ACID/BASE THEORY

For the ionization of water : $H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$
an equilibrium expression can be written:

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \quad (@\ 25^\circ C)$$

— *or* —

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

$\boldsymbol{K_w}$ is know as the **ion product** of water. In **ANY** aqueous solution this equilibrium is always present. One can always obtain the concentration of $\mathbf{H^+}$ from the concentration of $\mathbf{OH^-}$ and vice versa. A strong acid will <u>set</u> the value of $[H^+]$ and then $[OH^-]$ can be calculated. A strong base will <u>set</u> the value of $[OH^-]$ and then $[H^+]$ can be calculated.

The pH of a solution can be determine by the expression: $pH = -\log[H^+]$ which means that $[H^+] = 10^{-pH}$

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 pOH, pK_w, etc...

This leads to the "-log" form of the above equation: 14 = pH + pOH

The ONLY time that $[H^+] = [OH^-]$ (definition of neutral water) is when they each equal 1.0×10^{-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 H_2O (note this is only at 25˚C, however). pH values lower than 7 are acidic solutions while pH values higher than 7 are basic.

NOTE: ALL the questions in chapters 18 and 19 will fit into these equilibria and equations
compare and contrast the similarities and differences in these two columns

––––A C I D S ––––– ⟷ **––––B A S E S –––––**

ACIDS

For a weak acid (HA) in water:

$HA + H_2O \rightleftharpoons H_3O^+ + A^-$

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

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

For a <u>conjugate</u> acid (BH^+) of a weak base(B) in water:

$BH^+ + H_2O \rightleftharpoons H_3O^+ + B$

$$K_a = \frac{[H^+][B]}{[BH^+]} \qquad [H^+] = K_a\frac{[BH^+]}{[B]}$$

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

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:

$B + H_2O \rightleftharpoons BH^+ + OH^-$

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

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

For a <u>conjugate</u> base (A^-) of a weak acid (HA) in water:

$A^- + H_2O \rightleftharpoons HA + OH^-$

$$K_b = \frac{[OH^-][HA]}{[A^-]} \qquad [OH^-] = K_b\frac{[A^-]}{[HA]}$$

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

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

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