

### More Acid/Base

This worksheet should help you identify how we can determine the pH of salt solutions in the laboratory. It is intended for you to work through it in order. (Don't skip ahead.)

Two salt solutions are made:

- A: 4.68 g of NaBrO is added to 300.0 mL of pure H<sub>2</sub>O
- B: 2.94 g of NaClO is added to 300.0 mL of pure H<sub>2</sub>O

What are the molarities of the two solutions?

$$4.68 \text{ g NaBrO} \left| \begin{array}{l} 1 \text{ mol NaBrO} \\ 118.8 \text{ g NaBrO} \end{array} \right. = \frac{0.03936 \text{ mol}}{0.300 \text{ L}} = 0.131 \text{ M NaBrO}$$

$$2.94 \text{ g NaClO} \left| \begin{array}{l} 1 \text{ mol NaClO} \\ 74.49 \text{ g NaClO} \end{array} \right. = \frac{0.03949 \text{ mol}}{0.300 \text{ L}} = 0.132 \text{ M NaClO}$$

Will the solutions be acidic or basic?

Identify each ion in the two salts.



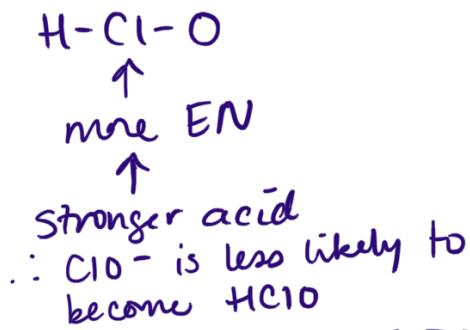
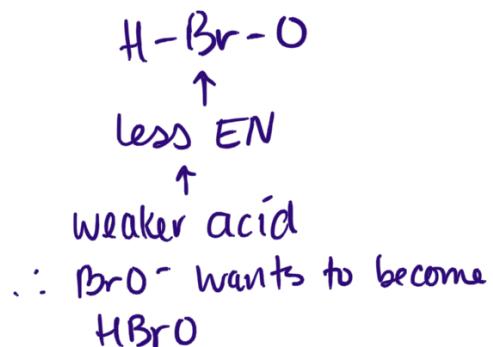
Write out the chemical equation of each ion in water.



Based on these chemical equations, will the solutions be acidic or basic?

Since both  $\text{BrO}^-$  and  $\text{ClO}^-$  are conjugate bases, they will pull  $\text{H}^+$  from water generating weak acids. The byproduct will be  $\text{OH}^-$ , thus both solutions would be acidic. basic  
 $(\text{Na}^+ \text{ does not react w/ water in either case, no effect on pH})$

Based on your decision of acidic or basic, now compare the two solutions. Hypothesize which is stronger? Draw molecular structures to support your statements.



Since HBrO is weaker than HClO the HBrO is more favored. More BrO<sup>-</sup> will go to HBrO than ClO<sup>-</sup> to HClO. so more OH<sup>-</sup> in BrO<sup>-</sup> case.

Consider that a K<sub>b</sub> of  $5.0 \times 10^{-6}$  is associated with salt A, while a K<sub>b</sub> of  $3.4 \times 10^{-7}$  is associated with salt B.

What is the pH of solution A?

$$K_b = \frac{[\text{HBrO}][\text{OH}^-]}{[\text{BrO}^-]}$$

	BrO <sup>-</sup>	OH <sup>-</sup>	HBrO
I	0.131	0	0
C	-x	+x	+x
E	0.131-x	x	x

$$5.0 \times 10^{-6} = \frac{(x)(x)}{0.131-x}$$

assume small x

$$x = 8.09 \times 10^{-4} = [\text{OH}^-]$$

$$\text{pOH} = -\log(8.09 \times 10^{-4})$$

$$\text{pOH} = 3.1$$

$$\text{pH} = 10.9$$

What is the pH of solution B?

$$K_b = \frac{[\text{HClO}][\text{OH}^-]}{[\text{ClO}^-]}$$

	ClO <sup>-</sup>	OH <sup>-</sup>	HClO
I	0.132	0	0
C	-x	+x	+x
E	0.132-x	x	x

$$3.4 \times 10^{-7} = \frac{(x)(x)}{0.132-x}$$

assume small x

$$x = 2.12 \times 10^{-4} = [\text{OH}^-]$$

$$\text{pOH} = 3.67$$

$$\text{pH} = 10.3$$

Do the two calculated pH values match your prediction of acidic or basic?

Yes! Both salts are basic  $\text{pH} > 7$

Do the two calculated pH values match your prediction of which salt was stronger?

Yes! The pH of the NaBrO solution is higher than the NaClO solution.

Let's try a few more example! Using the problem-solving method from above determine the pH of the three solutions below...you will need to look up  $K_a/K_b$  information.

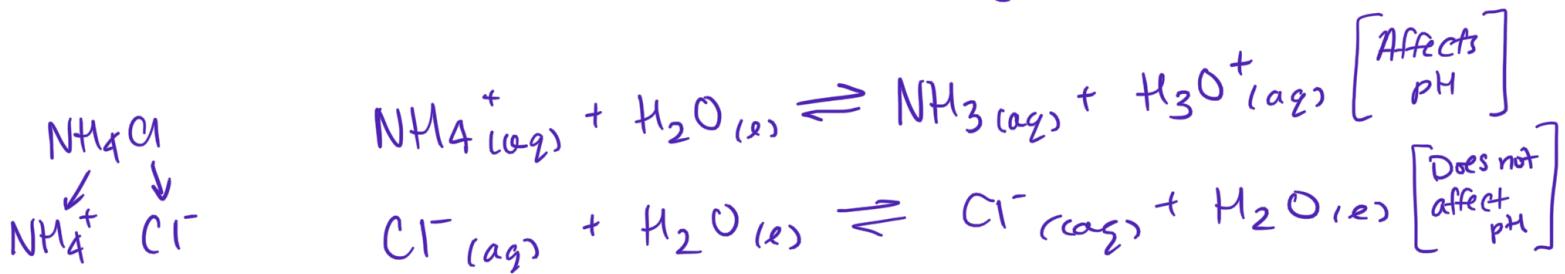
X. 2.43 g of NH<sub>4</sub>Cl in 300.0 mL of pure water.

Y. 1.54 g of KCl in 300.0 mL of pure water.

Z. 4.01 g of NH<sub>4</sub>HCO<sub>3</sub> in 300.0 mL of pure water.

SALT X

$$2.43 \text{ g NH}_4\text{Cl} \left| \begin{array}{l} 1 \text{ mole NH}_4\text{Cl} \\ 53.491 \text{ g NH}_4\text{Cl} \end{array} \right. = \frac{0.0454 \text{ moles NH}_4\text{Cl}}{0.300 \text{ L}} = 0.151 \text{ M}$$



$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$$

	NH <sub>4</sub> <sup>+</sup>	NH <sub>3</sub>	H <sub>3</sub> O <sup>+</sup>
I	0.151	0	0
C	-x	x	x
E	0.151-x	x	x

$$5.56 \times 10^{-10} = \frac{x^2}{0.151 - x} \quad \begin{array}{l} \text{ignore } x \\ K_a < 10^{-5} \end{array}$$

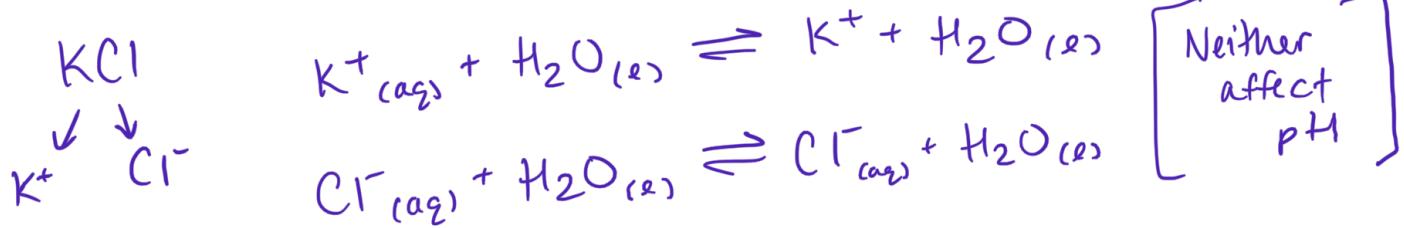
$$8.396 \times 10^{-11} = x^2$$

$$9.16 \times 10^{-6} = x = [\text{H}_3\text{O}^+]$$

$$\begin{aligned} \text{pH} &= -\log(9.16 \times 10^{-6}) \\ &= 5.04 \end{aligned}$$

SALT Y

$$1.54 \text{ g KCl} \left| \begin{array}{l} 1 \text{ mole KCl} \\ 74.55 \text{ g KCl} \end{array} \right. = \frac{0.0206 \text{ moles KCl}}{0.3 \text{ L}} = 0.0687 \text{ M}$$

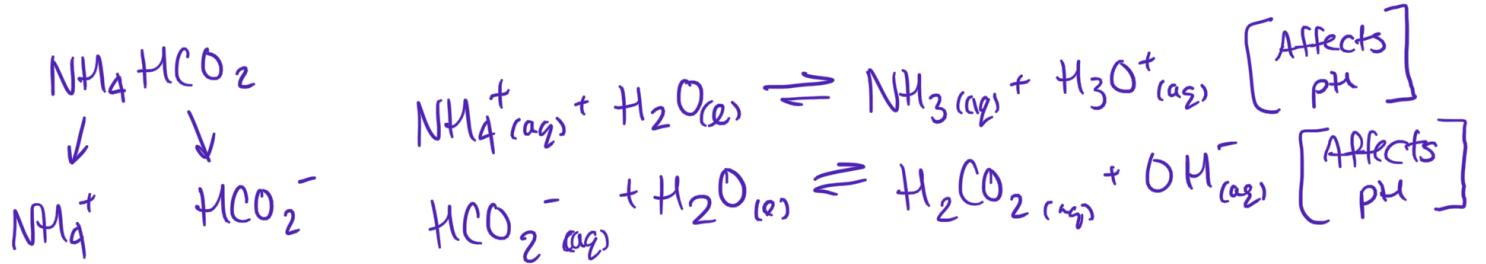


$$\text{pH} = 7$$

## SALT 2

$$4.01 \text{ g } \text{NH}_4\text{HCO}_3 \quad | \quad \begin{array}{l} 1 \text{ mole } \text{NH}_4\text{HCO}_3 = \frac{0.0636 \text{ mole } \text{NH}_4\text{HCO}_3}{0.300 \text{ L}} \\ 63.0559 \text{ g } \text{NH}_4\text{HCO}_3 \end{array}$$

$$= 0.212 \text{ M } \text{NH}_4\text{HCO}_3$$



$$K_a \text{ NH}_4^+ \quad 5.56 \times 10^{-10}$$

$$K_b \text{ HCO}_3^- \quad K_b = \frac{K_w}{K_a} \quad K_b = \frac{1.0 \times 10^{-14}}{2.1 \times 10^{-4}} = 4.76 \times 10^{-11}$$

formic  
acid

$$K_a \text{ NH}_4^+ > K_b \text{ HCO}_3^- \quad \therefore \text{solution pH} < 7$$

(For this class a salt that acts as an acid and a base, you only need to state pH generally, not exactly)