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. EffectofaDissolvedSoluteontheBPofaSolution

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(solution) > BP (pure solvent) → B.P.Elevation

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

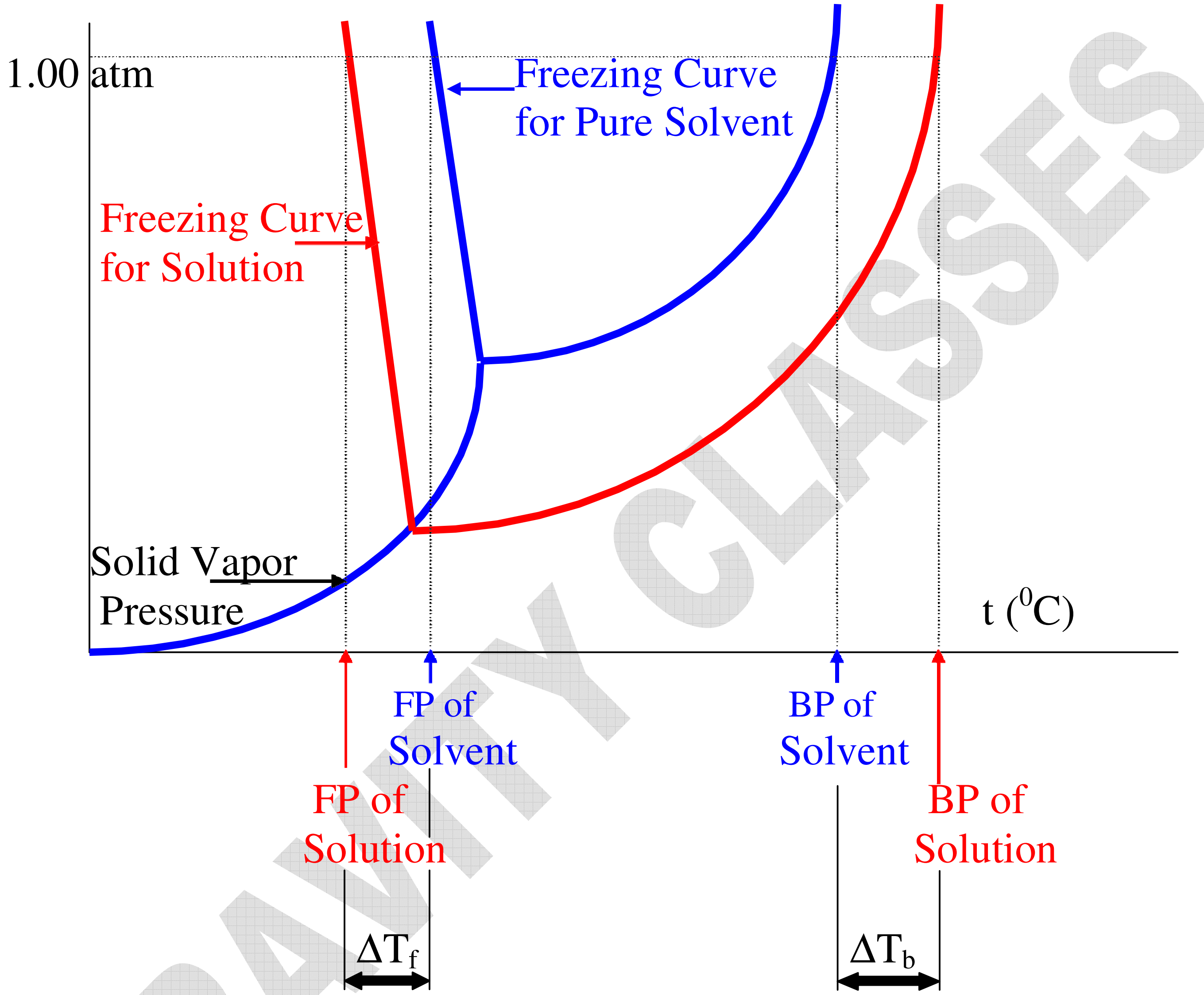
ΔT_b = BP (solution) – BP (solvent)

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

$$\boldsymbol{\Delta T_b = K_b C_m}$$

C_m = Concentration in molality
K_b = Boiling Point Elevation Constant (units of °C/m)

PHASE DIAGRAM FOR WATER AND AN AQUEOUS SOLUTION OF ETHYLENE GLYCOL



Examples:

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

$$\Delta T_b = (0.512\ ^oC/m)\ (5.00\ m) = 2.56\ ^oC$$

$$BP_{(solution)} = 100.00\ ^oC + 2.56\ ^oC = 102.56\ ^oC$$

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 ?

BP(diethyl ether) = 34.6 ^oC Kb(ether) = 2.02 $^oC/m$)

$$\Delta T_b = (2.02\ ^oC/m)\ (0.575\ m) = 1.16\ ^oC$$

$$BP(solution) = 34.6\ ^oC + 1.16\ ^oC = 35.8\ ^oC$$

FREEZOMG 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

ΔT_f = FP(solvent) – FP (solution)

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

$$\mathbf{\Delta T_f = K_f C_m}$$

C_m = Concentration in molality
K_f = Freezing Point Depression Constant (units of °C/m)

<u>PracticalApplications</u>

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 thetemperature 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 $-0.150\ ^oC$? ($K_f = 1.858\ °C/m$)

$$\Delta T_f = FP\ (solvent) - FP(solution) = 0.000\ ^oC - (-0.150\ ^oC) = 0.150\ ^oC$$

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

$$\mathbf{?\ g\ EG = 37.8\ \cancel{gwater} \times \frac{0.0807\ \cancel{mol}\ EG}{1000\ \cancel{gwater}} \times \frac{62.08\ g}{1\ \cancel{mol}} = 189\ g}$$

2. An aqueous solution of a molecular compound freezes at -0.086^0C. What is the molality of the solution? ($K_f = 1.858\ ^0C/m$)

$$\Delta T_f = 0.000^0C - (-0.086^0C) = 0.086^0C$$

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

$$m = \frac{\Delta T_f}{K_f} = \frac{0.086^0C}{1.858\ ^0C/m} = \mathbf{4.6 \times 10^{-2}\ m}$$