

XIII) Mixtures of Strong Acids & Bases

January 17, 2018 11:08 PM

Assignment 10:

- 1) Hebden p. 139 #51, 52
- 2) Water at a certain temperature has a K_w of 4.4×10^{-15} .
 - a) Is the water at a temperature above or below 25°C ?
 - b) What is the pK_w ?
 - c) What would the pH scale be at this temperature?
 - d) Find the $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$.
 - e) Find the pH and pOH.
 - f) Is water at this temperature acidic, basic, or neutral?

XIII) Mixtures of Strong Acids and Bases

Mixing an acid solution with a basic solution produces a solution that can be acidic, basic, or neutral depending on the moles of acid compared to the moles of base mixed. H_3O^+ ions react with OH^- ions to make $2\text{H}_2\text{O}$ molecules, known as neutralization. But if there are more of one ion than the other, the resulting solution will not be neutral.

$$\text{Molarity} = \frac{\text{moles}}{\text{volume}}$$

$$[\text{H}_3\text{O}^+] > [\text{OH}^-] = \text{acidic}$$

$$[\text{H}_3\text{O}^+] < [\text{OH}^-] = \text{basic}$$

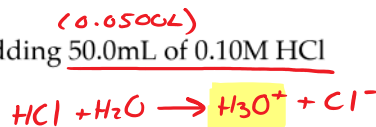
Example:

1. Calculate the pH of a solution obtained by adding 50.0mL of 0.10M HCl to 80.0mL of 0.15M NaOH.



$$\text{mol OH}^- = (0.15\text{M})(0.0800\text{L}) = 0.012 \text{ mol OH}^-$$

Both the strong acid + strong base dissociate 100%.



$$\text{mol H}_3\text{O}^+ = M \cdot V = (0.10\text{M})(0.0500\text{L})$$

$$\text{mol H}_3\text{O}^+ = 0.0050 \text{ mol}$$

Compare mol of OH^- to mol of H_3O^+ \therefore excess of OH^- in sol'n

$$\begin{array}{r} \text{excess OH}^- \\ 0.012 \text{ mol OH}^- \\ - 0.0050 \text{ mol H}_3\text{O}^+ \\ \hline 0.007 \text{ mol OH}^- \end{array}$$

Convert mol of excess $\text{OH}^- \rightarrow [\text{OH}^-]$ *Add the volumes!

$$M = \frac{n}{V} \quad [\text{OH}^-] = \frac{0.007 \text{ mol}}{0.1300 \text{ L}}$$

$$50.0\text{mL} + 80.0\text{mL} = 130.0\text{mL} = 0.1300\text{L}$$

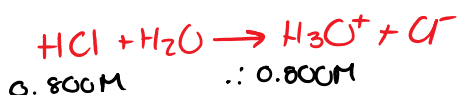
$$[\text{OH}^-] = 0.053846\text{M}$$

$$\therefore \text{pOH} = -\log(0.053846\text{M}) \text{ 1 s.f. in conc}$$

$$\text{pOH} = 1.2688 \text{ 1 dp in pOH/pH}$$

$$\therefore \text{pH} = 14 - 1.2688 = 12.7 \text{ Basic } \uparrow [\text{OH}^-]$$

2. Calculate the pH of a solution obtained by adding 1.00g of $\text{Ca}(\text{OH})_2$ to 650.0mL of 0.800M HCl.



0.800M

\therefore 0.800M

$$\text{mol H}_3\text{O}^+ = (0.800\text{M})(0.6500\text{L}) = 0.520 \text{ mol H}_3\text{O}^+$$

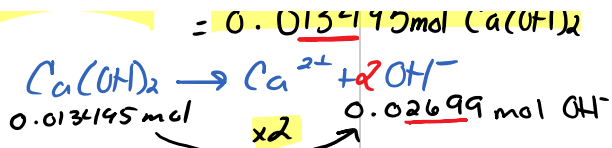
$$\begin{array}{r} \text{3 s.f.} \\ 1.00\text{g Ca}(\text{OH})_2 \quad | \quad 1 \text{ mol} \\ \hline 74.1\text{g} \end{array}$$

$$= 0.013495 \text{ mol Ca}(\text{OH})_2$$



0.0001

$$\text{mol H}_3\text{O}^+ = (0.800\text{M})(0.6500\text{L}) = 0.520 \text{ mol H}_3\text{O}^+$$



H₃O⁺ mols in excess

$$0.520 - 0.02699$$

$$0.493 \text{ mol H}_3\text{O}^+ \rightarrow [\text{H}_3\text{O}^+] \rightarrow \text{pH}$$

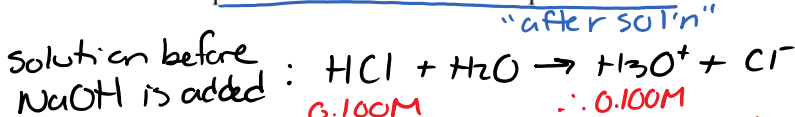
$$[\text{H}_3\text{O}^+] = \frac{0.493 \text{ mol}}{0.6500\text{L}} = 0.758477\text{M}$$

$$\therefore \text{pH} = -\log(0.758477)$$

$$\text{pH} = 0.120$$

very! Acidic (strong acid) in excess

3. What mass of NaOH would have to be added to 500.0 mL of 0.100 M HCl in order to produce a solution with a pH of 3.200?



$$\text{mol of H}_3\text{O}^+ = (0.100\text{M})(0.5000\text{L}) = 0.0500 \text{ mol H}_3\text{O}^+$$

After NaOH is added: known pH = 3.200 \rightarrow [H₃O⁺] \rightarrow mol H₃O⁺

$$[\text{H}_3\text{O}^+] = 10^{-(3.200)} = 6.30957 \times 10^{-4}\text{M}$$

* Assume the volume has not changed

$$\text{mol H}_3\text{O}^+ = (6.30957 \times 10^{-4}\text{M})(0.5000\text{L}) = 3.15479 \times 10^{-4} \text{ mol}$$

Difference in moles represents the H₃O⁺ "lost" by reacting with OH⁻ when NaOH is added.

$$0.0500 \text{ mol H}_3\text{O}^+ - 0.000315479 \text{ mol H}_3\text{O}^+ = 0.04968 \text{ mol H}_3\text{O}^+$$

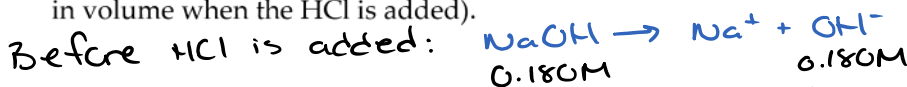
Means that 0.04968 mol of OH⁻ reacted to use up (1:1 ratio)

Therefore

0.04968 mol OH ⁻	40.0 g	= 1.987 g
	1 mol	

= 1.99 g of NaOH

4. How many moles of HCl must be added to 40.0 mL of 0.180 M NaOH to produce a solution having a pH of 12.500? (Assume that there is no change in volume when the HCl is added).



$$\text{mol OH}^- = (0.180\text{M})(0.0400\text{L}) = 0.00720 \text{ mol OH}^-$$

After HCl is added, the pH = 12.500 (2 sig figs)

if pH = 12.500 \therefore pOH = 14 - 12.500 = 1.500

$$[\text{OH}^-] = 10^{-(1.500)} = 0.031623\text{M} \quad \therefore \text{mol OH}^- = (0.031623\text{M})(0.0400\text{L}) = 1.2649 \times 10^{-3} \text{ mol OH}^-$$

* The difference in mols OH⁻ is how much reacted with H₃O⁺ (HCl)

$$0.00720 - 1.2649 \times 10^{-3} = 0.005935 \text{ mol OH}^-$$

react @ 1:1 ratio OH⁻ : H₃O⁺

\therefore the amount of H₃O⁺ added (1:1) was 0.005935 mol

$$\therefore 0.00594 \text{ moles of HCl were added}$$

$\therefore 0.00594$ moles of HCl were added

Assignment 11: Hebden p. 143 # 58, 60, 62, 65, 67

XIV) Titrations

A **titration** is a laboratory technique that is most often used to find the concentration (molarity) of a solution. Acid/base titrations are a common type of titration in which a base is used to find an unknown acid concentration, or *visa versa*.

Suppose you are cleaning up the lab and you find a large container labeled 'hydrochloric acid', but the concentration is not given. A _____ can be done to find the unknown concentration.

Titration is a process (procedure and calculations) for determining the concentration of a substance accurately and precisely using a measurable volume of a standardized solution. A _____ is simply a reactant of known concentration.