True or False

([✓ T] / F) The solubility of a slightly soluble salt can be expressed in units of moles per liter

([✓ T] / F) The solubility product of a slightly soluble salt is independent of the presence of a common ion

(T / [✓ F]) Titrating a strong acid with a strong base can produce a solution with buffer capabilities

([✓ T] / F) **In an acid/base equilibrium, the side that has the acid with the higher pK_a is the side that equilibrium favors.** The higher the pK_a , the lower the K_a. In lecture, we compared K_as and determined that the acid with the larger K_a was more reactive, and would cause the equilibrium to shift to the opposite side. This is essentially saying that equilibrium will favor the side with the comparatively weaker acid.

([✓ T] / F) **The molarity of a monoprotic strong acid is the same as its hydronium concentration.** Since strong acids dissociate completely, each mole of acid will produce one mole of hydronium.

(T / [✓ F]) **HF is a weak acid because the H-F bond is weak.** Fluorine is small and electronegative, making the H-F bond short and strong and not easily dissociable.

(T / [✓ F]) **In a 25ºC solution, it is impossible to have a negative pH.** Highly concentrated acids at 25ºC can have negative pH. For example, having a 2 M solution of HCl. -log(2) = -0.30.

1. A buffer contains 0.10 mol of benzoic acid and 0.13 mol of sodium benzoate in 1.00 L of solution. K_a of benzoic acid = 6.3×10^{-5}

   a. What is the pH of this buffer? (1)

$$pH = pK_a + \log\left(\frac{[base]}{[acid]}\right)$$

$$pH = -\log(6.3 * 10^{-5}) + \log\left(\frac{0.13\ M}{0.10\ M}\right) = 4.31$$

   b. What is the pH of the buffer after the addition of 0.03 mol of solid KOH? (3)

$$C_6H_5COOH(aq) + OH^-(aq) \leftrightarrow C_6H_5COO^-(aq) + H_2O(l)$$

	$[C_6H_5COOH]$	$[OH^-]$	$[C_6H_5COO^-]$
I	0.10	0	0.13
ADDITION OF OH-	0	+0.03	0
C	-0.03	-0.03	+0.03
E	0.07	0	0.16

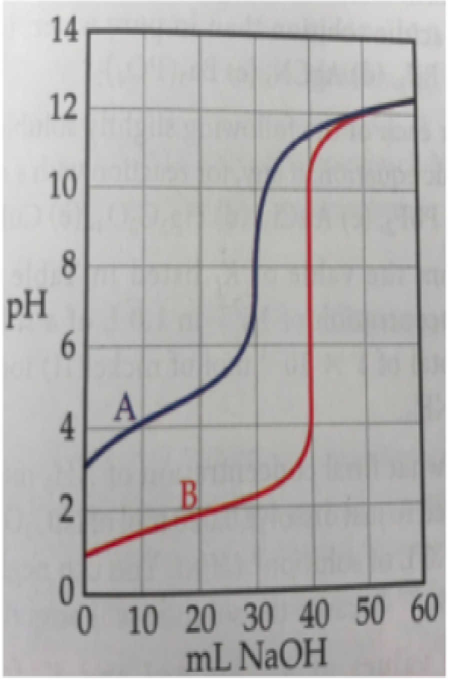
$$pH = pK_a + \log\left(\frac{[base]}{[acid]}\right)$$

$$pH = -\log(6.3 * 10^{-5}) + \log\left(\frac{0.16\ M}{0.07\ M}\right) = 4.56$$

2. You titrate 182 mL of 0.45 M acetic acid with KOH solution, and you see that you used 211 mL of the KOH solution to reach the equivalence point. What is the original KOH solution's pH? (3)

To reach the equivalence point, all moles of acetic acid must have been consumed. Since there were 0.182 L * 0.45 M = 0.0819 moles of acetic acid, there must have been the same number of moles of OH- in the KOH solution. Since there must have been 0.0819 moles of KOH in 211 mL of solution, the molarity and [OH-] of the original KOH solution is 2.41. Hence, pOH is 0.4 and pH is 13.6.

3 This graph shows the titration curves for two monoprotic acids.



a. Which curve is that of a strong acid? (1)

B

b. What is the approximate pH at the equivalence point of each titration? (1)

A – 8
B – 7

c. The same volume of each acid was titrated with 0.100 M base. Which acid is more concentrated? (1)

B, since it required a greater volume of NaOH solution to reach the equivalence point.

d. Estimate the pK_a of the weak acid. (1)

pH = pK_a at the ½ equivalence point. Since 30 mL reaches the equivalence point in curve A, the weak acid, pH is roughly 4.5 at 15 mL, making pKa 4.5

4. Calculate the concentration of an aqueous solution of NaOH that has a pH of 11.50. (2)

If pH is 11.50, pOH is 14 – 11.50 = 2.5

[OH-] = $10^{-2.5}$ = 3.2×10^{-3} M

Since NaOH is a strong base and dissociates completely in water, the $[OH^-]$ functions as a proxy for [NaOH], which is 3.2×10^{-3} M

5. Provide an example of a neutral salt and explain why by comparing relative K_a and K_b values. No actual K_a or K_b values needed. (3)

NaCl. It is neutral because it can be thought of as being made from the acid base reaction between NaOH and HCl, which are a strong base and acid, respectively. This means their K_b and K_a are, respectively, very high. Their conjugate acid and base will