

AP CHEMISTRY

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

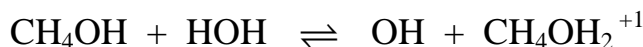
Day 94:

- Titrations
- pH curve
- Indicators

$$K_b = \frac{[OH^{-1}][CH_4OH_2^{+1}]}{[CH_4OH]} = 4.57 \times 10^{-5}$$

Methylamine, CH_4OH , is a weak base in water. The K_b expression is shown above.

- a) Write a chemical equation showing how CH_4OH behaves as an base in water.



- b) Calculate the pH of a 0.588 M solution of CH_4OH .

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

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

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

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

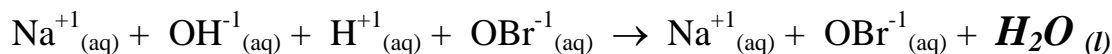
$$x = [H_3O^{+1}] = [H^{+1}] = 1.87 \times 10^{-5} M$$

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

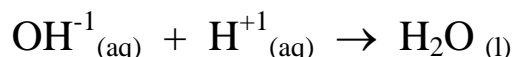
- c) Write the net ionic equation for the reaction between the weak base, $CH_4OH_{(aq)}$ and the strong acid, $HNO_{3(aq)}$.



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, 42.0 mL of 0.588 M CH₃OH_(aq) is placed in a flask and titrated with 10.0 mL of 0.400 M HNO_{3(aq)}.

(i) Calculate the number of moles of HNO_{3(aq)} added.

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

(ii) Calculate [OH⁻¹] in the flask after the HNO_{3(aq)} has been added.

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

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

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

	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

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

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

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

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

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

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

(iii) Calculate $[H^{+1}]$ in the flask after the $HNO_3_{(aq)}$ 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 $HNO_3_{(aq)}$ has been added.

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