Chemical Equilibrium

Practice Problems Set 1

For the following three reactions,

a) write the $K_c$ expression (i.e. the law of mass action) in terms of concentration, $K_c$.

b) given the equilibrium concentrations, state whether each equilibrium is product-favored, reactant-favored, or fairly even ($[\text{products}] \approx [\text{reactants}]$).

c) calculate the value of $K_c$.

1. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
   At equilibrium: $[N_2] = 1.50 \text{ M}$; $[H_2] = 2.00 \text{ M}$; $[NH_3] = 0.01 \text{ M}$

2. $HF(aq) \rightleftharpoons H^+(aq) + F^-(aq)$
   At equilibrium: $[HF] = 0.55 \text{ M}$; $[H^+] = 0.001 \text{ M}$; $[F^-] = 0.001 \text{ M}$

3. $Fe^{3+}(aq) + SCN^-(aq) \rightleftharpoons FeSCN^{2+}(aq)$
   At equilibrium: $[Fe^{3+}] = 0.55 \text{ M}$; $[SCN^-] = 0.001 \text{ M}$; $[FeSCN^{2+}] = 0.001 \text{ M}$

Summarize:

Fill in the blanks with product-favored, reactant-favored, and approximately equal $K_c$ states of equilibrium

<table>
<thead>
<tr>
<th>$K_c$</th>
<th>state of equilibrium</th>
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<tbody>
<tr>
<td>$K_c \gg 1$</td>
<td>product-favored</td>
</tr>
<tr>
<td>$K_c \ll 1$</td>
<td>reactant-favored</td>
</tr>
<tr>
<td>$K_c \approx 1$</td>
<td>approximately equal</td>
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</table>

4. Knowing that pure water has a density of 1g/1mL calculate the mass of 1.00 Liter of water.

Calculate the number of moles in 1.00 L of H$_2$O.

What is the concentration (M) of water in water?

At this temperature, can you get more moles of water into this Liter of water?

The $[H_2O]$ ________ (is / is not) constant.
Important Note:

Since the concentrations of solids and liquids are constant, they are not incorporated into the equilibrium constant, $K_{eq}$. That means, just leave them out of the $K_c$ or $K_p$ expression. Only include (g) and (aq) reactants and products!

5. Write equilibrium expressions (i.e. the law of mass action) for each of the following reactions:

a) $\text{CaCO}_3(s) \rightleftharpoons \text{CaO}(s) + \text{CO}_2(g)$

b) $\text{Ni}(s) + 4\text{CO}(g) \rightleftharpoons \text{Ni(CO)}_4(g)$

c) $5\text{CO}(g) + \text{I}_2\text{O}_5(s) \rightleftharpoons \text{I}_2(g) + 5\text{CO}_2(g)$

d) $\text{Ca(HCO}_3)_2(\text{aq}) \rightleftharpoons \text{CaCO}_3(s) + \text{H}_2\text{O}(l) + \text{CO}_2(g)$

e) $\text{AgCl}(s) \rightleftharpoons \text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq})$

6. Write the equilibrium expression (i.e. the law of mass action) in terms of partial pressures ($K_p$) for each of the following reactions.

Rate the reactions in order of their increasing tendency to proceed toward completion:

(a) $4\text{NH}_3(g) + 3\text{O}_2(g) \rightleftharpoons 2\text{N}_2(g) + 6\text{H}_2\text{O}(g)$ \hspace{1cm} $K_p = 1 \times 10^{228}$

(b) $\text{N}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{NO}(g)$ \hspace{1cm} $K_p = 5 \times 10^{-31}$

(c) $2\text{HF}(g) \rightleftharpoons \text{H}_2(g) + \text{F}_2(g)$ \hspace{1cm} $K_p = 1 \times 10^{-13}$

(d) $2\text{NOCl}(g) \rightleftharpoons 2\text{NO}(g) + \text{Cl}_2(g)$ \hspace{1cm} $K_p = 4.7 \times 10^{-4}$

7. (a) Write the $K_c$ expression for $2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g)$ Calculate the value of $K_c$:

At equilibrium: $[\text{SO}_2] = 1.50 \ M$; $[\text{O}_2] = 1.25 \ M$; $[\text{SO}_3] = 3.50 \ M$

(b) If we reverse the equation, it is: $2\text{SO}_3(g) \rightleftharpoons 2\text{SO}_2(g) + \text{O}_2(g)$

Write the $K_c$ expression for this equation and calculate the new value of $K_c$:

How does the expression and the value of $K_c$ in 7(b) compare with those in 7(a)?

(c) If we now multiply all of the coefficients by $\frac{1}{2}$: $\text{SO}_3(g) \rightleftharpoons \text{SO}_2(g) + \frac{1}{2} \text{O}_2(g)$

Write the $K_c$ expression for this equation and calculate the new value of $K_c$:

How do they compare with 7(b)?

(d) What would happen to the $K_c$ expression and its value if we doubled the coefficients?
Chemical Equilibrium

**PRACTICE PROBLEM SET #2**

1. Consider the equilibrium: \(2 \text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2 \text{SO}_3(g)\) \[K_c = 4.36 \text{ M}^{-1}\]

   Calculate the value of “\(Q\)” for a situation in which the concentrations are \([\text{SO}_2] = 2.00 \text{ M}, [\text{O}_2] = 1.50 \text{ M}, \text{ and } [\text{SO}_3] = 1.25 \text{ M}].

   Does this mixture shift toward the reactants or products to reach equilibrium? _________________

2. Study the discussion in your textbook about converting \(K_c\) and \(K_p\). Write the \(K_p\) expression for the reaction in question 1 and calculate its value at 0°C. Remember, \(R = 0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K}\).

3. Consider the equilibrium \(\text{PCl}_3(g) + \text{Cl}_2(g) \rightleftharpoons \text{PCl}_5(g)\).

   How would the following changes affect the partial pressures of each gas at equilibrium?
   
   (\(\uparrow\) = increase; \(\downarrow\) = decrease; \(=\) unchanged)

   \[
   \text{PCl}_3(g) + \text{Cl}_2(g) \rightleftharpoons \text{PCl}_5(g)
   \]

   a) addition of \(\text{PCl}_3\) \(\uparrow\) \(\downarrow\) \(=\) 

   b) removal of \(\text{Cl}_2\) \(\uparrow\) \(\downarrow\) \(=\) 

   c) removal of \(\text{PCl}_5\) \(\uparrow\) \(\downarrow\) \(=\) 

   d) decrease in the volume of the container \(\uparrow\) \(\downarrow\) \(=\) 

   e) addition of He without change in volume \(\uparrow\) \(\downarrow\) \(=\) 

4. How will each of the changes in question 3 affect the value for the \(K_{eq}\)?

   (\(\uparrow\) = increase; \(\downarrow\) = decrease; \(=\) unchanged)

   a \(\uparrow\) b \(\downarrow\) c \(=\) d \(\uparrow\) e \(\uparrow\)

5. Indicate how each of the following changes affects the amount of each gas in the system below, for which \(\Delta H_{\text{reaction}} = +9.9 \text{ kcal}\). (\(\uparrow\) = increase; \(\downarrow\) = decrease; \(=\) unchanged)

   \[
   \text{H}_2(g) + \text{CO}_2(g) \rightleftharpoons \text{H}_2\text{O}(g) + \text{CO}(g)
   \]

   a) addition of \(\text{CO}_2\) \(\uparrow\) \(\downarrow\) \(=\) 

   b) addition of \(\text{H}_2\text{O}\) \(\uparrow\) \(\downarrow\) \(=\) 

   c) addition of a catalyst \(\uparrow\) \(\downarrow\) \(=\) 

   d) increase in temperature \(\uparrow\) \(\downarrow\) \(=\) 

   e) decrease in the volume of the container \(\uparrow\) \(\downarrow\) \(=\) 

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Summarize:

<table>
<thead>
<tr>
<th>Equation</th>
<th>(K_c) expression &amp; Value</th>
</tr>
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<tbody>
<tr>
<td>doubled</td>
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<td>reversed</td>
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<td>halved</td>
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</table>
6. How will each of the changes in question 5 affect the equilibrium constant? 
   (↑ = increase; ↓ = decrease; = unchanged)
   
   a ____ b ____ c ____ d ____ e ____

7. Consider the equilibrium: \(2 \text{N}_2\text{O}(g) + \text{O}_2(g) \rightleftharpoons 4\text{NO}(g)\)
   
   How will the amount of chemicals at equilibrium be affected by
   \(2 \text{N}_2\text{O}(g) + \text{O}_2(g) \rightleftharpoons 4\text{NO}(g)\)
   
   a) adding \(\text{N}_2\text{O}\) ____ ____ ____
   b) removing \(\text{O}_2\) ____ ____ ____
   c) increasing the volume of the container ____ ____ ____
   d) adding a catalyst ____ ____ ____

8. For the reaction, \(4\text{NH}_3(g) + 3\text{O}_2(g) \rightleftharpoons 2\text{N}_2(g) + 6\text{H}_2\text{O}(l)\)

   How will the concentration of each chemical be affected by
   a) adding \(\text{O}_2\) to the system ____ ____ ____ ____ ____
   b) adding \(\text{N}_2\) to the system ____ ____ ____ ____ ____
   c) removing \(\text{H}_2\text{O}\) from the system ____ ____ ____ ____ ____
   d) decreasing the volume of the container ____ ____ ____ ____ ____

9. Consider the equilibrium: \(2\text{N}_2\text{O}(g) + \text{O}_2(g) \rightleftharpoons 4\text{NO}(g)\)

   3.00 moles of \(\text{NO}(g)\) are introduced into a 1.00-Liter evacuated flask. When the system comes to equilibrium, 1.00 mole of \(\text{N}_2\text{O}(g)\) has formed. Determine the equilibrium concentrations of each substance. Calculate the \(K_c\) for the reaction based on these data.

<table>
<thead>
<tr>
<th></th>
<th>2 (\text{N}_2\text{O})</th>
<th>(\text{O}_2)</th>
<th>4 (\text{NO})</th>
</tr>
</thead>
<tbody>
<tr>
<td>initial</td>
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<tr>
<td>change</td>
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<tr>
<td>equilibrium</td>
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</table>

Remember: The “ice” box may be used with **moles**, **molarity**, or **Liters** (for gaseous equilibria)... never grams.
Chemical Equilibrium

1. Write the expressions for the equilibrium constant $K_c$ for the following reactions:
   a. $4 \text{NH}_3(g) + 7 \text{O}_2(g) \rightleftharpoons 4 \text{NO}_2(g) + 6 \text{H}_2\text{O}(l)$
   b. $\text{HCN(aq)} + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{CN}^-(aq)$
   c. $\text{PCl}_5(g) \rightleftharpoons \text{PCl}_3(g) + \text{Cl}_2(g)$
   d. $\text{CaCO}_3(s) \rightleftharpoons \text{CaO}(s) + \text{CO}_2(g)$
   e. $3 \text{O}_2(g) \rightleftharpoons 2 \text{O}_3(g)$
   f. $2 \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{OH}^-(aq)$
   g. $3 \text{Zn}(s) + 2 \text{Fe}^{3+}(aq) \rightleftharpoons 2 \text{Fe}(s) + 3 \text{Zn}^{2+}(aq)$

2. Write the equilibrium constant expressions for the following reactions. How are they related to one another?
   a. $2 \text{N}_2\text{O}(g) + 3 \text{O}_2(g) \rightleftharpoons 4 \text{NO}_2(g)$
   b. $\text{N}_2\text{O}(g) + \frac{3}{2} \text{O}_2(g) \rightleftharpoons 2 \text{NO}_2(g)$
   c. $4 \text{NO}_2(g) \rightleftharpoons 2 \text{N}_2\text{O}(g) + 3 \text{O}_2(g)$

3. Calculate the value of the equilibrium constant for the following system, given the data shown:
   $$\text{H}_2(g) + \text{CO}_2(g) \rightleftharpoons \text{H}_2\text{O}(g) + \text{CO}(g)$$
   Concentrations at equilibrium:
   
   \begin{align*}
   [\text{H}_2] &= 1.5 \text{ mol L}^{-1} \quad [\text{H}_2\text{O}] = 0.5 \text{ mol L}^{-1} \\
   [\text{CO}_2] &= 2.5 \text{ mol L}^{-1} \quad [\text{CO}] = 3.0 \text{ mol L}^{-1}
   \end{align*}

4. Chlorine molecules will dissociate at high temperatures into chlorine atoms. At 3000°C, for example, $K_c$ for the equilibrium shown is 0.55. If the partial pressure of chlorine molecules is 1.5 atm, calculate the partial pressure of the chlorine atoms:
   $$\text{Cl}_2(g) \rightleftharpoons 2 \text{Cl}(g)$$

5. Suppose that 0.50 moles of hydrogen gas, 0.50 moles of iodine gas, and 0.75 moles of hydrogen iodide gas are introduced into a 2.0 Liter vessel and the system is allowed to reach equilibrium.
   $$\text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2 \text{HI}(g)$$
   Calculate the concentrations of all three substances at equilibrium. At the temperature of the experiment, $K_c$ equals $2.0 \times 10^{-2}$.

6. If the mechanism of a chemical equilibrium consists of two reversible elementary steps, each with its own equilibrium constant $K_{c1}$ and $K_{c2}$, what expression relates the equilibrium constant $K_c$ for the overall equilibrium to the two constants $K_{c1}$ and $K_{c2}$?

7. When 2.0 mol of carbon disulfide and 4.0 mol of chlorine are placed in a 1.0 Liter flask, the following equilibrium system results. At equilibrium, the flask is found to contain 0.30 mol of carbon tetrachloride. What quantities of the other components are present in this equilibrium mixture?
   $$\text{CS}_2(g) + 3 \text{Cl}_2(g) \rightleftharpoons \text{S}_2\text{Cl}_2(g) + \text{CCl}_4(g)$$

8. 3.0 moles each of carbon monoxide, hydrogen, and carbon are placed in a 2.0 Liter vessel and allowed to come to equilibrium according to the equation: $\text{CO}(g) + \text{H}_2(g) \rightleftharpoons \text{C}(s) + \text{H}_2\text{O}(g)$
   If the equilibrium constant at the temperature of the experiment is 4.0, what is the equilibrium concentration of water vapor?
9. Nitrosyl chloride NOCl decomposes to nitric oxide and chlorine when heated:

\[ 2 \text{NOCl}(g) \rightleftharpoons 2 \text{NO}(g) + \text{Cl}_2(g) \]

At 600K, the equilibrium constant \( K_p \) is 0.060. In a vessel at 600K, there is a mixture of all three gases. The partial pressure of NOCl is 675 torr, the partial pressure of NO is 43 torr and the partial pressure of chlorine is 23 torr.

a. What is the value of the reaction quotient?
b. Is the mixture at equilibrium?
c. In which direction will the system move to reach equilibrium?
d. When the system reaches equilibrium, what will be the partial pressures of the components in the system?

10. Sulfuryl chloride decomposes at high temperatures to produce sulfur dioxide and chlorine gases:

\[ \text{SO}_2\text{Cl}_2(g) \rightleftharpoons \text{SO}_2(g) + \text{Cl}_2(g) \]

At 375°C, the equilibrium constant \( K_c \) is 0.045. If there are 2.0 grams of sulfuryl chloride, 0.17 gram of sulfur dioxide, and 0.19 gram of chlorine present in a 1.0 Liter flask,

a. What is the value of the reaction quotient?
b. Is the system at equilibrium?
c. In which direction will the system move to reach equilibrium?

11. Ammonium chloride is placed inside a closed vessel where it comes into equilibrium at 400°C according to the equation shown. Only these three substances are present inside the vessel. If \( K_p \) for the system at 400°C is 0.640, what is the pressure inside the vessel?

\[ \text{NH}_4\text{Cl}(s) \rightleftharpoons \text{NH}_3(g) + \text{HCl}(g) \]

12. Bromine and chlorine react to produce bromine monochloride according to the equation. \( K_c = 36.0 \) under the conditions of the experiment. \( \text{Br}_2(g) + \text{Cl}_2(g) \rightleftharpoons 2 \text{BrCl}(g) \)

If 0.180 moles of bromine gas and 0.180 moles of chlorine gas are introduced into a 3.0 Liter flask and allowed to come to equilibrium, what is the equilibrium concentration of the bromine monochloride? How much BrCl is produced?

13. When ammonia is dissolved in water, the following equilibrium is established. If the equilibrium constant is \( 1.8 \times 10^{-5} \), calculate the hydroxide ion concentration in the solution if 0.100 mol of ammonia is dissolved in sufficient water to make 500 mL of solution.

\[ \text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq) \]

14. The following reaction is exothermic:

\[ \text{Ti}(s) + 2 \text{Cl}_2(g) \rightleftharpoons \text{TiCl}_4(g) \]

List all the ways the yield of the product TiCl\(_4\) could be increased.
Acid - Base Equilibria

**Practice Problems**

1. For the following aqueous equilibria, designate the Brønsted-Lowry conjugate acid-base pairs and establish the weaker side:
   a. \( \text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-\text{(aq)} \)
   b. \( \text{HCN}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{CN}^-\text{(aq)} \)
   c. \( \text{NH}_4^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightleftharpoons \text{NH}_3(\text{aq}) + \text{HCO}_3^-\text{(aq)} \)

2. Write the formula for the conjugate bases of the following:
   a. \( \text{HNO}_2 \)
   b. \( \text{H}_2\text{SO}_4 \)
   c. \( \text{H}_2\text{PO}_4^- \)
   d. \( \text{HF} \)
   e. \( \text{CH}_3\text{CO}_2\text{H} \)

3. Complete the Brønsted-Lowry equilibria, label the components acid or base, and pair up the conjugate acid-base pairs:
   a. \( \text{HSO}_4^- + \text{H}_2\text{O} \rightleftharpoons \)
   b. \( \text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \)
   c. \( \text{CN}^- + \text{H}_2\text{O} \rightleftharpoons \)
   d. \( \text{H}^- + \text{H}_2\text{O} \rightleftharpoons \)
   e. \( \text{HClO}_4 + \text{H}_2\text{O} \rightleftharpoons \)

4. Is the monohydrogenphosphate ion \( \text{HPO}_4^{2-} \) amphiprotic?
   If so, write the formulas of its conjugate acid and its conjugate base.

5. Of the following acids, determine
   a. The strongest acid
   b. The acid that produces the lowest concentration of hydronium ions per mole of acid
   c. The acid with the strongest conjugate base
   d. The diprotic acid
   e. The strong acid
   f. The acid with the weakest conjugate base

\[
\begin{align*}
\text{HNO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) & \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{NO}_3^-\text{(aq)} & K_a & = \text{very large} \\
\text{HSO}_4^-\text{(aq)} + \text{H}_2\text{O}(\text{l}) & \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{SO}_4^{2-}\text{(aq)} & K_a & = 1.2 \times 10^{-2} \\
\text{HCN}(\text{aq}) + \text{H}_2\text{O}(\text{l}) & \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{CN}^-\text{(aq)} & K_a & = 4.0 \times 10^{-10} \\
\text{H}_2\text{CO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) & \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{HCO}_3^-\text{(aq)} & K_a & = 4.2 \times 10^{-7} \\
\text{NH}_4^+(\text{aq}) + \text{H}_2\text{O}(\text{l}) & \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{NH}_3(\text{aq}) & K_a & = 5.6 \times 10^{-10} \\
\text{HF}(\text{aq}) + \text{H}_2\text{O}(\text{l}) & \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{F}^-\text{(aq)} & K_a & = 7.2 \times 10^{-4}
\end{align*}
\]
6. Write net ionic acid-base reactions for:
   a. The reaction of acetic acid with aqueous ammonia solution
   b. The reaction of hydrofluoric acid with sodium hydroxide
   c. The reaction of ammonium chloride with potassium hydroxide
   d. The reaction of sodium bicarbonate with sulfuric acid
   e. The reaction of chlorous acid with aqueous ammonia solution
   f. The reaction of disodium hydrogen phosphate with acetic acid

7. What is the pH of
   a. 0.0010 M HCl solution?
   b. 0.15 M KOH solution?
   c. 10⁻⁸ M HNO₃ solution?

8. List the following substances in order of increasing acid strength: *(Look up the Kₐ's in the appendix of your textbook)*
   H₂O, H₂SO₃, HCN, H₂PO₄⁻, NH₄⁺, [Cu(H₂O)₆]²⁺, NH₃, H₃O⁺, HCO₂H, HCl.

9. Complete the table for each aqueous solution at 25°C.
   State whether the solutions are acidic or basic.

<table>
<thead>
<tr>
<th>[H₃O⁺]</th>
<th>[OH⁻]</th>
<th>pH</th>
<th>pOH</th>
<th>acidic or basic</th>
</tr>
</thead>
<tbody>
<tr>
<td>2.0 x 10⁻⁵</td>
<td></td>
<td>6.25</td>
<td></td>
<td></td>
</tr>
<tr>
<td>5.6 x 10⁻²</td>
<td></td>
<td>9.20</td>
<td></td>
<td></td>
</tr>
<tr>
<td>8.7 x 10⁻¹⁰</td>
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</table>

10. What is the pH of a solution that contains 2.60 grams of NaOH in 250 mL of aqueous solution?

11. If the pH of a sample of rainwater is 4.62, what is the hydronium ion concentration [H₃O⁺] and the hydroxide ion concentration [OH⁻] in the rainwater?

12. A 0.12 M solution of an unknown weak acid has a pH of 4.26 at 25°C. What is the hydronium ion concentration in the solution and what is the value of its Kₐ?

13. Hydroxylamine is a weak base with a Kₐ = 6.6 x 10⁻⁹. What is the pH of a 0.36 M solution of hydroxylamine in water at 25°C?

14. Suppose you dissolved benzoic acid in water to make a 0.15 M solution. Kₐ for benzoic acid = 6.3 x 10⁻⁵ at 25°C. What is:
   a. the concentration of benzoic acid?
   b. the concentration of hydronium ion?
   c. the concentration of benzoate anion?
   d. the pH of the solution?
15. Which of the following salts, when dissolved in water to produce 0.10 M solutions, would have the lowest pH?
   a. sodium acetate
   b. potassium chloride
   c. sodium bisulfate
   d. magnesium nitrate
   e. potassium cyanide

16. For each of the following salts, predict whether an aqueous solution would be acidic, basic, or neutral.
   a. sodium nitrate NaNO₃
   b. ammonium iodide NH₄I
   c. sodium bicarbonate NaHCO₃
   d. ammonium cyanide NH₄CN
   e. sodium hypochlorite NaOCl
   f. potassium acetate KCH₃CO₂

17. a. Cyanic acid HOCN has a $K_a = 3.5 \times 10^{-4}$, what is the $K_b$ for the cyanate ion $\text{OCN}^-$?
   b. Phenol is a relatively weak acid, $K_a = 1.3 \times 10^{-10}$. How does the strength of its conjugate base compare with the strength of ammonia, the acetate ion, and sodium hydroxide?

18. a. What is the pH of a 0.80 M solution of sulfurous acid?
   b. What is the concentration of sulfite ion in a 0.80 M solution of sulfurous acid?
   c. What happens to the concentration of sulfite ion $\text{SO}_3^{2-}$ if the concentration of sulfurous acid is halved?

19. Identify the Lewis acid and the Lewis base in the following reactions:
   a. Boron trichloride reacts with chloride ion to produce $[\text{BCl}_4^-]$
   b. Nickel reacts with carbon monoxide to produce nickel tetracarbonyl $[\text{Ni(CO)}_4]$.
   c. Ammonia reacts with acetic acid to produce ammonium acetate.
   d. Sodium ions are solvated by water to produce $\text{Na}^+(\text{aq})$

20. Calculate the pH of a 0.35 M solution of potassium cyanide. $K_a$ for HCN = 4.0 x $10^{-10}$.

Note: Question #20.
In the answers, the authors forgot to change the $K_a$ into the $K_b$.
The $K_b$ should be $2.5 \times 10^{-5}$. The "x" = [OH⁻] = .002958; pOH = 2.53; pH = 11.47
1. HCN is a weak acid \((K_a = 6.2 \times 10^{-10})\). NH\(_3\) is a weak base \((K_b = 1.8 \times 10^{-5})\). A 1.0 M solution of NH\(_4\)CN would be
   (A) strongly acidic  (C) neutral
   (B) weakly acidic  (D) weakly basic

2. How many moles of HCOONa must be added to 1.0 L of 0.10 M HCOOH to prepare a buffer solution with a pH of 3.4? (HCOOH \(K_a = 2 \times 10^{-4}\))
   (A) 0.01  (C) 0.1
   (B) 0.05  (D) 0.2

3. The acid-base indicator methyl red has a \(K_a\) of 1 \(x\) \(10^{-4}\). Its acidic form is red while its alkaline form is yellow. If methyl red is added to a colorless solution with a pH = 7, the color will be
   (A) pink  (C) orange
   (B) red  (D) yellow

4. Which mixture forms a buffer when dissolved in 1.0 L of water?
   (A) 0.2 mol NaOH + 0.2 mol HBr
   (B) 0.2 mol NaCl + 0.3 mol HCl
   (C) 0.4 mol HNO\(_2\) + 0.2 mol NaOH
   (D) 0.5 mol NH\(_3\) + 0.5 mol HCl

5. A buffer solution is prepared in which the concentration of NH\(_3\) is 0.30 M and the concentration of NH\(_4\) is 0.20 M. What is the pH of this solution? The equilibrium constant, \(K_b\) for NH\(_3\) equals 1.8 \(x\) \(10^{-5}\).
   (A) 8.73  (C) 9.43
   (B) 9.08  (D) 11.72

6. For which titration would the use of phenolphthalein introduce a significant error?
   \(K_{\text{indicator}}\) for phenolphthalein = 1 \(x\) \(10^{-9}\)
   (A) \(\text{pH}\) 7  (B) \(\text{pH}\) 7
   (C) \(\text{pH}\) 7  (D) \(\text{pH}\) 7

7. The titration curves labeled 1 and 2 were obtained by titrating equal volumes of two different acid samples with portions of the same sodium hydroxide solution.
What conclusions can be drawn about the relative concentrations and strengths of acids 1 and 2 from these curves?
(A) The concentrations are the same but acid 1 is weaker than acid 2.
(B) The concentrations are the same but acid 1 is stronger than acid 2.
(C) Acid 1 is the same strength as acid 2, but it is less concentrated.
(D) Acid 1 is the same strength as acid 2, but it is more concentrated.

8. A 0.100 M solution of acetic acid ($K_a = 1.8 \times 10^{-5}$) is titrated with a 0.1000 M solution of NaOH. What is the pH when 50% of the acid has been neutralized?
(A) 2.38  (B) 4.74  (C) 5.70  (D) 7.00

9. The $pK_a$ values for several acid-base indicators are given in the table. Which indicator should be used in the titration of a weak base with a strong acid?

<table>
<thead>
<tr>
<th>Indicator</th>
<th>$pK_a$</th>
</tr>
</thead>
<tbody>
<tr>
<td>2,4-dintrophenol</td>
<td>3.5</td>
</tr>
<tr>
<td>bromthymol blue</td>
<td>7.0</td>
</tr>
<tr>
<td>cresol red</td>
<td>8.0</td>
</tr>
<tr>
<td>alizarin yellow R</td>
<td>11.0</td>
</tr>
</tbody>
</table>

(A) 2,4-dintrophenol  (B) bromthymol blue  (C) cresol red  (D) alizarin yellow R
Precipitation Reactions

Practice Problems

1. For an ionic compound with ions in a 1:1 ratio (e.g. NaCl), $K_{sp} = s^2$
   What is the relationship for $\text{Al}_2\text{S}_3$? $K_{sp} =$
   What is the relationship for $\text{Ag}_3\text{PO}_4$? $K_{sp} =$

2. The $K_{sp}$ of CuCN is $3.2 \times 10^{-20}$. What is the molar solubility of CuCN?

3. The $[\text{F}^-]$ in a saturated solution of BaF$_2$ is $1.5 \times 10^{-2}$ M. What is the $K_{sp}$ of BaF$_2$?

4. The solubility of lead(II) carbonate is $2.7 \times 10^{-7}$ mol L$^{-1}$. What is its $K_{sp}$?
   (A) $5.2 \times 10^{-4}$  (B) $2.7 \times 10^{-7}$  (C) $7.3 \times 10^{-14}$  (D) $3.9 \times 10^{-20}$

5. The solubility of PbI$_2$ is $0.0013$ mol L$^{-1}$. Use this information to find the $K_{sp}$ for PbI$_2$.
   (A) $1.7 \times 10^{-6}$  (B) $6.8 \times 10^{-6}$  (C) $2.2 \times 10^{-9}$  (D) $8.8 \times 10^{-9}$

6. The $K_{sp}$ of CuCl is $1.9 \times 10^{-7}$ at 25 °C. What is the solubility of CuCl in mol L$^{-1}$?
   (A) $3.6 \times 10^{-14}$  (B) $1.9 \times 10^{-7}$  (C) $4.4 \times 10^{-4}$  (D) $8.8 \times 10^{-4}$

7. The $K_{sp}$ of PbI$_2$(s) is $1.4 \times 10^{-8}$ at 25 °C. What is the solubility of PbI$_2$ in moles per liter?
   (A) $1.2 \times 10^{-4}$  (B) $1.5 \times 10^{-3}$  (C) $1.9 \times 10^{-3}$  (D) $2.4 \times 10^{-3}$