# 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 $4.4 \times 10^{-15}$.

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 $[H_3O^+]$ and $[OH^-]$.
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 $2H_2O$ molecules, known as neutralization. But if there are more of one ion than the other, the resulting solution will not be neutral.

$[H_3O^+] > [OH^-]$ = acidic

$[H_3O^+] < [OH^-]$ = basic

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

Example:

1. Calculate the pH of a solution obtained by adding 50.0mL of 0.10M HCl (0.0500L) to 80.0mL of 0.15M NaOH. (0.0800L)

Both the strong acid + strong base dissociate 100%

$NaOH \rightarrow Na^+ + OH^-$

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$

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

mol $H_3O^+$ = M·V = (0.10M)(0.0500L)
$\boxed{\text{mol } H_3O^+} = 0.0050$ mol

compare mol of $OH^-$ to mol of $H_3O^+$ ∴ excess of $OH^-$ in sol'n

excess $OH^-$
  0.012 mol $OH^-$
− 0.0050 mol $H_3O^+$
―――――――
0.007 mol $OH^-$   sig figs (3 d.p.)

Convert mol of excess $OH^- \longrightarrow [OH^-]$ *Add the volumes!

50.0mL + 80.0mL = 130.0mL
= 0.1300 L

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

$[OH^-] = 0.053846M$



∴ pOH = −log (0.053846M)   1 s.f. in [conc]

pOH = 1.2688   1 dp in pOH/pH

∴ pH = 14 − 1.2688 = 12.7   Basic ↑$[OH^-]$

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

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$
0.800M   ∴ 0.800M

mol $H_3O^+$ = (0.800M)(0.6500L)
= 0.520 mol $H_3O^+$

3 s.f.

$\frac{1.00g\ Ca(OH)_2}{} \Big| \frac{1 \text{ mol}}{74.1g}$

= 0.013495 mol $Ca(OH)_2$

$Ca(OH)_2 \rightarrow Ca^{2+} + 2OH^-$

$= 0.013495 mol\ Ca(OH)_2$

mol $H_3O^+$ = (0.800M)(0.6500L)
= 0.520 mol $H_3O^+$

$Ca(OH)_2 \rightarrow Ca^{2+} + 2OH^-$
0.013495 mol $\xrightarrow{\times 2}$ 0.02699 mol $OH^-$

$H_3O^+$ mols in excess
0.520
− 0.02699
0.493 mol $H_3O^+$ → $[H_3O^+]$ → pH

$[H_3O^+] = \frac{0.493 mol}{0.6500L} = 0.758477M$   ∴ pH = −log(0.758477)

$\boxed{pH = 0.120}$ very! Acidic (strong acid in excess)

(g)

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

"before sol'n"   "after sol'n"

Solution before NaOH is added: $HCl + H_2O \rightarrow H_3O^+ + Cl^-$
0.100M   ∴ 0.100M
↳ mol of $H_3O^+$ = (0.100M)(0.5000L)
= 0.0500 mol $H_3O^+$

After NaOH is added: $\boxed{\text{known pH} = 3.200}$ → $[H_3O^+]$ → mol $H_3O^+$

$[H_3O^+] = 10^{\wedge(-3.200)} = 6.30957 \times 10^{-4} M$

*Assume the volume has not changed

mol $H_3O^+$ = $(6.30957 \times 10^{-4}M)(0.5000L)$
mol $H_3O^+$ = $3.15479 \times 10^{-4}$ mol

Difference in moles represents the $H_3O^+$ "lost" by reacting with $OH^-$ when NaOH is added.

0.0500 mol $H_3O^+$
− 0.000315479 mol $H_3O^+$
0.04968 mol $H_3O^+$

Means that 0.04968 mol of $OH^-$ reacted to use up   $OH^-:H_3O^+$ (1:1 ratio)

Therefore $0.04968 mol\ OH^- \times \frac{40.0g}{1 mol} = 1.987g$

$\boxed{= 1.99g \text{ of NaOH}}$

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).

Before HCl is added: $NaOH \rightarrow Na^+ + OH^-$
0.180M   0.180M
↳ mol $OH^-$ = (0.180M)(0.0400L)
= 0.00720 mol $OH^-$

After HCl is added, the pH = 12.500 (3 sig figs)
if pH = 12.500 ∴ pOH = 14 − 12.500   pOH = 1.500
∴ $[OH^-] = 10^{\wedge(-1.500)} = 0.031623M$   ∴ mol $OH^-$ = (0.031623M)(0.0400L)
$= 1.26491 \times 10^{-3}$ mol $OH^-$

*The difference in mols $OH^-$ is how much reacted with $H_3O^+$ (HCl) ∴ 0.00720 − $1.2649 \times 10^{-3}$ = 0.005935 mol $OH^-$
3 d.p.
react @ 1:1 ratio $OH^- : H_3O^+$

∴ The amount of $H_3O^+$ added (1:1) was 0.005935 mol

$\boxed{\text{∴ 0.00594 moles of HCl were added}}$

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

$\therefore$ 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.