

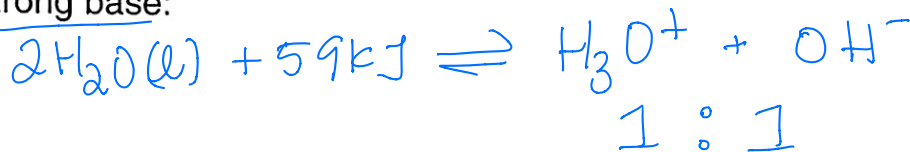
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Chemistry 12
ACID BASE PART II Lesson # 15
Mixing STRONG acids and Bases

$$M_1 V_1 = M_2 V_2$$

mixing

Recall the NET IONIC EQUATION for the NEUTRALIZATION reaction of a strong acid with a strong base:



When you MIX A STRONG ACID WITH A STRONG BASE there are THREE POSSIBLE OUTCOMES:

1. $[H_3O^+] = [OH^-] \Rightarrow$ neutral
2. $[H_3O^+] > [OH^-]$ OR $[OH^-] < [H_3O^+] \Rightarrow$ ACIDIC
3. $[H_3O^+] < [OH^-]$ OR $[OH^-] > [H_3O^+] \Rightarrow$ BASIC

TO SOLVE THESE PROBLEMS THE FOLLOWING EQUATIONS ARE USED:

A. IF THE SOLUTION IS BASIC (pH > 7)

$$[OH^-]_{XS} = [OH^-]_{ST} - [H_3O^+]_{ST}$$

B. IF THE SOLUTION IS ACIDIC (pH < 7)

$$[H_3O^+]_{XS} = [H_3O^+]_{ST} - [OH^-]_{ST}$$

Example 1. If 10.0 mL of 0.100 M HCl is mixed with 90.0 mL of 0.100 M NaOH, What is the pH of the resulting mixture?

1. HCl is a strong acid, $[HCl] = [H_3O^+]$ NaOH is a strong base; $[NaOH] = [OH^-]$

$$\therefore 0.100 \text{ M } H_3O^+ (M_1)$$

$$\therefore 0.100 \text{ M } OH^- (M_2)$$

2. Dilution calculations for $[H_3O^+]$ and $[OH^-]$:

$$[H_3O^+]_{ST} = \frac{0.100 \text{ M} \cdot 0.0100 \text{ L}}{0.1000 \text{ L}}$$

$$= 0.0100 \text{ M } H_3O^+$$

$$[OH^-]_{ST} = \frac{0.100 \text{ M} \cdot 0.0900 \text{ L}}{0.1000 \text{ L}}$$

$$= 0.0900 \text{ M } OH^-$$

3. Which ion has the larger concentration? $[H_3O^+]$ or $[OH^-]$?

$$0.0900 \text{ M } OH^- > 0.0100 \text{ M } H_3O^+$$

4. Solve for the XS of OH^- .

$$[OH^-]_{XS} = [OH^-]_{ST} - [H_3O^+]_{ST}$$

$$= 0.0900 \text{ M} - 0.0100 \text{ M} = \boxed{0.0800 \text{ M } OH^-}$$

5. Solve for the pH

$$pOH = -\log(0.0800)$$

$$= 1.097$$

$$pH = 14.000 - 1.097$$

$$\boxed{pH = 12.903}$$

Example 2. Calculate the pH that results when 50.0 mL of 0.150 M LiOH is mixed with 50.0 mL of 0.200 M HNO₃.

1. HNO₃ is a strong acid; [HNO₃] = [H₃O⁺] LiOH is a strong base; [LiOH] = [OH⁻]

∴ 0.200 M H₃O⁺ ∴ 0.150 M OH⁻

2. Dilution calculations for [H₃O⁺] and [OH⁻]:

$$[H_3O^+]_{ST} = \frac{0.200 M \cdot 0.0500 L}{0.1000 L} = 0.100 M H_3O^+$$

$$[OH^-]_{ST} = \frac{0.150 M \cdot 0.0500 L}{0.1000 L} = 0.0750 M OH^-$$

3. Which ion has the larger concentration? [H₃O⁺] or [OH⁻]?

$$0.100 M H_3O^+ > 0.0750 M OH^-$$

4. Solve for the XS of H₃O⁺.

$$[H_3O^+]_{XS} = [H_3O^+]_{ST} - [OH^-]_{ST} = 0.100 M - 0.0750 M = 0.025 M H_3O^+$$

5. Solve for the pH

$$pH = -\log(0.025)$$

$$pH = 1.60$$

Example 3. How many grams of NaOH must be added to 40.0 mL of 0.180 M HCl to produce a solution having a pH of 12.500. Assume NO CHANGE in volume when the NaOH is added.

1. pH of 12.500 means the solution is BASIC. There is an excess of OH⁻!

2. Use pH to solve for the excess concentration of [OH⁻]_{XS}

$$pOH = 14.000 - 12.500 = 1.500$$

$$[OH^-]_{XS} = \text{antilog}(-1.500) = 0.0316 M = 3.16 \cdot 10^{-2} M OH^-$$

3. Rearrange the XS equation to solve for [OH⁻]_{ST}

$$[OH^-]_{XS} = [OH^-]_{ST} - [H_3O^+]_{ST}$$

$$0.0316 M = [X]_{ST} - 0.180 M + 0.180 M \rightarrow [OH^-]_{ST} = 0.212 M$$

4. Now convert concentration into grams!

$$0.0400 L \cdot \frac{0.212 \text{ mol NaOH}}{1 L} \left(\frac{40.0 \text{ g NaOH}}{1 \text{ mol NaOH}} \right) = 0.339 \text{ g NaOH}$$

$$1 Na = 23.0$$

$$1 O = 16.0$$

$$1 H = 1.0$$

SEATWORK/HOMEWORK: Exercises 58 - 68 pgs 143-144 in Hebden

PLO's: P5