**Chemical Thermodynamics** **Name**_____

## Spontaneity and Gibbs Free Energy

Read from **Lesson 3: Spontaneity and Gibbs Free Energy** in the **Chemistry Tutorial** Section**, Chapter 17** of **The Physics Classroom.**
**Part a:** Predicting Spontaneity with Gibbs Free Energy **Part b:** Gibbs Free Energy and Equilibrium

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

### 1. Gibbs Free Energy

- **Definition:** Gibbs free energy (G) is a thermodynamic function combining enthalpy (H), entropy (S), and temperature (T): **$G = H - T \cdot S$**
- **Change in Gibbs Free Energy: $\Delta G = \Delta H - T\Delta S$**
- **Driving Forces:**
  - **$\Delta H$ (enthalpy change):** Reflects energy spread; negative $\Delta H$ favors spontaneity.
  - **$T\Delta S$ (entropy term):** Reflects matter spread; positive $\Delta S$ favors spontaneity.
- **Spontaneity Criteria:**
  - $\Delta G < 0 \rightarrow$ spontaneous process
  - $\Delta G > 0 \rightarrow$ non-spontaneous (reverse is spontaneous)
  - $\Delta G = 0 \rightarrow$ equilibrium
- **Gibbs free energy predicts spontaneity** and connects thermodynamics with equilibrium. $\Delta G$ tells us **not only** if a reaction **occurs** but also how far it **proceeds.**

### 2. Temperature Dependence

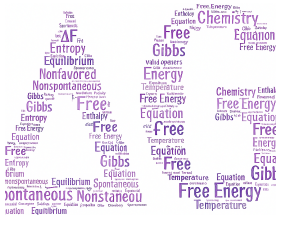
- When $\Delta H$ and $\Delta S$ are in opposition, **temperature determines spontaneity**:
  - $\Delta H < 0$, $\Delta S < 0 \rightarrow$ spontaneous at low T
  - $\Delta H > 0$, $\Delta S > 0 \rightarrow$ spontaneous at high T
  - $\Delta H < 0$, $\Delta S > 0 \rightarrow$ spontaneous at all T
  - $\Delta H > 0$, $\Delta S < 0 \rightarrow$ non-spontaneous at all T
- **Threshold Temperature: $T_{threshold} = \Delta H/\Delta S$**
- At this temperature, **$\Delta G = 0$ (equilibrium point).**

### 3. Magnitude of $\Delta G$

- **Large negative $\Delta G$** → reaction **strongly favors products** (nearly complete).
- **Small negative $\Delta G$** → reaction favors **products** but reaches equilibrium before completion.
- **Positive $\Delta G$** → reaction **favors reactants**.

### 4. Gibbs Free Energy and Equilibrium Constant

- At equilibrium: $\Delta G = 0$ and $Q = K$.
- Relationship between $\Delta G^\circ$ and K: **$\Delta G^\circ = - R \cdot T \cdot \ln K$** and **$K = e^{-\Delta G^\circ/RT}$**
- **Interpretation:**
  - **$\Delta G^\circ < 0 \rightarrow K > 1 \rightarrow$ products favored.**
  - **$\Delta G^\circ > 0 \rightarrow K < 1 \rightarrow$ reactants favored.**
  - **$\Delta G^\circ = 0 \rightarrow K \approx 1 \rightarrow$ balanced equilibrium.**




### 5. Non-Standard Conditions

- General equation: **$\Delta G = \Delta G^\circ + R \cdot T \cdot \ln Q$**
- Q (reaction quotient) reflects current concentrations/pressures.
- As reaction proceeds, Q changes, and $\Delta G$ shifts until equilibrium ($\Delta G = 0$, $Q = K$).

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## Questions

1. How does the magnitude of **ΔG** relate to the "extent" of a reaction?

2. For each of the following hypothetical reactions, analyze the signs of ΔH (enthalpy change) and ΔS (entropy change). State whether the reaction is spontaneous at "**low temperatures**," "**high temperatures**," "**all temperatures**," or "**no temperatures**":
   a) ΔH = +150 kJ, ΔS = +250 J/K

   b) ΔH = –75 kJ, ΔS = –120 J/K

   c) ΔH = –200 kJ, ΔS = +50 J/K

   d) ΔH = +10 kJ, ΔS = –5 J/K

3. Consider the formation of magnesium oxide from its elements: **$Mg(s) + \frac{1}{2} O_2(g) \rightarrow MgO(s)$**
   Given that **ΔH = -601.6 kJ** and that **ΔS = 26.9 J/K**:

   a. Is this reaction exothermic or endothermic?

   b. Does it increase or decrease the disorder of the system?

   c. Calculate the **$\Delta G^\circ$** for the reaction at **25 °C**. Is the reaction spontaneous at this temperature?

   d. Calculate the **$\Delta G^\circ$** for the reaction at **596 K**. Is the reaction spontaneous at this temperature?

4. **Octane** (**$C_8H_{18}$**) is a straight-chain alkane and a component of gasoline.
   a. Write the balanced equation for the combustion of octane with oxygen to produce carbon dioxide gas and liquid water.

   b. Without using thermodynamic data, predict whether **$\Delta G^\circ$** for this reaction is more negative or less negative than **$\Delta H^\circ$**. Explain your reasoning

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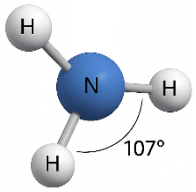
5. **Reaction A** has a **large negative ΔH** and a **small negative ΔS**. **Reaction B** has a **small negative ΔH** and a **large positive ΔS**.
   a. Under what temperature conditions could **Reaction B** be spontaneous while **Reaction A** is not?

   b. Explain your reasoning in terms of **ΔG**.

6. Consider the phase change: **$NH_3(l) \rightarrow NH_3(g)$.** The **ΔH = +23.3 kJ/mol** and **ΔS = +97.4 J/mol·K.**
   a. How much heat is released or absorbed when **53.47 g** of $NH_3$ ($l$) evaporates?




   b. Calculate the **$\Delta G^\circ$** for the reaction at **250 K.** Is the reaction spontaneous at this temperature?

   c. Calculate the **$\Delta G^\circ$** for the reaction at **550 K.** Is the reaction spontaneous at this temperature?

   d. At which temperature is the process at equilibrium?

   e. What is the normal boiling point (in **$^\circ$C**) of ammonia?

7. Suppose a reaction is at nonstandard conditions, and **Q** (the reaction quotient) > **K**. Answer these questions about this reaction and explain your reasoning.
   a. What sign will **ΔG** have?

   b. Will the reaction proceed in the forward or reverse direction to reach equilibrium?