ACID/BASE THEORY

For the ionization of water:

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

an equilibrium expression can be written:

\[
K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14} \quad (@ \text{ 25}^\circ\text{C})
\]

\[-\text{or}-
\[
K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14} \quad (@ \text{ 25}^\circ\text{C})
\]

\(K_w\) is known as the ion product of water. In ANY aqueous solution this equilibrium is always present. One can always obtain the concentration of \(\text{H}^+\) from the concentration of \(\text{OH}^-\) and vice versa. A strong acid will set the value of \([\text{H}^+]\) and then \([\text{OH}^-]\) can be calculated. A strong base will set the value of \([\text{OH}^-]\) and then \([\text{H}^+]\) can be calculated.

The pH of a solution can be determined by the expression:

\[
\text{pH} = -\log[\text{H}^+] \quad \text{which means that} \quad [\text{H}^+] = 10^{-\text{pH}}
\]

In general, the "p" in pH is really a function: \(p(X) = -\log(X)\) but instead of writing \(p(X)\) we write \(pX\).

So you can easily calculate values for \(p\text{OH}\), \(pK_a\), etc...

This leads to the "-log" form of the above equation:

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

The ONLY time that \([\text{H}^+] = [\text{OH}^-]\) (definition of neutral water) is when they each equal \(1.0 \times 10^{-7}\) M which corresponds to a pH (and pOH for that matter) of 7.00. This is neutral pH and is the pH of pure \(\text{H}_2\text{O}\) (note this is only at \(25^\circ\text{C}\), however). pH values lower than 7 are acidic solutions while pH values higher than 7 are basic.

NOTE: ALL the questions in chapters 6 and 7 will fit into these equilibria and equations compare and contrast the similarities and differences in these two columns

———ACIDS———  
For a weak acid (HA) in water:

\[
\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^-
\]

\[
K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} \quad [\text{H}^+] = K_a \frac{[\text{HA}]}{[\text{A}^-]}
\]

\[
\text{pH} = pK_a + \log \frac{[\text{A}^-]}{[\text{HA}]}
\]

For a conjugate acid (BH\(^+\)) of a weak base(B) in water:

\[
\text{BH}^+ + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{B}
\]

\[
K_a = \frac{[\text{H}^+][\text{B}]}{[\text{BH}^+]} \quad [\text{H}^+] = K_a \frac{[\text{BH}^+]}{[\text{B}]}
\]

\[
\text{pH} = pK_a + \log \frac{[\text{B}]}{[\text{BH}^+]}
\]

NOTICE! You can NOT look up the \(K_a\) of BH\(^+\).
You have to calculate it from the \(K_b\) of B:

\[
K_a = \frac{K_w}{K_b}
\]

———BASES———  
For a weak base (B) in water:

\[
\text{B} + \text{H}_2\text{O} \rightleftharpoons \text{BH}^+ + \text{OH}^-
\]

\[
K_b = \frac{[\text{OH}^-][\text{BH}^+]}{[\text{B}]} \quad [\text{OH}^-] = K_b \frac{[\text{B}]}{[\text{BH}^+]}
\]

\[
\text{pOH} = pK_b + \log \frac{[\text{BH}^+]}{[\text{B}]}
\]

For a conjugate base (A\(^-\)) of a weak acid (HA) in water:

\[
\text{A}^- + \text{H}_2\text{O} \rightleftharpoons \text{HA} + \text{OH}^-
\]

\[
K_b = \frac{[\text{OH}^-][\text{HA}]}{[\text{A}^-]} \quad [\text{OH}^-] = K_b \frac{[\text{A}^-]}{[\text{HA}]}
\]

\[
\text{pOH} = pK_b + \log \frac{[\text{HA}]}{[\text{A}^-]}
\]

NOTICE! You can NOT look up the \(K_b\) of A\(^-\).
You have to calculate it from the \(K_a\) of HA:

\[
K_b = \frac{K_w}{K_a}
\]