An **exothermic** reaction occurs when the temperature of a system increases due to the evolution of heat. This heat is released into the surroundings, resulting in an overall negative quantity for the heat of reaction ($q_{rxn} < 0$). An **endothermic** reaction occurs when the temperature of an isolated system decreases while the surroundings of a non-isolated system gains heat. Endothermic reactions result in an overall positive heat of reaction ($q_{rxn} > 0$).

Exothermic and endothermic reactions cause energy level differences and therefore differences in enthalpy ($\Delta H$), the sum of all potential and kinetic energies. $\Delta H$ is determined by the system, not the surrounding environment in a reaction. A system that **releases** heat to the surroundings, an exothermic reaction, has a **negative** $\Delta H$ by convention, because the enthalpy of the products is lower than the enthalpy of the reactants of the system.

\[
\ce{C(s) + O2(g) -> CO2 (g)} \tag{\Delta H = –393.5 \text{ kJ}}
\]
\[
\ce{ H2 (g) + 1/2 O2 (g) -> H2O(l)} \tag{\Delta H = –285.8 \text{ kJ}}
\]

The enthalpies of these reactions are **less** than zero, and are therefore exothermic reactions. A system of reactants that **absorbs** heat from the surroundings in an endothermic reaction has a **positive** ($\Delta H$), because the enthalpy of the products is higher than the enthalpy of the reactants of the system.

\[
\ce{N2(g) + O2(g) -> 2NO(g)} \tag{\Delta H = +180.5 \text{ kJ}}
\]
\[
\ce{ C(s) + 2S(s) -> CS2(l)} \tag{\Delta H = +92.0 \text{ kJ}}
\]

Because the enthalpies of these reactions are **greater** than zero, they are endothermic reactions.

---

**The Equilibrium Constant**

The equilibrium constant ($K_c$) defines the relationship among the concentrations of chemical substances involved in a reaction at equilibrium. The [Le Chatelier's principle](https://en.wikipedia.org/wiki/Le_Chatelier%27s_principle) states that if a stress, such as changing temperature, pressure, or concentration, is inflicted on an equilibrium reaction, the reaction will shift to restore the equilibrium. For exothermic and endothermic reactions, this added stress is a change in temperature. The equilibrium constant shows how far the reaction will progress at a specific temperature by determining the ratio of products to reactions using equilibrium concentrations.

The equilibrium expression for the following equation

\[[aA + bB \rightleftharpoons cC + dD]\]

is given below:

\[
K_c = \dfrac{[C]^c[D]^d}{[A]^a[B]^b} \label{Equation:Kc}
\]

where

- $K_c$ is the equilibrium constant (for concentrations)
- $[A], [B], [C], [D]$ are concentrations
• a, b, c, and d are the stoichiometric coefficients of the balanced equation

<table>
<thead>
<tr>
<th>Exothermic Reactions</th>
<th>Endothermic Reactions</th>
</tr>
</thead>
<tbody>
<tr>
<td>If $K_c$ decreases with an increase in temperature, the reaction shifts to the left.</td>
<td>If $K_c$ increases with an increase in temperature, the reaction shifts to the right.</td>
</tr>
<tr>
<td>If $K_c$ increases with a decrease in temperature, the reaction shifts to the right.</td>
<td>If $K_c$ decreases with a decrease in temperature, the reaction shifts to the left.</td>
</tr>
</tbody>
</table>

If the products dominate in a reaction, the value for $K$ is greater than 1. The larger the $K$ value, the more the reaction will tend toward the right and thus to completion.

• If $K=1$, neither the reactants nor the products are favored. Note that this is not the same as both being favored.

• If the reactants dominate in a reaction, then $K<1$. The smaller the $K$ value, the more the reaction will tend toward the left.

Example \(\PageIndex{1}\) : The Haber Process

Suppose that the following reaction is at equilibrium and that the concentration of $N_2$ is 2 M, the concentration of $H_2$ is 4 M, and the concentration of $NH_3$ is 3 M. What is the value of $K_c$?

\[
\ce{ N2 + 3H2 <=> 2NH3}
\]

The coefficients and the concentrations are plugged into the \(K_c\) expression (Equation \ref{Equation:Kc}) to calculate its value.

\[
\begin{align*}
K_c &= \dfrac{[NH_3]^2}{[N_2]^1[H_2]^3} \\
&= \dfrac{[3]^2}{[2]^1[4]^3} \\
&= 0.07
\end{align*}
\]

Exercise \(\PageIndex{1}\))

Determine \(K_c\) for the following chemical reaction at equilibrium if the molar concentrations of the molecules are:

• 0.20 M \(\langle\text{ce}{H2}\rangle\),
• 0.10 M \(\langle\text{ce}{NO}\rangle\),
• 0.20 M \(\langle\text{ce}{H2O}\rangle\), and
• 0.10M \(\langle\text{ce}{N2}\rangle\).

\[
\ce{2H2 (g) + 2NO (g) <=> 2H2O (g) + N2 (g)}
\]

Answer

Using the \(K_c\) expression (Equation \ref{Equation:Kc}) and plugging in the concentration values of each molecule:

\[
\begin{align*}
K_c &= \dfrac{[H2O]^2[N2]^1}{[H2]^2[NO]^2} \\
&= \dfrac{0.20^2 \cdot 0.1}{0.20^2 \cdot 0.10^2} \\
&= 10
\end{align*}
\]
Exercise \(\PageIndex{2}\))

For the previous equation, does the equilibrium favor the products or the reactants?

**Answer**

Because \(K_c = 10 > 1\), the reaction favors the products.

Exercise \(\PageIndex{3}\))

In the following reaction, the temperature is increased and the \(K_c\) value decreases from 0.75 to 0.55. Is this an exothermic or endothermic reaction?

\[
\ce{N_2 (g) + 3H_2 \rightleftharpoons 2NH_3 (g)}
\]

**Answer**

Because the K value decreases with an increase in temperature, the reaction is an exothermic reaction.

Exercise \(\PageIndex{4}\))

In the following reaction, in which direction will the equilibrium shift if there is an increase in temperature and the enthalpy of reaction is given such that \(\Delta H\) is -92.5 kJ?

\[
\ce{PCl_3(g) + Cl_2(g) \rightleftharpoons PCl_5(g)}
\]

**Answer**

In the initial reaction, the energy given off is negative and thus the reaction is exothermic. However, an increase in temperature allows the system to absorb energy and thus favor an endothermic reaction; the equilibrium will shift to the left.

---

**Contributors and Attributions**

- Alyson Salmon, Nikita Patel (UCD), Deepak Nallur (UCD)