

Chem 116 POGIL Worksheet - Week 3 - Solutions
Intermolecular Forces, Liquids, Solids, and Solutions

Key Questions

1. Is the average kinetic energy of molecules greater or lesser than the energy of intermolecular forces of attraction in (a) solids, (b) liquids, and (c) gases?
 - (a) In solids, kinetic energy is less than intermolecular energy.
 - (b) In liquids, kinetic energy is less than intermolecular energy, but the disparity is less than in solids.
 - (c) In gases, the kinetic energy is much greater than the energy of intermolecular forces of attraction.

2. Why does increasing the temperature cause a substance to change in succession from a solid to a liquid to a gas?

Kinetic energy increases with temperature. As the kinetic energy rises with temperature, the intermolecular forces of attraction are overcome by increasing molecular motion.

3. Why do substances with high surface tension also tend to have high viscosities?

Liquids with stronger intermolecular forces of attraction hold the molecules closer together, which causes stronger surface tension and greater resistance to flowing (viscosity).

4. Why do surface tension and viscosity decrease with increasing temperature?

The increased kinetic energy with rising temperature overcomes the cohesive intermolecular forces of attraction.

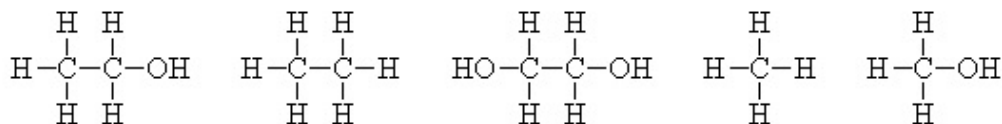
5. Name the kind or kinds of intermolecular forces that must be overcome to convert the following from liquid or solid to gas: (a) Br₂, (b) CH₃OH, (c) CO₂, (d) HCN, (e) NH₃

- (a) Br₂ – London dispersion
- (b) CH₃OH – London dispersion, dipole-dipole, hydrogen bonding
- (c) CO₂ – London dispersion (CO₂ is linear and therefore nonpolar.)
- (d) HCN – London dispersion, dipole-dipole (HCN is linear but polar.)
- (e) NH₃ – London dispersion, dipole-dipole, hydrogen bonding

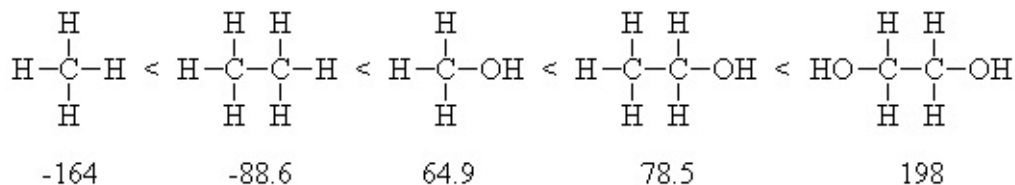
6. Normal alkanes are hydrocarbons with unbranched carbon chains, having a general formula C_nH_{2n+2}. At room temperature, ethane, C₂H₆, is a gas; hexane, C₆H₁₄, is a liquid; and octadecane, C₁₈H₃₈, is a solid. Describe the intermolecular forces present in each substance and explain the differences in their room-temperature phases.

All are nonpolar and therefore only have London dispersion forces. However, London dispersion forces rise with molecular weight, as the numbers of electrons increase, which in turn cause the polarizabilities to increase. Thus, the order of increasing intermolecular forces is $C_2H_6 < C_6H_{14} < C_{18}H_{38}$. As the London dispersion forces increase the tendency to be in a condensed phase increases.

7. Arrange the following in order of increasing boiling point:



The compounds that have -OH groups will have hydrogen bonding and therefore much higher boiling points than those that do not. Within the two groups (hydrogen bonding or no hydrogen bonding), the ordering follows increasing London dispersion forces (increasing molecular weight) and additional hydrogen bonding. Note that the last compound, ethylene glycol, has twice as many sites for hydrogen bonding per molecule as the two alcohols, CH_3OH and C_2H_5OH . The observed boiling point ($^{\circ}C$) is shown for each compound.



8. How much heat is required to heat 10.0 g of ice at $-5.00^{\circ}C$ to become liquid water at $+7.00^{\circ}C$? In this temperature range, the heat capacity of $H_2O(s)$ is $37.7 \text{ J/mol}\cdot\text{K}$, and the heat capacity of $H_2O(l)$ is $75.8 \text{ J/mol}\cdot\text{K}$. The molar heat of fusion of ice is 6.01 kJ/mol .

Calculate the heat to warm the ice, then the heat to melt the ice, and finally the heat to warm the liquid water. The sum is the total heat required.

$$\text{mol } H_2O = 10.0 \text{ g} / 18.02 \text{ g}\cdot\text{mol}^{-1} = 0.555 \text{ mol}$$

$$\text{Heat ice to } 0^{\circ}C: q_1 = C_p\Delta T = (37.7 \text{ J/mol}\cdot\text{K})(0.555 \text{ mol})(5.00 \text{ K}) = 105 \text{ J}$$

$$\text{Melt ice at } 0^{\circ}C: q_2 = \Delta H_{\text{melt}} \times \text{moles} = (6.01 \text{ kJ/mol})(0.555 \text{ mol}) = 3.34 \text{ kJ} = 3340 \text{ J}$$

$$\text{Heat liquid to } 7^{\circ}C: q_3 = C_p\Delta T = (75.8 \text{ J/mol}\cdot\text{K})(0.555 \text{ mol})(7.00 \text{ K}) = 294 \text{ J}$$

$$\text{Total heat: } q_t = q_1 + q_2 + q_3 = 105 \text{ J} + 3340 \text{ J} + 294 \text{ J} = 3739 \text{ J} = 3.74 \text{ kJ}$$

9. Explain how each of the following affects the vapor pressure of a liquid: (a) the volume of the liquid, (b) the volume of the container, (c) the surface area of the liquid, (d) the temperature, (e) intermolecular forces of attraction, (f) the density of the liquid.
- (a) No effect
 - (b) No effect
 - (c) No effect
 - (d) Higher vapor pressure at higher temperature
 - (e) Lower vapor pressure with stronger intermolecular forces of attraction
 - (f) Density is mass per unit volume. It tends to increase as molecular weight increases. London dispersion forces also increase with molecular weight, and this would cause a decrease in volatility, resulting in lower vapor pressure. Thus, denser liquids tend to have lower vapor pressures.
10. Describe the phases and/or phase transitions experienced by CO_2 under the following conditions:
- (a) Heating from $-100\text{ }^\circ\text{C}$ to $30\text{ }^\circ\text{C}$ at 1.0 atm
Solid to gas, with a normal sublimation point at $-78\text{ }^\circ\text{C}$.
 - (b) Heating from $-100\text{ }^\circ\text{C}$ to $50\text{ }^\circ\text{C}$ at 70 atm
Solid to liquid to gas
 - (c) A sample at $35\text{ }^\circ\text{C}$ and 100 atm
A point above the critical point, so a supercritical fluid
 - (d) A sample at $-50\text{ }^\circ\text{C}$ and 6.0 atm
A point in the liquid range, just above the triple point
11. Does carbon dioxide have a normal boiling point? Explain.
The normal boiling point is the temperature at which the vapor pressure above a liquid reaches exactly 1 atm. At 1 atm there is no liquid-vapor equilibrium for CO_2 , so it does not have a normal boiling point.
12. Describe the conditions under which liquid carbon dioxide boils.
At a pressure and temperature just above the triple point ($-57\text{ }^\circ\text{C}$, 5.1 atm), liquid CO_2 would exist. This could be boiled by raising the temperature to achieve a point on the liquid-vapor line.
13. Identify the principal type of solute-solvent interaction that is responsible for forming the following solutions: (a) KNO_3 in water; (b) Br_2 in benzene, C_6H_6 ; (c) glycerol, $\text{CH}_2(\text{OH})\text{CH}(\text{OH})\text{CH}_2\text{OH}$, in water; (d) HCl in acetonitrile, CH_3CN [HCl does not form ions in CH_3CN].
- (a) ion-dipole (b) London dispersion (c) hydrogen bonding (d) dipole-dipole

14. For the following carboxylic acids, predict whether solubility will be greater in water or carbon tetrachloride, and give your reasoning: (a) acetic acid, $\text{CH}_3\text{CO}_2\text{H}$, (b) stearic acid, $\text{CH}_3(\text{CH}_2)_{16}\text{CO}_2\text{H}$.
- (a) Acetic acid's -OH groups make hydrogen bonding possible, which is compatible with solvent water. Carbon tetrachloride has only London dispersion forces, which are less compatible. Therefore, acetic acid is more soluble in water than carbon tetrachloride.
- (b) Stearic acid has a very long chain and much higher London dispersion forces than acetic acid. This makes it more compatible with carbon tetrachloride, despite the potential for hydrogen bonding (which is largely mitigated by the long chain getting in the way). Therefore, stearic acid is more soluble in carbon tetrachloride than water.
15. Hexane (C_6H_{14}) and heptane (C_7H_{16}) are miscible in all proportions with $\Delta H_{\text{soln}} \approx 0$.
- (a) Why are these two liquids miscible with each other?
- They have very similar London dispersion forces, owing to their similar molar masses.
- (b) Why is $\Delta H_{\text{soln}} \approx 0$ for this pair of liquids?
- The intermolecular forces of attraction in the neat liquids are so similar to each other that little change occurs on mixing. It is the change in intermolecular attraction strength that is principally responsible for the sense and magnitude of ΔH_{soln} .
- (b) Why do they spontaneously form solutions, given that $\Delta H_{\text{soln}} \approx 0$?
- Mixing is a more disordered state than exists in the separate neat liquids. The increase in entropy is the driving factor in making solution formation spontaneous in this case.
16. The solubility of N_2 at $p(\text{N}_2) = 1 \text{ atm}$ is $1.75 \times 10^{-3} \text{ g/100 mL}$ of water. What is the solubility in water at an air pressure of 2.51 atm, the pressure at 50 ft below the surface of the water? Air is 78.1 vol-% N_2 . [Hint: What is the partial pressure of $\text{N}_2(\text{g})$ when the air pressure is 2.51 atm?]

From Dalton's Law of Partial Pressures, at air pressure of 2.51 atm, the partial pressure of N_2 is

$$P_{\text{N}_2} = (0.781)(2.51 \text{ atm}) = 1.96 \text{ atm}$$

From Henry's Law, the solubility is

$$C_{\text{N}_2} = \left(\frac{1.75 \times 10^{-3} \text{ g/100 mL}}{\text{atm}} \right) (1.96 \text{ atm}) = 3.43 \times 10^{-3} \text{ g/100 mL}$$

17. Calculate the molality of ethanol, C₂H₅OH (m.w. = 46.06) in a solution prepared by dissolving 5.00 g of ethanol in 25.00 g of water.

$$m = \left(\frac{5.00 \text{ g ethanol}}{25.00 \text{ g H}_2\text{O}} \right) \left(\frac{\text{mol ethanol}}{46.06 \text{ g ethanol}} \right) \left(\frac{1000 \text{ g H}_2\text{O}}{\text{kg H}_2\text{O}} \right) = \frac{4.34 \text{ mol ethanol}}{\text{kg H}_2\text{O}} = 4.34 \text{ } m$$

18. Calculate the total molality of all ions in a solution prepared by dissolving 20.0 g of (NH₄)₂SO₄ in 95.0 g of water. [f.w. (NH₄)₂SO₄ = 132 u]



$$\text{mol } (\text{NH}_4)_2\text{SO}_4 = (20.0 \text{ g } (\text{NH}_4)_2\text{SO}_4) \left(\frac{\text{mol } (\text{NH}_4)_2\text{SO}_4}{132 \text{ g } (\text{NH}_4)_2\text{SO}_4} \right) = 0.151_{515} \text{ mol } (\text{NH}_4)_2\text{SO}_4$$

$$\text{mol ions} = (0.1515 \text{ mol } (\text{NH}_4)_2\text{SO}_4) \left(\frac{3 \text{ mol ions}}{\text{mol } (\text{NH}_4)_2\text{SO}_4} \right) = 0.454_{545} \text{ mol ions}$$

$$m_{\text{ions}} = \left(\frac{0.4545 \text{ mol ions}}{95 \text{ g H}_2\text{O}} \right) \left(\frac{1000 \text{ g H}_2\text{O}}{\text{kg H}_2\text{O}} \right) = 4.78 \text{ } m$$

19. Consider a 2.00 *m* solution of sugar in water at 25.00 °C.

- (a) What is the value of the mole fraction of water in this solution? [Hint: Imagine that the solution was made up with exactly 1 Kg of water.] (m.w. H₂O = 18.02 u)

To calculate mole fraction, we need the numbers of moles of water and of sugar. If we assume that exactly 1000 g of water were used, then we already know that the solution contains 2.00 moles of sugar. All we need is the number of moles in 1000 g of water, and then we can calculate $\chi(\text{H}_2\text{O})$.

$$\text{mol sugar} = 2.00 \text{ mol sugar}$$

$$\text{mol H}_2\text{O} = (1000 \text{ g H}_2\text{O})(1 \text{ mole H}_2\text{O}/18.02 \text{ g H}_2\text{O}) = 55.49 \text{ mol H}_2\text{O}$$

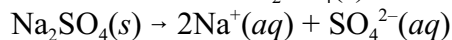
$$\chi(\text{H}_2\text{O}) = 55.49 \text{ mol}/(55.49 + 2.00) \text{ mol} = 55.49/57.49 = 0.9652$$

- (b) Calculate the vapor pressure above a 2.00 *m* solution of sugar in water at 25.00 °C, given that the vapor pressure of pure water at this temperature is 23.76 mm Hg.

$$P_{\text{soln}} = \chi(\text{H}_2\text{O}) \times P^\circ(\text{H}_2\text{O}) = 0.9652 \times 23.76 \text{ mm Hg} = 22.93 \text{ mm Hg}$$

20. Calculate the expected vapor pressure above a 2.00 *m* solution of Na₂SO₄ in water at 25.00 °C. Compare this result to what you found in part a of the preceding Key Question.

We must take account of the dissociation of Na₂SO₄(s):



Again, assume a solution prepared with exactly 1000 g of water.

$$\text{mol ions} = 3 \times \text{mol Na}_2\text{SO}_4 = 3 \times 2.00 \text{ mol} = 6.00 \text{ mol}$$

$$\chi(\text{H}_2\text{O}) = 55.49 \text{ mol} / (55.49 + 6.00) \text{ mol} = 55.49 / 61.49 = 0.9024$$

$$P_{\text{soln}} = \chi(\text{H}_2\text{O}) \times P^\circ(\text{H}_2\text{O}) = 0.9024 \times 23.76 \text{ mm Hg} = 21.44 \text{ mm Hg}$$

The amount by which the vapor pressure of water has been lowered is almost three times greater in the Na₂SO₄ solution, compared to the sugar solution.

21. What are the partial pressures and total vapor pressure above a solution at 20.0 °C made by mixing 12.5 g benzene (C₆H₆) with 44.2 g toluene (C₆H₅CH₃). At 20.0 °C, $P^\circ(\text{C}_6\text{H}_6) = 74.7$ torr and $P^\circ(\text{C}_6\text{H}_5\text{CH}_3) = 22.3$ torr. [m.w. C₆H₆ = 78.11; m.w. C₆H₅CH₃ = 92.14]

$$\text{mol C}_6\text{H}_6 = (12.5 \text{ g C}_6\text{H}_6) / (\text{mol C}_6\text{H}_6 / 78.11 \text{ g C}_6\text{H}_6) = 0.160 \text{ mol C}_6\text{H}_6$$

$$\text{mol C}_6\text{H}_5\text{CH}_3 = (44.2 \text{ g C}_6\text{H}_5\text{CH}_3) / (\text{mol C}_6\text{H}_5\text{CH}_3 / 92.14 \text{ g C}_6\text{H}_5\text{CH}_3) = 0.480 \text{ mol C}_6\text{H}_5\text{CH}_3$$

$$\text{total moles} = 0.160 \text{ mol} + 0.480 \text{ mol} = 0.640 \text{ mol}$$

$$\chi(\text{C}_6\text{H}_6, \text{soln}) = 0.160 \text{ mol} / 0.640 \text{ mol} = 0.250$$

$$\chi(\text{C}_6\text{H}_5\text{CH}_3, \text{soln}) = 1 - 0.250 = 0.750$$

$$P(\text{C}_6\text{H}_6) = (0.250)(74.5 \text{ torr}) = 18.7 \text{ torr} \quad P(\text{C}_6\text{H}_5\text{CH}_3) = (0.750)(22.3 \text{ torr}) = 16.7 \text{ torr}$$

$$P_t = (18.7 + 16.7) \text{ torr} = 35.4 \text{ torr}$$

22. In terms of mole fractions, what is the composition of the vapor above the previously described benzene-toluene mixture?

We previously found

$$P(\text{C}_6\text{H}_6) = (0.250)(74.5 \text{ torr}) = 18.7 \text{ torr}$$

$$P(\text{C}_6\text{H}_5\text{CH}_3) = (0.750)(22.3 \text{ torr}) = 16.7 \text{ torr}$$

$$P_t = (18.7 + 16.7) \text{ torr} = 35.4 \text{ torr}$$

Using these values, we can calculate the mole fractions in the vapor as follows:

$$\chi(\text{C}_6\text{H}_6, \text{vap}) = 18.7 \text{ torr}/35.4 \text{ torr} = 0.528$$

$$\chi(\text{C}_6\text{H}_5\text{CH}_3, \text{vap}) = 16.7 \text{ torr}/35.4 \text{ torr} = 1 - 0.528 = 0.473$$

Notice that more volatile benzene has an increased mole fraction in the vapor.

23. Pure benzene has a freezing point of 5.5 °C and a boiling point of 80.1 °C. What are the expected freezing point and boiling point for a 0.15 *m* solution of a nonvolatile solute in benzene? For benzene, $K_f = 5.12 \text{ °C}/m$ and $K_b = 2.53 \text{ °C}/m$.

$$\Delta T_f = (5.12 \text{ °C}/m)(0.15 m) = 0.77 \text{ °C} \quad T_f^{\text{soln}} = T_f - \Delta T_f = (5.5 - 0.77) \text{ °C} = 4.7 \text{ °C}$$

$$\Delta T_b = (2.53 \text{ °C}/m)(0.15 m) = 0.38 \text{ °C} \quad T_b^{\text{soln}} = T_b + \Delta T_b = (80.1 + 0.38) \text{ °C} = 80.5 \text{ °C}$$

24. When 45.0 g of an unknown nonvolatile nonelectrolyte is dissolved in 500.0 g of water, the resulting solution freezes at -0.930 °C. What is the molar mass of the unknown substance? $K_f = 1.86 \text{ °C}/m$ for water.

$$\Delta T_f = 0.930 \text{ °C} \quad m = \Delta T_f/K_f = 0.930 \text{ °C}/1.86 \text{ °C}/m = 0.500 m = 0.500 \text{ mol X}/\text{kg H}_2\text{O}$$

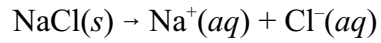
$$\text{m.w.} = \left(\frac{45.0 \text{ g X}}{500.0 \text{ g H}_2\text{O}} \right) \left(\frac{10^3 \text{ g H}_2\text{O}}{0.500 \text{ mol X}} \right) = 180 \text{ g X}/\text{mol X}$$

25. What is the osmotic pressure of a 0.100 M glucose solution in torr at 25.0 °C?

$$\begin{aligned} \pi &= MRT = (0.100 \text{ mol/L})(0.0821 \text{ L}\cdot\text{atm}/\text{K}\cdot\text{mol})(298 \text{ K}) = 2.45 \text{ atm} \times (760 \text{ torr}\cdot\text{atm}^{-1}) \\ &= 1860 \text{ torr} \end{aligned}$$

26. Sea water is approximately 0.60 M NaCl. What is the minimum applied pressure that must be exceeded to achieve water purification by reverse osmosis at 25 °C?

We must use the molarity of ions, not the stated molarity of NaCl (analytical concentration of NaCl).



$$\text{molarity of ions} = 2 \times C_{\text{NaCl}} = 2 \times 0.60 \text{ M} = 1.20 \text{ M}$$

$$\pi = MRT = (1.20 \text{ mol/L})(0.0821 \text{ L}\cdot\text{atm/K}\cdot\text{mol})(298 \text{ K}) = 29.3_6 \text{ atm} = 29.4 \text{ atm}$$