Melting point and freezing point of water

Learning Objectives Explain what the term "colligative" means, and list the colligative properties. Indicate what happens to the boiling point and the freezing point of a solvent when a solute is added to it. Calculate boiling point elevations and freezing point depressions for a solution. People who live in colder climates have seen trucks put salt on the roads when snow or ice is forecast. Why is this done? As a result of the information you explore in this section, you will understand why these events occur. You will also learn to calculate exactly how much of an effect a specific solute can have on the boiling point or freezing point of a solution.The example given in the introduction is an example of a colligative property. Colligative properties are properties that differ based on the concentration of solute in a solvent, but not on the type of solute. What this means for the example above is that people in colder climates do not necessarily need salt to get the same effect on the roads—any solute will work. However, the higher the concentration of solute, the more these properties will change. Water boils at 100^oC at $1 \: \text{atm}$ of pressure, but a solution of saltwater does not . When table salt is added to water, the resulting solution has a higher boiling point than the water did by itself. The ions form an attraction with the solvent particles that prevents the water molecules from going into the gas phase. Therefore, the saltwater solution will not boil at 100^oC. In order for the saltwater solution to boil, the temperature must be raised about 100^oC. This is true for any solute added to a solvent; the boiling point will be higher than the boiling point of the pure solvent (without the solute). In other words, when anything is dissolved in water, the solution will boil at a higher temperature than pure water would. The boiling point elevation due to the presence of a solute is also a colligative property. That is, the amount of change in the boiling point is related to the number of particles of solute in a solution and is not related to the chemical composition of the solute. A $0.20 \: \text{m}$ solution of table salt and a $0.20 \: \text{m}$ solution of hydrochloric acid would have the same effect on the boiling point. The effect of adding a solute to a solvent has the opposite effect on the freezing point of a solution as it does on the boiling point. A solution will have a lower freezing point than a pure solvent. The freezing point is the temperature at which the liquid changes to a solid. At a given temperature, if a substance is added to a solvent (such as water), the solute-solvent interactions prevent the solvent from going into the solid phase. The solute-solvent interactions require the temperature to decrease further in order to solidify the solution. A common example is found when salt is used on icy roadways. Salt is put on roads so that the water on the roads will not freeze at the normal 0^oC but at a lower temperature, as low as -9^oC. The de-icing of planes is another common example of freezing point depression in action. A number of solutions are used, but commonly a solution such as ethylene glycol, or a less toxic monopropylene glycol, is used to de-ice an aircraft. The aircrafts are sprayed with the solution when the temperature is predicted to drop below the freezing point. The freezing point depression is the difference in the freezing points of the solution from the pure solvent. This is true for any solute added to a solvent; the freezing point of the solution will be lower than the freezing point of the pure solvent (without the solute). Thus, when anything is dissolved in water, the solution will freeze at a lower temperature than pure water would. The freezing point depression due to the presence of a solute is also a colligative property. That is, the amount of change in the freezing point is related to the number of particles of solute in a solution and is not related to the chemical composition of the solute. A $0.20 \: \text{m}$ solution of table salt and a $0.20 \: \text{m}$ solution of hydrochloric acid would have the same effect on the freezing point. Figure $\PageIndex{1}$: Comparison of boiling and freezing points of a pure liquid (right side) with a solution (left side). Recall that covalent and ionic compounds do not dissolve in the same way. Ionic compounds break up into cations and anions when they dissolve. Covalent compounds typically do not break up. For example a sugar/water solution stays as sugar and water, with the sugar molecules staying as molecules. Remember that colligative properties are due to the number of solute particles in the solution. Adding 10 molecules of sugar to a solvent will produce 10 solute particles in the solution. When the solute is ionic, such as $\ce{NaCl}$ however, adding 10 formulas of solute to the solution will produce 20 ions (solute particles) in the solution. Therefore, adding enough $\ce{NaCl}$ solute to a solvent to produce a $0.20 \: \text{m}$ solution will have twice the effect of adding enough sugar to a solvent to produce a $0.20 \: \text{m}$ solution. Colligative properties depend on the number of solute particles in the solution. "i" is the number of particles that the solute will dissociate into upon mixing with the solvent. For example, sodium chloride, $\ce{NaCl}$, will dissociate into two ions so for $\ce{NaCl}$, $i = 2$; for lithium nitrate, $\ce{LiNO_3}$, $i = 2$; and for calcium chloride, $\ce{CaCl_2}$, $i = 3$. For covalent compounds, i is always equal to 1. By knowing the molality of a solution and the number of particles a compound will dissolve to form, it is possible to predict which solution in a group will have the lowest freezing point. To compare the boiling or freezing points of solutions, follow these general steps: Label each solute as ionic or covalent. If the solute is ionic, determine the number of ions in the formula. Be careful to look for polyatomic ions. Multiply the original molality (m) of the solution by the number of particles formed when the solution dissolves. This will give you the total concentration of particles dissolved. Compare these values. The higher total concentration will result in a higher boiling point and a lower freezing point. Example $\PageIndex{1}$ Rank the following solutions in water in order of increasing (lowest to highest) freezing point: $0.1 \: \text{m} \: \ce{NaCl}$ $0.1 \: \text{m} \: \ce{C_6H_{12}O_6}$ $0.1 \: \text{m} \: \ce{CaI_2}$ Solution To compare freezing points, we need to know the total concentration of all particles when the solute has been dissolved. $0.1 \: \text{m} \: \ce{NaCl}$: This compound is ionic (metal with nonmetal), and will dissolve into 2 parts. The total final concentration is: $\left(0.1 \: \text{m} \right) \left(2 \right) = 0.2 \: \text{m}$ $0.1 \: \text{m} \: \ce{C_6H_{12}O_6}$: This compound is covalent (nonmetal with nonmetal), and will stay as 1 part. The total final concentration is: $\left(0.1 \: \text{m} \right) \left(1 \right) = 0.1 \: \text{m}$ $0.1 \: \text{m} \: \ce{CaI_2}$: This compound is ionic (metal with nonmetal), and will dissolve into 3 parts. The total final concentration is: $\left(0.1 \: \text{m} \right) \left(3 \right) = 0.3 \: \text{m}$ Remember, the greater the concentration of particles, the lower the freezing point will be. $0.1 \: \text{m} \: \ce{CaI_2}$ will have the lowest freezing point, followed by $0.1 \: \text{m} \: \ce{NaCl}$, and the highest of the three solutions will be $0.1 \: \text{m} \: \ce{C_6H_{12}O_6}$, but all three of them will have a lower freezing point than pure water. The boiling point of a solution is higher than the boiling point of a pure solvent, and the freezing point of a solution is lower than the freezing point of a pure solvent. However, the amount to which the boiling point increases or the freezing point decreases depends on the amount of solute that is added to the solvent. A mathematical equation is used to calculate the boiling point elevation or the freezing point depression. The boiling point elevation is the amount that the boiling point temperature increases compared to the original solvent. For example, the boiling point of pure water at $1.0 \: \text{atm}$ is 100^oC while the boiling point of a 2% saltwater solution is about 102^oC. Therefore, the boiling point elevation would be 2^oC. The freezing point depression is the amount that the freezing temperature decreases. Both the boiling point elevation and the freezing point depression are related to the molality of the solution. Looking at the formula for the boiling point elevation and freezing point depression, we see similarities between the two. The equation used to calculate the increase in the boiling point is: $$\Delta T_b = k_b \cdot \text{m} \cdot i \label{BP}$$ Where: $\Delta T_b =$ the amount the boiling point increases. $k_b =$ the boiling point elevation constant which depends on the solvent (for water, this number is $0.515^\text{o} \text{C/m}$). $\text{m} =$ the molality of the solution. $i =$ the number of particles formed when that compound dissolves (for covalent compounds, this number is always 1). The following equation is used to calculate the decrease in the freezing point: $$\Delta T_f = k_f \cdot \text{m} \cdot i \label{FP}$$ Where: $\Delta T_f =$ the amount the freezing temperature decreases. $k_f =$ the freezing point depression constant which depends on the solvent (for water, this number is 1.86^oC/m). $\text{m} =$ the molality of the solution. $i =$ the number of particles formed when that compound dissolves (for covalent compounds, this number is always 1). Example $\PageIndex{2}$: Adding Antifreeze to Protein Engines Antifreeze is used in automobile radiators to keep the coolant from freezing. In geographical areas where winter temperatures go below the freezing point of water, using pure water as the coolant could allow the water to freeze. Since water expands when it freezes, freezing coolant could crack engine blocks, radiators, and coolant lines. The main component in antifreeze is ethylene glycol, $\ce{C_2H_4(OH)_2}$. What is the concentration of ethylene glycol in a solution of water, in molality, if the freezing point dropped by 2.64^oC? The freezing point constant, k_f, for water is 1.86^oC/m. Solution Use the equation for freezing point depression of solution (Equation $\ref{FP}$): $$\Delta T_f = k_f \cdot \text{m} \cdot i$$ Substituting in the appropriate values we get: $$2.64^\text{o} \text{C} = \left(1.86^\text{o} \text{C/m} \right) \left(\text{m} \right) \left(1 \right)$$ Solve for m by dividing both sides by 1.86^oC/m. $$\text{m} = 1.42$$ Example $\PageIndex{3}$: Adding Salt to Elevate Boiling Temperature A solution of $10.0 \: \text{g}$ of sodium chloride is added to $100.0 \: \text{g}$ of water in an attempt to elevate the boiling point. What is the boiling point of the solution? k_b for water is 0.52^oC/m. Solution Use the equation for boiling point elevation of solution (Equation $\ref{BP}$): $$\Delta T_b = k_b \cdot \text{m} \cdot i$$ We need to be able to substitute each variable into this equation. $k_b = 0.52^\text{o} \text{C/m}$ m: We must solve for this using stoichiometry. Given: $10.0 \: \text{g} \: \ce{NaCl}$ and $100.0 \: \text{g} \: \ce{H_2O}$ Find: $\text{mol} \: \ce{NaCl}/\text{kg} \: \ce{H_2O}$. Ratios: molar mass of $\ce{NaCl}$, $1000 \: \text{g} = 1 \: \text{kg}$ $$\dfrac{10.0 \: \cancel{\text{g} \: \ce{NaCl}}}{100.0 \: \cancel{\text{g} \: \ce{H_2O}}} \cdot \dfrac{1 \: \text{mol} \: \ce{NaCl}}{58.45 \: \cancel{\text{g} \: \ce{NaCl}}} \cdot \dfrac{1000 \: \cancel{\text{g} \: \ce{H_2O}}}{1 \: \text{kg} \: \ce{H_2O}} = 1.71 \: \text{m}$$ For $\ce{NaCl}$, $i = 2$ Substitute these values into the equation $\Delta T_b = k_b \cdot \text{m} \cdot i$. We get: $$\Delta T_b = \left(0.52 \dfrac{^\text{o} \text{C}}{\cancel{\text{m}}} \right) \left(1.71 \: \cancel{\text{m}} \right) \left(2 \right) = 1.78^\text{o} \text{C}$$ Water normally boils at 100^oC, but our calculation shows that the boiling point increased by 1.78^oC. Our new boiling point is 101.78^oC. Note: Since sea water contains roughly 28.0 g of NaCl per liter, this saltwater solution is approximately four times more concentrated than sea water (all for a 2° C rise of boiling temperature). This page was constructed from content via the following contributor(s) and edited (topically or extensively) by the LibreTexts development team to meet platform style, presentation, and quality: CK-12 Foundation by Sharon Bewick, Richard Parsons, Therese Forsythe, Shonna Robinson, and Jean Dupon. Melting Point, Freezing Point, Boiling Point Melting Point and Freezing Point Boiling Point Melting Point and Freezing Point Pure, crystalline solids have a characteristic melting point, the temperature at which the solid melts to become a liquid. The transition between the solid and the liquid is so sharp for small samples of a pure substance that melting points can be measured to 0.1oC. The melting point of solid oxygen, for example, is -218.4oC. Liquids have a characteristic temperature at which they turn into solids, known as their freezing point. In theory, the melting point of a solid should be the same as the freezing point of the liquid. In practice, small differences between these quantities can be observed. It is difficult, if not impossible, to heat a solid above its melting point because the heat that enters the solid at its melting point is used to convert the solid into a liquid. It is possible, however, to cool some liquids to temperatures below their freezing points without forming a solid. When this is done, the liquid is said to be supercooled. An example of a supercooled liquid can be made by heating solid sodium acetate trihydrate (NaCH3CO2 3 H2O). When this solid melts, the sodium acetate dissolves in the water that was trapped in the crystal to form a solution. When the solution cools to room temperature, it should solidify. But it often doesn't. If a small crystal of sodium acetate trihydrate is added to the liquid, however, the contents of the flask solidify within seconds. A liquid can become supercooled because the particles in a solid are packed in a regular structure that is characteristic of that particular substance. Some of these solids form very easily; others do not. Some need a particle of dust, or a seed crystal, to act as a site on which the crystal can grow. In order to form crystals of sodium acetate trihydrate, Na+ ions, CH3CO2- ions, and water molecules must come together in the proper orientation. It is difficult for these particles to organize themselves, but a seed crystal can provide the framework on which the proper arrangement of ions and water molecules can grow. Because it is difficult to heat solids to temperatures above their melting points, and because pure solids tend to melt over a very small temperature range, melting points are often used to help identify compounds. We can distinguish between the three sugars known as glucose (MP = 150oC), fructose (MP = 103-105oC), and sucrose (MP = 185-186oC), for example, by determining the melting point of a small sample. Measurements of the melting point of a solid can also provide information about the purity of the substance. Pure, crystalline solids melt over a very narrow range of temperatures, whereas mixtures melt over a broad temperature range. Mixtures also tend to melt at temperatures below the melting points of the pure solids. Boiling Point When a liquid is heated, it eventually reaches a temperature at which the vapor pressure is large enough that bubbles form inside the body of the liquid. This temperature is called the boiling point. Once the liquid starts to boil, the temperature remains constant until all of the liquid has been converted to a gas. The normal boiling point of water is 100oC. But if you try to cook an egg in boiling water while camping in the Rocky Mountains at an elevation of 10,000 feet, you will find that it takes longer for the egg to cook because water boils at only 90oC at this elevation. In theory, you shouldn't be able to heat a liquid to temperatures above its normal boiling point. Before microwave ovens became popular, however, pressure cookers were used to decrease the amount of time it took to cook food. In a typical pressure cooker, water can remain a liquid at temperatures as high as 120oC, and food cooks in as little as one-third the normal time. To explain why water boils at 90oC in the mountains and 120oC in a pressure cooker, even though the normal boiling point of water is 100oC, we have to understand why a liquid boils. By definition, a liquid boils when the vapor pressure of the gas escaping from the liquid is equal to the pressure exerted on the liquid by its surroundings, as shown in the figure below. Liquids boil when their vapor pressure is equal to the pressure exerted on the liquid by its surroundings. The normal boiling point of water is 100oC because this is the temperature at which the vapor pressure of water is 760 mmHg, or 1 atm. Under normal conditions, when the pressure of the atmosphere is approximately 760 mmHg, water boils at 100oC. At 10,000 feet above sea level, the pressure of the atmosphere is only 526 mmHg. At these elevations, water boils when its vapor pressure is 526 mmHg, which occurs at a temperature of 90oC. Pressure cookers are equipped with a valve that lets gas escape when the pressure inside the pot exceeds some fixed value. This valve is often set at 15 psi, which means that the water vapor inside the pot must reach a pressure of 2 atm before it can escape. Because water doesn't reach a vapor pressure of 2 atm until the temperature is 120oC, it boils in this container at 120oC. Liquids often boil in an uneven fashion, or bump. They tend to bump when there aren't any scratches on the walls of the container where bubbles can form. Bumping is easily prevented by adding a few boiling chips to the liquid, which provide a rough surface upon which bubbles can form. When boiling chips are used, essentially all of the bubbles that rise through the solution form on the surface of these chips.