**Entropy and Gibbs Free Energy**

*Entropy* is the degree of randomness in a substance. The symbol for change in entropy is $\Delta S$.

Solids are very ordered and have low entropy. Liquids and aqueous ions have more entropy because they move about more freely, and gases have an even larger amount of entropy. According to the Second Law of Thermodynamics, nature is always proceeding to a state of higher entropy. Also, when there are moles of products than moles of reactants, entropy is increasing. When total moles are reduced, entropy is decreasing.

Determine whether the following reactions show an *increase or decrease* in entropy and write the $\Delta S$ as (+) or (−).

1. $\text{KClO}_3(s) \rightarrow 2\text{KCl}(s) + 3\text{O}_2(g)$  
   entropy increases $\quad +\Delta S$

2. $\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(s)$  
   entropy decreases $\quad -\Delta S$

3. $\text{N}_2(g) + 3 \text{H}_2(g) \rightarrow 2 \text{NH}_3(g)$  
   entropy decreases $\quad -\Delta S$

4. $\text{NaCl}(s) \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq)$  
   entropy increases $\quad +\Delta S$

5. $\text{KCl}(s) \rightarrow \text{KCl}(l)$  
   entropy increases $\quad +\Delta S$

6. $\text{CO}_2(s) \rightarrow \text{CO}_2(g)$  
   entropy increases $\quad +\Delta S$

7. $\text{H}^+(aq) + \text{C}_2\text{H}_3\text{O}_2(aq) \rightarrow \text{HC}_2\text{H}_3\text{O}_3(l)$  
   entropy decreases $\quad -\Delta S$

8. $\text{C}(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g)$  
   too close to tell

9. $\text{H}_2(g) + \text{Cl}_2(g) \rightarrow 2\text{HCl}(g)$  
   no change

10. $\text{Ag}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{AgCl}(s)$  
    entropy decreases $\quad -\Delta S$

11. $2\text{N}_2\text{O}_5(g) \rightarrow 4\text{NO}_2(g) + \text{O}_2(g)$  
    entropy increases $\quad +\Delta S$

12. $2\text{Al}(s) + 3\text{I}_2(s) \rightarrow 2\text{AlI}_3(s)$  
    entropy decreases $\quad -\Delta S$

13. $\text{H}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}(l)$  
    entropy decreases $\quad -\Delta S$

14. $2\text{NO}(g) \rightarrow \text{N}_2(g) + \text{O}_2(g)$  
    no change

15. $\text{H}_2\text{O}(g) \rightarrow \text{H}_2\text{O}(l)$  
    entropy decreases $\quad -\Delta S$
Gibbs Free Energy

For a reaction to be spontaneous, the sign of \( \Delta G \) (Gibbs Free Energy) must be negative. The mathematical formula for this value is:

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

Where \( \Delta H \) = change in enthalpy or heat of reaction, \( T \) = temperature in Kelvin, \( \Delta S \) = change in entropy

Complete the table for the sign of \( \Delta G \); (+) OR (-) OR undetermined. When conditions allow for an undetermined sign of \( \Delta G \), temperature will decide spontaneity.

<table>
<thead>
<tr>
<th>( \Delta H )</th>
<th>( \Delta S )</th>
<th>( \Delta G )</th>
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Answer the questions below.

1. The conditions in which \( \Delta G \) is always negative is when \( \Delta H \) is exothermic and \( \Delta S \) is increasing.
2. The conditions in which \( \Delta G \) is always positive is when \( \Delta H \) is endothermic and \( \Delta S \) is decreasing.
3. When the situation is indeterminate, a low temperature favors the (entropy / enthalpy) factor, and a high temperature favors the (entropy / enthalpy) factor.

Answer Problems 4-6 with always, sometimes, or never.

4. The reaction: \( \text{Na(OH)}(s) \rightarrow \text{Na}^{+}(aq) + \text{OH}^{-}(aq) + \text{energy} \) will always be spontaneous.
5. The reaction: \( \text{energy} + 2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(l) \) will never be spontaneous.
6. The reaction: \( \text{energy} + \text{H}_2\text{O}(s) \rightarrow \text{H}_2\text{O}(l) \) will sometimes be spontaneous.
7. What is the value of \( \Delta G \) if \( \Delta H = -32.0 \text{ kJ}, \Delta S = +25.0 \text{ kJ/K} \) and \( T = 293 \text{ K} \)? Tell whether this reaction is spontaneous.

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

\[
\Delta G = -7357 \text{ kJ}
\]

spontaneous

8. What is the value of \( \Delta G \) if \( \Delta H = +12.0 \text{ kJ}, \Delta S = -5.00 \text{ kJ/K} \) and \( T = 290 \text{ K} \)? Tell whether this reaction is spontaneous.

\[
\Delta G = 1500 \text{ kJ}
\]

nonspontaneous

Use the Table of Thermodynamic Properties for the following problems.

9. Calculate the standard heat of reaction (\( \Delta H \)) for the reaction: \( \text{NH}_4\text{Cl}(s) \rightarrow \text{HCl}(g) + \text{NH}_3(g) \)

\[
\Delta H = 176.9 \text{ kJ}
\]

10. Calculate the change in entropy (\( \Delta S \)) for: \( \text{NH}_4\text{Cl}(s) \rightarrow \text{HCl}(g) + \text{NH}_3(g) \)

\[
\Delta S = 285 \text{ J/K} \text{ or } 0.285 \text{ kJ/K}
\]
11. Calculate the standard $\Delta G$ for the reaction in (9) and (10). Remember standard means the temperature is 25°C or 298 K. Use $\Delta G = \Delta H - T \Delta S$. Identify the reaction as spontaneous or non-spontaneous.

** convert temp to Kelvin
 ** convert $\Delta S$ from Joules to kJ $\Delta G = 92.0 kJ$ nonspontaneous

12. Use the table values for $\Delta G$ of formation to re-calculate the $\Delta G$ for the reaction in (9) and (10). Compare your answer to your answer for (11).

$\Delta G = 91.0 kJ$ nonspontaneous

13. At what temperature is the reaction in (9) and (10) spontaneous? Assume $\Delta G = 0$ is spontaneous.

At high temps

14. Calculate the standard heat of reaction ($\Delta H$) for: $4 \text{NH}_3(g) + 5 \text{O}_2(g) \rightarrow 4 \text{NO}(g) + 6 \text{H}_2\text{O}(g)$

$\Delta H = -905.6 \text{ kJ}$

15. Calculate the change in entropy ($\Delta S$) for the reaction in (14).

$\Delta S = 181 \text{ J/K or 0.181 kJ/K}$

16. Calculate the $\Delta G$ for the reaction in (14) and determine if this reaction is spontaneous or non-spontaneous.

$\Delta G = -960. \text{ kJ} \text{ spontaneous}$

17. Determine the $\Delta G$ for the reaction in which aluminum oxidizes to form aluminum oxide. Is it spontaneous or non-spontaneous?

$4 \text{Al(s)} + 3 \text{O}_2(g) \rightarrow 2 \text{Al}_2\text{O}_3(s)$

$\Delta G = -3160 \text{ kJ} \text{ spontaneous}$

18. Determine the $\Delta G$ for the following reaction:

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

$\Delta G = -1480 \text{ kJ} \text{ spontaneous}$