The Brønsted-Lowry definition of an acid is a substance capable of donating a proton (H\(^+\)), and a base is a substance capable of accepting a proton. For example, the weak acid, HF, can be dissolved in water, giving the reaction:

\[
\text{HF (aq)} + \text{H}_2\text{O (l)} \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{F}^- (aq)
\]

In this reaction, HF is the species losing the proton (H\(^+\)), making it the acid. Water is the species accepting the proton, to form the hydronium ion, H\(_3\)O\(^+\), making it the base. The F\(^-\) (aq) is called the conjugate base of HF. It can gain a proton in the reverse reaction. H\(_3\)O\(^+\) is the conjugate acid of H\(_2\)O, since it can lose a proton in the reverse reaction. The stronger an acid, the weaker its conjugate base will be and the stronger the base, the weaker its conjugate acid.

The equilibrium concentrations of these species will be determined by the relative strengths of the acids and bases. The strongest acid will dissociate to the greatest extent. H\(_3\)O\(^+\) (H\(^+\)) is the strongest acid that can exist in an aqueous system. So, the equilibrium in this system will favor the reactants, HF and H\(_2\)O, the weaker acid and base. The equilibrium state is described by an equilibrium constant, \(K_a\), in the case of acids, and \(K_b\) in the case of bases. These are related by the expression \(K_w = K_a \times K_b = 1 \times 10^{-14}\).

1. Classify the following as Brønsted acids, bases or both.
   If the compound has a hydrogen, it can be an acid; if it has any lone pairs, it can be a base
   a) H\(_2\)O  b) OH\(^-\)  c) NH\(_3\)
   both   both   both
   d) NH\(_4\)^+  e) NH\(_2\)^-  f) CO\(_3\)^{2-}
   acid   both   base

2. What is the conjugate base of the following acids?
   a) HClO\(_4\)  b) NH\(_4\)^+  c) H\(_2\)O  d) HCO\(_3\)^-
   ClO\(_4\)^-  NH\(_3\)  OH\(^-\)  CO\(_3\)^{2-}
   (Remove one hydrogen and reduce the charge by 1)

3. What is the conjugate acid of the following bases?
   a) CN\(^-\)  b) SO\(_4^{2-}\)  c) H\(_2\)O  d) HCO\(_3\)^-
   HCN   HSO\(_4\)^-   H\(_3\)O\(^+\)   H\(_2\)CO\(_3\)
   (Add one hydrogen and increase the charge by 1)
4. \(\text{HSO}_3^-\) is amphoteric; it can behave as either an acid or a base. In the examples above, you may have noticed that \(\text{H}_2\text{O}\) and \(\text{HCO}_3^-\) can also exist as acids or bases.

a) Write the equation for the reaction of \(\text{HSO}_3^-\) with \(\text{H}_2\text{O}\) in which it acts like an acid and identify the acid-base pairs. Circle the strongest acid. Then write an expression for \(K_a\). The value of \(K_a\) at 25\(^\circ\) C is 1.23 x 10\(^{-7}\). Use an arrow to indicate which side is favored by the equilibrium.

\[
\text{HSO}_3^- + \text{H}_2\text{O} \rightleftharpoons \text{SO}_3^{2-} + \text{H}_3\text{O}^+
\]

\[
K_a = \frac{[\text{SO}_3^{2-}][\text{H}_3\text{O}^+]}{[\text{HSO}_3^-]} = 1.23 \times 10^{-7}
\]

b) Write the equation for the reaction of \(\text{HSO}_3^-\) with water in which it acts like a base and identify the acid-base pairs. Circle the strongest base. Then write an expression for \(K_b\). What is the value of \(K_b\) at 25\(^\circ\) C, given that the \(K_a\) for \(\text{H}_2\text{SO}_3\) = 1.58 x 10\(^{-2}\)? Use an arrow to indicate which side is favored by the equilibrium.

\[
\text{HSO}_3^- + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{SO}_3 + \text{OH}^-
\]

\[
K_b = \frac{[\text{H}_2\text{SO}_3][\text{OH}^-]}{[\text{HSO}_3^-]}
\]

\[
K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.58 \times 10^{-2}} = 6.33 \times 10^{-13}
\]

c) If \(\text{HSO}_3^-\) is placed in water, is the resulting solution acidic or basic?

\(K_a > K_b\), so \(\text{HSO}_3^-\) will produce more \(\text{H}_3\text{O}^+\) than \(\text{OH}^-\) and the solution overall will be acidic.

d) Compute a value for the equilibrium constant for the reaction shown below:

\[
\text{SO}_3^{2-} + \text{H}_3\text{O}^+ \rightleftharpoons \text{HSO}_3^- + \text{H}_2\text{O} \quad K = 8.13 \times 10^6
\]

\[
K = \frac{[\text{HSO}_3^-]}{[\text{SO}_3^{2-}][\text{H}_3\text{O}^+]} = \frac{1}{K_a}\text{ from part a} = \frac{1}{1.23 \times 10^{-7}} = 8.13 \times 10^6
\]

Circle the strongest acid and indicate which side is favored by the equilibrium.
5. Identify the conjugate acid-base pairs in the following reactions. Circle the strongest acid and indicate which side is favored at equilibrium:

a) \( \text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+ \) 
- \( \text{NH}_4^+ / \text{NH}_3 \) (acid/base)
- \( \text{H}_2\text{O} / \text{H}_3\text{O}^+ \) (base/acid)

b) \( \text{HCOOH} + \text{CN}^- \rightleftharpoons \text{HCOO}^- + \text{HCN} \) 
- \( \text{HCOOH} / \text{HCOO}^- \) (acid/base)
- \( \text{CN}^- / \text{HCN} \) (base/acid)

- \( K_a \text{HCOOH} = 1.8 \times 10^{-5} \)
- \( K_a \text{HCN} = 6.2 \times 10^{-10} \)

Kb \( \text{NH}_3 = 1.8 \times 10^{-5} \)
- \( K_b \text{CO}_3^{2-} = 1.8 \times 10^{-4} \)

c) \( \text{NH}_4^+ + \text{CO}_3^{2-} \rightleftharpoons \text{NH}_3 + \text{HCO}_3^- \) 
- \( \text{NH}_4^+ / \text{NH}_3 \) (acid/base)
- \( \text{CO}_3^{2-} / \text{HCO}_3^- \) (base/acid)

Kb \( \text{NH}_3 = 1.8 \times 10^{-5} \)
- \( K_b \text{CO}_3^{2-} = 1.8 \times 10^{-4} \)

d) \( \text{H}_2\text{PO}_4^- + \text{OH}^- \rightleftharpoons \text{HPO}_4^{2-} + \text{H}_2\text{O} \) 
- \( \text{H}_2\text{PO}_4^- / \text{HPO}_4^{2-} \) (acid/base)
- \( \text{OH}^- / \text{H}_2\text{O} \) (base/acid)

6. Name 6 strong acids. Name 6 strong bases.

**Strong acids:** HCl, HBr, HI, HNO₃, HClO₄, HClO₃, H₂SO₄

**Strong bases:** LiOH, NaOH, KOH, RbOH, CsOH, Ca(OH)₂, Sr(OH)₂, Ba(OH)₂

7. Complete the following table, which deals with pH, \([H^+]\), acidity, neutrality and basicity.

<table>
<thead>
<tr>
<th>pH</th>
<th>([H^+])</th>
<th>Solution is:</th>
</tr>
</thead>
<tbody>
<tr>
<td>&lt; 7</td>
<td>&gt; 1.0 x 10⁻⁷</td>
<td>acidic</td>
</tr>
<tr>
<td>&gt; 7</td>
<td>&lt; 1.0 x 10⁻⁷</td>
<td>basic</td>
</tr>
<tr>
<td>7</td>
<td>= 1.0 x 10⁻⁷</td>
<td>neutral</td>
</tr>
</tbody>
</table>
8. Calculate the pH of the solution formed when 100 mL of 1.00 M HCl is added to 1.00 L of water.

$$0.100 \text{ L} \times \frac{1.00 \text{ mol HCl}}{1 \text{ L}} \times \frac{1 \text{ mol H}^+}{1 \text{ mol HCl}} = 0.100 \text{ mol H}^+$$

$$[H^+] = \frac{0.100 \text{ mol H}^+}{1.00 + 0.100 \text{ L}} = 0.0909 \text{ M}$$

$$\text{pH} = -\log(0.0909) = 1.04$$

9. Calculate the pH of the following solutions of strong acids and bases:

a) 0.001 M HCl

$$[H^+] = 0.001 \text{ M}$$

$$\text{pH} = -\log(0.001) = 3$$

b) 0.76 M KOH

$$[OH^-] = 0.76 \text{ M}$$

$$\text{pOH} = -\log(0.76) = 0.119$$

$$\text{pH} = 14 - 0.119 = 13.9$$

c) 2.8 \times 10^{-4} \text{ M Ba(OH)_2} \quad \text{(Think carefully about this one!)}$

When Ba(OH)_2 dissociates, two OH^- ions are produced, so the total concentration of OH^- is double the concentration of Ba(OH)_2

$$[OH^-] = 2(2.8 \times 10^{-4} \text{ M}) = 5.6 \times 10^{-4} \text{ M}$$

$$\text{pOH} = -\log(5.6 \times 10^{-4}) = 3.25$$

$$\text{pH} = 14 - 3.25 = 10.75$$