

## 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?

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2. Given the reaction:  $\text{N}_2 + \text{O}_2 \rightleftharpoons 2\text{NO}$  for which the  $K_{\text{eq}}$  at 2273 K is  $1.2 \times 10^{-4}$

a. Write the equilibrium constant expression for the reaction.

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

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c. What is the concentration of NO if  $[\text{N}_2] = 0.166\text{M}$  and  $[\text{O}_2] = 0.145\text{M}$ ?

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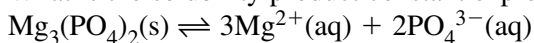
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3. What is the solubility product constant?

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4. What is the solubility product constant expression for the reaction:



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5. Given the equilibrium  $\text{BaSO}_4(\text{s}) \rightleftharpoons \text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$ , what is the solubility product constant expression?

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6. The solubility product constant for  $\text{BaSO}_4$  at 298 K is  $1.1 \times 10^{-10}$ . Calculate the solubility of  $\text{BaSO}_4$  in mol/L at 298 K.

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**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  $K_{sp}$ ?

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

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

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8. Predict whether a precipitate of AgBr will form if 100 mL of 0.0025M AgNO<sub>3</sub> and 100 mL of 0.0020M NaBr are mixed.

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9. Explain briefly why Ag<sub>3</sub>PO<sub>4</sub> might be more soluble in water than in the same volume of a solution containing Na<sub>3</sub>PO<sub>4</sub>.

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