Chapter 17 Acid-Base Equilibrium Systems

Solutions to Practice Problems

1.

Problem

Name and write the formula of the conjugate base of each molecule or ion:

a) $HF(aq)$

b) $HCO_3^-(aq)$

c) $H_2SO_4(aq)$

d) $N_2H_5^+$ (aq)

Solution

The conjugate base for each molecule or ion will have one less proton, H^+ (aq), than its acid.

a) The conjugate base of $HF(aq)$ is the fluoride ion, $F^-(aq)$.

b) The conjugate base of $HCO_3^-(aq)$ is the carbonate ion, $CO_3^{2-}(aq)$.

c) The conjugate base of $H_2SO_4(aq)$ is the hydrogen sulfate ion, $HSO_4^-(aq)$.

d) The conjugate base of $N_2H_5^+(aq)$ is hydrazine, $N_2H_4(aq)$.

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) $NO_3^-(aq)$

b) OH^- (aq)

c) $H_2O(l)$

d) $HCO_3^-(aq)$

Solution

The conjugate acid for each molecule or ion will have one more proton, $H^+(aq)$, than its base.

a) The conjugate acid of $NO_3^-(aq)$ is nitric acid, $HNO_3(aq)$.

b) The conjugate acid of OH^- (aq) is water, $H_2O(l)$.

c) The conjugate acid of $H_2O(l)$ is hydronium ion, $H_3O^+(aq)$.

d) The conjugate acid of $HCO_3^-(aq)$ is carbonic acid, H_2CO_3 (aq).

Check Your Solution

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

3.

Problem

When perchloric acid dissolves in water, the following reaction occurs:

$$HClO_4(aq) + H_2O(l) \rightarrow H_3O^+(aq) + ClO_4^-(aq)$$

Identify the conjugate acid-base pairs.

What is Required?
You must identify the conjugate acid-base pairs.

What is Given?
The balanced chemical equation is given.

Plan Your Strategy
Identify the proton donor on the left side of the equation as the acid and the proton acceptor on the left side as the base. The conjugate acid and base on the right side of the equation will differ by a single proton from the acid and base on the left side.

Act on Your Strategy
The conjugate acid-base pairs are $HClO_4(aq)/ClO_4^-(aq)$ and $H_2O(l)/H_3O^+(aq)$.

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

4.
Problem
Identify the conjugate acid-base pairs in the following reactions:
a) $HS^-(aq) + H_2O(l) \leftrightarrow H_2S(aq) + OH^-(aq)$
b) $O^{2-}(aq) + H_2O(l) \leftrightarrow 2OH^-(aq)$
c) $H_2S(aq) + NH_3(aq) \leftrightarrow NH_4^+(aq) + HS^-(aq)$
d) $H_2SO_4(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + HSO_4^-(aq)$

What is Required?
You must identify the conjugate acid-base pairs in each reaction.

What is Given?
The balanced chemical equation is given.

Plan Your Strategy
Identify the proton donor on the left side of the equation as the acid and the proton acceptor on the left side as the base. The conjugate acid and base on the right side of the equation will differ by a single proton from the acid and base on the left side.

Act on Your Strategy
a) The conjugate acid-base pairs are $H_2O(l)/OH^-(aq)$ and $HS^-(aq)/H_2S(aq)$.
b) The conjugate acid-base pairs are $H_2O(l)/OH^-(aq)$ and $O^{2-}(aq)/OH^-(aq)$.
c) The conjugate acid-base pairs are $H_2S(aq)/HS^-(aq)$ and $NH_3(aq)/NH_4^+(aq)$.
d) The conjugate acid-base pairs are $H_2SO_4(aq)/HSO_4^-(aq)$ and $H_2O(l)/H_3O^+(aq)$.

Check Your Answer
Each acid-base conjugate pair differ by one proton. The acid has one more proton than the base.

5.

Problem

Hydrogen sulfide is a gas at room temperature. It has a very unpleasant smell and is not very soluble in water, forming a dilute solution of hydrosulfuric acid, $H_2S(aq)$.

a) Hydrosulfuric acid is diprotic. Write two chemical equations to show the ionization of $H_2S(aq)$ in water.

b) The hydrogen sulfide ion is amphiprotic. Write a chemical equation to show $HS^{-}(aq)$ acting as a base in water.

Solution

a) Ionization of an acid in aqueous solution occurs when the acid donates a proton, $H^{+}(aq)$, to a water molecule, $H_2O(l)$.

The first ionization is represented as:

$H_2S(aq) + H_2O(l) \leftrightarrow HS^{-}(aq) + H_3O^{+}(aq)$

The second ionization is represented as

$HS^{-}(aq) + H_2O(l) \leftrightarrow S^{2-}(aq) + H_3O^{+}(aq)$

b) To act as a base, $HS^{-}(aq)$ must accept a proton, $H^{+}(aq)$.

The equation representing $HS^{-}(aq)$ acting as a base is:

$HS^{-}(aq) + H_2O(l) \leftrightarrow H_2S(aq) + OH^{-}(aq)$

Check Your Solution

Each acid-base conjugate pair differ by one proton. The acid has one more proton than the base.

6.

Problem

a) Write a chemical equation to show the hydrogen carbonate ion, $HCO_3^{-}(aq)$, acting as an acid in the presence of $OH^{-}(aq)$.

b) Write a chemical equation to show $HCO_3^{-}(aq)$ acting as a base in the presence of $HF(aq)$.

What is Required?

a) You must show the hydrogen carbonate ion, $HCO_3^{-}(aq)$, acting as an acid in the presence of $OH^{-}(aq)$.

b) You must show $HCO_3^{-}(aq)$ acting as a base in the presence of $HF(aq)$.

What is Given?

The names and formulas of the starting acid-base combination are given.

Plan Your Strategy

a) For $HCO_3^{-}(aq)$ to act as an acid, it must donate a proton to $OH^{-}(aq)$.

b) For $HCO_3^{-}(aq)$ to act as a base, it must accept a proton from $HF(aq)$.

Act on Your Strategy

a) $HCO_3^{-}(aq) + OH^{-}(aq) \leftrightarrow CO_3^{2-}(aq) + H_2O(l)$

b) $HCO_3^{-}(aq) + HF(aq) \leftrightarrow H_2CO_3(aq) + F^{-}(aq)$

Check Your Solution

Each acid-base conjugate pair differ by one proton, $H^+(aq)$. The acid has one more proton than the base.

7.

Problem

Sodium dihydrogen phosphate, $NaH_2PO_4(s)$, is soluble in water. Write chemical equations to show that $H_2PO_4^-(aq)$ is amphiprotic. Label each equation as an acid or a base reaction, as appropriate.

What is Required?

You must show that $H_2PO_4^-(aq)$ is amphiprotic, that is, it can either accept or donate a proton.

What is Given?

The formula for the dihydrogen phosphate ion is $H_2PO_4^-(aq)$ and the acid-base reaction occurs in water, $H_2O(l)$.

Plan Your Strategy

When $H_2PO_4^-(aq)$ donates a proton, it acts as an acid. When $H_2PO_4^-(aq)$ accepts a proton, it acts as a base.

Act on Your Strategy

Donate a proton (acid): $H_2PO_4^-(aq) + H_2O(l) \leftrightarrow HPO_4^{2-}(aq) + H_3O^+(aq)$
Accept a proton (base): $H_2PO_4^-(aq) + H_2O(l) \leftrightarrow H_3PO_4(aq) + OH^-(aq)$

Check Your Solution

Each acid/base conjugate pair differ by one proton. The acid has one more proton than the base.

8.

Problem

Predict the direction for the following reactions. State whether reactants or products are favoured, and give reasons to support your decision:
a) $NH_4^+(aq) + H_2PO_4^-(aq) \leftrightarrow NH_3(aq) + H_3PO_4(aq)$
b) $H_2O(l) + HS^-(aq) \leftrightarrow OH^-(aq) + H_2S(aq)$
c) $HF(aq) + SO_4^{2-}(aq) \leftrightarrow F^-(aq) + HSO_4^-(aq)$

What is Required?

You must determine whether reactants or products are favoured for each reaction.

What is Given?

The chemical equations are given and Table 17.1 lists the relative strengths of the acids and bases.

Plan Your Strategy

Identify the acid on each side of the equation.
Refer to Table 17.1 and determine which acid is stronger. The reaction will proceed from the stronger acid to the weaker acid.

Act on Your Strategy

a) $NH_4^+(aq) + H_2PO_4^-(aq) \leftrightarrow NH_3(aq) + H_3PO_4(aq)$
acid 1 acid 2
weaker stronger

The reaction will go to the left and reactants are favoured.

b) $H_2O(\ell) + HS^-(aq) \leftrightarrow OH^-(aq) + H_2S(aq)$
acid 1 acid 2
weaker stronger

The reaction will go to the left and reactants are favoured.

c) $HF(aq) + SO_4^{2-}(aq) \leftrightarrow F^-(aq) + HSO_4^-(aq)$
acid 1 acid 2
weaker stronger

The reaction will go to the left and reactants are favoured.

Check Your Solution

In each case the reaction proceeds from the stronger acid to the weaker acid.

9.

Problem

In which direction will the following reactions proceed? In each case, explain your decision:

a) $HPO_4^{2-}(aq) + NH_4^+(aq) \leftrightarrow H_2PO_4^-(aq) + NH_3(aq)$
b) $H_2SO_4(aq) + H_2O(l) \rightarrow HSO_4^-(aq) + H_3O^+(aq)$
c) $H_2S(aq) + NH_3(aq) \leftrightarrow HS^-(aq) + NH_4^+(aq)$

What is Required?

You must determine whether reactants or products are favoured for each reaction.

What is Given?

The chemical equations are given and Table 17.1 lists the relative strengths of the acids and bases.

Plan Your Strategy

Identify the base on each side of the equation.
Refer to Table 17.1 and determine which base is stronger. The reaction will proceed from the stronger base to the weaker base.

Act on Your Strategy

a) $HPO_4^{2-}(aq) + NH_4^+(aq) \leftrightarrow H_2PO_4^-(aq) + NH_3(aq)$
base 1 base 2
weaker stronger

The reaction will go to the left and reactants are favoured.

b) $H_2SO_4(aq) + H_2O(l) \rightarrow HSO_4^-(aq) + H_3O^+(aq)$
base 1 base 2
stronger weaker

The reaction will go to the right and products are favoured.

c) $H_2S(aq) + NH_3(aq) \leftrightarrow HS^-(aq) + NH_4^+(aq)$

base 1	base 2
stronger	weaker

The reaction will go to the right and products are favoured.

Check Your Solution

In each case the reaction proceeds from the stronger base to the weaker base.

10.

Problem

Write equilibrium equations for each of the following reactions. State whether the reactants or the products are favoured. (Consult Table 17.1 or the table in the Appendix.)

a) aqueous carbonic acid is combined with ammonia

b) sodium hydrogen sulfite is dissolved in water (**Hint:** the sodium hydrogen sulfite dissociates completely into sodium ions and hydrogen sulfite ions. What happens to the hydrogen sulfite ion in water?)

c) hydrofluoric acid is mixed with potassium nitrate (consider the hint in part b.)

d) all possible reactions between water and phosphoric acid

What is Required?

You must write the equilibrium equations for the following reactions: aqueous carbonic acid combined with ammonia; sodium hydrogen sulfite dissolved in water; hydrofluoric acid mixed with potassium nitrate; and all possible reactions between water and phosphoric acid.

What is Given?

The reactions are given.

Plan Your Strategy

Write the balanced equations for each reaction.

Determine the relative strength of each acid from Table 17.1. Reactions will proceed from the strong acid toward the weak acid. Then determine if the reactants or products are favoured.

Act on Your Strategy

a) $H_2CO_3(aq) + NH_3(aq) \leftrightarrow NH_4^+(aq) + HCO_3^-(aq)$

From Table 17.3, we know that $H_2CO_3(aq)$ is a strong acid and $NH_4^+(aq)$ is a weak acid; therefore, the products are favoured.

b) $NaHSO_3(aq) + H_2O(l) \leftrightarrow Na^+(aq) + SO_3^-(aq) + H_3O^+(aq)$

From Table 17.3, we know that $NaHSO_3(aq)$ is a weak acid and $H_2SO_3(aq)$ is a strong acid; therefore, the reactants are favoured.

c) $HF(aq) + KNO_3(aq) \leftrightarrow K^+(aq) + HNO_3\ (aq) + F^-(aq)$

From Table 17.3, we know that $HNO_3(aq)$ is a strong acid and $HF(aq)$ is a weak acid; therefore, the reactants are favoured.

d) $H_3PO_4(aq) + H_2O(l) \leftrightarrow H_2PO_4^-(aq) + H_3O^+(l)$

$H_2PO_4^-(aq) + H_2O(l) \leftrightarrow HPO_4^{2-}(aq) + H_3O^+(l)$

$HPO_4^{2-}(aq) + H_2O(l) \leftrightarrow PO_4^{3-}(aq) + H_3O^+(l)$

From Table 17.3, we know that H_3PO_4(aq) is a stronger acid than $H_2PO_4^-$(aq) so the products are favoured. We also know that $H_2PO_4^-$(aq) is a stronger acid than HPO_4^{2-}(aq) so the products are favoured. PO_4^{3-}(aq) is a strong base than HPO_4^{2-}(aq) as an acid so the reactants are favoured.

Check Your Solution
The direction of reaction favoured for each equation is consistent with the reaction proceeding from the stronger acid toward the weaker acid or the stronger base toward the weaker base. The answers seem reasonable.

11. a) $K_a = \dfrac{[NO_2^-][H_3O^+]}{[HNO_2]} = 5.6 \times 10^{-4}$

b) $K_a = \dfrac{[C_6H_5COO^-][H_3O^+]}{[C_6H_5COOH]} = 6.3 \times 10^{-5}$

c) $K_a = \dfrac{[H_2C_6H_5O_7^-][H_3O^+]}{[H_3C_6H_5O_7]} = 7.4 \times 10^{-4}$

d) $K_a = \dfrac{[C_6H_5O_7^{3-}][H_3O^+]}{[HC_6H_5O_7^{2-}]} = 1.7 \times 10^{-5}$

12.
Problem
Carbonic acid, H_2CO_3(aq), is a weak diprotic acid. Write an acid ionization expression for each of the ionizations that take place in an aqueous solution of carbonic acid.

Solution
The ionization of an acid in aqueous solution occurs when the acid donates a proton, H^+(aq), to a water molecule, $H_2O(l)$.

The first ionization is represented as:
$H_2CO_3(aq) + H_2O(l) \leftrightarrow HCO_3^-(aq) + H_3O^+(aq)$

The second ionization is represented as:
$HCO_3^-(aq) + H_2O(l) \leftrightarrow CO_3^{2-}(aq) + H_3O^+(aq)$

Check Your Solution
Each acid-base conjugate pair differ by one proton. The acid has one more proton than the base.

13.
Problem
Phenol, C_6H_6O(aq), is a weak monoprotic acid used as a disinfectant. The K_a expression is $K_a = \dfrac{[C_6H_5O^-][H_3O^+]}{[C_6H_6O]}$. Write the ionization reaction for phenol in an aqueous solution.

What is Required?
You must write the ionization reaction for phenol in an aqueous solution.

What is Given?

The expression for the K_a for phenol is given.

Plan Your Strategy

The ionization constant is a special case of the equilibrium constant. The ionization of a weak acid can be represented as

$$HA(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + A^-(aq)$$

This is a heterogeneous equilibrium containing a pure liquid, water. Pure liquids (and pure solids) have constant concentrations and for dilute solutions, the concentration of water is almost constant. The concentration of water is included in the equilibrium constant and is not written in the equilibrium expression.

The acid ionization constant, K_a for this reaction is

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

Follow this pattern for the phenol solution.

Act on Your Strategy

The reactants are $C_6H_6O(aq)$ and $H_2O(l)$ and the products are $H_3O^+(aq)$ and $C_6H_5O^-(aq)$.

The ionization equation is

$C_6H_6O(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + C_6H_5O^-(aq)$

Check Your Solution

The products are in the numerator and the reactant (acid) is in the denominator.

14.

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 $HC_4H_3N_2O_3(s)$. A chemist prepares a 0.10 mol/L solution of barbituric acid. The chemist measured the pH of the solution and recorded the value as 2.50. What is the acid ionization constant for barbituric acid? What percentage of its molecules were ionized?

What is Required?

You need to find K_a and the percent ionization for barbituric acid.

What is Given?

The initial molar concentration: $[HC_4H_3N_2O_3(aq)] = 0.10$ mol/L

pH = 2.50

Plan Your Strategy

Step 1 Write the equation for the ionization of barbituric acid in water.

Step 2 Set up an ICE table for this ionization. Let x represent the change in the concentration of the acid. This will also be $[H_3O^+(aq)]$.

Step 3 Calculate $[H_3O^+(aq)]$ using $[H_3O^+(aq)] = 10^{-pH}$. Use this value of $[H_3O^+(aq)]$ and the stoichiometry of the equation to calculate the equilibrium concentrations of each species.
Step 4 Write the expression for the K_a for the acid. Substitute the equilibrium concentrations into the equation to determine the value of K_a.
Step 5 Express the concentration $[H_3O^+(aq)]$ as a fraction of the original acid concentration and calculate the percent ionization.

Act on Your Strategy
Step 1: $HC_4H_3N_2O_3(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + C_4H_3N_2O_3^-(aq)$

Step 2: ICE Table

Concentration (mol/L)	$HC_4H_3N_2O_3(aq)$ +	$H_2O(l)$ ↔	$H_3O^+(aq)$ +	$C_4H_3N_2O_3^-(aq)$
initial	0.10		≈ 0.00	0.00
change	− x		+ x	+ x
equilibrium	0.10 − x		x	x

Step 3: $[H_3O^+(aq)] = 10^{-pH} = 10^{-2.50}$

$x = [H_3O^+(aq)] = 3.16 \times 10^{-3}$ mol/L $= [C_4H_3N_2O_3^-(aq)]$
Therefore
$[HC_4H_3N_2O_3(aq)] = 0.10$ mol/L $- 3.16 \times 10^{-3}$ mol/L $= 9.69 \times 10^{-2}$ mol/L

$$K_a = \frac{[H_3O^+][C_4H_3N_2O_3^-]}{[HC_4H_3N_2O_3]} = \frac{(3.16\times10^{-3})^2}{9.69\times10^{-2}} = 1.0 \times 10^{-4}$$

$$\text{percent ionization} = \frac{[H_3O^+]}{[HC_4H_3N_2O_3]}\times100 = \frac{3.16\times10^{-4}}{0.10}\times100 = 3.2\%$$

Check Your Answer
The answer for the K_a and the percent ionization are reasonable for a weak acid. The answers have the correct number of significant digits (2).

15.
Problem
The word “butter” comes from the Greek butyros. Butanoic acid, $HC_4H_7O_2(aq)$, (common name butyric acid) gives rancid butter its distinctive odour. If the pH of a 1.00×10^{-2} mol/L solution of butanoic acid is 3.41, calculate the acid ionization constant for butyric acid. What percentage of butyric acid molecules in this solution is ionized?

What is Required?
You need to find K_a and the percent ionization for butanoic acid.

What is Given?

The initial molar concentration: $[HC_4H_7O_2(aq)] = 1.00 \times 10^{-2}$ mol/L
pH = 3.41

Plan Your Strategy

Step 1 Write the equation for the ionization of butanoic acid in water.
Step 2 Set up an ICE table for this ionization. Let x represent the change in the concentration of the acid. This will also be $[H_3O^+(aq)]$.
Step 3 Calculate $[H_3O^+(aq)]$ using $[H_3O^+(aq)] = 10^{-pH}$. Use this value of $[H_3O^+(aq)]$ and the stoichiometry of the equation to calculate the equilibrium concentrations of each species.
Step 4 Write the expression for the K_a for the acid. Substitute the equilibrium concentrations into the equation to determine the value of K_a.
Step 5 Express the concentration $[H_3O^+(aq)]$ as a fraction of the original acid concentration and calculate the percent ionization.

Act on Your Strategy

Step 1: $HC_4H_7O_2\,(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + C_4H_7O_2{}^-(aq)$

Step 2: ICE Table

concentration (mol/L)	$HC_4H_7O_2\,(aq)$ +	$H_2O(l) \leftrightarrow$	$H_3O^+(aq)$ +	$C_4H_7O_2{}^-(aq)$
initial	1.0×10^{-2}		≈ 0.00	0.00
change	$-$ x		+ x	+ x
equilibrium	$1.0 \times 10^{-2} - x$		x	x

Step 3: $[H_3O^+(aq)] = 10^{-pH} = 10^{-3.41}$

$x = [H_3O^+(aq)] = 3.89 \times 10^{-4}$ mol/L $= [C_4H_7O_2{}^-(aq)]$
Therefore,
$[HC_4H_7O_2(aq)] = 0.0100$ mol/L $- 3.89 \times 10^{-4}$ mol/L $= 9.61 \times 10^{-3}$ mol/L

$$K_a = \frac{[H_3O^+][C_4H_7O_2{}^-]}{[HC_4H_7O_2]} = \frac{(3.89\times10^{-4})^2}{9.61\times10^{-3}} = 1.6 \times 10^{-5}$$

$$\text{percent ionization} = \frac{[H_3O^+]}{[HC_4H_7O_2]}\times100 = \frac{3.89\times10^{-4}}{0.010}\times100 = 3.9\%$$

Check Your Answer

The answer for the K_a and the percent ionization are reasonable for a weak acid. The answers have the correct number of significant digits (2).

16.
Problem

Wild almonds taste bitter (and are dangerous to eat!) because they contain hydrocyanic acid, HCN(aq). When a chemist prepared a 0.75 mol/L solution of HCN(aq), the pH was found to be 4.67. What is the acid ionization constant, K_a?

What is Required?
You need to find K_a for hydrocyanic acid.

What is Given?
The initial molar concentration: [HCN(aq)] = 0.75 mol/L
pH = 4.67

Plan Your Strategy
Step 1 Write the equation for the ionization of hydrocyanic acid in water.
Step 2 Set up an ICE table for this ionization. Let x represent the change in the concentration of the acid. This will also be $[H_3O^+(aq)]$
Step 3 Calculate $[H_3O^+(aq)]$ using $[H_3O^+(aq)] = 10^{-pH}$. Use this value of $[H_3O^+(aq)]$ and the stoichiometry of the equation to calculate the equilibrium concentrations of each species.
Step 4 Write the expression for the K_a for the acid. Substitute the equilibrium concentrations into the equation to determine the value of K_a.

Act on Your Strategy
Step 1: $HCN(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + CN^-(aq)$

Step 2: ICE Table

concentration (mol/L)	$HCN(aq)$ +	$H_2O(l)$ ↔	$H_3O^+(aq)$ +	$CN^-(aq)$
initial	0.75		≈ 0.00	0.00
change	– x		+ x	+ x
equilibrium	0.75 – x		x	x

Step 3: $[H_3O^+(aq)] = 10^{-pH} = 10^{-4.67}$

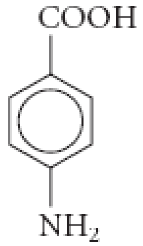
$x = [H_3O^+(aq)] = 2.14 \times 10^{-5}$ mol/L $= [CN^-(aq)]$
$[HCN(aq)] = 0.75$ mol/L $- 2.14 \times 10^{-5}$ mol/L $= 0.75$ mol/L

$$K_a = \frac{[H_3O^+][CN^-]}{[HCN]} = \frac{(2.14\times10^{-5})^2}{7.5\times10^{-1}} = 6.1 \times 10^{-10}$$

Check Your Answer
The answer for K_a is reasonable for a weak acid. The answer has the correct number of significant digits (2).

17.
Problem

Many sunscreen lotions contain salts of para-aminobenzoic acid (PABA). The structure of PABA is shown here:



A saturated solution of PABA was prepared by dissolving 4.7 g in 1.0 L solution. The pH of the solution was found to be 3.19. Calculate the acid ionization constant, K_a, for PABA.

What is Required?
You must find K_a for PABA, para-aminobenzoic acid, $C_7H_7NO_2(aq)$.

What is Given?
initial concentration $C_7H_7NO_2(aq) = 4.7$ g/L
pH = 3.41

Plan Your Strategy
Determine the molar mass of PABA and calculate the molar concentration of this compound.

$$nC_7H_7NO_2(aq) = \frac{m}{M}$$

$V = 1.0$ L

Step 1 Write the equation for the ionization of PABA in water.
Step 2 Set up an ICE table for this ionization. Let x represent the change in the concentration of the acid. This will also be $[H_3O^+(aq)]$
Step 3 Calculate $[H_3O^+(aq)]$ using $[H_3O^+(aq)] = 10^{-pH}$. Use this value of $[H_3O^+(aq)]$ and the stoichiometry of the equation to calculate the equilibrium concentrations of each species.
Step 4 Write the expression for the K_a for the acid. Substitute the equilibrium concentrations into the equation to determine the value of K_a.

Act on Your Strategy
$MC_7H_7NO_2(aq) = 137.15$ g/mol

$$nC_7H_7NO_2(aq) = \frac{4.7\,\text{g}}{137.15\,\text{g/mol}} = 0.0343 \text{ mol}$$

initial concentration $[C_7H_7NO_2(aq)] = 0.0343$ mol/L

Step 1: $C_7H_7O_2(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + C_7H_6O_2^-(aq)$

Step 2: ICE Table

Concentration (mol/L)	$C_7H_7O_2(aq)$	$+ H_2O(l) \leftrightarrow$	$H_3O^+(aq)$	$+ C_7H_6O_2^-(aq)$
initial	0.0343		≈ 0.00	0.00
change	$-x$		$+x$	$+x$
equilibrium	$0.0343 - x$		x	x

Step 3: $[H_3O^+(aq)] = 10^{-pH} = 10^{-3.19}$

$x = [H_3O^+(aq)] = 6.46 \times 10^{-4}$ mol/L $= [C_7H_6NO_2^-(aq)]$

Therefore

$[C_7H_7NO_2(aq)] = 0.0343$ mol/L $- 6.46 \times 10^{-4}$ mol/L $= 3.37 \times 10^{-2}$ mol/L

$$K_a = \frac{[H_3O^+][C_4H_7O_2^-]}{[HC_4H_7O_2]} = \frac{(6.46\times10^{-4})^2}{3.37\times10^{-2}} = 1.2 \times 10^{-5}$$

Check Your Answer

The answer for the K_a is reasonable for a weak acid. The answer has the correct number of significant digits (2).

18.

Problem

Aspirin™ (acetylsalicylic acid) is a monoprotic acid with molecular mass of 180. An aqueous solution containing 3.3 g/L was found to have pH = 2.62. What percentage of acetylsalicylic acid molecules were ionized in the solution?

What is Required?

You must find the percent ionization in an aqueous solution of ASA.

What is Given?

initial concentration ASA = 3.3 g/L
pH = 2.62
molar mass = 180 g/mol

Plan Your Strategy

Determine the molar concentration of ASA.

$$nC_7H_7NO_2(aq) = \frac{m}{M}$$

$V = 1.0$ L

Calculate $[H_3O^+(aq)]$ using $[H_3O^+(aq)] = 10^{-pH}$
Express the concentration $[H_3O^+(aq)]$ as a fraction the original acid concentration and calculate the percent ionization.

Act on Your Strategy

$$nC_7H_7NO_2(aq) = \frac{3.3\,\text{g}}{180\,\text{g/mol}} = 0.0183 \text{ mol}$$

initial concentration [ASA] = 0.0183 mol/L

$[H_3O^+(aq)] = 10^{-pH} = 10^{-2.62}$

$x = [H_3O^+(aq)] = 2.40 \times 10^{-3}$ mol/L

$$\text{percent ionization} = \frac{2.40\times10^{-3}}{0.0183} \times 100 = 13\%$$

Check Your Answer
The answer for percent ionization is reasonable for a weak acid. The answer has the correct number of significant digits (2).

19.
Problem
Write the chemical equation for each base ionizing in an aqueous solution:
a) ammonia, $NH_3(aq)$
b) trimethylamine, $(CH_3)_3N(aq)$
c) hydrogen sulfite ion, $HSO_3^-(aq)$
d) carbonate ion, $CO_3^{2-}(aq)$

Solution
A weak base, B, will react with water to form an equilibrium solution of ions.

$$B(aq) + H_2O(l) \leftrightarrow HB^+(aq) + OH^-(aq)$$

a) $NH_3(aq) + H_2O(l) \leftrightarrow NH_4^+(aq) + OH^-(aq)$
b) $(CH_3)_3N(aq) + H_2O(l) \leftrightarrow (CH_3)_3NH^+(aq) + OH^-(aq)$
c) $HSO_3^-(aq) + H_2O(l) \leftrightarrow H_2SO_3(aq) + OH^-(aq)$
d) $CO_3^{2-}(aq) + H_2O(l) \leftrightarrow HCO_3^-(aq) + OH^-(aq)$

20.
Problem
Write the K_b expression for each base in the previous question.

Solution

a) $K_b = \dfrac{[NH_4^{+}][OH^{-}]}{[NH_3]}$

b) $K_b = \dfrac{[(CH_3)_3NH^{+}][OH^{-}]}{[(CH_3)_3N]}$

c) $K_b = \dfrac{[H_2SO_3][OH^{-}]}{[HSO_3^{-}]}$

d) $K_b = \dfrac{[HCO_3^{-}][OH^{-}]}{[CO_3^{2-}]}$

Check Your Solution

The equilibrium constant, K_b, has the form:

$$K_b = \frac{[HB^{+}][OH^{-}]}{[B]}$$

21.

Problem

When a 0.25 mol/L aqueous solution of methylamine was prepared, the pH of the solution was found to be 10.04. What percentage of methylamine molecules ionized in the solution?

What is Required?

You must find the percentage of methylamine molecules ionized in the solution.

What is Given?

initial concentration methylamine = 0.25 mol/L
pH = 10.04

Plan Your Strategy

Calculate the pOH using the expression pH + pOH = 14.00
Calculate $[OH^{-}(aq)]$ using $[OH^{-}(aq)] = 10^{-pOH}$
Express the concentration $[OH^{-}(aq)]$ as a fraction the original base concentration and calculate the percent ionization.

Act on Your Strategy

pH + pOH = 14.00

pOH = 14 – 10.04 = 3.96

$[OH^{-}(aq)] = 10^{-pOH} = 10^{-3.96}$

$[OH^-(aq)] = 1.10 \times 10^{-4}$ mol/L

$$\text{Percent ionization} = \frac{1.10\times10^{-4}}{0.25}\times100 = 0.044\ \%$$

Check Your Answer
The answer for percent ionization is reasonable for a weak base. The answer has the correct number of significant digits (2).

22.
Problem
Codeine, $C_{18}H_{21}NO_3(s)$, is added to some cough medicines. When a 0.020 mol/L aqueous solution of codeine was prepared, the pH of the solution was found to be 10.26. Calculate K_b for codeine.

What is Required?
You must find K_b for codeine.

What is Given?
The initial molar concentration: $[C_{18}H_{21}NO_3\ (aq)] = 0.020$ mol/L
pH = 10.26

Plan Your Strategy
Step 1 Write the equation for the ionization of codeine in water.
Step 2 Set up an ICE table for this ionization. Let x represent the change in the concentration of the base. This will also be $[OH^-(aq)]$.
Step 3 Calculate the pOH using the expression pH + pOH = 14.00
Calculate $[OH^-(aq)]$ using $[OH^-(aq)] = 10^{-pOH}$. Use $[OH^-(aq)]$ and the stoichiometry of the equation to calculate the equilibrium concentrations of each species.
Step 4 Write the expression for the K_b for the base. Substitute the equilibrium concentrations into the equation to determine the value of K_b.

Act on Your Strategy
Step 1: $C_{18}H_{21}NO_3\ (aq) + H_2O(l) \leftrightarrow C_{18}H_{22}NO_3^+(aq) + OH^-(aq)$

Step 2: ICE Table
Concentration (mol/L) $C_{18}H_{21}NO_3\ (aq) + H_2O(l) \leftrightarrow\ \ OH^-(aq) + C_{18}H_{22}NO_3^+(aq)$

initial	0.020		≈ 0.00	0.00
change	– x		+ x	+ x
equilibrium	0.020 – x		x	x

Step 3: pH + pOH = 14.00

pOH = 14.00 – 10.26 = 3.74
$[OH^-(aq)] = 10^{-pOH} = 10^{-3.74}$
$[OH^-(aq)] = 1.82 \times 10^{-4}$ mol/L = $[C_{18}H_{22}NO_3^+(aq)]$
$[C_{18}H_{21}NO_3\ (aq)]$ = 0.020 mol/L – 1.82×10^{-4} mol/L = 1.98×10^{-2} mol/L

Step 4: $K_b = \frac{[C_{18}H_{22}NO_3^+][OH^-]}{[C_{18}H_{21}NO_3]} = \frac{(1.82\times10^{-4})^2}{1.98\times10^{-2}} = 1.7 \times 10^{-6}$

Check Your Answer
The answer for the K_b is reasonable for a weak base. The answer has the correct number of significant digits (2).

23.
Problem
A Material Safety Data Sheet (MSDS) describes pyridine,$C_5H_5N(l)$, as a clear liquid with a putrid odour. A 16 g/L solution of pyridine has pH = 9.23. Use these data to calculate K_b for pyridine.

What is Required?
You must find K_b for pyridine, $C_5H_5N(l)$.

What is Given?
initial concentration $C_5H_5N(l)$ =16 g/L
pH = 9.23

Plan Your Strategy
Determine the molar mass of pyridine, $C_5H_5N(l)$ and calculate the molar concentration of this compound.

$$nC_5H_5N(l) = \frac{m}{M}$$

$V = 1.0$ L

Step 1 Write the equation for the ionization of $C_5H_5N(l)$ in water.
Step 2 Set up an ICE table for this ionization. Let x represent the change in the concentration of the acid. This will also be $[OH^-(aq)]$
Step 3 Calculate $[OH^-(aq)]$ using $[OH^-(aq)] = 10^{-pOH}$. Use this value of $[OH^-(aq)]$ and the stoichiometry of the equation to calculate the equilibrium concentrations of each species.
Step 4 Write the K_b expression for the base. Substitute the equilibrium concentrations into the equation to determine the value of K_b.

Act on Your Strategy
$MC_5H_5N(l)$ = 79.11 g/mol

$$nC_5H_5N(l) \;=\; \frac{16\,\text{g}}{79.11\,\text{g/mol}} = 0.202 \text{ mol}$$

initial concentration $[C_5H_5N(l)] = 0.202$ mol/L

Step 1: $C_5H_5N(l) + H_2O(l) \leftrightarrow C_5H_6N^+(aq) + OH^-(aq)$
Step 2: ICE Table
Concentration (mol/L) $C_5H_5N(l)$ + $H_2O(l)$ $\leftrightarrow$ $C_5H_6N^+(aq)$ + $OH^-(aq)$

initial	0.202		0.00	≈ 0.00
change	– x		+ x	+ x
equilibrium	0.202 – x		x	x

Step 3: Step 3: pH + pOH = 14.00
pOH = 14.00 – 9.23 = 4.77
$[OH^-(aq)] = 10^{-pOH} = 10^{-4.77}$
$[OH^-(aq)] = 1.70 \times 10^{-5}$ mol/L $= [C_5H_6N^+(aq)]$
$[C_5H_5N(aq)] = 0.202$ mol/L $- 1.70 \times 10^{-5}$ mol/L $= 0.202$ mol/L

Step 4: $K_b = \dfrac{[C_5H_6N^+][OH^-]}{[C_5H_5N]} = \dfrac{(1.70\times10^{-5})^2}{0.202} = 1.4 \times 10^{-9}$

Check Your Answer
The answer for the K_b is reasonable for a weak base. The answer has the correct number of significant digits (2).

24.
Problem
Calculate the pH of a sample of vinegar that contains 0.83 mol/L ethanoic acid. What is the percent ionization of the ethanoic acid?

What is Required?
You must calculate the pH and the percent ionization of ethanoic acid in a vinegar solution.

What is Given?
The ethanoic acid concentration $[CH_3COOH(aq)] = 0.83$ mol/L
K_a for ethanoic acid $= 1.8 \times 10^{-5}$ (from Tables)

Plan Your Strategy
Step 1 Write the equation for the ionization of ethanoic acid in water.
Step 2 Set up an ICE table for this ionization. Let x represent the change in the concentration of the acid. This will also be $[H_3O^+(aq)]$

Step 3 Calculate the ratio of acid to K_a. If this ratio is greater than 1000, the equilibrium concentration of acid is essentially the same as the initial acid concentration.
Step 4 Write the expression for K_a of the acid. Substitute the equilibrium terms into the equation and solve for $[H_3O^+(aq)]$.
Step 5 Solve for pH using the expression $pH = -\log [H_3O^+(aq)]$
Step 6 Express the concentration $[H_3O^+(aq)]$ as a fraction of the original acid concentration and calculate the percent ionization.

Act on Your Strategy
Step 1: $CH_3COOH(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + CH_3COO^-(aq)$

Step 2: ICE Table
Concentration (mol/L) $CH_3COOH(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + CH_3COO^-(aq)$

initial	0.83		≈ 0.00	0.00
change	− x		+ x	+ x
equilibrium	0.83 − x		x	x

Step 3: $\dfrac{CH_3COOH}{K_a} = \dfrac{0.83}{1.8 \times 10^{-5}} >> 1000$

Therefore, at equilibrium, $[CH_3COOH(aq)] \approx 0.83$ mol/L

Step 4: $K_a = \dfrac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} = 1.8 \times 10^{-5}$

$$\frac{x^2}{0.83} = 1.8 \times 10^{-5}$$

$x = 3.87 \times 10^{-3}$

$[H_3O^+(aq)] = 3.87 \times 10^{-3}$ mol/L

Step 5: $pH = -\log [H_3O^+(aq)] = -\log 3.87 \times 10^{-3} = 2.41$

Step 6: percent ionization $= \dfrac{[H_3O^+]}{[CH_3COOH]} \times 100 = \dfrac{3.87 \times 10^{-3}}{0.83} \times 100 = 0.47\ \%$

Check Your Answer
The answer for the pH and the percent ionization are reasonable for a weak acid. The answers have the correct number of significant digits (2).

25.

Problem

A solution of hydrofluoric acid has a molar concentration 1.00 mol/L. What is the pH of the solution? What percentage of its molecules ionize?

What is Required?

You must calculate the pH and the percent ionization of a solution of hydrofluoric acid.

What is Given?

The hydrofluoric acid concentration [HF(aq)] = 1.00 mol/L
K_a for HF = 6.3×10^{-4} (from Tables)

Plan Your Strategy

Step 1 Write the equation for the ionization of HF in water.
Step 2 Set up an ICE table for this ionization. Let x represent the change in the concentration of the acid. This will also be $[H_3O^+(aq)]$
Step 3 Calculate the ratio of acid to K_a. If this ratio is greater than 1000, the equilibrium concentration of acid is essentially the same as the initial acid concentration.
Step 4 Write the expression for K_a of the acid. Substitute the equilibrium terms into the equation and solve for $[H_3O^+(aq)]$.
Step 5 Solve for pH using the expression pH = $-\log [H_3O^+(aq)]$
Step 6 Express the concentration $[H_3O^+(aq)]$ as a fraction of the original acid concentration and calculate the percent ionization.

Act on Your Strategy

Step 1: $HF(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + F^-(aq)$

Step 2: ICE Table

Concentration (mol/L)	$HF(aq)$	$+ H_2O(l) \leftrightarrow$	$H_3O^+(aq)$ +	$F^-(aq)$
initial	1.00		≈ 0.00	0.00
change	– x		+ x	+ x
equilibrium	1.00 – x		x	x

Step 3: $\dfrac{HF}{K_a} = \dfrac{1.00}{6.3 \times 10^{-4}} > 1000$

Therefore, at equilibrium, [HF (aq)] ≈ 1.00 mol/L

Step 4: $K_a = \dfrac{[H_3O^+][F^-]}{[HF]} = 6.3 \times 10^{-4}$

$$\frac{x^2}{1.00} = 6.3 \times 10^{-4}$$

$x = 2.51 \times 10^{-2}$

$[H_3O^+(aq)] = 2.51 \times 10^{-2}$ mol/L

Step 5: $pH = -\log [H_3O^+(aq)] = -\log 2.51 \times 10^{-2} = 1.60$

Step 6: percent ionization $= \dfrac{[H_3O^+]}{[HF]} \times 100 = \dfrac{2.51\times 10^{-2}}{1.00} \times 100 = 2.5\%$

Check Your Answer
The answer for the pH and the percent ionization are reasonable for a weak acid. The answers have the correct number of significant digits (2).

26.
Problem
Hypochlorous acid, HOCl(aq), is used to create bleach. A chemist finds that 0.027% of hypochlorous acid molecules are ionized in a 0.40 mol/L solution. What is the value of K_a for the acid?

What is Required?
You must find K_a for hypochlorous acid, HOCl(aq).

What is Given?
The initial molar concentration: [HOCl(aq)] = 0.40 mol/L
The acid is 0.027% ionized.

Plan Your Strategy
Step 1 Write the equation for the ionization of HOCl(aq) water.
Step 2 Calculate $[H_3O^+(aq)]$ from the percent ionization.
Step 3 Set up an ICE table for this ionization. Use the value of $[H_3O^+(aq)]$ and the stoichiometry of the equation to calculate the equilibrium concentrations of each species.
Step 4 Write the expression for the K_a for the acid. Substitute the equilibrium concentrations into the equation to determine the value of K_a.

Act on Your Strategy
Step 1: $HOCl(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + OCl^-(aq)$

Step 2: $[H_3O^+(aq)] = \dfrac{0.027}{100} \times 0.40$ mol/L $= 1.08 \times 10^{-4}$ mol/L

Step 3: ICE Table

Concentration (mol/L)	$HOCl(aq)$	+ $H_2O(l)$ ↔	$H_3O^+(aq)$ +	$OCl^-(aq)$
initial	0.40		≈ 0.00	0.00
change	-1.08×10^{-4}		$+1.08 \times 10^{-4}$	$+1.08 \times 10^{-4}$
equilibrium	$0.40 - 1.08 \times 10^{-4}$ ≈ 0.40		1.08×10^{-4}	1.08×10^{-4}

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

Check Your Answer

The answer for the K_a is reasonable for a weak acid and it has the correct number of significant digits (2).

27.

Problem

Hexanoic acid, commonly known as caproic acid, $C_5H_{11}COOH(s)$, 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 hexanoic acid has pH of 2.94. How much acid was dissolved to make 100 mL of this solution?

What is Required?

You must find the mass of hexanoic acid that was dissolved to make 100 mL solution?

What is Given?

K_a for hexanoic acid = 1.3×10^{-5}.
pH of the acid solution = 2.94
The chemical formula for hexanoic acid is $C_5H_{11}COOH(s)$.

Plan Your Strategy

Step 1 Write the equation for the ionization of $C_5H_{11}COOH(s)$ in water.
Step 2 Calculate $[H_3O^+(aq)]$ from the pH:
$pH = -\log[H_3O^+(aq)]$
Step 3 Set up an ICE table for this ionization. Let x = initial acid concentration. Use the value of $[H_3O^+(aq)]$ and the stoichiometry of the equation to calculate the equilibrium concentrations of each species.
Step 4 Write the expression for the K_a of the acid. Substitute the equilibrium concentrations and the K_a into this equation to calculate the initial acid concentration.
Step 5 Determine the molar mass of $C_5H_{11}COOH$ and the mass of $C_5H_{11}COOH(aq)$ in 1.0 L of solution.
Step 6 Calculate the mass of $C_5H_{11}COOH$ in 100 mL of solution.

Act on Your Strategy

Step 1: $C_5H_{11}COOH(s) + H_2O(l) \leftrightarrow H_3O^+(aq) + C_5H_{11}COO^-(aq)$
Step 2: $pH = -\log[H_3O^+(aq)] = 2.94$

$[H_3O^+(aq)] = 10^{-2.94} = 1.15 \times 10^{-3}$ mol/L

Step 3: ICE Table

Concentration (mol/L) $C_5H_{11}COOH(s) + H_2O(l) \leftrightarrow H_3O^+(aq) + C_5H_{11}COO^-(aq)$

initial	x		≈ 0.00	0.00
change	-1.15×10^{-3}		$+1.15 \times 10^{-3}$	$+1.15 \times 10^{-3}$
equilibrium	$x - 1.15 \times 10^{-3}$		1.15×10^{-3}	1.15×10^{-3}

Step 4: $K_a = \dfrac{[H_3O^+][C_5H_{11}COO^-]}{[C_5H_{11}COOH]}$

$$1.3 \times 10^{-5} = \frac{(1.15 \times 10^{-3})^2}{(x - 1.15 \times 10^{-3})}$$

$$1.3 \times 10^{-5}\,x - (1.3 \times 10^{-5})(1.15 \times 10^{-3}) = (1.15 \times 10^{-3})^2$$

$x = 0.101$

$[C_5H_{11}COOH(aq)] = 0.101$ mol/L

Step 5: $MC_5H_{11}COOH = 116.18$ g/mol

$mC_5H_{11}COOH(s)$ in 1.0 L $= n \times M = 0.101 \text{ mol} \times 116.18 \text{ g/mol} = 11.7$ g

Step 6: Therefore in 100 mL of solution, the mass of $C_5H_{11}COOH(s) = 1.17$ g or 1.2 g

Check Your Answer

The mass of acid is small which is reasonable for a weak acid and it has the correct number of significant digits (2).

28.

Problem

An aqueous solution of household ammonia, $NH_3(aq)$, has a molar concentration of 0.105 mol/L. Calculate the pH of the solution.

What is Required?

You must calculate the pH of a household ammonia, $NH_3(aq)$ solution.

What is Given?

The initial concentration of ammonia, $[NH_3(aq)] = 0.105$ mol/L

K_b for $NH_3(aq) = 1.8 \times 10^{-5}$ (from Tables)

Plan Your Strategy

Step 1 Write the equation for the ionization of $NH_3(aq)$ in water.
Step 2 Set up an ICE table for this ionization. Let x represent the change in the concentration of the $NH_3(aq)$. This will also be $[OH^-(aq)]$. Use the value of $[OH^-(aq)]$ and the stoichiometry of the equation to determine the equilibrium concentrations of each species.
Step 3 Calculate the ratio of base to K_b. If this ratio is greater than 1000, the equilibrium concentration of base is essentially the same as the initial base concentration.
Step 4 Write the expression for K_b of the base. Substitute the equilibrium terms into the equation and solve for $[OH^-(aq)]$.
Step 5 Solve for pOH using the expression $pOH = -\log [OH^-(aq)]$
Step 6 Solve for the pH using the relationship: $pH + pOH = 14.00$

Act on Your Strategy
Step 1: $NH_3(aq) + H_2O(l) \leftrightarrow NH_4^+(aq) + OH^-(aq)$

Step 2: ICE Table

Concentration (mol/L)	$NH_3(aq)$	$+ H_2O(l) \leftrightarrow$	$NH_4^+(aq)$ +	$OH^-(aq)$
initial	0.105		0.00	≈ 0.00
change	– x		+ x	+ x
equilibrium	0.105 – x		x	x

Step 3: $\frac{NH_3}{K_b} = \frac{0.105}{1.8 \times 10^{-5}} >> 1000$

Therefore, at equilibrium, $[NH_3(aq)] \approx 0.105$ mol/L

Step 4: $K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = 1.8 \times 10^{-5}$

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

$x = 1.37 \times 10^{-3}$

$[OH^-(aq)] = 1.37 \times 10^{-3}$ mol/L

Step 5: $pOH = -\log [OH^-(aq)] = -\log 1.37 \times 10^{-3} = 2.86$

Step 6: $pH = 14.00 - pOH = 14.00 - 2.86 = 11.14$

Check Your Answer
The answer for the pH is reasonable for a weak acid. The answer has the correct number of significant digits (2).

29.
Problem
Hydrazine, $N_2H_4(l)$, has been used as a rocket fuel. The concentration of an aqueous solution of hydrazine is 5.9×10^{-2} mol/L. Calculate the pH of the solution.

What is Required?
You must calculate the pH of a hydrazine, $N_2H_4(l)$ solution.

What is Given?
The initial concentration of hydrazine, $[N_2H_4(aq)] = 5.9 \times 10^{-2}$ mol/L
K_b for $N_2H_4(l) = 1.7 \times 10^{-6}$ (from Tables)

Plan Your Strategy
Step 1 Write the equation for the ionization of $N_2H_4(l)$ in water.
Step 2 Set up an ICE table for this ionization. Let x represent the change in the concentration of the $N_2H_4(l)$. This will also be $[OH^-(aq)]$. Use the value of $[OH^-(aq)]$ and the stoichiometry of the equation to determine the equilibrium concentrations of each species.
Step 3 Calculate the ratio of base to K_b. If this ratio is greater than 1000, the equilibrium concentration of base is essentially the same as the initial base concentration.
Step 4 Write the expression for K_b of the base. Substitute the equilibrium terms into the equation and solve for $[OH^-(aq)]$.
Step 5 Solve for pOH using the expression $pOH = -\log [OH^-(aq)]$
Step 6 Solve for the pH using the relationship $pH + pOH = 14.00$

Act on Your Strategy
Step 1: $N_2H_4(l) + H_2O(l) \leftrightarrow N_2H_5^+(aq) + OH^-(aq)$

Step 2: ICE Table

Concentration (mol/L)	$N_2H_4(l)$ +	$H_2O(l)$ $\leftrightarrow$	$N_2H_5^+(aq)$ +	$OH^-(aq)$
initial	5.9×10^{-2}		0.00	≈ 0.00
change	$-$ x		+ x	+ x
equilibrium	$5.9 \times 10^{-2} - x$		x	x

Step 3: $\dfrac{[N_2H_4]}{K_b} = \dfrac{5.9\times10^{-2}}{1.7\times10^{-6}} >> 1000$

Therefore, at equilibrium, $[N_2H_4(aq)] \approx 5.9 \times 10^{-2}$ mol/L

Step 4: $K_b = \dfrac{[N_2H_5^{+}][OH^{-}]}{[N_2H_4]} = 1.7 \times 10^{-6}$

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

$x = 3.17 \times 10^{-4}$

$[OH^{-}(aq)] = 3.17 \times 10^{-4}$ mol/L

Step 5: pOH = –log $[OH^{-}(aq)]$ = – log 3.17×10^{-4} = 3.50

Step 6: pH = 14.00 – pOH = 14.00 – 3.50 = 10.50

Check Your Answer
The answer for the pH is reasonable for a weak base. The answer has the correct number of significant digits (2).

30.
Problem
Morphine, $C_{17}H_{19}NO_3(s)$, is a naturally occurring base that is used to control pain. A 4.5×10^{-3} mol/L solution has pH of 9.93. Calculate K_b for morphine.

What is Required?
You must calculate the K_b of morphine, $C_{17}H_{19}NO_3(s)$.

What is Given?
The initial concentration of morphine $[C_{17}H_{19}NO_3(aq)] = 4.5 \times 10^{-3}$ mol/L
pH of the solution = 9.93

Plan Your Strategy
Step 1 Write the equation for the ionization of $C_{17}H_{19}NO_3(s)$ in water.
Step 2 Use the expression: pH + pOH = 14.00 to determine the pOH of the solution.
Step 3 Calculate the $[OH^{-}(aq)]$ using the relationship: $[OH^{-}(aq)] = 10^{-pOH}$
Step 4 Set up an ICE table for this ionization. Use the value of $[OH^{-}(aq)]$ and the stoichiometry of the equation to determine the equilibrium concentrations of each species.
Step 5 Write the expression for K_b of the base. Substitute the equilibrium terms into the equation and solve for K_b.

Act on Your Strategy
Let the base morphine be represented by M.
Step 1: $M(aq) + H_2O(l) \leftrightarrow MH^{+}(aq) + OH^{-}(aq)$
Step 2: pOH = 14.00 – pH = 14.00 – 9.93 = 4.07
Step 3: $[OH^{-}(aq)] = 10^{-pOH} = 8.51 \times 10^{-5}$

Step 4: ICE Table

Concentration (mol/L) $M(aq) + H_2O(l) \leftrightarrow MH^+(aq) + OH^-(aq)$

initial	4.5×10^{-3}		0.00	≈ 0.00
change	-8.51×10^{-5}		$+8.51 \times 10^{-5}$	$+8.51 \times 10^{-5}$
equilibrium	$4.5 \times 10^{-3} - 8.51 \times 10^{-5}$ $= 4.41 \times 10^{-3}$		8.51×10^{-5}	8.51×10^{-5}

Step 5: $K_b = \dfrac{[MH^+][OH^-]}{[M]} = \dfrac{(8.51 \times 10^{-5})^2}{4.41 \times 10^{-3}} = 1.6 \times 10^{-6}$

Check Your Answer

The answer for the K_b is reasonable for a weak base. The answer has the correct number of significant digits (2).

31.

Problem

Methylamine, $CH_3NH_2(g)$, is a fishy-smelling gas at room temperature. It is used in the manufacture of several prescription drugs, including methamphetamine. Calculate $[OH^-(aq)]$ and pOH of a 1.5 mol/L aqueous solution of methylamine.

What is Required?

You must calculate $[OH^-(aq)]$ and pOH of a 1.5 mol/L aqueous solution of methylamine.

What is Given?

The initial concentration $[CH_3NH_2(g)] = 1.5$ mol/L

K_b for methylamine, $CH_3NH_2(g)$, $= 4.4 \times 10^{-4}$ (from Tables)

Plan Your Strategy

Step 1 Write the equation for the ionization of $CH_3NH_2(aq)$ in water.

Step 2 Set up an ICE table for this ionization. Let x represent the change in the concentration of the methylamine, $CH_3NH_2(g)$. This will also be $[OH^-(aq)]$. Use the value of $[OH^-(aq)]$ and the stoichiometry of the equation to determine the equilibrium concentrations of each species.

Step 3 Calculate the ratio of base to K_b. If this ratio is greater than 1000, the equilibrium concentration of base is essentially the same as the initial base concentration.

Step 4 Write the expression for K_b of the base. Substitute the equilibrium terms into the equation and solve for $[OH^-(aq)]$.

Step 5 Solve for pOH using the expression $pOH = -\log [OH^-(aq)]$

Act on Your Strategy

Step 1: $CH_3NH_2(g) + H_2O(l) \leftrightarrow OH^-(aq) + CH_3NH_3^+(aq)$

Step 2: ICE Table

Concentration (mol/L) $CH_3NH_2(g) + H_2O(l) \leftrightarrow OH^-(aq) + CH_3NH_3^+(aq)$

initial	1.5		≈ 0.00	0.00
change	– x		+ x	+ x
equilibrium	1.5 – x		x	x

Step 3: $\frac{[CH_3NH_2]}{K_b} = \frac{1.5}{4.4\times10^{-4}} >> 1000$

Therefore, at equilibrium, $[N_2H_4(aq)] = 1.5$ mol/L – x

Step 4: $K_b = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]} = 4.4\times10^{-4}$

$$\frac{x^2}{1.5-x} = 4.4\times10^{-4}$$

$$x^2 + 4.4\times10^{-4}x - 6.6\times10^{-4} = 0$$

Solve for x using the standard formula for a quadratic equation $ax^2 + bx + c = 0$

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

$$x = \frac{-4.4\times10^{-4} \pm \sqrt{(4.4\times10^{-4})^2 - 4(1)(-6.6\times10^{-4})}}{2}$$

$x = 2.547\times10^{-2}$ or to 2 significant digits 2.5×10^{-2} mol/L

Step 5: $pOH = -\log[OH^-(aq)] = -\log 1.03\times10^{-2} = 1.98$

Check Your Solution

The answer for the $[OH^-(aq)]$ and pH is reasonable for a weak acid. The answer has the correct number of significant digits (2).

32.

Problem

At room temperature, trimethylamine, $(CH_3)_3N(g)$, is a gas with a strong ammonia-like odour. Calculate $[OH^-(aq)]$ and the percentage of trimethylamine molecules that react with water in a 0.22 mol/L aqueous solution.

What is Required?
You must calculate $[OH^-(aq)]$ and the percentage of trimethylamine molecules, $(CH_3)_3N(g)$, in a 0.22 mol/L aqueous solution.

What is Given?
The initial concentration of trimethylamine, $[(CH_3)_3N(g)] = 0.22$ mol/L
K_b for $(CH_3)_3N(g) = 6.5 \times 10^{-5}$ (from Tables)

Plan Your Strategy
Step 1 Write the equation for the ionization of $(CH_3)_3N(aq)$ in water.
Step 2 Set up an ICE table for this ionization. Let x represent the change in the concentration of the $(CH_3)_3N(g)$. This will also be $[OH^-(aq)]$. Use the value of $[OH^-(aq)]$ and the stoichiometry of the equation to determine the equilibrium concentrations of each species.
Step 3 Calculate the ratio of base to K_b. If this ratio is greater than 1000, the equilibrium concentration of base is essentially the same as the initial base concentration.
Step 4 Write the expression for K_b of the base. Substitute the equilibrium terms into the equation and solve for $[OH^-(aq)]$.
Step 5 Express the concentration $[OH^-(aq)]$ as a fraction of the original base concentration and calculate the percent ionization.

Act on Your Strategy
Step 1: $(CH_3)_3N(aq) + H_2O(l) \leftrightarrow OH^-(aq) + (CH_3)_3NH^+(g)$

Step 2: ICE Table
Concentration (mol/L) $(CH_3)_3N(aq)$ + $H_2O(l)$ $\leftrightarrow$ $OH^-(aq)$ + $(CH_3)_3NH^+(g)$

initial	0.22		≈ 0.00	0.00
change	− x		+ x	+ x
equilibrium	0.22 − x		x	x

Step 3: $\dfrac{[(CH_3)_3N]}{K_b} = \dfrac{0.22}{6.5 \times 10^{-5}} > 1000$

Therefore, at equilibrium, $[(CH_3)_3N(g)] \approx 0.22$ mol/L

Step 4: $K_b = \dfrac{[(CH_3)_3NH^+][OH^-]}{[(CH_3)_3N]} = 6.5 \times 10^{-5}$

$$\frac{x^2}{0.22} = 6.5 \times 10^{-5}$$

$x = 3.8 \times 10^{-3}$

$[OH^{-}(aq)] = 3.8 \times 10^{-3}$ mol/L

$$\text{percent ionization} = \frac{3.8\times10^{-3}}{0.22}\times100 = 1.7\%$$

Check Your Answer
The answers for the $[OH^{-}(aq)]$ and percent ionization are reasonable for a weak acid. The answers have the correct number of significant digits (2).

33.
Problem
An aqueous solution of ammonia has a pH of 10.85. What is the concentration of the solution?

What is Required?
You must calculate concentration of an ammonia, $NH_3(aq)$, solution.

What is Given?
K_b for $NH_3(aq) = 1.8 \times 10^{-5}$
pH of the $NH_3(aq) = 10.85$

Plan Your Strategy
Step 1 Write the equation for the ionization of $NH_3(aq)$.
Step 2 Determine the pOH using the expression pH + pOH = 14.00
Step 3 Calculate $[OH^{-}(aq)]$ from the pOH: $[OH^{-}(aq)] = 10^{-pOH}$
Step 5 Set up an ICE table for this ionization. Let x = initial acid concentration. Use the value of $[OH^{-}(aq)]$ and the stoichiometry of the equation to calculate the equilibrium concentrations of each species.
Step 6 Write the expression for the K_b of the acid. Substitute the equilibrium concentrations and the K_b into this equation to calculate the initial acid concentration.

Act on Your Strategy
Step 1: $NH_3(aq) + H_2O(l) \leftrightarrow NH_4^{+}(aq) + OH^{-}(aq)$

Step 2: pOH = 14.00 −10.85 = 3.15
$[OH^{-}(aq] = 10^{-3.15} = 7.08 \times 10^{-4}$ mol/L

Step 3: ICE Table
Concentration (mol/L) $NH_3(aq) + H_2O(l) \leftrightarrow NH_4^{+}(aq) + OH^{-}(aq)$

initial	x		0.00	≈ 0.00
change	-7.08×10^{-4}		$+7.08 \times 10^{-4}$	$+7.08 \times 10^{-4}$

equilibrium	$x - 7.08 \times 10^{-4}$		7.08×10^{-4}	7.08×10^{-4}

Step 4: $K_b = \dfrac{[NH_4^+][OH^-]}{[NH_3]} = 1.8 \times 10^{-5}$

$$1.8 \times 10^{-5} = \frac{(7.08 \times 10^{-4})^2}{(x - 7.08 \times 10^{-4})}$$

$$1.8 \times 10^{-5}\, x - (1.8 \times 10^{-5})(7.08 \times 10^{-4}) = (7.08 \times 10^{-4})^2$$

$x = 0.028$

$[NH_3(aq)] = 0.028$ mol/L

Check Your Answer
The concentration of $NH_3(aq)$ reasonable for a weak base and it has the correct number of significant digits (2).

34.
Problem
For each of the following acids, write the name and formula of its conjugate base. Use Table 1 in Appendix G to calculate K_b for the conjugate base:
a) hydrocyanic acid, $HCN(aq)$
b) nitrous acid, $HNO_2(aq)$

Solution
The acid-base conjugate pair will differ by one more proton, $H^+(aq)$.

$$K_b = \frac{K_w}{K_a}$$

a) The conjugate base of $HCN(aq)$ is $CN^-(aq)$.

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{6.2 \times 10^{-10}} = 1.6 \times 10^{-5}$$

b) The conjugate base of $HNO_2(aq)$ is $NO_2^-(aq)$.

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-3}} = 1.8 \times 10^{-12}$$

Check Your Solution

The acid and its conjugate base differ by one proton $H^+(aq)$. The stronger acid will have a weaker conjugate base, $HNO_2(aq)/NO_2^-(aq)$, and the weaker acid will have a stronger base, $HCN(aq)/CN^-(aq)$.

35.

Problem

Write the formula for the conjugate acid of methylamine, $CH_3NH_2(aq)$, and pyridine, $C_5H_5N(aq)$. Then use the data in Table 2 in Appendix G to decide which of these conjugate acids is stronger. Explain your choice.

Solution

The acid-base conjugate pair will differ by one proton, $H^+(aq)$. The acid with the larger K_a is the stronger acid.

The conjugate acid of methylamine, $CH_3NH_2(aq)$ is $CH_3NH_3^+(aq)$.
K_a for $CH_3NH_3^+(aq)$ is 2.3×10^{-11}.
The conjugate acid of $C_5H_5N(aq)$ is $C_5H_5NH^+(aq)$
K_a for $C_5H_5NH^+(aq)$ is 7.1×10^{-6}
The stronger acid is $C_5H_5NH^+(aq)$.

Check Your Solution

The K_a for $C_5H_5NH^+(aq)$ is greater than the K_a for $CH_3NH_3^+(aq)$.

36.

Problem

Which is the stronger Brønsted-Lowry base, $CN^-(aq)$ or $F^-(aq)$? Give a reason for your choice.

Solution

The stronger base will have the weaker conjugate acid.
The conjugate acid of $CN^-(aq)$ is HCN. K_a for HCN is 6.2×10^{-10}. The conjugate acid of $F^-(aq)$ is HF. K_a for HF is 6.3×10^{-4}. Therefore, $CN^-(aq)$ is the stronger base.

Check Your Solution

The conjugate acid of $CN^-(aq)$ is weaker (smaller K_a) than the conjugate acid of $F^-(aq)$. Therefore, $CN^-(aq)$ is the stronger base.

37.

Problem

Hydrosulfuric acid is a weak diprotic acid:

$$H_2S(aq) + H_2O(l) \leftrightarrow HS^-(aq) + H_3O^+(aq)\ K_{a1} = 8.9 \times 10^{-8}$$
$$HS^-(aq) + H_2O(l) \leftrightarrow S^{2-}(aq) + H_3O^+(aq)\ K_{a2} = 1.0 \times 10^{-19}$$

To calculate K_b for the sulfide ion, should you use K_{a1} or K_{a2} for hydrosulfuric acid? Explain. Calculate K_b for $S^{2-}(aq)$.

Solution

K_{a2} should be used to calculate K_b for the sulfide ion.
Since K_{a2} for $HS^-(aq)$ (1.0×10^{-19}) << K_{a1} for $H_2S(aq)$ (8.9×10^{-8}), the conjugate base of $HS^-(aq)$ is stronger than the conjugate base of $H_2S(aq)$.

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-19}} = 1.0 \times 10^{5}$$

K_b for $S^{2-} = 1.0 \times 10^{-5}$

Check Your Solution

There is an inverse relationship between the strength of an acid and its conjugate base. Strong acids have weaker conjugate bases and weak acids have stronger conjugate bases. $HS^-(aq)$ is a very weak acid, therefore the conjugate base $S^{2-}(aq)$ is a strong base.

38.

Problem

Predict the direction for the reaction between dihydrogen phosphate ions and aqueous ammonia:

$$H_2PO_4^-(aq) + NH_3(aq) \leftrightarrow NH_4^+(aq) + HPO_4^{2-}(aq)$$

State whether reactants or products are favoured, and give reasons to support your decision.
K_a for $H_2PO_4^-(aq) = 6.2 \times 10^{-8}$.
K_b for $NH_3(aq) = 1.8 \times 10^{-5}$.

What is Required?

You must determine whether reactants or products are favoured in the given reactions.

What is Given?

The acid-base reaction is: $H_2PO_4^-(aq) + NH_3(aq) \leftrightarrow NH_4^+(aq) + HPO_4^{2-}(aq)$
K_a for $H_2PO_4^-(aq) = 6.2 \times 10^{-8}$.
K_b for $NH_3(aq) = 1.8 \times 10^{-5}$.

Plan Your Strategy

The reaction will proceed from the stronger acid to the weaker acid.
Identify the acid-base conjugate pairs in the given reaction.

Calculate K_a for the acid on the right side of the equation using $K_a = \frac{K_w}{K_b}$

Compare the strength of the acids on both sides of the equation.

Act on Your Strategy

$H_2PO_4^-(aq) + NH_3(aq) \leftrightarrow NH_4^+(aq) + HPO_4^{2-}(aq)$
acid 1 base 1 acid 2 base 2

$$K_a(NH_4^+) = \frac{K_w}{K_b(NH_3)} = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

Since K_a ($H_2PO_4^-$(aq)) > K_a (NH_4^+(aq)) the reaction is proceeds to the right and products are favoured.

Check Your Solution

There is an inverse relationship between the strength of an acid and its conjugate base. The reaction will be favoured to the side having the weaker acid.