Molal Freezing Point Depression Constants

Calculating K_f Values

Introduction

The *freezing point* of a liquid is the temperature at which the forces of attraction among molecules are just enough to cause a phase change from the liquid state to the solid state. It is the temperature at which the liquid and solid phases exist in equilibrium. The solvent molecules in a solution are somewhat more separated from each other than they are in a pure solvent due to the interference of solute particles. Therefore, the temperature of a solution must be lowered below the freezing point of the pure solvent in order to freeze it. The freezing point depression (ΔT_f) of solutions of nonelectrolytes has been found to be equal to the molality (m) of the solute times a proportionality constant called the *molal freezing point depression constant* (K_f). Therefore,

$$\Delta T_f = K_f \times m, \qquad \textit{Equation 1}$$

where m is the number of moles of solute per kilograms of solvent.

Concepts

- Freezing point depression
- Colligative properties

Discussion

The determination of the molecular mass of a substance from the freezing point depression is included in many high school and most college laboratory manuals. In a typical experiment, the freezing point of the pure solvent is measured followed by the freezing point of a solution of known molality. Dividing the freezing point depression (ΔT_f) by the solution molality (m) gives the molal freezing point constant (K_f) of the solvent, according to Equation 1. The second part of the lab consists of using the K_f value determined in part one to determine the molality, and ultimately the molecular mass of an unknown substance. Since many schools have limited laboratory time, teachers often eliminate the first part (determination of the molal freezing point constant) entirely, and limit their solvent choice to those with constants given in various handbooks. The goal of this handout is to allow teachers to calculate the K_f value for any solvent; hence they may use the solvent of choice and still save lab time.

The molal freezing depression constant of a solvent is related to the latent heat of fusion for that substance. It may be calculated from a derivative of the Clausius-Clapeyron equation shown below:

$$K_f = \frac{R \cdot T_{fp}^2}{1000 \cdot \Delta H_{fusion}} \qquad \textit{Equation 2}$$

where, K_f = molal freezing point depression constant

R = ideal gas constant, 8.31 J/mol·K or 1.987 cal/mol·K

T_{fp} = freezing point of the pure solvent in Kelvin, where K = °C + 273.2

ΔH_{fusion} = the heat of fusion of the solvent in Joules/gram or calories/gram

For example, water has a latent heat of fusion of 79.71 cal/gram and freezes at 0 °C (273.2 K). By substituting these values into Equation 2, the freezing point depression constant becomes 1.86 °C/m, as shown below.

$$K_f = \frac{1.987 \dfrac{\text{cal}}{\text{mol·K}} \times (273.2\ \text{K})^2}{1000 \dfrac{\text{grams}}{\text{kg}} \times 79.71 \dfrac{\text{cal}}{\text{gram}}} = 1.86\ °\text{C/molal}$$

While this method gives the theoretical freezing point depression constant for nonvolatile nonelectrolytes in pure solvents, the actual values may be somewhat different when working with impure samples in normal high school laboratories. These calculated values are most consistent with the experimental values in dilute solutions where both the solvent and solute have similar chemical and physical properties. The molal freezing point depression constant for benzene varies from 5.09 °C/m

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for a 0.118 molal solution containing carbon tetrachloride to 4.82 °C/*m* for a 1.17 molal solution (Ref. 3).

Some other useful solvents and their freezing point constants are listed below.

Molal Freezing Point Depression Constants for Some Common Cryoscopic Solvents (Ref. 2, 4)

Compound	Formula	Freezing Point, °C	Heat of Fusion cal/gram	Molal Freezing Point Constant (literature)	Molal Freezing Point Constant (calculated)
t-Butanol	C_4H_9OH	25.82 °C	21.88 cal/g	8.37 °C/*m*	8.12 °C/*m*
Camphor	$C_{10}H_{16}O$	178.75 °C	10.74 cal/g	37.7 °C/*m*	37.7 °C/*m*
Cetyl Alcohol	$C_{16}H_{33}OH$	49.27 °C	39.18 cal/g	- - -	5.27 °C/*m*
Cyclohexane	C_6H_{12}	6.54 °C	7.47 cal/g	20.0 °C/*m*	20.8 °C/*m*
Lauric Acid	$C_{11}H_{23}COOH$	43.22 °C	43.72 cal/g	- - -	4.53 °C/*m*
Naphthalene	$C_{10}H_8$	80.29 °C	35.39 cal/g	6.94 °C/*m*	7.01 °C/*m*
Stearic Acid	$C_{17}H_{35}COOH$	68.82 °C	27.39 cal/g	4.5 °C/*m*	4.27 °C/*m*
Water	H_2O	0 °C	79.71 cal/g	1.86 °C/*m*	1.86 °C/*m*

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References

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Dean, John, A. *Lange's Handbook of Chemistry*, 13th ed.; McGraw-Hill: New York, 1985; pp 10–80.

Glasstone, Samuel. *Textbook of Physical Chemistry*, 2nd ed.; D.Van Nostrand: New York, 1946; p 645.

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Materials for *Molal Freezing Point Depression Constants* are available from Flinn Scientific, Inc.

Catalog No.	Description
B0179	t-Butyl Alcohol, 100 mL
C0354	Camphor, 25 g
C0200	Cetyl Alcohol, 100 g
C0113	Cyclohexane, 500 g
L0052	Lauric Acid, 100 g
N0065	Naphthalene, 250 g
S0335	Stearic Acid, 100 g

Consult your *Flinn Scientific Catalog/Reference Manual* for current prices.