalgam — Hg in Na), solid-in-solid

(alloys — brass = Cu/Zn).
Most laboratory solutions are **solid-in-liquid** or **liquid-in-liquid**. Air is the most familiar **gas-in-gas** solution.

Dilute vs concentrated; saturated vs unsaturated

A **dilute** solution contains a small amount of solute relative to solvent; a **concentrated** solution contains a large amount. A **saturated** solution is one in which no more solute can dissolve at a given temperature — the dissolved solute is in dynamic equilibrium with the undissolved solute. **Unsaturated** solutions can take up more solute. **Supersaturated** solutions hold more than the equilibrium amount and crystallise on disturbance.

2 Expressing Concentration of Solutions (NCERT 1.2)

NCERT Section 1.2 lists **seven** quantitative ways to express concentration. Mass % and mole fraction are dimensionless; molarity and molality have units; ppm is used for very dilute solutions (pollutants, trace metals).

Mass percentage (w/w) & Volume percentage (V/V)

$$\text{Mass }\%(A) = \frac{\text{Mass of } A}{\text{Mass of solution}} \times 100$$

$$\text{Volume }\%(A) = \frac{\text{Volume of } A}{\text{Volume of solution}} \times 100$$

Used for commercial reagents (e.g., 35% w/w HCl, 35% v/v ethanol antifreeze). Both are dimensionless; the choice depends on whether components are easier to mass or to measure by volume.

Mass by volume % & Parts per million

$$\text{Mass–volume }\% = \frac{\text{Mass of solute (g)}}{\text{Volume of solution (mL)}} \times 100$$

$$\text{ppm} = \frac{\text{Number of parts of component}}{\text{Total number of parts of all components}} \times 10^6$$

Mass–volume % is common in medicine and pharmacy. **ppm** is used for trace concentrations (e.g., fluoride in drinking water ≈ 1 ppm). ppm can be mass/mass, volume/volume, or mass/volume.

Mole fraction (x)

$$x_A = \frac{n_A}{n_A + n_B}, \quad x_B = \frac{n_B}{n_A + n_B}$$

For binary solution: $x_A + x_B = 1$; in general: $\sum_i x_i = 1$.

$n_A = \frac{w_A}{M_A}$, where w_A = mass and M_A = molar mass of A.

Mole fraction is **temperature-independent**. It is the natural variable in Raoult's law, partial pressures, and any equation involving the relative numbers of molecules.

Molarity (M)

$$M = \text{(Moles of solute) / (Volume of solution in L)} = (w_B/M_B)\,/\,V_{\text{soln}}(\text{L})$$

Units: mol L^{-1} (also written M).

Depends on temperature because the volume of a solution expands/contracts with T. Most volumetric (titration) calculations use molarity. Convert: $w_B = M \times M_B \times V$.

Molality (m)

$\boldsymbol{m}$ = (Moles of solute) / (Mass of solvent in kg) = $(\boldsymbol{w_B}/\boldsymbol{M_B})$ / $\boldsymbol{w_A}$(kg)

Units: mol kg^{-1} (also written $\boldsymbol{m}$).

Independent of temperature (mass doesn't change with T). Preferred in colligative-property formulas ($\boldsymbol{\Delta T_b, \Delta T_f}$) where T-stability matters. For dilute aqueous solutions $\boldsymbol{M \approx m}$.

Inter-conversion shortcut

For a solution of density d (g mL^{-1}) and solute molar mass M_B: $M = \dfrac{m \cdot d \cdot 1000}{1000 + m \cdot M_B}$ and $x_B = \dfrac{m \cdot M_A}{1000 + m \cdot M_A}$

Quick check: at low m, $M \approx m \cdot d$ and $x_B \approx m \cdot M_A/1000$.

3 Solubility (NCERT 1.3)

NCERT Section 1.3 defines **solubility** as the maximum amount of a substance that can be dissolved in a specified amount of solvent at a specified temperature. The two sub-sections cover solids dissolved in liquids (1.3.1) and gases dissolved in liquids (1.3.2 — this is where **Henry's law** appears).

Solubility of a solid in a liquid (NCERT 1.3.1)

"Like dissolves like": polar solutes (NaCl, sugar) dissolve in polar solvents (water); non-polar solutes (naphthalene, anthracene) dissolve in non-polar solvents (benzene). At saturation, Solute (solid) $\rightleftharpoons$ Solute (in solution).

Effect of T: if dissolution is endothermic ($\Delta_{\text{sol}}H > 0$), solubility *increases* with T (e.g., KNO_3, $NaNO_3$); if exothermic ($\Delta_{\text{sol}}H < 0$), solubility *decreases* with T (e.g., $Ce_2(SO_4)_3$).

Effect of P: negligible for solids/liquids in liquids (incompressible).

Henry's law (gas in liquid, NCERT 1.3.2)

Statement: At constant temperature, the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas above the liquid.

Form (mole fraction, NCERT eq. 1.11): $\boldsymbol{p = K_H \cdot x}$

where $\boldsymbol{p}$ = partial pressure of the gas above the solution (bar or Pa); $\boldsymbol{x}$ = mole fraction of gas in solution; $\boldsymbol{K_H}$ = Henry's law constant (same units as $\boldsymbol{p}$, since $\boldsymbol{x}$ is dimensionless).

Larger $\boldsymbol{K_H}$ at a given $\boldsymbol{T}$ $\Rightarrow$ *lower* solubility (rearrange: $\boldsymbol{x = p/K_H}$). $\boldsymbol{K_H}$ *increases* with temperature, so gases are less soluble in hot water — this is why aquatic life suffers in summer and why warm soda goes flat faster.

Typical K_H values (NCERT Table 1.2)

He (293 K): $\mathbf{144.97}$ kbar H$_2$ (293 K): $\mathbf{69.16}$ kbar N$_2$: $\mathbf{76.48}$ kbar (293 K), $\mathbf{88.84}$ kbar (303 K)
O$_2$: $\mathbf{34.86}$ kbar (293 K), $\mathbf{46.82}$ kbar (303 K) Argon (298 K): $\mathbf{40.3}$ kbar CO$_2$ (298 K): $\mathbf{1.67}$ kbar
Methane (298 K): $\mathbf{0.413}$ kbar Vinyl chloride (298 K): $\mathbf{0.611}$ kbar HCHO (298 K): $\mathbf{1.83 \times 10^{-5}}$ kbar

HCHO has the lowest $\boldsymbol{K_H}$ $\Rightarrow$ extremely high water solubility (commercial formalin is 37–40% HCHO). Among common gases, CO$_2$ ($\boldsymbol{K_H = 1.67}$ kbar) is more soluble than O$_2$, N$_2$ or H$_2$, which is why it can be packed into soft drinks under pressure.

Applications of Henry's law (NCERT)

(i) Soda water / soft drinks: bottled under high p_{CO_2} to keep CO_2 dissolved.

(ii) Scuba diving (the bends): at depth, high pressure forces extra N_2 to dissolve in blood; on rapid ascent it bubbles out (*bends*). To prevent this, divers' tanks are filled with **air diluted with helium**

(11.7% He, 56.2% N_2, 32.1% O_2).
(iii) High altitude (anoxia): low p_{O_2} $\Rightarrow$ less O_2 dissolved in blood $\Rightarrow$ climbers feel weak and unable to think clearly.

Henry's law — units of K_H

K_H has *pressure units* (bar or Pa) when written as $p = K_H\,x$, because x is dimensionless. Tabulated values in NCERT Table 1.2 are in kbar (1 kbar = 10^3 bar). Higher $K_H \Rightarrow$ *lower* solubility; never read K_H as a direct measure of solubility.

4 Vapour Pressure of Liquid Solutions (NCERT 1.4)

NCERT Section 1.4 develops Raoult's law in three sub-sections: 1.4.1 binary liquid solutions, 1.4.2 Raoult's law as a special case of Henry's law, 1.4.3 solid solutes in liquid solvent (relative lowering of vapour pressure). This section gathers the vapour-pressure relations themselves; the colligative consequences are in Section 6 of this sheet.

Raoult's law for binary liquid solutions (NCERT 1.4.1)

For a solution of two volatile liquids A and B:

$$p_A = p_A^\circ\,x_A, \quad p_B = p_B^\circ\,x_B \quad \text{(partial pressures)}$$

$$p_{\text{total}} = p_A + p_B = p_A^\circ x_A + p_B^\circ x_B = p_B^\circ + (p_A^\circ - p_B^\circ)x_A$$

where p_A°, p_B° = vapour pressures of pure A and B; x_A, x_B = mole fractions in the *liquid* phase.

p_{total} varies **linearly** with x_A between p_B° (at $x_A = 0$) and p_A° (at $x_A = 1$). This is the defining linear plot of an ideal solution.

Vapour-phase composition (Dalton's law)

$$y_A = \frac{p_A}{p_{\text{total}}} = \frac{p_A^\circ x_A}{p_{\text{total}}}, \quad y_B = \frac{p_B}{p_{\text{total}}}$$

where y_A, y_B = mole fractions in the *vapour* phase.

The vapour is always **richer in the more volatile component** (the one with higher p°). This is the principle behind **fractional distillation**.

Raoult's law as a special case of Henry's law (NCERT 1.4.2)

Henry's law: $p_{\text{gas}} = K_H\,x_{\text{gas in solution}}$
Raoult's law: $p_A = p_A^\circ\,x_A$

Both say partial pressure is proportional to mole fraction in solution. Raoult's law is the special case where $K_H = p_A^\circ$.

Equivalently: **Henry** $\rightarrow$ **Raoult** as the solute becomes the major component. For very dilute solutions of a gas, $K_H \neq p_{\text{gas}}^\circ$ because gas–solvent interactions differ from gas–gas interactions.

Vapour pressure of solutions with non-volatile solute (NCERT 1.4.3)

If the solute B is **non-volatile** ($p_B^\circ = 0$), then only solvent contributes to the vapour:

$$p_{\text{soln}} = p_A = p_A^\circ\,x_A$$

Since $x_A < 1$, $p_{\text{soln}} < p_A^\circ$ — the solvent's vapour pressure is *lowered*.

This lowering is the parent quantity behind *all four* colligative properties (relative lowering, ΔT_b, ΔT_f, Π). Magnitude depends only on the **number of solute particles**, not their identity.

5 Ideal and Non-ideal Solutions (NCERT 1.5)

NCERT Section 1.5 classifies solutions by how well they obey Raoult's law. **Ideal** solutions (1.5.1) obey it at every composition; **non-ideal** solutions (1.5.2) deviate, either positively or negatively, and the sign of deviation predicts whether an azeotrope forms.

Ideal solution conditions (NCERT 1.5.1)

A solution is ideal if it obeys Raoult's law over the entire composition range AND:

(i) $\Delta_{\text{mix}}H = 0$ (no heat absorbed/released on mixing)
(ii) $\Delta_{\text{mix}}V = 0$ (no volume change on mixing)
(iii) A–B intermolecular interactions $\approx$ A–A $\approx$ B–B interactions

Examples: n-hexane + n-heptane; bromoethane + chloroethane; benzene + toluene; CCl_4 + $SiCl_4$. All have similar shapes/polarities.

Non-ideal solutions — positive deviation (NCERT 1.5.2)

$p_A > p_A^\circ x_A$, $p_B > p_B^\circ x_B$, p_{total} shows a **maximum** above the Raoult line.
Cause: A–B interactions *weaker* than A–A, B–B.
Sign of thermodynamics: $\Delta_{\text{mix}}H > 0$ (endothermic), $\Delta_{\text{mix}}V > 0$ (expansion).
Examples: ethanol + acetone; CS_2 + acetone; water + ethanol; CCl_4 + $CHCl_3$.

Ethanol in pure form forms H-bonds among ethanol molecules; adding acetone breaks these bonds, so escape into vapour becomes easier $\Rightarrow$ higher p.

Non-ideal solutions — negative deviation (NCERT 1.5.2)

$p_A < p_A^\circ x_A$, $p_B < p_B^\circ x_B$, p_{total} shows a **minimum** below the Raoult line.
Cause: A–B interactions *stronger* than A–A, B–B (often new H-bonding).
Sign of thermodynamics: $\Delta_{\text{mix}}H < 0$ (exothermic), $\Delta_{\text{mix}}V < 0$ (contraction).
Examples: phenol + aniline; $CHCl_3$ + acetone ($CHCl_3$ H-bonds to C=O); HNO_3 + H_2O; HCl + H_2O.

Strong A–B attractions trap molecules in the liquid phase $\Rightarrow$ fewer molecules escape to vapour $\Rightarrow$ lower p.

Azeotropes

An **azeotrope** is a binary mixture that boils at a single constant temperature and distils *without* change in composition — so it cannot be separated by fractional distillation.
Minimum-boiling azeotrope (forms at composition of maximum p): from **positive deviation** systems. Example (NCERT): *ethanol–water $\approx$ 95% ethanol by volume* (the fermentation limit).
Maximum-boiling azeotrope (forms at composition of minimum p): from **negative deviation** systems. Example (NCERT): *HNO_3–water at 68% HNO_3 + 32% water by mass, b.p. 393.5 K* (this is concentrated lab nitric acid).

Spotting deviation type in a question

If the question describes mixing molecules that can form *new* H-bonds with each other (e.g., chloroform + acetone, phenol + aniline) $\Rightarrow$ **negative** deviation.
If the question describes mixing molecules that *disrupt* existing H-bonds (e.g., ethanol + acetone, water + ethanol) $\Rightarrow$ **positive** deviation.

6 Colligative Properties (NCERT 1.6)

NCERT Section 1.6 introduces **colligative properties** — properties that depend only on the *number* of solute particles, not their identity. There are four: relative lowering of vapour pressure (1.6.1), elevation of boiling point (1.6.2), depression of freezing point (1.6.3), osmotic pressure (1.6.4). All follow from the lowering of solvent vapour pressure derived in Section 1.4.3.

Relative lowering of vapour pressure (NCERT 1.6.1)

$\Delta p/p_A^\circ \;=\; (p_A^\circ - p_A)/p_A^\circ \;=\; x_B \;=\; n_B/(n_A + n_B)$

For **dilute solutions** ($n_B \ll n_A$): $(p_A^\circ - p_A)/p_A^\circ \;\approx\; n_B/n_A \;=\; (w_B \cdot M_A)/(w_A \cdot M_B)$

where w_A, w_B = masses; M_A, M_B = molar masses of solvent, solute.

Rearranging gives **Ostwald–Walker method** for molar-mass determination: $M_B = (w_B\, M_A\, p_A^\circ)\,/\,(w_A\,(p_A^\circ - p_A))$.

Elevation of boiling point (NCERT 1.6.2)

$\Delta T_b = T_b - T_b^\circ = K_b \cdot m$

where ΔT_b = boiling-point elevation (K); K_b = **ebullioscopic** (molal boiling-point elevation) constant of solvent (K kg mol^{-1}); m = molality (mol kg^{-1}).

Determination of molar mass: $m = (w_B/M_B)/(w_A/1000) \Rightarrow M_B = 1000 \cdot K_b \cdot w_B/(\Delta T_b \cdot w_A)$ (w_A, w_B in g).

Adding a non-volatile solute lowers p at every T, so the solution must be heated *higher* than T_b° to bring its p up to p_{atm}. K_b depends only on solvent: water 0.52, benzene 2.53, $CHCl_3$ 3.63 K kg mol^{-1}.

Theoretical K_b from solvent properties

$K_b \;=\; (R \cdot T_b^2 \cdot M_A)\,/\,(1000 \cdot \Delta_{\text{vap}}H)$

where R = gas constant; T_b = solvent's normal boiling point (K); M_A = molar mass of solvent; $\Delta_{\text{vap}}H$ = molar enthalpy of vaporisation.

Often quoted but rarely *calculated* in CBSE Q's — it explains *why* K_b differs across solvents (water's high $\Delta_{\text{vap}}H$ gives a small K_b = 0.52, hence small ΔT_b).

Depression of freezing point (NCERT 1.6.3)

$\Delta T_f = T_f^\circ - T_f = K_f \cdot m$

where ΔT_f = freezing-point depression (K); K_f = **cryoscopic** (molal freezing-point depression) constant (K kg mol^{-1}); m = molality.

Determination of molar mass: $M_B = 1000 \cdot K_f \cdot w_B/(\Delta T_f \cdot w_A)$

At T_f, the solid solvent is in equilibrium with the liquid solvent. Lowering p_{liquid} shifts this equilibrium $\Rightarrow$ freezing happens at a lower T. K_f values: water 1.86, benzene 5.12, $CHCl_3$ 4.79, camphor 39.7 K kg mol^{-1}. Camphor's huge K_f makes it ideal for **Rast's method**.

Theoretical K_f from solvent properties

$K_f \;=\; (R \cdot T_f^2 \cdot M_A)\,/\,(1000 \cdot \Delta_{\text{fus}}H)$

where T_f = solvent's normal freezing point; $\Delta_{\text{fus}}H$ = molar enthalpy of fusion.

Same logic as for K_b: high $\Delta_{\text{fus}}H$ (water) gives small K_f; low $\Delta_{\text{fus}}H$ (camphor) gives huge K_f.

Osmotic pressure (NCERT 1.6.4)

$\Pi = C \cdot R \cdot T = (n_B/V) \cdot RT = w_B \cdot RT/(M_B \cdot V)$

where Π = osmotic pressure (bar, atm or Pa); C = molar concentration (mol L^{-1}); R = gas constant in matching units; T = absolute temperature (K); V = volume of solution (L).

Determination of molar mass: $M_B = w_B \cdot RT/(\Pi \cdot V)$

Best for macromolecules (proteins, polymers) because (i) measurements are at room T so the biomolecule isn't denatured, (ii) Π stays measurable even at tiny C, (iii) molarity (not molality) is the natural concentration variable.

Osmosis, semipermeable membranes & isotonic solutions (NCERT 1.6.4)

Osmosis is the net flow of *solvent* from a region of *lower* solute concentration to a region of *higher* solute concentration through a semipermeable membrane (SPM). The pressure that just stops this flow is the **osmotic pressure** Π.
Isotonic solutions have the *same* Π at the same T (e.g., 0.9% NaCl saline w.r.t. blood). **Hypertonic** ($\Pi >$ that of cell) shrinks cells; **hypotonic** ($\Pi <$) makes them swell/burst.
SPMs: natural — pig bladder, parchment; synthetic — cellulose acetate (used in reverse-osmosis units).

Reverse osmosis (NCERT 1.6.5)

Applying $P_{\text{ext}} > \Pi$ on the higher-concentration side forces solvent through the SPM in the *opposite* direction — pure solvent comes out.

Use: desalination of seawater (typical $\Pi_{\text{seawater}} \approx 30$ atm $\Rightarrow$ pumps work at $\sim$ 50 atm).

The membrane is usually cellulose acetate, which passes water but rejects Na^+, Cl^-, and most organic impurities. RO units in homes and ships are based on this principle.

Choosing the right colligative property

For **small molecules** ($M_B \lesssim 1000$ g/mol): use ΔT_b or ΔT_f — temperature changes are large enough to measure.
For **macromolecules** ($M_B \gtrsim 10^4$ g/mol): use Π — the only property that stays measurable at very low molar concentration, and it doesn't require heating.

ΔT_f sign convention

ΔT_f is always defined *positive*: $\Delta T_f = T_f^\circ - T_f$ (pure $-$ solution). The freezing point of the solution is *lower*, but the depression itself is reported as a positive number. Many marks lost by writing $\Delta T_f = T_f - T_f^\circ$ and getting a negative answer.

7 Abnormal Molar Masses & van't Hoff Factor (NCERT 1.7)

NCERT Section 1.7 handles solutes that dissociate (NaCl, KCl, electrolytes) or associate (benzoic acid in benzene, acetic acid in benzene) in solution. The *actual* number of particles differs from what the formula suggests, so observed colligative effects are abnormal. The **van't Hoff factor** i corrects all four colligative-property formulas.

van't Hoff factor (i) — definition

$$\begin{aligned} i &= \text{(Observed colligative property) / (Theoretical colligative property)} \\ &= \text{(Normal } M_B\text{) / (Observed } M_B\text{)} \\ &= \text{(Actual particles in solution) / (Particles before dissociation)} \end{aligned}$$

$i > 1 \Rightarrow$ **dissociation** (more particles); $i < 1 \Rightarrow$ **association** (fewer particles); $i = 1 \Rightarrow$ no change (e.g., glucose, urea in water).

Modified colligative-property equations (with i)

$$\frac{p_A^\circ - p_A}{p_A^\circ} = i\,x_B$$

$$\Delta T_b = i\,K_b\,m$$

$$\Delta T_f = i\,K_f\,m$$

$$\Pi = i\,C\,RT$$

Always use the i-corrected forms for ionic solutes (NaCl, KCl, $CaCl_2$, $MgSO_4$) and for solutes that dimerise in low-polarity solvents.

Dissociation: α from i

For $A_n \rightleftharpoons n\,A$ with degree of dissociation α:

$$i = 1 + (n-1)\,\alpha \quad \Rightarrow \quad \alpha = \frac{i-1}{n-1}$$

Examples (ideal, complete): $NaCl \rightarrow Na^+ + Cl^-$, $n = 2$, $i = 2$; $K_2SO_4 \rightarrow 2\,K^+ + SO_4^{2-}$, $n = 3$, $i = 3$.

Real values are slightly less than these due to ion pairing in solution (the more concentrated, the lower the apparent i).

Association: α from i

For $n\,A \rightleftharpoons A_n$ with degree of association α:

$$i = 1 - \alpha + \frac{\alpha}{n} = 1 - \alpha\left(\frac{n-1}{n}\right) \quad \Rightarrow \quad \alpha = \frac{(1-i)\,n}{n-1}$$

Examples: benzoic acid in benzene forms dimers ($n = 2$, $i \approx 0.5$); ethanoic (acetic) acid in benzene similarly dimerises through H-bonding.

Association is favoured in **low-polarity solvents** (benzene, $CHCl_3$) where carboxylic acid H-bonding pairs are not disrupted by solvent.

i values to remember

Non-electrolytes (glucose, urea, sucrose) in water: $i = 1$.
Strong 1:1 electrolytes (NaCl, KCl, $AgNO_3$): $i \to 2$.
Strong 1:2 / 2:1 (Na_2SO_4, $BaCl_2$, K_2SO_4): $i \to 3$.
Carboxylic acid dimers in benzene: $i \to 0.5$.

Experimental i values (NCERT Table 1.4)

Salt	i (0.1 m)	i (0.01 m)	i (0.001 m)	i (complete)
NaCl	1.87	1.94	1.97	2.00
KCl	1.85	1.94	1.98	2.00
$MgSO_4$	1.21	1.53	1.82	2.00
K_2SO_4	2.32	2.70	2.84	3.00

i approaches its "complete dissociation" limit as the solution is diluted (less ion pairing). $MgSO_4$ lags behind the alkali halides because Mg^{2+}–SO_4^{2-} ion pairing is strong (2+/2— Coulombic attraction).

i in Π for ionic vs molecular solutes

For Q's like *"calculate the osmotic pressure of 0.1 M NaCl"*, you must use $\Pi = iCRT$ with $i \approx 2$ — doubling the predicted Π. Forgetting i gives half the true answer and is one of the most common 1-mark deductions in board exams.

8 Master Reference Table (all colligative properties)

Property	Formula (with i)	Molar-mass formula	Best suited for
Relative lowering of p	$(p_A^\circ - p_A)/p_A^\circ = i\,x_B$	$M_B = i\,w_B M_A p_A^\circ/[w_A(p_A^\circ - p_A)]$	Low-volatility solutes
Elevation of BP	$\Delta T_b = i\,K_b\,m$	$M_B = 1000\,i\,K_b\,w_B/(\Delta T_b\,w_A)$	Thermostable solutes
Depression of FP	$\Delta T_f = i\,K_f\,m$	$M_B = 1000\,i\,K_f\,w_B/(\Delta T_f\,w_A)$	Cryoscopy; Rast's method
Osmotic pressure	$\Pi = i\,C\,RT$	$M_B = i\,w_B\,RT/(\Pi\,V)$	Macromolecules (proteins, polymers)

Solvent	T_b° (K)	K_b (K kg/mol)	T_f° (K)	K_f (K kg/mol)
Water	373.15	0.52	273.15	1.86
Benzene	353.3	2.53	278.6	5.12
Chloroform	334.4	3.63	209.6	4.79
Carbon tetrachloride	350.0	5.03	250.5	31.8
Cyclohexane	353.74	2.79	279.55	20.00
Acetic acid	391.1	2.93	290.0	3.90
Camphor	—	—	451.15	39.7
Ethanol	351.5	1.20	155.7	1.99

Remember: colligative $\propto$ *number*, not identity

"Same particles, same property." 1 mol of glucose, 0.5 mol of NaCl ($i = 2$), and 0.33 mol of K_2SO_4 ($i = 3$) all give roughly the *same* ΔT_f in 1 kg water — because each contributes 1 mol of *particles*. The chemical identity is irrelevant.