Section 18.3 Using Equilibrium Constants

In your textbook, read about calculating equilibrium concentrations.

Answer the following questions.

1. What can you use the equilibrium constant to do?

____

____

2. Given the reaction: $N_2 + O_2 \rightleftharpoons 2NO$ for which the K_{eq} at 2273 K is 1.2×10^{-4}

a. Write the equilibrium constant expression for the reaction.

____

b. Write the equation that would allow you solve for the concentration of NO.

____

c. What is the concentration of NO if $[N_2] = 0.166M$ and $[O_2] = 0.145M$?

____

____

____

3. What is the solubility product constant?

____

____

4. What is the solubility product constant expression for the reaction:
$Mg_3(PO_4)_2(s) \rightleftharpoons 3Mg^{2+}(aq) + 2PO_4{}^{3-}(aq)$

____

5. Given the equilibrium $BaSO_4(s) \rightleftharpoons Ba^{2+}(aq) + SO_4{}^{2-}(aq)$, what is the solubility product constant expression?

____

6. The solubility product constant for $BaSO_4$ at 298 K is 1.1×10^{-10}. Calculate the solubility of $BaSO_4$ in mol/L at 298 K.

____

____

____

____

Section 18.3 ***continued***

In your textbook, read about predicting precipitates.

The solubility product constant can be used to determine if a precipitate will form when two aqueous solutions are mixed together. First, calculate the concentrations of the ions in the final solution. Use the solubility product constant expression to calculate the ion product (Q_{sp}) for the substance that might precipitate. Compare the result with the K_{sp} of the substance.

7. What can you say about a solution when

a. Q_{sp} is greater than K_{sp}?

____

____

b. Q_{sp} is equal to K_{sp}?

____

____

c. Q_{sp} is less than K_{sp}?

____

____

8. Predict whether a precipitate of AgBr will form if 100 mL of 0.0025*M* $AgNO_3$ and 100 mL of 0.0020*M* NaBr are mixed.

____

____

____

____

____

9. Explain briefly why Ag_3PO_4 might be more soluble in water than in the same volume of a solution containing Na_3PO_4.

____

____

____

____

____