

AP CHEMISTRY

TOPIC 7: ACIDS & BASES, PART F, PART II

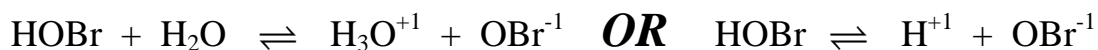
Day 94:

- Titrations
- pH curve
- Indicators

$$K_a = \frac{[H_3O^{+1}][OBr^{-1}]}{[HOBr]} = 2.00 \times 10^{-9}$$

Hypobromous acid, HOBr, is a weak acid in water. The K_a expression for HOBr is shown above.

a) Write a chemical equation showing how HOBr behaves as an acid in water.



b) Calculate the pH of a 0.175 M solution of HOBr.

	[HOBr]	\rightleftharpoons	[OBr ⁻¹]	+	[H ⁺¹]
I	0.175		0		0
C	-x		+x		+x
E	0.175 - x		x		x

$$K_a = \frac{[H^{+1}][OBr^{-1}]}{[HOBr]} = \frac{x^2}{(0.175 - x)} = \frac{x^2}{(0.175)} = 2.00 \times 10^{-9}$$

$$x^2 = (2.00 \times 10^{-9})(0.175) = 3.50 \times 10^{-10}$$

$$x = \sqrt{3.50 \times 10^{-10}} = 1.87 \times 10^{-5}$$

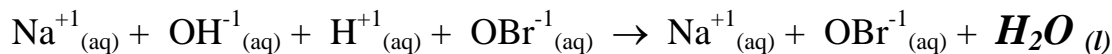
$$x = [H_3O^{+1}] = [H^{+1}] = 1.87 \times 10^{-5} M$$

$$\text{pH} = -\log(1.87 \times 10^{-5}) = \mathbf{4.73}$$

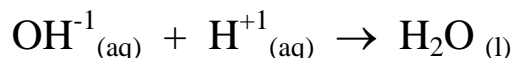
c) Write the net ionic equation for the reaction between the weak acid, HOBr_(aq) and the strong base NaOH_(aq).



Ionic Equation:



Net Ionic Equation:



Net Ionic Equation (Definition): A net ionic equation is written by dropping out the spectator ions and showing only those chemical species that are involved in the chemical reaction.

This reaction is an important (**neutralization**) reaction that occurs when an acid (H^{+1}) reacts with a base (OH^{-1}).

d) In an experiment, 20.00 mL of 0.175 M HOBr_(aq) is placed in a flask and titrated with 6.55 mL of 0.435 M NaOH_(aq).

(i) Calculate the number of moles of NaOH_(aq) added.

$$\text{Number of moles (NaOH)} = \frac{6.55 \text{ mL}}{1000 \text{ mL}} \times \frac{1 \text{ L}}{1} \times \frac{0.435 \text{ mol}}{L} = 0.00285 \text{ mol NaOH} = OH^{-1}$$

(ii) Calculate $[H_3O^{+1}]$ (or $[H^{+1}]$) in the flask after the NaOH_(aq) has been added.

$$\text{Number of moles (HOBr)} = \frac{20.00 \text{ mL}}{1000 \text{ mL}} \times \frac{1 \text{ L}}{1} \times \frac{0.175 \text{ mol}}{L} = 0.00350 \text{ mol HOBr}$$

$$\text{Total Volume: } 20.0 \text{ mL} + 6.55 \text{ mL} = 26.55 \text{ mL} = 0.02655 \text{ L}$$

$$[HOBr] = \frac{0.00350 \text{ mol}}{0.02655 \text{ L}} = 0.1318 \text{ M} \quad ; \quad [OH^{-1}] = \frac{0.00285 \text{ mol}}{0.02655 \text{ L}} = 0.107 \text{ M}$$

	OBr ⁻¹	+	H ₂ O	⇌	HOBr	+	OH ⁻¹
I	~ 0		-		0.1318		0.107
C	+ 0.107		-		- 0.107		- 0.107
E	0.107		-		0.0248		~ 0

$$K_b = \frac{K_w}{K_a} = \frac{1.00 \times 10^{-14}}{2.00 \times 10^{-9}} = 5.00 \times 10^{-6}$$

$$K_b = \frac{[HOBr][OH^{-1}]}{[OBr^{-1}]} = \frac{(0.0248)[OH^{-1}]}{(0.107)} = 5.00 \times 10^{-6}$$

$$\frac{(5.00 \times 10^{-6})(0.107)}{(0.0248)} = 2.15 \times 10^{-5} = [OH^{-1}]$$

$$K_w = [H^{+1}][OH^{-1}] = 1.00 \times 10^{-14}$$

$$[H^{+1}] = \frac{1.00 \times 10^{-14}}{[OH^{-1}]} = \frac{1.00 \times 10^{-14}}{2.15 \times 10^{-5}} = 4.64 \times 10^{-10}$$

OR, you could find the pH and then calculate the Antilog (- pH) to get the $[H^{+1}]$.

(iii) Calculate $[OH^{-1}]$ in the flask after the $NaOH_{(aq)}$ has been added.

$$K_w = [H^{+1}] [OH^{-1}] = 1.00 \times 10^{-14}$$

$$[OH^{-1}] = \frac{1.00 \times 10^{-14}}{[H^{+1}]} = \frac{1.00 \times 10^{-14}}{4.64 \times 10^{-10}} = 2.15 \times 10^{-5} M$$

(iv) Calculate the pH of the titration solution after the $NaOH_{(aq)}$ has been added.

$$\mathbf{pH} = - \log (4.64 \times 10^{-10}) = \mathbf{9.33}$$

(v) Is the titration solution at the equivalence point? Justify your answer.

No, At the equivalence point, the number of moles of HOBr (acid) will be equal to the number of moles of OH^{-1} (base).