

Unit 3: States of Matter**Chapter 5: Gases**

pg. 209 to 217

2. Gases take the shape and volume of the container. They have very low boiling points. Gases can be compressed much more easily than liquids and solids. Gases have much lower densities (g/L) than liquids and solids (g/mL). Two or more different gases will mix readily and thoroughly when placed in the same container to form a gaseous mixture.
3. Pressure is the amount of force exerted per unit of area. The metric unit for pressure is Pascal (Pa).
4. A barometer is an instrument used to measure atmospheric pressures. It can measure the pressure the pressure exerted by the atmosphere by using water, air, or mercury. The setup involves a dish of liquid and an inverted vacuum tube in it. As the atmospheric pressure pushes down on the dish, the liquid in it will get drawn up and this height can be translated into pressure. A manometer is a pressure measuring device of a container of gas. It is a U-tube with liquid inside. As one end of the tube is hooked up to the gas container, a closed-ended manometer will register the pressure of the gas as a difference in height.
5. Mercury is a liquid of choice for most barometer and manometer because it has a high density and its volume is not easily affected by temperature.
6. The height of the liquid is inversely proportional to its density, not the cross-sectional area of the tube.
7. It would be easier to drink from a straw on Mt. Everest since the atmospheric pressure at that altitude is lowered. Thus, less force is needed to create the difference in pressure needed to draw the liquid into the mouth.
9. A gas is a state of matter where the molecules or atoms have nearly no interaction with each other. A vapour is a substance in a gas phase that exists lower than its boiling point. It can easily condense back into a liquid if the partial pressure is increased sufficiently.
11. There is a fixed amount of gas in the atmosphere. In order to keep this amount constant, a decrease in atmospheric pressure at one part of the planet means an increase in the pressure on another part of the planet.
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 n and a constant V . Avogadro's Law states that at constant pressure and temperature, the volume of a gas is directly proportional to the number of moles 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).

86. The gas would behave most ideally under (a) High temperature and low pressure.
87. $P_{\text{real}} = \frac{nRT}{V - nb} - a\left(\frac{n}{V}\right)^2$ where $(V - nb)$ is the volume correction and $a\left(\frac{n}{V}\right)^2$ is the pressure correction.
88. (a) The corrected volume is always less than the volume of the container, V .
(b) Since $a_{\text{ammonia}} > a_{\text{neon}}$, one can conclude that ammonia has a stronger attractive force with each other compared to that of neon.
89. $P_{\text{real}} = 18.0 \text{ atm}$; $P_{\text{ideal}} = 18.5 \text{ atm}$
93. $V_{\text{product}} = 1.7 \times 10^2 \text{ L}$; $P_{\text{H}_2\text{O}} = 0.41 \text{ atm}$; $P_{\text{N}_2} = 0.25 \text{ atm}$; $P_{\text{O}_2} = 0.041 \text{ atm}$ 95. 0.273 g NH_4NO_2
107. O_2 and NO_2 remain; $P_{\text{NO}_2} = 0.333 \text{ atm}$; $P_{\text{O}_2} = 0.166 \text{ atm}$ 111. 33.1 % by mass of Na_2CO_3
123. 86.0 g/mol – C_6H_{12} 135. $P_{\text{NO}_2} = 0.53 \text{ atm}$; $P_{\text{N}_2\text{O}_4} = 0.45 \text{ atm}$

Chapter 11: Intermolecular Forces and Liquids and Solids

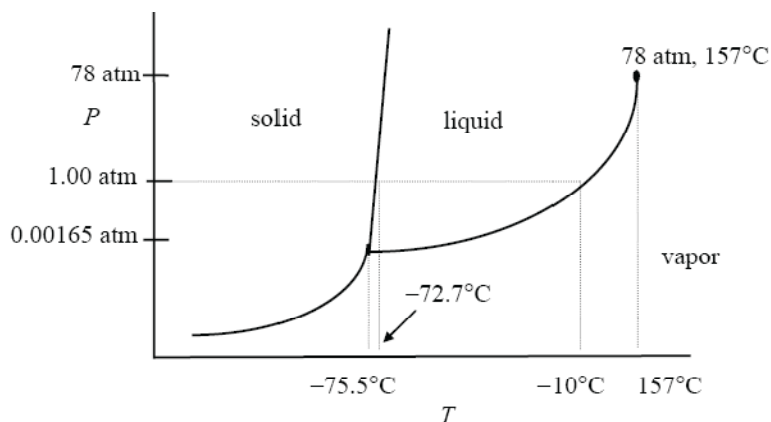
pg. 494 to 500

2. “Polarizability” means the ability to create a molecule with high polarity. Bonds that consist of atoms with large dipole moment (large difference in electronegativities) tend to have high “polarizability”. The stronger the molecule’s overall “polarizability”, the stronger are their intermolecular forces.
3. Temporary dipole moment, such as London dispersion force, is the result of electron “sloshing” in their mixed orbitals. Due to their non-permanent nature, the intermolecular forces results are not strong. Permanent dipole moments of polar molecules are the result large difference in electronegativities of the adjacent bond atoms, and their molecular geometry. The resulting intermolecular forces are very strong as these dipole moments are permanent.
6. Hydrogen bonds will result when H atoms are bonded with O, N, F or Cl atoms.
7. ICl has a dipole moment and Br_2 does not. The dipole moment increases the intermolecular attractions between ICl molecules and causes that substance to have a higher melting point than bromine.
9. All are tetrahedral (AB_4 type) and are non-polar. Therefore, the only intermolecular forces possible are dispersion forces. Without worrying about what causes dispersion forces, you only need to know that the strength of the dispersion force increases with the number of electrons in the molecule (all other things being equal). As a consequence, the magnitude of the intermolecular attractions and of the boiling points should increase with increasing molar mass.
13. CO_2 is a non-polar molecular compound. The only intermolecular force present is a relatively weak dispersion force (small molar mass). CO_2 will have the lowest boiling point. CH_3Br is a polar molecule. Dispersion forces (present in all matter) and dipole-dipole forces will be present. This compound has the next highest boiling point. CH_3OH is polar and can form hydrogen bonds, which are especially strong dipole-dipole attractions. Dispersion forces and hydrogen bonding are present to give this substance the next highest boiling point. RbF is an ionic compound. Ion-ion attractions are much stronger than any intermolecular force. RbF has the highest boiling point. (mp: $\text{CO}_2 < \text{CH}_3\text{Br} < \text{CH}_3\text{OH} < \text{RbF}$)
15. (a) O_2 : it has more electrons than N_2 (both are non-polar) and therefore has stronger dispersion forces.
(b) SO_2 : it is polar (most important) and also has more electrons than CO_2 (non-polar). More electrons imply stronger dispersion forces.
(c) HF : although HI has more electrons and should therefore exert stronger dispersion forces, HF is capable of hydrogen bonding and HI is not. Hydrogen bonding is the stronger attractive force
17. (a) CH_4 has a lower boiling point because NH_3 is polar and can form hydrogen bonds; CH_4 is non-polar and can only form weak attractions through dispersion forces.
(b) KCl is an ionic compound. Ion-ion forces are much stronger than any intermolecular forces. I_2 is a non-polar molecular substance; only weak dispersion forces are possible.

- (f) SiO_2 (quartz) is a hard, high melting nonmetallic compound; it forms covalent crystals like boron and C (diamond).
- (g) LiCl is an ionic compound; it forms ionic crystals.
- (h) Cr (chromium) is a metal and forms metallic crystals.
57. Amorphous solids are solids that are formed rapidly (when liquid is cooled quickly), and their atoms or molecules do not have time to align themselves and may become locked in positions other than those of a regular crystal. In contrast, crystalline solids may have regular lattice structures like those in ionic solids.
59. Phase change is a transformation from one phase of matter into another. It is usually characterized by the addition or removal of heat. Melting is from solid to liquid, while the reverse is called freezing. Evaporation refers to the transformation of liquid to vapour. The opposite of that transformation is called condensation. As a solid change directly to a vapour and bypassing the liquid phase, it is termed sublimation. If the vapour is converted to solid directly, it is called deposition.
60. The equilibrium vapour pressure of a liquid occurs when the rate of evaporation is equalled to the rate of condensation at a particular temperature. At this state, it is called dynamic equilibrium. As temperature increases, the equilibrium vapour pressure increases as well. Hence, as the atmospheric pressure decreases, boiling point tends to decrease.
61. For example, at the standard atmospheric pressure of 1 atm, water boils at 100°C . This means at 100°C , water's vapour pressure is also at 1 atm. Hence, the rate of condensation of water vapour to liquid water is the same as the rate of evaporation of liquid water to water vapour.
64. The larger the value of the molar heat of vaporization, the stronger is the intermolecular force between the liquid molecules.
66. 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 5.3 on pg. 152 of the textbook, water has a boiling point of 65°C when the vapour pressure is 187.5 mm Hg.
68. Critical temperature is a temperature of a gas above which it can not be liquefy no matter how great is the applied pressure. This is also the highest temperature a substance can exist as a liquid.
69. The stronger is the intermolecular force of a liquid, the higher is its boiling point and critical temperature. The boiling point is higher for liquid with strong intermolecular force because more energy is required to separate them completely when it moves from a liquid phase to a gaseous phase. The critical temperature is higher for water because it ha stronger intermolecular force. It takes a large amount of energy to cause these molecules to completely separate from each other to a point where no amounts of pressure can liquefy it.
70. Water is very polar and has hydrogen bonds. Carbon tetrachloride, on the other hand, is non-polar. Hence, it takes more energy to separate water molecules compared to carbon tetrachloride at a given pressure. Therefore, water has a higher boiling point and melting point than CCl_4 at the same pressure. This makes CCl_4 very volatile at low temperature because it has a higher vapour pressure than water at the same temperature.
71. 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 11.41 on pg. 489 of the textbook, liquid carbon dioxide can only appear at pressure above 5.2 atm and at temperature above -57°C .
72. In a closed container, the vapour pressure of a liquid is depended on temperature only. (c)
73. From Figure 11.35 on pg. 481 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.
74. 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 ($P_{\text{wet clothes}} \leq P_{\text{vapour, air}}$).

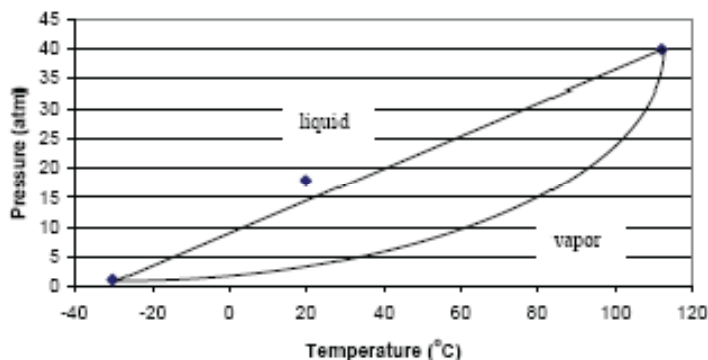
76. No, adding another Bunsen burner would not change the boiling point of the water. At 1 atm, the boiling point of water remains at 100°C. This is because the boiling point of a liquid is related to its intermolecular forces. It is the temperature where the dynamic equilibrium of the vapour pressure is achieved. Adding another Bunsen burner will change the rate (how fast) the water boils, but it will not change the temperature at which it will boil.
79. (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.
81. 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.
85. $\Delta H_{\text{vap}} = 60.1 \text{ kJ/mol}$
87. Application of the Clausius-Clapeyron, Equation (11.5) of the text, predicts that the more the vapor pressure rises over a temperature range, the smaller the heat of vaporization will be. Considering the equation below, if the vapor pressure change is greater, then $\frac{P_1}{P_2}$ is a smaller number and therefore ΔH is smaller. Thus, the molar heat of vaporization of **X** < **Y**.
89. A phase diagram is a diagram that shows phases of a substance at various pressures and temperatures. In addition of determining the phases of a substance at these different conditions, we can also see how changes in pressures and temperatures may change its phases. Furthermore, the phase diagram indicates the triple point and the critical temperature and pressure of a substance.
90. The phase diagram of water consists of a solid liquid line that has a negative slope. Most other substances tend to have a positive slope for their solid-liquid line. This means as pressure increases, the melting (freezing point) of water decreases. This is because water has strong intermolecular forces and its V-shaped molecular structure gives rise to a hexagonal solid crystalline configuration as it freezes. As a result, water expands as it freezes. As the pressure increases, more energy are needed be extracted from the molecules (they need to slow down even more) in order for these arrangements to form. Hence, the freezing point of water lowers with increasing pressures.
91. The pressure exerted by the blades on the ice lowers the melting point of the ice. A film of liquid water between the blades and the solid ice provides lubrication for the motion of the skater. The main mechanism for ice skating, however, is due to friction. See Chemistry in Action on p. 490 of the text.

93.



95. (a) Boiling liquid ammonia requires breaking hydrogen bonds between molecules. Dipole-dipole and dispersion forces must also be overcome.
 (b) P_4 is a nonpolar molecule, so the only intermolecular forces are of the dispersion type.
 (c) CsI is an ionic solid. To dissolve in any solvent ion-ion interparticle forces must be overcome.
 (d) Metallic bonds must be broken.
99. Reading directly from the graph: (a) solid; (b) vapor.
101. CCl_4 . Generally, the larger the number of electrons and the more diffuse the electron cloud in an atom or a molecule, the greater its polarizability. Recall that polarizability is the ease with which the electron distribution in an atom or molecule can be distorted.

103. As the vacuum pump is turned on and the pressure is reduced, the liquid will begin to boil because the vapour pressure of the liquid is greater than the external pressure (approximately zero). The heat of vaporization is supplied by the water, and thus the water cools. Soon the water loses sufficient heat to drop the temperature below the freezing point. Finally the ice sublimates under reduced pressure.
115. LiF, ionic bonding and dispersion forces; BeF₂, ionic bonding and dispersion forces; BF₃, dispersion forces; CF₄, dispersion forces; NF₃, dipole-dipole interaction and dispersion forces; OF₂, dipole-dipole interaction and dispersion forces; F₂, dispersion forces.
131. (a) Pumping allows Ar atoms to escape, thus removing heat from the liquid phase. Eventually the liquid freezes.
(b) The solid-liquid line of cyclohexane is positive. Therefore, its melting point increases with pressure.
(c) These droplets are super-cooled liquids.
(d) When the dry ice is added to water, it sublimates. The cold CO₂ gas generated causes nearby water vapour to condense, hence the appearance of fog.
133. The ice condenses the water vapor inside. Since the water is still hot, it will begin to boil at reduced pressure. (Be sure to drive out as much air in the beginning as possible.)
- 139.



Plotting the three points, and connecting the boiling point to the critical point with both a straight line and a curved line, we see that the point (20°C, 18 atm) lies on the liquid side of the phase boundary. The gas will condense under these conditions. The curved line better represents the liquid/vapor boundary for a typical phase diagram. See Figures 11.40 and 11.41 of the text.

Chapter 12: Physical Properties of Solutions

pg. 534 to 538

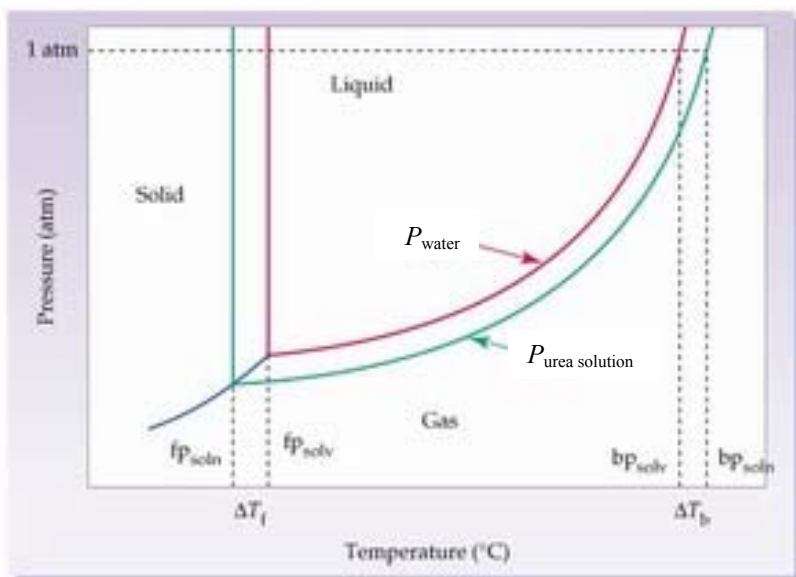
1. An unsaturated solution is a solution where it hasn't reached its maximum solubility at a given temperature. A saturated solution is a solution at its maximum solubility at that given temperature. A supersaturated solution is a solution that has exceeded its maximum solubility at a given temperature.
2. Crystallization and precipitation occur when a supersaturated solution is suddenly disturbed or its solvent is evaporated over time. Crystallization tends to occur slowly as the larger solid crystals need time to form properly. Precipitation occurs when solvent boils off quickly and the solid that forms are relatively much smaller in size.
3. 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.

4. "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).
5. Solvation refers to the dissolving process. There are many factors that can affect solvation. They are temperature, particle size, agitation, and pressure. Salt dissolves in water by first breaking themselves into ions. The water molecules then have to loosen their intermolecular forces (hydrogen bonds). Finally, new intermolecular forces (ion-dipole forces) are formed as different ends of the water molecules are attracted to the cations and anions of the solute. In another example, sugar molecules, a molecular substance, dissolves in water by first overcoming their own intermolecular forces. Again, water molecules have to then loosen their own hydrogen bonds. At the end, sugar molecules are fitted between the spacing of the water molecules as new intermolecular forces are formed between sugar and water molecules.
6. Solution process that are extremely endothermic usually consists of solute and solvent molecules that are not alike (one is polar and the other one is non-polar). Hence, the hydration energy is much smaller than ΔH_1 or E_{lattice} . This means that it takes much more energy to loosen the intermolecular forces of the solute and solvent and not enough energy is given off as they tried to form new intermolecular forces with each other. On the other hand, solution process that are exothermic or slightly endothermic have much higher hydration energy (energy given off as solute and solvent molecules form new intermolecular forces) compared to the energy required to loosen their cohesive forces. Hence, solute and solvent that are alike (have the same degree of polarity) tends to be very soluble with each other.
9. CsF is an ionic solid; the ion-ion attractions are too strong to be overcome in the dissolving process in benzene. The ion-induced dipole interaction is too weak to stabilize the ion. Non-polar naphthalene molecules form a molecular solid in which the only interparticle forces are of the weak dispersion type. The same forces operate in liquid benzene causing naphthalene to dissolve with relative ease. Like dissolves like.
11. The order of increasing solubility is: $\text{O}_2 < \text{Br}_2 < \text{LiCl} < \text{CH}_3\text{OH}$. Methanol is miscible with water because of strong hydrogen bonding. LiCl is an ionic solid and is very soluble because of the high polarity of the water molecules. Both oxygen and bromine are non-polar and exert only weak dispersion forces. Bromine is a larger molecule and is therefore more polarizable and susceptible to dipole-induced dipole attractions.
13. Percent by Mass (%) = $\frac{\text{mass of solute in grams}}{\text{mass of solution (solute + solvent) in grams}} \times 100\%$
 Mole Fraction (unitless – between 0 and 1) = $\frac{\text{moles of solute}}{\text{moles of solution (solute + solvent)}}$
 Molarity (mol/L or M) = $\frac{\text{moles of solute}}{\text{Volume of solution (solute + solvent) in L}}$
 Molality (mol/kg or *m*) = $\frac{\text{moles of solute}}{\text{mass of solvent in kg}}$
- Molarity is most affected by temperature as volume of the solvent changes with temperature. Mole fraction, molality, and percent by mass are concentrations units that are not affected by the change in temperature. However, percent by mass is commonly used by the largely scientific illiterate public as they have no idea what a mole is.
15. (a) 7.03% (b) 16.9% (c) 13% 17. (a) 0.0618 *m* (b) 2.03 *m*
 19. (a) 1.74 *m* (b) 0.87 *m* (c) 6.09 *m* 21. Ethanol Mass = 3.0×10^2 g
 23. Molarity = 17.3 M ; Molality = 25.1 *m*
 25. Most ionic compounds (because they are solids in room temperature) usually have increase solubility in water as temperature rises. 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.
 27. Solubility of Salt = 35.2 g salt / 100 g H₂O 29. 44 g KClO₃
 30. 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. As a result, 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 ($\text{NH}_3(g) + \text{H}_2\text{O}(l) \rightarrow \text{NH}_4^+(aq) + \text{OH}^-(aq)$) or compounds in the existing solutions (haemoglobin in blood binding with oxygen molecules).

31. 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.
32. 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. 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 ($\text{NH}_3(g) + \text{H}_2\text{O}(l) \rightarrow \text{NH}_4^+(aq) + \text{OH}^-(aq)$) and haemoglobin in blood binding with oxygen molecules.
33. 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.
34. The solubility of gases in water usually decreases with increasing temperature. See Figure 12.5 of the text. Boiling the water decreased the oxygen content of the water. Cooling the water quickly did not allow enough time for oxygen gas from the atmosphere to re-dissolve in the water to reach the equilibrium concentration at the lower temperature. The oxygen concentration in the water was too low for the goldfish to survive.
35. When a dissolved gas is in dynamic equilibrium with its surroundings, the number of gas molecules entering the solution (dissolving) is equal to the number of dissolved gas molecules leaving and entering the gas phase. When the surrounding air is replaced by helium, the number of air molecules leaving the solution is greater than the number dissolving. As time passes the concentration of dissolved air becomes very small or zero, and the concentration of dissolved helium increases to a maximum.
37. 1.0×10^{-5} mol/L
39. Colligative properties (or collective properties) are properties that depend only on the number of solute particles in solution and not on the nature of the solute particles. These properties are bound together by the fact that they all depend on the number of solute particles present, regardless of whether they are atoms, ions or molecules. The colligative properties are lowering of vapour pressures, boiling-point elevation, freezing-point depression, and osmotic pressure.
40. 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. Raoult's Law states that the partial pressure of a solvent over a solution P_1 , is given by the vapour pressure of the pure solvent, P_1° , times the mole fraction of the solvent in the solution, χ_1 . Hence, the Raoult's Law equation is $P_1 = \chi_1 P_1^\circ$.
42. An ideal solution is any solution that obeys Raoult's law. It is usually a solution with relatively small amount of interactions between solute and solvent molecules. An example of an ideal solution is benzene in toluene. Both substances have similar molar mass and are non-polar. This means they have nearly no molecular interactions. Therefore, there are no other factors that will affect the vapour pressure of the solution mixture.
43. Boiling-point elevation: $\Delta T_b = K_b m$ ($\Delta T_b =$ Change in solvent's boiling point in $^\circ\text{C}$,
 $K_b =$ molal boiling point constant in $^\circ\text{C}/m$, $m =$ molality of the solute)
 Freezing-point elevation: $\Delta T_f = K_f m$ ($\Delta T_f =$ Change in solvent's freezing point in $^\circ\text{C}$,
 $K_f =$ molal freezing point constant in $^\circ\text{C}/m$, $m =$ molality of the solute)
44. 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.
- 45.



47. Osmotic Pressure: $\Pi = CRT$ ($\Pi =$ Osmotic pressure in atm, $C =$ Molarity of Solution in mol/L
 $R =$ Gas constant $= 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$, $T =$ temperature in Kelvin)

50. Molar Mass Determination by Freezing-point Depression:

$$\Delta T_f = K_f m = K_f \times \frac{n_{\text{solute}}}{\text{kg of solvent}} = K_f \times \frac{\left(\frac{m_{\text{solute}}}{\text{Molar Mass}_{\text{solute}}}\right)}{\text{kg of solvent}} = K_f \times \frac{m_{\text{solute}}}{(\text{Molar Mass}_{\text{solute}})(\text{kg of solvent})}$$

$$\text{Molar Mass}_{\text{solute}} = \frac{K_f \times m_{\text{solute}}}{\Delta T_f (\text{kg of solvent})}$$

Molar Mass Determination by Osmotic Pressure:

$$\Pi = CRT = \frac{n_{\text{solute}}}{V} RT = \frac{\left(\frac{m_{\text{solute}}}{M_{\text{solute}}}\right)}{V} RT = \frac{m_{\text{solute}} RT}{M_{\text{solute}} V} \quad M_{\text{solute}} = \frac{m_{\text{solute}} RT}{\Pi V}$$

Molar Mass is commonly determined by Freezing-point depression or by Osmotic pressure because they are relatively safe and easy to set up in the lab. We do not need to boil a solution as required by boiling-point elevation, nor do we need expensive equipment like the mass spectrometer.

51. $P_{\text{solution}} = 30.8 \text{ mm Hg}$

53. 88.6 mm Hg

57. New $T_b = 86.4^\circ\text{C}$; New $T_f = -7.1^\circ\text{C}$ 65. 32.9 atm 67. $\text{C}_{15}\text{H}_{20}\text{O}_{10}\text{N}_5$
69. 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. An ion-pair is made up of one or more cations and one or more anions held together by electrostatic forces. The presence of an ion pair reduces the number of particles in the solution, causing a reduction in the colligative properties.
- (a) As the charges on the ions increase, the effect of ion-pair increases.
(b) As the size of the ions increase, the effect of ion-pair increases.
(c) As the solvent becomes more non-polar, the effect of ion-pair increases.
(d) As the concentrations of the ions increase, the effect of ion-pair increases.
71. (a) Na_2SO_4 (b) MgSO_4 (c) KBr
72. 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.
73. CaCl_2 is an ionic compound and is therefore an electrolyte in water. Assuming that CaCl_2 is a strong electrolyte and completely dissociates (no ion pairs, van't Hoff factor $i = 3$), the total ion concentration will be $3 \times 0.35 = 1.05 m$, 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 **the urea** solution will be higher.
(c) The CaCl_2 solution will have a larger vapour pressure lowering.
75. Assume that all the salts are completely dissociated. Calculate the molality of the ions in the solutions.
(a) $0.10 m \text{ Na}_3\text{PO}_4$: $0.10 m \times 4 \text{ ions/unit} = 0.40 m$ (b) $0.35 m \text{ NaCl}$: $0.35 m \times 2 \text{ ions/unit} = 0.70 m$
(c) $0.20 m \text{ MgCl}_2$: $0.20 m \times 3 \text{ ions/unit} = 0.60 m$ (d) $0.15 m \text{ C}_6\text{H}_{12}\text{O}_6$: non-electrolyte, $0.15 m$
(e) $0.15 m \text{ CH}_3\text{COOH}$: weak electrolyte, slightly greater than $0.15 m$
- The solution with the lowest molality will have the highest freezing point (smallest freezing point depression): (d) > (e) > (a) > (c) > (b).
77. (a) New $T_b = 102.8^\circ\text{C}$; New $T_f = -10.0^\circ\text{C}$ (b) New $T_b = 102.0^\circ\text{C}$; New $T_f = -7.14^\circ\text{C}$
79. 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.
81. The temperature and molarity of the two solutions are the same. If we divide Equation (12.12) of the text for one solution by the same equation for the other, we can find the ratio of the van't Hoff factors in terms of the osmotic pressures ($i = 1$ for urea).
- $$\frac{\pi_{\text{CaCl}_2}}{\pi_{\text{urea}}} = \frac{iMRT}{MRT} = i = \frac{0.605 \text{ atm}}{0.245 \text{ atm}} = 2.47$$