Chapter 8

Acids, Bases, and pH

Solutions for Practice Problems
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1. Problem
Name and write the formula of the conjugate base of each molecule or ion.
(a) HCl
(b) HCO$_3^-$
(c) H$_2$SO$_4$
(d) N$_2$H$_5^+$

Solution
A conjugate base differs from the molecule or ion by having one less proton.
(a) chloride ion, Cl$^-$
(b) carbonate ion, CO$_3^{2-}$
(c) hydrogen sulfate ion, HSO$_4^-$
(d) hydrazine, N$_2$H$_4$

Check Your Solution
In each case, the conjugate base has one less proton than its acid.

2. Problem
Name and write the formula of the conjugate acid of each molecule or ion.
(a) NO$_3^-$
(b) OH$^-$
(c) H$_2$O
(d) HCO$_3^-$

Solution
A conjugate acid differs from the molecule or ion by having one more proton.
(a) nitric acid, HNO$_3$
(b) water, H$_2$O
(c) hydronium ion, H$_3$O$^+$
(d) carbonic acid, H$_2$CO$_3$

Check Your Solution
In each case, the conjugate acid has one more proton than its base.

3. Problem
Identify the conjugate acid-base pairs in each reaction.
(a) HS$^-$(aq) + H$_2$O(l) $\rightleftharpoons$ H$_2$S(aq) + OH$^-$ (aq)
(b) O$^{2-}$(aq) + H$_2$O(l) $\rightarrow$ 2OH$^-$ (aq)

Solution
On the left side of the equation, the acid is the molecule or ion that donates a proton. The base is the molecule or ion that accepts the proton. On the right side of the equation, you can identify the conjugate acid and base by the difference of a single proton from the base and acid on the left side.
(a) base conjugate acid
\[ \text{HS}^- (aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{S}(aq) + \text{OH}^- (aq) \]
acid conjugate base
The conjugate acid-base pairs are HS⁻/H₂S and H₂O/OH⁻.

(b) base conjugate acid
\[ \text{O}_2\text{⁻} (aq) + \text{H}_2\text{O}(l) \rightleftharpoons 2\text{OH}^- (aq) \]
acid conjugate base
The conjugate acid-base pairs are O₂⁻/OH⁻ and H₂O/OH⁻.

Check Your Solution
The conjugate acid-base pairs differ by one proton. The acid has one more proton than its conjugate base.

4. Problem
Identify the conjugate acid-base pairs in each reaction.
(a) \[ \text{H}_2\text{S}(aq) + \text{NH}_3(aq) \rightleftharpoons \text{NH}_4^+ (aq) + \text{HS}^- (aq) \]
(b) \[ \text{H}_2\text{SO}_4(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{HSO}_4^- (aq) \]

Solution
On the left side of the equation, the acid is the molecule or ion that donates a proton. The base is the molecule or ion that accepts the proton. On the right side of the equation, you can identify the conjugate acid and base by the difference of a single proton from the base and acid on the left side.
(a) base conjugate acid
\[ \text{H}_2\text{S}(aq) + \text{NH}_3(aq) \rightleftharpoons \text{NH}_4^+ (aq) + \text{HS}^- (aq) \]
acid conjugate base
The conjugate acid-base pairs are H₂S/HS⁻ and NH₃/NH₄⁺.

(b) base conjugate acid
\[ \text{H}_2\text{SO}_4(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{HSO}_4^- (aq) \]
acid conjugate base
The conjugate acid-base pairs are H₂SO₄/HSO₄⁻ and H₂O/H₃O⁺.

Check Your Solution
The conjugate acid-base pairs differ by one proton. The acid has one more proton than its conjugate base.

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5. Problem
Calculate the concentration of hydronium ions in each solution.
(a) 4.5 mol/L HCl(aq)
(b) 30.0 mL of 4.50 mol/L HBr(aq) diluted to 100.0 mL
c) 18.6 mL of 2.60 mol/L HClO₄(aq) added to 24.8 mL of 1.92 mol/L NaOH(aq)
d) 17.9 mL of 0.175 mol/L HNO₃(aq) added to 35.4 mL of 0.0160 mol/L Ca(OH)₂(aq)

What Is Required?
You must calculate the [H₃O⁺] in each solution.

(a) What Is Given?
\([\text{HCl}] = 4.5 \text{ mol/L}\)

Solution
HCl is a strong acid. Therefore, \([\text{HCl}] = [\text{H}_3\text{O}^+]\).
\([\text{H}_3\text{O}^+] = 4.5 \text{ mol/L}\)
Check Your Solution
The $[\text{H}_3\text{O}^+]$ is equal to the concentration of a strong acid, provided it is not very dilute.

(b) What Is Given?
$[\text{HBr}] = 4.50 \text{ mol/L}$ and $30.0 \text{ mL}$ is diluted to $100.0 \text{ mL}$.

Solution
Calculate the concentration of HBr after dilution. HBr is a strong acid.
Therefore, $[\text{HBr}] = [\text{H}_3\text{O}^+]$.
Amount of HBr $= 4.50 \text{ mol/L} \times 0.0300 \text{ L} = 0.135 \text{ mol}$
After dilution, the volume is $0.100 \text{ L}$
$[\text{HBr}]$ after dilution $= \frac{0.135 \text{ mol}}{0.100 \text{ L}} = 1.35 \text{ mol/L}$
$[\text{H}_3\text{O}^+] = 1.35 \text{ mol/L}$

Check Your Solution
The acid is diluted a little more than three times, so the concentration of the diluted solution should be a little less than one-third the original concentration. The $[\text{H}_3\text{O}^+]$ is equal to the concentration of a strong acid, provided it is not very dilute.

(c) What Is Given?
Volume of perchloric acid $= 18.6 \text{ mL}$
$[\text{HClO}_4] = 2.60 \text{ mol/L}$
Volume of sodium hydroxide $= 24.8 \text{ mL}$
$[\text{NaOH}] = 1.92 \text{ mol/L}$

Plan Your Strategy
Step 1 Write the chemical equation for the reaction.
Step 2 Calculate the amount of each reactant using the following equation:
Amount (mol) = Concentration (mol/L) $\times$ Volume (L)
Step 3 Determine the limiting reagent.
Step 4 The reagent in excess is a strong acid or base. Thus, the excess amount results in the same amount of $\text{H}_3\text{O}^+$ or $\text{OH}^-$.
Step 5 Calculate the concentration of the excess ion using the amount in excess and the total volume of solution.

Act on Your Strategy
Step 1 $\text{HClO}_4(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{NaClO}_4(\text{aq}) + \text{H}_2\text{O}(\ell)$
Step 2 Amount of $\text{HClO}_4 = 2.60 \text{ mol/L} \times 0.0186 \text{ L} = 0.04836 \text{ mol}$
Amount of $\text{NaOH} = 1.92 \text{ mol/L} \times 0.0248 \text{ L} = 0.04762 \text{ mol}$
Step 3 The reactants combine in a 1:1 ratio. The amount of $\text{NaOH}$ is less, so this reactant must be the limiting reactant.
Step 4 Amount of excess
$\text{HClO}_4 = 0.04836 \text{ mol} - 0.04762 \text{ mol}$ (carrying one more digit)
$= 7.4 \times 10^{-4} \text{ mol}$ (again, carrying one more digit)
Therefore, the amount of $\text{H}_3\text{O}^+ (\text{aq}) = 7.4 \times 10^{-4} \text{ mol}$
Step 5 Total volume of solution $= 18.6 \text{ mL} + 24.8 \text{ mL} = 43.4 \text{ mL}$
$[\text{H}_3\text{O}^+] = \frac{7.4 \times 10^{-4}}{0.0434 \text{ L}}$
$= 0.017 \text{ mol/L}$
Thus, $[\text{H}_3\text{O}^+] = 0.02 \text{ mol/L}$.
Check Your Solution
The chemical equation has a 1:1 ratio between reactants. The amount of acid is greater than the amount of base, so the resulting solution should be acidic. Subtracting the amount of base from the amount of acid gives an answer with one decimal place. The final answer should have one significant digit.

(d)
What Is Given?
Volume of nitric acid = 17.9 mL
[HNO₃] = 0.175 mol/L
Volume of calcium hydroxide = 35.4 mL
[Ca(OH)₂] = 0.0160 mol/L

Plan Your Strategy
Step 1  Write the chemical equation for the reaction.
Step 2  Calculate the amount of each reactant using the following equation.
\[ \text{Amount (mol)} = \text{Concentration (mol/L)} \times \text{Volume (L)} \]
Step 3  Determine the limiting reagent.
Step 4  The reagent in excess is a strong acid or base. Thus, the excess amount results in the same amount of H₃O⁺ or OH⁻.
Step 5  Calculate the concentration of the excess ion using the amount in excess and the total volume of solution.

Act on Your Strategy
Step 1  \( 2\text{HNO}_3(\text{aq}) + \text{Ca(OH)}_2(\text{aq}) \rightarrow \text{Ca(NO}_3)_2(\text{aq}) + 2\text{H}_2\text{O(ℓ)} \)
Step 2  Amount of \( \text{HNO}_3 = 0.175 \text{ mol/L} \times 0.0179 \text{ L} = 0.003132 \text{ mol} \)
Amount of \( \text{Ca(OH)}_2 = 0.0160 \text{ mol/L} \times 0.0354 \text{ L} = 0.000566 \text{ mol} \)
Step 3  The reactants combine in a 2:1 ratio. The amount of \( \text{HNO}_3 \) that will react with 0.000566 mol of \( \text{Ca(OH)}_2 \) is:
\[ \text{Amount } \text{HNO}_3 = 2 \times 0.000566 \text{ mol} = 0.00113 \text{ mol} \]
The amount of acid is greater, therefore \( \text{Ca(OH)}_2 \) must be the limiting reactant.
Step 4  Amount of excess
\[ \text{HNO}_3 = 0.003132 - 0.00113 \text{ mol (carrying one more digit)} \]
\[ = 0.00200 \text{ mol} \]
Therefore, the amount of \( H_3O^+ \) is 0.00200 mol.
Step 5  Total volume of solution = 17.9 mL + 35.4 mL = 53.3 mL
\[ [H_3O^+] = \frac{0.00200 \text{ mol}}{0.0533 \text{ L}} \]
\[ = 0.0375 \text{ mol/L} \]

Check Your Solution
The chemical equation has a 2:1 ratio between reactants. The amount of acid is greater than the amount of base that reacts, so the resulting solution should be acidic.

6. Problem
Calculate the concentration of hydroxide ions in each solution.
(a) 3.1 mol/L KOH(\text{aq})
(b) 21.0 mL of 3.1 mol/L KOH diluted to 75.0 mL
(c) 23.2 mL of 1.58 mol/L HCl(\text{aq}) added to 18.9 mL of 3.50 mol/L NaOH
(d) 16.5 mL of 1.50 mol/L H₂SO₄(\text{aq}) added to 12.7 mL of 5.50 mol/L NaOH

What Is Required?
You must calculate [OH⁻] in each solution.
(a) What Is Given?

\[ [\text{KOH}] = 3.1 \text{ mol/L} \]

**Solution**

KOH is a strong base. Therefore, \([\text{KOH}] = [\text{OH}^-] \).

\[ [\text{OH}^-] = 3.1 \text{ mol/L} \]

**Check Your Solution**

The \([\text{OH}^-]\) is equal to the concentration of a strong base, provided it is not very dilute.

(b) What Is Given?

\[ [\text{KOH}] = 3.1 \text{ mol/L} \]

21.0 mL is diluted to 75.0 mL

**Solution**

Calculate the concentration of KOH after dilution. KOH is a strong base. Therefore, \([\text{KOH}] = [\text{OH}^-] \).

Amount of KOH = \(3.1 \text{ mol/L} \times 0.0210 \text{ L} = 0.0651 \text{ mol} \)

After dilution, the volume is 0.0750 L.

\[ [\text{KOH}] \text{ after dilution} = \frac{0.00651 \text{ mol}}{0.075 \text{ L}} = 0.868 \text{ mol/L} \]

\[ [\text{OH}^-] = 0.87 \text{ mol/L} \]

**Check Your Solution**

The \([\text{OH}^-]\) is equal to the concentration of a strong base, provided it is not very dilute. The base is diluted to almost a quarter the original concentration. Therefore, the \([\text{OH}^-]\) should be just greater than one-quarter the original concentration of the base.

(c) What Is Given?

Volume of hydrochloric acid = 23.2 mL

\[ [\text{HCl}] = 1.58 \text{ mol/L} \]

Volume of sodium hydroxide = 18.9 mL

\[ [\text{NaOH}] = 3.50 \text{ mol/L} \]

**Plan Your Strategy**

**Step 1** Write the chemical equation for the reaction.

**Step 2** Calculate the amount of each reactant using the following equation:

\[ \text{Amount (mol)} = \text{Concentration (mol/L)} \times \text{Volume (L)} \]

**Step 3** Determine the limiting reagent.

**Step 4** The reagent in excess is a strong acid or base. Thus, the excess amount results in the same amount of \(\text{H}_3\text{O}^+\) or \(\text{OH}^-\).

**Step 5** Calculate the concentration of the excess ion using the amount in excess and the total volume of solution.

**Act on Your Strategy**

**Step 1** \(\text{HCl}_{(aq)} + \text{NaOH}_{(aq)} \rightarrow \text{NaCl}_{(aq)} + \text{H}_2\text{O}(l)\)

**Step 2** Amount of HCl = \(1.58 \text{ mol/L} \times 0.0232 \text{ L} = 0.03666 \text{ mol} \)

Amount of NaOH = \(3.50 \text{ mol/L} \times 0.0189 \text{ L} = 0.06615 \text{ mol} \)

**Step 3** The reactants combine in a 1:1 ratio. The amount of HCl is less, so this reactant must be the limiting reactant.
**Step 4**  
Amount of excess NaOH = 0.06615 mol − 0.03666 mol  
= 0.02949 mol (carrying one more digit)  
Therefore, the amount of OH\(^{-}\)\(_{(aq)}\) is 0.02949 mol.

**Step 5**  
Total volume of solution = 23.2 mL + 18.9 mL = 42.1 mL  
\[ [\text{OH}^-] = \frac{0.0249 \text{ mol}}{0.0421 \text{ L}} = 0.700 \text{ mol/L} \]

**Check Your Solution**  
The chemical equation has a 1:1 ratio between reactants. The amount of base is greater than the amount of acid, so the resulting solution should be basic.

**What Is Given?**
Volume of sulfuric acid = 16.5 mL  
\([\text{H}_2\text{SO}_4]\) = 1.50 mol/L  
Volume of sodium hydroxide = 12.7 mL  
\([\text{NaOH}]\) = 5.50 mol/L

**Plan Your Strategy**

**Step 1**  
Write the chemical equation for the reaction.

**Step 2**  
Calculate the amount of each reactant using the following equation:  
Amount (mol) = Concentration (mol/L) × Volume (L)

**Step 3**  
Determine the limiting reagent.

**Step 4**  
The reagent in excess is a strong acid or base. Thus, the excess amount results in the same amount of \(\text{H}_3\text{O}^+\) or \(\text{OH}^-\).

**Step 5**  
Calculate the concentration of the excess ion using the amount in excess and the total volume of solution.

**Act on Your Strategy**

**Step 1**  
\(\text{H}_2\text{SO}_4(aq) + 2\text{NaOH(aq)} \rightarrow \text{Na}_2\text{SO}_4(aq) + 2\text{H}_2\text{O(ℓ)}\)

**Step 2**  
Amount of \(\text{H}_2\text{SO}_4\) = 1.50 mol/L × 0.0165 L = 0.02475 mol  
Amount of \(\text{NaOH}\) = 5.50 mol/L × 0.0127 L = 0.06985 mol

**Step 3**  
The reactants combine in a 1:2 ratio. The amount of \(\text{NaOH}\) that will react with 0.02475 mol of \(\text{H}_2\text{SO}_4\) is:  
Amount of \(\text{NaOH}\) = 2 × 0.02475 = 0.0495 mol  
The amount of base is greater, therefore \(\text{H}_2\text{SO}_4\) must be the limiting reactant.

**Step 4**  
Amount of excess \(\text{NaOH}\) = 0.06985 − 0.0495 mol = 0.0204 mol  
Therefore, the amount of \(\text{OH}^-\)\(_{(aq)}\) is 0.0204 mol.

**Step 5**  
Total volume of solution = 16.5 mL + 12.7 mL = 29.2 mL  
\[ [\text{OH}^-] = \frac{0.204 \text{ mol}}{0.0292 \text{ L}} = 0.699 \text{ mol/L} \]

**Check Your Solution**  
The chemical equation has a 1:2 ratio between reactants. The amount of base is greater than the amount of acid that reacts, so the resulting solution should be basic.

### 7. Problem
Determine whether reacting each pair of solutions results in an acidic or basic solution. Then calculate the concentration of the ion that causes the solution to be acidic or basic. (Assume that the volumes in part (a) are additive. Assume that the volumes in part (b) stay the same.)

(a) 31.9 mL of 2.75 mol/L \(\text{HCl(aq)}\) added to 125 mL of 0.0500 mol/L \(\text{Mg(OH)}_2(aq)\)

(b) 4.87 g of \(\text{NaOH(aq)}\) added to 80.0 mL of 3.50 mol/L \(\text{HBr(aq)}\)
What Is Required?
You must determine the ion in excess and its concentration.

(a)

What Is Given?
Volume of hydrochloric acid = 31.9 mL
\([\text{HCl}] = 2.75 \text{ mol/L}\)
Volume of magnesium hydroxide = 125 mL
\([\text{Mg(OH)}_2] = 0.0500 \text{ mol/L}\)

Plan Your Strategy

Step 1  Write the chemical equation for the reaction.
Step 2  Calculate the amount of each reactant using the following equation:
\[
\text{Amount (mol)} = \text{Concentration (mol/L)} \times \text{Volume (L)}
\]
Step 3  Determine the limiting reactant.
Step 4  The reactant in excess is a strong acid or base. Thus, the excess amount results in the same amount of \(\text{H}_3\text{O}^+\) or \(\text{OH}^-\).
Step 5  Calculate the concentration of the excess ion using the amount in excess and the total volume of solution.

Act on Your Strategy

Step 1  \(2\text{HCl(aq)} + \text{Mg(OH)}_2\text{(aq)} \rightarrow \text{MgCl}_2\text{(aq)} + 2\text{H}_2\text{O(}\ell\text{)}\)
Step 2  Amount of HCl = \(2.75 \text{ mol/L} \times 0.0319 \text{ L} = 0.08772 \text{ mol}\)
Amount of Mg(OH)_2 = \(0.0500 \text{ mol/L} \times 0.125 \text{ L} = 0.00625 \text{ mol}\)
Step 3  The reactants combine in a 2:1 ratio. The amount of HCl that will react with 0.00625 mol of Mg(OH)_2 is:
Amount of HCl = \(2 \times 0.00625 = 0.0125 \text{ mol}\)
The amount of acid is greater, therefore Mg(OH)_2 must be the limiting reactant.
Step 4  Amount of excess HCl = \(0.08772 - 0.0125 \text{ mol} = 0.0752 \text{ mol}\)
Therefore, the amount of \(\text{H}_3\text{O}^+\text{(aq)}\) = 0.0752 mol
Step 5  Total volume of solution = 31.9 mL + 125 mL = 157 mL
\[
[\text{H}_3\text{O}^+] = \frac{0.0752 \text{ mol}}{0.157 \text{ L}} = 0.479 \text{ mol/L}
\]

Check Your Solution

The chemical equation has a 2:1 ratio between reactants. The amount of acid is greater than the amount of base that reacts, so the resulting solution should be acidic.

(b)

What Is Given?
Mass of sodium hydroxide = 4.87 g
Volume of hydrobromic acid = 80.0 mL
\([\text{HBr}] = 3.50 \text{ mol/L}\)

Plan Your Strategy

Step 1  Write the chemical equation for the reaction.
Step 2  Calculate the amount of each reactant using the following equations:
\[
\text{Amount (mol)} = \frac{\text{Mass (g)}}{\text{Molar mass (g/mol)}}
\]
\[
\text{Amount (mol)} = \text{Concentration (mol/L)} \times \text{Volume (L)}
\]
Step 3  Determine the limiting reactant.
Step 4  The reactant in excess is a strong acid or base. Thus, the excess amount results in the same amount of \(\text{H}_3\text{O}^+\) or \(\text{OH}^-\).
Step 5  Calculate the concentration of the excess ion using the amount in excess and the total volume of solution.
Act on Your Strategy

**Step 1**  \( \text{HBr(aq)} + \text{NaOH(aq)} \rightarrow \text{NaBr(aq)} + \text{H}_2\text{O(ℓ)} \)

**Step 2**  
Amount of \( \text{HBr} \) = \( 3.50 \text{ mol/L} \times 0.0800 \text{ L} = 0.280 \text{ mol} \)  
Molar mass of \( \text{NaOH} \) = \( 40.0 \text{ g/mol} \)  
Amount of \( \text{NaOH} \) = \( \frac{4.87 \text{ g}}{40.0 \text{ g/mol}} = 0.1218 \text{ mol} \)

**Step 3**  
The reactants combine in a 1:1 ratio. The amount of acid is greater, therefore \( \text{NaOH} \) must be the limiting reactant.

**Step 4**  
Amount of excess \( \text{HBr} \) = \( 0.280 - 0.122 \text{ mol} = 0.158 \text{ mol} \)  
Therefore, the amount of \( \text{H}_3\text{O}^+(aq) \) = 0.158 mol

**Step 5**  
\[ [\text{H}_3\text{O}^+] = \frac{0.158 \text{ mol}}{0.0800 \text{ L}} = 1.98 \text{ mol/L} \]

Check Your Solution

The chemical equation has a 1:1 ratio between reactants. The amount of acid is greater than the amount of base that reacts, so the resulting solution should be acidic.

8. Problem

2.75 g of \( \text{MgO(s)} \) is added to 70.0 mL of 2.40 mol/L \( \text{HNO}_3(aq) \). Is the solution that results from the reaction acidic or basic? What is the concentration of the ion that is responsible for the character of the solution?

**What Is Given?**

- Mass of magnesium oxide = 2.75 g
- Volume of nitric acid = 70.0 mL
- \([\text{HNO}_3]\) = 2.40 mol/L

**Plan Your Strategy**

**Step 1**  
Write the chemical equation for the reaction.

**Step 2**  
Calculate the amount of each reactant using the following equations:

\[
\text{Amount (mol)} = \frac{\text{Mass (g)}}{\text{Molar mass (g/mol)}}
\]

\[
\text{Amount (mol)} = \text{Concentration (mol/L)} \times \text{Volume (L)}
\]

**Step 3**  
Determine the limiting reactant.

**Step 4**  
The reactant in excess is a strong acid or base. Thus, the excess amount results in the same amount of \( \text{H}_3\text{O}^+ \) or \( \text{OH}^- \).

**Step 5**  
Calculate the concentration of the excess ion using the amount in excess and the total volume of solution.

Act on Your Strategy

**Step 1**  
\( 2\text{HNO}_3(aq) + \text{MgO(s)} \rightarrow \text{Mg(NO}_3)_2(aq) + \text{H}_2\text{O(ℓ)} \)

**Step 2**  
Amount of \( \text{HNO}_3 \) = \( 2.40 \text{ mol/L} \times 0.0700 \text{ L} = 0.168 \text{ mol} \)  
Molar mass of \( \text{MgO} \) = \( 40.3 \text{ g/mol} \)  
Amount of \( \text{MgO} \) = \( \frac{2.75 \text{ g}}{40.3 \text{ g/mol}} = 0.06824 \text{ mol} \)

**Step 3**  
The reactants combine in a 2:1 ratio. The amount of acid that will combine with 0.06824 mol \( \text{MgO} \) is:

\[
\text{Amount of } \text{HNO}_3 = 2 \times 0.06824 \text{ mol} = 0.1365 \text{ mol}
\]

This is less than the amount of acid. Therefore, \( \text{MgO} \) is the limiting reactant.

**Step 4**  
Amount of excess \( \text{HNO}_3 \) = \( 0.168 - 0.136 = 0.032 \text{ mol} \)  
(Note: the subtraction gives 3 decimal places, but only 2 significant digits.)  
Therefore, the amount of \( \text{H}_3\text{O}^+(aq) \) is 0.032 mol.
Step 5 \[ [H_3O^+] = \frac{0.032 \text{ mol}}{0.0700 \text{ L}} \]
\[ = 0.46 \text{ mol/L} \]

**Check Your Solution**
The chemical equation has a 2:1 ratio between reactants. The amount of acid is greater than the amount of base that reacts, so the resulting solution should be acidic.

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9. **Problem**
Determine \([H_3O^+]\) and \([OH^-]\) in each solution.
(a) 0.45 mol/L hydrochloric acid
(b) 1.1 mol/L sodium hydroxide

**Solution**
You know that hydrochloric acid is a strong acid and sodium hydroxide is a strong base. Since both dissociate completely in aqueous solutions, you can use their molar concentrations to determine \([H_3O^+]\) or \([OH^-]\). You can find the concentration of the other ion using \(K_w\), as shown below:
\[ K_w = 1.0 \times 10^{-14} = [H_3O^+][OH^-] \]

(a) \([HCl] = 0.45 \text{ mol/L}, \) so \([H_3O^+] = 0.45 \text{ mol/L} \)
\[ [OH^-] = \frac{1.0 \times 10^{-14} \text{ mol/L}}{0.45} \]
\[ = 2.2 \times 10^{-14} \text{ mol/L} \]

(b) \([NaOH] = 1.1 \text{ mol/L} \)
Therefore, \([OH^-] = 1.1 \text{ mol/L} \)
\[ [H_3O^+] = \frac{1.0 \times 10^{-14} \text{ mol/L}}{1.1} \]
\[ = 9.1 \times 10^{-15} \text{ mol/L} \]

**Check Your Solution**
Solution (a) is a strong acid. Therefore, \([H_3O^+]\) should be greater than \(1.0 \times 10^{-7}\), and \([OH^-]\) should be less than \(1.0 \times 10^{-7}\). For a solution of a strong base, as in (b), \([OH^-]\) should be greater than \(1.0 \times 10^{-7}\), and \([H_3O^+]\) should be less than \(1.0 \times 10^{-7}\).

10. **Problem**
Determine \([H_3O^+]\) and \([OH^-]\) in each solution.
(a) 0.95 mol/L hydrobromic acid
(b) 0.012 mol/L calcium hydroxide

**Solution**
You know that hydrobromic acid is a strong acid and calcium hydroxide is a strong base. Since both dissociate completely in aqueous solutions, you can use their molar concentrations to determine \([H_3O^+]\) or \([OH^-]\). You can find the concentration of the other ion using \(K_w\), as shown below:
\[ K_w = 1.0 \times 10^{-14} = [H_3O^+][OH^-] \]

(a) \([HBr] = 0.95 \text{ mol/L}, \) so \([H_3O^+] = 0.95 \text{ mol/L} \)
\[ [OH^-] = \frac{1.0 \times 10^{-14} \text{ mol/L}}{0.95} \]
\[ = 1.1 \times 10^{-15} \text{ mol/L} \]

Check Your Solution
Solution (a) is a strong acid. Therefore, \([H_3O^+]\) should be greater than \(1.0 \times 10^{-7}\), and \([OH^-]\) should be less than \(1.0 \times 10^{-7}\). For a solution of a strong base, as in (b), \([OH^-]\) should be greater than \(1.0 \times 10^{-7}\), and \([H_3O^+]\) should be less than \(1.0 \times 10^{-7}\).
(b) \([\text{Ca(OH)}_2] = 0.12 \text{ mol/L}\)

Each mole of \(\text{Ca(OH)}_2\) in solution forms two moles of \(\text{OH}^-\) ions.
Therefore, \([\text{OH}^-] = 2 \times 0.012 \text{ mol/L} = 0.024 \text{ mol/L}\)

\[
[H_3O^+] = \frac{1.0 \times 10^{-14} \text{ mol/L}}{0.024} \\
= 4.2 \times 10^{-13} \text{ mol/L}
\]

**Check Your Solution**

Solution (a) is a strong acid. Therefore, \([H_3O^+]\) should be greater than \(1.0 \times 10^{-7}\),
and \([\text{OH}^-]\) should be less than \(1.0 \times 10^{-7}\). For a solution of a strong base, as in (b),
[\(\text{OH}^-\)] should be greater than \(1.0 \times 10^{-7}\), and \([H_3O^+]\) should be less than \(1.0 \times 10^{-7}\).

11. **Problem**

[\(\text{OH}^-\)] is \(5.6 \times 10^{-14}\) mol/L in a solution of hydrochloric acid. What is the molar concentration of the \(\text{HCl(aq)}\)?

**Solution**

You know [\(\text{OH}^-\)] = \(5.6 \times 10^{-14}\) mol/L. You can find \([H_3O^+]\) using \(K_w\) as shown below:

\[K_w = 1.0 \times 10^{-14} = [H_3O^+][\text{OH}^-]\]

Hydrochloric acid is a strong acid. Therefore, \([H_3O^+] = [\text{HCl}]\)

\[
[H_3O^+] = \frac{1.0 \times 10^{-14} \text{ mol/L}}{5.6 \times 10^{-14}} \\
= 0.18 \text{ mol/L}
\]

Therefore, \([\text{HCl}] = 0.18 \text{ mol/L}\)

**Check Your Solution**

For a solution with [\(\text{OH}^-\)] less than \(1.0 \times 10^{-7}\), the \([H_3O^+]\) must be greater than \(1.0 \times 10^{-7}\).

12. **Problem**

[\(H_3O^+\)] is \(1.7 \times 10^{-14}\) in a solution of calcium hydroxide. What is the molar concentration of the \(\text{Ca(OH)}_2(aq)\)?

**Solution**

You know \([H_3O^+] = 1.7 \times 10^{-14}\) mol/L. You can find \([\text{OH}^-]\) using \(K_w\), as shown below:

\[K_w = 1.0 \times 10^{-14} = [H_3O^+][\text{OH}^-]\]

Calcium hydroxide is a strong base. Each mole of \(\text{Ca(OH)}_2\) forms two moles of \(\text{OH}^-\) ions. Therefore, \([\text{Ca(OH)}_2] = \frac{1}{2} \times [\text{OH}^-]\)

\[
[\text{OH}^-] = \frac{1.0 \times 10^{-14} \text{ mol/L}}{1.7 \times 10^{-14}} \\
= 0.588 \text{ mol/L}
\]

Therefore, \([\text{Ca(OH)}_2] = \frac{1}{2} \times 0.588 \text{ mol/L} = 0.29 \text{ mol/L}\)

**Check Your Solution**

For a solution with \([H_3O^+]\) less than \(1.0 \times 10^{-7}\), the \([\text{OH}^-]\) must be greater than \(1.0 \times 10^{-7}\).
Solutions for Practice Problems
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13. Problem

[\text{H}_3\text{O}^+] \text{ of a sample of milk is found to be } 3.98 \times 10^{-7} \text{ mol/L. Is the milk acidic, neutral, or basic? Calculate the pH and [OH\text{\textsuperscript{-}}] of the sample.}

Solution

Compare [\text{H}_3\text{O}^+] \text{ in the milk with } [\text{H}_3\text{O}^+] \text{ in neutral water.}

[\text{H}_3\text{O}^+] = 3.98 \times 10^{-7} \text{ mol/L, which is greater than } 1.0 \times 10^{-7} \text{ mol/L. Therefore, the milk is acidic.}

Use the equation \( \text{pH} = -\log[\text{H}_3\text{O}^+] \) to find the pH.

\[
\text{pH} = -\log 3.98 \times 10^{-7} = 6.400
\]

Use the following equation to find the hydroxide ion concentration:

\[
[\text{OH}^-] = \frac{1.0 \times 10^{-14} \text{ mol/L}}{3.98 \times 10^{-7}} = 2.51 \times 10^{-8} \text{ mol/L}
\]

Check Your Solution

The ion product is

\[
[\text{H}_3\text{O}^+] \times [\text{OH}^-] = 3.98 \times 10^{-7} \times 2.51 \times 10^{-8} = 9.99 \times 10^{-15}
\]

This is equal to the value of \( K_w, 1.0 \times 10^{-14}, \text{ within the error introduced by mathematical rounding.} \)

14. Problem

A sample of household ammonia has a pH of 11.9. What is the pOH and [OH\text{\textsuperscript{-}}] of the sample?

Solution

Use the following equation to find the pOH:

\[
\text{pH} + \text{pOH} = 14.0
\]

Then calculate [OH\text{\textsuperscript{-}}] using the following equation:

\[
[\text{OH}^-] = 10^{-\text{pOH}}
\]

pOH = 14.0 − 11.9

\[
= 2.1
\]

[OH\text{\textsuperscript{-}}] = 10^{-2.1}

\[
= 8 \times 10^{-3} \text{ mol/L}
\]

Check Your Solution

The ammonia solution has pH greater than 7, and is therefore basic. The [OH\text{\textsuperscript{-}}] is greater than 1 \times 10^{-7} \text{ mol/L.}

15. Problem

Phenol, \( \text{C}_6\text{H}_5\text{OH} \), is used as a disinfectant. An aqueous solution of phenol was found to have a pH of 4.72. Is phenol acidic, neutral, or basic? Calculate [\text{H}_3\text{O}^+], [\text{OH}^-], and pOH of the solution.

Solution

Compare the pH of the solution with the pH of neutral water. pH = 4.72, which is less than 7. Therefore, the solution is acidic.

Use the following equation to find the [\text{H}_3\text{O}^+]:
\[ [\text{H}_3\text{O}^+] = 10^{-\text{pH}} \]
\[ [\text{H}_3\text{O}^+] = 10^{-4.72} \]
\[ = 1.9 \times 10^{-5} \text{ mol/L} \]

Use the following equation to find the hydroxide ion concentration:
\[ [\text{OH}^-] = \frac{1.0 \times 10^{-14} \text{ mol/L}}{[\text{H}_3\text{O}^+]} \]
\[ [\text{OH}^-] = \frac{1.0 \times 10^{-14} \text{ mol/L}}{1.9 \times 10^{-5}} \]
\[ = 5.3 \times 10^{-10} \text{ mol/L} \]

Calculate pOH using the following equation:
\[ \text{pOH} = -\log [\text{OH}^-] \]
\[ \text{pOH} = -\log 5.3 \times 10^{-10} \]
\[ = 9.28 \]

Check Your Solution
\[ \text{pH} + \text{pOH} = 4.72 + 9.28 = 14.00 \]

16. Problem
At normal body temperature, 37˚C, the value of \( K_w \) for water is \( 2.5 \times 10^{-14} \).
Calculate \([\text{H}_3\text{O}^+]\) and \([\text{OH}^-]\) at this temperature. Is pure water at 37˚C acidic, neutral, or basic?

Solution
\[ \text{H}_2\text{O}(l) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{OH}^-(aq) \]
Therefore, \([\text{H}_3\text{O}^+] = [\text{OH}^-]\)
Let \( x \) represent \([\text{H}_3\text{O}^+]\).
Find the hydronium ion concentration using:
\[ K_w = [\text{H}_3\text{O}^+] [\text{OH}^-] \]
\[ 2.5 \times 10^{-14} = x^2 \]
\[ x = \pm 1.6 \times 10^{-7} \]
The negative root is not physically possible.
Therefore, \([\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.6 \times 10^{-7} \text{ mol/L} \)

Check Your Solution
\( K_w \) is an equilibrium constant. Every equilibrium constant depends on temperature.
Since the water is pure, it must be neutral.

17. Problem
A sample of baking soda was dissolved in water and the pOH of the solution was found to be 5.81 at 25˚C. Is the solution acidic, basic or neutral? Calculate the pH, \([\text{H}_3\text{O}^+]\), and \([\text{OH}^-]\) of the solution.

Solution
Compare the pOH of the solution with the pOH of neutral water at 25˚C. The pOH is 5.81, which is less than 7. Therefore, the solution is basic.

Use the following equation to find the pH:
\[ \text{pH} + \text{pOH} = 14.00 \]
\[ \text{pH} = 14.00 - 5.81 \]
\[ = 8.19 \]
\[ [\text{H}_3\text{O}^+] = 10^{-\text{pH}} \]
\[ = 10^{-8.19} \]
\[ = 6.5 \times 10^{-9} \text{ mol/L} \]
Use the following equation to find the hydroxide ion concentration:

\[ [\text{OH}^-] = \frac{1.0 \times 10^{-14} \text{ mol/L}}{[\text{H}_3\text{O}^+]} \]

\[ [\text{OH}^-] = \frac{1.0 \times 10^{-14} \text{ mol/L}}{6.5 \times 10^{-9}} \]

\[ = 1.5 \times 10^{-6} \text{ mol/L} \]

**Check Your Solution**

\[[\text{H}_3\text{O}^+][\text{OH}^-] = (6.5 \times 10^{-9}) \times (1.5 \times 10^{-6})

\[ = 9.8 \times 10^{-15} \]

This is equal to the value of \( K_w \), \( 1.0 \times 10^{-14} \), within the error introduced by mathematical rounding.

18. Problem

A chemist dissolved some Aspirin\textsuperscript{TM} in water. The chemist then measured the pH of the solution and found it to be 2.73 at 25°C. What are the [H\textsubscript{3}O\textsuperscript{+}] and [OH\textsuperscript{-}] of the solution?

**Solution**

Find the hydronium ion concentration using the following equation:

\[ [\text{H}_3\text{O}^+] = 10^{-\text{pH}} \]

\[ [\text{H}_3\text{O}^+] = 10^{-2.73} = 1.9 \times 10^{-3} \text{ mol/L} \]

Use the following equation to find the hydroxide ion concentration:

\[ [\text{OH}^-] = \frac{1.0 \times 10^{-14} \text{ mol/L}}{[\text{H}_3\text{O}^+]} \]

\[ [\text{OH}^-] = \frac{1.0 \times 10^{-14} \text{ mol/L}}{1.9 \times 10^{-3}} \]

\[ = 5.3 \times 10^{-12} \text{ mol/L} \]

**Check Your Solution**

\[[\text{H}_3\text{O}^+][\text{OH}^-] = (1.9 \times 10^{-3}) \times (5.3 \times 10^{-12})

\[ = 1.0 \times 10^{-14} \]

This is equal to the value of \( K_w \) at 25°C.

**Solutions for Practice Problems**

**Student Textbook page 400**

19. Problem

Calculate the pH of a sample of vinegar that contains 0.83 mol/L acetic acid. What is the percent dissociation of the vinegar?

**What Is Required?**

You need to calculate the pH of the solution and the percent dissociation for acetic acid.

**What Is Given?**

Initial [CH\textsubscript{3}COOH] = 0.83 mol/L

Acid dissociation constant for acetic acid (listed in Table 8.2):

\[ K_a = 1.8 \times 10^{-5} \]

**Plan Your Strategy**

**Step 1** Check the value of \( \frac{[\text{CH}_3\text{COOH}]}{K_a} \) to see whether or not the amount that dissociates is negligible compared with the initial concentration of the acid.
Step 2 Write the equation for the dissociation equilibrium of acetic acid in water. Then set up an ICE table.

Step 3 Write the equation for the acid dissociation constant. Substitute equilibrium concentrations into the equilibrium expression. Solve the equilibrium expression for \( x \). If the amount that dissociates is not negligible compared with the initial concentration of acid, you will need to solve a quadratic equation.

Step 4 \( \text{pH} = -\log[H_3O^+] \)

Step 5 Calculate the percent dissociation by expressing the fraction of molecules that dissociate out of 100.

Act on Your Strategy

Step 1

\[
\frac{[\text{CH}_3\text{COOH}]}{K_a} = \frac{0.83}{1.8 \times 10^{-5}} = 4.6 \times 10^4
\]

Since this value is greater than 500, the amount that dissociates can be neglected compared to the initial concentration of the acid.

Step 2

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>CH(<em>3)COOH(</em>{(aq)})</th>
<th>H(<em>2)O(</em>{(l)})</th>
<th>CH(<em>3)COO(^-)(</em>{(aq)})</th>
<th>H(<em>3)O(^+)(</em>{(aq)})</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.83</td>
<td>0</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>(-x)</td>
<td>(+x)</td>
<td>(+x)</td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.83 (-x) 0.83</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Step 3

\[
K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]} = \frac{(x)(x)}{0.83} = 1.8 \times 10^{-5}
\]

\[
x = \sqrt{1.49 \times 10^{-5}} = \pm 3.9 \times 10^{-3}
\]

Only the positive root is chemically important. Therefore, \([H_3O^+] = 3.9 \times 10^{-3} \text{ mol/L} \)

Step 4

\[
\text{pH} = -\log 3.9 \times 10^{-3} = 2.41
\]

Step 5

Percent dissociation = \[
\frac{3.9 \times 10^{-3} \text{ mol/L}}{0.83 \text{ mol/L}} \times 100 = 0.47%
\]

Check Your Solution

The pH indicates an acidic solution, as expected. Data given in the problem has two significant digits, and the pH has two digits following the decimal place. The assumption that \( x \) is negligible compared with the initial concentration is valid because, using the rules for subtracting measured quantities, \( 0.83 - 3.9 \times 10^{-3} = 0.83 \). Next, check the equilibrium values:

\[
\frac{(3.9 \times 10^{-3})^2}{0.83} = 1.8 \times 10^{-5}
\]

This is equal to \( K_a \).

20. Problem

In low doses, barbiturates act as sedatives. Barbiturates are made from barbituric acid, a weak monoprotic acid that was first prepared by the German Chemist Adolph von Baeyer in 1864. The formula of barbituric acid is \( \text{C}_4\text{H}_3\text{N}_2\text{O}_3 \). A chemist prepares a 0.10 mol/L solution of barbituric acid. The chemist finds the pH of the solution to be 2.50. What is the acid dissociation constant for barbituric acid? What percent of its molecules dissociate?
What Is Required?
You need to find $K_a$ and the percent dissociation for barbituric acid.

What Is Given?
Initial $[\text{C}_4\text{H}_4\text{N}_2\text{O}_3] = 0.10 \text{ mol/L}$
$pH = 2.50$

Plan Your Strategy
Step 1 Write the equation for the dissociation equilibrium of barbituric acid in water. Then set up an ICE table.
Step 2 Write the equation for the acid dissociation constant. Substitute equilibrium terms into the equation.
Step 3 Calculate $[\text{H}_3\text{O}^+]$ using $[\text{H}_3\text{O}^+] = 10^{-pH}$
Step 4 Use the stoichiometry of the equation and $[\text{H}_3\text{O}^+]$ to substitute for the unknown term, $x$, and calculate $K_a$.
Step 5 Calculate the percent dissociation by expressing the fraction of molecules that dissociate out of 100.

Act on Your Strategy
Step 1 The equation for the dissociation equilibrium of propanoic acid in water is given in the ICE table.

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>$\text{C}_4\text{H}_4\text{N}_2\text{O}_3(\text{aq})$</th>
<th>$\text{H}_2\text{O}(\ell)$</th>
<th>$\text{C}_4\text{H}_3\text{N}_2\text{O}_3^- (\text{aq})$</th>
<th>$\text{H}_3\text{O}^+ (\text{aq})$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.10</td>
<td>0</td>
<td>$-0$</td>
<td></td>
</tr>
<tr>
<td>Change</td>
<td>$-x$</td>
<td>$+x$</td>
<td>$+x$</td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>$0.10 - x$</td>
<td>$x$</td>
<td>$x$</td>
<td></td>
</tr>
</tbody>
</table>

Step 2

$K_a = \frac{[\text{C}_4\text{H}_3\text{N}_2\text{O}_3^-][\text{H}_3\text{O}^+]}{[\text{C}_4\text{H}_4\text{N}_2\text{O}_3]}$

$= \frac{(x)(x)}{(0.10 - x)}$

Step 3

The value of $x$ is equal to $[\text{H}_3\text{O}^+]$ and $[\text{C}_4\text{H}_3\text{N}_2\text{O}_3^-]$. $[\text{H}_3\text{O}^+] = 10^{-2.50}$

$= 3.16 \times 10^{-3} \text{ mol/L}$

Step 4

$K_a = \frac{(3.16 \times 10^{-3})^2}{0.10 - (3.16 \times 10^{-3})}$

$= 1.0 \times 10^{-4}$

Step 5

Percent dissociation $= \frac{3.16 \times 10^{-3} \text{ mol/L}}{0.10 \text{ mol/L}} \times 100$

$= 3.2\%$

Check Your Solution
The value of $K_a$ and the percent dissociation are reasonable for a weak acid.

21. Problem
A solution of hydrofluoric acid has a molar concentration of 0.0100 mol/L. What is the pH of this solution?

What Is Required?
You need to calculate the pH of the solution.

What Is Given?
Initial $[\text{HF}] = 0.0100 \text{ mol/L}$
The acid dissociation constant for hydrofluoric acid (listed in Table 8.2) is: $K_a = 6.6 \times 10^{-4}$
Plan Your Strategy

**Step 1** Check the value of $\frac{[HF]}{K_a}$ to see whether or not the amount that dissociates is negligible compared with the initial concentration of the acid.

**Step 2** Write the equation for the dissociation equilibrium of hydrofluoric acid in water. Then set up an ICE table.

**Step 3** Write the equation for the acid dissociation constant. Substitute equilibrium concentrations into the equilibrium expression. Solve the equilibrium expression for $x$. If the amount that dissociates is not negligible compared with the initial concentration of acid, you will need to solve a quadratic equation.

**Step 4** Calculate the pH using $\text{pH} = -\log[H_3O^+]$.

Act on Your Strategy

**Step 1**

$$\text{pH} = \frac{0.0100}{6.6 \times 10^{-4}} = 152$$

Since this value is less than 500, the amount that dissociates cannot be neglected compared to the initial concentration of the acid.

**Step 2**

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>$[HF]_{eq}$</th>
<th>$[H_2O]$</th>
<th>$[F^-]_{eq}$</th>
<th>$[H_3O^+]_{eq}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.0100</td>
<td>0</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>$-x$</td>
<td>$+x$</td>
<td>$+x$</td>
<td>(0.0100 - x)</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.0100 - $x$</td>
<td>$x$</td>
<td>$x$</td>
<td>$\approx 0$</td>
</tr>
</tbody>
</table>

**Step 3**

$$K_a = \frac{[F^-][H_3O^+]}{[HF]} = \frac{(x)(x)}{0.0100 - x}$$

This equation must be rearranged into a quadratic equation.

$$x^2 + (6.6 \times 10^{-4})x - (6.6 \times 10^{-6}) = 0$$

Recall that a quadratic equation of the form $ax^2 + bx + c = 0$ has the solution

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$x = \frac{6.6 \times 10^{-4} \pm \sqrt{4.36 \times 10^{-7} - 2.64 \times 10^{-5}}}{2}$$

$$x = -2.92 \times 10^{-3} \text{ and } x = 2.26 \times 10^{-3}$$

The value $x = -2.92 \times 10^{-3}$ is not physically possible. It would result in negative concentrations of $F^-_{(aq)}$ and $H_3O^+$ at equilibrium.

Therefore, $[H_3O^+]_{eq} = 2.26 \times 10^{-3}$ mol/L

**Step 4**

$$\text{pH} = -\log 2.26 \times 10^{-3}$$

$$= 2.65$$

Check Your Solution

The pH indicates an acidic solution, as expected. The value for $K_a$ has two significant digits, and the pH has two digits following the decimal place. Next, check the equilibrium values:

$$\frac{(2.26 \times 10^{-3})^2}{0.0100 - 2.26 \times 10^{-3}} = 6.6 \times 10^{-4}$$

This is equal to $K_a$. 

Chapter 8 Acids, Bases, and pH • MHR 142
22. Problem
Hypochlorous acid, HOCl, is used as a bleach, and a germ-killer. A chemist finds that 0.027% of hypochlorous acid molecules are dissociated in a 0.40 mol/L solution of the acid. What is the value of $K_a$ for the acid?

What Is Required?
You need to calculate the value of $K_a$ for hypochlorous acid.

What Is Given?
You know that 0.027% of HOCl molecules dissociate in a 0.40 mol/L solution.

Plan Your Strategy
Step 1 Calculate $[H_3O^+]$ using the following equation:

\[
\text{Percent dissociation} = \frac{[H_3O^+]}{[HOCl]} \times 100
\]

Step 2 Write the equation for the dissociation equilibrium of hypochlorous acid in water. Then set up an ICE table.

Step 3 Write the equation for the acid dissociation constant. Substitute equilibrium terms into the equation and calculate $K_a$.

Act on Your Strategy
Step 1
\[
0.027 = \frac{[H_3O^+]}{0.40} \times 100
\]

Therefore, $[H_3O^+] = 1.08 \times 10^{-4}$

\[= [OCl^-]\]

Step 2

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>HOCl(_{\text{aq}})</th>
<th>H(<em>2)O(</em>{\text{l}})</th>
<th>H(<em>3)O(</em>{\text{aq}}^+)</th>
<th>OCl(^-)(_{\text{aq}})</th>
<th>H(<em>3)O(</em>{\text{aq}}^+)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Equilibrium</td>
<td>(0.40 – 1.08 \times 10^{-4}) \approx 0.40</td>
<td>\approx 1.08 \times 10^{-4}</td>
<td>\approx 1.08 \times 10^{-4}</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Step 3
\[
K_a = \frac{[OCl^-][H_3O^+]}{[HOCl]}
\]
\[= \frac{(1.08 \times 10^{-4})^2}{0.40}
\]
\[= 2.9 \times 10^{-8}
\]

Check Your Solution
The value of $K_a$ is reasonable for a weak acid.

23. Problem
The word “butter” comes from the Greek butyros. Butanoic acid (common name: butyric acid) gives rancid butter its distinctive odour. Calculate the pH of a 1.00 \times 10^{-2} mol/L solution of butanoic acid ($K_a = 1.51 \times 10^{-5}$).

What Is Required?
You need to calculate the pH of the solution.

What Is Given?
Initial $[\text{CH}_3\text{CH}_2\text{CH}_2\text{COOH}] = 1.00 \times 10^{-2}$ mol/L.
The acid dissociation constant for butanoic acid is $1.51 \times 10^{-5}$.

Plan Your Strategy
Butanoic acid is monoprotic, with the formula $\text{CH}_3\text{CH}_2\text{CH}_2\text{COOH}$.

Step 1 Check the value of $\frac{[\text{CH}_3\text{CH}_2\text{CH}_2\text{COOH}]}{K_a}$ to see whether or not the amount that dissociates is negligible compared with the initial concentration of the acid.

Step 2 Write the equation for the dissociation equilibrium of butanoic acid in water. Then set up an ICE table.
Step 3 Write the equation for the acid dissociation constant. Substitute equilibrium concentrations into the equilibrium expression. Solve the equilibrium expression for $x$.

Step 4 Use the equation $pH = -\log[H_3O^+]$.

Act on Your Strategy

Step 1 

\[
\frac{[CH_3CH_2CH_2COOH]}{K_a} = \frac{1.00 \times 10^{-2}}{1.51 \times 10^{-5}} = 662
\]

Since this value is greater than 500, the amount that dissociates can be neglected compared to the initial concentration of the acid.

Step 2

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>CH$_3$CH$_2$CH$<em>2$COOH$</em>{aq}$</th>
<th>H$<em>2$O$</em>{aq}$</th>
<th>CH$_3$CH$_2$CH$<em>2$COO$^-$$</em>{aq}$</th>
<th>H$<em>3$O$^+$</em>{aq}</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>1.00 $\times$ 10$^{-2}$</td>
<td>0</td>
<td>-0</td>
<td></td>
</tr>
<tr>
<td>Change</td>
<td>-$x$</td>
<td>+$x$</td>
<td>+$x$</td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>1.00 $\times$ 10$^{-2}$ - $x$</td>
<td>$x$</td>
<td>$x$</td>
<td>1.00 $\times$ 10$^{-2}$</td>
</tr>
</tbody>
</table>

\[
K_a = \frac{[CH_3CH_2CH_2COO^-][H_3O^+]}{[CH_3CH_2CH_2COOH]} = 1.51 \times 10^{-5}
\]

\[
x = \sqrt{1.51 \times 10^{-7}} = 3.89 \times 10^{-4}
\]

The value $x = -3.89 \times 10^{-4}$ is not physically possible. It would result in negative concentrations of CH$_3$CH$_2$CH$_2$COO$^-$ and H$_3$O$^+$ at equilibrium.

Therefore, [H$_3$O$^+$] = 3.89 $\times$ 10$^{-4}$ mol/L.

Step 4

\[
pH = -\log 3.89 \times 10^{-4} = 3.411
\]

Check Your Solution

The given data has three significant digits, and the pH has three digits following the decimal place. The assumption that the amount of acid that dissociates can be neglected is valid because:

\[
1.00 \times 10^{-2} - 3.89 \times 10^{-4} = 9.61 \times 10^{-3}
\]

The error introduced by the assumption is 4 $\times$ 10$^{-4}$, or 4%. This is less than the typical experimental error in $K_a$ values, which is 5%. Next, check the equilibrium values:

\[
\frac{(3.98 \times 10^{-4})^2}{1.00 \times 10^{-2}} = 1.51 \times 10^{-5}
\]

This is equal to $K_a$.

24. Problem

Caproic acid, C$_5$H$_{11}$COOH, occurs naturally in coconut and palm oil. It is a weak monoprotic acid, with $K_a = 1.3 \times 10^{-5}$. A certain aqueous solution of caproic acid has pH = 2.94. How much acid was dissolved to make 100 mL of this solution?

What Is Required?

You need to calculate how much caproic acid was dissolved to make 100 mL of solution.

What Is Given?

The acid dissociation constant for caproic acid is $1.3 \times 10^{-5}$.

pH = 2.94
Plan Your Strategy

**Step 1**  Calculate $[\text{H}_3\text{O}^+]$ using the equation $[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$.

**Step 2**  Write the equation for the dissociation equilibrium of caproic acid in water. Then show the equilibrium concentration below each chemical.

**Step 3**  Write the equation for the acid dissociation constant. Substitute equilibrium concentrations into the equilibrium expression. Solve the equilibrium expression for $[\text{C}_5\text{H}_{11}\text{COOH}]$.

**Step 4**  Calculate the amount of caproic acid dissolved in 100 mL of solution using the following equation:

$$[\text{C}_5\text{H}_{11}\text{COOH}] = \frac{\text{C}_5\text{H}_{11}\text{COOH (mol)}}{\text{Volume (L)}}$$

**Step 5**  Calculate the mass of caproic acid using the following equation:

$$\text{Amount (mol)} = \frac{\text{Mass (g)}}{\text{Molar mass (g/mol)}}$$

Act on Your Strategy

**Step 1**  $[\text{H}_3\text{O}^+] = 10^{-2.94}$

$$= 1.15 \times 10^{-3} \text{ mol/L}$$

**Step 2**  Concentration (mol/L)  |  $\text{C}_5\text{H}_{11}\text{COOH}_{(aq)}$  +  $\text{H}_2\text{O}_{(l)}$  $\rightleftharpoons$  $\text{C}_5\text{H}_{11}\text{COO}^-_{(aq)}$  +  $\text{H}_3\text{O}^+_{(aq)}$
--- | --- | --- | ---
Equilibrium  | $[\text{C}_5\text{H}_{11}\text{COOH}]$  −  $1.15 \times 10^{-3} = \frac{1.15 \times 10^{-3}}{[\text{C}_5\text{H}_{11}\text{COOH}]}$

**Step 3**  $$K_a = \frac{[\text{C}_5\text{H}_{11}\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{C}_5\text{H}_{11}\text{COOH}]}$$

$$1.3 \times 10^{-5} = \frac{(1.15 \times 10^{-3})^2}{[\text{C}_5\text{H}_{11}\text{COOH}]}$$

Therefore, $[\text{C}_5\text{H}_{11}\text{COOH}] = 0.102 \text{ mol/L}$

**Step 4**  $0.102 \text{ mol/L} = \frac{\text{C}_5\text{H}_{11}\text{COOH (mol)}}{0.100 \text{ L}}$

Therefore, the amount of $\text{C}_5\text{H}_{11}\text{COOH} (\text{mol}) = 1.02 \times 10^{-2} \text{ mol}$

**Step 5**  Molar mass of $\text{C}_5\text{H}_{11}\text{COOH} (\text{mol}) = 116 \text{ g/mol}$

$$1.02 \times 10^{-2} \text{ mol} = \frac{\text{Mass (g)}}{116 \text{ g/mol}}$$

Therefore, mass = $1.2 \text{ g}$

Check Your Solution

The data given in the problem has two significant digits, and the final answer has two significant digits.

Solutions for Practice Problems

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25. Problem

Carbonated beverages contain a solution of carbonic acid. Carbonic acid is also important in forming the ions that are present in blood.

$\text{CO}_2_{(aq)} + \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{H}_2\text{CO}_3_{(aq)}$

$\text{H}_2\text{CO}_3_{(aq)} + \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{HCO}_3^-_{(aq)} + \text{H}_3\text{O}^+_{(aq)} \quad K_{a1} = 4.5 \times 10^{-7}$

$\text{HCO}_3^-_{(aq)} + \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{CO}_3^{2-}_{(aq)} + \text{H}_3\text{O}^+_{(aq)} \quad K_{a2} = 4.7 \times 10^{-11}$

Calculate the pH of a solution of $5.0 \times 10^{-4} \text{ mol/L}$ carbonic acid. What is $[\text{CO}_3^{2-}]$ in the solution?
What Is Required?
You need to find pH and $[\text{CO}_3^{2-}]$.

What Is Given?
You know that $[\text{H}_2\text{CO}_3]$ is $5.0 \times 10^{-4}$ mol/L. From data tables, you can find $K_{a1} = 4.5 \times 10^{-7}$ and $K_{a2} = 4.7 \times 10^{-11}$.

Plan Your Strategy

Step 1 Determine whether or not the dissociation of $\text{H}_2\text{CO}_3$ is negligible, compared with the initial concentration.

Step 2 Write the equation for the dissociation equilibrium of carbonic acid in water. Then set up an ICE table.

Step 3 Write the dissociation equation for $K_{a1}$.

Step 4 Solve the equation for $x$.

Step 5 Calculate the pH of the solution using the equation $\text{pH} = -\log[\text{H}_3\text{O}^+]$.

Step 6 Use the equation $[\text{CO}_3^{2-}] = K_{a2}$.

Act on Your Strategy

Step 1

$$\frac{[\text{H}_2\text{CO}_3]}{K_{a1}} = \frac{5.0 \times 10^{-4}}{4.5 \times 10^{-7}} = 1.1 \times 10^3$$

Therefore, the amount that dissociates is probably negligible compared with $5.0 \times 10^{-4}$.

Step 2

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>$\text{H}_2\text{CO}_3$</th>
<th>$\text{H}<em>2\text{O}</em>(l)$</th>
<th>$\text{HCO}_3^-$</th>
<th>$\text{H}_3\text{O}^+$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>$5.0 \times 10^{-4}$</td>
<td>$0$</td>
<td>$0$</td>
<td>$0$</td>
</tr>
<tr>
<td>Change</td>
<td>$-x$</td>
<td>$+x$</td>
<td>$+x$</td>
<td>$+x$</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>$5.0 \times 10^{-4} - x = x$</td>
<td>$x$</td>
<td>$x$</td>
<td>$x$</td>
</tr>
</tbody>
</table>

Step 3

$$K_{a1} = \frac{[\text{HCO}_3^-][\text{H}_3\text{O}^+]}{[\text{H}_2\text{CO}_3]}$$

Step 4

$$4.5 \times 10^{-7} = \frac{(x)(x)}{5.0 \times 10^{-4}}$$

$$x = \sqrt{2.25 \times 10^{-10}}$$

$$x = +1.5 \times 10^{-5} \text{ or } x = -1.5 \times 10^{-5}$$

The negative root is not chemically possible.

Therefore, $[\text{H}_3\text{O}^+] = 1.5 \times 10^{-5}$ mol/L.

Step 5

$$\text{pH} = -\log 1.5 \times 10^{-5} = 4.82$$

Step 6

$$[\text{CO}_3^{2-}] = K_{a2}$$

Therefore, $[\text{CO}_3^{2-}] = 4.7 \times 10^{-11}$ mol/L.

Check Your Solution

The first dissociation of carbonic acid determines $[\text{H}_3\text{O}^+]$, because the second dissociation is much weaker. The assumption that the amount of carbonic acid that dissociates is negligible is valid because: $5.0 \times 10^{-4} - 1.5 \times 10^{-5} = 4.85 \times 10^{-4}$ (an error of only 3%)

Check the value of $K_{a1}$:

$$K_{a1} = \frac{(1.5 \times 10^{-5})^2}{5.0 \times 10^{-4}} = 4.5 \times 10^{-7}$$

This is equal to the value of $K_{a1}$ for carbonic acid.
26. Problem
Adipic acid is a diprotic acid that is used to manufacture nylon. Its formula can be abbreviated to H$_2$Ad. The acid dissociation constants for adipic acid are $K_{a1} = 3.9 \times 10^{-5}$ and $K_{a2} = 3.87 \times 10^{-6}$. What is the pH of a 0.085 mol/L solution of adipic acid?

What Is Required?
You need to find the pH of the solution.

What Is Given?
You know that [H$_2$Ad] = 0.085 mol/L. You are given $K_{a1} = 3.9 \times 10^{-5}$ and $K_{a2} = 3.87 \times 10^{-6}$.

Plan Your Strategy
Step 1 Determine whether or not the dissociation of H$_2$Ad is negligible, compared with the initial concentration.
Step 2 Write the equation for the dissociation equilibrium of adipic acid in water. Then set up an ICE table.
Step 3 Write the dissociation equation for $K_{a1}$ and solve the equation for $x$.
Step 4 Calculate the pH of the solution using pH = $-\log$[H$_3$O$^+$].

Act on Your Strategy
Step 1 $\frac{[H_2Ad]}{K_{a1}} = \frac{0.085}{3.9 \times 10^{-5}} = 2.2 \times 10^3$
Therefore, the amount that dissociates is probably negligible compared with 0.085.

Step 2 Concentration (mol/L)  
<table>
<thead>
<tr>
<th></th>
<th>H$<em>2$Ad$</em>{aq}$</th>
<th>+</th>
<th>H$<em>3$O$</em>{aq}$</th>
<th>⇌</th>
<th>HAd$^-_{aq}$</th>
<th>+</th>
<th>H$<em>3$O$^+</em>{aq}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.085</td>
<td></td>
<td>0</td>
<td></td>
<td>-</td>
<td></td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>$-x$</td>
<td></td>
<td>+x</td>
<td></td>
<td>+x</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>$0.085 - x = 0.085$</td>
<td></td>
<td>$x$</td>
<td></td>
<td>$x$</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Step 3 $K_{a1} = \frac{[HAd^-][H_3O^+]}{[H_2Ad]}$
$3.9 \times 10^{-5} = \frac{(x)(x)}{0.085}$
$x = \sqrt{3.32 \times 10^{-6}}$
$x = +1.82 \times 10^{-3}$ or $x = -1.82 \times 10^{-3}$
Only the positive root is chemically possible, therefore [H$_3$O$^+$] = $1.8 \times 10^{-3}$ mol/L.

Step 4 pH = $-\log 1.8 \times 10^{-3}$
= 2.74

Check Your Solution
The first dissociation of adipic acid determines [H$_3$O$^+$], because the second dissociation is much weaker. The assumption that the amount of adipic acid that dissociates is negligible is valid because:
$0.085 - 1.8 \times 10^{-3} = 0.083$ (an error of less than 3%)
The data given in the problem has two significant digits, and the pH has two digits after the decimal. Check the value of $K_{a1}$:
$K_{a1} = \frac{(1.8 \times 10^{-3})^2}{0.085}$
= $3.8 \times 10^{-5}$
This is equal to the value of $K_{a1}$ for adipic acid, within rounding errors.
27. Problem
Hydrosulfuric acid, \( H_2S(aq) \), is a weak diprotic acid that is sometimes used in analytical work. It is used to precipitate metal sulfides, which tend to be very insoluble. Calculate the pH and \([\text{HS}^-]\) of a \( 7.5 \times 10^{-3} \) mol/L solution.

What Is Required?
You need to find pH and \([\text{HS}^-]\).

What Is Given?
You know that \([H_2S] = 7.5 \times 10^{-3} \) mol/L. Look in Appendix E, on page 597, to find \( K_{a1} = 8.9 \times 10^{-8} \).

Plan Your Strategy
Step 1  Determine whether or not the dissociation of \( H_2S \) is negligible, compared with the initial concentration.

Step 2  Write the equation for the dissociation equilibrium of hydrosulfuric acid in water. Then set up an ICE table.

Step 3  Write the dissociation equation for \( K_{a1} \) and solve the equation for \( x \).

Step 4  Calculate the pH of the solution using \( \text{pH} = -\log[H_3O^+] \).

Act on Your Strategy
Step 1  
\[
\frac{[H_2S]}{K_{a1}} = \frac{7.5 \times 10^{-3}}{8.9 \times 10^{-8}} = 8.4 \times 10^4
\]
Therefore, the amount that dissociates is negligible compared with \( 7.5 \times 10^{-3} \).

Step 2  
<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>( H_2S(aq) )</th>
<th>( H_2O(l) )</th>
<th>( \text{⇌} )</th>
<th>( \text{HS}^-_{(aq)} )</th>
<th>( H_3O^+_{(aq)} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>( 7.5 \times 10^{-3} )</td>
<td>0</td>
<td>(-0)</td>
<td>( \text{≈} )</td>
<td>( \text{≈} )</td>
</tr>
<tr>
<td>Change</td>
<td>(-x)</td>
<td>(+x)</td>
<td>(+x)</td>
<td>( \approx )</td>
<td>( \approx )</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>( 7.5 \times 10^{-3} - x =)</td>
<td>( x)</td>
<td>( x)</td>
<td>( 7.5 \times 10^{-3} )</td>
<td></td>
</tr>
</tbody>
</table>

Step 3  
\[
K_{a1} = \frac{[\text{HS}^-][H_3O^+]}{[H_2S]}
\]
\[
8.9 \times 10^{-8} = \frac{(x)(x)}{7.5 \times 10^{-3}}
\]
\[
x = \sqrt{6.68 \times 10^{-10}}
\]
\[
x = +2.6 \times 10^{-5} \text{ or } x = -2.6 \times 10^{-5}
\]
The negative root is not chemically possible.

Therefore, \([H_3O^+] = [\text{HS}^-] = 2.6 \times 10^{-5} \) mol/L

Step 4  
\[
\text{pH} = -\log 2.6 \times 10^{-5} = 4.59
\]

Check Your Solution
The first dissociation of hydrosulfuric acid determines \([H_3O^+]\), because the second dissociation is much weaker. The assumption that the amount of hydrosulfuric acid that dissociates is negligible is valid because, using the rules for subtracting numbers:
\[
7.5 \times 10^{-3} - 2.6 \times 10^{-5} = 7.5 \times 10^{-3}
\]
Check the value of \( K_{a1} \):
\[
K_{a1} = \frac{(2.6 \times 10^{-5})^2}{7.5 \times 10^{-3}}
\]
\[
= 9.0 \times 10^{-8}
\]
This is equal to the value of \( K_{a1} \) for hydrosulfuric acid, within the error introduced by mathematical rounding.
28. Problem
What is the value of $K_a$ when water acts as a Brønsted-Lowry acid? Write the expression for $K_a$ if water acts as a diprotic acid.

What Is Required?
You need to find the value of $K_a$ for water. Then write the expression for $K_{a2}$ for water.

What Is Given?
$K_w = 1.0 \times 10^{-14}$

Plan Your Strategy
Step 1 Write the equation for the dissociation equilibrium of water.
Step 2 Write the dissociation equation for $K_{a1}$. Find the value, using $K_w = 1.0 \times 10^{-14}$.
Step 3 Write the equation for the dissociation equilibrium of $\text{OH}^-(aq)$.
Step 4 Write the dissociation equation for $K_{a2}$.

Act on Your Strategy
Step 1 $\text{H}_2\text{O}(l) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{OH}^-(aq)$
Step 2 This is a heterogeneous equilibrium, and the concentration of pure liquids are always included in the value for the equilibrium constant. $K_{a1} = [\text{H}_3\text{O}^+] \cdot [\text{OH}^-] = K_w$
Therefore, $K_a$ for water $= 1.0 \times 10^{-14}$ at 25˚C
Step 3 $\text{OH}^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{O}_2^-(aq)$
Step 4 $K_{a2} = \frac{[\text{O}_2^-][\text{H}_3\text{O}^+]}{[\text{OH}^-]}$

Check Your Solution
The value for $K_{a1}$ indicates a weak acid. The expression for $K_{a2}$ has product concentration terms in the numerator, and reactant concentration terms in the denominator.

Solutions for Practice Problems
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29. Problem
An aqueous solution of household ammonia has a molar concentration of 0.105 mol/L. Calculate the pH of the solution.

What Is Required?
You need to find the pH of the solution.

What Is Given?
From Table 8.3, $K_b = 1.8 \times 10^{-5}$
Concentration of ammonia solution = 0.105 mol/L

Plan Your Strategy
Step 1 Check the value of $\frac{[\text{NH}_3]}{K_b}$ to see if the amount that dissociates can be neglected compared to the initial $[\text{NH}_3]$.
Step 2 Write the equation for the equilibrium reaction of ammonia in water. Then set up an ICE table.
Step 3 Write the equation for the base dissociation constant. Substitute equilibrium terms into the equation and solve for $x$.
Step 4 Use the following equations:
$p\text{OH} = -\log[\text{OH}^-]$ $p\text{H} = 14.00 - p\text{OH}$
Act on Your Strategy

Step 1
\[
\frac{0.105}{1.8 \times 10^{-5}} = 5.8 \times 10^4
\]
Since this value is greater than 500, the amount that dissociates can be neglected compared to the initial concentration of ammonia.

Step 2

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>NH₃(aq)</th>
<th>H₂O(l)</th>
<th>NH₄⁺(aq)</th>
<th>OH⁻(aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.105</td>
<td>0</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>−x</td>
<td>+x</td>
<td>+x</td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.105 − x</td>
<td>0.105</td>
<td>x</td>
<td>x</td>
</tr>
</tbody>
</table>

\[
K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}
\]

\[
1.8 \times 10^{-5} = \frac{x^2}{0.105}
\]

\[
x = \sqrt{1.89 \times 10^{-6}}
\]

\[
x = \pm 1.37 \times 10^{-3}
\]

The negative root is not reasonable.

Therefore, \(x = 1.37 \times 10^{-3}\) mol/L = [OH⁻]

Step 5

\[\text{pOH} = −\log[OH^-]\]

\[= 2.86\]

\[\text{pH} = 14.00 − 2.86\]

\[= 11.14\]

Check Your Solution

The pH of the solution is greater than 7, as expected for a basic solution.

30. Problem

Hydrazine, N₂H₄, has been used as a rocket fuel. The concentration of an aqueous solution of hydrazine is \(5.9 \times 10^{-2}\) mol/L. Calculate the pH of the solution.

What Is Required?

You need to find the pH of the solution.

What Is Given?

From Table 8.3, \(K_b = 1.7 \times 10^{-6}\).

\([N₂H₄] = 5.9 \times 10^{-2}\) mol/L

Plan Your Strategy

Step 1
Check the value of \(\frac{[N₂H₄]}{K_b}\) to see if the amount that dissociates can be neglected compared to the initial \([N₂H₄]\).

Step 2
Write the equation for the equilibrium reaction of hydrazine in water. Then set up an ICE table.

Step 3
Write the equation for the base dissociation constant. Substitute equilibrium terms into the equation and solve for \(x\).

Step 4
Use the following equations:

\[\text{pOH} = −\log[OH^-]\]

\[\text{pH} = 14.00 − \text{pOH}\]

Act on Your Strategy

Step 1

\[
\frac{5.9 \times 10^{-2}}{1.7 \times 10^{-6}} = 3.5 \times 10^4
\]

Since this value is greater than 500, the amount that dissociates can be neglected compared to the initial concentration of hydrazine.
Step 2

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>(N_2H_4(aq))</th>
<th>(H_2O(l))</th>
<th>(N_2H_5^+(aq))</th>
<th>(OH^-)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>(5.9 \times 10^{-2})</td>
<td>0</td>
<td>(-0)</td>
<td></td>
</tr>
<tr>
<td>Change</td>
<td>(-x)</td>
<td>(+x)</td>
<td>(+x)</td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>(5.9 \times 10^{-2} - x)</td>
<td>(x)</td>
<td>(x)</td>
<td></td>
</tr>
</tbody>
</table>

\[K_b = \frac{[N_2H_5^+][OH^-]}{[N_2H_4]}\]

\[1.7 \times 10^{-6} = \frac{x^2}{5.9 \times 10^{-2}}\]

\[x = \sqrt{1.00 \times 10^{-7}}\]

\[x = \pm 3.17 \times 10^{-4}\]

The negative root is not reasonable.
Therefore, \(x = 3.17 \times 10^{-4}\) mol/L = [OH⁻]

Step 5

\[pOH = -\log 3.17 \times 10^{-4}\]

\[= 3.50\]

\[pH = 14.00 - 3.50\]

\[= 10.50\]

Check Your Solution

The pH of the solution is greater than 7, as expected for a basic solution.

31. Problem

Morphine, \(C_{17}H_{19}NO_3\), is a naturally occurring base that is used to control pain. A \(4.5 \times 10^{-3}\) mol/L solution has a pH of 9.93. Calculate \(K_b\) for morphine.

What Is Required?

You need to find \(K_b\).

What Is Given?

\([C_{17}H_{19}NO_3] = 4.5 \times 10^{-3}\) mol/L

\(pH = 9.93\)

Plan Your Strategy

Step 1

Write the equation for the equilibrium reaction of morphine in water. Then set up an ICE table.

Step 2

Write the equation for the base dissociation constant. Substitute equilibrium terms into the equation.

Step 3

Use the equation \(pOH = 14.00 - pH\).

Step 4

Use the equation \([OH^-] = 10^{-pOH}\).

Step 5

Substitute for \(x\) into the equilibrium equation. Calculate the value of \(K_b\).

Act on Your Strategy

Step 1

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>(C_{17}H_{19}NO_3)</th>
<th>(H_2O)</th>
<th>(C_{17}H_{19}NO_3H^+)</th>
<th>(OH^-)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>(4.5 \times 10^{-3})</td>
<td>0</td>
<td>(-0)</td>
<td></td>
</tr>
<tr>
<td>Change</td>
<td>(-x)</td>
<td>(+x)</td>
<td>(+x)</td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>(4.5 \times 10^{-3} - x)</td>
<td>(x)</td>
<td>(x)</td>
<td></td>
</tr>
</tbody>
</table>

\[K_b = \frac{[C_{17}H_{19}NO_3H^+][OH^-]}{[C_{17}H_{19}NO_3]}\]

\[= \frac{(x)(x)}{(4.5 \times 10^{-3} - x)}\]

Step 3

\[pOH = 14.00 - 9.93 = 4.07\]
Step 4 \[ [\text{OH}^-] = 10^{-4.07} \]
\[ = 8.5 \times 10^{-5} \text{ mol/L} \]

Step 5 \[ 4.5 \times 10^{-3} - 8.5 \times 10^{-5} = 4.4 \times 10^{-3} \]
\[ K_b = \frac{(8.5 \times 10^{-5})^2}{4.4 \times 10^{-3}} \]
\[ = 1.6 \times 10^{-6} \]

Check Your Solution
The value for $K_b$ is reasonable for a weak organic base. The final answer has two significant digits, consistent with the two decimal places in the given pH.

32. Problem
Methylamine, CH$_3$NH$_2$, is a fishy-smelling gas at room temperature. It is used to manufacture several prescription drugs, including methamphetamine. Calculate [OH$^-$] and pOH of a 0.25 mol/L aqueous solution.

What Is Required?
You need to find [OH$^-$] and the pOH of the solution.

What Is Given?
From Table 8.3, $K_b = 4.4 \times 10^{-4}$
[CH$_3$NH$_2$] = 0.25 mol/L

Plan Your Strategy
Step 1 Check the value of $\frac{[\text{CH}_3\text{NH}_2]}{K_b}$ to see if the amount that dissociates can be neglected compared to the initial [CH$_3$NH$_2$].
Step 2 Write the equation for the equilibrium reaction of methylamine in water. Then set up an ICE table.
Step 3 Write the equation for the base dissociation constant. Substitute equilibrium terms into the equation and solve for $x$.
Step 4 Use the equation $\text{pOH} = -\log[\text{OH}^-]$.

Act on Your Strategy
Step 1 \[ \frac{0.25}{4.4 \times 10^{-4}} = 568 \]
Since this value is greater than 500, the amount that dissociates can probably be neglected compared to the initial concentration of methylamine.

Step 2 Concentration (mol/L)

<table>
<thead>
<tr>
<th>Initial</th>
<th>CH$_3$NH$_2$(aq)</th>
<th>+</th>
<th>H$_2$O(l)</th>
<th>⇌</th>
<th>CH$_3$NH$_3$$^+$</th>
<th>+</th>
<th>OH$^-$</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.25</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
</tr>
</tbody>
</table>

Change  
| -x      | +x              | +x |

Equilibrium  
| 0.25 - x | 0.25 | x | x |

Step 3 \[ K_b = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]} \]
\[ 4.4 \times 10^{-4} = \frac{x^2}{0.25} \]
\[ x = \sqrt{1.1 \times 10^{-4}} \]
\[ x = \pm 1.0 \times 10^{-2} \]
The negative root is not reasonable.
Therefore, $x = 1.0 \times 10^{-2}$ mol/L = [OH$^-$]

Step 4 \[ \text{pOH} = -\log 1.0 \times 10^{-2} \]
\[ = 1.98 \]
Check Your Solution
The pOH of the solution is less than 7, as expected for a basic solution.

33. Problem
At room temperature, trimethylamine, \((\text{CH}_3)_3\text{N}\), is a gas with a strong ammonia-like odour. Calculate \([\text{OH}^-]\) and the percent of trimethylamine molecules that react with water in a 0.22 mol/L aqueous solution.

What Is Required?
You need to find \([\text{OH}^-]\) and the percent dissociation.

What Is Given?
From Table 8.3, \(K_b = 6.5 \times 10^{-5}\)
\([\text{CH}_3\text{N}] = 0.22\ \text{mol/L}\)

Plan Your Strategy
Step 1 Check the value of \(\frac{[\text{CH}_3\text{N}]}{K_b}\) to see if the amount that dissociates can be neglected compared to the initial \([\text{CH}_3\text{N}].\)
Step 2 Write the equation for the equilibrium reaction of trimethylamine in water. Then set up an ICE table.
Step 3 Write the equation for the base dissociation constant. Substitute equilibrium terms into the equation and solve for \(x\).
Step 4 Calculate the percent of molecules that dissociated.

Act on Your Strategy
Step 1 \(\frac{0.22}{6.5 \times 10^{-5}} = 3.4 \times 10^3\)
Since this value is greater than 500, the amount that dissociates can be neglected compared to the initial concentration of trimethylamine.
Step 2

<table>
<thead>
<tr>
<th>Concentration (mol/L)</th>
<th>((\text{CH}_3)<em>3\text{N}</em>{(aq)} + \text{H}<em>2\text{O}</em>{(l)} \rightleftharpoons (\text{CH}<em>3)<em>3\text{NH}^+</em>{(aq)} + \text{OH}^-</em>{(aq)})</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.22</td>
</tr>
<tr>
<td>Change</td>
<td>(-x)</td>
</tr>
<tr>
<td>Change</td>
<td>(+x)</td>
</tr>
<tr>
<td>Change</td>
<td>(+x)</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.22 (- x = 0.22) (x) (x)</td>
</tr>
</tbody>
</table>

Step 3 \(K_b = \frac{[\text{(CH}_3\text{N})\text{NH}^+][\text{OH}^-]}{[\text{CH}_3\text{N}]}\)
\(6.5 \times 10^{-5} = \frac{x^2}{0.22}\)
\(x = \sqrt{1.43 \times 10^{-5}}\)
\(x = \pm 3.8 \times 10^{-3}\)
The negative root is not reasonable.
Therefore, \([\text{OH}^-] = 3.8 \times 10^{-3}\ \text{mol/L}\).
Step 4 Percent dissociation = \(\frac{3.8 \times 10^{-3}}{0.22} \times 100\)
= 1.7%

Check Your Solution
The \([\text{OH}^-]\) of the solution is greater than \(1.0 \times 10^{-7}\), as expected for a basic solution. The percent dissociation is small, consistent with trimethylamine being a weak base.

34. Problem
An aqueous solution of ammonia has a pH of 10.85. What is the concentration of the solution?
What Is Required?
You need to find $[\text{NH}_3]$.

What Is Given?
You know the pH of the solution is 10.85. $K_b$ for aqueous ammonia = $1.8 \times 10^{-5}$.

Plan Your Strategy

Step 1 Calculate $[\text{OH}^-]$ using the following equations:
\[
pOH = 14.00 - \text{pH} \\
[\text{OH}^-] = 10^{-pOH}
\]

Step 2 Write the equation for the equilibrium reaction of ammonia in water. Then write down the equilibrium concentrations.

Step 3 Write the equation for the base dissociation constant. Substitute equilibrium terms into the equation and solve for $[\text{NH}_3]$.

Act on Your Strategy

\[
pOH = 14.00 - 10.85 \\
= 3.15 \\
[\text{OH}^-] = 10^{-3.15} = 7.1 \times 10^{-4} \text{ mol/L}
\]

\begin{tabular}{ l c c }
  Equilibrium & $[\text{NH}_3] - 7.1 \times 10^{-4} = [\text{NH}_3]$ & $7.1 \times 10^{-4}$ & $7.1 \times 10^{-4}$
\end{tabular}

Step 3

\[
K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} \\
1.8 \times 10^{-5} = \frac{(7.1 \times 10^{-4})^2}{[\text{NH}_3]} \\
\text{Therefore, } [\text{NH}_3] = 2.8 \times 10^{-2} \text{ mol/L}
\]

Check Your Solution

The assumption that the amount of $\text{NH}_3$ that dissociates is negligible compared with the initial concentration is valid because:

$2.8 \times 10^{-2} - 7.1 \times 10^{-4} = 2.7 \times 10^{-2}$ (an error of 4%)

Check the equilibrium values:

\[
\frac{(7.1 \times 10^{-4})^2}{2.8 \times 10^{-2}} = 1.8 \times 10^{-5}
\]

This is equal to $K_a$ for aqueous ammonia.

Solutions for Practice Problems

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35. Problem

Use the table of $K_b$ values in Appendix E to list the conjugate bases of the following acids in order of increasing base strength: formic acid, HCOOH; hydrofluoric acid, HF$(aq)$; benzoic acid, $C_6H_5COOH$; phenol, $C_6H_5OH$.

What Is Required?
You need to determine the $K_b$ values for each conjugate base and interpret the values.

What Is Given?

<table>
<thead>
<tr>
<th>Acid Form</th>
<th>$K_b$ Value</th>
</tr>
</thead>
<tbody>
<tr>
<td>formic acid, HCOOH</td>
<td>$K_b = 1.8 \times 10^{-4}$</td>
</tr>
<tr>
<td>hydrofluoric acid, HF$(aq)$</td>
<td>$K_b = 6.3 \times 10^{-4}$</td>
</tr>
<tr>
<td>benzoic acid, $C_6H_5COOH$</td>
<td>$K_b = 6.3 \times 10^{-5}$</td>
</tr>
<tr>
<td>phenol, $C_6H_5OH$</td>
<td>$K_b = 1.0 \times 10^{-10}$</td>
</tr>
</tbody>
</table>
Plan Your Strategy
Calculate $K_b$ for each acid, using the equation $K_a K_b = K_w$.

Act on Your Strategy
$K_a K_b = K_w$
$K_b = \frac{K_w}{K_a}$

For formic acid, $K_b$ for $\text{HCOO}^- (aq) = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-4}} = 5.6 \times 10^{-11}$

Similarly, for the other conjugate bases:
$K_b$ for $\text{F}^- (aq) = 1.6 \times 10^{-11}$
$K_b$ for $\text{C}_6\text{H}_5\text{COO}^- = 1.6 \times 10^{-10}$
$K_b$ for $\text{C}_6\text{H}_5\text{O}^- = 1.0 \times 10^{-4}$

Check Your Solution
The order of increasing $K_b$ values is the order of increasing conjugate base strength. Therefore, $\text{F}^- < \text{HCOO}^- < \text{C}_6\text{H}_5\text{COO}^- < \text{C}_6\text{H}_5\text{O}^-$

36. Problem
$K_b$ for ammonia, $\text{NH}_3$, is $1.8 \times 10^{-5}$. $K_b$ for trimethylamine, $(\text{CH}_3)_3\text{N}$, is $6.5 \times 10^{-5}$. Which is the stronger acid, $\text{NH}_4^+$ or $(\text{CH}_3)_3\text{NH}^+$?

What Is Required?
You need to find the $K_a$ values for $\text{NH}_4^+$ and $(\text{CH}_3)_3\text{NH}^+$. Then determine which is the stronger acid.

What Is Given?
For $\text{NH}_3$, $K_b = 1.8 \times 10^{-5}$
For $(\text{CH}_3)_3\text{N}$, $K_b = 6.5 \times 10^{-5}$

Plan Your Strategy
Determine $K_a$ for each acid.

Act on Your Strategy
$K_a = \frac{K_w}{K_b}$
$K_a$ for $\text{NH}_4^+ (aq) = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$
$K_a$ for $(\text{CH}_3)_3\text{NH}^+ (aq) = \frac{1.0 \times 10^{-14}}{6.5 \times 10^{-5}} = 1.5 \times 10^{-10}$

Check Your Solution
$K_a$ for $\text{NH}_4^+ (aq)$ is greater than $K_a$ for $(\text{CH}_3)_3\text{NH}^+ (aq)$. Therefore, $\text{NH}_4^+ (aq)$ is the stronger acid.

37. Problem
Sodium benzoate is used as a food preservative. Calculate the pH of a 1.0 mol/L aqueous solution of sodium benzoate. (Only the benzoate ion affects the pH of the solution.)

What Is Required?
You need to calculate the pH.

What Is Given?
From Appendix E, $K_a$ for benzoic acid = $6.3 \times 10^{-5}$
$[\text{C}_6\text{H}_5\text{COO}^-] = 1.0 \text{ mol/L}$
Plan Your Strategy

**Step 1**  Find $K_b$ using the relationship $K_aK_b = K_w$.

**Step 2**  Determine whether or not the dissociation of the benzoate ions is negligible compared with its initial concentration.

**Step 3**  Write the chemical equation for the benzoate ion acting as a base. Then set up an ICE table.

**Step 4**  Write the dissociation equation for $K_b$ and substitute equilibrium values. Solve the equation for $x$.

**Step 5**  Calculate $[\text{H}_3\text{O}^+]$ using the following equation:

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

**Step 6**  Calculate pH from the following equation:

$$pH = -\log[\text{H}_3\text{O}^+]$$

Act on Your Strategy

**Step 1**  $K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.6 \times 10^{-5}}$

$K_b$ for $\text{C}_6\text{H}_5\text{COO}^- = 1.6 \times 10^{-10}$

**Step 2**  

$$\frac{[\text{C}_6\text{H}_5\text{COO}^-]}{K_b} = \frac{1.0}{1.6 \times 10^{-10}} = 6.2 \times 10^9$$

This is much greater than 500. Therefore, $x$ can be neglected compared to the initial concentration.

**Step 3**  Concentration (mol/L)  

<table>
<thead>
<tr>
<th></th>
<th>$\text{C}_6\text{H}_5\text{COO}^-(aq)$</th>
<th>$\text{H}_2\text{O}(l)$</th>
<th>$\text{C}_6\text{H}_5\text{COOH}(aq)$</th>
<th>$\text{OH}^-(aq)$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>1.0</td>
<td>0</td>
<td>-0</td>
<td></td>
</tr>
<tr>
<td>Change</td>
<td>$-x$</td>
<td>$+x$</td>
<td>$+x$</td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>$1.0 - x \approx 1.0$</td>
<td>$x$</td>
<td>$x$</td>
<td></td>
</tr>
</tbody>
</table>

**Step 4**  

$1.6 \times 10^{-10} = \frac{(x)(x)}{1.0}$

$x = \sqrt{1.6 \times 10^{-10}}$

$x = \pm 1.3 \times 10^{-5}$

Only the positive root is chemically possible.

Therefore, $[\text{OH}^-] = 1.3 \times 10^{-5}$ mol/L

**Step 5**  

$$[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14}}{1.3 \times 10^{-5}} = 7.7 \times 10^{-10} \text{ mol/L}$$

**Step 6**  

$pH = -\log 7.7 \times 10^{-10} = 9.11$

Check Your Solution

The solution is weakly basic. The benzoate ion is a very weak base, so the answer is reasonable.

38. Problem

The hydrogen sulfite ion, $\text{HSO}_3^-$, is amphoteric. Write chemical equations to show how it acts first as an acid and then as a base.

What Is Required?

One equation is required, which shows the hydrogen sulfite ion acting as an acid. A second equation is required, which shows the hydrogen sulfite ion acting as a base.
What Is Given?
The formula of the hydrogen sulfite ion is $\text{HSO}_3^-$.

Plan Your Strategy
Write out the chemical equations.

Act on Your Strategy
As an acid,
$\text{HSO}_3^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{SO}_3^{2-}(aq) + \text{H}_3\text{O}^+(aq)$

As a base,
$\text{HSO}_3^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{SO}_3(aq) + \text{OH}^-(aq)$

Check Your Solution
The acid-base conjugate pairs differ by one proton.