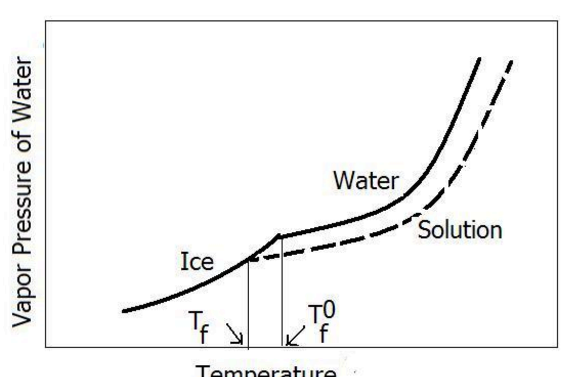
Chemistry of Changing Seasons: Melting Snow and Ice

In the wintertime, during the snowfall and/or after the snowfall, we have witnessed that roads are covered with salt (such NaCl or $CaCl_2$) and sand mixture to melt the snow and the ice to make the roads clear and passable. The mixture of salt and sand has dual effects: the salt melts the snow and the ice, and the sand increases the coefficient of friction between the road and the tires on cars that in essence reduces slippery conditions meaning improves tire traction. However, after the roads are dry, sand creates hazardous conditions for motorists, especially for motorcyclists and bicyclist. The salt works by lowering the melting or freezing point of water [In reality, the melting and freezing points are basically the same. For example, the ice melts at 32^0 F (0^0C) or water freezes at 32^0C (0^0F)] depending upon how we say it. We say water freezes (going from liquid state to solid sate) and ice melts (going from solid state to liquid state]. When salt is added, the freezing point of water decreases that is known as '*freezing point depression',* which of course depends on the amount of solute (salt) added.

How the *freezing point depression* works? When the salt is added to water, the salt dissolves in water because both salt and water are polar in nature. This process introduces foreign dissolved particles in water that lowers the freezing point of water. In general, solution freezes at lower temperature than pure solvent of the same solution. The lowering of freezing point continues until the added salt stops dissolving. This temperature is -6 ^{0}F (-21^0 C) for table salt (NaCl) in water under controlled laboratory conditions. However, in the real world, on roads and sidewalks, NaCl can melt ice down to 15^0 F (-9^0 C) (See the table at the end for more chemicals).

Theory Behind Freezing Point Depression

Freezing point depression is a colligative property, which depends on the number of dissolved particles in a solvent. It is s direct result of the lowering of the solvent vapor pressure by the solute. All liquid solvents when solutes are dissolved exhibit this property. Other colligative properties are vapor pressure lowering, boiling point elevation, and osmotic pressure. The following phase diagram illustrates the concept of freezing point depression of aqueous solution:



This diagram indicates that lowering the vapor pressure of the solution lowers the freezing point of solution. Let T_f^0 be the freezing point of the pure solvent and T_f be the freezing point of the solution. Then, the freezing point depression, ΔT_f, is expressed as

$$\Delta T_f = T_f^0 - T_f$$

Further, the ΔT_f is proportional to molal concentration (m) of the solution [molality (m) is the number of moles of solute dissolved in 1 kg (1000 g) of solvent] and hence

$$\Delta T_f \propto m$$

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

where K_f is the *molal freezing-point depression constant* that varies from solvent to solvent. Following table lists some of the common liquids.

Solvent	**Normal Freezing Point (^{0}C)#**	**K_f (^{0}C/m)**
Water	0	1.86
Benzene	5.5	5.12
Cyclohexane	6.6	20.0
Acetic acid	16.6	3.90
Ethanol (Ethyl Alcohol)	-117.3	1.99

Measured at 1 atm.

Note that the same principle, depression in freezing point, also applies in using the antifreeze in cars and trucks in the wintertime and adding salt to make ice cream.

Qualitatively the freezing point depression can be explained in terms of order-disorder state. In general, freezing involves a transition from the disordered state (solvent) to the ordered state (solid) (for e.g. water → ice). For this to take place, the energy must be expelled from the system:

Disordered state → Ordered state + energy

A solution has greater disorder than the solvent. Hence, more energy needs to be expelled from it to create an ordered state compared to pure solvent. This is the reason why the solution has a lower freezing point than the solvent.

Chemicals Used to Melt Snow and Ice

There are number of chemicals in usage to melt the ice and Snow in the wintertime. Following is the information taken from About.com

Name	**Chemical Formula**	**Lowest Practical Temperature**	**Pros**	**Cons**
Sodium Chloride (Rock Salt & Halite)	$NaCl$	15 ^{0}F (-9 ^{0}C)	Keeps sidewalks dry	Corrosive, damages concrete & vegetation
Potassium Chloride	KCl	20 ^{0}F (-7 ^{0}C)	Fertilizer	Damages concrete
Calcium Chloride	$CaCl_2$	-20 ^{0}F(-29 ^{0}C)	Melts ice faster than NaCl	Attracts moisture, surfaces become slippery below 0 ^{0}F (-18 ^{0}C)
Magnesium Chloride	$MgCl_2$	5 ^{0}F (-15 ^{0}C)	Melts ice faster than NaCl	Attracts moisture
Potassium Acetate	CH_3COOK	15 ^{0}F (-9 ^{0}C)	Biodegradable	Corrosive
Calcium Carbonate, Magnesium Carbonate, & Acetic Acid (CMA)	$CaCO_3$, $MgCO_3$, & $HC_2H_3O_2$	15 ^{0}F (-9 ^{0}C)	Safest for concrete & vegetation	Works better to prevent re-icing than as ice remover
Ammonium Sulfate	$(NH_4)_2SO_4$	20 ^{0}F (-7 ^{0}C)	Fertilizer	Damages concrete
Urea	NH_2-CO-NH_2	20 ^{0}F (-7 ^{0}C)	Fertilizer	Agricultural grade is corrosive

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