

Classroom Activity Chapter 17 Number 1

20 March 2017

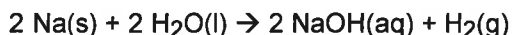
1. Which of the following processes are spontaneous?

- (a) dissolving salt in water
- (b) dissolving NaCl in a saturated solution of NaCl(aq)
- (c) paper burning
- (d) iron rusting
- (e) ice melting in a freezer
- (f) a reaction at equilibrium shifting right or left
- (g) water evaporating at 25 °C in a container in which the relative humidity is 50%. (The vapor pressure of water is 24.0 mmHg at 25 °C. If the partial pressure of water is 24.0 mmHg, that would be 100% humidity. A 50% humidity at 25 °C corresponds to a partial pressure of water = 12.0 mmHg.)
- (h) cheese becoming moldy

2. Which of the following processes are associated with an increase in entropy?

- (a) $\text{CO(g)} + \text{O}_2\text{(g)} \rightarrow 2 \text{CO}_2\text{(g)}$
- (b) $2 \text{NO}_2\text{(g)} \rightarrow \text{N}_2\text{O}_4\text{(g)}$
- (c) $2 \text{Na(s)} + 2 \text{H}_2\text{O(l)} \rightarrow 2 \text{NaOH(aq)} + \text{H}_2\text{(g)}$
- (d) iron rusting
- (e) water freezing in a freezer
- (f) water evaporating

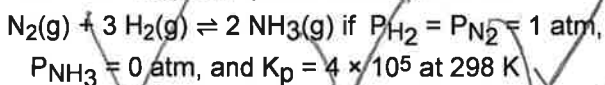
3. Sodium reacts violently with water according to the equation:



The resulting solution has a higher temperature than the water prior to the addition of sodium. What are the signs of ΔH° and ΔS° ?

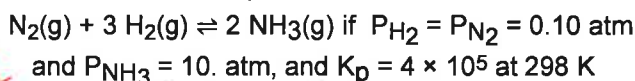
$$\Delta H^\circ < 0 \quad \Delta S^\circ > 0$$

4a. Is this reaction spontaneous at 298 K?



~~$Q_p = \frac{P_{\text{NH}_3}^2}{P_{\text{N}_2} \cdot P_{\text{H}_2}^3} = \frac{0^2}{1 \cdot 1^3} = 0$~~ ~~$\Delta G = \Delta G^\circ + RT \ln Q$~~
 ~~$\Delta G^\circ = -RT \ln K_p$~~

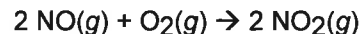
4b. Is this reaction spontaneous at 298 K?



Be longer in Sect 17.10

$Q_p = \frac{P_{\text{NH}_3}^2}{P_{\text{N}_2} \cdot P_{\text{H}_2}^3} = \frac{10^2}{0.1 \cdot 0.1^3} = 1 \times 10^6$ Q_p
 $\Delta G^\circ = -RT \ln K_p$
 $\Delta G = \Delta G^\circ + RT \ln Q_p = -0.008314 \cdot 298 \cdot \ln(1 \times 10^6)$
 $= -32.0 + RT \ln(1 \times 10^6) = -32.0 \text{ kJ}$ no net spont

5a. Calculate ΔS° for the following reaction.



	ΔH_f° kJ/mol	ΔG_f° kJ/mol	S° J/mol K
NO(g)	91.3	87.6	210.7
O ₂ (g)	0	0	205.0
NO ₂ (g)	33.2	51.3	240.0

$2 \text{NO} + \text{O}_2 \rightarrow 2 \text{NO}_2$
 $2(210.7) \quad 2(205 \text{ J/mol K}) \quad 2(240 \text{ J/mol K})$
 $\times -2 \quad \times -1 \text{ mol} \quad \times 2 \text{ mol}$
 $\Delta S^\circ = -421.4 \text{ J/K} - 205 \text{ J/K} + 480 \text{ J/K} = -146.4 \text{ J/K}$

5b. Calculate ΔH° for the reaction above.

$2 \text{NO} + \text{O}_2 \rightarrow 2 \text{NO}_2$
 $2(91.3 \text{ kJ/mol}) \quad 0 \text{ kJ/mol} \quad 2(33.2 \text{ kJ/mol})$
 $\times -2 \text{ mol} \quad \times -1 \text{ mol} \quad \times 2 \text{ mol}$
 $\Delta H^\circ = -182.6 \text{ kJ} + 0 + 66.4 \text{ kJ} = -116.2 \text{ kJ}$

5c. Calculate ΔG° for the reaction using the data.

$2 \text{NO} + \text{O}_2 \rightarrow 2 \text{NO}_2$
 $2(87.6 \text{ kJ/mol}) \quad 0 \text{ kJ/mol} \quad 2(51.3 \text{ kJ/mol})$
 $\times -2 \text{ mol} \quad \times -1 \text{ mol} \quad \times 2 \text{ mol}$
 $\Delta G^\circ = -175.2 \text{ kJ} + 0 + 102.6 \text{ kJ} = -72.6 \text{ kJ}$

5d. Calculate ΔG° for the reaction using the Gibbs-Helmholtz equation: $\Delta G^\circ = \Delta H^\circ - T \Delta S^\circ$.

$\Delta G^\circ = \Delta H^\circ - T \Delta S^\circ$
 $= -116.2 \text{ kJ} - 298 \text{ K} \cdot (-0.1464 \text{ kJ/K})$
 $= -72.6 \text{ kJ}$

5d. Estimate ΔG for the reaction at 350 K using the Gibbs-Helmholtz equation: $\Delta G = \Delta H^\circ - T \Delta S^\circ$.

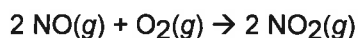
$\Delta G \approx \Delta H^\circ - T \Delta S^\circ$
 $= -116.2 \text{ kJ} - 350 \text{ K} \cdot (-0.1464 \text{ kJ/K})$
 $= -65 \text{ kJ}$

Q_p

Classroom Activity Chapter 17 Number 2

24 March 2017

1. In the previous worksheet, you calculated $\Delta H^\circ = -116.2 \text{ kJ}$, $\Delta S^\circ = -146.4 \text{ J/K}$ and $\Delta G^\circ = -72.6 \text{ kJ}$ for the reaction.



1a. Is this reaction spontaneous at...

(a) all temperatures

(b) never

(c) spontaneous at high temperatures; non-spontaneous at low temperatures

(d) spontaneous at low temperatures; non-spontaneous at high temperatures

$\Delta S < 0$
entropy not favored



1b. At what temperature does this reaction become spontaneous?

$$T = \frac{\Delta H^\circ}{\Delta S^\circ} = \frac{-116.2 \text{ kJ}}{-0.1464 \text{ kJ/K}} = 794 \text{ K}$$

1c. Estimate ΔG at a temperature of 25 degrees lower than the temperature calculated in the previous problem. Is the reaction spontaneous at this temperature?

$$\Delta G = \Delta H^\circ - T\Delta S^\circ = -116.2 \text{ kJ} - 769 \text{ K} * (-0.1464 \text{ kJ/K}) = -3.7 \text{ kJ}$$

1d. Estimate ΔG at a temperature of 25 degrees higher than the temperature calculated in Problem 1b. Is the reaction spontaneous at this temperature?

$$\Delta G = \Delta H^\circ - T\Delta S^\circ = -116.2 - 819 * (-0.1464) = +3.7 \text{ kJ}$$

1e. In our next lecture we will tie ΔG° to the equilibrium constant, K_p . From what you know already, how does ΔG , Q_p , and K_p relate to one another? In each case, circle one: $<$, $=$, or $>$

If $\Delta G < 0$ then Q_p (\ominus) = (\ominus) $>$ K_p

If $\Delta G = 0$ then Q_p (\ominus) = (\ominus) $>$ K_p

If $\Delta G > 0$ then Q_p (\ominus) = (\ominus) $>$ K_p

2. Phase changes always have the same signs for ΔH° and ΔS° . What are the signs for ΔH° and ΔS° for these phase changes?

	ΔH	ΔS
2a. vaporization	+	+
2b. freezing	-	-
2c. condensation	-	-
2d. fusion (melting)	+	+
2e. sublimation	+	+
2f. deposition	-	-

3. Determine ΔH° and ΔS° for the vaporization of mercury and then estimate the boiling point.

	ΔH_f° kJ/mol	ΔG_f° kJ/mol	S° J/mol K
Hg(l)	0	0	76.0
Hg(g)	61.3	31.9	174.8

$$\text{Hg}(l) \rightarrow \text{Hg}(g) \quad T = \frac{\Delta H}{\Delta S} = \frac{61.3}{0.0988} = 620 \text{ K}$$

$$\Delta S^\circ = -76.0 + 174.8 = 98.8 \text{ J/K}$$

4. Predict ΔH° , ΔS° and ΔG° for these processes and classify each as

	ΔG	ΔH	ΔS
(a) spontaneous at all temperatures,	-	-	+
(b) never spontaneous,	+	+	-
(c) spontaneous at high temperatures; non-spontaneous at low temperatures, or		+	+
(d) spontaneous at low temperatures; non-spontaneous at high temperatures		-	-

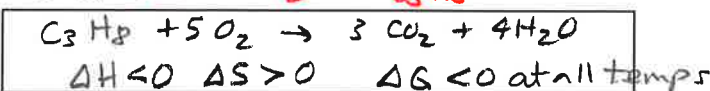
4a. $\text{O}_2(g) \rightarrow 2 \text{ O}(g)$

$$\Delta G (?) \quad \Delta H > 0 \quad \Delta S > 0$$

4b. $2 \text{ NO}(g) + \text{ O}_2(g) \rightarrow 2 \text{ NO}_2(g)$ (exothermic)

$$\Delta G (?) \quad \Delta H < 0 \quad \Delta S < 0$$

4c. combustion of propane (bp = -42°C) to form $\text{H}_2\text{O}(g)$ and CO_2



4d. $\text{H}_2\text{O}_2(l) \rightarrow \text{H}_2\text{O}(l) + \text{ O}_2(g)$ (exothermic)

$$\Delta H < 0 \quad \Delta S > 0 \quad \Delta G < 0 \text{ always}$$

Classroom Activity Chapter 17 Number 3

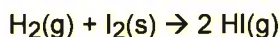
27 March 2017

$$\Delta G = \Delta G^\circ + RT \ln(Q_p)$$

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

$$\text{where } R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$$

- 1a. Calculate ΔG for this reaction at 298 K given $\Delta G^\circ = +3.4 \text{ kJ}$, $P_{\text{H}_2} = 1.0 \text{ atm}$, and $P_{\text{HI}} = 5.0 \text{ atm}$.



$$Q_p = \frac{P_{\text{HI}}^2}{P_{\text{H}_2}} = \frac{5.0^2}{1.0} = 25$$

$$\Delta G = \Delta G^\circ + RT \ln Q_p$$

$$= +3.4 + (0.008314)(298) \ln 25$$

$$= +11.4 \text{ kJ}$$

- 1b. Repeat the calculation at 298 K if $P_{\text{H}_2} = 5.0 \text{ atm}$, and $P_{\text{HI}} = 1.0 \text{ atm}$.

$$Q_p = \frac{1^2}{5} = 0.20$$

$$\Delta G = +3.4 + (0.008314)(298) \ln(0.20)$$

$$= -0.59 \text{ kJ}$$

- 1c. What is the equilibrium constant, K_p , for this reaction at 298 K?

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

$$3.4 = -(0.008314)(298) \ln K_p$$

$$K_p = 0.25$$

2. What is the relationship between ΔG° and the equilibrium constant, K_p ?

If $\Delta G^\circ < 0$ then $K_p (< \textcircled{>}) > 1$

If $\Delta G^\circ = 0$ then $K_p (< \textcircled{=} >) = 1$

If $\Delta G^\circ > 0$ then $K_p (< \textcircled{<} >) < 1$

3. What is the vapor pressure of water at 298 K? Express your answer in mmHg. Given: $\Delta G_f^\circ = -237.2 \text{ kJ}$ for $\text{H}_2\text{O}(\text{l})$ and -228.6 kJ for $\text{H}_2\text{O}(\text{g})$

$$\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_2\text{O}(\text{g}) \quad K_p = P_{\text{H}_2\text{O}}$$

ΔG_f°	-237.2	-228.6
	* -1	* 1

$$\Delta G_{\text{vap}}^\circ = +237.2 - 228.6 = 8.6 \text{ kJ}$$

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

$$8.6 = -(0.008314)(298) \ln K_p$$

$$K_p = 0.0311 \text{ atm} \rightarrow 23.6 \text{ mmHg}$$

4. Consider the reaction $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightarrow 2 \text{NH}_3(\text{g})$ for which $\Delta H^\circ = -92.2 \text{ kJ}$, $\Delta S^\circ = -198.9 \text{ J/K}$, and $\Delta G^\circ = -33.0 \text{ kJ}$.

- 4a. At what temperature does this reaction become spontaneous?

$$T = \frac{\Delta H^\circ}{\Delta S^\circ} = \frac{-92.2}{-0.1989} = 464 \text{ K}$$

- 4b. What is the equilibrium constant, K_p , at 298 K?

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

$$-33.0 = -(0.008314)(298) \ln K_p$$

$$K_p = 6.1 \times 10^5$$

- 4c. The reaction is kinetically too slow at 298 K. Increasing the temperature increases the rates of every reaction. What is the trade-off in terms of kinetics vs. thermodynamics if we increase the temperature?

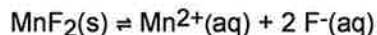
This is a rxn that becomes non-spont at 464 K. can't raise temp too much

- 4d. Does the LeChatelier principle support reducing the volume (increasing the pressure) to shift the mixture towards more ammonia? Circle: Yes or No

5. As a reaction proceeds undisturbed, which of these things happen? Circle all that apply.

- | | |
|--------------------------------|---------------------------------------|
| $\Delta G^\circ \rightarrow 0$ | $\Delta G \rightarrow 0$ |
| $Q_p \rightarrow K_p$ | $Q_p \rightarrow 0$ |
| $K_p \rightarrow 1$ | $\Delta G \rightarrow \Delta G^\circ$ |

6. The solubility of manganese(II) fluoride in water is 6.6 g/mL at 40 °C and 4.8 g/L at 100 °C. Based on these data, what is the sign of ΔH° and ΔS° for the process below?



- (A) ΔH° is negative but the sign of ΔS° cannot be determined from this information.

- B. ΔH° is negative and ΔS° is definitely negative.

- C. ΔH° is positive but the sign of ΔS° cannot be determined from this information.

- D. ΔH° is positive and ΔS° is definitely negative.

A is the best answer

incr temp favors the endothermic direction

K_{sp} decr at temp ↑
∴ $\Delta H < 0$
 $\Delta S > 0$