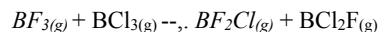


Multiple-Choice Questions

- 19.35 ΔG°_f refers to the free-energy change associated with which of the following?
- the formation of 1 mol of a compound from its constituent elements
 - the decomposition of 1 mole of a compound into its constituent elements
 - the freezing of 1 mol of a substance
 - the melting of 1 mol of a substance
 - all of the above
- 19.36 Under what conditions can we absolutely say a system is at equilibrium at constant pressure and temperature?
- $\Delta H = 0$
 - $\Delta H > 0$
 - $\Delta G = 0$
 - $\Delta G > 0$
 - $\Delta S > 0$
- 19.37 Which of the following substances would you expect to have the highest entropy at 1 atm of pressure?
- $\text{H}_2\text{O}_{(s)}$ (1 mol)
 - $\text{H}_2\text{O}_{(l)}$ (1 mol)
 - $\text{H}_2\text{O}_{(g)}$ (1 mol)
 - 0.5 mol of ethanol in 0.5 mol of $\text{H}_2\text{O}_{(l)}$
 - a gaseous solution of 0.5 mol of $\text{H}_2\text{O}_{(g)}$ and 0.5 mol of $\text{CH}_4_{(g)}$
- 19.38 For any *isolated* spontaneous process that is endothermic, which of the following conditions are true: (1) $\Delta G_{\text{system}} < 0$; (2) $\Delta G_{\text{system}} > 0$; (3) $\Delta S_{\text{system}} < 0$; (4) $\Delta S_{\text{system}} > 0$; (5) $\Delta S_{\text{universe}} < 0$; (6) $\Delta S_{\text{universe}} > 0$?
- (1), (3), and (4)
 - (1), (4), and (6)
 - (2), (4), and (6)
 - (2), (4), and (6)
 - (1), (4), and (5)
- 19.39 For which of the following changes would you expect ΔS to be less than zero?
- $\text{H}_2\text{O}_{(l)} \rightarrow \text{H}_2\text{O}_{(g)}$
 - $\text{CO}_{(g)} + \text{O}_{2(g)} \rightarrow \text{CO}_{2(g)}$
 - $2\text{SO}_{3(g)} \rightarrow 2\text{SO}_{2(g)} + \text{O}_{2(g)}$
 - $\text{CO}_{(s)} \rightarrow \text{CO}_{2(g)}$
 - $2\text{AgCl}_{(s)} \rightarrow 2\text{Ag}_{(s)} + \text{Cl}_{2(g)}$
- 19.40 When KNO_3 dissolves in water at room temperature, ΔH is positive for the dissolution process. Given this information, what can you conclude?
- $\Delta G > 0$ for the dissolution process.
 - $\Delta G = 0$ for the dissolution process.
 - The dissolving of salts in water is always a spontaneous process.
 - $\Delta S > 0$ for the dissolution process.
 - $\Delta S < 0$ for the dissolution process.
- 19.41 Which of the following is true for the sublimation of solid CO_2 at room temperature and pressure?
- ΔH and ΔS are positive.
 - ΔH and ΔS are negative.
 - ΔH is negative, and ΔS is positive.
 - ΔG becomes more positive as temperature increases.
 - ΔG is positive.
- 19.42 Which of the following can be used to calculate directly the value of an equilibrium constant for a reaction at a given temperature?
- ΔH
 - H°
 - ΔS
 - ΔG
 - ΔG°
- 19.43 Calculate ΔG° for the following reaction at 25°C
- $$4\text{NO}_{(g)} + 6\text{H}_2\text{O}_{(g)} \rightarrow 5\text{O}_{2(g)} + 4\text{NH}_{3(g)}$$
- given the following ΔG°_f values: NO , 86.71 kJ/mol; $\text{H}_2\text{O}_{(g)}$, -228.61 kJ/mol; O_2 , 0.0 kJ/mol; NH_3 , -16.66 kJ/mol.
- 958.2 kJ
 - 2145.14 kJ
 - 2145.14 kJ
 - 1318.18 kJ
 - 1318.18 kJ
- 19.44 Estimate the normal boiling point for HCl, given the following information:
- $$\Delta H_{\text{vaporization}} = 16.13 \text{ kJ/mol} \quad \Delta S = 85.77 \text{ J/mol}$$
- 188°C
 - 188 K
 - 0.189°C
 - 0.189 K
 - 100°C

CH 19 - CHALLENGE PROBS

19.31 The bond enthalpies for B-F and B-Cl are 645 kJ/mol and 445 kJ/mol, respectively. (a) Would you expect the following reaction to occur based on its enthalpy of reaction?



(b) Given that the reaction in (a) does tend to occur, and given your answer in (a), explain this tendency.

19.33 Diamond and graphite are allotropic forms of carbon. The standard enthalpies of formation for diamond and graphite are 1.88 kJ/mol and a kJ/mol, respectively; the standard molar entropies are 2.43 J/mol-K and 5.69 J/mol-K, respectively. (a) What does the term allotropic mean? (b) Which allotropic form exists in a more "disordered" state? (c) Can the transformation of diamond to graphite at 298 K and 1 atm pressure occur spontaneously? Does this agree with your knowledge of diamonds?

19.34 In Chapter 13 of the text, you learned about colligative properties. One of the topics was boiling point elevation: The addition of a nonvolatile solute to a volatile solvent raises the boiling point of the solvent (the escaping tendency of solvent molecules decreases). If a liquid and vapor are at equilibrium, what is true about their free energies? If a nonvolatile solute is added to the liquid, what happens to the free-energy of the liquid state? Given your conclusion, why must the temperature of the solution be raised to restore equilibrium?

Practice 19.2 (pg 2 of 2) Thermodynamics

ANSWERS

19.35 (a). 19.36 (c). 19.37 (e). A mixture of gases shows the greatest random motion of particles. 19.38 (d). 19.39 (b) A mixture of gases with 1~ particles reacting yields a single gas with fewer particles of gas, one particle. 19.40 (d). 19.41 (a). Sublimation of CO₂ is a spontaneous process at room temperature and pressure. This process requires energy ($\Delta H > 0$), and a solid-to-gas conversion always shows $\Delta S > 0$. 19.42 (e). 19.43 (a). 19.44 (b).

19.31. (a) The enthalpy of reaction can be estimated by summing the bond enthalpies of the reactants and subtracting the bond enthalpies of the product. However, since all bonds in both reactants and products are either B-F or B-Cl, the number and types of bonds does not change. This should result in a enthalpy of reaction that is zero (or very close to zero since bond enthalpies are slightly different depending on the environment in the molecule). In any event the enthalpy of reaction is not a significant driving force for the reaction. The problem says that the reaction tends to occur, which means $\Delta G_{rx} < 0$. Thus, the $T\Delta S$ term in $\Delta G = \Delta H - T\Delta S$ must be positive; therefore, ΔS must also be positive. Inspection of the reaction shows that the bonds in the products are arranged in a more random manner than those in the reactants. Thus, this supports the idea that ΔS is indeed greater than zero and therefore the entropy change drives the reaction.

19.33 (a) Allotropes are different forms of the same element existing in the same physical state. (b) Graphite should exist in a more disordered solid state form than diamond because it has a larger standard molar entropy value. Diamond is a covalently bonded giant molecule with each carbon atom bonded to four others in an extensive network. Graphite has carbon atoms bonded to three other carbon atoms to form a sheet of hexagonal units bonded together. The sheets are able to slide with respect to each other. (c) To answer this question you need to know the value of ΔG° for one mole of diamond converting to one mole of graphite.

$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ solve for ΔG° after calculating ΔH° and ΔS° values from products minus reactants, use T as 298 K

$$\Delta G^\circ = -2.85 \text{ kJ}$$

The conversion should be spontaneous as the sign of the free-energy change is negative. Our experience with diamonds should tell us that we do not observe this transformation. Diamond is considered a durable gem. This is an example of the difference between thermodynamics and kinetics.

Although the application of thermodynamic principles tells us the transformation is spontaneous, our observations should convince us that the kinetics of the transformation are extremely slow.

19.34 If a liquid and vapor are in equilibrium, then $\Delta G^\circ = 0$ and their free energies must be the same, that is $G(\text{liquid}) = G(\text{vapor})$. When a nonvolatile solute is added to the liquid, the concentration of the solvent decreases and the escaping tendency of the solvent molecules decreases. This means that $G(\text{liquid}) < G(\text{vapor})$. If the pure liquid was at the boiling point, this is no longer the case. The temperature must be raised again to have the two free energies be equivalent: $\Delta G = \Delta G^\circ + RT \ln Q$. This raises the boiling point.