Representative Exam Questions
On The Topic of Equilibrium
(Includes Acid / Base Equilibria)

Detailed Answer Key – Answers and Explanations

1. If a chemical equilibrium very much favors the products over the reactants, what would we expect its equilibrium constant to be like?

   * a) We would expect the equilibrium constant to be a very large number.
   b) We would expect the equilibrium constant to be a very small number.
   c) We would not be able to make any prediction about the size of the equilibrium constant, because its numerical value conveys no information on what is favored in the equilibrium.

   The larger the equilibrium constant, the more the right (product) side is favored, and the smaller the equilibrium constant, the more the left (reactant) side is favored.

2. If a chemical reaction reaches equilibrium very quickly, what would we expect its equilibrium constant to be like?

   a) We would expect the equilibrium constant to be a very large number.
   b) We would expect the equilibrium constant to be a very small number.
   * c) We would not be able to make any prediction about the size of the equilibrium constant, because its numerical value conveys no information on the time required to reach equilibrium.

   The equilibrium constant tells us about the position of equilibrium, but not about the time required to get there.
3. What is always true about a chemical reaction when it is at equilibrium?

a) The concentrations of all the reactants are zero.

b) All substances shown in the reaction have equal concentrations.

* c) The forward reaction rate is equal to the reverse reaction rate.

d) Both the forward and reverse reaction have come to a stop.

Equilibrium is the name given to the condition of having two opposing processes occurring at the same rate. Although all concentrations remain constant at equilibrium, they are not necessarily equal to each other.

4. Chemical equilibrium is said to be

a) static  

b) dynamic

The forward and reverse reactions continue to occur when the system is at equilibrium.

5. A negative equilibrium constant is obtained

a) when a reaction is elementary

b) when a reaction is non-elementary

c) when the reaction is exothermic in the forward direction

d) when you reverse the direction of a reaction that initially had a positive equilibrium constant

* e) under no circumstances

Equilibrium constants are calculated by multiplying and dividing concentrations (for \( K_c \)) or pressures (for \( K_p \)) raised to various powers. Concentrations and pressures are always positive numbers, and there is no way to obtain a negative answer when positive numbers are multiplied, divided, or raised to powers.
6. For any chemical reaction that reaches an equilibrium, it is possible to write a balanced equilibrium equation and a corresponding equilibrium constant expression. On what factors will the numerical value of the equilibrium constant depend?

a) None. A given reaction has a characteristic numerical value for its equilibrium constant that applies under all conditions.

b) Temperature

c) The choice of coefficients used to balance the equation describing the reaction.

d) Whether the reaction is elementary or non-elementary

e) Both b and c

* e) Both b and c

Except for reactions that are neither endothermic nor exothermic (a very rare situation) changing the temperature shifts the position of equilibrium. Heating the system drives the reaction in the endothermic direction, and cooling the system drives the reaction in the exothermic direction. If a chemical equation is multiplied by a constant, the equilibrium constant is raised to that power. For example, multiplying a reaction by 2 results in the equilibrium constant being squared. Also, reversing the direction of a chemical reaction takes the reciprocal of the equilibrium constant. For example, a reaction that has an equilibrium constant of 2 when written in one direction has an equilibrium constant of 0.5 when written in the reverse direction.

7. Given the equilibrium reaction

\[ 2A(g) + 2BC(g) \rightleftharpoons AB_2(g) + AC_2(g) \quad K_C = 4 \]

what is the numerical value of the equilibrium constant \( K_C \) for the reaction

\[ AB_2(g) + AC_2(g) \rightleftharpoons 2A(g) + 2BC(g) \quad K_C = ? \]

a) -4  b) -2  *c) 0.25  d) 0.50  e) 4

The original reaction has been reversed. Therefore, we must take the reciprocal of the original equilibrium constant. \( \frac{1}{4} = 0.25 \)
8. Given the equilibrium reaction
\[ \text{A(g)} + 2\text{B(g)} \rightleftharpoons \text{AB}_2(g) \quad K_C = 9 \]

what is the numerical value of the equilibrium constant (K_C) for the reaction
\[ 2\text{A(g)} + 4\text{B(g)} \rightleftharpoons 2\text{AB}_2(g) \quad K_C = ? \]

a) 3  b) 4.5  c) 9  d) 18  *e) 81

The original reaction was multiplied by 2, so the equilibrium constant must be raised to the power 2, that is, squared. \( 9^2 = 81 \).

9. Given the equilibrium reactions
\[ \text{AB(g)} \rightleftharpoons \text{A(g)} + \text{B(g)} \quad K_C = 0.5 \]
\[ 2\text{A(g)} + \text{B}_2(g) \rightleftharpoons \text{A}_2\text{B(g)} + \text{B(g)} \quad K_C = 6 \]

what is the equilibrium constant (K_C) for the following reaction?
\[ 3\text{A(g)} + \text{B}_2(g) \rightleftharpoons \text{A}_2\text{B(g)} + \text{AB(g)} \quad K_C = ? \]

a) 3  b) 5.5  c) 6.5  *d) 12  e) 36

You can get the target reaction by reversing the first reaction and then adding the second reaction. Reversing the first reaction changes its equilibrium constant from 0.5 to 2. Recall that reversing a reaction takes the reciprocal of its equilibrium constant. \( 1 / 0.5 = 2 \). When chemical reactions are added, their equilibrium constants are multiplied. \( 2 \times 6 = 12 \).
10. The reaction

\[ 2\text{NH}_3(g) \rightleftharpoons \text{N}_2(g) + 3\text{H}_2(g) \]

has \( K_C = 6.46 \times 10^{-3} \) at 300 °C. What is the value of \( K_P \) for the reaction at this same temperature? Note that \( R = 0.0820584 \) (L atm) / (K mol). Temperatures must be in Kelvin units.

\[ K_P = K_C \cdot (RT)^{\Delta n} \quad \Delta n = (1+3) - 2 = 4 - 2 = 2 \]

\[ K_P = 6.46 \times 10^{-3} \cdot (0.0820584 \times 573.15)^2 = 14.3 \]

11. The reaction

\[ \text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO}_2(g) \]

has \( K_P = 0.145 \) at 25 °C. What is \( K_C \) for the reaction at this same temperature? Note that \( R = 0.0820584 \) (L atm) / (K mol). Temperatures must be in Kelvin units.

\[ K_C = K_P \cdot (RT)^{-\Delta n} \quad \Delta n = 2 - 1 = 1 \]

\[ K_C = 0.145 \cdot (0.0820584 \times 298.15)^{-1} = 5.93 \times 10^{-3} \]
12. Sulfur trioxide gas, SO₃, was put in a heated reaction vessel and maintained at a temperature of 700 °C. Some of the SO₃ decomposed into SO₂ and O₂ gases, and an equilibrium was established between the SO₃, SO₂, and O₂. The chemical equation describing this equilibrium is

\[ 2 \text{SO}_3(g) \rightleftharpoons 2\text{SO}_2(g) + \text{O}_2(g) \]

The equilibrium concentrations of these three gases were determined and found to be

\[ [\text{SO}_3]_{eq} = 1.14 \times 10^{-2} \text{ M} \]
\[ [\text{SO}_2]_{eq} = 8.64 \times 10^{-3} \text{ M} \]
\[ [\text{O}_2]_{eq} = 4.32 \times 10^{-3} \text{ M} \]

What is the numerical value of the equilibrium constant for this reaction at 700 °C?

a) $1.56 \times 10^{-4}$  

b) $2.48 \times 10^{-3}$  

c) $5.81 \times 10^{-2}$  

d) $4.72 \times 10^2$  

e) $6.38 \times 10^3$

\[
K_C = \frac{[\text{SO}_2]_{eq}^2 \times [\text{O}_2]_{eq}}{[\text{SO}_3]_{eq}^2} \\
K_C = \frac{(8.64 \times 10^{-3})^2 \times (4.32 \times 10^{-3})}{(1.14 \times 10^{-2})^2} = 2.48 \times 10^{-3}
\]
13. At 395 °C, an equilibrium mixture of CO, Cl₂, and COCl₂ was found to have the following concentrations:

\[
\begin{align*}
[\text{CO}]_{\text{eq}} &= 6.5 \times 10^{-3} \text{ M} \\
[\text{Cl}_2]_{\text{eq}} &= 1.4 \times 10^{-3} \text{ M} \\
[\text{COCl}_2]_{\text{eq}} &= \text{unspecified}
\end{align*}
\]

The equilibrium reaction for this chemical system is as follows:

\[
\text{CO(g)} + \text{Cl}_2(g) \rightleftharpoons \text{COCl}_2(g) \quad K_C = 1.2 \times 10^3 \text{ at 395 } ^\circ\text{C}
\]

What is the concentration of COCl₂ in this mixture at 395 °C?

a) \(7.4 \times 10^{-9} \text{ M}\)  
b) \(8.8 \times 10^{-6} \text{ M}\)  
c) \(3.5 \times 10^{-3} \text{ M}\)  
d) \(6.9 \times 10^{-3} \text{ M}\)  
\[\text{**e) } 1.1 \times 10^{-2} \text{ M}\]

\[
K_C = \frac{[\text{COCl}_2]_{\text{eq}}}{[\text{CO}]_{\text{eq}} \cdot [\text{Cl}_2]_{\text{eq}}}
\]

Solving this for \([\text{COCl}_2]_{\text{eq}}\) gives

\[
[\text{COCl}_2]_{\text{eq}} = K_C \cdot [\text{CO}]_{\text{eq}} \cdot [\text{Cl}_2]_{\text{eq}}
\]

\[
= 1.2 \times 10^3 \times (6.5 \times 10^{-3}) \times (1.4 \times 10^{-3})
\]

\[
= 1.1 \times 10^{-2}
\]
14. Into a 10.00 L reaction vessel, a chemist placed 0.381 moles of HI, 0.492 moles of H₂, and 0.773 moles of I₂ at 425 °C. The chemical equation for the equilibrium of HI, H₂, and I₂ is as follows:

\[ 2 \text{HI(g)} \quad \text{<---------->} \quad \text{H}_2(g) \quad + \quad \text{I}_2(g) \quad \text{K}_C = 1.84 \text{ at } 425 \, ^\circ \text{C} \]

What can be said about this chemical system?

a) This is an equilibrium mixture. No net reaction will occur.

*b) This is not an equilibrium mixture. Some of the H₂ and I₂ will combine to form HI until equilibrium is reached.

b) This is not an equilibrium mixture. Some of the HI will decompose into H₂ and I₂ until equilibrium is reached.

\[ [\text{HI}] = \frac{0.381 \text{ mol}}{10.00 \text{ L}} = 0.0381 \text{ mol L}^{-1} \]

\[ [\text{H}_2] = \frac{0.492 \text{ mol}}{10.00 \text{ L}} = 0.0492 \text{ mol L}^{-1} \]

\[ [\text{I}_2] = \frac{0.773 \text{ mol}}{10.00 \text{ L}} = 0.0773 \text{ mol L}^{-1} \]

\[ Q_C = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = \frac{(0.0492)(0.0773)}{(0.0381)^2} = 2.62 \]

\[ Q_C > K_C \text{ so the reaction will proceed to the left. This means some of the H}_2 \text{ and I}_2 \text{ will react to form HI.} \]
A 5.00 L reaction vessel at 3900 °C initially contained 0.200 moles each of N₂ and O₂ gases. These reacted to form NO as described by the following equation:

\[ \text{N}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{NO}(g) \quad K_C = 0.0123 \text{ at 3900 °C} \]

How many moles of NO will be present in the container when the system has reached equilibrium?

a) 8.91 x 10⁻³ mol
b) 2.10 x 10⁻² mol
c) 4.23 x 10⁻² mol
d) 7.66 x 10⁻² mol
e) 1.85 x 10⁻¹ mol

\[
\begin{array}{c|c|c|c}
\text{Initial} & \text{[N}_2\text{]} & \text{[O}_2\text{]} & \text{[NO]} \\
\text{Change} & -X & -X & +2X \\
\text{Equilibrium} & 0.0400 - X & 0.0400 - X & 2X \\
\end{array}
\]

\[
\frac{[\text{NO}]^2_{\text{eq}}}{[\text{N}_2]_{\text{eq}} \cdot [\text{O}_2]_{\text{eq}}} = K_C
\]

\[
\frac{(2X)^2}{(0.0400 - X)(0.0400 - X)} = 0.0123
\]

\[
\frac{(2X)^2}{(0.0400 - X)^2} = 0.0123
\]

2X

\[
\frac{2X}{(0.0400 - X)} = \pm 0.1109 \quad \text{Note: A square root has both a positive and a negative root}
\]

2X = \pm 0.1109 (0.0400 - X)
Case 1: positive root  Case 2: negative root

\[
2X = 0.1109 \ (0.0400 - X)
\]

\[
2X = -0.1109 \ (0.0400 - X)
\]

\[
2X = 0.004436 - 0.1109X
\]

\[
2X = -0.004436 + 0.1109X
\]

\[
2X + 0.1109X = 0.004436
\]

\[
2X - 0.1109X = -0.004436
\]

\[
(2 + 0.1109)X = 0.004436
\]

\[
(2 - 0.1109)X = -0.004436
\]

\[
2.1109X = 0.004436
\]

\[
1.8891X = -0.004436
\]

\[
X = 0.00210
\]

\[
X = -0.00235
\]

Since none of the equilibrium concentrations in the “ICE” table can be negative, the acceptable values for X are in the following range:

\[
0 < X < 0.0400
\]

Therefore, the X value obtained in Case 2 (-0.00235) must be rejected. The value of X obtained in Case 1 (0.00210) falls within the acceptable range. The equilibrium concentrations are now known:

\[
[N_2]_{eq} = 0.0400 - X = 0.0400 - 0.00210 = 0.0379
\]

\[
[O_2]_{eq} = 0.0400 - X = 0.0400 - 0.00210 = 0.0379
\]

\[
[NO]_{eq} = 2X = 2 \cdot (0.00210) = 0.00420
\]

To confirm that these are the equilibrium concentrations, substitute them back into the equilibrium constant expression and confirm that the original $K_c$ is obtained.

\[
K_c = \frac{(0.00420)^2}{(0.0379)(0.0379)} = 0.0123
\]

The equilibrium concentration of NO is 0.00420 mol L$^{-1}$. The total number of moles of NO in the 5.00 L container is

\[
0.00420 \text{ mol L}^{-1} \times 5.00 \text{ L} = 0.0210 \text{ mol}
\]

\[
= 2.10 \times 10^{-2} \text{ mol}
\]
16. What is the pH of a 3.88 M NH₃ solution? NH₃ is a weak base that establishes the following equilibrium when dissolved in water:

\[
\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+ (\text{aq}) + \text{OH}^- (\text{aq}) \quad K_b = 1.8 \times 10^{-5}
\]

The math is no worse than a quadratic, even if you don't approximate, but it should be ok to approximate in this problem.

\[a) 2.1 \quad b) 8.6 \quad c) 9.2 \quad d) 11.9 \quad e) 12.8\]

\[\text{[NH}_3\text{]}_{\text{initial}} = 3.88 \text{ mol L}^{-1}\]

\[\text{[NH}_4^+\text{]}_{\text{initial}} = 0 \text{ mol L}^{-1}\]

\[\text{[OH}^-\text{]}_{\text{initial}} \approx 0 \text{ mol L}^{-1}\]

Note: In an aqueous system, a small amount of OH⁻ is always present from the self-ionization of H₂O.

AN “ICE” TABLE

<table>
<thead>
<tr>
<th></th>
<th>[NH₃]</th>
<th>[NH₄⁺]</th>
<th>[OH⁻]</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>3.88</td>
<td>0</td>
<td>≈ 0</td>
</tr>
<tr>
<td>Change</td>
<td>-X</td>
<td>+X</td>
<td>+X</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>3.88 - X</td>
<td>X</td>
<td>X</td>
</tr>
</tbody>
</table>

\[
\frac{[\text{NH}_4^+]_{\text{eq}} [\text{OH}^-]_{\text{eq}}}{[\text{NH}_3]_{\text{eq}}} = K_b
\]

\[
\frac{(X) (X)}{(3.88 - X)} = 1.8 \times 10^{-5}
\]

Since the initial concentration of NH₃ is fairly large, and the extent of ionization is small (as indicated by the small value of \(K_b\)), the concentration of NH₃ is probably not significantly reduced by the ionization. That is,

\[3.88 - X \approx 3.88\]

This makes it possible to ignore the subtraction of X in the denominator. This greatly simplifies the equation. It can now be written as follows:
\[
\frac{(X)}{(X)} \quad = \quad 1.8 \times 10^{-5} \\
\frac{3.88}{3.88} \\
X^2 \\
\frac{3.88}{3.88} \\
\frac{X^2}{3.88} = 1.8 \times 10^{-5} \\
X = 8.4 \times 10^{-3} = [OH^-]_{eq} \\
\left[ H_3O^+ \right] [OH^-] = K_w = 1.0 \times 10^{-14} \quad \text{Numerical value applies at 25 °C} \\
\left[ H_3O^+ \right] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{8.4 \times 10^{-3}} = 1.2 \times 10^{-12} \\
pH = -\log \left[ H_3O^+ \right] = -\log (1.2 \times 10^{-12}) = 11.9
17. What is the $[\text{H}_3\text{O}^+]$ in a 2.50 M HClO solution? Hypochlorous acid, HClO, is a weak acid with the following equilibrium equation:

$$\text{HClO(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{ClO}^-(aq) \quad K_a = 3.5 \times 10^{-8}$$

The math is no worse than a quadratic, even if you don't approximate, but it should be ok to approximate in this problem.

\[
\begin{array}{ccc}
\text{a) } 3.4 \times 10^{-11} & \text{b) } 5.5 \times 10^{-8} & \text{c) } 5.1 \times 10^{-6} \\
*\text{d) } 3.0 \times 10^{-4} & \text{e) } 2.3 \times 10^{-2} \\
\end{array}
\]

$[\text{HClO}]_{\text{initial}} = 2.50 \text{ mol L}^{-1}$

$[\text{H}_3\text{O}^+]_{\text{initial}} \approx 0 \text{ mol L}^{-1}$ \text{ Note: There is always a small amount of H}_3\text{O}^+ \text{ present in an aqueous system due to the self-ionization of water.}$

$[\text{ClO}^-]_{\text{initial}} = 0 \text{ mol L}^{-1}$

\begin{tabular}{|c|c|c|}
\hline
 & $[\text{HClO}]$ & $[\text{H}_3\text{O}^+]$ & $[\text{ClO}^-]$ \\
\hline
Initial & 2.50 & $\approx 0$ & 0 \\
Change & $- X$ & $+X$ & $+X$ \\
Equilibrium & $2.50 - X$ & $X$ & $X$ \\
\hline
\end{tabular}

\[
\frac{[\text{H}_3\text{O}^+]_{\text{eq}} \cdot [\text{ClO}^-]_{\text{eq}}}{[\text{HClO}]_{\text{eq}}} = K_a
\]

\[
\frac{(X) \cdot (X)}{(2.50 - X)} = 3.5 \times 10^{-8}
\]

Since the initial concentration of HClO is fairly large and the extent of ionization is very small (as indicated by the numerical value of $K_a$), the concentration of HClO is probably not significantly reduced by the ionization. That is,

$$2.50 - X \approx 2.50$$
With this approximation, we can ignore the subtraction of $X$ in the denominator, which greatly simplifies the equation.

\[
\frac{X}{(X)} \quad \frac{X}{2.50} = 3.5 \times 10^{-8}
\]

\[
\frac{X^2}{2.50} = 3.5 \times 10^{-8}
\]

\[
x^2 = 2.50 (3.5 \times 10^{-8}) = 8.8 \times 10^{-8}
\]

\[
x = 3.0 \times 10^{-4} = [H_3O^+]
\]
18. What is the [OH\textsuperscript{-}] in a 0.0100 M HCl solution? Note that HCl is a strong acid. It ionizes completely when dissolved in water:

\[
\text{HCl(aq) } \rightarrow\text{ H}^+(aq) + \text{Cl}^-(aq)
\]

Enough HCl was added to provide 0.0100 moles of HCl for each liter of solution, but once added to water, it exists entirely as ions -- no molecular HCl is found in the solution.

\*a) 1.0 \times 10^{-12}  \quad b) 1.0 \times 10^{-10}  \quad c) 1.0 \times 10^{-8}

d) 1.0 \times 10^{-4}  \quad e) 1.0 \times 10^{-2}

Note that the equation is written to show H\textsuperscript{+} rather than H\textsubscript{2}O\textsuperscript{+}. You will find that H\textsuperscript{+} and H\textsubscript{3}O\textsuperscript{+} are often used interchangeably. If you prefer, you can add H\textsubscript{2}O to both sides of the equation and write it as

\[
\text{HCl(aq) } + \text{H}_2\text{O}(l) \rightarrow\text{H}_3\text{O}^+(aq) + \text{Cl}^-(aq)
\]

HCl is a strong acid. We regard the ionization of a strong acid to be 100\% complete, rather than having the ions in equilibrium with the unionized acid. We can still set up an “ICE” table, but there is no unknown X in it. Rather, numerical values can be obtained for every cell in the table.

[HCl\textsubscript{initial}] = 0.0100 mol L\textsuperscript{-1}

[H\textsubscript{3}O\textsuperscript{+}\textsubscript{initial}] \approx 0 mol L\textsuperscript{-1} \quad \text{Note: There is always a small amount of H\textsubscript{3}O\textsuperscript{+} present in an aqueous system due to the self-ionization of water.}

[Cl\textsuperscript{-}\textsubscript{initial}] = 0 mol L\textsuperscript{-1}

\begin{center}
\begin{tabular}{ | l | l | l |}
\hline
& [HCl] & [H\textsubscript{3}O\textsuperscript{+}] & [Cl\textsuperscript{-}] \\
\hline
Initial & 0.0100 & \approx 0 & 0 \\
Change & -0.0100 & +0.0100 & +0.0100 \\
Equilibrium & 0 & 0.0100 & 0.0100 \\
\hline
\end{tabular}
\end{center}
\[ [\text{H}_3\text{O}^+]_{eq} = 0.0100 \text{ mol L}^{-1} \]

\[ [\text{H}_3\text{O}^+]_{eq} [\text{OH}^-]_{eq} = K_W = 1.0 \times 10^{-14} \text{ Numerical value applies at } 25 \, ^\circ \text{C} \]

\[ [\text{OH}^-]_{eq} = \frac{K_W}{[\text{H}_3\text{O}^+]_{eq}} = \frac{1.0 \times 10^{-14}}{0.0100} = 1.0 \times 10^{-12} \]

19. What is the pOH of a 4.20 M hydrocyanic acid (HCN) solution? Note that hydrocyanic acid is a weak acid that establishes the following equilibrium in water:

\[
\text{HCN(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{CN}^- (aq) \quad K_a = 6.2 \times 10^{-10}
\]

The math is no worse than a quadratic, even if you don't approximate, but it should be ok to approximate in this problem.

- a) 1.6
- b) 4.3
- c) 8.1
- d) 9.7
- e) 12.4

\[ [\text{HCN}]_{initial} = 4.20 \text{ mol L}^{-1} \]

\[ [\text{H}_3\text{O}^+]_{initial} \approx 0 \text{ Note: There is always a small amount of H}_3\text{O}^+ \text{ present in an aqueous system due to the self-ionization of water.} \]

\[ [\text{CN}^-]_{initial} = 0 \]

AN “ICE” TABLE

<table>
<thead>
<tr>
<th></th>
<th>[HCN]</th>
<th>[H_3O^+]</th>
<th>[CN^-]</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>4.20</td>
<td>\approx 0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-X</td>
<td>+X</td>
<td>+X</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>4.20 - X</td>
<td>X</td>
<td>X</td>
</tr>
</tbody>
</table>

\[ \frac{[\text{H}_3\text{O}^+]_{eq} [\text{CN}^-]_{eq}}{[\text{HCN}]_{eq}} = K_a \]

\[ \frac{(X) (X)}{4.20 - X} = 6.2 \times 10^{-10} \]
Since the initial concentration of HCN is fairly large and the extent of ionization is very small (as indicated by the small value of $K_a$), the concentration of HCN is probably not significantly reduced by the ionization. That is,

$$4.20 - X \approx 4.20$$

With this approximation, the subtraction of $X$ in the denominator can be ignored. This greatly simplifies the equation.

$$\frac{(X) (X)}{4.20} = 6.2 \times 10^{-10}$$

$$\frac{X^2}{4.20} = 6.2 \times 10^{-10}$$

$$X^2 = 4.20 \times (6.2 \times 10^{-10}) = 2.6 \times 10^{-9}$$

$$X = 5.1 \times 10^{-5} = [H_3O^+]_{eq}$$

$$[H_3O^+]_{eq} [OH^-]_{eq} = K_w = 1.0 \times 10^{-14} \quad \text{Numerical value applies at 25 °C}$$

$$[OH^-]_{eq} = \frac{K_w}{[H_3O^+]_{eq}} = \frac{1.0 \times 10^{-14}}{5.1 \times 10^{-5}} = 2.0 \times 10^{-10}$$

$$pOH = -\log [OH^-]_{eq} = -\log (2.0 \times 10^{-10}) = 9.7$$
What is the [H$_3$O$^+$] in a 3.00 M sodium acetate (NaHC$_2$H$_3$O$_2$) solution? Note that sodium acetate is a salt that undergoes hydrolysis after being dissolved in water. The salt is freely soluble, and dissolves without establishing an equilibrium:

NaHC$_2$H$_3$O$_2$(aq) $\rightleftharpoons$ Na$^+$ (aq) + C$_2$H$_3$O$_2^-$ (aq)

The Na$^+$, being the conjugate acid of a strong base (NaOH) is too weak an acid to react with water. The C$_2$H$_3$O$_2^-$, being the conjugate base of a weak acid (HC$_2$H$_3$O$_2$) has enough base strength to react with water:

C$_2$H$_3$O$_2^-$ (aq) + H$_2$O(l) $\rightleftharpoons$ H$_3$O$^+$ (aq) + C$_2$H$_3$O$_2$ (aq)

The numerical value of the equilibrium constant for the above reaction is normally not published in tables, because it can be obtained from the equilibrium constants of two other reactions that ARE published:

HC$_2$H$_3$O$_2$(aq) + H$_2$O(l) $\rightleftharpoons$ H$_3$O$^+$ (aq) + C$_2$H$_3$O$_2^-$ (aq) \[ K_a = 1.8 \times 10^{-5} \]

2H$_2$O(l) $\rightleftharpoons$ H$_3$O$^+$ (aq) + OH$^-$ (aq) \[ K_w = 1.0 \times 10^{-14} \]

By combining these two reactions in the appropriate way, you can produce the reaction for which the numerical value of the equilibrium constant is desired. Once you have that, you can do an ICE table calculation in the usual way. It should be possible to make an approximation in your solution to the ICE table calculation to avoid having to solve the quadratic formula.

a) $1.6 \times 10^{-13}$  
*b) $2.4 \times 10^{-10}$  
c) $2.5 \times 10^{-7}$

d) $4.1 \times 10^{-5}$  
e) $1.8 \times 10^{-3}$

If we reverse the acetic acid ionization reaction and add the water ionization reaction, we obtain the desired reaction:

H$_3$O$^+$ (aq) + C$_2$H$_3$O$_2^-$ (aq) $\rightleftharpoons$ HC$_2$H$_3$O$_2$ (aq) + H$_2$O(l) \[ \frac{1}{K_a} = 5.6 \times 10^4 \]

2H$_2$O(l) $\rightleftharpoons$ H$_3$O$^+$ (aq) + OH$^-$ (aq) \[ K_w = 1.0 \times 10^{-14} \]

C$_2$H$_3$O$_2^-$ (aq) + H$_2$O(l) $\rightleftharpoons$ HC$_2$H$_3$O$_2$ (aq) + OH$^-$ (aq) \[ K_b = 5.6 \times 10^{-10} \]
Recall that when we reverse the direction of a chemical reaction, we take the reciprocal of its equilibrium constant, and when we add chemical reactions, we multiply their equilibrium constants. Using these principles, we have determined that the desired reaction has an equilibrium constant of $5.6 \times 10^{-10}$. This equilibrium constant has been labeled $K_b$ since the $C_2H_3O_2^-$ is acting as a base. We know that it acts as a base because it generates $OH^-$ ions in its reaction with water. We refer to $C_2H_3O_2^-$ as the conjugate base of $HC_2H_3O_2$ because $C_2H_3O_2^-$ is what is left over after $HC_2H_3O_2$, acting as an acid, donates an $H^+$ to $H_2O$ to form $H_3O^+$.

Notice that what has been shown above is that

$$\frac{1}{K_a} \cdot K_W = K_b$$

This can be rewritten in the form

$$K_a \cdot K_b = K_W$$

This is a general relationship between the equilibrium constants of any acid and its conjugate base. Since $K_W$ is a known constant, we can calculate $K_b$ if we know $K_a$ and we can calculate $K_a$ if we know $K_b$. Keeping this formula in mind, it will not be necessary to actually set up the chemical arithmetic each time, as I have done above. Once you know the pattern, you can quickly and easily find the needed equilibrium constant.

Now that we have the equilibrium constant for the desired reaction, we can calculate the equilibrium concentrations in the usual manner.

Since the $NaC_2H_3O_2$ is 100% ionized in water, a 3.00 M solution of $NaC_2H_3O_2$ is really a solution that is 3.00 M in both $Na^+$ and $C_2H_3O_2^-$. Therefore, the initial concentration of $C_2H_3O_2^-$ is 3.00 M.
\[ [C_2H_3O_2^-]_{\text{initial}} = 3.00 \text{ mol L}^{-1} \]
\[ [HC_2H_3O_2^-]_{\text{initial}} = 0 \text{ mol L}^{-1} \]
\[ [OH^-]_{\text{initial}} \approx 0 \] Note: In an aqueous solution, there will always be a small amount of OH\(^-\) present from the self-ionization of water.

**AN “ICE” TABLE**

<table>
<thead>
<tr>
<th></th>
<th>([C_2H_3O_2^-])</th>
<th>([HC_2H_3O_2^-])</th>
<th>([OH^-])</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>3.00</td>
<td>0</td>
<td>(\approx 0)</td>
</tr>
<tr>
<td>Change</td>
<td>-X</td>
<td>+X</td>
<td>+X</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>3.00 - X</td>
<td>X</td>
<td>X</td>
</tr>
</tbody>
</table>

\[
\frac{[HC_2H_3O_2^-]_{\text{eq}} [OH^-]_{\text{eq}}}{[C_2H_3O_2^-]_{\text{eq}}} = K_b
\]

\[
\frac{(X) (X)}{3.00 - X} = 5.6 \times 10^{-10}
\]

Since the initial concentration of \(C_2H_3O_2^-\) is fairly large and the extent of reaction is small (as indicated by the small value of \(K_b\)), the concentration of \(C_2H_3O_2^-\) is probably not significantly reduced by the reaction. That is,

\[ 3.00 - X \approx 3.00 \]

With this approximation, we can ignore the subtraction of \(X\) in the denominator. This greatly simplifies the equation.
\[
\frac{(X)(X)}{3.00} = 5.6 \times 10^{-10}
\]

\[
\frac{X^2}{3.00} = 5.6 \times 10^{-10}
\]

\[
X^2 = 3.00 \times (5.6 \times 10^{-10}) = 1.7 \times 10^{-9}
\]

\[
X = 4.1 \times 10^{-5} = [OH^-]_{eq}
\]

\[
[H_3O^+]_{eq} [OH^-]_{eq} = K_W = 1.0 \times 10^{-14} \quad \text{Numerical value applies at } 25 \, ^\circ C
\]

\[
[H_3O^+]_{eq} = \frac{K_W}{[OH^-]_{eq}} = \frac{1.0 \times 10^{-14}}{4.1 \times 10^{-5}} = 2.4 \times 10^{-10}
\]