Worksheet 10: Ionization

Objectives

1. Use K_w to determine relative amounts of hydroxide and hydronium ions in solution and assess whether the solution is acidic, basic, or neutral

2. Calculate equilibrium concentrations and pH using any appropriate approximations

3. Determine the pH or pOH of a solution and identify the relationship between these quantities

4. Carry out all kinds of pH calculations and calculations using pH to find other quantities

Key Questions

1. Write the expressions for K_c and K_a of water. Recall that $[H_2O] \approx 55$ M. Using this information, the expression for the K_a of water, and the fact that water's K_a is 1.8×10^{-16}, calculate the value of $[H_3O^+] \times [OH^-]$.
   Start with writing the equation:
   $$H_2O + H_2O \longrightarrow H_3O^+ + OH^-$$
   Then the K_c expression:
   $$K_c = \frac{[H_3O^+][OH^-]}{[H_2O][H_2O]}$$
   To go from K_c to K_a, we remove a factor of $[H_2O]$:
   $$K_a = \frac{[H_3O^+][OH^-]}{[H_2O]}$$
   Finally, rearrange to solve for the desired $[H_3O^+][OH^-]$:
   $$K_a \times [H_2O] = [H_3O^+][OH^-] = 1.8 \times 10^{-16} \times 55 = 1.0 \times 10^{-14}$$

2. What is the name for the value calculated in the previous problem?
   K_w

3. Use the value of K_w to calculate the hydronium and hydroxide ion concentrations in pure water. Also calculate the pH and pOH of pure water.
   In pure water, $[H_3O^+] = [OH^-]$, so $K_w = [H_3O^+]^2$ and $[H_3O^+] = [OH^-] = (K_w)^{\frac{1}{2}}$
   $(K_w)^{\frac{1}{2}} = (1.0 \times 10^{-14})^{\frac{1}{2}} = 1.0 \times 10^{-7}$
   $pH = -\log[H_3O^+] = 7$
   $pOH = -\log[OH^-] = 7$

4. For the following concentrations, state whether the associated solution will be acidic, basic, or neutral, and calculate the corresponding hydroxide or hydronium concentration.
   For a neutral solution, $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$, so if $[H_3O^+] > 1.0 \times 10^{-7}$, the solution is acidic and if $[H_3O^+] < 1.0 \times 10^{-7}$, the solution is basic. Similarly, if $[OH^-] > 1.0 \times 10^{-7}$, the solution is basic and if $[OH^-] < 1.0 \times 10^{-7}$, the solution is acidic.
   Since $K_w = 1.0 \times 10^{-14} = [H_3O^+] \times [OH^-]$, $[H_3O^+] = K_w / [OH^-]$ and $[OH^-] = K_w / [H_3O^+]$.

   (a) $[H_3O^+] = 7.2 \times 10^{-4}$ M
   acidic, $[OH^-] = 1.4 \times 10^{-11}$

   (b) $[H_3O^+] = 5.8 \times 10^{-10}$ M
   basic, $[OH^-] = 1.7 \times 10^{-5}$

   (c) $[OH^-] = 1.8 \times 10^{-6}$ M
   basic, $[H_3O^+] = 5.6 \times 10^{-9}$

   (d) $[OH^-] = 1.0 \times 10^{-7}$ M
   neutral, $[H_3O^+] = 1.0 \times 10^{-7}$

5. Given an initial concentration of 0.5 M H_2S and its K_a of 1.1×10^{-7}, determine the equilibrium concentration of HS^-, the pH of the solution, and $[OH^-]$.

(R)	$H_2S + H_2O \rightleftharpoons H_3O^+ + HS^-$			
I	0.5	–	0	0
C	−x	–	+x	+x
E	0.5−x	–	x	x

$K_a = \frac{[H_3O^+][HS^-]}{[H_2S]} \implies 1.1\times10^{-7} = \frac{x^2}{0.5-x}$

Rearrange: $1.1\times10^{-7}(0.5 - x) = x^2 \implies 5.5\times10^{-8} - 1.1\times10^{-7}x = x^2$

$\implies x^2 - 5.5\times10^{-8} + 1.1\times10^{-7}x = 0$

Now, use quadratic equation to find $x = -0.000235, 0.000234$. The negative root gives a nonsensical negative concentration, so take the positive root. From the RICE table, we see that x is the equilibrium concentration of HS^- so that part of the answer is finished. pH is given by pH $= -\log[H_3O^+]$ and x is also $[H_3O^+]$, so pH $= -\log(0.000234) = 3.63$. As we saw previously, we can get $[OH^-]$ from $K_w/[H_3O^+]$, so $[OH^-] = 1.0\times10^{-14}$ / $2.34\times10^{-4} = 4.27\times10^{-11}$ M.

6. How many moles of NH_3 must be dissolved in 1.00 liters of aqueous solution to produce a solution with a pH of 11.47? The K_a of NH_4^+ is 5.8×10^{-10}.

This will ultimately be a RICE table problem, but first we have to use the pH to calculate the equilibrium concentration of OH^-. We want OH^- in this case because NH_3 is a base and will react with water in the reaction $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$. To go from pH to pOH, use the relation that pH + pOH = 14, so $14 - $ pH $=$ pOH $= 14 - 11.47 = 2.53$. Like pH, pOH $= -\log[OH^-]$, so $[OH^-] = 10^{-pOH} = 10^{-2.53} = 2.95\times10^{-3}$. This time, our unknown is actually the initial concentration of NH_3, but we know we start with none of our products but end up with $[OH^-] = 2.95\times10^{-3}$, so we know the change and can fill in our RICE table accordingly and actually end up with an easier problem than normal.

(R)	$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$			
I	x	–	0	0
C	-2.95×10^{-3}	–	$+2.95\times10^{-3}$	$+2.95\times10^{-3}$
E	x -2.95×10^{-3}	–	2.95×10^{-3}	2.95×10^{-3}

NH_3 is a base, so we need a K_b expression, despite being given a K_a value, so convert to K_b using $K_b = K_w$ / $K_a = $ 1.0e-14 / $5.8\times10^{-10} = 1.7\times10^{-5}$

Now set up the K_b expression:

$K_a = \frac{[NH_4^+][OH^-]}{[NH_3]} \implies 1.7\times10^{-5} = \frac{(2.95\times10^{-3})^2}{x - 2.95\times10^{-3}}$

Rearrange and solve for x to get x $= 0.515$. This is the initial concentration of NH_3 required to make a solution of pH $= 11.47$, but because the problem specifies the total volume is 1 L, it is also the number of moles desired.

7. Calculate the percent ionization of the weak acid, HA, given a 0.25 M HA solution and a K_a of 5.3×10^{-7}.

The percent ionization is given by the ratio of dissociated acid to the initial amount of acid. For every molecule of acid that dissociates, one molecule of H_3O^+ and one molecule of conjugate base is produced, assuming a 1:1 stoichiometry, so percent ionization is given by the equation $\%I = \frac{[H_3O^+]}{(HA)_0} \times 100\%$ We have the initial concentration of acid, $(HA)_0$, so we just need a RICE table to calculate the equilibrium hydronium concentration.

(R)	$HA + H_2O \rightleftharpoons H_3O^+ + A^-$			
I	0.25	–	0	0
C	−x	–	+x	+x
E	0.25−x	–	x	x

$K_a = 5.3 \times 10^{-7} = \frac{x^2}{0.25 - x}$, $x = -0.000364, 0.000364$, again discarding the negative root to get x = $3.64 \times 10^{-4} = [H_3O^+]$. Dividing this by the initial acid concentration and multiplying by 100% gives the percent ionization: %I = $[H_3O^+]$ / $(HA)_0 \times 100\% = 0.14\%$.