(6) 1. Chemists sometimes like to say that an increase in temperature of 10 \( \text{\( ^\circ \)C} \) leads to a doubling of the rate of a reaction. Suppose a reaction has a rate constant of \( 1.5 \times 10^{-2} \text{ s}^{-1} \) at 25 \( \text{\( ^\circ \)C} \) and shows an increase in rate constant to \( 3.0 \times 10^{-2} \text{ s}^{-1} \) at 35 \( \text{\( ^\circ \)C} \). What is the activation energy of such a reaction?

\[
\ln \left( \frac{1.5 \times 10^{-2} \text{ s}^{-1}}{3.0 \times 10^{-2} \text{ s}^{-1}} \right) = \frac{E_a}{8.314 \text{ J/molK}} \left( \frac{1}{308 K} - \frac{1}{298 K} \right) \Rightarrow -0.693 = (-1.31 \times 10^{-3} \text{ mol/J})E_a \Rightarrow
\]

\[
E_a = \frac{-0.693}{-1.31 \times 10^{-3} \text{ mol/J}} \Rightarrow E_a = 52900 \text{ J/mol}
\]

\[
k = Ae^{-E_a/RT} \quad \ln \frac{k_1}{k_2} = \frac{E_a}{R} \left( \frac{T_2}{T_1} - 1 \right)
\]

(4) 2. Consider the following equilibria and the given concentrations and equilibrium constants. For each equilibrium, determine whether or not the system is at equilibrium. If not, predict the direction it will shift (left or right) to attain equilibrium.

\[
\text{H}_2 (g) + I_2 (g) \rightleftharpoons 2 \text{HI} (g) \quad K_c = 50
\]

Concentrations: \([\text{H}_2] = 3.0 \times 10^{-2} \text{ M}; \quad [I_2] = 2.0 \times 10^{-2} \text{ M}; \quad [\text{HI}] = 0.20 \text{ M}\)

\[
Q = \frac{[\text{HI}]^2}{[\text{H}_2][I_2]} = \frac{(0.20)^2}{(3.0 \times 10^{-2})(2.0 \times 10^{-2})} = 67
\]

Since \( Q > K \), the reaction will shift to the left (\( \searrow \)) to come to equilibrium.

\[
\text{C (s) + H}_2\text{O (g) \rightleftharpoons CO (g) + H}_2\text{(g)} \quad K_p = 14
\]

Pressures: \( P_{\text{H}_2\text{O}} = 0.10 \text{ atm}; \quad P_{\text{CO}} = 0.20 \text{ atm}; \quad P_{\text{H}_2} = 8.0 \text{ atm}\)

\[
Q = \frac{[\text{CO}][\text{H}_2]}{[\text{H}_2\text{O}]} = \frac{(0.20)(8.0)}{0.10} = 16
\]

Since \( Q > K \), the reaction will shift to the left (\( \searrow \)) to come to equilibrium.
(4) 3. For each of the following equilibria, predict the direction the equilibrium will shift (right/left/no shift) if the indicated stress is applied.

<table>
<thead>
<tr>
<th>Stress</th>
<th>Direction</th>
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<tbody>
<tr>
<td>H₂(g) + F₂(g) → 2 HF(g) + heat</td>
<td>Increase Volume</td>
</tr>
</tbody>
</table>
* (Equal moles of gas on each side) |

2NO(g) + O₂(g) → NO₂(g) + heat
Decrease Temperature | Right |
* (Since heat is effectively a product, removing heat will cause a shift to the right) |

N₂(g) + 3 H₂(g) → 2 NH₃(g)
Liquefy ammonia | Right |
* (Liquefying the ammonia effectively removes it from the gas phase – the equilibrium will shift to make up for the loss of ammonia.) |

HF(aq) + HOH → F⁻(aq) + OH⁻(aq)
Put something in solution that precipitates the F⁻ ion | Right |
* (Precipitating the F⁻ out as a solid will remove F⁻ from the solution – the equilibrium shifts to make up for the loss.) |

(3) 4. Consider the equilibrium:

\[ 2 \text{NOBr} \rightarrow 2 \text{NO} + \text{Br}_2 \]

The equilibrium constant for this reaction is \( K'_e \). Write the equilibrium constant for each of the following systems in terms of \( K'_e \). For example, if the equilibrium constant for the expression should be the cube of the \( K'_e \), write \((K'_e)^3\).

\[ 2 \text{NO} + \text{Br}_2 \rightarrow 2 \text{NOBr} \]

\[ K_e = \frac{1}{K'_e} \] Reaction is reverse of original

\[ \text{NO} + \frac{1}{2} \text{Br}_2 \rightarrow \text{NOBr} \]

\[ K_e = \left(\frac{1}{K'_e}\right)^{1/2} \] Reaction is one-half of the reverse of the original

\[ \text{NOBr} \rightarrow \text{NO} + \frac{1}{2} \text{Br}_2 \]

\[ K_e = (K'_e)^{1/2} \] Reaction is one-half the original