Name ______________________________

Date _________________ Period _____

Calculating pH

pH is defined as the negative logarithm of the hydronium ion concentration (pH = $-\log[H_3O^+]$). For neutral substances, such as water, the hydronium ion concentration is10^{-7} *M*, and the pH is 7, because $-\log(10^{-7}) = 7$. In this case, the relationship between the exponent for the hydronium ion concentration and pH is straight forward. For less obvious examples, use a calculator. The pH of a solution with a hydronium ion concentration of 2.45×10^{-10} *M* is 9.61 because $-\log(2.45 \times 10^{-10}) = 9.61$. Check it! By the way, it is clear that since the hydronium ion concentration is between 10^{-9} *M* and 10^{-10} *M*, the pH is between 9 and 10.

pOH, on the other hand, is the negative logarithm of the hydroxide ion concentration (pOH = $-\log[OH^-]$). The equilibrium constant for water, K_w, is 10^{-14} ($K_w = [H_3O^+][OH^-] = 10^{-14}$), so pH + pOH = 14. As a result, it is possible to determine the pH if the hydroxide ion concentration is known (pH = 14 – pOH).

> **Sample Problem**
>
> **What is the pH of a solution with a hydroxide ion concentration of 1.45×10^{-9} *M*?**
> - pOH = $-\log[OH^-] = -\log(1.45 \times 10^{-9}) = 8.84$
> - pH = 14 – pOH = 5.16

Converting from pH to hydronium ion concentration or from pOH to hydroxide ion concentration is a matter of doing an antilog, again using a calculator. If the pH is 7.3, then $7.3 = -\log[H_3O^+]$. The hydronium ion concentration is 5×10^{-8} *M*. Again, you should be able to estimate that it is between 10^{-7} *M* and 10^{-8} *M*, because the pH is between 7 and 8.

pHs can also be calculated from the acid or base concentration. First consider strong acids and bases. Strong acids include HCl, HBr, HI, H_2SO_4, HNO_3, and $HClO_4$. Strong bases are hydroxides of group 1 and 2 metals. To calculate the pH of a strong acids or strong bases, assume the ions are 100 percent separated. This means, for example, that 0.15 *M* HCl has a hydronium ion concentration of 0.15 *M*, and a pH of 0.82. A solution of 0.010 *M* $Ca(OH)_2$ has a hydroxide ion concentration of 0.020 *M* because each mole of $Ca(OH)_2$ dissociates into 1 mol of calcium ions and 2 mol of hydroxide ions. As a result, the pOH is 1.7, and the pH is 12.3. Check these calculations!

The pH of weak acids and bases is a bit more complicated. It requires use of the equilibrium expression. For weak acids, write a balanced equation for the ionization of the acid, write the equilibrium expression, and write the algebraic expression substituting known values and variables for unknowns. Keep in mind that the balanced equation provides the mole ratios of the ions which are all integral multiples of the same unknown, *x*. Because the equilibrium constant is small, for weak acids, the change in the concentration of the acid when it ionizes is negligible and can be ignored. Weak bases work the same way. Find pOH instead of pH. Then subtract (pH = 14 – pOH).

> **Sample Problem**
>
> **What is the pH of 0.100 M HNO_2? ($k_a = 7.2 \times 10^{-4}$)**
> - Write a balanced equation for the ionization of the acid.
> $HNO_2 = H^+ + NO_2^-$
> - Write the equilibrium expression
> $$k_a = \frac{[H^+][NO_2^-]}{[HNO_2]}$$
> - Write the algebraic expression substituting known values and variables for unknowns.
> $$7.2 \times 10^{-4} = \frac{(x)(x)}{(0.100)}$$
> - Solve the expression for $[H^+]$ (or $[H_3O^+]$)
> $x^2 = 7.2 \times 10^{-5}$ $x = 8.5 \times 10^{-3}$
> - Calculate pH
>   pH = $-\log[H_3O^+] = -\log(8.5 \times 10^{-3}) = 2.1$

The pH of buffers is calculated using the *Henderson-Hasselbalch Equation* which says pH = $pK_a + \log([A^-]/[HA])$ where $pK_a = -\log K_a$

> **Sample Problem**
>
> **What is the pH of a solution with 0.1 *M* hydrofluoric acid (HF) and 0.01 sodium fluoride? ($K_a = 6.6 \times 10^{-4}$)**
> - pH = $-\log(6.6 \times 10^{-4}) + \log((0.01)/(0.1\ M))$
> - pH = $3.2 + \log(10^{-1}) = 3.2 + (-1) = 2.2$

(CONTINUE ON THE NEXT PAGE)

Answer the following questions based on your reading and your knowledge of chemistry.

1. Find the pH for each of the following:

   a. $[H_3O^+] = 0.0315\ M$

   b. $[OH^-] = 0.0067\ M$

   c. $0.0025\ M\ HNO_3$

   d. $0.00012\ M\ Ba(OH)_2$

   e. $0.0325\ M\ HIO_3$ ($K_a = 1.6 \times 10^{-1}$)

   f. $3.0\ M\ NH_3$ ($K_b = 1.8 \times 10^{-5}$)

2. Find the hydronium ion concentration for each of the following:

   a. A base with a pH of 8.2

   b. A base with a pOH of 3.4

3. What is the pH of a buffered solution of $0.3\ M\ HCH_3COO$ and $0.02\ M\ NaCH_3COO$? ($K_a = 1.8 \times 10^{-5}$)