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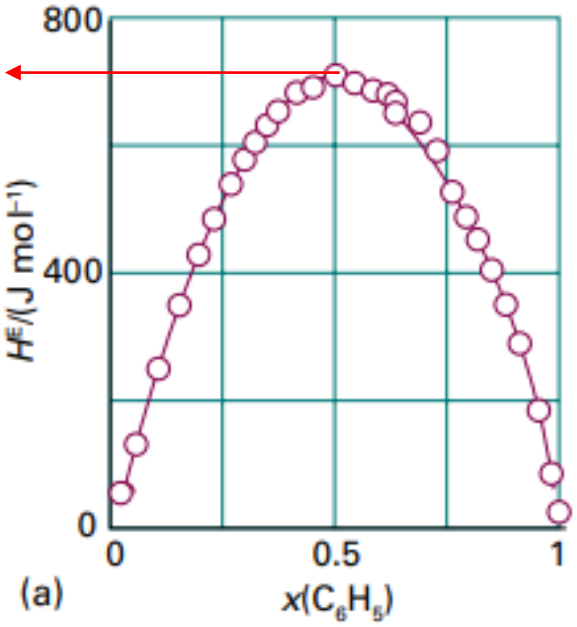

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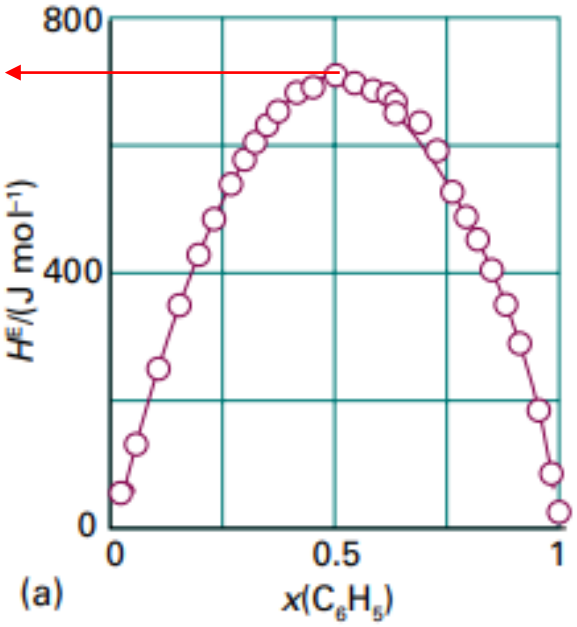

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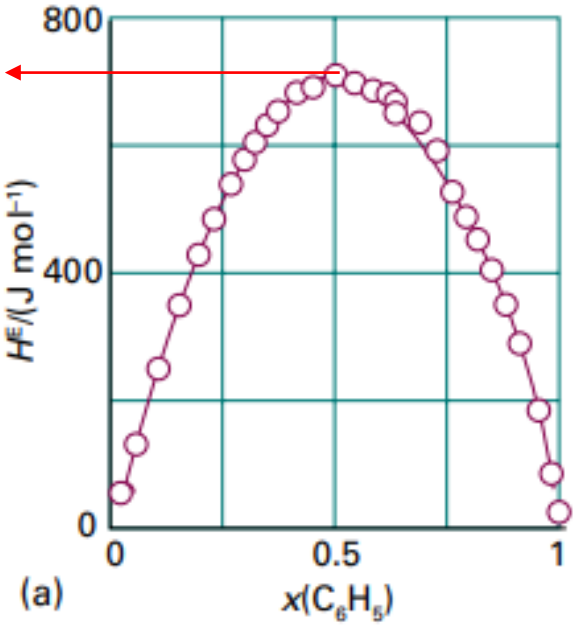


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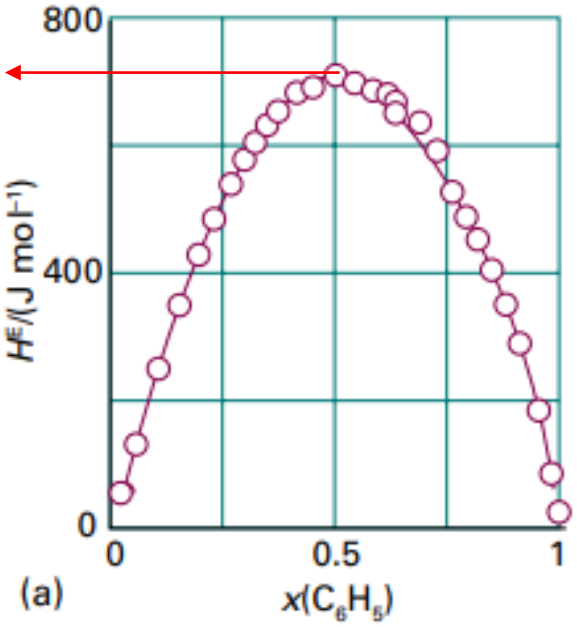


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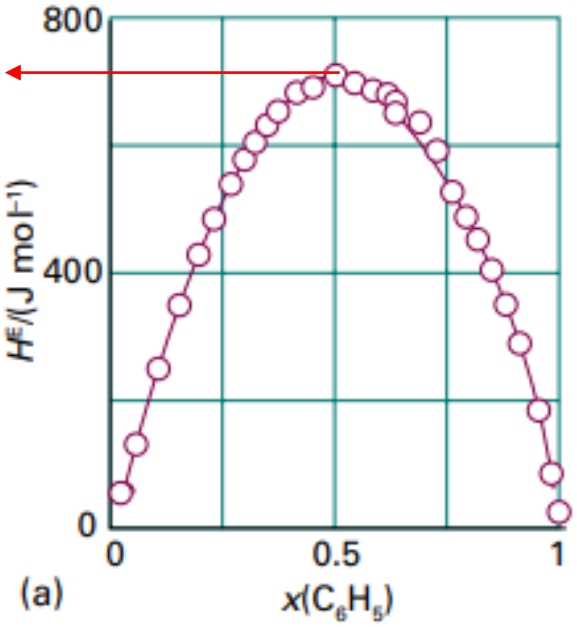


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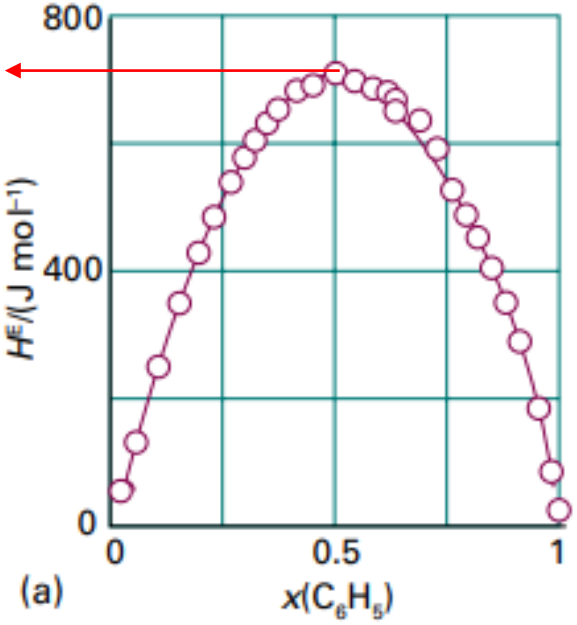


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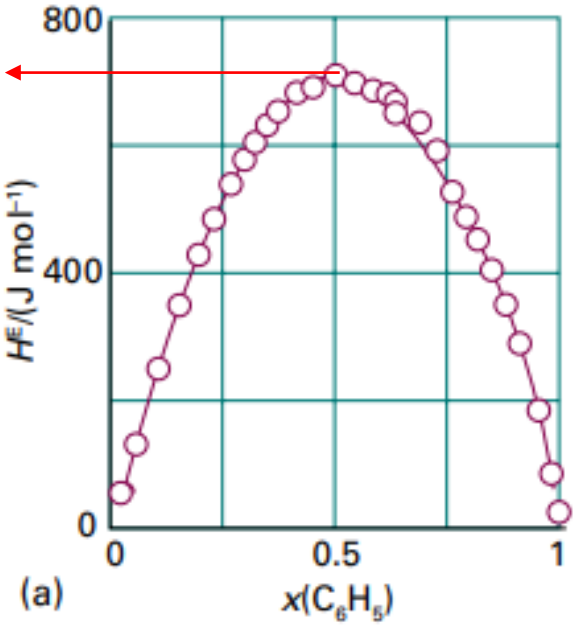


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$$= -1.02\,\mathrm{kJ\,mol^{-1}}$$

# Colligative properties

..depends on the relative number of solute particles present but not their chemical identity

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- vapor pressure lowering
- boiling point elevation
- freezing point depression
- osmotic pressure

# FP depression and BP elevation

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pure $\mu_A = \mu_A^*$

with solute B $\mu_A = \mu_A^* + RT \ln x_A$

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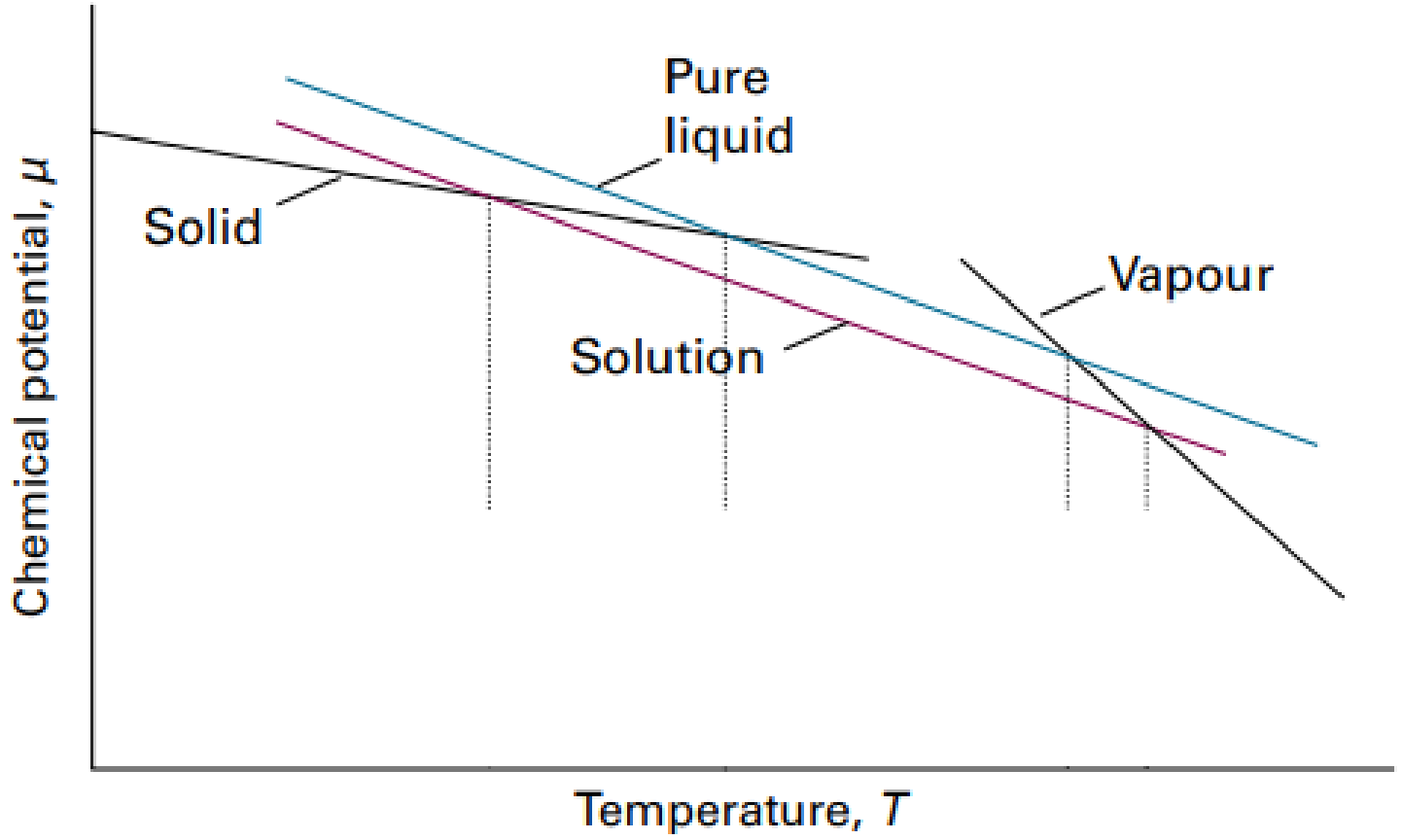


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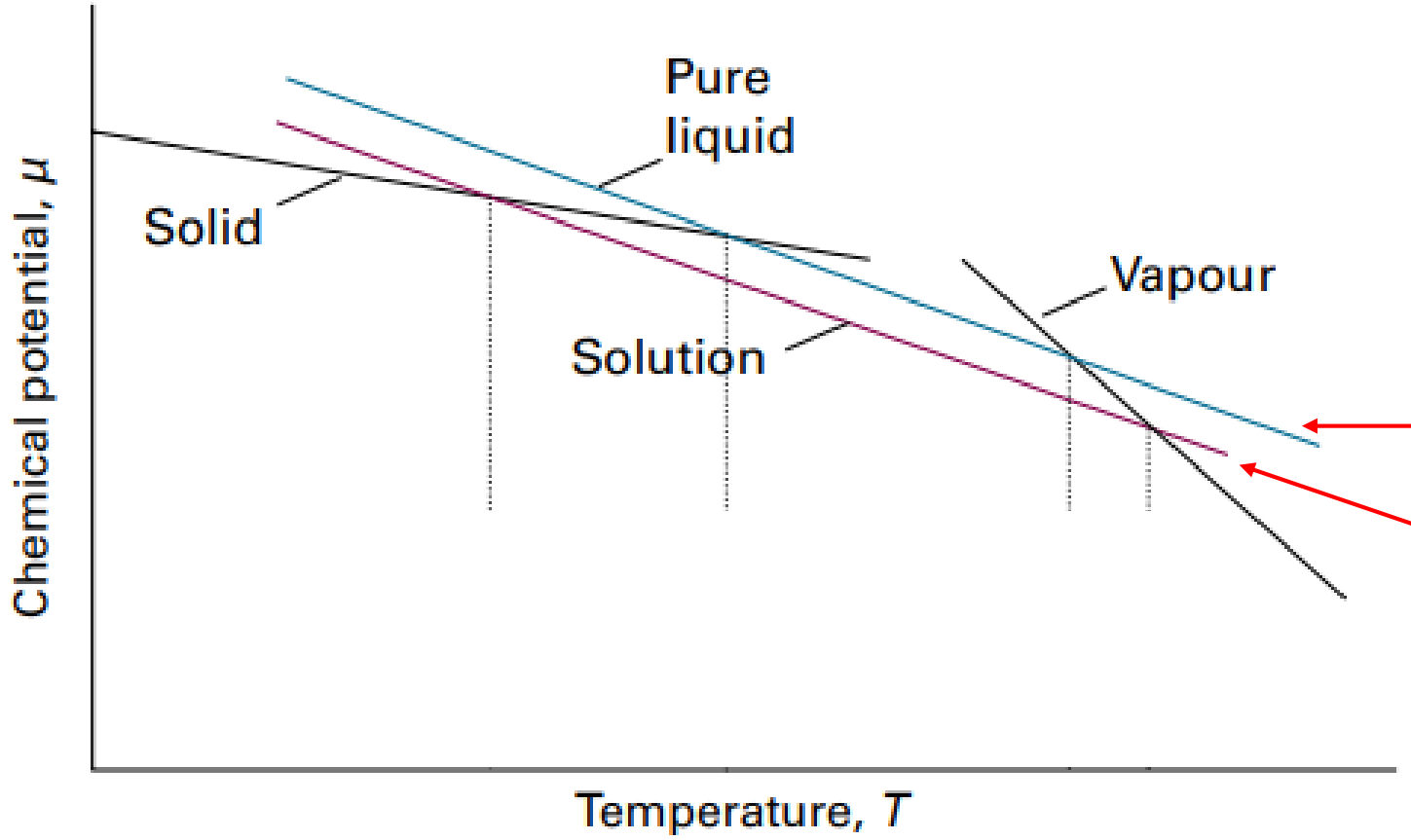


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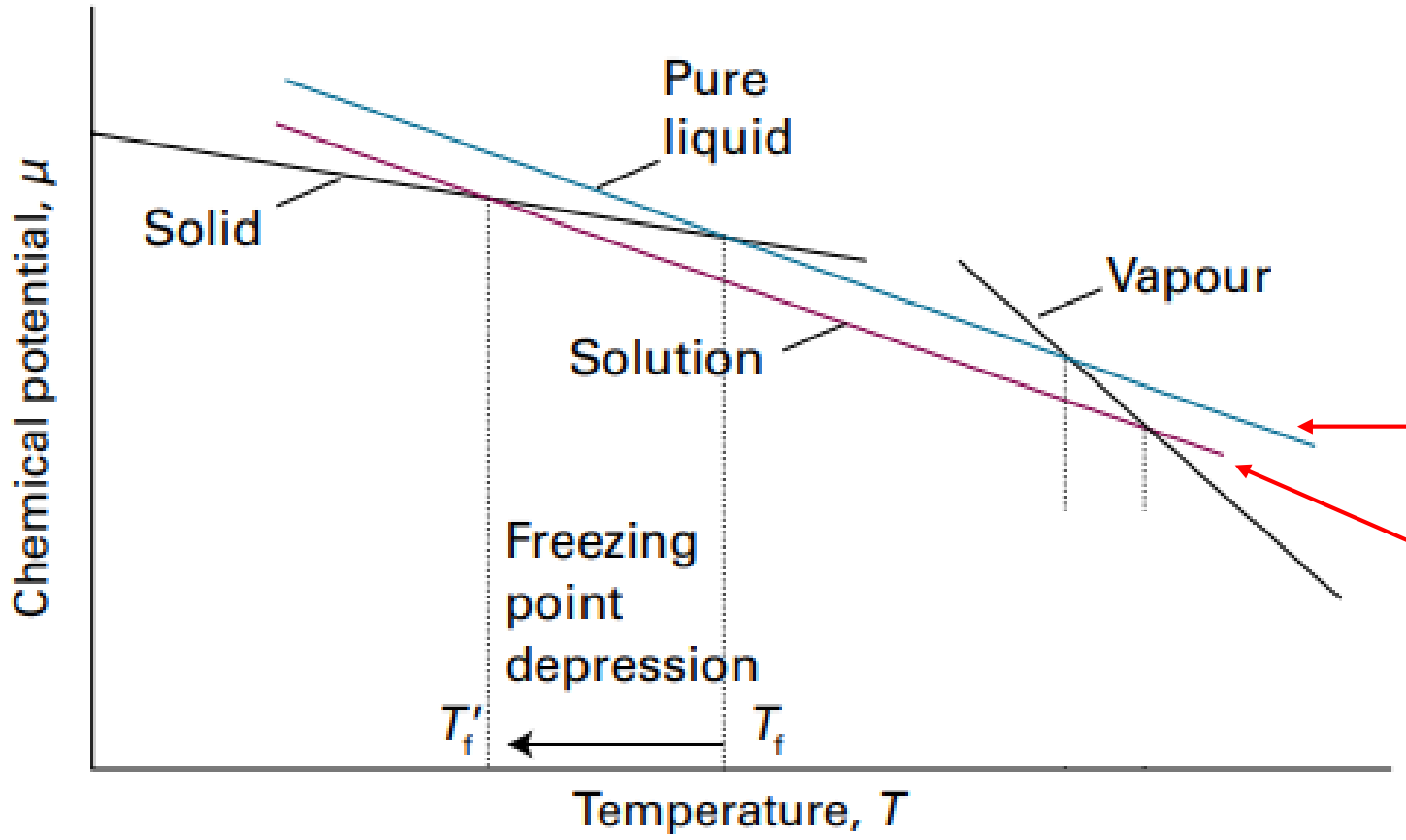


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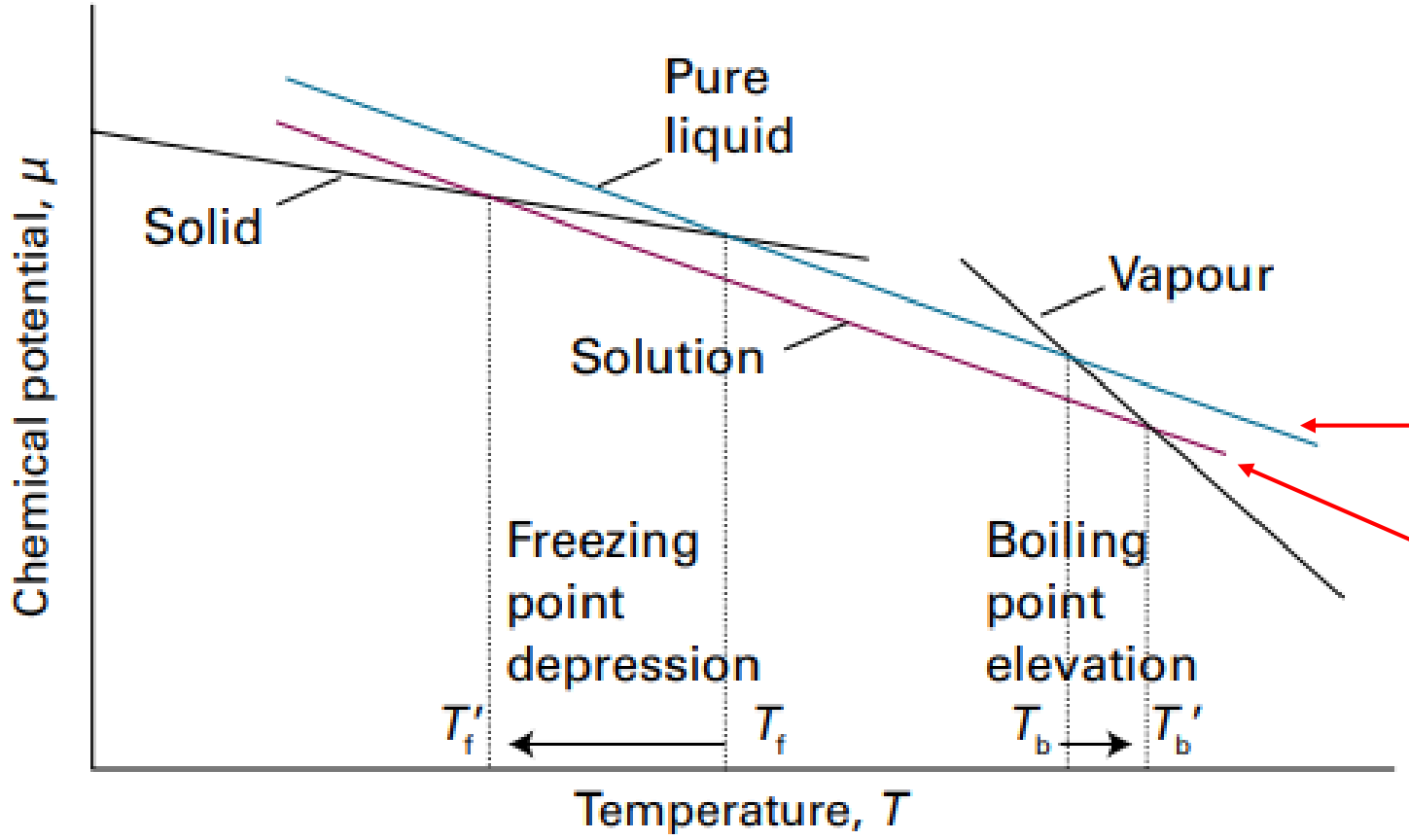


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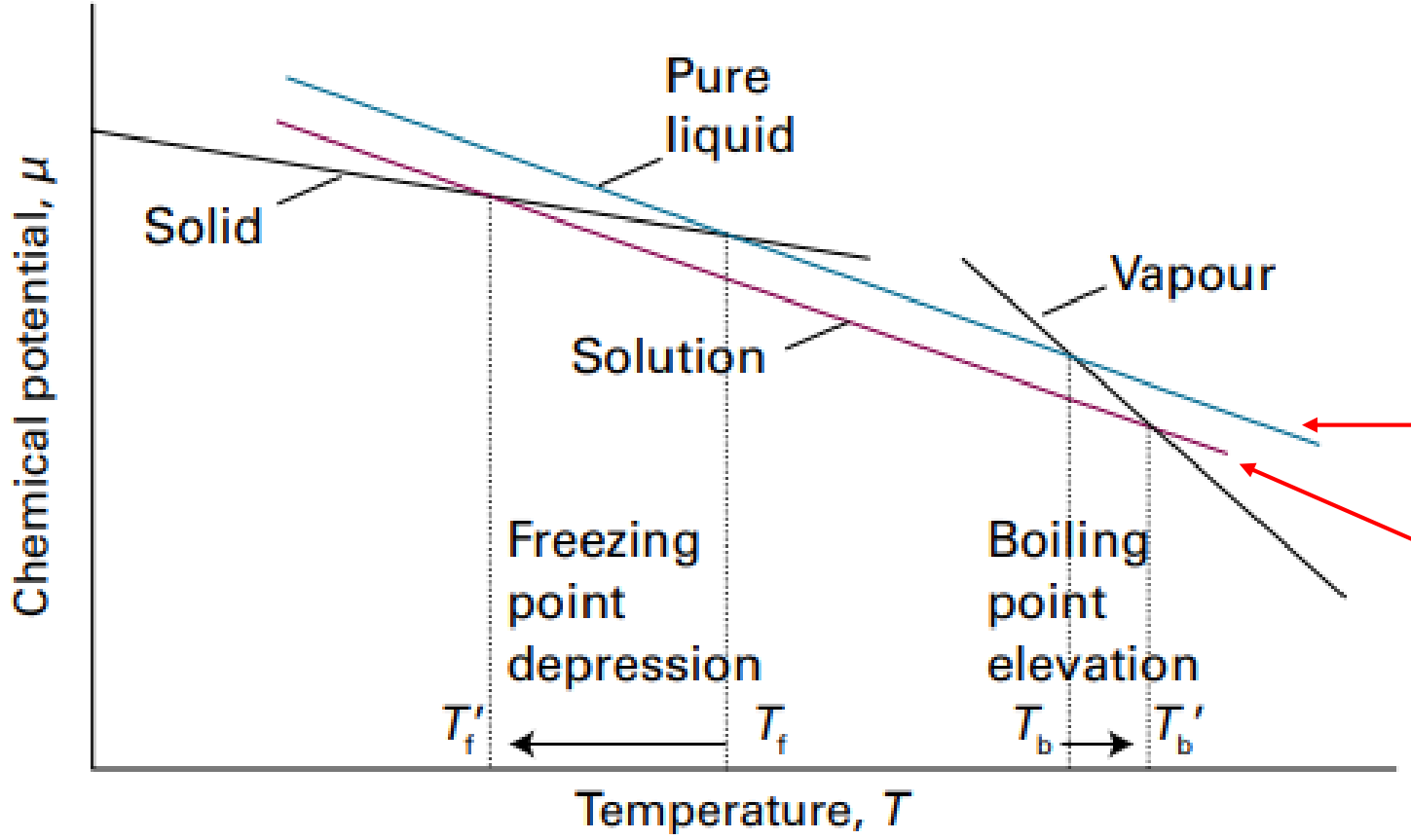
This occurs even in ideal solutions

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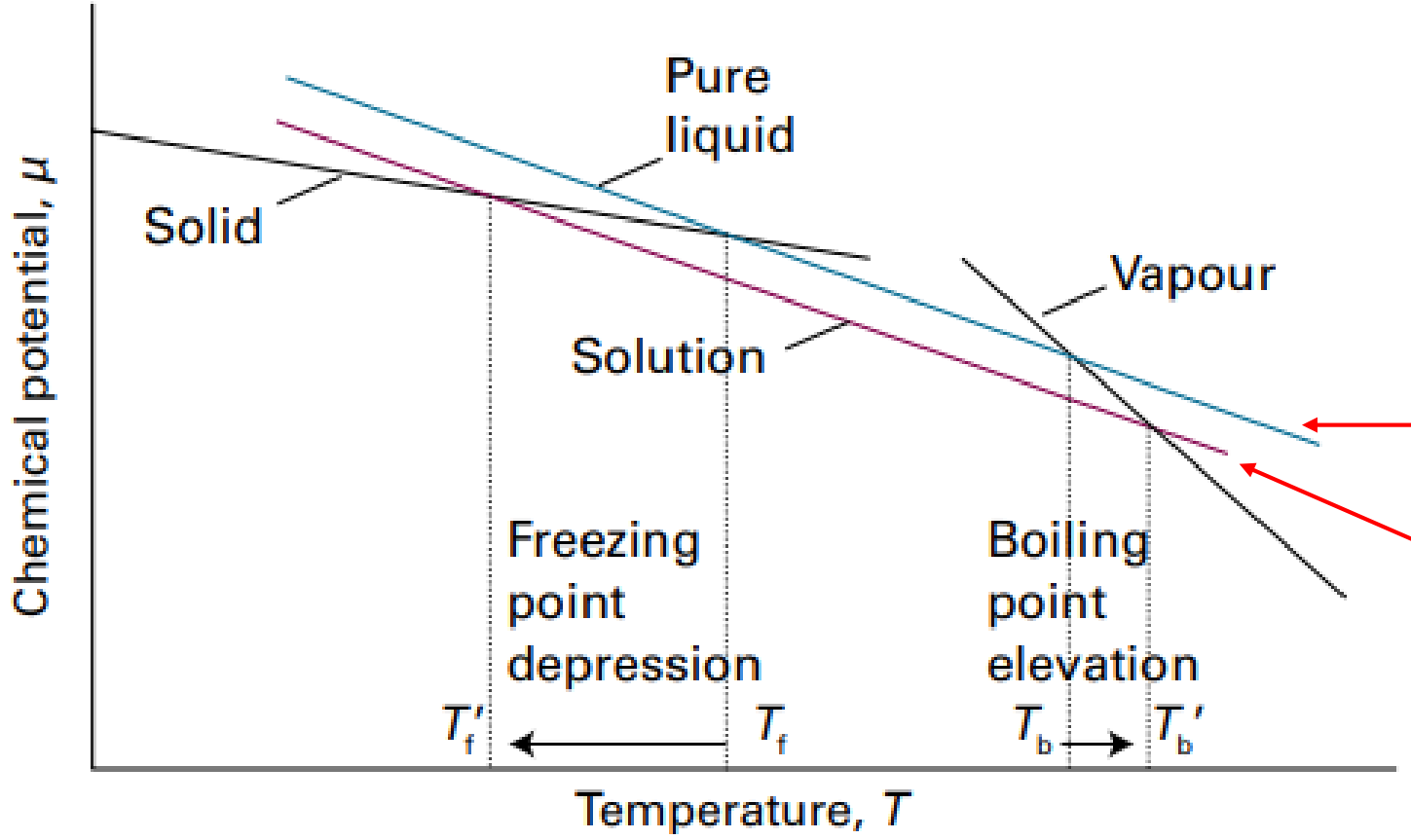


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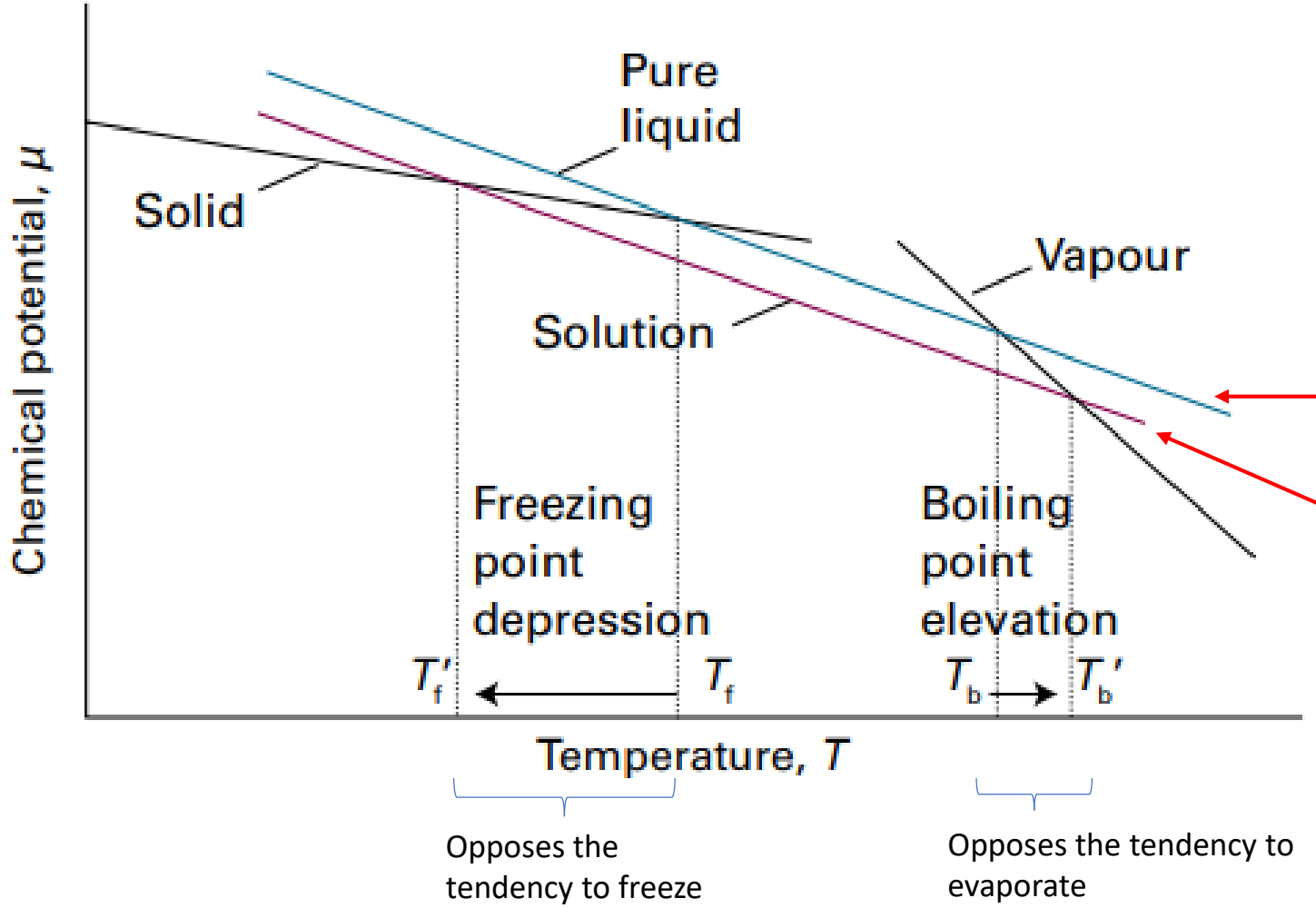


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Solute:  An enhanced molecular randomness in solution => opposes the tendency to freeze => requires lower temperature => lowers the FP

# Elevation of BP

The elevation in boiling point ($\Delta T$) is directly proportional to the molality of the solute ($m_B$)

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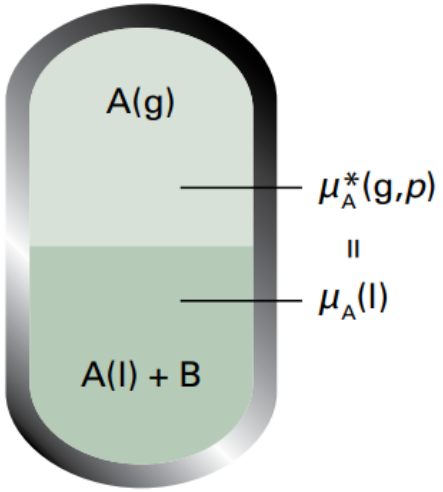


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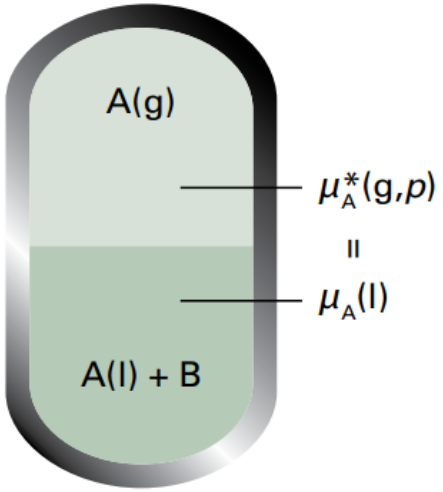


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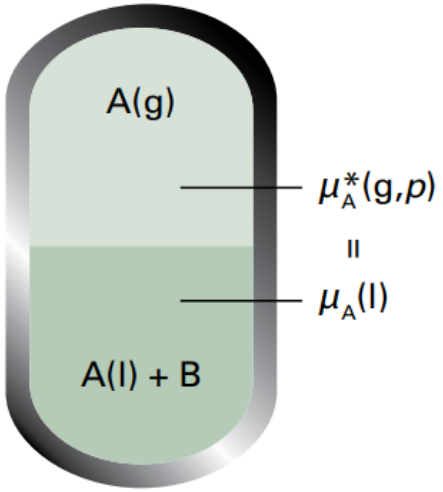

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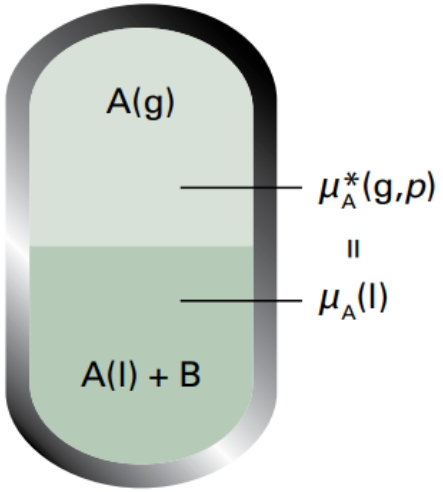

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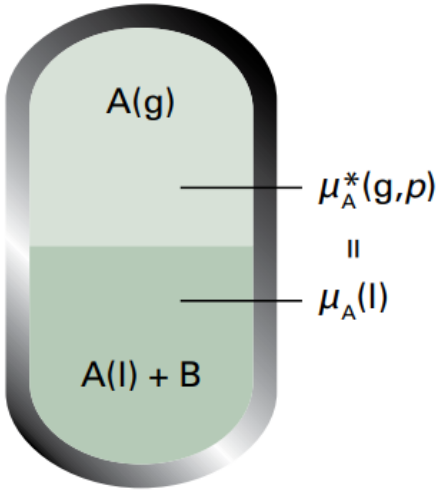


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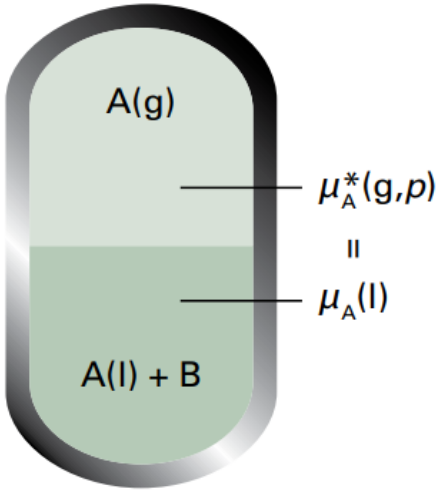


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Gibbs–Helmholtz equation

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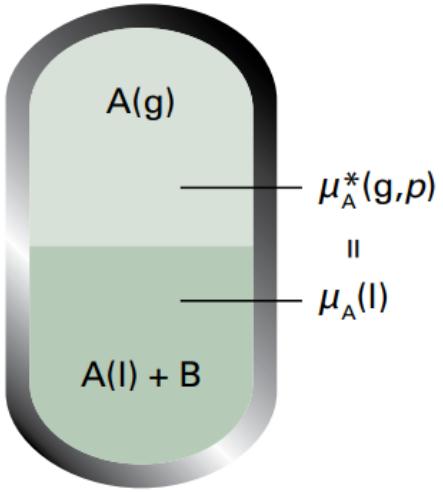


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Enthalpy of vaporization is assumed to be constant over the small range of temperature

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Suppose, $x_B << 1$; the approximation $\ln(1 - x) \approx -x$

The **Taylor series** for $\ln(1-x)$ around $x = 0$ is:

$$\ln(1-x) = -x - \frac{x^2}{2} - \frac{x^3}{3} - \frac{x^4}{4} - \cdots$$

So if $x$ is **small**, then:

- $x^2, x^3, \ldots$ are **even smaller**, and can be neglected.
- That means:

$$\ln(1-x) \approx -x$$

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$$\frac{1}{T^*}-\frac{1}{T}=\frac{T-T^*}{TT^*}\approx\frac{T-T^*}{T^{*2}}=\frac{\Delta T_b}{T^{*2}}$$

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$X_B << 1$, $X_B$ is proportional to its molality, b

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$$b=\frac{\text{mol}}{kg}$$

$b$ = molality

mol = moles of solute

kg = kilogram of solvent

When the solution is dilute, the mole fraction of solute is approximately proportional to molality:

$$x_B = \frac{n_B}{n_A + n_B}$$

$$\text{If } n_B \ll n_A, \text{ then } x_B \approx \frac{n_B}{n_A}$$

$$m = \frac{n_B}{M_A \cdot n_A}$$

$$\frac{n_B}{n_A} = m \cdot M_A$$

$$x_B \approx m \cdot M_A$$

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Suppose, $x_B << 1$; the approximation $\ln(1 - x) \approx -x$

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$$\Delta T_b=K x_B \quad K=\frac{RT^{*2}}{\Delta_{vap}H}$$

increase in the boiling point is small

$X_B << 1$, $X_B$ is proportional to its molality, b

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$$b=\frac{\text{mol}}{kg}$$

$b$ = molality

mol = moles of solute

kg = kilogram of solvent

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# Elevation of BP

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$K_b$: empirical **boiling-point constant**

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No reference to the identity of the solute; thus, a colligative property!

# Elevation of BP

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empirical **boiling-point constant**

The change in bp depends on the solvent.

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What is the elevation of the BP of

1) Water **0.051 K**
2) Benzene **0.25 K**

When a solute is present at a molality of 0.10 mol kg$^{-1}$

# Depression of Freezing Point

$$\mu_{\mathrm{A}}^{\star}(\mathrm{s}) = \mu_{\mathrm{A}}^{\star}(\mathrm{l}) + RT \ln x_{\mathrm{A}}$$

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$T^\star$ is the freezing point of the pure liquid

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$\Delta_{fus}H$ is the enthalpy of fusion of the solvent

When the solution is dilute,

$$\Delta T_f = K_f b$$

$K_f$ is the empirical freezing-point constant
b is the molality of the solute

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1) Water **0.186 K**
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BP

**0.051 K**

**0.25 K**

When a solute is present at a molality of 0.10 mol $kg^{-1}$

Freezing requires precise molecular ordering (forming a solid lattice), which is more easily disrupted by solute particles than the process of molecules escaping to vapor.