

AP Chemistry: Acids & Bases
Honors Review

Fill in the blank. (30 points)

Directions: Read each question carefully then write the correct answer to the question in the blank provided.

1. A Lewis base donates an electron pair.
2. A Bronsted acid donates a hydrogen ion.
3. A(n) acid has a sour taste.
4. A(n) base is slippery to touch.
5. An acid with 3 hydrogen ions is called a triprotic acid.
6. A substance that is able to act as an acid or a base like water is called amphoteric.
7. A weak acid only ionizes slightly in water.
8. A strong acid is a strong electrolyte.
9. When an acid reacts with a base the reaction produces salt and water.
10. The pH of a base is greater than 7.

MATCHING. (20 points)

	Column A	Column B
<u>A</u> 11.	A weak monoprotic acid	A. $\text{HC}_2\text{H}_3\text{O}_2$
<u>E</u> 12.	A weak diprotic acid	B. HCl
<u>C</u> 13.	A weak triprotic acid	C. H_3PO_3
<u>B</u> 14.	A strong monoprotic acid	D. H_2SO_4
<u>D</u> 15.	A strong diprotic acid	E. H_2SO_3

PROBLEMS. (30 points)

Directions: Solve the following problems. Show ALL of your work.

16. The hydroxide concentration of a solution is
- 4.5×10^{-9}
- M.

a. What is the pH?

$$\text{pH} = 14 - 8.3 = 5.7$$

b. What is the pOH?

$$\text{pOH} = -\log(4.5 \times 10^{-9}) = 8.35$$

↙ 2 sig figs, so 2 decimal places ↘

c. Acid or Base? acid

17. The pH of a solution is 10.00.

a. What is the pOH?

$$\text{pOH} = 4.00$$

b. What is the $[\text{OH}^-]$?

$$10^{-4} = 1.0 \times 10^{-4}$$

c. Acid or Base? base

18. The pOH of a solution is 3.70.

a. What is the pH?

$$\text{pH} = 14 - 3.70 = 10.3$$

b. What is the $[\text{OH}^-]$?

$$10^{-3.70} = 2.0 \times 10^{-4} \text{ M}$$

c. Acid or Base? base

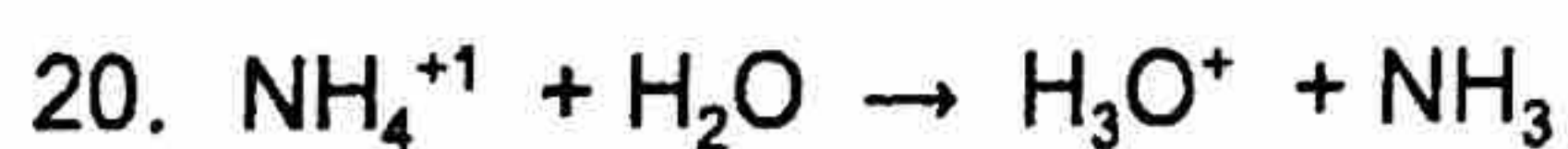
IDENTIFY. (20 points)

Directions: For the following acid-base reactions, identify the following:

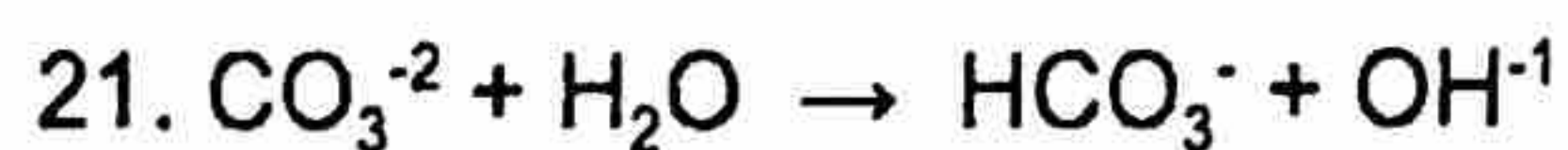
Acid (A), Base (B), Conjugate Acid (CA), Conjugate Base (CB)



A B CA CB



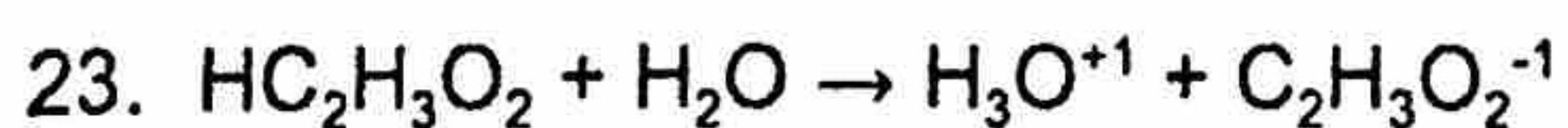
A B CA CB



B A CA CB



B A CA CB



A B CA CB

Conjugate Pairs Practice Questions

1. Identify the acid, base, conjugate acid and conjugate base for each of the following.

- a) $\underset{A}{\text{HClO}_4}(\text{aq}) + \underset{B}{\text{H}_2\text{O}}(\text{l}) \rightleftharpoons \underset{C}{\text{H}_3\text{O}^+}(\text{aq}) + \underset{D}{\text{ClO}_4^-}(\text{aq})$
- b) $\underset{A}{\text{H}_2\text{SO}_3}(\text{aq}) + \underset{B}{\text{H}_2\text{O}}(\text{l}) \rightleftharpoons \underset{C}{\text{H}_3\text{O}^+}(\text{aq}) + \underset{D}{\text{HSO}_3^-}(\text{aq})$
- c) $\underset{A}{\text{HC}_2\text{H}_3\text{O}_2}(\text{aq}) + \underset{B}{\text{H}_2\text{O}}(\text{l}) \rightleftharpoons \underset{C}{\text{H}_3\text{O}^+}(\text{aq}) + \underset{D}{\text{C}_2\text{H}_3\text{O}_2^-}(\text{aq})$
- d) $\underset{A}{\text{H}_2\text{S}}(\text{g}) + \underset{B}{\text{H}_2\text{O}}(\text{l}) \rightleftharpoons \underset{C}{\text{H}_3\text{O}^+}(\text{aq}) + \underset{D}{\text{HS}^-}(\text{aq})$
- e) $\underset{A}{\text{HSO}_3^-}(\text{aq}) + \underset{B}{\text{H}_2\text{O}}(\text{l}) \rightleftharpoons \underset{C}{\text{H}_3\text{O}^+}(\text{aq}) + \underset{D}{\text{SO}_3^{2-}}(\text{aq})$
- f) $\underset{B}{\text{NH}_3}(\text{g}) + \underset{A}{\text{H}_2\text{O}}(\text{l}) \rightleftharpoons \underset{C}{\text{NH}_4^+}(\text{aq}) + \underset{D}{\text{OH}^-}(\text{aq})$
- g) $\underset{A}{\text{HF}}(\text{aq}) + \underset{B}{\text{HSO}_3^-}(\text{aq}) \rightleftharpoons \underset{C}{\text{F}^-}(\text{aq}) + \underset{D}{\text{H}_2\text{SO}_3}(\text{aq})$
- h) $\underset{A}{\text{HNO}_2}(\text{aq}) + \underset{B}{\text{HS}^-}(\text{aq}) \rightleftharpoons \underset{C}{\text{NO}_2^-}(\text{aq}) + \underset{D}{\text{H}_2\text{S}}(\text{aq})$

2. Calculate pH and pOH for the following solutions AND Identify as acidic or basic:

- A a) $[\text{H}^+] = 1.0 \times 10^{-5} \text{ M}$ pH = 5.00 pOH = 9.00 b) $[\text{OH}^-] = 3.0 \times 10^{-8} \text{ M}$ pH = 6.5 pOH = 7.5 A
- A c) $[\text{H}^+] = 2.5 \times 10^{-2} \text{ M}$ pH = 1.60 pOH = 12.40 d) $[\text{OH}^-] = 7.5 \times 10^{-3} \text{ M}$ pH = 11.9 pOH = 2.1 B
- B e) $[\text{H}^+] = 1.2 \times 10^{-14} \text{ M}$ pH = 13.92 pOH = 0.08 f) $[\text{H}^+] = 0.600 \text{ M}$ pH = 0.223 pOH = 13.8 A

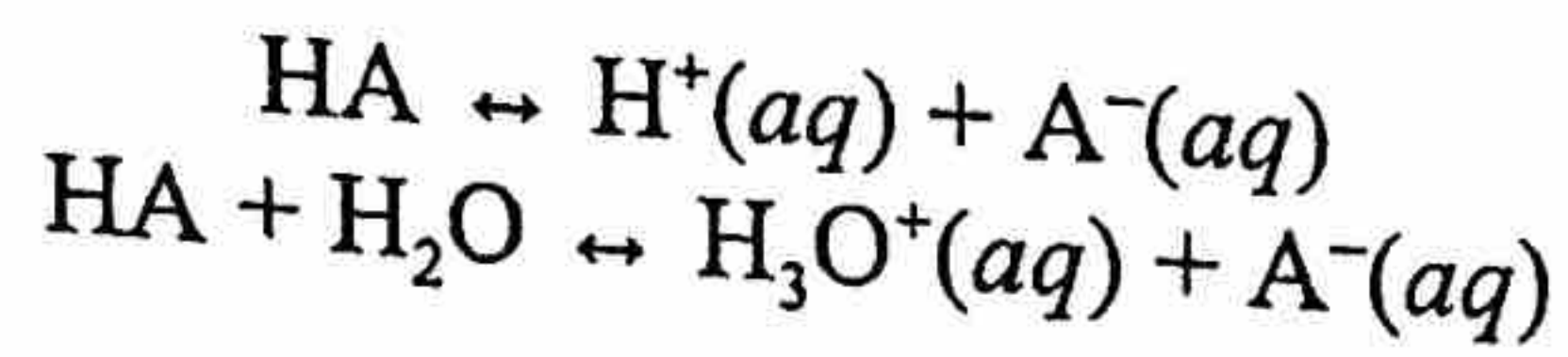
3. Calculate $[\text{H}^+]$ and $[\text{OH}^-]$ for the following AND Identify as acidic or basic:

- A a) pH = 3.0 $[\text{H}^+] = 0.001 \text{ M}$ $[\text{OH}^-] = 1 \times 10^{-11} \text{ M}$
- B b) pOH = 2.60 $[\text{H}^+] = 4.0 \times 10^{-12} \text{ M}$ $[\text{OH}^-] = 0.0025 \text{ M}$
- B c) pOH = 5.63 $[\text{H}^+] = 4.3 \times 10^{-9} \text{ M}$ $[\text{OH}^-] = 2.3 \times 10^{-6} \text{ M}$
- B d) pH = 7.51 $[\text{H}^+] = 3.1 \times 10^{-8} \text{ M}$ $[\text{OH}^-] = 3.2 \times 10^{-7} \text{ M}$
- B e) pOH = 1.13 $[\text{H}^+] = 1.4 \times 10^{-13} \text{ M}$ $[\text{OH}^-] = 0.074 \text{ M}$
- A f) pH = 0.03 $[\text{H}^+] = 0.93 \text{ M}$ $[\text{OH}^-] = 1.1 \times 10^{-14} \text{ M}$

Chapter 14: Weak Acids & Bases
Calculating pH and $[H^+]$ for Weak Acids and Bases

Weak acids and bases are usually *less than 5% ionized*. The equilibrium constant for a weak acid equilibrium is the *acid ionization constant* K_a , and for a weak base equilibrium is the *base ionization constant* K_b .

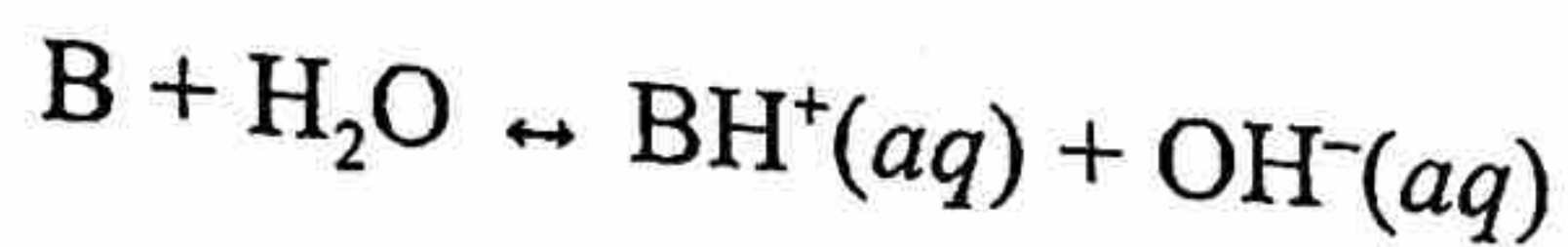
A typical monoprotic weak acid equilibrium can be written in two forms, the second of which emphasizes the Brønsted acid-base nature of the reaction:



In Eq(9) the Brønsted acid HA donates a proton H^+ to the Brønsted base H_2O to form H_3O^+ and the conjugate base A^- . The acid ionization constant (using the second form) is

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

A typical weak base equilibrium is



In Eq(10) the Brønsted base B accepts a proton H^+ from the Brønsted base H_2O to form the conjugate acid BH^+ and OH^- . The base ionization constant is

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

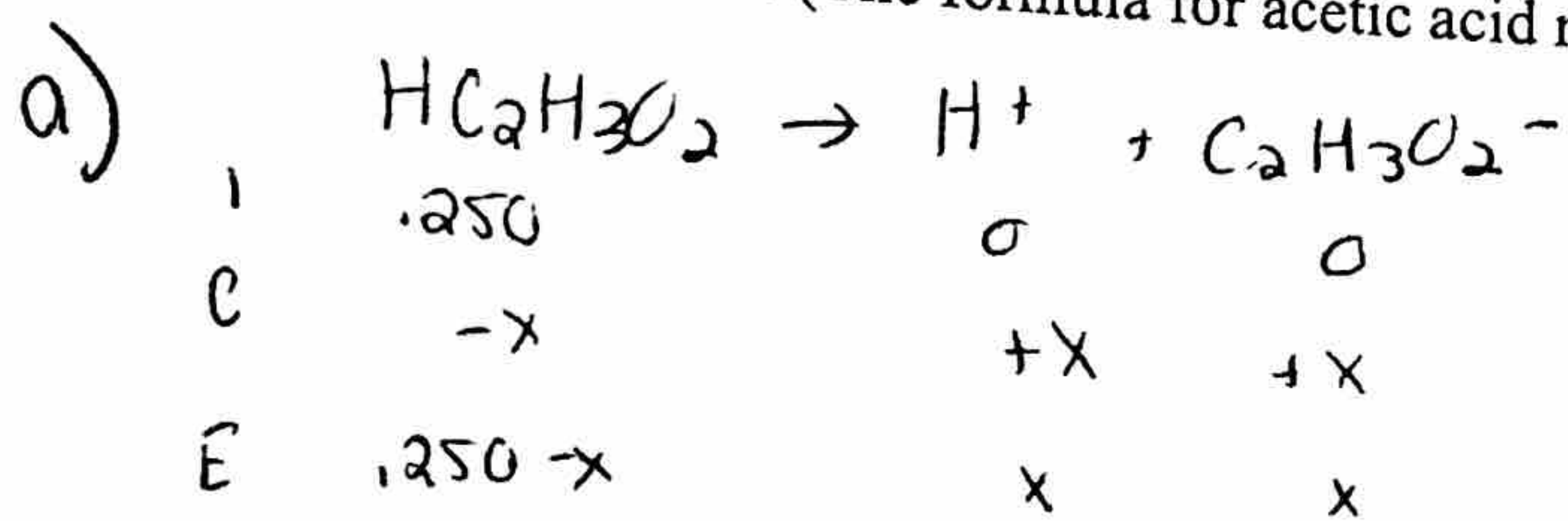
Percent Ionization:

$$\% \text{ Ionization of Weak Acid} = \frac{[H_3O^+]}{[HA]_0} \times 100$$

$$\% \text{ Ionization of Weak Base} = \frac{[OH^-]}{[B]_0} \times 100$$

Example 1:

Calculate (a) the pH and (b) the percent ionization of a 0.250 M $\text{HC}_2\text{H}_3\text{O}_2$ solution. $K_a(\text{HC}_2\text{H}_3\text{O}_2) = 1.8 \times 10^{-5}$. (The formula for acetic acid may also be written as CH_3COOH .)



$$1.8 \times 10^{-5} = \frac{x^2}{.25}$$

$$x = .00212 \text{ M}$$

$$\text{pH} = -\log(.00212)$$

$$\text{pH} = 2.67$$

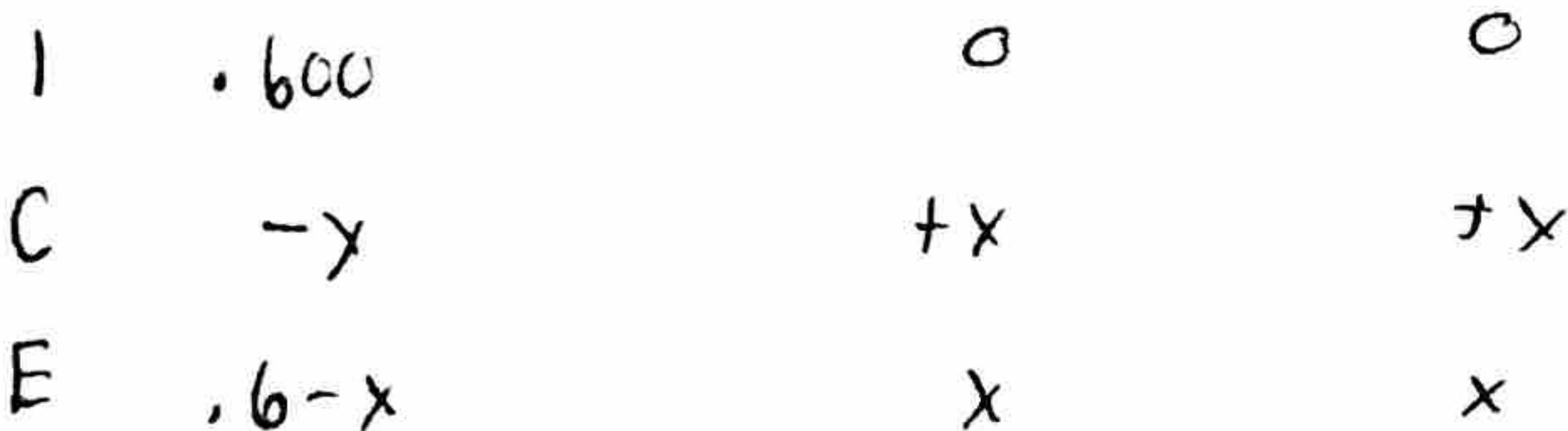
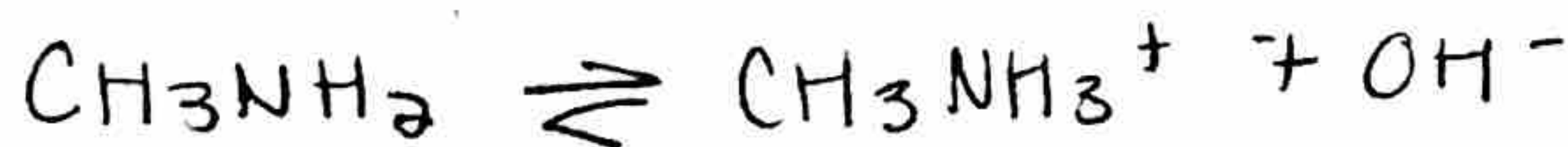
b)

$$\frac{.00212}{.250} \times 100$$
$$.85\%$$

Answer: (a) $x = 2.12 \times 10^{-3} = [\text{H}^+]$. $\text{pH} = 2.67$. (b) 0.085%

Practice Problem 1.

Calculate the pH and % ionization of a 0.600 M solution of methylamine (CH_3NH_2). $K_b = 4.4 \times 10^{-4}$.
HINT: Methylamine is a weak base. First write the equation for the reaction following. Then fill out the equilibrium table below.



$$4.4 \times 10^{-4} = \frac{x^2}{.6}$$

$$x = .0162 \text{ M}$$

$$[\text{OH}^-] = .0162 \text{ M}$$

pOH

$$\text{pOH} = 1.79$$
$$\text{pH} = 12.2$$

$$\% \text{ ionization} = \frac{.0162}{.6} \times 100$$

$$2.7\%$$

Answer: $x = 1.62 \times 10^{-2} = [\text{OH}^-]$, and $\text{pOH} = 1.79$.

Example Problem 2.

The pH of a 0.10 M solution of a weak base is 9.67. What is the K_b of the base?

$$pOH = 4.33$$

$$[OH^-] = 4.677 \times 10^{-5} M$$

$$K_b = \frac{(4.677 \times 10^{-5})^2}{.1}$$

$$K_b = 2.2 \times 10^{-8}$$

Answer: $K_b = 2.2 \times 10^{-8}$

Practice Problem 2.

The pH of a 0.012M solution of a weak acid is 5.66. What is the K_a of the acid?

$$[H^+] = 2.19 \times 10^{-6}$$

$$K_a = \frac{(2.19 \times 10^{-6})^2}{.012}$$

$$K_a = 3.99 \times 10^{-10}$$