According to Bronsted-Lowry Theory, which of the following is not an acid?

(a) HCOOH        (b) HS⁻
(c) NH₄⁺          (d) CH₄
3.3-020

According to Bronsted-Lowry Theory, which of the following is a base?

(a) HCN  
(b) CH₄  
(c) H₂S  
(d) NH₃
The equilibrium constant for which of the following reactions in water is an acidity constant?

(a) $\text{NH}_4^+ + \text{H}_2\text{O} \rightarrow \text{NH}_3 + \text{H}_3\text{O}^+$
(b) $\text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{NH}_4^+ + \text{OH}^-$
(c) $\text{HCOO}^- + \text{H}_3\text{O}^+ \rightarrow \text{HCOOH} + \text{H}_2\text{O}$
(d) $\text{HCN} + \text{OH}^- \rightarrow \text{CN}^- + \text{H}_2\text{O}$
K_a(HCOOH) is the equilibrium constant for which of the following reactions in water?

(a) \[ \text{HCOOH} + \text{NH}_3 \rightarrow \text{HCOO}^- + \text{NH}_4^+ \]

(b) \[ \text{HCOOH} + \text{H}_2\text{O} \rightarrow \text{HCOO}^- + \text{H}_3\text{O}^+ \]

(c) \[ \text{HCOO}^- + \text{H}_3\text{O}^+ \rightarrow \text{HCOOH} + \text{H}_2\text{O} \]

(d) \[ \text{HCOOH} + \text{OH}^- \rightarrow \text{HCOO}^- + \text{H}_2\text{O} \]
3.3-050

$K_a(\text{HS}^-)$ is the equilibrium constant for which of the following reactions in water?

(a) $\text{HS}^- + \text{OH}^- \rightarrow \text{H}_2\text{O} + \text{S}^{2-}$
(b) $\text{HS}^- + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{S}^{2-}$
(c) $\text{HS}^- + \text{H}_2\text{O} \rightarrow \text{OH}^- + \text{H}_2\text{S}$
(d) $\text{HS}^- + \text{H}_3\text{O}^+ \rightarrow \text{H}_2\text{O} + \text{H}_2\text{S}$
K\textsubscript{b}(HS\textsuperscript{-}) is the equilibrium constant for which of the following reactions in water?

(a) \( HS^- + OH^- \rightarrow H_2O + S^{2-} \)

(b) \( HS^- + H_2O \rightarrow H_2S + OH^- \)

(c) \( HS^- + H_3O^+ \rightarrow H_2O + H_2S \)

(d) \( HS^- + H_2O \rightarrow H_3O^+ + S^{2-} \)
The equilibrium constant for the reaction

$$\text{NH}_4^+ + \text{OH}^- \rightarrow \text{NH}_3 + \text{H}_2\text{O}$$

is

(a) $K_a(\text{NH}_4^+)$  
(b) $K_b(\text{NH}_4^+)$  
(c) $1 / K_a(\text{NH}_3)$  
(d) $1 / K_b(\text{NH}_3)$
Which of the following is the strongest acid?

(a) $H_3AsO_4$  (b) $H_2AsO_4^-$
(c) $HAsO_4^{2-}$  (d) $AsO_4^{3-}$
3.3-120

The following figure shows the fractions of the protonated (HA) and deprotonated (A⁻) forms of two monoprotic acids.

Which line represents the protonated form of the stronger acid?
The following figure shows the fractions of the four forms of a triprotic acid H₃A.

Which line represents A³⁻?
The following figure shows the fractions of the protonated and deprotonated forms of two bases.

Which curve represents the protonated form of the stronger base?
If a 1M solution was prepared of each of the following acids, which would have the highest $[\text{H}_3\text{O}^+]$?

(a) CH$_3$COOH ($K_a = 1.76 \times 10^{-5}$ M)
(b) HF ($K_a = 7.2 \times 10^{-4}$ M)
(c) HCOOH ($K_a = 1.8 \times 10^{-4}$ M)
(d) HOCl ($K_a = 3.5 \times 10^{-8}$ M)
When an acid HA is dissolved in water, which is not true?
(a) The greater the total concentration of the acid, the greater \([A^-]\).
(b) The greater the total concentration of the acid, the greater \([HA]\).
(c) The greater the total concentration of the acid, the greater \([H_3O^+]\).
(d) The greater the total concentration of the acid, the greater the degree of dissociation.
An acid is fully dissociated when the total concentration of acid in all forms is
(a) $>> K_a$  (b) $\sim K_a$  (c) $<< K_a$
In a 1 M solution of HClO₂ ($K_a = 1 \times 10^{-2}$ M), $[\text{H}_3\text{O}^+]$ will be

(a) 1 M  
(b) $1 \times 10^{-1}$ M  
(c) $1 \times 10^{-2}$ M
Which of the following acid solutions has the higher $[H_3O^+]$?

(a) $1 \times 10^{-3}$ M HIO$_3$ ($K_a = 1.6 \times 10^{-1}$ M)

(b) 1 M HClO$_2$ ($K_a = 1.1 \times 10^{-2}$ M)
Which of the following acid solutions has the higher \([\text{H}_3\text{O}^+]\) ?

(a) \(1 \times 10^{-3} \text{ M} \ \text{HClO} \ (K_a = 3.0 \times 10^{-8} \text{ M})\)

(b) \(1 \times 10^{-4} \text{ M} \ \text{HClO}_2 \ (K_a = 1.1 \times 10^{-2} \text{ M})\)
The ion product of water varies with temperature as follows:

<table>
<thead>
<tr>
<th>$T$ (°C)</th>
<th>$K_w$ (M²)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.0</td>
<td>$0.12 \times 10^{-14}$</td>
</tr>
<tr>
<td>25.0</td>
<td>$1.00 \times 10^{-14}$</td>
</tr>
<tr>
<td>60.0</td>
<td>$9.60 \times 10^{-14}$</td>
</tr>
</tbody>
</table>

We can conclude that as water is heated,
(a) $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ both decrease
(b) $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ both increase
(c) $[\text{H}_3\text{O}^+]$ increases and $[\text{OH}^-]$ decreases
(d) $[\text{H}_3\text{O}^+]$ decreases and $[\text{OH}^-]$ increases
At 25°C, a 0.010 M solution of HCl has a [OH⁻] of

(a) 0.010 x 10^{-14} M
(b) 100.0 x 10^{-14} M
(c) 100.0 x 10^{+14} M
(d) 0.010 x 10^{+14} M
At 25°C, a 1.1 M solution of HClO₂ \((K_a = 1.1 \times 10^{-2} \text{ M})\) has a \([\text{OH}^-]\) of

(a) 9 \times 10^{-14} \text{ M}
(b) 9 \times 10^{-13} \text{ M}
(c) 9 \times 10^{-12} \text{ M}
(d) 8 \times 10^{-12} \text{ M}
A solution has pH=10.82. The [H\(^+\)] in this solution is

(a) \(1.5 \times 10^{-11}\) M  
(b) \(6.6 \times 10^{-10}\) M  
(c) \(1.5 \times 10^{-5}\) M  
(d) \(6.6 \times 10^{-4}\) M

The [OH\(^-\)] in this solution is

(a) \(1.5 \times 10^{-11}\) M  
(b) \(6.6 \times 10^{-10}\) M  
(c) \(1.5 \times 10^{-5}\) M  
(d) \(6.6 \times 10^{-4}\) M
A solution has $[\text{OH}^-] = 4.8 \times 10^{-3}$ M. Its pH is

(a) 11.7
(b) 8.4
(c) 4.8
(d) 2.3
The pH of a 0.250 M HA solution is 5.00. The $K_a$ for HA is

(a) $0.250 \times 10^{-10} \text{ M}$
(b) $0.250 \times 10^{-5} \text{ M}$
(c) $4.00 \times 10^{-10} \text{ M}$
(d) $4.00 \times 10^{-5} \text{ M}$
3.3-310

The conjugate base of HS\(^{-}\) is

(a) \(S^{2-}\)          (b) \(H_2S\)
The conjugate acid of $\text{H}_2\text{PO}_4^-$ is

(a) $\text{HPO}_4^{2-}$  
(b) $\text{H}_3\text{PO}_4$
All of the following are acid base conjugate pairs except:

(a) $\text{HONO, NO}_2^-$
(b) $\text{CH}_3\text{NH}_3^+, \text{CH}_3\text{NH}_2$
(c) $\text{C}_6\text{H}_5\text{COOH}, \text{C}_6\text{H}_5\text{COO}^-$
(d) $\text{H}_3\text{O}^+, \text{OH}^-$
$K_b(\text{HC}_2\text{O}_4^-) = $

(a) $\frac{1}{K_a(\text{H}_2\text{C}_2\text{O}_4)}$

(b) $\frac{1}{K_a(\text{C}_2\text{O}_4^{2-})}$

(c) $\frac{K_w}{K_a(\text{H}_2\text{C}_2\text{O}_4)}$

(d) $\frac{K_w}{K_a(\text{C}_2\text{O}_4^{2-})}$
Which of the following is the strongest base?

(a) $H_3AsO_4$  (b) $H_2AsO_4^-$
(c) $HAsO_4^{2-}$  (d) $AsO_4^{3-}$
Which of the following acids has the strongest conjugate base?

(a) $\text{HClO}_4$    (b) $\text{HClO}_3$
(c) $\text{HClO}_2$    (d) $\text{HClO}$
Given that
\[ K_a(C_5H_5NH^+) = 5.6 \times 10^{-6} \text{ M} \]
\[ K_a(HCOOH) = 1.77 \times 10^{-4} \text{ M} \]
the salt pyridinium formate
\((C_5H_5NH^+HCOO^-)\) is
(a) acidic  (b) basic  (c) neutral
Given that

\[ K_a(H_3PO_4) = 7.5 \times 10^{-3} \text{ M} \]
\[ K_a(H_2PO_4^-) = 6.23 \times 10^{-8} \text{ M} \]
\[ K_a(HPO_4^{2-}) = 2.2 \times 10^{-13} \text{ M} \]

the salt NaH_2PO_4 is

(a) acidic  (b) basic  (c) neutral

the salt Na_2H_2PO_4 is

(a) acidic  (b) basic  (c) neutral
NH\(_4\)NO\(_3\) is

(a) acidic  (b) basic  (c) neutral

\([H_3O^+]\) in a 0.1 M aqueous solution of this salt is obtained by solving

(a) \(\frac{x^2}{0.1M-x} = K_a(\text{NH}_4^+)\)

(b) \(\frac{x^2}{0.1M-x} = K_a(\text{HNO}_3)\)

(c) \(\frac{x^2}{0.1M-x} = K_b(\text{NH}_3)\)

(d) \(\frac{x^2}{0.1M-x} = K_b(\text{NO}_3^-)\)
KCN is
(a) acidic  (b) basic  (c) neutral

[OH\(^{-}\)] in a 0.5 M aqueous solution of this salt is obtained by solving

(a) \( \frac{x^2}{0.5M-x} = K_a(K^+) \)
(b) \( \frac{x^2}{0.5M-x} = K_a(HCN) \)
(c) \( \frac{x^2}{0.5M-x} = K_b(KOH) \)
(d) \( \frac{x^2}{0.5M-x} = K_b(CN^-) \)
Given that

\[ K_a(\text{CH}_3\text{NH}_3^+) = 2.38 \times 10^{-11} \text{ M} \]
\[ K_a(\text{HCN}) = 4.9 \times 10^{-10} \text{ M} \]
\[ K_a(\text{HOAc}) = 1.8 \times 10^{-5} \text{ M} \]
\[ K_a(\text{NH}_4^+) = 5.7 \times 10^{-10} \text{ M} \]

which of the following solutions has the highest pH?

(a) 0.1M CH\(_3\)NH\(_3\)Cl
(b) 0.1M NaCN
(c) 0.1M NH\(_4\)NO\(_3\)
(d) 0.1M NaOAc

The lowest pH?
The equilibrium constant for the reaction

\[ \text{HONO (aq) + CN}^- (aq) \rightarrow \text{HCN (aq) + ONO}^- (aq) \]

is $1.1 \times 10^6$ M. This value indicates that

(a) CN$^-$ is a stronger base than ONO$^-$

(b) HCN is a stronger acid than HONO

(c) the conjugate base of HONO is ONO$^-$

(d) the conjugate acid of CN$^-$ is HCN
The equilibrium constant for the reaction

$$\text{HClO}_2 \text{ (aq)} + \text{NO}_2^- \text{ (aq)} \rightarrow \text{HNO}_2 \text{ (aq)} + \text{ClO}_2^- \text{ (aq)}$$

is

(a) $K_a \text{ (HClO}_2) \cdot K_b \text{ (NO}_2^-)$

(b) $K_a \text{ (HNO}_2) \cdot K_b \text{ (ClO}_2^-)$

(c) $K_a \text{ (HClO}_2) / K_a \text{ (HNO}_2)$

(d) $K_b \text{ (ClO}_2^-) / K_b \text{ (NO}_2^-)$
[NH₃] and [NH₄⁺] will be equal when [H₃O⁺] is

(a)  Kₐ(NH₄⁺)
(b)  K₉(NH₃)
(c)  1 / Kₐ(NH₄⁺)
(d)  1 / K₉(NH₃)
Methyl orange has a pK$_a$ of 3.7. It will change color when the [H$_3$O$^+$] is in the vicinity of

(a) 3.7 M  (c) 10$^{-3.7}$ M
(b) 10.3 M  (d) 10$^{-10.3}$ M
Bromthymol blue changes color when the pH is in the vicinity of 7.1. Its $K_a$ is

(a) $7.1 \text{ M}$  
(b) $6.9 \text{ M}$

(c) $10^{-7.1} \text{ M}$  
(d) $10^{-6.9} \text{ M}$
A flask contains 100 mL of 0.1 M HOAc. To prepare a buffer with pH=pK_a(HOAc) which of the following samples should be added to the flask?

(a) 50 mL of 0.2 M Ba(OAc)_2
(b) 200 mL of 0.1 M Ba(OAc)_2
(c) 25 mL of 0.2 M Ba(OAc)_2
(d) 100 mL of 0.1 M Ba(OAc)_2
Which mixture has a different pH from the others?

(a) 1 L of 1 M NaOH + 1 L of 2 M HCN
(b) 1 L of 1 M NaCN + 1 L of 1 M HCN
(c) 1 L of 1 M NaOH + 2 L of 1 M HCN
(d) 1 L of 1 M NaOH + 1 L of 1 M HCN
3.3-550

Which of the following mixtures will be a buffer solution when dissolved in 500 mL of water?

(a) 0.2 mol $\text{C}_6\text{H}_5\text{NH}_2 + 0.2$ mol $\text{HCl}$
(b) 0.2 mol $\text{C}_6\text{H}_5\text{NH}_2 + 0.2$ mol $\text{NaOH}$
(c) 0.2 mol $\text{C}_6\text{H}_5\text{NH}_2 + 0.1$ mol $\text{HCl}$
(d) 0.2 mol $\text{C}_6\text{H}_5\text{NH}_2 + 0.1$ mol $\text{NaOH}$
Which of the following mixtures is a buffer with pH close to $pK_a(\text{NH}_4^+)$

(a) 30 mL of 0.1 M NH$_4$Cl 
    + 10 mL of 0.1 M NaOH

(b) 30 mL of 0.1 M NH$_4$Cl 
    + 15 mL of 0.1 M NaOH

(c) 30 mL of 0.1 M NH$_4$Cl 
    + 20 mL of 0.1 M NaOH

(d) 30 mL of 0.1 M NH$_4$Cl 
    + 30 mL of 0.1 M NaOH
A weak base, B, has a basicity constant $K_b = 2 \times 10^{-5} \text{ M}$. The pH of any solution in which $[B] = [BH^+]$ is

(a) 4.7  (b) 5.3
(c) 8.7  (d) 9.3
The following figure shows the composition diagram of $\text{H}_3\text{PO}_4$.

Which pH cannot be buffered well by phosphate?

(a) 2  (b) 5  (c) 7  (d) 13
The following figure shows the composition diagram of $\text{H}_3\text{PO}_4$.

To make a pH 7 buffer which mixture would you use?

(a) $\text{Na}_3\text{PO}_4$ and $\text{Na}_2\text{HPO}_4$
(b) $\text{Na}_2\text{HPO}_4$ and $\text{NaH}_2\text{PO}_4$
(c) $\text{NaH}_2\text{PO}_4$ and $\text{H}_3\text{PO}_4$

For a pH 2 buffer?
A buffer is prepared that contains 1 mol NaOAc and 1 mol HOAc. When 0.5 mol of NaOH is added, the pH

(a) doubles
(b) increases by $\log 2 = 0.301$
(c) increases by $\log 3 = 0.477$
(d) depends on the volume
100 mL of a solution of a solution of X is titrated with a 0.1 M solution of Y giving

Y is a

(a) strong base  (c) weak base
(b) strong acid   (d) weak acid

The initial concentration of X is
(a) 1 M  (b) 0.1 M  (c) 0.05 M  (d) 0.01 M

X is a

(a) strong base  (c) weak base
(b) strong acid   (d) weak acid
100 mL samples of two different acids are titrated with 0.1 M NaOH giving

Which is the more concentrated acid?

Which is the weaker acid?

The molarity of acid a is
(a) 0.01 (b) 0.1 (c) 1 (d) 10

The molarity of acid b is
(a) 0.01 (b) 0.02 (c) 0.05 (d) 0.2
In titrating 0.10 M HCOOH (formic acid) with 0.10 M NaOH, the solution at the equivalence point is

(a) 0.10 M HCOONa
(b) 0.05 M HCOOH
(c) 0.05 M HCOONa
(d) 0.05 M NaOH
A 50 mL sample of 30 mM Ba(OH)$_2$ is titrated with 30 mM HCl. The solution at the equivalence point is

(a) 60 mM BaCl$_2$
(b) 30 mM BaCl$_2$
(c) 15 mM BaCl$_2$
(d) 10 mM BaCl$_2$
A 10 mL sample of a 0.2 M solution of the weak diprotic acid H₂X is titrated with 0.2 M KOH. The solution at the second equivalence point is

(a) 0.2 M KHX
(b) 0.2 M K₂X
(c) 0.1 M KHX
(d) 0.0667 M K₂X