

BOILING POINT ELEVATION

- Normal Boiling Point (BP) of a liquid is the temperature at which the vapor pressure of the liquid equals 1.00 atm

I. Effect of a Dissolved Solute on the BP of a Solution

What happens if a nonvolatile solute such as ethylene glycol ($C_2H_6O_2$) is added to pure water ?

Vapor Pressure of Solution < Vapor Pressure of Pure Solvent (water)

$$P(A + B) < P_A$$

- It follows:
 1. The normal BP of solution (A+B) is the temperature at which $P_{(A+B)}$ is equal to 1.00 atm
 2. This requires a temperature higher than for the pure liquid (water)
 3. $BP(\text{solution}) > BP(\text{pure solvent}) \rightarrow$ **B.P. Elevation**

$$\Delta T_b = \text{Boiling Point Elevation}$$

$$\Delta T_b = BP(\text{solution}) - BP(\text{solvent})$$

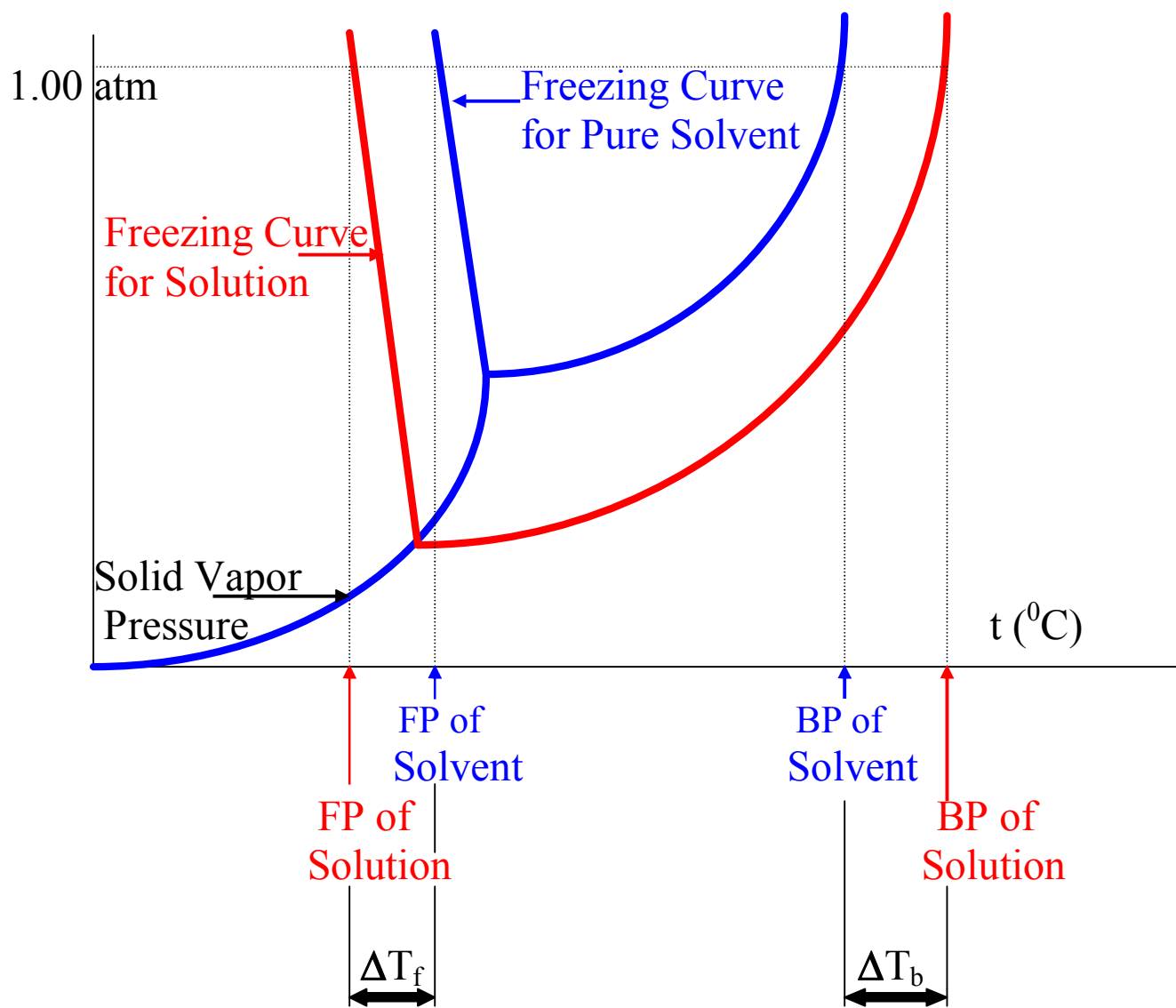
- Boiling point elevation is a Colligative Property
 - depends on the concentration of the Solution
 - does not depend on the nature of the Solute

$$\Delta T_b = K_b C_m$$

C_m = Concentration in molality

K_b = Boiling Point Elevation Constant (units of $^{\circ}C/m$)

**PHASE DIAGRAM FOR WATER AND AN
AQUEOUS SOLUTION OF ETHYLENE GLYCOL**



Examples:

1. At what temperature would a 5.00 molal solution of ethylene glycol boil? ($K_{b(\text{water})} = 0.512 \text{ }^\circ\text{C/m}$)

$$\Delta T_b = (0.512 \text{ }^\circ\text{C/m}) (5.00 \text{ m}) = 2.56 \text{ }^\circ\text{C}$$

$$\text{BP}_{(\text{solution})} = 100.00 \text{ }^\circ\text{C} + 2.56 \text{ }^\circ\text{C} = 102.56 \text{ }^\circ\text{C}$$

2. A solution was made of eugenol in diethyl ether (ether). If the solution was 0.575 m eugenol in ether, what was the Boiling Point of the solution ?

$$\text{BP}(\text{diethyl ether}) = 34.6 \text{ }^\circ\text{C} \qquad K_b(\text{ether}) = 2.02 \text{ }^\circ\text{C/m}$$

$$\Delta T_b = (2.02 \text{ }^\circ\text{C/m}) (0.575 \text{ m}) = 1.16 \text{ }^\circ\text{C}$$

$$\text{BP}(\text{solution}) = 34.6 \text{ }^\circ\text{C} + 1.16 \text{ }^\circ\text{C} = 35.8 \text{ }^\circ\text{C}$$

FREEZING POINT DEPRESSION

- From the phase diagram, the vapor pressure curve for the solid solution is unchanged.
- As the temperature of a solution is lowered, the pure solvent freezes out of solution. (Sea ice is almost pure water)
- FP of the solution shifts toward a lower temperature (**F. P. Depression**)

ΔT_f = Freezing Point Depression

$$\Delta T_f = \text{FP}(\text{solvent}) - \text{FP}(\text{solution})$$

- Freezing point depression is a Colligative Property
 - depends on the concentration of the Solution
 - does not depend on the nature of the Solute

$$\Delta T_f = K_f C_m$$

C_m = Concentration in molality

K_f = Freezing Point Depression Constant (units of °C/m)

Practical Applications

1. Use of ethylene glycol as antifreeze in car radiators:
 - lower the FP of coolant
 - elevates the BP of coolant (prevents it from boiling away)
2. NaCl poured over icy roads lowers the FP of ice below the temperature of the surrounding air. As a result, the salted ice melts.
3. To obtain Molecular Weights of compounds

Examples:

1. How many grams of ethylene glycol ($C_2H_6O_2$) must be added to 37.8 g of water to give a Freezing Point of $-0.150\text{ }^\circ\text{C}$? ($K_f = 1.858\text{ }^\circ\text{C/m}$)

$$\Delta T_f = \text{FP (solvent)} - \text{FP(solution)} = 0.000\text{ }^\circ\text{C} - (-0.150\text{ }^\circ\text{C}) = 0.150\text{ }^\circ\text{C}$$

$$\Delta T_f = K_f C_m \qquad C_m = \frac{\Delta T_f}{K_f} = \frac{0.150\text{ }^\circ\text{C}}{1.858\text{ }^\circ\text{C/m}} = 0.0807\text{ m}$$

$$? \text{ g EG} = 37.8 \text{ g water} \times \frac{0.0807 \text{ mol EG}}{1000 \text{ g water}} \times \frac{62.08 \text{ g}}{1 \text{ mol}} = 189 \text{ g}$$

2. An aqueous solution of a molecular compound freezes at $-0.086\text{ }^\circ\text{C}$. What is the molality of the solution? ($K_f = 1.858\text{ }^\circ\text{C/m}$)

$$\Delta T_f = 0.000\text{ }^\circ\text{C} - (-0.086\text{ }^\circ\text{C}) = 0.086\text{ }^\circ\text{C}$$

$$\Delta T_f = K_f \times m$$

$$m = \frac{\Delta T_f}{K_f} = \frac{0.086\text{ }^\circ\text{C}}{1.858\text{ }^\circ\text{C/m}} = 4.6 \times 10^{-2} \text{ m}$$