

Unit 2: Matter as Solutions and Gases**Chapter 4: Reactions in Aqueous Solutions** (pg. 124 to 129)

- Solute is the substance being dissolved. Solvent is the substance doing the dissolving. When an ionic solid is dissolved in a liquid like water, the negative pole of the water molecule is attracted to the cations of the ionic solid whereas the positive pole of the water molecule is attracted to the anions of the ionic solid. This action slowly breaks apart the ionic bond and the ionic solid is “dissolved”. When a polar molecular solid is dissolved in a liquid like water, the negative pole of the water molecules is attracted to the positive pole of the solute and the positive pole of the water molecule is attracted to the negative pole of the solute. This action breaks apart any intermolecular forces between the molecules of the solid solute. It is important to note that the chemical bond of a molecular solute is not broken during the dissolving process, but the ionic bonds of the ionic solid are broken when it is dissolved.
- Nonelectrolytes are solutes that do not conduct electricity when dissolved in a solvent (molecular compound solutes). Electrolytes are solutes that conduct electricity when dissolved in a solvent (ionic compound solutes). When an ionic compound dissolves completely in water, all ions are broken apart from each other. Hence, electricity conduct strongly and these ions are called strong electrolytes. Weak electrolytes are ionic solids that are slightly soluble in water and very small amount of ions are broken apart from the solid solute. Therefore, they can only conduct electricity mildly.
- Hydration is a process where ions are surrounded by water molecules when being dissolved. The polar nature of the water molecule allows this to happen.
- This is because some dissolved ions are present in many water supplies. Besides, our body sweat glands produce sweat that contains ions. Thereby, they allow electric conduction to occur and we can easily get electrocuted if we operate electrical appliances when our hands are wet.
- $\text{LiF}_{(s)}$ dissociates into Li^+ and F^- when dissolves in water.
- Ionic compounds, strong acids, and strong bases (metal hydroxides) are strong electrolytes (completely broken up into ions of the compound). Weak acids and weak bases are weak electrolytes. Molecular substances other than acids or bases are nonelectrolytes.
 - very weak electrolyte
 - strong electrolyte (ionic compound)
 - strong electrolyte (strong acid)
 - weak electrolyte (weak acid)
 - nonelectrolyte (molecular compound - neither acid nor base)
- Since solutions must be electrically neutral, any flow of positive species (cations) must be balanced by the flow of negative species (anions). Therefore, the correct answer is **(d)**.
- Measure the conductance to see if the solution carries an electrical current. If the solution is conducting, then you can determine whether the solution is a strong or weak electrolyte by comparing its conductance with that of a known strong electrolyte.
- The advantage of writing net-ionic equations is to allow us to see what is the main event of the reaction, be it precipitation, neutralization, gas formation, and reduction-oxidation reactions.
- AgCl is insoluble in water. It will precipitate from solution. NaNO_3 is soluble in water and will remain as Na^+ and NO_3^- ions in solution. Diagram **(c)** best represents the mixture.
- (a)** $\text{Ca}_3(\text{PO}_4)_2$ (*insoluble*) **(b)** $\text{Mn}(\text{OH})_2$ (*insoluble*) **(c)** AgClO_3 (*soluble*) **(d)** K_2S (*soluble*)
- (a)** Complete Ionic: $2 \text{Ag}^+ + 2 \text{NO}_3^- + 2 \text{Na}^+ + \text{SO}_4^{2-} \rightarrow \text{Ag}_2\text{SO}_4(s) + 2 \text{Na}^+ + 2 \text{NO}_3^-$
 Net ionic: $2 \text{Ag}^+ + \text{SO}_4^{2-} \rightarrow \text{Ag}_2\text{SO}_4(s)$
- (b)** Complete Ionic: $\text{Ba}^{2+} + 2 \text{Cl}^- + \text{Zn}^{2+} + \text{SO}_4^{2-} \rightarrow \text{BaSO}_4(s) + \text{Zn}^{2+} + 2 \text{Cl}^-$
 Net Ionic: $\text{Ba}^{2+} + \text{SO}_4^{2-} \rightarrow \text{BaSO}_4(s)$
- (c)** Complete Ionic: $2 \text{NH}_4^+ + \text{CO}_3^{2-} + \text{Ca}^{2+} + 2 \text{Cl}^- \rightarrow 2 \text{NH}_4^+ + 2 \text{Cl}^- + \text{CaCO}_3(s)$
 Net Ionic: $\text{CO}_3^{2-} + \text{Ca}^{2+} \rightarrow \text{CaCO}_3(s)$
- (a)** Both reactants are soluble ionic compounds. The other possible ion combinations, Na_2SO_4 and $\text{Cu}(\text{NO}_3)_2$, are also soluble.

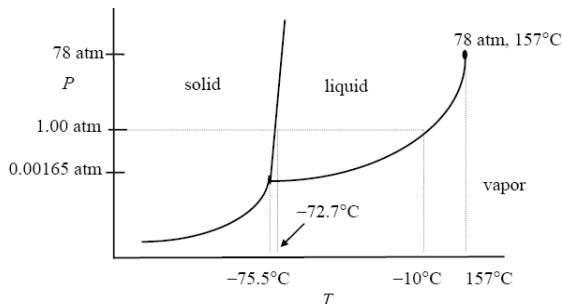
- (b) Both reactants are soluble. Of the other two possible ion combinations, KCl is soluble, but BaSO₄ is insoluble and will precipitate.
$$\text{Ba}^{2+} + \text{SO}_4^{2-} \rightarrow \text{BaSO}_4(s)$$
26. Arrhenius stated that acid is a H⁺ producer and a base is a OH⁻ producer. Brønsted stated that acid is a H⁺ donor and a base is a H⁺ acceptor. Brønsted's definition describes bases that are not ionic compound that contains OH⁻. We can understand why some molecular compound (like NH₃) and anions (like F⁻ or CN⁻) as bases.
27. Answers may vary. Monoprotic acid: HNO₃ (aq); Diprotic acid: H₂SO₄ (aq); Triprotic acid: H₃PO₄ (aq)
28. Acid-base neutralization will produce water and an acid-base salt.
29. Salt are ionic compound where the anion is a super weak base. KBr
30. Strong Acid: H₂SO₄ Strong Base: LiOH, Ba(OH)₂
Weak Acid: H₃PO₄, HCOOH, HF Weak Base: NH₃
31. (a) HI dissolves in water to produce H⁺ and I⁻, so HI is a **Brønsted acid**.
(b) CH₃COO⁻ can accept a proton to become acetic acid CH₃COOH, so it is a **Brønsted base**.
(c) H₂PO₄⁻ can either accept a proton, H⁺, to become H₃PO₄ and thus behaves as a **Brønsted base**, or can donate a proton in water to yield H⁺ and HPO₄²⁻, thus behaving as a **Brønsted acid**.
(d) HSO₄⁻ can either accept a proton, H⁺, to become H₂SO₄ and thus behaves as a **Brønsted base**, or can donate a proton in water to yield H⁺ and SO₄²⁻, thus behaving as a **Brønsted acid**.
33. Recall that strong acids and strong bases are strong electrolytes. They are completely ionized in solution. An ionic equation will show strong acids and strong bases in terms of their free ions. A net ionic equation shows only the species that actually take part in the reaction.
- (a) Molecular Equation: $\text{HBr}_{(aq)} + \text{NH}_3_{(aq)} \rightarrow \text{NH}_4\text{Br}_{(aq)}$
Complete Ionic: $\text{H}^+ + \text{Br}^- + \text{NH}_3_{(aq)} \rightarrow \text{NH}_4^+_{(aq)} + \text{Br}^-$
Net Ionic: $\text{H}^+ + \text{NH}_3_{(aq)} \rightarrow \text{NH}_4^+_{(aq)}$
- (b) Molecular Equation: $3 \text{Ba}(\text{OH})_2_{(aq)} + 2 \text{H}_3\text{PO}_4_{(aq)} \rightarrow \text{Ba}_3(\text{PO}_4)_3(s) + 6 \text{H}_2\text{O}_{(l)}$
Complete Ionic: $3 \text{Ba}^{2+} + 6 \text{OH}^- + 2 \text{H}_3\text{PO}_4_{(aq)} \rightarrow \text{Ba}_3(\text{PO}_4)_3(s) + 6 \text{H}_2\text{O}_{(l)}$
Net Ionic: $3 \text{Ba}^{2+} + 6 \text{OH}^- + 2 \text{H}_3\text{PO}_4_{(aq)} \rightarrow \text{Ba}_3(\text{PO}_4)_3(s) + 6 \text{H}_2\text{O}_{(l)}$
- (c) Molecular Equation: $2 \text{HClO}_4_{(aq)} + \text{Mg}(\text{OH})_2_{(aq)} \rightarrow 2 \text{H}_2\text{O}_{(l)} + \text{Mg}(\text{ClO}_4)_2_{(aq)}$
Complete Ionic: $2 \text{H}^+ + 2 \text{ClO}_4^- + \text{Mg}^{2+} + 2 \text{OH}^- \rightarrow 2 \text{H}_2\text{O}_{(l)} + \text{Mg}^{2+} + 2 \text{ClO}_4^-$
Net Ionic: $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}_{(l)}$
52. Calculate the amount of solute needed. Dissolve it in half the volume of distilled deionized water required using a beaker. Pour the content of the beaker into the volumetric flask. Wash the beaker, stirring rod and funnel with distilled deionized water. Top it up to the mark. Shake to mix completely.
53. 232 g KI 55. 6.00×10^{-3} mol MgCl₂
57. (a) 1.15 M (b) 0.608 M (c) 1.78 M
59. (a) 136 mL solution (b) 62.2 mL solution (c) 47 mL solution
61. Pipet the needed volume of the concentrated solution into the volumetric flask. Add distilled deionized water to the mark. Shake to mix completely.
63. To prepare the 0.646 M solution, you would dilute 323 mL of the 2.00 M HCl solution to a final volume of 1.00 L.
65. To prepare the 0.200 M solution, you would dilute 3.00 mL of the 4.00 M HNO₃ solution to a final volume of 60.0 mL.
67. 1.41 M
70. There can be a lot of ions in non-distilled, non-deionized water that will precipitate with the test solution.
71. 0.215 g AgCl 73. 0.165 g NaCl 77. 0.1106 M
79. (a) 42.78 mL (b) 158.5 mL (c) 79.23 mL
91. 0.773 L 93. 1145 g/mol 95. 43.4 g BaSO₄
107. (a) 1.40 M (b) 4.96 g Cl⁻

Chapter 5: Gases (pg. 163 to 167)

2. Gases assume the volume and shape of their containers. They are most compressible of all the states of matter. Gases will mix evenly and completely when confined to the same container. They have much lower densities than liquids and solids. **13.** 0.739 atm
15. Boyle's Law states that the pressure of a fixed amount of gas maintained at constant temperature is inversely proportional to the volume of the gas. ($P_1V_1 = P_2V_2$) Pressures can be in any units (atm, kPa, mmHg or torr). Volumes can be in mL or L. It is applicable when the gas has a fixed n and a constant T . Charles's Law states that the volume of a fixed amount of gas maintained at constant pressure is directly proportional to the absolute temperature of the gas. ($\frac{V_1}{T_1} = \frac{V_2}{T_2}$) Volumes can be in mL or L, but temperatures must be in Kelvin. Charles's law is applicable when the gas has a fixed n and a constant P . Guy-Lussac's Law states that the pressure of a fixed amount of gas maintained at constant volume is directly proportional to the absolute temperature of the gas. ($\frac{P_1}{T_1} = \frac{P_2}{T_2}$) Pressures can be in any units, but temperatures must be in Kelvin. It is applicable when the gas has a fixed and a constant volume. Avogadro's Law states that at constant pressure and temperature, the volume of a gas is directly proportional to the number of the gas present. ($\frac{V_1}{n_1} = \frac{V_2}{n_2}$) Volumes can be in mL or L, and amounts are in moles. Avogadro's law is applicable when the gas has a constant pressure and temperature.
16. A helium balloon expands as it rises in the air because as the air pressure drops with increasing altitude. According to Boyle's law, when pressure decreases, volume increases. Hence, the balloon expands.
17. (a) If the final temperature of the sample is above the boiling point, it would still be in the gas phase. The diagram that best represents this is choice (d).
(b) If the final temperature of the sample is below its boiling point, it will condense to a liquid. The liquid will have a vapor pressure, so some of the sample will remain in the gas phase. The diagram that best represents this is choice (b).
19. 1.30 L **21.** 457 mm Hg **23.** 44.1 L
25. The balanced equation is: $4 \text{NH}_3(\text{g}) + 5 \text{O}_2(\text{g}) \rightarrow 4 \text{NO}(\text{g}) + 6 \text{H}_2\text{O}(\text{g})$
Recall that Avogadro's Law states that the volume of a gas is directly proportional to the number of moles of gas at constant temperature and pressure. The ammonia and nitric oxide coefficients in the balanced equation are the same, so **one volume** of nitric oxide must be obtained from **one volume** of ammonia.
27. An ideal gas follow the kinetic molecular theory that assumes molecules in the gaseous state do not exert any force, either attractive or repulsive, on one another. It also assumes that the volume of the molecules is negligibly small compared with that of the container.
29. STP (Standard Temperature and Pressure) has a condition of 0°C and 1 atm (101.325 kPa or 760 mmHg). One mole of ideal gas at STP has a volume of 22.414 L.
30. The density of a gas is much lower than that of a liquid or solid under atmospheric condition because gases have much more spaces between the molecules compared to the solid and liquid counterparts. We tend to use the unit g/L to express density of a gas.
55. Dalton's Law of Partial Pressures states that the total pressure of a mixture of gases is the sum of the pressures that each gas would exert if it were present alone. The mole fraction is a dimensionless quantity that expresses the ratio of the number of moles of one component to the number of moles of all components present.
56. $P_{\text{total}} = 1 \text{ atm}$; $\chi_{\text{nitrogen}} = 0.80 \text{ atm}$; $\chi_{\text{oxygen}} = 0.20 \text{ atm}$ **61.** 0.45 g Na **63.** 4.8 %
65. The kinetic molecular theory assumes molecules in the gaseous state do not exert any force, either attractive or repulsive, on one another. It also assumes that the volume of the molecules is negligibly small compared with that of the container.
76. The gas would behave most ideally under (a) High temperature and low pressure.
83. C_6H_6 **87.** (a) 9.53 atm
89. 1.30×10^{22} molecules of gas. The most common gases present in exhaled air are: CO_2 , O_2 , N_2 , and H_2O .

Chapter 12: Intermolecular Forces and Liquids and Solids (pg. 418 to 423)

21. Unlike gases, molecules in liquid phase have significantly less spaces between them. Hence, there is not enough room to compact these molecules to a smaller volume. Therefore, liquids are incompressible.
22. Surface tension is the amount of energy required to stretch or increase the surface of a liquid by a unit area. The stronger the intermolecular forces molecules have with each other, the higher is its surface tension. This is because the cohesive forces inside the liquid (polar-polar forces) become much greater than the adhesive forces between the liquid (polar) and the air molecules (non-polar) outside. Such difference in these forces bind the liquid molecules closer together thus creating a “skin” with a tension force on the surface of the liquid.
24. Water has a higher cohesive force with adjacent molecules than the adhesive force between the glass or plastic tube that it is in. Therefore, water in a narrow tube would have a concave meniscus. Mercury, however, is non-polar. Thus, it lacks any cohesive forces between its atoms. The adhesive force between the mercury atoms and the plastic or glass tube is the only force in this situation. As a result, mercury in a narrow tube would produce a convex meniscus.
27. Viscosity is the resistance for a liquid to flow. The stronger the intermolecular forces in a liquid, the more viscose it is. This is because the molecules with stronger intermolecular forces with have a more difficult time rolling pass each other as they are more “sticky” with each other. Hence, they would have more resistance to flow.
28. As temperature increases, the molecules move faster and have more spaces between each other. This in turn loosens the intermolecular forces between them. As a result, they will have a easier time to roll pass each other and experience less resistance to flow (ie. less viscose).
29. Ice is less dense than water because water molecules lined up in a hexagonal structure as they crystallize into ice. This arrangement was brought on by the hydrogen bonding (a type of strong intermolecular force) water molecules have due to its V-shape structure with two sets of exposed lone-pairs, which makes water molecules very polar. As ice forms, the hexagonal structure takes up more space than its liquid counterpart. As such, ice has a lower density and floats on water readily.
30. This is because water expands as it freezes (ice has a higher density which means more volume for the same mass). If outdoor or unheated pipes are not drained completely or properly insulated from the cold, water left inside would expand and crack the piping causing damage and possibly flooding.
55. As a liquid boils into a gas, the vapour pressure of liquid is in dynamic equilibrium with the vapour pressure of the gas (on Earth, it would be the air pressure of a particular location). This means that equal number of water molecules are vaporizing as well as condensing at the same time.
60. Boiling point is the temperature of a liquid at which its vapour pressure is equalled to the of the external air pressure outside. From Table 5.2 on pg. 152 of the textbook, water has a boiling point of 65°C when the vapour pressure is 187.5 mm Hg.
65. Solid carbon dioxide is called dry ice because it does not pass through a liquid phase as it sublimates at normal atmospheric pressure. According to Figure 12.31 on pg. 416 of the textbook, liquid carbon dioxide can only appear at pressure above 5.2 atm and at temperature above -57°C.
67. From Figure 12.26 on pg. of the textbook, at 0.5 atm, the boiling points of dimethyl ether, water, and mercury are approximately 0°C, 80°C, and 300°C respectively.
68. During hot, dry summer days, there is not as much water vapour in the air, which means that the vapour pressure is low for that specific temperature. Hence, the water in the wet clothes will evaporate more quickly into the air in order to establish a dynamic equilibrium ($P_{\text{wet clothes}} > P_{\text{vapour, air}}$). During hot, humid summer days, the air is saturated with water molecules so the vapour pressure is high for that specific temperature. As such, the water in the wet clothes could not evaporate as quickly into the air ($P_{\text{wet clothes}} \leq P_{\text{vapour, air}}$).
73. (a) Other factors being equal, liquids evaporate faster at higher temperatures.
 (b) The greater the surface area, the greater the rate of evaporation.
 (c) Weak intermolecular forces imply a high vapor pressure and rapid evaporation.
75. The substance with the lowest boiling point will have the highest vapor pressure at some particular temperature. Thus, butane will have the highest vapor pressure at -10°C and toluene the lowest.
77. The solid ice turns to gas (sublimation). The temperature is too low for melting to occur.
- 87.



97. It has reached the critical point; the point of critical temperature (T_c) and critical pressure (P_c).

Chapter 13: Physical Properties of Solutions (pg. 448 to 452)

1. When ionic solids dissolve in a liquid solvent like water, it first has to break the ionic bonds of the solute and overcome (or loosen) the intermolecular forces of the solvent by making the solvent molecules not as “sticky”. Then, the solvent molecules have to form new intermolecular forces between them and the solute ions. When molecular solids dissolve in a liquid solvent like water, it first has to break the intermolecular forces of the solute and overcome (or loosen) the intermolecular forces of the solvent by making the solvent molecules not as “sticky”. Then, the solvent molecules have to form new intermolecular forces between them and the solute molecules. It is like fitting the solute molecules in the spacing between the solvent molecules.
2. “Like dissolves like” means the molecules with similar type of intermolecular forces dissolves well with each other. Examples are alcohol dissolving well in water (both have polar molecules), and Vaseline dissolving well in cooking oil (both have non-polar molecules).
6. Some factors that affect the solubility of solid in a liquid are their polar or non-polar nature of their molecules, temperature, particle size as well as the their molecular structures. Two liquids are considered miscible when the solute and solvent can dissolve infinitely into each other.
13. (a) 7.03% (b) 16.9% (c) 13%
15. (a) 0.0618 *m* or 0.618 mol/kg (b) 2.03 *m* or 2.03 mol/kg
17. (a) 1.74 *m* or 1.74 mol/kg (b) 0.87 *m* or 0.87 mol/kg (c) 6.99 *m* or 6.99 mol/kg
19. 3.0×10^2 g or 0.30 kg
21. 17.3 M or 17.3 mol/L
23. Most ionic compounds (because they are solids in room temperature) usually have increase solubility in water as temperature rises.
24. Henry’s law predicts that as pressure increases, the solubility of gases increase as well. However, external pressure has very little effect on the solubility of solid and liquid solutes.
25. 35.2 g salt / 100 g H₂O
27. With decreasing temperature, the solubility of gases increases. This is because gas molecules tend to be moving very fast at any given temperature compared to a liquid solvent. Most are also non-polar where water is very polar. Hence, lowering the temperature will make the solute molecules move slower, as well as decreasing the spacing between the liquid solvent molecules. Hence, the solute molecules can be better “trap” between the solvent molecules. Thus, allowing the increase in solubility of gases. This is of course, assuming that the gases does not react with water ($\text{NH}_3(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$) or compounds in the existing solutions (hemoglobin in blood binding with oxygen molecules).
28. Thermal pollution refers to adding unwanted or excessive heat into the environment. In the aquatic environment, increasing the temperature of bodies of water will decrease the oxygen level due to decrease gas solubilities. Hence, marine life that depends on the supply of oxygen will suffer greatly.
29. Henry law states that the solubility of a gas (molar concentration) is directly proportional to external pressure. The formula, $C = kP$ or when comparing two conditions of the same gas-liquid solution, $\frac{C_1}{P_1} = \frac{C_2}{P_2}$ can be used to calculate the solubility of gas at a constant temperature. C is molar concentration or solubility of gas in mol/L, P is pressure in atm, kPa, mmHg or torr. The constant, k , is in mol / (L• atm) or mol / (L• kPa) or mol / (L• mmHg) or mol / (L• torr) is specific to the type of gas and temperature. As pressure increases, liquid solvent has less vapour pressure above it’s surface meaning that the gaseous solute molecules are now forced into the spacing between the liquid solvent molecules. Thus, it increases the solubility of gases.
30. Henry’s law cannot be applied with gas solutes that have a high solubility in the liquid solvent due to the principle of “like dissolves like” or if they react readily with the solvent. Two examples are ammonia in water ($\text{NH}_3(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$) and hemoglobin in blood binding with oxygen molecules

31. The solubility of gases in water usually decreases with increasing temperature (see Figure 13.4 on pg. 433 of the text). When the water in the beaker is heated to 30°C, bubbles of air form because the air is less soluble in the water at the higher temperature. At 100°C, the normal boiling point of water, a phase transition is occurring from liquid water to water vapor. The bubbles will predominantly contain water vapor escaping from the liquid, although there will also be a small amount of air present in the bubbles at the higher temperature.
35. 1.0×10^{-5} mol/L or 0.010 mmol/L
38. Two examples of volatile liquids are gasoline and toluene (both non-polar substances have relatively low vapour pressures at a given temperature). Two examples of non-volatile liquids are water and ethylene glycol (both polar substances have relatively high vapour pressures at a given temperature).
41. As vapour pressure is lowered, it means molecules in liquid have to attain higher kinetic energy (temperature) to evaporate off its surface. Therefore, the corresponding boiling point will have to increase as well.
55. $\Delta T_b = 6.25^\circ\text{C}$ so the new boiling point is $80.1^\circ\text{C} + 6.25^\circ\text{C} = 86.4^\circ\text{C}$
 $\Delta T_f = 12.6^\circ\text{C}$ so the new freezing point is $5.5^\circ\text{C} - 12.6^\circ\text{C} = -7.1^\circ\text{C}$
59. 3.93 L of ethylene glycol is needed. $\Delta T_b = 5.6^\circ\text{C}$ so the new boiling point is $100.0^\circ\text{C} + 5.6^\circ\text{C} = 105.6^\circ\text{C}$
67. Electrolytic solutions consist of at least two moles of ions per mole of solute. Hence, the colligative properties like boiling point elevation and freezing point depression is amplified compared to their non-electrolytic counterparts.
70. van't Hoff factor is defined as the number of moles of ions dissociated per mole of solute dissolved. We need to multiply the van't Hoff factor into each of the boiling point elevation and freezing point depression formula to find the new boiling point and freezing point of an electrolytic solution.
71. CaCl_2 is an ionic compound and is therefore an electrolyte in water. Assuming that CaCl_2 is a strong electrolyte and completely dissociates (van't Hoff factor $i = 3$), the total ion concentration will be $3 \times 0.35 \text{ m} = 1.05 \text{ m}$, which is larger than the urea (non-electrolyte) concentration of 0.90 m.
(a) The CaCl_2 solution will show a larger boiling point elevation.
(b) The CaCl_2 solution will show a larger freezing point depression. The freezing point of urea solution will be higher. (Read the question carefully.)
(c) The CaCl_2 solution will have a larger vapor pressure lowering.
73. The solution with the lowest molality will have the highest freezing point (smallest freezing point depression): (d) > (e) > (a) > (c) > (b).
75. (a) $\Delta T_b = 2.8^\circ\text{C}$ so the new boiling point is $100.0^\circ\text{C} + 2.8^\circ\text{C} = 102.8^\circ\text{C}$
 $\Delta T_f = 10.0^\circ\text{C}$ so the new freezing point is $0.0^\circ\text{C} - 10.0^\circ\text{C} = -10.0^\circ\text{C}$
(b) $\Delta T_b = 2.0^\circ\text{C}$ so the new boiling point is $100.0^\circ\text{C} + 2.0^\circ\text{C} = 102.0^\circ\text{C}$
 $\Delta T_f = 7.14^\circ\text{C}$ so the new freezing point is $0.00^\circ\text{C} - 7.14^\circ\text{C} = -7.14^\circ\text{C}$
77. Both NaCl and CaCl_2 are strong electrolytes. Urea and sucrose are non-electrolytes. The NaCl or CaCl_2 will yield more particles per mole of the solid dissolved, resulting in greater freezing point depression. Also, sucrose and urea would make a mess when the ice melts.