

AP Self-Test Review Sheet for Acids/Bases Section

NOTE: This worksheet is NOT a substitute for studying all of your notes on acids and bases. You should know something about every item on the acid/base checklist, and work through all problems in the notes before using this self-test sheet.

- A beaker contains 100. mL of a solution of hypochlorous acid, HOCl, of unknown concentration
 - The solution was titrated with 0.100 molar NaOH solution and the equivalence point was reached when 40.0 mL of NaOH was added. What was the original concentration of HOCl solution?
 - If the original solution had a pH of 4.46, what is the value of K_a for HOCl?
 - What percent of the HOCl molecules were ionized in the original solution?
 - What is the concentration of OCl^- ions in the solution at the equivalence point reached in (a)?
- Use the principles of acid-base theory to answer the following questions.
 - Predict whether a 0.1 M solution of sodium acetate, $NaC_2H_3O_2$, will be acidic or basic, and give a reaction occurring with water that supports your conclusion.
 - Predict whether a 0.1 M solution of ammonium chloride, NH_4Cl , will be acidic or basic, and give a reaction occurring with water that supports your conclusion.
 - Explain why buffer solutions are made with weak acids instead of strong acids.
 - HCN is a weak acid with an acid-dissociation constant, $K_a = 4.9 \times 10^{-10}$. Write the equation for the acid-dissociation and also the equation for the base-dissociation of its conjugate base. Calculate K_b for the base-dissociation.
- Hypobromous acid, HOBr, is a weak acid that dissociates in water, as represented by the equation below;



- Calculate the value of $[H^+]$ in an HOBr solution that has a pH of 4.95.
- Write the equilibrium constant expression for the ionization of HOBr in water, then calculate the concentration of $HOBr_{(aq)}$ in an HOBr solution that has $[H^+] = 1.8 \times 10^{-5}$ M.
- A solution of $Ba(OH)_2$ is titrated into a solution of HOBr.
 - Calculate the volume of 0.115 M $Ba(OH)_2_{(aq)}$ needed to reach the equivalence point when titrated into a 65.0 mL sample of 0.146 M $HOBr_{(aq)}$.
 - Indicate whether the pH at the equivalence point is less than 7, equal to 7, or greater than 7. Explain.
- Calculate the number of moles of $NaOBr_{(s)}$ that would have to be added to 125 mL of 0.160 M HOBr to produce a buffer solution with $[H^+] = 5.00 \times 10^{-9}$ M. Assume that volume change is negligible.

Answers:

- 1(a) First find out how many moles of NaOH were added; molarity (M) = moles / volume (L)
So, moles NaOH = (0.100 M) (0.040 L) = 0.0040 moles
At the equivalence point, every H^+ has been neutralized by an OH^- , so for a monoprotic acid, the number of moles of base added = the initial number of moles of acid
So, 0.0040 moles = initial moles of acid
There was 100. mL of the original acid solution, so we can calculate the molarity of the acid;
molarity (M) = moles / volume (L) molarity (M) = 0.0040 / 0.100 = **0.040 M**
- 1(b) First calculate $[H^+]$ from pH = 4.46; pH = $-\log[H^+]$ $[H^+] = 10^{(-4.46)} = 3.47 \times 10^{-5}$ M
Since $[H^+]$ is very small, HOCl must be a weak acid. Therefore it is almost entirely present in the form of HOCl molecules. Therefore, the concentration of HOCl, $[HOCl] = 0.040$ M (i.e. you can ignore the small amount of dissociation of HOCl into H^+ and OCl^- , which would reduce the equilibrium concentration of HOCl).

Next, write the acid-dissociation equation; $\text{HOCl} \rightleftharpoons \text{H}^+ + \text{OCl}^- \quad K_a = ?$

From this, $K_a = \frac{[\text{H}^+][\text{OCl}^-]}{[\text{HOCl}]}$ so, $K_a = \frac{(3.47 \times 10^{-5} \text{ M})(3.47 \times 10^{-5} \text{ M})}{(0.040 \text{ M})} = \underline{\underline{3.00 \times 10^{-8}}}$

1(c) The percent of molecules ionized is calculated as follows;

% ionized = $\frac{\text{(moles split into H}^+ \text{ and OCl}^-)}{\text{(moles remaining as HOCl)}} \times 100$ % ionized = $\frac{[\text{H}^+]}{[\text{HOCl}]} \times 100$

% ionized = $(3.47 \times 10^{-5} \text{ M}) / (0.040 \text{ M}) \times 100 = \underline{\underline{0.087 \%}}$

1(d) At the equivalence point, all of the HOCl initially present has been converted into OCl⁻ ions. From 1(a) we know there were originally 0.004 moles of HOCl present, so there will be 0.004 moles of OCl⁻ ions. However, after adding 40 mL of base, there is now 140 mL of solution, so the concentration (molarity) of the OCl⁻ ions, $[\text{OCl}^-] = \text{moles} / \text{volume (L)} = (0.004) / (0.14 \text{ L})$
 $[\text{OCl}^-] = \underline{\underline{0.029 \text{ M}}}$

2(a) The solution will be basic. A salt composed of the conjugate of a strong base (Na⁺ is conjugate of NaOH) and the conjugate of a weak acid (CH₃CO₂⁻ is conjugate of CH₃CO₂H) will create a basic solution in water.

Na₂CH₃CO₂ dissociates into Na⁺ and CH₃CO₂⁻. Na⁺ does not react with water, but CH₃CO₂⁻ does;
 $\text{CH}_3\text{CO}_2^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{CO}_2\text{H} + \text{OH}^-$

2(b) The solution will be acidic. A salt composed of the conjugate of a strong acid (Cl⁻ is conjugate of HCl) and the conjugate of a weak base (NH₄⁺ is conjugate of NH₃) will create an acidic solution in water.

NH₄Cl dissociates into NH₄⁺ and Cl⁻. Cl⁻ does not react with water, but NH₄⁺ does;

$\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$

3(a) pH = -log[H⁺] [H⁺] = 10^(-4.95) = 1.1 × 10⁻⁵ M

3(b) $K_a = \frac{[\text{H}^+][\text{OBr}^-]}{[\text{HOBr}]}$ If [H⁺] = 1.8 × 10⁻⁵ M, then [OBr⁻] = 1.8 × 10⁻⁵ M

$$2.3 \times 10^{-9} = (1.8 \times 10^{-5} \text{ M})(1.8 \times 10^{-5} \text{ M}) / [\text{HOBr}]$$

So, [HOBr] = 0.14 M

3(c) (i) Ba(OH)₂ puts 2 OH⁻ into solution for every Ba(OH)₂. At the equivalence point, moles OH⁻ equals the moles of HOBr. For HOBr, molarity = moles / vol (L), 0.146 M = moles / 0.065 L
 So, moles HOBr = 9.49 × 10⁻³

Therefore, moles of OH⁻ at equivalence point = 9.49 × 10⁻³, and moles of Ba(OH)₂ = 4.745 × 10⁻³

For Ba(OH)₂, molarity = moles / vol (L), 0.115 M = 4.745 × 10⁻³ / vol (L)

So, volume = 0.0413 L = 41.3 mL

(ii) The pH is greater than 7. At the equivalence point, the OH⁻ has neutralized all the HOBr, leaving just OBr⁻ behind in solution. We know from the fact that HOBr is a weak acid that OBr⁻ will hydrolyze water to re-form HOBr, leaving OH⁻ in solution – hence the solution is basic.

3(d) A solution with [H⁺] = 5.00 × 10⁻⁹ has a pH = -log (5.00 × 10⁻⁹) = 8.301

And, pK_a = -log K_a = -log (2.3 × 10⁻⁹) = 8.638

Using the Henderson-Hasselbalch equation; pH = pK_a + log ([OBr⁻] / [HOBr])

So, 8.301 = 8.638 + log ([OBr⁻] / (0.160 M))

Which gives; -0.337 = log ([OBr⁻] / (0.160 M)) 10^(-0.337) = [OBr⁻] / (0.160 M)

So, 0.460 = [OBr⁻] / (0.160 M) [OBr⁻] = 0.0736 M

And, [NaOBr] = [OBr⁻] = 0.0736 M

so, molarity = moles / vol (L) gives, 0.0736 M = moles / 0.125 L

Finally, moles NaOBr = (0.0736 M)(0.125 L) = 9.2 × 10⁻³ moles.