

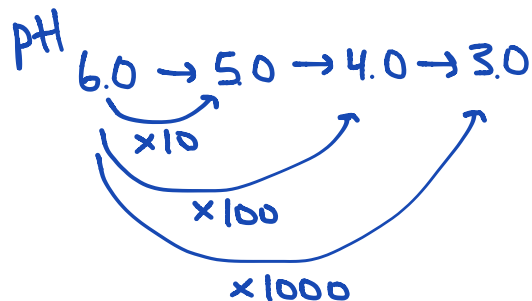
# Acid Base Part 2 Review Package

Name: Key  
 Date:  
 Block:

## I. Multiple Choice

1. Consider the following data:

Solution	Initial pH	Final pH
1	1.0	4.0
2	2.0	6.0
3	6.0	3.0
4	9.0	3.0



In which solution has the  $[H_3O^+]$  increased 1000 times?

- A. 1
- B. 2

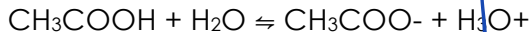
- C. 3
- D. 4

2. A solution is amber with neutral red and colourless with phenolphthalein. The approximate pH of the solution is:

- A. 4
- B. 6

- C. 8
- D. 10

3. Consider the following equilibrium for an acetic acid buffer solution:



When a small amount of acid is added to this system, and equilibrium is reestablished,

- A.  $[CH_3COO^-]$  and pH have both increased
- B.  $[CH_3COOH]$  and pH have both decreased
- C.  $[CH_3COO^-]$  has decreased and pH remains relatively constant
- D.  $[CH_3COOH]$  has decreased and pH remains relatively constant.

4. The stoichiometric point of a titration is reached when 35.50 mL 0.40 M HBr is added to a 25.00 mL sample of LiOH. The original  $[LiOH]$  is:

- A. 0.014 M
- B. 0.024 M

- C. 0.28 M
- D. 0.57 M

$$0.03550 \text{ L HBr} \times \frac{0.40 \text{ mol HBr}}{1 \text{ L HBr}} \times \frac{1 \text{ mol LiOH}}{1 \text{ mol HBr}} \times \frac{1}{0.02550 \text{ L}} = 0.56 \text{ M LiOH}$$

5. A buffer solution can be prepared by mixing equal numbers of moles of:

- A. ~~NH<sub>4</sub>Cl and HCl~~ SA
- B. ~~NaCl and NaOH~~ SB
- C. Na<sub>2</sub>CO<sub>3</sub> and NaHCO<sub>3</sub> ✓
- D. ~~NaCH<sub>3</sub>COO and NaOH~~ SB

Weak acid  
 &  
Conjugate base

not conjugates!

6. Which of the following, when dissolved in water, produces a **basic** solution?

- A. ~~KCl~~  
 B. ~~NaClO<sub>4</sub>~~  
 C. ~~Na<sub>2</sub>CO<sub>3</sub>~~  
 D. ~~NH<sub>4</sub>NO<sub>3</sub>~~
- Spec  
 Spec  
 WA Spec

- C. ~~Na<sub>2</sub>CO<sub>3</sub>~~  
 D. ~~NH<sub>4</sub>NO<sub>3</sub>~~
- Spec  
 WA Spec

7. The net ionic equation for the hydrolysis of Na<sub>2</sub>CO<sub>3</sub> is:

- A. ~~H<sub>2</sub>O + Na<sup>+</sup> ⇌ NaOH + H<sup>+</sup>~~  
 B. ~~H<sub>2</sub>O + 2Na<sup>+</sup> ⇌ Na<sub>2</sub>O + 2H<sup>+</sup>~~  
 C. ~~H<sub>2</sub>O + CO<sub>3</sub><sup>2-</sup> ⇌ H<sub>2</sub>CO<sub>3</sub> + O<sup>2-</sup>~~  
 D. ~~H<sub>2</sub>O + CO<sub>3</sub><sup>2-</sup> ⇌ HCO<sub>3</sub><sup>-</sup> + OH<sup>-</sup>~~
- no spec  
 spec

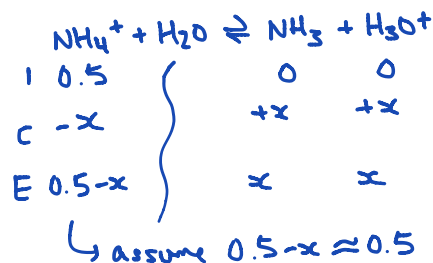
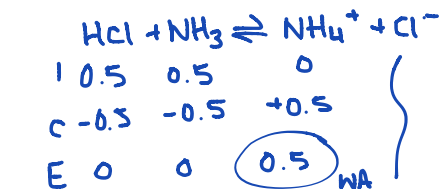
8. When 1.0 M NH<sub>3</sub> is titrated with 1.0 M HCl, the most suitable indicator is:

- A. methyl violet.  
 B. indigo carmine.

- C. phenolphthalein.  
 D. bromcresol green.
- $5.6 \times 10^{-10} = \frac{x^2}{0.5}$   
 $x = 1.7 \times 10^{-5} = [\text{H}_3\text{O}^+]$   
 pH = 4.78

9. All buffer solutions are able to

- A. maintain pH at 7.00.  
 B. neutralize acidic solutions only.  
 C. maintain a relatively constant pH.



10. The most appropriate indicator for the titration of 0.50 M CH<sub>3</sub>COOH with 0.50 M NaOH is (ICE table calculation necessary!)

- A. methyl red.  
 B. indigo carmine.  
 C. phenolphthalein.  
 D. bromcresol green.

(See # 8 for how to calculate)

11. An indicator is useful for a titration when its K<sub>a</sub> is close to

- A. K<sub>w</sub> / K<sub>b</sub>  
 B. the pH at the equivalence point.  
 C. the [H<sup>+</sup>] at the equivalence point.  
 D. the concentration of the acid form of the indicator.

$$K_a = \frac{[\text{Ind}^-][\text{H}_3\text{O}^+]}{[\text{HInd}]}$$

12. In order to change the pH of a solution from 2.0 to 4.0 the [H<sub>3</sub>O<sup>+</sup>] must:

- A. decrease by a factor of 100  
 B. decrease by a factor of 2  
 C. increase by a factor of 2  
 D. increase by a factor of 100

more basic

13. Which of the following represents the predominant reaction between HCO<sub>3</sub><sup>-</sup> and water?

- A. ~~2HCO<sub>3</sub><sup>-</sup> + H<sub>2</sub>O ⇌ H<sub>3</sub>O<sup>+</sup> + CO<sub>3</sub><sup>2-</sup> + OH<sup>-</sup> + CO<sub>2</sub>~~  
 B. ~~HCO<sub>3</sub><sup>-</sup> + H<sub>2</sub>O ⇌ H<sub>3</sub>O<sup>+</sup> + CO<sub>3</sub><sup>2-</sup>~~  
 C. ~~HCO<sub>3</sub><sup>-</sup> + H<sub>2</sub>O ⇌ H<sub>2</sub>CO<sub>3</sub> + OH<sup>-</sup>~~  
 D. ~~2HCO<sub>3</sub><sup>-</sup> ⇌ H<sub>2</sub>O + 2CO<sub>2</sub>~~

$K_a = 5.6 \times 10^{-11}$

$K_b = \frac{K_w}{4.7 \times 10^{-7}} = 2.1 \times 10^{-8}$

14. Which of the following applies at the transition point for all indicators, HInd?

- A.  $[HInd] = [Ind^-]$
- B.  $[H_3O^+] = [OH^-]$
- C.  $[Ind^-] = [H_3O^+]$
- D.  $[HInd] = [H_3O^+]$

$$K_a = \frac{[Ind^-][H_3O^+]}{[HInd]}$$

15. Which of the following 1.0M solutions would have a pH greater than 7.00?

- A. HCN <sup>WA</sup>
- B. ~~NH<sub>4</sub>Cl~~ <sup>WA</sup>
- C. ~~KNO<sub>3</sub>~~
- D. NaCH<sub>3</sub>COO <sup>WB</sup>

basic ↘

16. What is the pH of the solution formed when 0.040 mol NaOH<sub>(s)</sub> is added to 2.00 L of 0.020M HCl?

- A. 7.00
- B. 1.70
- C. 1.40
- D. 0.00

$$SB [NaOH] = \frac{0.040 \text{ mol}}{2.00 \text{ L}} = 0.020 \text{ M}$$

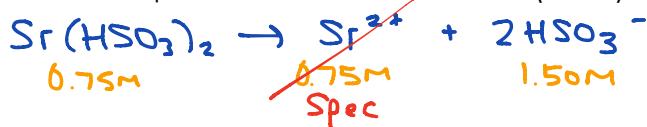
$$SA [HCl] = 0.020 \text{ M} \quad \text{equal!}$$

17. A buffer solution may contain equal moles of

- A. strong acid and its conjugate base.
- B. weak acid and strong base.
- C. weak acid and its conjugate base.
- D. strong acid and strong base.

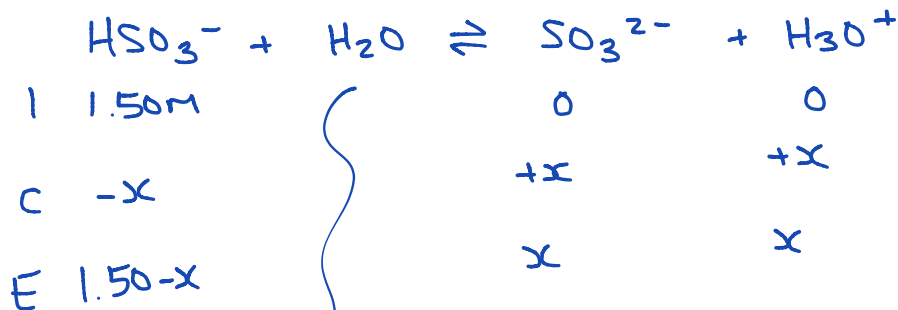
## I. Problems

1. Determine the pH of a 0.75M solution of  $\text{Sr}(\text{HSO}_3)_2$ .



$$K_a = 1.0 \times 10^{-7}$$

$$K_b = \frac{1.00 \times 10^{-14}}{1.5 \times 10^{-2}} = 6.7 \times 10^{-13}$$



↳ assume  $1.50 - x \approx 1.50$

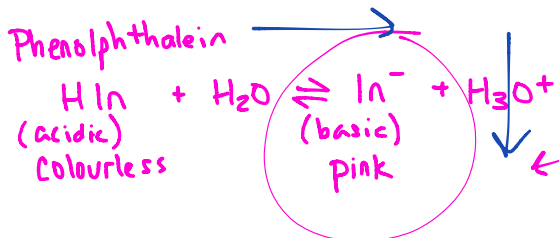
$$1.0 \times 10^{-7} = \frac{x^2}{1.50}$$

$$x = \sqrt{(1.0 \times 10^{-7})(1.50)}$$

$$= 3.9 \times 10^{-4} \text{ M} = [\text{H}_3\text{O}^+]$$

pH = 3.41

2. Consider the salt sodium oxalate ( $\text{Na}_2\text{C}_2\text{O}_4$ ). When a few drops 1.0M solution of sodium oxalate is added to phenolphthalein indicator the solution turns pink. Write a hydrolysis equation and an equilibrium expression to explain the shift that caused the indicator to change colour.



Adding  $\text{OH}^-$  results in the equilibrium to shift right, causing  $[\text{In}^-] \uparrow$  and expressing the pink colour

3. a) Calculate the pH of the solution produced when 9.00 mL of 0.200 M NaOH has been added to 20.0 mL of 0.200 M HCOOH.

WA

TB

$$[\text{HCOOH}]$$

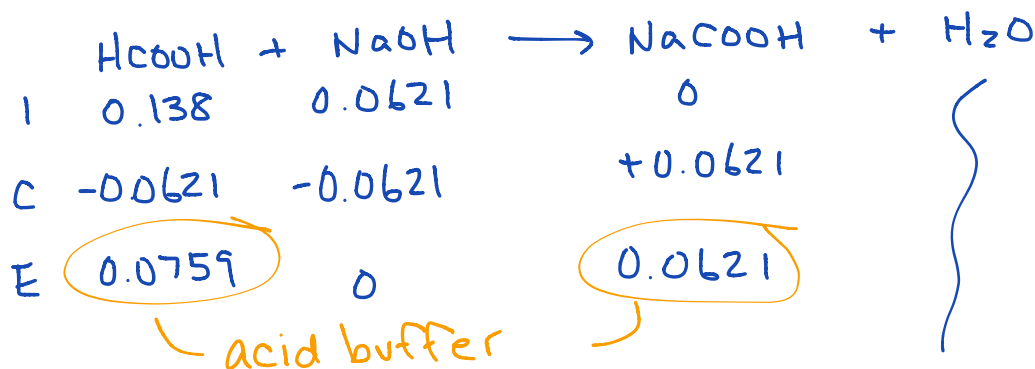
$$(0.200)(20.0) = C_2(29.0)$$

$$C_2 = 0.138\text{M}$$

$$[\text{NaOH}]$$

$$(0.200)(9.00) = C_2(29.0)$$

$$C_2 = 0.0621\text{M}$$



$$[\text{H}_3\text{O}^+] = K_a \left( \frac{[\text{acid}]}{[\text{base}]} \right)$$

$$= 1.8 \times 10^{-4} \left( \frac{0.0759}{0.0621} \right) = 2.2 \times 10^{-4}$$

$\text{pH} = 3.67$

b) How many mL of NaOH have been added at equivalence point?

$$0.0200 \text{ L HCOOH} \times \frac{0.200 \text{ mol HCOOH}}{1 \text{ L}} \times \frac{1 \text{ mol NaOH}}{1 \text{ mol HCOOH}} \times \frac{1 \text{ L NaOH}}{0.200 \text{ mol}}$$

$$= 0.0200 \text{ L NaOH} = \boxed{20.0 \text{ mL NaOH}}$$

c) Calculate the pH at the equivalence point of this titration.

$$[\text{HCOOH}]$$

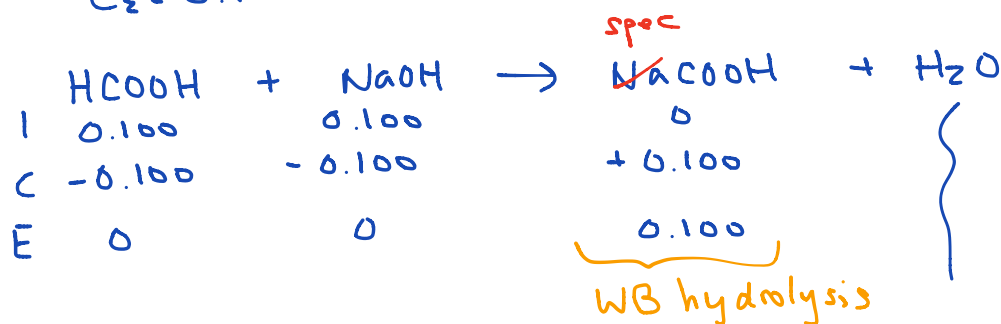
$$(0.200)(20.0) = C_2(40.0)$$

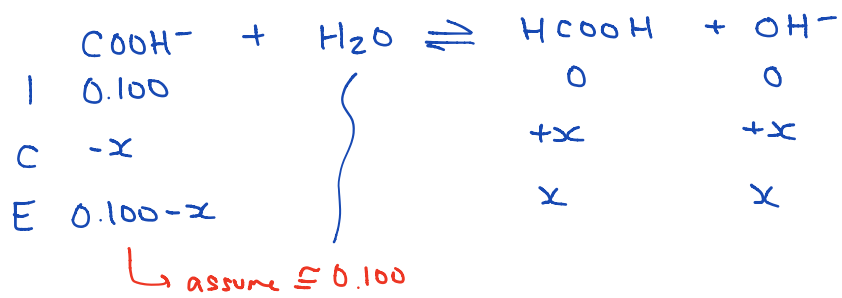
$$C_2 = 0.100\text{M}$$

$$[\text{NaOH}]$$

$$(0.200)(20.0) = C_2(40.0)$$

$$C_2 = 0.100\text{M}$$





$$K_b = \frac{K_w}{1.8 \times 10^{-4}} = 5.6 \times 10^{-11} = \frac{x^2}{0.100}$$

$$\begin{aligned}
 x &= \sqrt{(5.6 \times 10^{-11})(0.100)} \\
 &= 2.4 \times 10^{-6} = [\text{OH}^-]
 \end{aligned}$$

$$\text{pOH} = 5.63$$

$$\boxed{\text{pH} = 8.37}$$