Q18.1a

Explain how you know if each of the following change in states result in an increase or decrease in entropy.

a. solid to liquid
b. liquid to gas
c. gas to liquid

S18.1a

a. Increase in entropy. Liquids have more entropy than solids do.
b. Increase in entropy. Gas has more entropy than liquids do.
c. Decrease in entropy. Liquids have less entropy than gases.

Q18.1b

Determine whether each of the following is an increase or decrease in entropy.

a. fusion
b. freezing
c. combustion

S18.1b

a. increase
b. decrease
c. increase

Q18.2a

Which of the following changes in states represent the smallest Delta S. Which one has the largest Delta S? How do you know?
a. freezing of water to ice
b. melting of ice to liquid water
c. sublimation of ice to gas

S18.2a

The freezing of water is a decrease in entropy so it has the smallest Delta S because solids have less entropy than liquids. The other two options represent an increase in Entropy. The sublimation of ice is a large increase in Entropy because gas has more Entropy than solids. The sublimation of ice to gas is more of an increase than the melting of ice to water because gas has more entropy than liquid and therefore the change must have a larger Delta S.

Q18.2c

Arrange the compounds in order of increasing entropy (\(\Delta S\)) and justify your order.

- \(\text{CH}_3\text{OH}_{(s)}\)
- \(\text{CH}_3\text{OH}_{(l)}\)
- \(\text{CH}_3\text{OH}_{(g)}\)
- \(\text{CH}_3\text{CH}_2\text{OH}_{(g)}\)

A- \(\text{CH}_3\text{OH}_{(s)}<\text{CH}_3\text{OH}_{(l)}<\text{CH}_3\text{OH}_{(g)}<\text{CH}_3\text{CH}_2\text{OH}_{(g)}\)

The bigger the compound, the greater the entropy as there are more degrees of freedom.

Q18.2d

Given the reaction of diamond converting to graphite

\[2\text{C}_{(s \; \text{diamond})} \rightarrow 2 \text{C}_{(s \; \text{graphite})}\]

Determine \(\Delta G\) at 298 K and determine if this reaction is spontaneous or not. What does \(\Delta G\) say about the rate of this reaction?

- \(\Delta H^\circ_{f} (\text{C}_{(s \; \text{diamond})})=1.9; \; \text{kJ/mol}\)
- \(S^\circ (\text{C}_{(s \; \text{diamond})})=2.38; \; \text{J/(mol \; K)}\)
- \(S^\circ (\text{C}_{(s \; \text{graphite})})=5.74; \; \text{J/(mol \; K)}\)

\[\Delta G=\Delta H-T\Delta S\]
\[ \Delta H = (2 \text{mol C(s)}_{\text{graphite}} \times \Delta H^\circ_{\text{C(s) graphite}}) - (2 \text{mol C(s)}_{\text{diamond}} \times \Delta H^\circ_{\text{C(s) diamond}}) \]
\[ \Delta H = 2(0) - 2(1.9 \text{ kJ/mol}) = -3.8 \text{ kJ} \]

\[ \Delta S = (2 \text{mol C(s)}_{\text{graphite}} \times S_{\text{C(s) graphite}}) - (2 \text{mol C(s)}_{\text{diamond}} \times S_{\text{C(s) diamond}}) \]
\[ \Delta S = 2(5.74 \text{ J/(mol K)}) - 2(2.38 \text{ J/(mol K)}) = +6.72 \text{ J/mol} \]
\[ \Delta G = -3.8 \times 10^3 \text{ J} - 298.15(6.72 \text{ J/mol K}) = -5.51 \text{ kJ} \]

This reaction is spontaneous. \( \Delta G \) says nothing about the kinetics; while this reaction is thermodynamics favored, it occurs on a 100 million year timescale. Diamonds are kinetically meta-stable materials.

Q18.2

Given these reactions reactions, determine whether the \( \Delta S \) increases or decreases or stays the same at 25 °C;

a. \( \text{H}_2\text{SO}_4 (l, \; 1 \text{ atm}) \rightarrow \text{H}_2\text{SO}_4 (s) \; (1 \text{ atm}) \)
b. \( \text{H}_2\text{O} (l, \; 1 \text{ atm}) \rightarrow \text{H}_2\text{O} (g, \; 1 \text{ atm}) \)
c. \( \text{CH}_4 (g) + 2\text{O}_2 (g) \rightleftharpoons \text{CO}_2 (g) + 2\text{H}_2\text{O} (l) \)

S18.2

a. decreases because of a phase change from a liquid to a solid.
b. increases because of a phase change from a liquid to a gas.
c. decreases because there are more moles of gas in the reactants.

Q18.3

"Spontaneous reactions are faster than non spontaneous reactions". Access this statement and explain why it is true or false.

S18.3

This statement is false because spontaneity does not determine speed; there are spontaneous reactions that range from very slow like melting an ice cube in cold water, to very fast like melting an ice cube in hot water.

Q18.25a

Given the dissolving of \( \text{CaCl}_2 (s) \)
\[ \text{CaCl}_2 (s) \rightarrow \text{Ca}^{2+}_{(aq)} + 2\text{Cl}^-_{(aq)} \]

- the \( \Delta G^\circ_{\text{f}} \) of \( \text{CaCl}_2 (s) \) = -748.1 kJ/mol,
• the \(\Delta H\) of \(\text{CaCl}_2\) = -795.8 kJ
• the \(\Delta \text{S}^\circ\)
  - \(\text{Ca}^{2+}\) = -53.1 J/Kmol,
  - \(\text{CaCl}_2\) = 104.6 J/Kmol,
  - \(\text{Cl}^-\) = 56.5

Calculate the temperature of this reaction.

**S18.25a**

\[
\Delta G^\circ = \Delta H - T \Delta S
\]

\[
T = \frac{\Delta H - \Delta G^\circ}{\Delta S}
\]

\[
T = (0 - [-795.8]) - (-748.1)) \left(2 \times 56.5\right) - (-53.1) \times 104.6
\]

\[
T = -34.539 \text{ K}
\]

**Q18.25b**

At what temperature would the following reaction occur spontaneously? (Delta H\(^\circ\) = 1256.4 kJ; Delta S\(^\circ\) = 587 J/K)

**S18.25b**

\[
2\text{Cr}_2\text{O}_3 \rightarrow 4\text{Cr} + 3\text{O}_2
\]

\[
\Delta G^\circ = \Delta H^\circ - T \Delta S^\circ = 0
\]

\[
1256.4 \text{kJ} = T \times 0.587 \text{kJ/K} \quad T = 2140.37 \text{K}
\]

**Q18.27a**

Calculate \(\Delta G^\circ\) for

\[
\text{CS}_{(l)} + 2 \text{O}_{(g)} \rightarrow \text{CO}_{(g)} + 2 \text{SO}_{(g)}
\]

using only the following \(\Delta G^\circ\) values for the reaction:

- \(\text{C}(_{(s)}) + \text{O}_{(g)} \rightarrow \text{CO}_{(g)}\) with \(\Delta G^\circ = -394.39\; \text{kJ/mol}\)
- \(\text{S}(_{(s)}) + \text{O}_{(g)} \rightarrow \text{SO}_{(g)}\) with \(\Delta G^\circ = -300.13\; \text{kJ/mol}\)
- \(\text{C}(_{(s)}) + 2\text{S}(_{(s)}) \rightarrow \text{CS}_{(l)}\) with \(\Delta G^\circ = 67.1\; \text{kJ/mol}\)
S18.27a

\[
\begin{align*}
A\text{-} \text{C(s)} + \text{O}_2 \text{(g)} &\rightarrow \text{CO}_2 \text{(g)}; \Delta G = -394.39 \text{ kJ/mol} \\
2[\text{S(s)} + \text{O}_2 \text{(g)} &\rightarrow \text{SO}_2 \text{(g)}]; \Delta G = 2[-300.13 \text{ kJ/mol}] \\
\text{CS}_2 \text{(l)} &\rightarrow \text{C(s)} + 2\text{S(s)}; \Delta G = -(67.1) \text{ kJ/mol}
\end{align*}
\]

\[
\Delta G = -394.39 \text{ kJ/mol} + 2[-300.13 \text{ kJ/mol}] + -(67.1) \text{ kJ/mol} = -1061.75
\]

Q18.27b

Use the given standard Gibbs energy changes for these equations:

a. \(2\text{Fe}_2\text{O}_3 \text{(s)} \rightarrow 4\text{Fe}(\text{s}) + 3\text{O}_2 \text{(g)} \Delta \text{G}^\circ = -742.2 \text{ kJ/mol}\)

b. \(\text{Fe(OH)}_3 \text{(s)} \rightarrow 3\text{Fe}(\text{s}) + \text{OH}^- \text{(g)} \Delta \text{G}^\circ = -696.5 \text{ kJ/mol}\)

c. \(\text{Fe}_3\text{O}_4 \text{(s)} \rightarrow 3\text{Fe}(\text{s}) + 2\text{O}_2 \text{(g)} \Delta \text{G}^\circ = -1015 \text{ kJ/mol}\)

Identify the \(\Delta \text{G}^\circ\) for the following reaction
\[
\text{[6Fe}_2\text{O}_3\text{(s)} \rightarrow \text{O}_2 \text{(g)} + 4\text{Fe}_3\text{O}_4\text{(s)}]]
\]

S18.27b

\[
\begin{align*}
6\text{Fe}_2\text{O}_3(\text{s}) &\rightarrow \text{O}_2(\text{g}) + 4\text{Fe}_3\text{O}_4(\text{s}) \\
\text{Fe}_3\text{O}_4(\text{s}) &\rightarrow 3\text{Fe}(\text{s}) + 2\text{O}_2(\text{g}) \Delta \text{G}^\circ = 4(-1015) \text{ kJ/mol} \\
2\text{Fe}_2\text{O}_3(\text{s}) &\rightarrow 4\text{Fe}(\text{s}) + 3\text{O}_2(\text{g}) \text{ -3(-742.2) kJ/mol}
\end{align*}
\]

net: \(6\text{Fe}_2\text{O}_3(\text{s}) \rightarrow \text{O}_2(\text{g}) + 4\text{Fe}_3\text{O}_4(\text{s}) \Delta \text{G}^\circ = (-3(-742.2)) + 4(-1015) = 3837.4 \text{ kJ/mol}\)

Q18.27c

Find \(\Delta \text{G}^\circ\) at 298.15 K for reaction
\[
\text{[2CO + O}_2 \rightarrow 2\text{CO}_2]\]

\[
\text{DH}^\circ = -128.3 \text{ kJ DS}^\circ = - 159.5 \text{ J K}^{-1}
\]

S18.27c

\[
\text{DG}^\circ = \text{DH}^\circ - \text{TDG}^\circ
\]
Delta G° = (-128.3 kJ) – [298.15K x (-159.5 JK⁻¹ x (1 J/1000kJ))] = -80.75 kJ

Q18.29a

Write a the chemical equation for the complete combustion of methane, \(\text{CH}_4\) gas and calculate \(\Delta G°\) at 298.15 K.

\[
[5\text{CH}_4(g) + 5\text{O}_2(g) \rightleftharpoons 5\text{CO}_2(g) + \text{H}_2\text{O}(g)]
\
\]

From Tables T1 or T2

- \(\Delta G°\) of \(\text{CO}_2(g)\) is -393.3 kJ/mol
- \(\Delta G°\) of \(\text{H}_2\text{O}(l)\) is -228.6 kJ/mol

So the standard products minus reactions approach:

\[
\Delta G° = (5\times(-394.4) + (-228.6)) - (5\times(-50.72)) \Delta G° = -1947 \text{ kJ}
\]

Q18.29

Solve for \(\Delta G°\) at 298 K for the combustion of \(\text{B}_2\text{H}_6(g)\) in \(\text{B}_2\text{O}_3(s)\) and \(\text{H}_2\text{O}(l)\).

First the balanced reaction must be identified

\[
[\text{B}_2\text{H}_6(g) + 3\text{O}_2(g) \rightleftharpoons \text{B}_2\text{O}_3(s) + 3\text{H}_2\text{O}(l)]
\
\]

Using the parameters from Table T1 or T2

\[
\Delta G° = [3\times(-237.1) + (-1194)] - [86.7+3\times0] = -11,992; \text{ kJ/mol}
\]

The reaction is (strongly) spontaneous since \(\Delta G° < 0\) and essentially goes to completion (irreversible reaction).

Q18.31a

Assess the feasibility of the reaction below by evaluating \((\Delta S°_{rxn}), (\Delta H°_{rxn})\) and \((\Delta G°_{rxn})\) for this reaction at 25°C. If the reaction is spontaneous, is it enthalpically favored, entropically favored or both?

\[
[(\text{NH})_4\text{Cl}_{(s)} \rightarrow \text{NH}_{3(g)} + \text{HCl}_{(g)}]
\
\]

\[
\Delta S° = \sum(\Delta S°_{products}) - \sum(\Delta S°_{reactants})
\
\]

S18.31a
\[(192.5 + 186.9) - (94.6) = 284.8 \text{ kJ/mol K}\]

\[\Delta H^o_{\text{rxn}} = \sum(\Delta H^o) \text{ (products)} - \sum(\Delta H^o) \text{ (reactants)}\]

\[(-46.11 + -92.31) - (-314.4) = 175.98 \text{ kJ/mol} \]

The Gibbs energy can be also tabulated,

\[\Delta G^o_{\text{rxn}} = \sum(\Delta G^o) \text{ (products)} - \sum(\Delta G^o) \text{ (reactants)}\]

but can be calculate directly from \(\Delta S^o_{\text{rxn}}\) and \(\Delta H^o_{\text{rxn}}\)

\[\Delta G^o_{\text{rxn}} = \Delta H^o_{\text{rxn}} - T\Delta S^o_{\text{rxn}}\]

\[\Delta G^o_{\text{rxn}} = 175.98 \text{ kJ/mol} - 298 \text{ K} (284.8) = -84,737 \text{ kJ/mol}\]

The thermal decomposition of \((\text{NH})_4\text{Cl}_{(s)}\) is a strongly spontaneous process that is entropically driven, but not enthalpically driven.

---

**Q18.31a**

Given the reaction at 298 K;

\[4\text{NH}_{3\text{(g)}} + 5\text{O}_{2\text{(g)}} \rightarrow 4\text{NO}_{(g)} + 6\text{H}_2\text{O}_{\text{(l)}}\]

Determine the S

---

**S18.31a**

\[\Delta S^o = \Delta S^o_{\text{products}} - \Delta S^o_{\text{reactants}}\]

\[\Delta S^o = (69.91 + 210.8) - (111.3)\]

\[\Delta S^o = 169.41 \text{ J/molK}\]

---

**Q18.31b**

Assess the feasibility of the reaction:

\[\text{ClNO}_2 \text{ (g)} + \text{NO} \text{ (g)} \rightarrow \text{NO}_2 \text{ (g)} + \text{ClNO} \text{ (g)}\]

by calculating each of the following for the reaction at 25 °C.

a. \(\Delta S^o\) (standard molar entropy for \(\text{ClNO}_2 \text{ (g)}\) is 272.23 J/mol*K; entropy for \(\text{NO} \text{ (g)}\) is 210.8 J/mol*K; entropy for \(\text{NO}_2 \text{ (g)}\) is 240.1 J/mol*K; entropy for \(\text{ClNO} \text{ (g)}\) is 261.58 J/mol*K.)

b. \(\Delta H^f\) (\(\Delta H^f\) for \(\text{ClNO}_2 \text{ (g)}\) is 12.5 kJ/mol; \(\Delta H^f\) for \(\text{NO} \text{ (g)}\) is 90.25 kJ/mol; \(\Delta H^f\) for \(\text{NO}_2 \text{ (g)}\) is 33.18 kJ/mol; \(\Delta H^f\) for \(\text{ClNO} \text{ (g)}\) is 51.71 kJ/mol.)
c. $\Delta G^\circ$

S18.31b

a) $\Delta S^\circ = (240.1 + 261.58) - (272.23 + 210.8) = 18.65 \text{ J/mol} \cdot \text{K}$

b) $\Delta H^\circ = (33.18 + 51.71) - (12.5 + 90.25) = -17.86 \text{ kJ/mol}$

c) $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = -17860 \text{ J/mol} - (298.15 \text{ K})(18.65 \text{ J/mol} \cdot \text{K}) = -23420 \text{ J/mol} = -23.42 \text{ kJ/mol}$

$\Delta G^\circ < 0$... therefore... reaction is spontaneous.

Q18.37

Determine the $K_p$ at 298.15 K for the reaction:

\[
2\text{CO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{CO}_2(\text{g})
\]

S18.37

\[
\Delta G^\circ = \sum (\Delta G^\circ)_{\text{products}} - \sum (\Delta G^\circ)_{\text{reactants}}
\]

\[
(2\times-137.2) - (2\times-394.4) = 514.4 \text{ kJ/mol}
\]

\[
\ln K_p = \frac{\Delta G^\circ}{-RT}
\]

\[
\ln K_p = \frac{-514.4 \text{ kJ/mol}}{(8.3145\times10^{-3}\times298.15 \text{ K})} = -207.5
\]

\[
K_p = e^{-207.5} = (7.65\times10^{-91})
\]

Q18.37

Given the reaction at 298 K and the thermodynamic values in Table T1

\[
2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})
\]

estimate $K_p$.

S18.37

\[
\Delta G^\circ = (2\times-371.10) - (2\times-300.2) = -141.8 \text{ kJ/mol}
\]
\[
\ln K_p = -(-141.8)/((8.3145E-3)(298))\]

\[
\ln K_p = 57.23
\]

\[
K_p = e^{57.23} = 7.156 \times 10^{24}
\]

**Q18.39**

Determine values of \(\Delta G^\circ\) and \(K\) at 298.15 K for the following reaction:

\[
(2\text{SO}_2 (g) + \text{O}_2 (g) \rightleftharpoons 2\text{SO}_3 (g))
\]

\[
\Delta G^\circ = \sum (\Delta G^\circ)_{\text{products}} - \sum (\Delta G^\circ)_{\text{reactants}}
\]

\[
(2\times(-371.1)) - (2\times - 300.2) = -141.8 \text{ kJ/mol}
\]

\[
\ln K = -\Delta G/RT = ((141.8\times10)^3 \text{ J/mol})/((8.3145 \text{ J/mol})(298.15 \text{ K})) = 57.2
\]

\[
K = e^{57.2} = (6.94\times10)^{24}
\]

**Q18.39**

Given the reaction at 298 K and the thermodynamic values in Table T1

\[
4\text{NH}_3(g) + 5\text{O}_2(g) \rightleftharpoons 4\text{NO(g)} + 6\text{H}_2\text{O(l)}
\]

solve for \(K_p\).

\[
\Delta G^\circ = (6(-237.6) + (4(86.55)) - (4(-16.45))
\]

\[
= -1010.6 \text{kJ/mol}
\]

\[
= -RT\ln K_p
\]

\[
\ln K_p = -(-1010.6)/((8.3145E-3)(298))
\]

\[
K_p = e^{407.87}
\]

**Video Solution**
Q18.43

\[ \Delta G^\circ \text{ for the reaction below is } 31.4 \text{ kJ/mol at 298.15 K. Use the thermodynamics quantity to decide in which direction the reaction is spontaneous when the concentrations of } \{H_2SO_{4(aq)}\}, \{HSO^-_{4(aq)}\}, \text{ and } \{H_3O^+_{(aq)}\} \text{ are } 0.2 \text{ M, } 0.1 \text{ M, and } 0.1 \text{ M respectively.} \]

\[ \{H_2SO_{4 (aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HSO^-_{4(aq)}\} \]

\[ Q_C = \frac{[H_3O^+] [HSO_4^-]}{[H_2SO_4]} = \frac{(0.1 \text{ M})^2}{0.2 \text{ M}} = 0.05 \]

\[ \Delta G = 31.4 \text{ kJ} + ((8.3145 \times 10)^3 \text{ kJ/(mol*K)}) \times (298.15 \text{ K}) \times \ln(0.05) = 23.97 \text{ kJ} \]

Because \( \Delta G \) is positive, the reaction is not spontaneous going in the forward direction under these conditions.

S18.43

\[ \{Q_C= \frac{[H_3O^+] [HSO_4^-]}{[H_2SO_4]} \} = \frac{(0.1 \text{ M})^2}{0.2 \text{ M}} = 0.05 \]

\[ [\Delta G = 31.4 \text{ kJ} + ((8.3145 \times 10)^3 \times \text{kJ/(mol*K)}) \times (298.15 \text{ K}) \times \ln 0.05 = 23.97 \text{ kJ}] \]

Because \( \Delta G \) is positive, the reaction is not spontaneous going in the forward direction under these conditions.

Q18.43

Given the ionization of ethanol in water at 298 K;

\[ \{CH_3COH_{(g)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + CH_3CO^-_{(aq)}\} \]

when the \( \Delta G^\circ = 28 \text{ kJ/mol} \) and the concentration of CH3CHO(g)=0.1 M, H3O+(aq)= 0.09129 M CH3CO-(aq)=0.09129 M consider in which direction the spontaneous reaction will go.

S18.43

\[ Q_c=0.091292(0.1-0.09129) \]

\[ Q_c=0.949 \]

\[ \Delta G = 28 + (8.314E-3)(298)(\ln 0.949) = 27.87 \]

spontaneous rxn happens in backward direction .

Q18.47

Kp=0.12520.3352 Kp=0.139

\[ \Delta G^\circ 1000k=-RT\ln K =-(8.3145J/molK)(1000K)\ln 0.139 =1.64 \text{ kJ/mol} \]

\[ Q_p=(0.032)(0.022)(0.071)(0.095) \]

\[ Q_p=0.104 \]

Because \( Q_p<K_p \), the spontaneous reaction will go to the right to form more products.

S18.45a

Kp can be calculated by using the partial pressures of each element in the system. However, since partial pressures are
not given, one can use the mole fractions of products over reactants, which in this case is the same as moles of products over reactants. This is because A, B, C, and D are all divided by the same total moles.

\[
K_p = \frac{(0.6)(0.6)}{(0.4)(0.4)} = 2.25
\]

\[\Delta G^\circ_{500K} = -RT\ln K_p = -3.371 \text{ KJ}\]

\[Q_p = (0.2 \times 0.017)/(0.01 \times 0.3) = 1.133 < 2.25, \text{ which implies that the reaction will proceed to the right to reach equilibrium.}\]

Q18.45b

For the reaction

\[\text{NH}_3\text{ (aq)} + \text{H}_2\text{O (l)} \rightleftharpoons \text{NH}_4\text{ (aq)}^+ + \text{OH}^-\text{(aq)}\]

which equation is invalid.

a. \(\Delta G = \Delta G^\circ + RT\ln Q\)
b. \(K = K_p\)
c. \(K_p = e^{\Delta G^\circ/RT}\)

S18.45b

C is incorrect because the correct equation is \(G = -RT\ln K_p\), thus correct format would be \(K_p = e^{-G/RT}\)

Q18.45c

For reactions involving only gases, which of the following statements is false?

a. \(\Delta G^\circ 1/(-RT) = \ln k\)
b. \(T\Delta S = \Delta H^\circ - \Delta G\)
c. \(K_c = ([\text{products}])/([\text{reactants}] )\)
d. \(K = e^{\Delta G^\circ 1/RT}\)

Q18.45c

D is incorrect because the exponent, \(\Delta G^\circ 1/RT\), is supposed to be multiplied by negative 1.

Q18.47

In the reaction at 1000K:

\[\text{NO}_2\text{(g)} + \text{H}_2\text{(g)} \rightleftharpoons \text{NO}_\text{(g)} + \text{H}_2\text{O}_\text{(g)}\]
calculate $K_p$ when there are 0.335 mol NO$_2$(g), 0.335 mol H$_2$(g), 0.125 mol NO(g), and 0.125 mol H$_2$O(g). Solve for $\Delta G^*$ at 1000K. Evaluate in what direction the spontaneous reaction will go in if there are 0.071 mol NO$_2$(g), 0.095 mol H$_2$(g), 0.032 mol NO(g), and 0.022 mol H$_2$O(g).

Q18.47

The imaginary reaction $A + B= C + D$ is at equilibrium at 500 K when there is .4 mol A, .4 mol B, .6 mol C, and .6 mol D.

a. Calculate $K_p$ at 500 K.

b. Calculate $\Delta G^*$ at 500K

c. What direction would a spontaneous reaction occur if the following were the new number of moles present in the system; .01 mol A, .3 mol B, 0.2 mol C, .017 mol D?

Q18.49a

Calculate $\Delta G^*$ for each of the following equilibrium reactions at the given temperature. (Kc values not verified for accuracy. Just for practice purposes).

a. $2H_2S(g)+ CH_4(g)= 4H_2(g) +Cs_2(g)$ Kc=5.27X10$^{-8}$ at 400K

b. $SO_3(g) + NO(g)\rightarrow SO_2(g) + NO_2(g)$ Kc= 2.4X10$^{-6}$ at 300 K

c. $C_3H_8(g)+CH_4(g)\rightarrow 2C_2H_6(g)$ Kc= 3.6X10$^{-4}$ 550 K

S18.49a

$\Delta G^*= -RT \ln K_p$

a. $\Delta G^*=-(8.3145)(400\;K) \ln(5.27 \times 10^{-8})= 55.7 \;kJ/mol$

b. $\Delta G^*=-(8.3145)(300\;K) \ln(2.4 \times 10^{-6})=32.2 \;kJ/mol$

c. $\Delta G^*=-(8.3145)(550 \;K) \ln(3.6 \times 10^{-4})= 36.3 \;kJ/mol$

Q18.49

Given the following equations solve for the

a. $2SO_2(g)+2H_2(g)\rightarrow 2SO_3(g)$ 125 C and $\langle K_c=28 \rangle$

b. $4NH_3(g)+5O_2(g)\rightarrow 4NO(g)+6H_2O(l)$, 332 C and $\langle K_c=32 \rangle$

S18.49

a. $\Delta G^*=-(8.3145E-3 \;kJ/mol\;K)(125+273K)(\ln28)=-11.02 \;kJ/mol$

b. $\Delta G^*=-(8.3145E-3 \;kJ/mol\;K)(332+273K)(\ln32)=-17.43 \;kJ/mol$
Q18.53A

Consider the following imaginary reaction:

\[ A(s) + B(g) \rightleftharpoons C(s) \]

At 30 degrees, $\Delta G^\circ = 40.00$ KJ

Prove that the partial pressure of $B(g)$ is very large.

What conditions could be changed to decrease the $B(g)$ partial pressure.

S18.53

$\Delta G^\circ = -RT \ln K_p$

$40$ KJ = $-0.0083145 \times 303 \times \ln K_p$

$K_p = e^{-15.8774}$

Partial pressure of $B(g) = 1/(e^{-15.8774})$ which is a very large number.

To decrease the partial pressure of $B(g)$, one must make $K_p$ larger which can occur by decreasing the temperature.

Q18.55

Use data from Appendix D to calculate at 298 K for the following reaction:

$H_2O_2(g) + CH_4(g) = CH_3OH (g) + H_2O(g)$

$\Delta S^\circ \ b) \ \Delta H^\circ \ c) \ \Delta G^\circ \ d) \ K$

S18.55

$\Delta S^\circ = 239.8 \text{ J} + 188.8 \text{ J} - 186.3 \text{ J} + 188.8 \text{ J} = 53.5 \text{ J}$

$\Delta H^\circ = -200.7 - 241.8 + 136.3 + 74.81 = -231.39 \text{ KJ}$

$\Delta G^\circ = -162 - 228.6 + 105.6 + 50.72 = -234.28 \text{ KJ}$

$\Delta G^\circ = -RT \ln K_p$

$-234.28 \text{ KJ} = -0.0083145 \times 298 \times \ln K$

$K = e^{944.555}$
Q18.57A

What is the temperature of the following reaction if $\Delta G^\circ=-53.2 \text{ kJ}$, $\Delta H^\circ=-21.5 \text{ kJ}$, and $\Delta S^\circ=14.7 \text{ J/K}$?

$$[2\text{NO(g)}+\text{O}_2 \text{(g)} \rightarrow (2\text{NO})_2 \text{(g)}]$$

S18.57A

In this problem we are asked to calculate the temperature of the given reaction. This can be done after rearranging $\Delta G^\circ=\Delta H^\circ-T\Delta S^\circ$ in order to find $T$.

$$\Delta G^\circ=\Delta H^\circ-T\Delta S^\circ$$

$$T\Delta S^\circ=\Delta H^\circ-\Delta G^\circ$$

$$T=(\Delta H^\circ-\Delta G^\circ)/(\Delta S^\circ)$$

Find $T$:

$$T=(-21.5 \times (10)^3 \text{ J}-(-53.2 \times (10)^3 \text{ J})/((14.7 \text{ J})/\text{K})=2.6 \times (10)^3 \text{ K}$$

Q18.57

What must be the temperature (K) if the following reaction has $\Delta G^\circ= -38.5 \text{ kJ}$, $\Delta H^\circ= -14.6 \text{ kJ}$, and $\Delta S^\circ= 0.152 \text{ kJ/ K}$?

$\text{Fe}_2\text{O}_3 \text{(s)} + 3\text{CO (g)} \rightarrow 2\text{Fe (s)} + 3\text{CO}_2 \text{(g)}$

S18.57

$$\Delta G^\circ=\Delta H^\circ-T\Delta S^\circ$$

$$(-38.5 \text{ kJ}) = (-14.6 \text{ kJ}) - T(0.152 \text{ kJ/K})$$

$$-157.24 = -T$$

$$T=157.24 \text{ K}$$

Q18.53

For the decomposition reaction

$$[2\text{MgO}_{(s)} \rightleftharpoons 2\text{Mg}_{(s)}+\text{O}_{2\;(g)}]$$

where at 298 K

- $\Delta H = -601.7 \text{ kJ/mol}$ and
\( G = -569.4 \text{ kJ/mol} \)
solve for \( K_p \).

**Q18.55**

Using the equation FeCO₃(s) Fe(s)+CO₃(aq), solve for H & S at 298K.

**Q18.54**

Determine the equilibrium constant at 298 K for the following reaction:

\[
\text{Mg(OH)₂(s)} + 2\text{H}^+(aq) \rightleftharpoons \text{Mg}^{2+}(aq) + 2\text{H}_2\text{O(l)}
\]

- \( \Delta G°: [\text{H}_2\text{O(l)}] = -237.1 \text{ kJ/mol}, \)
- \( \Delta G°: [\text{Mg}^{2+}(aq)] = -454.8 \text{ kJ/mol}, \)
- \( \Delta G°: [\text{Mg(OH)₂(s)}] = -833.5 \text{ kJ/mol} \)

\[\Delta G° = [2[\text{H}_2\text{O(l)}] + [\text{Mg}^{2+}(aq)] - [\text{Mg(OH)₂(s)}] \]
\[\Delta G° = [2[-237.1 \text{ kJ/mol}] + [-454.8 \text{ kJ/mol}] - [-833.5 \text{ kJ/mol}] \]
\[\Delta G° = -95.5 \text{ kJ/mol} = -95.5 \times 10^3 \text{ J/mol} \]
\[\Delta G° = -RT\ln K \]
\[\ln K = -\Delta G°/ RT = -(-95.5 \times 10^3 \text{ J/mol})/(8.3145 \text{ J/mol K} \times 298 \text{K}) = 38.5 \]
\[K = e^{38.5} = 5 \times 10^{16} \]

**S18.54**

\[\Delta G° = 2[\text{H}_2\text{O(l)}] + [\text{Mg}^{2+}(aq)] - [\text{Mg(OH)₂(s)}] \]
\[\Delta G° = 2[-237.1 \text{ kJ/mol}] + [-454.8 \text{ kJ/mol}] - [-833.5 \text{ kJ/mol}] \]
\[\Delta G° = -95.5 \text{ kJ/mol} = -95.5 \times 10^3 \text{ J/mol} \]
\[\Delta G° = -RT\ln K \]
\[\ln K = -\Delta G°/ RT = -(-95.5 \times 10^3 \text{ J/mol})/(8.3145 \text{ J/mol K} \times 298 \text{K}) = 38.5 \]
\[K = e^{38.5} = 5 \times 10^{16} \]

**Q57A**

What must the temperature be if the following reaction has \( \Delta G° = -53.2 \text{ kJ}, \Delta H° = -21.5 \text{ kJ}, \text{and } \Delta S° = 14.7 \text{J/K?} \)

\[\text{2NO}_{(g)} + O_{(2 \text{ g})} \rightarrow (2\text{NO})_{2 \text{ (g)}} \]

**S57A**

In this problem we are asked to calculate the temperature of the given reaction. This can be done after rearranging \( \Delta G° = \Delta H° - T\Delta S° \) in order to find \( T \).

\[\Delta G° = \Delta H° - T\Delta S° \]
The synthesis of water occurs by the reaction $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(l)$ at 25°C. Using the following information, and assuming that $\Delta H^*$ and $\Delta S^*$ are essentially unchanged, estimate $K_p$ at 25°C.

\[
\begin{align*}
\Delta H^*(\text{H}^+) &= 0; \quad \Delta H^*(\text{OH}^-) = -230.0 \text{ kJ/mol}; \quad \Delta H^*(\text{H}_2\text{O}) = -285.8 \text{ kJ/mol} \\
\Delta S^*(\text{H}^+) &= 0; \quad \Delta S^*(\text{OH}^-) = -10.75 \text{ J/(mol K)}; \quad \Delta S^*(\text{H}_2\text{O}) = 69.91 \text{ J/(mol K)}
\end{align*}
\]

First, determine the value of $\Delta G^*$ at 25°C, from the $\Delta H^*$ and $\Delta S^*$ values.

\[
\begin{align*}
\Delta H^* &= \Delta(H^*)_f [\text{H}_2\text{O}(l)] - \Delta(H^*_f [\text{H}^+(aq)] - \Delta(H^*_f [(\text{OH})^- (aq)] \\
&= (-285.8 \text{ kJ/mol}) - (0) - (-230.0 \text{ kJ/mol}) = -55.8 \text{ kJ/mol H}^+ \\
\Delta S^* &= S^*[\text{H}_2\text{O}(l)] - S^*[\text{H}^+(aq)] - S^*[\text{OH}^-(aq)] \\
&= (69.91 \text{ J/(mol K)}) - (0) - (-10.75 \text{ J/(mol K)}) = 80.66 \text{ J/(mol K)}
\end{align*}
\]

\[
\Delta G^* = \Delta H^* - T\Delta S^* = -55.8 \text{ kJ/mol -298Kx}(0.08066 \text{ kJ/(mol K)}) = -79.84 \text{ kJ/mol} = -RT \ln (K_p) \\
\ln (K_p) = (-\Delta G^*)/RT = (79.84 \times 10^3 \text{ J/mol})/(8.3145 J/(\text{mol K}) \times 298 K) = 32.2; \quad K_p = e^{32.2} = 9.6 \times 10^{13}
\]
\[ (-285.8 \text{ kJ/mol}) - (0) - (-230.0 \text{ kJ/mol}) = -55.8 \text{ kJ/mol} \]
\[ \Delta S^\circ = S^\circ \text{[H}_2 \text{O(l)]} - S^\circ \text{[H}^+ \text{(aq)]} - S^\circ \text{[(OH)}^- \text{(aq)]} \]
\[ = (69.91 \text{ J/(mol K)}) - (0) - (10.75 \text{ J/(mol K)}) = 80.66 \text{ J/(mol K)} \]
\[ \Delta G^\circ = \Delta H^\circ - T \Delta S^\circ = -55.8 \text{ kJ/mol} - 298 \text{K} \times (0.08066 \text{ kJ/(mol K)}) = -79.84 \text{ kJ/mol} = -RT \ln (K_p) \]
\[ \ln (K_p) = (-\Delta G^\circ) / RT = (79.84 \times (10)^3 \text{ J/mol}) / (8.3145 \text{ J/(mol K)} \times 298 \text{ K}) = 32.2 \]
\[ K_p = e^{32.2} = 9.6 \times (10)^{13} \]

**Q61**

Determine the temperature at which the \( K_p = 1.0 \times (10)^5 \) for the reaction \( \text{[N}_2 (\text{g})] + 3\text{[H}_2 (\text{g})] \rightarrow 2\text{[NH}_3 (\text{g})] \). Using the following information:

- \( \Delta H^\circ (\text{N}_2) = 0 \) ; \( 3\Delta H^\circ (\text{H}_2) = 0 \) ; \( 2\Delta H^\circ (\text{NH}_3) = -46.11 \text{ kJ/mol} \)
- \( \Delta S^\circ (\text{N}_2) = 191.6 \text{ J/(mol K)} \) ; \( 3\Delta S^\circ (\text{H}_2) = 130.7 \text{ J/(mol K)} \) ; \( 2\Delta S^\circ (\text{NH}_3) = 192.5 \text{ J/(mol K)} \)

\[ \Delta H^\circ = 2\Delta (H^\circ)_f \text{[(NH}_3 (\text{g})]} - \Delta (H^\circ)_f \text{[N}_2 (\text{g})] - 3\Delta (H^\circ)_f \text{[H}_2 (\text{g})] \]
\[ = 2(-46.11 \text{ kJ/mol}) - (0) - (0) = -92.22 \text{ kJ/mol} \]
\[ \Delta S^\circ = 2\Delta S^\circ \text{[(NH}_3 (\text{g})]} - \Delta S^\circ \text{[N}_2 (\text{g})] - 3\Delta S^\circ \text{[H}_2 (\text{g})] \]
\[ = 2(192.5 \text{ J/(mol K)}) - (191.6 \text{ J/(mol K)}) - 3(130.7 \text{ J/(mol K)}) = -198.7 \text{ J/(mol K)} \]
\[ \Delta G^\circ = \Delta H^\circ - T \Delta S^\circ = -RT \ln K \]
\[ \Delta H^\circ - T \Delta S^\circ = RT \ln K \]
\[ T = \Delta H^\circ / (\Delta S^\circ - R \ln K) \]
\[ T = (-92.22 \times (10)^3 \text{ J/mol}) / (-198.7 \text{ J/(mol K)} - 8.3145 \text{ J/(mol K) \ln (1.0 \times (10)^5)}) = 313 \text{ K} \]

**Q61**

Determine the temperature at which the \( (K_p = 1.0 \times 10^5) \) for the reaction:

\[ \text{[N}_2 (\text{g})] + 3\text{[H}_2 (\text{g})] \rightleftharpoons 2\text{[NH}_3 (\text{g})] \]

Using the following information:

- \( \Delta H^\circ (\text{N}_2) = 0 \) ; \( 3\Delta H^\circ (\text{H}_2) = 0 \) ; \( 2\Delta H^\circ (\text{NH}_3) = -46.11 \text{ kJ/mol} \)
- \( \Delta S^\circ (\text{N}_2) = 191.6 \text{ J/(mol K)} \) ; \( 3\Delta S^\circ (\text{H}_2) = 130.7 \text{ J/(mol K)} \) ; \( 2\Delta S^\circ (\text{NH}_3) = 192.5 \text{ J/(mol K)} \)
\[ \Delta H^\circ = 2\Delta (H^\circ)_{f\ [(NH)_3\ (g)]} - \Delta (H^\circ)_{f\ [N_2\ (g)]} - 3\Delta (H^\circ)_{f\ [H_2\ (g)]} = 2(-46.11\ \text{kJ/mol}) - (0) - (0) = -92.22\ \text{kJ/mol} \]

\[ \Delta S^\circ = 2S^\circ[(NH)_3\ (g)] - S^\circ[N_2\ (g)] - 3S^\circ[H_2\ (g)] = 2(192.5\ \text{J/(mol K)}) - (191.6\ \text{J/(mol K)}) - 3(130.7\ \text{J/(mol K)}) = -198.7\ \text{J/(mol K)} \]

\[ \Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = -RT\ln K \]

\[ \Delta H^\circ - T\Delta S^\circ - RT\ln K \]

\[ T = \frac{\Delta H^\circ}{\Delta S^\circ - R\ln K} \]

\[ T = \frac{-92.22 \times (10)^3\ \text{J/mol}}{-198.7\ \text{J/(mol K)} - 8.3145\ \text{J/(mol K)} \ln (1.0 \times (10)^5)} = 313\ \text{K} \]

**Q61**

Kp = 1e6 for the following reaction:

\[ 2\text{SO}_2\ (g) + \text{O}_2\ (g) \leftrightarrow 2\text{SO}_3\ (g) \]

Determine the temperature at which this Kp occurs.

T1 = ?, K1 = 1e6; T2 = 800K, K2 = 9.1e2, and \( \Delta H^\circ = -1.8e5\ \text{J/mol} \)

\[ \ln \left( \frac{K_2}{K_1} \right) = \frac{\Delta H^\circ}{R} \left( \frac{1}{T_1} \right) \left( \frac{1}{T_2} \right) \]

\[ \ln \left( \frac{9.1e2}{1e6} \right) = -1.8e5/ 8.3145\left( \frac{1}{T_1}\right)\left( \frac{1}{800K} \right) \]

\[ 1/T_1 = (3.2e-4) + (1.25e-3) = 1.57e-3 \]

\[ T_1 = 1/1.57e-3 = 6.37e2\ \text{K} \]

**Q63A**

For the reaction \( \text{C}_6\text{H}_6(g) \rightleftharpoons 3\text{C}_2\text{H}_2(g) \), \( \Delta H^\circ = 66.3\ \text{kJ/mol} \) and Kp = 0.214 at 298K.

a. What is Kp at 0°C?

b. At what temperature will Kp = 2.50?
a) 
\[ \ln \frac{K_2}{K_1} = \frac{\Delta H^\circ}{R} \left( \frac{1}{T_1} - \frac{1}{T_2} \right) = \frac{66.3 \times 10^3 \text{ J/mol}}{8.3145 \text{ J/(mol K)}} \left( \frac{1}{298} - \frac{1}{273} \right) = -2.45 \]

\[ \frac{K_2}{K_1} = e^{-2.45} = 0.0863 \]

\[ K_2 = 0.0863 \times 0.214 = 0.018 \text{ at 273K} \]

b) 
\[ \ln \frac{K_2}{K_1} = \frac{\Delta H^\circ}{R} \left( \frac{1}{T_1} - \frac{1}{273} \right) = \frac{66.3 \times 10^3 \text{ J/mol}}{8.3145 \text{ J/(mol K)}} \left( \frac{1}{T_1} - \frac{1}{298} \right) = \ln 0.214 / 2.50 = -2.458 \]

\[ \frac{1}{T_1} - \frac{1}{298} = -2.458 \times 8.3145 \times \left( \frac{1}{66.3 \times 10^3} \right) \]

\[ K^{-1} = -3.08 \times 10^{-4} \text{ K}^{-1} \]

\[ 1/T_1 = 1/298 - 3.08 \times 10^{-4} = 3.05 \times 10^{-3} - 3.08 \times 10^{-4} = 2.74 \times 10^{-3} \text{ K}^{-1}; T_1 = 365 \text{ K} \]

**Q63**

At what temperature does this reaction become spontaneous? (\(\Delta H^\circ = 1118.4 \text{ kJ}, \Delta S^\circ = 347.2 \text{ J/K}\))

\[
\text{\[Fe}_3\text{O}_4(\text{s}) \rightarrow 3\text{Fe}_\text{(s)} + 2\text{O}_2(\text{g})\]
\]

\[ \Delta G^\circ = \Delta H^\circ - T \Delta S^\circ = 0 \]

\[-1118.4 \text{ kJ} = -T(0.3472 \text{ kJ/K}) \]

\[ T = 3221 \text{ K} = 2948 \text{°C} \]

**Q71**

When referencing a table, what would you predict the normal boiling point of mercury and the vapor pressure of mercury is at 25°C?

After referencing a table with both phases of mercury as gas and liquid, we can conclude that the normal boiling point of liquid mercury (Hg) is 356.73°C = 629.88 K, and vapor pressure of mercury is about 0.002 mmHg at room temperature.

**Q74**

While mercury is useful in barometers, mercury vapor can be toxic. Given that mercury has Hvap of 56.9 kJ/mol and its normal boiling point is 356.7°C, calculate the vapor pressure in mmHg at room temperature, 25°C.

**S18.74**

\[ \Delta Hvap / R \times (1/T1) - (1/T2) \ln P \]
\[
\left(\frac{56.9 \text{e3 J/mol}}{8.314 \text{ J/mol K}}\right) \times \left(\frac{1}{298} - \frac{1}{629.7 \text{K}}\right) = \ln(12.10) = e^{\ln(12.10)} = 179,433 \text{ mmHg}
\]

**Q83**

Consider the following reaction at 25°C:

\[
\text{H}_2\text{O}(g) \rightleftharpoons \text{H}_2\text{O}(l)
\]

\[\Delta G_f^\circ (\text{kJ/mol}): \text{H}_2\text{O}(l) = -237.129, \text{H}_2\text{O}(g) = -228.572\]

\[\Delta H_f^\circ (\text{kJ/mol}): \text{H}_2\text{O}(l) = -285.83, \text{H}_2\text{O}(g) = -241.818\]

\[S^\circ (\text{J/mol K}): \text{H}_2\text{O}(l) = 69.91, \text{H}_2\text{O}(g) = 188.825\]

Calculate the \(K_{eq}\) for the given reaction.

**S18.83**

\[\Delta H^\circ = [-285.83] - [-241.818] = -44.012 \text{kJ/mol}\]

\[\Delta S^\circ = [69.91] - [188.825] = -118.915 \text{J/mol K}\]

\[\Delta G^\circ = \Delta H^\circ - T \Delta S^\circ = -44012 \text{ J/mol} - (298 \text{K})(-118.915 \text{ J/mol K}) = -8575 \text{ J/mol} = -8.575 \text{ kJ/mol}\]

\[\Delta G^\circ = -RT\ln K_{eq}\]

\[\ln K = -\Delta G^\circ / RT = -(-8557 \text{J/mol}) / (8.3145 \text{ J/mol K} \times 298 \text{K}) = 38.5\]

\[K = e^{38.5} = 31.6\]

**Q88**

Calculate \(\Delta G^\circ\) (in kJ) for the following reaction in which occurs at 25.0°C.

\[
\text{Fe}_3\text{O}_4(s) \rightarrow 3\text{Fe(s)} + 2\text{O}_2(g)
\]

\[\Delta H_f (\text{kJ/mol}): \text{Fe}_3\text{O}_4(s) = -1118.4\]

\[S^\circ (\text{J/mol K}): \text{Fe}_3\text{O}_4(s) = 146.4, \text{Fe(s)}= 27.8, \text{O}_2 (g) = 205.1\]

\[\Delta H^\circ = [0] - [-1118.4] = 1118.4 \text{ kJ/mol}\]

\[\Delta S^\circ = [3(27.8) + 2(205.1)] - [146.4] = 347.2 \text{ J/mol K}\]
ΔG°=ΔH°-T ΔS° = 1118.4 kJ - (298K)(0.3472 kJ/mol K) = 1015 kJ/ mol

Q104
If ΔG°> 0 for a reaction, it must also be true that: ________________________________

S18.104
Forward reaction is nonspontaneous at standard state and Keq<1.

Q1
Indicate whether each of the following changes represents an increase or decrease in entropy in a system, and explain your reasoning;

a. melting ice
b. burning gasoline
c. freezing liquid bromine.

S1
a. increases entropy, molecular bonds breaking -> greater disorder
b. increases entropy, molecules and compounds released in the burning of anything
c. decreases entropy, frozen molecules are more ordered

Q2
Which of the following processes has the highest ΔS at 25C:

\[\text{CO}_2(\text{s, 1 atm}) \rightarrow \text{CO}_2(\text{g, 10\;mmHg})\]
on\[\text{H}_2\text{O}_2(\text{s, 1atm}) \rightarrow \text{H}_2\text{O}_2(\text{g, 10\;mmHg})\]

S2
Converting CO2(s, 1 atm)--CO2(g, 10 mmHg) would have a much higher ΔS because it takes much more energy to convert a substance directly from a solid state to a gaseous state.
Q3  
Why is it incorrect to say that the entropy of the world decreases and the total energy of the world fluctuates?

S3  
First law of thermodynamics states that energy is neither created nor destroyed, so there is a fixed amount of energy in the universe at a given time. Also, the entropy of the universe increases for all spontaneous, naturally occurring processes, so the entropy of the universe is boundless.

Q4  
Calculate the \(\Delta{H}\) and \(\Delta{S}\) for the following reaction at 25C:

\[
4\text{NO}_\text{(g)} \rightarrow 2\text{N}_2\text{O}_\text{(g)} + \text{O}_\text{2(g)}
\]

S3  
From Table T1, the following properties are identified:

- \(\text{NO}_\text{(g)}\): \(\Delta{H}=90.9\ \text{kJ/mol}\), \(S=210.76\ \text{kJ/mol}\)
- \(\text{N}_2\text{O}_\text{(g)}\): \(\Delta{H}=82.05\ \text{kJ/mol}\), \(S=219.85\ \text{kJ/mol}\)
- \(\text{O}_\text{2(g)}\): \(\Delta{H}=0\ \text{kJ/mol}\), \(S=205.14\ \text{kJ/mol}\)

\[
\Delta{H} = \text{Sum of } \Delta{H} \text{ (products)} - \text{Sum of } \Delta{H} \text{ (reactants)}
\]

\[
\Delta{S} = \text{Sum of } \Delta{S} \text{ (products)} - \text{Sum of } \Delta{S} \text{ (reactants)}
\]

\[
\Delta{H} = [(2\times82.05 \text{ kJ/mol}) + (0 \text{ kJ/mol})] - [(4\times90.9 \text{ kJ/mol})] = -199.5 \text{ kJ/mol}
\]

\[
\Delta{S} = [(2\times219.85 \text{ kJ/mol}) + (205.14 \text{ kJ/mol})] - [(4\times210.76 \text{ kJ/mol})] = -198.2 \text{ kJ/mol}
\]

Q5  
Determine \(\Delta{G}\) at 265.25 K for the reaction:

\[
2\text{NO}_\text{(g)} + \text{O}_\text{2(g)} \rightarrow 2\text{NO}_\text{2(g)}
\]

S5  
From Table T1, the following properties are identified:

- \(\Delta{H}= -114.1 \text{ kJ}\)
- \(\Delta{S}= -146.5 \text{ J/K}\).
\[ \Delta G = \Delta H - T \Delta S \]

\[ \Delta G = (-114.1 \ \text{kJ}) - (265.25 \ \text{K})(0.1465 \ \text{kJ/K}) \]

\[ \Delta G = -153.0 \ \text{kJ} \]

**Q6**

What is the Gibbs energy change \( \Delta G \) for the following reaction at 25°C:

\[ \text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3 \]

**S6**

From Table T1, the following properties are identified:

- \( \Delta H = -91.8 \ \text{kJ} \) and
- \( \Delta S = -198.0 \ \text{J/K} \).

\[ \Delta G = \Delta H - T \Delta S \]

\[ \Delta G = (-91.8 \ \text{kJ}) - (298 \ \text{K})(0.1980 \ \text{kJ/K}) = -32.8 \ \text{kJ} \]

**Q7**

Which of the following is spontaneous?

a. \( \Delta H > 0 \) and \( T \Delta S > 0 \)

b. \( \Delta H < 0 \) and \( T \Delta S > 0 \)

c. \( \Delta H > 0 \) and \( T \Delta S < 0 \)

d. \( \Delta H < 0 \) and \( T \Delta S < 0 \)

**S7**

*Answer: B*

**Q8**

At what temperatures are the following reaction spontaneous:

\[ \text{Br}_2(\ell) \rightarrow \text{Br}_2(g) \]

**S8**

- \( \Delta H = 30.91 \ \text{kJ/mol} \),
For the reaction to be spontaneous:

\[
\Delta G = \Delta H - T \Delta S < 0
\]

\[
(30.91 \text{ kJ/mol}) - T(0.0932 \text{ J/molK}) < 0
\]

\[T = 331.6 \text{ K}\]

Q9

Determine \(\Delta G\) for the following reaction:

\[
[C_2H_4(g) + H_2O(l) \rightarrow C_2H_5OH(l)]
\]

S9

From Table T1, the following properties are identified:

\(\Delta G\) C\text{C}_2\text{H}_5\text{OH}(l) = -175 \text{ kJ/mol}\)

\(\Delta G\) C\text{H}_4(g) = 68 \text{ kJ/mol}\)

\(\Delta G\) H\text{O}(l) = -237 \text{ kJ/mol}\)

\[
\Delta G = \text{Sum of } \Delta G \text{ products} - \text{Sum of } \Delta G \text{ reactants}
\]

\[
\Delta G = -175 \text{ kJ} - 68 \text{ kJ} - (-237 \text{ kJ})
\]

\[
\Delta G = -6 \text{ kJ}
\]

\(-6 \text{ kJ} < 0\), therefore, spontaneous

Q10

Is the following reaction spontaneous under standard conditions?

\[
[4\text{KClO}_3(s) \rightarrow 3 \text{KClO}_4(s) + \text{KCl}(s)]
\]

S10

From Table T1, the following properties are identified:

- \(\Delta G\) \text{KClO}_3(s):
  - \(\Delta H_f^o = -397.7 \text{ kJ/mol}\)
  - \(S^o = 143.1 \text{ J/mol K}\)
Now calculate the enthalpy of the reaction (change of entropy) under standard conditions:

heat of reaction = 3(-432.8 kJ) + (-436.7 kJ) - 4(397.7 kJ) = -144 kJ

Now calculate the entropy of the reaction under standard conditions:

\( \Delta S = 3S(KClO_4) + S(KCl) - 4S(KClO_3) \)

\( \Delta S = 3(151.0 \text{ J/K}) + (82.6 \text{ J/K}) - 4(143.1 \text{ J/K}) = -36.8 \text{ J/K} \)

Now combine them together under standard conditions:

\[ \Delta G = \Delta H - T \Delta S \]

\( \Delta G = -144 \text{ kJ} - (298 \text{ K})(-38.6 \text{ J/K})(1 \text{ kJ/1000 J}) = -133 \text{ kJ} \)

since \( \Delta G < 0 \), the reaction is spontaneous under standard conditions.

Q11
For what temperatures will this reaction be spontaneous?

\[ 4 \text{KClO}_3(s) \rightarrow 3 \text{KClO}_4(s) + \text{KCl}(s) \]

S11
The reaction will be spontaneous when \( \Delta G < 0 \). So all temperatures \( T \) such that this inequality is satisfied:

\( \Delta G = \Delta H - T \Delta S < 0 \)

or solving for \( T \):

\[ \frac{\Delta H}{\Delta S} < T \]

\[ \frac{-144 \text{ kJ}}{-38.6 \text{ J/K}} < 3731 \text{ K} \]

Q12
The following reaction occurring at 1 atm is spontaneous at ~331 K:

\[ \text{Br}_2(l) \rightarrow \text{Br}_2(g) \]
If the temperature increased, will the reaction?

a. remain spontaneous  
b. become non-spontaneous  
c. approach equilibrium  
d. can not say

S12

*Answer: A*

Q13

Use data from Appendix D to establish for the following reaction to solve for $\Delta{G}$:

$$2N_2O_4(g) + O_2(g) \rightarrow 2N_2O_5(g)$$

S13

$$\Delta{G} = [(2 \times 115.1 \text{ kJ/mol}) - (2 \times 97.54 \text{ kJ/mol})]$$

$$\Delta{G} = 35.12 \text{ kJ/mol}$$

Q14

Using the same reaction and $\Delta{G}$ calculated from the above equation, solve for $K_p$ at 298 K.

S14

$$\Delta{G} = -RT\ln K_p$$

$$\ln K_p = \frac{\Delta{G}}{-RT}$$

$$\ln K_p = \frac{35120 \text{ J}}{(-8.314 \text{ J/mol*K} \times 298 \text{ K})}$$

$$K_p = 6.98 \times 10^{-7}$$

Q15

The standard Gibbs energy change for the reaction

$$NH_3(aq) + H_2O(l) \rightarrow NH^+(aq) + OH^-(aq)$$

is 29.05 kJ/mol at 298 K. Use this thermodynamic quantity to decide in which direction the reaction is spontaneous when
the concentrations of $\text{NH}_3(\text{aq})$, $\text{NH}_4^+(\text{aq})$, and $\text{OH}^-(\text{aq})$ are 0.10 M, $1.0 \times 10^{-3}$ M, and $1.0 \times 10^{-3}$ M, respectively.

**S15**

$$Q_c = \frac{[\text{NH}_4][\text{OH}^-]}{[\text{NH}_3]} = \frac{(1.0 \times 10^{-3} \text{M}) \times (1.0 \times 10^{-3} \text{M})}{0.10 \text{M}}$$

$$Q_c = 1.0 \times 10^{-5}$$

$$\Delta G = \Delta G^o + RT \ln Q_c$$

$$\Delta G = (29050 \text{ kJ/mol}) + (8.314 \text{ J/mol*K})(298 \text{ K})(\ln 1.0 \times 10^{-5})$$

$$\Delta G = 525.9 \text{ kJ/mol}$$

**Q16**

For a process to occur spontaneously:

- a. the entropy of the system must increase
- b. the entropy of the surroundings must increase
- c. both the entropy of the system and the entropy of the surroundings must increase
- d. the net change in entropy of the system and surroundings considered together must be a positive quantity
- e. the entropy of the universe must remain constant

**S16**

*Answer: D*

**Q17**

The Gibbs energy change of a reaction can be used to assess

- a. how much work the system does on the surroundings
- b. the net direction in which the reaction occurs to reach equilibrium
- c. how much heat absorbed from surroundings
- d. the proportion of the heat evolved in an exothermic reaction that can be converted to various forms of work

**S17**

*Answer: B*
Q18.29B

Calculate ΔG for a reaction at 400K if the Keq at that temperature is 200.

\[ \Delta G = -400(8.314) \ln(200) = -17620 \text{ J} \]

Q18.37B

Determine \( K_p \) at 298 K for the reaction CO (g) + O\(_2\) (g) \( \rightarrow \) CO\(_2\) (g)

(\( \Delta G^\circ \) for CO (g) = -137.2 kJ / mol; \( \Delta G^\circ \) (CO\(_2\)) = -394.4 kJ / mol)

\[ \Delta G^\circ = -RT \ln K_p \ln K_p = - \Delta G^\circ / RT \]

\[ \Delta G^\circ_{\text{rxn}} = -394.4 + 137.2 = -257.2 \text{ kJ / mol} \]

\[ \ln K_p = 257.2 / (8.3145)(298) K_p = 1.11 \]

Q18.39B

is ΔG positive or negative for this reaction?

\[ [\text{HCl} + \text{NaOH} \rightarrow \text{H}_2\text{O} + \text{NaCl}] \]

S18.39B

\[ \Delta H = -110.27 \text{ KJ/mol} \]

\[ \Delta G = \Delta H (- \text{ sign}) - T (\text{ always positive})( \Delta S) \]

Because we go from aq to solid in this reaction we assume that \( \Delta S \) is decreasing (increasing in order, so therefore with the given information it is impossible to know)

Q18.47B

Calculate Delta G° for the reaction at 30°C

\[ [2\text{ZnO} \rightarrow 2\text{Zn} + \text{O}_2] \]

ZnO \( \Delta H^\circ = -1034 \text{ kJ/mol} \) \( S^\circ = 145 \text{ J/molK} \)

Zn \( \Delta H^\circ = 0 \) \( S^\circ = 68 \text{ J/molK} \)
\[ O_2 \ \Delta H_f^\circ = 0 \ S^\circ = 225 \ J/molK \]

**S18.47B**

\( \Delta H^\circ = (0) - (-1034) = 1034 \text{ kJ} \)

\( \Delta S^\circ = [2(68)+1(225)] - (145) = 216 \text{ J/K} \)

\( \Delta G^\circ = 1034 \text{ kJ} - (303K)(0.216 \text{kJ/K}) = 968.55 \text{ kJ} \)

**Q18.49B**

Determine \( K_c \) or \( K_p \) for the following reactions:

**a.** \( \text{Na}^+_{(aq)} + \text{Cl}^-_{(aq)} \rightleftharpoons \text{NaCl}_{(aq)} \) with \( \Delta G^\circ = -45.2 \text{ kJ/mol} \) at \( 267 \text{ K} \)

\( \ln K_c = \frac{-\Delta G^\circ}{RT} = \frac{45.2}{(8.3145)(267)} \)

\( K_c = 1.02 \)

**b.** \( \text{CN}_{(g)} + \text{H}_{2 (g)} \rightleftharpoons \text{HCN}_{(g)} \) with \( \Delta G^\circ = -324 \text{ kJ/mol} \) at \( 312 \text{ K} \)

\( \ln K_p = \frac{-\Delta G^\circ}{RT} = \frac{324}{(8.3145)(312)} \)

\( K_p = 1.13 \)

**c.** \( \text{SO}_{2(g)} + \frac{1}{2}\text{O}_{2 (g)} \rightarrow \text{SO}_{3 (g)} \) with \( \Delta G^\circ = -215 \text{ kJ/mol} \) at \( 417 \text{ K} \)

\( \ln K_c = \frac{-\Delta G^\circ}{RT} = \frac{215}{(8.3145)(417)} \)

\( K_c = 1.06 \)

**Q18.53B**

Calculate the \( \Delta S \) for the following fake reaction:

\[ [\text{A} + \text{B} \rightarrow \text{C} + 2\text{D}] \]

\( S^\circ: 160 \text{J}, 205 \text{J}, 213.6 \text{J}, 69.9 \text{J} \)

To solve this simply do products minus reactants:

\( (2(69.9) + 213.6) - (160 + 205) = \Delta S \)

\( \Delta = -11.6 \text{ J} \)
Q55B

Find \( \Delta S^o \), \( \Delta H^o \), and \( \Delta G^o \) for the following reaction. Use Appendix D.

\[
2\text{NaOH}_{(aq)} + \text{H}_2\text{SO}_4_{(aq)} \rightarrow \text{Na}_2\text{SO}_4_{(aq)} + \text{H}_2\text{O}_{(l)}
\]

S55B

\[
\Delta S^o = 138.1 + 69.9 - (2 \times 48.1 + 20.1) = 91.7 \text{ J/mol*K}
\]

\[
\Delta H^o = -1390 + -285.8 - (2 \times -470.1 + -909.3) = 1675.8 + 1849.5 = 173.7 \text{ kJ/mol}
\]

\[
\Delta G^o = -1268 + -237.1 - (2 \times -419.2 + 744.5) = -1505.1 + 1582.9 = 77.8 \text{ kJ/mol}
\]

Q18.57B

Using an equation for Gibbs free energy, solve to express \( T \) as a function of \( \Delta G, H, S \)

\[
T = (\Delta H - \Delta G) / \Delta S
\]

Q18.59

Find \( K_{eq} \) at 25°C for the oxidation of iron (rusting).

\[
4\text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3
\]

S18.59B

\[
\Delta H^f = -921 \text{ kJ/mol}
\]

\[
\Delta S^o = 31 \text{ 210 81 J/mol K}
\]

\[
\Delta G^o = -RT \ln K
\]

\[
\Delta G^o = \Delta H^o - T \Delta S^o
\]

\[
\Delta H^o = (-(921)2) -((0)4 + (0)3) = -1842 \text{ kJ -1.842 x 10}^6 \text{ J}
\]

\[
\Delta S^o = ((31)2) -((210)3) = -592 \text{ J/K}
\]

\[
\Delta G^o = \Delta H^o - T \Delta S^o
\]

\[
=- (1.842 x 10^6) - (298)(-592 \text{ J/K})=-1665548. \text{ (spontaneous because negative)}
\]

\[
K = e^{\Delta G^o / RT} = e^{(((-1665548)J)/(8.3145)(298))}
\]
Q18.61B

Determine the temperature, in Kelvin, at which $K_p = 2.3 \times 10^{-5}$ for the following reaction, given $K_p = 1.0 \times 10^3$ at 345 K, and $\Delta H_{\text{rxn}}^\circ = 3.2 \times 10^4$ J / mol.

\[
\text{SO}_3(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_4(\text{g})
\]

\[
\ln \left( \frac{2.3 \times 10^{-5}}{1.0 \times 10^3} \right) = \frac{(3.2 \times 10^4)}{8.3145} \left[ \frac{1}{345} - \frac{1}{T_1} \right]
\]

$T_1 = 134$ K

Q18.63B

For the reaction $\text{N}_2\text{O}_4 \rightarrow 2\text{NO}_2$ and $K_p = 0.25$ at 298 K

At what temperature does $K_p = 1.00$?

S18.63B

First solve for $K_c$.

\[
0.25 = K_c (8.314 \times 298)^2
\]

\[
K_c = 4.0 \times 10^{-8}
\]

Plug in values for $K_c$ and $K_p$

\[
1 = 4.0 \times 10^{-8} (8.314 \times T)^2
\]

$T = 0.000024$ K

Q71B

Using the Clausius-Clapyron equation, calculate the boiling point of benzene for a pressure of 260 mmHg. The normal boiling point of Benzene is 80.1 degC and has a heat of vaporization of 30.72 kJ/mole

S18.71B

The Clausius-Clapeyron Equation is:

\[
\ln(P_2/P_1) = \frac{(\Delta H/R)}{(T_1-1/T_2)}
\]

Substituting:
\[ \ln\left(\frac{260}{760}\right) = \left(\frac{30720}{8.314}\right)\left(\frac{1}{T_1} - \frac{1}{(80.1+273.15)}\right) \]

Solving for \( T_1 \) gives \( T_1 = 310\,\text{degK} - 273 = 37\,\text{degC} \)

---

**Q18.74B**

Define \( a \) in the equation \( S = S_{\text{initial}} - R\ln(a) \).

\( a = \frac{\text{(the effective concentration of a substance in a system)}}{\text{(the effective concentration of the substance in a standard reference state)}} \)

---

**Q18.83B**

What temperature will the equilibrium constant for formation of \( \text{COCl}_2 \) be \( (K_p = 1.6 \times 10^4) \)?

\[ \text{CO} + \text{Cl}_2 \rightarrow \text{COCl}_2 \]

---

**Q18.88B**

Consider the rusting of iron at 298 K:

\[ 4 \text{Fe}_\text{(s)} + 3 \text{O}_\text{2(g)} \rightleftharpoons 2 \text{Fe}_2\text{O}_3\text{(s)} \]

a. calculate \( \Delta G^\circ \) for this reaction at 298 K

b. determine whether the reaction proceeds spontaneously in the forward or reverse direction.
S18.88B

a) \( \Delta S_{\text{rxn}} = 2(67.4) - 4(45.3) = -46.4 \text{ J / mol} \cdot \text{K} \)

\( \Delta G_{\text{rxn}} = -349000 \text{ J / mol} - (-46.4)(298) = -21.1 \text{ kJ / mol} \)

b) reaction is spontaneous in the forward direction.

Q18.104B

if the \( K_{eq} \) is zero, does \( \Delta G = 0? \)

S18.104B

No, \( \Delta G \) would be zero if \( K_{eq} = 1 \), the equivalence point.

Q18.3D

What is Boltzmann’s equation for entropy? Explain in detail how he derived this equation and how it shows the relationship between entropy and the number of microstates particles can occupy.

S18.3D

\( S = K \ln W. \)

\( S \) is the entropy, \( K \) is Boltzmann's constant, and \( W \) is the number of microstates. He was able to derive this equation based on the fact that the more microstates a system has, the greater the entropy; so as the \( W \) increases, as will \( S \).
Q18.37D

Use data from Appendix D of Petrucci to determine \(K_p\) at 298 K for the reaction

\[2\text{CO}(g) + \text{O}_2(g) \rightleftharpoons 2\text{CO}_2(g).\]

S18.37D

\[K_p = K_c(RT)^{\Delta n}\]

\[\Delta G^\circ = -RT\ln K_c\]

\(-128.6 = (.0083145)(298 \text{ K})\ln K_c\) (\(\Delta G\) calculated from difference of values found in Appendix D)

\(-128.6/2.477 = \ln K_c\)

\(e^{-51.90} = K_c\)

\(K_c = 2.88 \times 10^{-23}\)

\(K_p = (2.88 \times 10^{-23})(.0821)(298)(-1)\)

\(K_p = -7.14 \times 10^{-20}\)

Q18.49D

For the following equilibrium reactions, calculate \(\Delta G^\circ\) at the indicated temperature. [Hint: How is each equilibrium constant related to a thermodynamic equilibrium constant, \(K\)?]

a. \(\text{HCl}(g) + \text{O}_2(g) \rightleftharpoons \text{H}_2\text{O}(g) + \text{Cl}_2(g)\) with \(K_c = 2.3 \times 10^{-12}\) at 25°C

b. \(\text{H}_2(g) + \text{Cl}_2(g) \rightleftharpoons 2\text{HCl}(g)\) with \(K_c = 4.5 \times 10^{-7}\) at 25°C

c. \(2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g)\) with \(K_c = 1.5 \times 10^{-13}\) at 25°C

S18.49D

Use equation \(\Delta G^\circ = -RT\ln K\)

a.) \(\Delta G^\circ = -(.008314)(298 \text{ K})\ln(2.3 \times 10^{-12})\)

\(\Delta G^\circ = 66.45 \text{ kJ/mol}\)

b.) \(\Delta G^\circ = -(.008314)(298 \text{ k})\ln(4.5 \times 10^{-7})\)
\[ \Delta G^\circ = 36.207 \text{ kJ/mol} \]

c.) \[ \Delta G^\circ = (0.008314)(298 \text{ K}) \ln(1.5 \times 10^{-13}) \]

\[ \Delta G^\circ = 73.15 \text{ kJ/mol} \]

Gibbs_Free_Energy

Q18.61D

Use the equations \[ \Delta G = \Delta H - T \Delta S \] and \[ \Delta G^\circ = -RT \ln K \], and Appendix D to estimate the temperature at which \( K_p = 2.6 \times 10^3 \) for the reaction \( 2\text{SO}_2(g) + \text{O}_2(g) \leftrightarrow 2\text{SO}_3(g) \).

S18.61D

Assume \( \Delta G^\circ = -70.9 \) (from appendix D, difference of both \( \Delta G^\circ \) values)

\[-70.9 = -(0.0083145)T(\ln(2.6 \times 10^3))\]

\( T \approx -10844.4 \text{ K} \)

Gibbs_Free_Energy

Q18.88D

The standard molar entropy of solid hydrazine at its melting point of 1.53°C is 67.15 J/mol*K. The enthalpy of fusion is 12.66 kJ/mol. For N\(_2\)H\(_4\)(l) in the interval from 1.53°C to 298.15K, the molar heat capacity at constant pressure is given by the expression \( C_p = 97.8 + 0.0586(T - 280) \). Determine the standard molar entropy of \( \text{N}_2\text{H}_4(l) \) at 500 K. [Hint: The heat absorbed to produce an infinitesimal change in temperature of a substance is \( \Delta q_{rev} = (C_p)(\Delta T) \)].

S18.88D

\( C_p = 97.8 + 0.0586(500-280) \)

\( C_p = 110.7 \)

\( \Delta q_{rev} = (110.7)(23.47 \text{ K}) = 2598.13 \)

\( \Delta S = q_{rev}/T \)

\( \Delta S = 2598.13/500 \text{ K} = 5.196 \text{ J/K} \)

http://chemwiki.ucdavis.edu/Physical_Chemistry/Thermodynamics/State_Functions/Entropy
Q18.1E
Is the entropy increasing or decreasing?

a. Tear open and start using a cold pack.
b. Boiling water.
c. Freezing water.

S18.1E
(a) Increasing (b) Increasing (c) Decreasing

Q18.2E
Put them in order of increasing $\Delta S$ at 25°C.

a. $\text{CO}_2 (g, 1\text{atm}) \leftrightarrow \text{CO} (s, 1\text{atm})$
b. $\text{CO}_2 (s, 1\text{atm}) \leftrightarrow \text{CO}_2 (l, 1\text{atm})$
c. $\text{H}_2\text{O} (s, 1\text{atm}) \leftrightarrow \text{H}_2\text{O} (g, 1\text{atm})$

S18.2E
Answer: a < b < c.

Q18.27E
Calculate $\Delta G^\circ$ for the reaction

$\left[\text{COS}_2(s) + \text{H}_2\text{O} (g) \rightleftharpoons \text{CO}_2(g) + \text{H}_2\text{S} (g)\right]$

given the known thermodynamic properties of the following reactions

a. $\left[\text{CO}_2(g) + \text{H}_2(g) \rightleftharpoons \text{COS}_2(g) + \text{H}_2\text{S} (g)\right] \Delta G^\circ = 14.7 \text{kJ}$
b. $\left[\text{CO}_2(g) + \text{H}_2\text{O} (g) \rightleftharpoons \text{CO}_2(g) + \text{H}_2\text{S} (g)\right] \Delta G^\circ = -28.6 \text{kJ}$

S18.27E
$\Delta G^\circ = -27.2 \text{kJ.}$

Q18.29E
What’s $\Delta G^\circ$ For the reaction
\[2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g)\]

\[\Delta G^\circ\] of \(\text{SO}_2\) is -300kJ/mol, and its \(-371\text{ kJ/mol}\) of \(\text{SO}_3\).

S18.29E

\[\Delta G^\circ = -142\text{kJ}.\]

Q18.31E

Find \(\Delta H^\circ\) if the reaction \(\text{Fe}_3\text{O}_4(s) \leftrightarrow 3\text{Fe}(s) + 2\text{O}_2(g)\) happens spontaneously. The temperature is 2948°C, \(\Delta S^\circ = 347.2\text{ J/K}\).

S18.31E

\[\Delta H^\circ = 1118.3\text{ kJ}.\]

Q18.37E

If the Keq for the reaction \(\text{CaCO}_3(s) \leftrightarrow \text{CaO}(s) + \text{CO}_2(g)\) is \(8.1 \times 10^{-5}\), find \(\Delta G\) when \(T=25^\circ\text{C}\). Given: \([\text{CaCO}_3] = 0.003\text{M}, \quad [\text{CaO}] = 0.0004\text{M}, \quad \text{and } [\text{CO}_2]=0.0005\text{M}.

S18.37E

\[\Delta G = -482.5\text{ J}.\]

Q18.39E

What’s the \(\Delta S^\circ\) of the reaction \(2\text{NO}(g) + \text{O}_2(g) \leftrightarrow 2\text{NO}_2(g)\)? If the \(\Delta S^\circ\) is 210.8 J/K, 205.1 J/K, and 240.1 J/K respectively.

S18.39E

\[\Delta S^\circ = -146.6\text{ J/K}.\]

Q18.43E

For the reaction, \(\text{N}_2\text{O}_2(g) + \frac{1}{2} \text{O}_2(g) \leftrightarrow 2\text{NO}_2(g)\), the partial pressure are 0.1atm, 0.5atm, and 0.8atm respectively. Is the reaction spontaneous?

S18.43E

It’s nonspontaneous.
Q18.45E

For reaction $3\text{Fe}_2\text{O}_3(s) + \text{H}_2(s) \leftrightarrow 2\text{Fe}_3\text{O}_4(s) + \text{H}_2\text{O}(g)$, which of the following is correct and why?

a. $\Delta G = \Delta G^\circ + RT \ln Q$.

b. $K = k_p$.

c. $K_p = e^{\Delta G^\circ / RT}$

---

S18.45E

A

---

Q18.47E

The gas concentrations of the reaction

$[2A_{(g)} + 3B_{(g)} \rightleftharpoons 2C_{(g)} + 4D_{(g)}]$ are $[A] = 1.8 \times 10^{-2} \text{M}$, $[B] = 3.4 \times 10^{-3} \text{M}$, $[C] = 5.6 \times 10^{-1} \text{M}$, $[D] = 2.8 \times 10^{-4} \text{M}$. Assume $T = 25 \degree C$, and $\ln(K_{eq}) = 4.42 \times 10^{-9}$. Find $Q$, $\Delta G^\circ$, and $\Delta G$.

---

S18.47E

$1.51 \times 10^{-4} \text{J}$, $4.77 \times 10^{4} \text{J}$, $25.9 \text{kJ}$

---

Q18.49E

At 298K, $\text{Mg(OH)}_2(s) + 2\text{H}^+(aq) \leftrightarrow \text{Mg}^{2+}(aq) + 2\text{H}_2\text{O}(l)$ has $\Delta G^\circ = -95.5 \text{kJ/mol}$. If the $Q$ is $2.389 \times 10^{-4}$, what’s $\Delta G$?

---

S18.49E

$\Delta G = 1.97 \times 10^6 \text{kJ}$.53.

---

Q18.53E

Find the $\ln(K)\), $\ln(K_{cl})$, and $\ln(K_{eq})$ of the reaction

$[\text{Si}(s) + 2\text{Cl}_2(g) \rightleftharpoons \text{SiCl}_4(g)]$

---

S18.53E

$K = [\text{SiCl}_4]/[\text{Cl}_2]^2 = K_c = K_{sp}$. 

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Q18.55E
If $T$ has a positive value, $\Delta H$ has a negative value, $\Delta S$ has a negative value, when will $\Delta G$ be positive or negative?

S18.55E
$\Delta G$ will be positive/negative when $\Delta H$ is smaller/greater than $T \cdot \Delta S$.

Q18.57E
If $T$ has a positive value, $\Delta H$ has a negative value, $\Delta S$ has a positive value, when will $\Delta G$ be positive or negative?

S18.57E
$\Delta G$ will be positive/negative when $\Delta H$ is greater/smaller than $T \cdot \Delta S$.

Q18.59E
What’s the temperature of the reaction

$$[H_2(g) + I(g) \rightleftharpoons 2HI (g)]$$

if $\Delta H^\circ = -30 kJ$, $\Delta S^\circ = 30 J/K$, and $K_{eq} = 2.5 \times 10^{-3}$?

Video Solution

T = 855.8K

Q18.61E
Given $K_2 = 2$, $K_1 = 3$. Temperature increases from 100°C to 120°C. What’s $\Delta H^\circ$?

S18.61E
$\Delta H^\circ = 358.8 J$.

Q18.63E
For the reaction

$$[2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)]$$
calculate $\Delta G$ at 25°C if it’s known that $\Delta G^\circ (SO_3) = -150kJ$, $\Delta G^\circ (SO_2) = -125kJ$, $Q = 2.0 \times 10^{-3}$

S18.63E
$\Delta G = 159.6 \text{ kJ}$.

Q18.71E
For problem 61, if the value of $Q$ is not given, instead, it says that SO$_2$ gas is at 2atm, O$_2$ is at 0.5atm, and SO$_3$ gas at 4 atm. Calculate the new $\Delta G$.

S18.71E
$\Delta G = 176.7 \text{ kJ}$.

Q18.74E
For the reaction, Br$_2$(l) $\leftrightarrow$ Br$_2$(g), what’s the normal boiling point of it says that $\Delta H^\circ = 31.0 \text{ kJ/mol}$, $\Delta S^\circ = 93.07 \text{ J/Kmol}$?

S18.74E
$\Delta H^\circ = 333 \text{ K}$.

Q18.83E
Calculate $\Delta G^\circ$ for the oxidation of iron:

$$[4 \text{Fe(s) + 3O}_2(g) \rightleftharpoons 2\text{Fe}_2\text{O}_3(s)]$$

given:

- $\Delta H^\circ$ of $\text{2Fe}_2\text{O}_3(s))$ is -826 kJ/mol,
- $S^\circ_m$ of $\text{Fe_{(s)}}$ is 27 J/mol K and 205 J/mol K for $\text{O}_2$ and 90 J/mol K for $\text{Fe}_2\text{O}_3(s)$.

S18.83E
$\Delta G^\circ = 10^{261} \text{ J}$.

Q18.88E
What would happen if increasing the temperature of the reaction

$$[\text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2\text{HI}(s)]$$
Given that \( K_p = 620 \) at 298 K and \( \Delta H^\circ = -9.5 \text{ kJ/mol} \).

**S18.88E**

products are favored.

**Q18.104E**

For the reaction

\[
2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})
\]

if it’s known that \( K_p = 500 \) at 25 °C and \( \Delta H^\circ = -7.23 \text{ kJ/mol} \), what would happen if we decrease the temperature?

**S18.104E**

the reactants are favored.

**Misc**

**Q1**

Using the following values for entropy determine if a reaction would be spontaneous.

I) \( \Delta S_{\text{sys}} = 30 \text{ J/K} \), \( \Delta S_{\text{surr}} = 50 \text{ J/K} \)

II) \( \Delta S_{\text{sys}} = 60 \text{ J/K} \), \( \Delta S_{\text{surr}} = -85 \text{ J/K} \)

A reaction will be spontaneous if the total entropy change is positive. It will not be spontaneous if the total entropy change is negative.

\[ \Delta S_{\text{total}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}} \]

where

\( \Delta S_{\text{total}} \) is the total entropy change

\( \Delta S_{\text{sys}} \) is the entropy change of the system

\( \Delta S_{\text{surr}} \) is the entropy change of the surroundings

I.) \( \Delta S_{\text{total}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}} \)

\( \Delta S_{\text{total}} = 30 \text{ J/K} + 50 \text{ J/K} \)
ΔStotal = 80 J/K

The total entropy change is positive, therefore the reaction will be spontaneous.

II.) ΔStotal = ΔSsys + ΔSsurr

ΔStotal = 60 J/K + -85 J/K

ΔStotal = -25 J/K

The total entropy change is negative, therefore the reaction will not be spontaneous.

Q2

Under what conditions are the following reactions spontaneous, and why?

a. \( 3N_2(g) \rightleftharpoons 2N_3(g) \)

b. \( H_2O(l) \rightleftharpoons H_2O(g) \) where \( \Delta H \) is negative

c. \( 2NH_4NO_2(s) \rightleftharpoons 2N_2(g) + 4H_2O(g) + O_2(g) \) where \( \Delta H \) is negative

S2

a. Non spontaneous at all temperatures because entropy is decreased

b. Spontaneous when temperature is high, above 100˚C at STP

c. This reaction is spontaneous at all temperatures because of negative \( \Delta H \) and increasing entropy of the system

Q25

Calculate the change in Gibbs free energy.

\[ \text{HCl + NaOCH}_3 \rightarrow \text{NaCl + HOCH}_3 \]

\[ pK_a \text{HCl} = -7 \]
\[ pK_a \text{HOCH}_3 = 16 \]
\[ T = 298 \text{K} \]

\[ K_{eq} = \frac{[\text{NaCl}][\text{HOCH}_3][\text{H}^+]}{[\text{HCl}][\text{NaOCH}_3][\text{H}^+]} \]

\[ K_{eq} = \frac{K_a \text{HCl}}{K_a \text{HOCH}_3} \]

\[ K_{eq} = 10^{7}/10^{16} = 10^{23} \]

\[ \Delta G = -RT \ln K \]

\[ \Delta G = -8.314 \times 298 \times \ln 10^{23} \]

\[ \Delta G = -13.1 \text{kJ} \quad \text{← favored towards products} \]
Q35

In the synthesis of gaseous methanol from carbon monoxide gas and hydrogen gas, the following equilibrium concentrations were determined at 444 K: [CO(g)] = 0.0977M, [H2O(g)] = 0.0799M, and [CH3OH(g)] = 0.00799M. Calculate the equilibrium constant and Gibbs energy for this reaction.

S35

\[ \text{CO(g) + 2H}_2(\text{g}) \rightarrow \text{CH}_3\text{OH(g)} \]

\[ \text{Keq} = \frac{[\text{CH}_3\text{OH}]}{([\text{CO}] [\text{H}_2]^2)} \]

\[ \text{Keq} = \frac{0.00799}{(0.0977)(0.0799)^2} \]

\[ \text{Keq} = 12.8 \]

\[ \Delta \text{G}^\circ = -RT \ln(\text{Keq}) \]

\[ \Delta \text{G}^\circ = -(8.314 \text{J/Kmol})(444 \text{K}) \ln(12.8) (\text{kJ/1000J}) \]

\[ \Delta \text{G}^\circ = -9.4 \text{kJ/mol} \]

Q37

For the following reaction, what would \( \Delta \text{G}^\circ \) be at 298K ?

\[ \text{Fe}_3\text{O}_4(\text{s}) \rightarrow 3 \text{ Fe (s)} + 2\text{O}_2(\text{g}) \]

\[ \Delta \text{H}^\circ (\text{kJ/mol}) = -1118.4 \]

\[ \Delta \text{S}^\circ (\text{J/molK}) = 146.427.8205.1 \]

\[ \Delta \text{H}^\circ 0 - (-1118.4) = +1118.4 \text{ kJ} \]

\[ \Delta \text{S}^\circ (3(27.8) + 2(205.1)) - 146.4 = 347.2 \text{ J/K} \]

\[ \Delta \text{G}^\circ = \Delta \text{H}^\circ - T\Delta \text{S}^\circ = 1118.4 \text{kJ} - (298.15 \text{K})(0.3472 \text{kJ/K}) = 1015 \text{kJ} \]

Q41

If \( \Delta \text{H} = 158 \text{ kJ} \) and \( \Delta \text{S} = 411 \text{ J/k.} \) At what temperature will this reaction be spontaneous?

S41

\[ \Delta \text{G} = \Delta \text{H}^\circ - T\Delta \text{S}^\circ \]
0> 158000J – T(411 J/K)

T(411 J/K)/ 411 J/K > 158000 J/ 411 J/K

T > 384 K

**Q55**

Calculate $\Delta G$ of this reaction

$\Delta H=-537.22kJ \Delta S=13.74J/K T=25^\circ C$

$H_2(g)+F_2(g) \rightarrow 2HF$

$\Delta G=\Delta H-T\Delta S$

$\Delta G=(-537.22kJ)-(298K)(13.74J/K)$

$\Delta G=(-537.22kJ)-(298K)(.01374kJ/K)$

$\Delta G=-541.31kJ \leftarrow $ SPONTANEOUS

**Q57**

What must be the temperature if the following reaction has $\Delta G^\circ= -52.9 \text{kJ}$, $\Delta H^\circ= -34.7 \text{kJ}$, and $\Delta S^\circ=12.4 \text{J/K}$?

$\text{Fe}_2\text{O}_3(s)+3\text{CO}(g) \rightarrow 2\text{Fe}(s)+3\text{CO}_2(g)$

$\Delta G^\circ=\Delta H^\circ-T\Delta S^\circ$

$(-52.9 \text{kJ})(1000\text{kJ})=(-34.7 \text{kJ})(1000\text{kJ})-T(12.4\text{J/K})$

$T=1470K$

**Q67**

Label which reaction is spontaneous or nonspontaneous, and compute the overall reaction, given that it is spontaneous.

$\text{Cu}_2\text{O} (s) \rightarrow 2 \text{Cu}(s) + \frac{1}{2} \text{O}_2 (g) \Delta G^\circ = 125 \text{kJ}$

$\text{C(s) + } \frac{1}{2} \text{O}_2(g) \rightarrow \text{CO(g) } \Delta G^\circ = -175$

Reaction 1: nonspontaneous; reaction 2: spontaneous

Net reaction:
Cu₂O (s) + C (s) → 2 Cu (s) + CO(g)

ΔG° = 125 + (-175) = -50 kJ

Net reaction is spontaneous.