Equilibrium Worksheet 5 Name Key

Solve each of the following problems. Show your work in the space provided. Write your final answer on the blank line.

Part A

1. Write an equilibrium expression for the following reaction:

$A_{(g)} + B_{(g)} <=====> AB_{(g)}$

Then calculate the value of K_{eq} given that [A] = 1.1×10^{-3}M, [B] = 4.4M, and [AB] = 1.5×10^{-8}M. Finally, tell whether reactants or products are favored, and why.

$$K_{eq} = \frac{[AB]}{[A][B]} = \frac{(1.5 \times 10^{-8})}{(1.1 \times 10^{-3})(4.4)} = 3.1 \times 10^{-6}$$

Reactants are favored because Keq is {small or less than 1}

2. Write an equilibrium expression for the following reaction:

$A_{2(g)} + B_{(s)} <=====> A_2B_{(g)}$

Then calculate the value of K_{eq} given that [A_2] = 1.9×10^{-3}M, and [A_2B] = 1.4×10^{-5}M. Finally, tell whether reactants or products are favored, and why.

$$K_{eq} = \frac{[A_2B]}{[A_2]} = \frac{1.4 \times 10^{-5}}{1.9 \times 10^{-3}} = 7.4 \times 10^{-3}$$

Reactants are favored because keq is {small or less than 1}

3. Write an equilibrium expression for the following reaction:

$2A_{(g)} + B_{(g)} <=====> A_2B_{(g)}$

Then calculate the value of K_{eq} given that [A] = 1.0×10^{-6}M, [B] = 2.2×10^{-4}M, and [A_2B] = 6.5×10^{-1}M. Finally, tell whether reactants or products are favored, and why.

$$K_{eq} = \frac{[A_2B]}{[A]^2[B]} = \frac{6.5 \times 10^{-1}}{(1.0 \times 10^{-6})^2(2.2 \times 10^{-4})} = 3.0 \times 10^{15}$$

Products are favored because keq is {greater than 1}

Key.

4. **Write an equilibrium expression for the following reaction:**

$A_{(s)} + B_{2(g)} <====> AB_{(s)} + B_{(g)}$

Then calculate the value of K_{eq} given that $[B_2] = 5.5 \times 10^{-4}M$, and $[B] = 3.9 \times 10^{-7}M$. Finally, tell whether reactants or products are favored, and why.

$K_{eq} = \frac{[B]}{[B_2]} = \frac{3.9 \times 10^{-7}}{5.5 \times 10^{-4}} = 7.1 \times 10^{-4}$

Reactants are favored because K_{eq} is less than 1.

5. Write an equilibrium expression for the following reaction:

$2A_{(g)} + 3B_{(g)} <====> A_2B_{3(s)}$

Then calculate the value of K_{eq} given that $[A] = 4.6 \times 10^{-3}M$, and $[B] = 1.5 \times 10^{-5}M$. Finally, tell whether reactants or products are favored, and why.

$K_{eq} = \frac{1}{[A]^2[B]^3} = \frac{1}{(4.6 \times 10^{-3})^2(1.5 \times 10^{-5})^3} = 1.4 \times 10^{19}$

Products are favored because K_{eq} is greater than 1.

Part B

6. Write an equilibrium expression for the following reaction:

$AB_{2(s)} + \text{energy} <====> A_{(s)} + B_{2(g)}$

Then calculate the value of K_{eq} given that $[B_2] = 1.3 \times 10^{-9}M$. Finally, predict the effect of increased temperature on the value of K_{eq} and explain your answer.

$K_{eq} = [B_2] = 1.3 \times 10^{-9}$

Increasing temperature will shift the eqm to the products, increasing $[B_2]$ and increasing K_{eq} because this reaction is endothermic

Key

7. Write an equilibrium expression for the following reaction:

$2A_{(g)} + B_{(g)} + \text{energy} <=====> A_2B_{(g)}$

Then calculate the value of K_{eq} given that [A] = 1.6×10^{-2}M, [B] = 1.4×10^{-4}M, and $[A_2B] = 3.6 \times 10^{-1}$ M. Finally, predict the effect of decreased temperature on the value of K_{eq} and explain your answer.

$$K_{eq} = \frac{[A_2B]}{[A]^2[B]} = \frac{3.6\times10^{-1}\,M}{(1.6\times10^{-2}M)^2(1.4\times10^{-4}M)} = 1.0\times10^{7}$$

This is an endothermic reaction, decreasing the temperature shifts the equilibrium towards the reactants and decreases K_{eq}.

Part C

8. Write an equilibrium expression for the following reaction:

$A_{2(g)} + B_{(s)} <=====> A_{(g)} + AB_{(s)}$

Then calculate the concentration of $A_{(g)}$ given that $K_{eq} = 1.5 \times 10^{-3}$M, and $[A_2] = 2.5 \times 10^{-4}$ M. Finally, predict the effect of adding some $A_{2(g)}$ on the values for [A], and explain your answer.

$$K_{eq} = \frac{[A]}{[A_2]} = 1.5\times10^{-3} = \frac{[A]}{2.5\times10^{-4}M} \qquad [A] = 3.8\times10^{-7}\,M$$

Adding some $A_{2(g)}$ will increase the value for [A] but will not change the value of K_{eq}.

9. Write an equilibrium expression for the following reaction:

$2A_{(g)} + B_{2(g)} <=====> A_2B_{(g)} + B_{(g)}$

Then calculate the concentration of A_2B, given that $K_{eq} = 7.1 \times 10^{+}$ M, [A] = 1.9×10^{-2}M, $[B_2] = 4.1 \times 10^{-3}$M, and [B] = 8.4×10^{-3} M. Finally, predict the effect of adding some $A_{(g)}$ on the values for $[B_2]$, $[A_2B]$, and [B], and explain your answer.

$$K_{eq} = \frac{[A_2B][B]}{[A]^2[B_2]} = 7.1\times10^{+} = \frac{[A_2B](8.4\times10^{-3}M)}{(1.9\times10^{-2}M)^2(4.1\times10^{-3}M)}$$

$$[A_2B] = 1.2\times10^{1} \sim 1.3\times10^{1}$$

Adding $A_{(g)}$ will shift the equilibrium to the products; $[B_2]$ will decrease, $[A_2B]$ and [B] will increase. K_{eq} will not change.