

3) Consider 500.0 mL of a 0.200 M aqueous solution of hypobromous acid at T = 25. °C.

a) What is the pH of the above solution? Note that for HOBr, $K_a = 2.5 \times 10^{-9}$.

This is just a weak acid problem.

The reaction is $\text{HOBr}(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{OBr}^-(\text{aq})$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{OBr}^-]}{[\text{HOBr}]} = 2.5 \times 10^{-9}$$

For our ICE table

	Initial	Change	Equilibrium
H_3O^+	0	x	x
OBr^-	0	x	x
HOBr	0.200	- x	0.200 - x

$$\frac{(x)(x)}{(0.200 - x)} = 2.5 \times 10^{-9} \quad \text{If we assume } x < 0.200, \text{ then we get}$$

$$\frac{x^2}{0.200} = 2.5 \times 10^{-9} \quad x^2 = (2.5 \times 10^{-9})(0.200) = 5.0 \times 10^{-10} \quad x = (5.0 \times 10^{-10})^{1/2} \\ = 2.2 \times 10^{-5}$$

So our assumption that $x \ll 0.200$ was good.

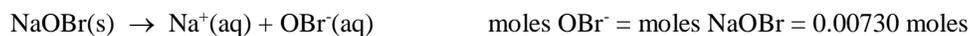
$$\text{So } [\text{H}_3\text{O}^+] = x = 2.2 \times 10^{-5} \text{ M} \quad \text{pH} = -\log_{10}(2.2 \times 10^{-5}) = 4.65$$

b) 0.868 g of solid sodium hypobromite (NaOBr) is added to the above solution. What is the new pH of the solution after the addition of NaOBr?

$$\text{MW}(\text{NaOBr}) = 118.9 \text{ g/mol}$$

$$\text{moles NaOBr} = 0.868 \text{ g} \frac{1 \text{ mol}}{118.9 \text{ g}} = 0.00730 \text{ mol NaOBr}$$

NaOBr is a soluble ionic compound and a strong electrolyte, and so



$$[\text{OBr}^-] = \frac{0.00730 \text{ mol}}{0.5000 \text{ L}} = 0.0146 \text{ mol/L}$$

Since we have an appreciable amount of both a weak acid and its conjugate base present, we can use the Henderson equation to find pH.

$$\text{pH} = \text{p}K_a + \log_{10}\left\{\frac{[\text{base}]}{[\text{acid}]}\right\} = -\log_{10}(2.5 \times 10^{-9}) + \log_{10}\left\{\frac{(0.0146)}{(0.200)}\right\} \\ = 8.60 + (-1.14) = 7.46$$

Note we could get this same result using an ICE table, but using the Henderson equation is faster. As long as we know how to use the equation correctly this is the preferred method.