Wessner

Chemistry

Week 7 and 8 work

May 11th - May 26th

Due Tuesday

May 26nd because of

Memorial day on

Monday

# Week 7 notes and work

## Chapter 15

## Acid – Base pH

- pH stands for "hydrogen power"
- 
- pH scale 0 – 14
- 0-6 acid  7 neutral  8-14 base
- pH is based on the concentration of $H_3O^+$ ions.
- The more $H_3O^+$ ions the more acidic the solution, the less $H_3O^+$ ions the more basic the solution.

- Acids form Hydronium ions($H_3O^+$) and Bases form Hydroxide ions ($OH^-$).
- 
- Water self-ionizes, in other words, 2 $H_2O$ molecules will combine to form $H_3O^+$ and $OH^-$
- $H_2O(l) + H_2O(l) \; ---\geq \; H_3O^+(aq) + OH^-(aq)$

- Concentrations of H3O+ and OH- in pure water are each
- H3O+ 1.0 X 10-7 mol/L
- OH- 1.0 X 10-7 mol/L
- at 25oC

- Water has a mathematical ionization constant at a constant temperature. At room temp. 25oC
- The ionization constant of water (Kw)
- Kw = [H3O+][OH-]
- = (1.0 X 10-7)(1.0 X10-7)
- Kw = 1.0 X 10-14

Neutral, Acidic, and Basic solutions

- When H3O+ is equal to the amount of OH- then the solution is neutral
- H3O+ > OH- = acidic solution

- If **H3O+** is less than 1.0 x 10-7
- Between (x 10-1 through x 10-6 )
- then the solution is acidic
- (this means OH- concentration is between
- x10-8 to x10-14)

- H3O+ < OH- = basic solution
- If **OH-** is less than 1.0 x 10-7 between (1.0x 10-1 through1.0 x 10-6 ) then solution is basic,
- (this means the H3O+ concentration is
- Between x 10-8 to x10-14 )

Calculating H3O+ and OH-

- Strong acids and bases are completely ionized in aqueous solutions.
- 1 mol of a strong base NaOH will yield
- 1 mol of OH- (aq) H2O
- NaOH (s) —— ≥ Na+(l) + OH-(l)
- 1 mol 1 mol 1mol
- Therefore……

a 1.0 x 10-2 M NaOH solution has an [OH-] of 1.0 x 10-2 M, as shown here:

- 1.0 x 10-2 x 1 mol OH- = 1.0 x 10-2 mol OH- =
- 1 L solution 1 mol NaOH 1 L solution

= 1.0 x 10-2 M OH-

OH- is less than 1.0 x 10-7 so this solution is basic.

# Using Kw – ionization constant of water

We can use the ionization constant of water Kw, when one ion concentration is known, to find the concentration of the other ion.

$Kw = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$

To find the $H_3O^+$ concentration, the equation would be

$$[H_3O^+] = \frac{Kw}{[OH^-]}$$

- $Kw = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$
- $OH^- = 1.0 \times 10^{-2}$
- $[H_3O^+] = \dfrac{1.0 \times 10^{-14}}{1.0 \times 10^{-2}} = 1.0 \times 10^{-12}$ M

To find $OH^-$ when you know the $H_3O^+$ is the same Kw equation.

$Kw = [H_3O^+][OH^-]$

You have a $2.0 \times 10^{-4}$ M HCl solution. Because HCl is a strong acid, $[H_3O^+] = 2.0 \times 10^{-4}$. what is the $[OH^-]$ concentration?

$$OH^- = \frac{1.0 \times 10^{-14}}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-4}}$$

$OH^- = 5.0 \times 10^{-11}$ M

- The pH scale is 0 – 14
- 0 – 6 equals an acidic solution
- 7 neutral
- 8 – 14 is a basic solution

- pH stands for "hydrogen power"
- pH indicates the hydronium ion($H_3O^+$) concentration of a solution.
- pH of a solution is defined as the negative of the common logarithm(log) of the hydronium ion concentration.
- Expressed:
- $pH = -\log[H_3O^+]$

- $pH = -\log[H_3O^+]$
- A neutral solution at 25°C has a $[H_3O^+]$ of $1 \times 10^{-7}$ M.
- pH = ?
- On your calculator
- Exp or EE takes the place of x 10 for scientific notation.
- So $1 \times 10^{-7}$ on your calculator is
- 1 exp/EE – 7
- pH = (1 exp/EE 7 - )log - = 7.0

- $pOH = -\log[OH^-]$
- If the $OH^-$ of a solution is $1 \times 10^{-7}$ M, what is the pOH?
- $pOH = -\log[OH^-]$ = ($1 \times 10^{-7}$)-log =
- pOH = 7
-

- pH + pOH = +14

- If pH + pOH = + 14 then a solution with a pH = 5, would have a pOH =?
- pOH = 14-5 = 9
-
- The solution above is acidic, it's pH is between 0 and 6

Determining pH

Acid-Base indicators can be used to determine pH of a solution. these weak acid or weak base indicator solutions are color sensitive, in other words they change to specific colors for specific pHs.

Weak acid indicator ,H*In* , can be represented by the equation

H*In* ≤≠≥ *H+* + *In-* ≥

H+ cation part of the indicator In- anion part of the indicator

## Determining pH and Titrations

- Acid base indicators are compounds whose colors are sensitive to pH
- They are weak acids or bases

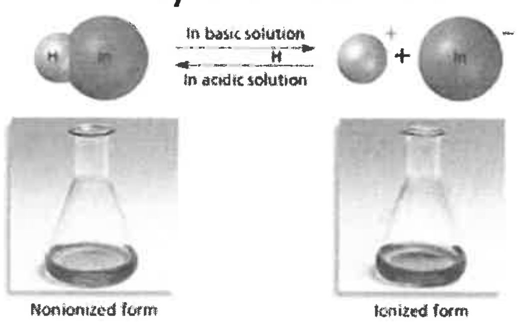
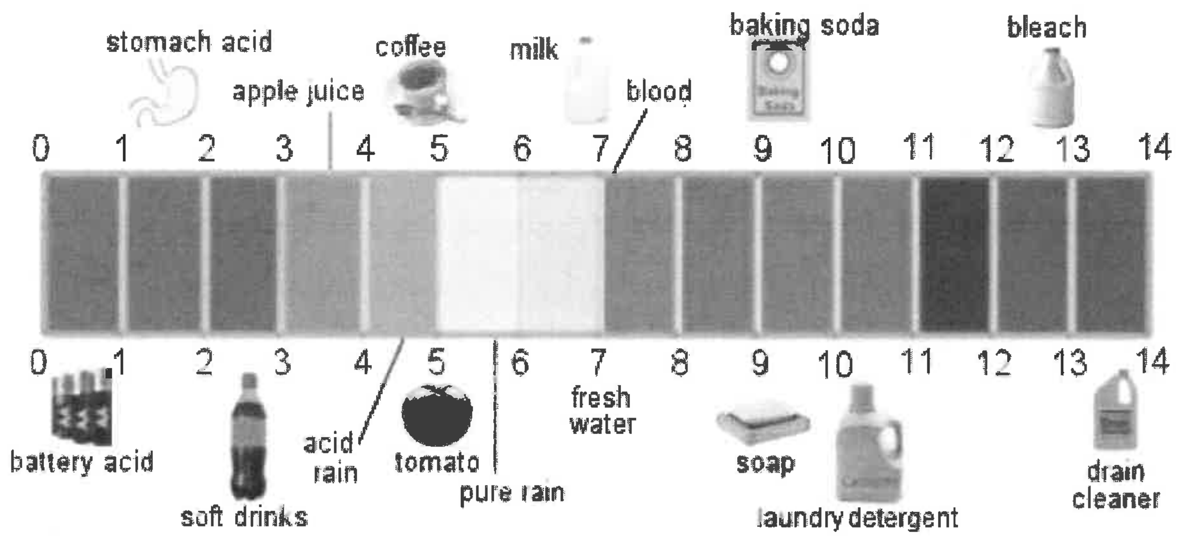



FIGURE 16-4 Basic solutions shift the equilibrium of litmus to the right. The ionized form, $In^-$, then predominates, and the litmus turns blue. Acidic solutions shift the equilibrium of the indicator litmus to the left. The nonionized form, H*In*, predominates, and the litmus turns red.

If we use a litmus paper indicator , the paper will turn red in the presence of an acid and will turn blue in the presence of a base




Indicators come in many different color ranges

The pH color change range of an indicator is called the transition interval

Different indicators change colors at different pH levels

For example:

Methyl reds transition interval is

pH 4.4 to pH 6.2

And will change from a red color to a yellow color

Phenol reds transition interval is

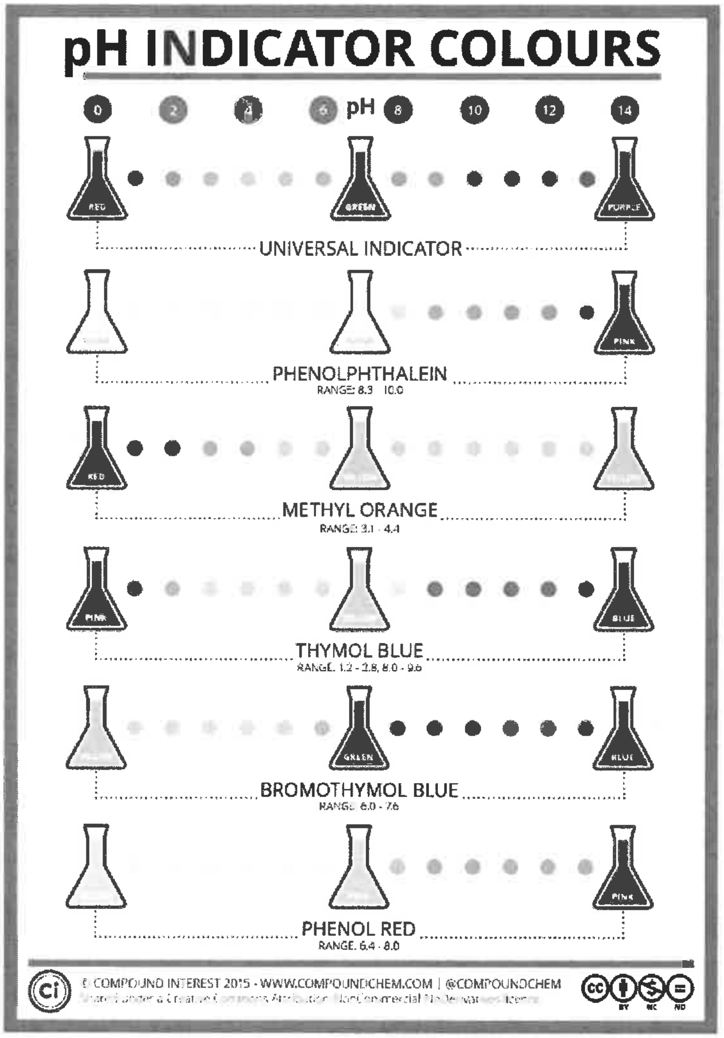
pH 6.4 to pH 8.0

And will change from a yellow to a red color.

Methyl oranges transition interval is pH 3.1 to pH 4.4 changing from an orange to a yellow color.

There are also universal indicators that will change to all colors of the rainbow because they are soaked with more than one indicator solution.

One can also use a pH meter. This is an electronic device which uses a sensitive wand and computer to measure the pH.



Titrations

Titrations are used to determine the exact pH concentration

Titrations are really just neutralizations of solutions, using a known pH solution

And an unknown pH solution.

Titration is a controlled addition and measurement of the amount of a solution of known concentration required to react completely with a measured amount of a solution of an unknown concentration.

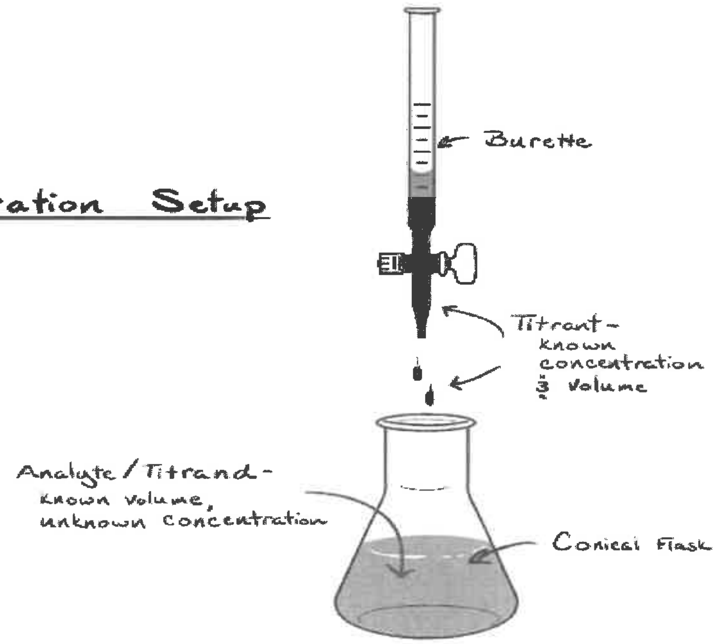
Equivalence point – the point at which the two solutions used in a titration are present in chemically equivalent amounts.

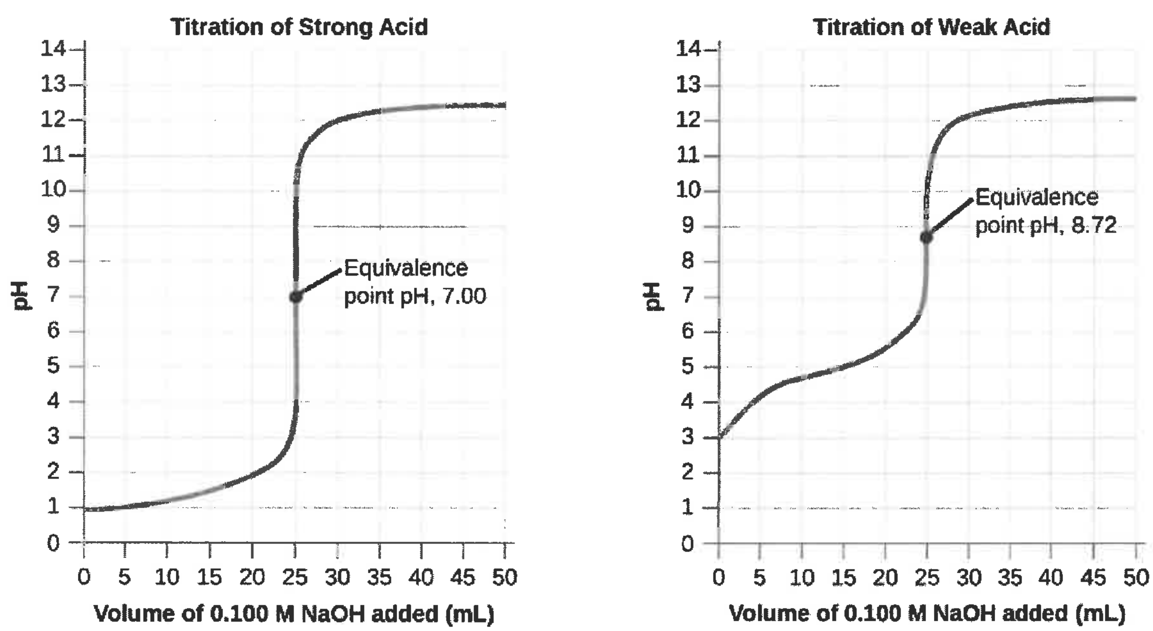
pH will change rapidly at the equivalence point is approached while doing a titration.

Indicators will need to be used that have the pH values of the equivalence point.

Below is a basic titration set up, the known titrant concentration is in the burette and will need to slowly drip into the known solution with the unknown concentration to determine the exact pH.

## Titration Setup




(a) (b)

The point in a titration at which an indicator changes color is called the end point of the indicator.

A standard solution – the solution that contains the –precisely known concentration of a solute

Primary standard – is a highly purified solid compound used to check the concentration of the known solution is a titration.

Titration problem steps:

- Start with the balanced equation for the neutralization reaction and determine the chemically equivalent amounts of the acid and base.
- Determine the moles of acid (or base) from the known solution used during the titration
- Determine the moles of solute of the unknown solution used during the titration
- Determine the molarity of the unknown solution.

20.0 mL of 5.0 x 10-3 M NaOH is required to reach the end point in a titration of 10.0 mL of HCl of an unknown concentration. What is the molarity of the HCl?

Balanced equation

- HCl + NaOH ——≥ NaCl + H2O

  1 mol 1 mol 1 mol 1 mol

- Calculate the number of moles of NaOH used in the titration

  $\frac{5.0 \times 10^{-3} \text{ mol NaOH}}{1 \text{ L}} \times \frac{1 \text{ L}}{1000 \text{mL}} \times 20.0 \text{ mL} = 1.0 \times 10^{-4}$ mol NaOH used

- 1 mol of NaOH is needed to neutralize 1 mol of HCl the amount of HCl in the titration must be 1.0 x 10-4 mol.

1.0 x 10-4 mol NaOH x <u>1 mol HCl</u> = 1.0 x 10-4 mol HCl

1 mol NaOH

- 10.0 mL of HCl is the amount of known solution with unknown concentration an be used to determine the molarity of HCl

$$\frac{1.0 \times 10\text{-}4 \text{mol HCl}}{10.0 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = \frac{1.0 \times 10\text{-}2 \text{ mol HCl}}{1\text{L}}$$

= 1.0 x10-2 M HCl

So our solution of HCl has a 1.0 x 10-2 molarity.

Example problem #2

In a titration 27.4 mL of 0.0154 M Ba(OH)2 is added to a 20.0 mL sample of HCl solution of unknown concentration. What is the molarity of the HCl?

Balance:

Ba(OH)2 + HCl ——≥ BaCL2 + H2O

__Mol __mol __mol __mol

Find amount of base used:

$$\frac{\_\_\_\_\text{ mol Ba(OH)2}}{1 \text{ L}} \times \frac{1 \text{ L}}{1000\text{mL}} \times \_\_\_\_\text{mL Ba(OH)2} =$$

Find moles of acid used:

$$\_\_ \text{ mol Ba(OH)2} \times \frac{\_\_\_\text{ mol HCl}}{\_\_ \text{ mol Ba(OH)2}} =$$

Find the molarity of the HCl:

Amount of solute in unknown solution mol  x  1000 mL =

Volume of unknown solution mL  1L

4.22 x 10-2 M HCl answer

# Titration work to be turned in:

- If it takes 54 mL of .1 M of NaOH to neutralize 125mL of an HCl solution. What is the concentration of the HCl?

- If it takes 25 mL of .05 M HCl to neutralize 345 mL of NaOH solution, what is the concentration of the NaOH solution?

- If it takes 50 mL of .5 M KOH solution to completely neutralize 125 mL of $H_2SO_4$, what is the concentration of the $H_2SO_4$ solution?

- What is the concentration of a sodium hydroxide solution if 14.5 mL of it are exactly neutralized by 30.0 mL of a 0.500 M hydrochloric acid solution?

- Phosphoric acid is neutralized by potassium hydroxide according to the following reaction:

  $KOH\ (aq) + H_3PO_4\ (aq) \rightarrow K_3PO_4\ (aq) + H_2O\ (l)$

  What is the concentration of a phosphoric acid solution if 25.0 mL are exactly neutralized by 20.0 mL of 2.000 M KOH solution?

- What is the concentration of a calcium hydroxide solution if 15.0 mL are exactly neutralized by 10.00 mL of 0.250 M HCl solution?

# Practice worksheets to be turned in:

CHAPTER 15 REVIEW

## *Acid-Base Titration and pH*

## SECTION 1

**SHORT ANSWER Answer the following questions in the space provided.**

1.Calculate the following values without using a calculator.

a.The $[H_3O^+]$ is $1 \times 10^{-6}$ M in a solution. Calculate the $[OH^-]$.

b.The $[H_3O^+]$ is $1 \times 10^{-9}$ M in a solution. Calculate the $[OH^-]$.

c.The $[OH^-]$ is $1 \times 10^{-12}$ M in a solution. Calculate the $[H_3O^+]$.

d.The $[OH^-]$ in part c is reduced by half, to
$0.5 \times 10^{-12}$ M. Calculate the $[H_3O^+]$.

e.The $[H_3O^+]$ and $[OH^-]$ are _____ (directly, inversely, or not) proportional in any system involving water.

2.Calculate the following values without using a calculator.

a.The pH of a solution is 2.0. Calculate the pOH.

b.The pOH of a solution is 4.73. Calculate the pH.

c.The $[H_3O^+]$ in a solution is $1 \times 10^{-3}$ M. Calculate the pH.

d. The pOH of a solution is 5.0. Calculate the [OH–].

e. The pH of a solution is 1.0. Calculate the [OH–].

3. Calculate the following values.

a. The [$H_3O^+$] is $2.34 \times 10^{-5}$ M in a solution. Calculate the pH.

b. The pOH of a solution is 3.5. Calculate the [$OH^-$].

c. The [$H_3O^+$] is $4.6 \times 10^{-8}$ M in a solution. Calculate the [$OH^-$].

**PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.**

4. [$H_3O^+$] in an aqueous solution = $2.3 \times 10^{-3}$ M.

a. Calculate [$OH^-$] in this solution.

## SECTION 1 continued

b. Calculate the pH of this solution.

c. Calculate the pOH of this solution.

d. Is the solution acidic, basic, or neutral? Explain your answer.

5. Consider a dilute solution of 0.025 M $Ba(OH)_2$ in answering the following questions.

a. What is the [$OH^-$] in this solution? Explain your answer.

b. What is the pH of this solution?

6. Vinegar purchased in a store may contain 6 g of $CH_3COOH$ per 100 mL of solution.

a. What is the molarity of the solute?

b. The actual $[H_3O^+]$ in the vinegar solution in part a is $4.2 \times 10^{-3}$ M. In this solution, has more than 1% or less than 1% of the acetic acid ionized? Explain your answer.

c. Is acetic acid strong or weak, based on the ionization information from part b?

d. What is the pH of this vinegar solution?

CHAPTER 15 REVIEW

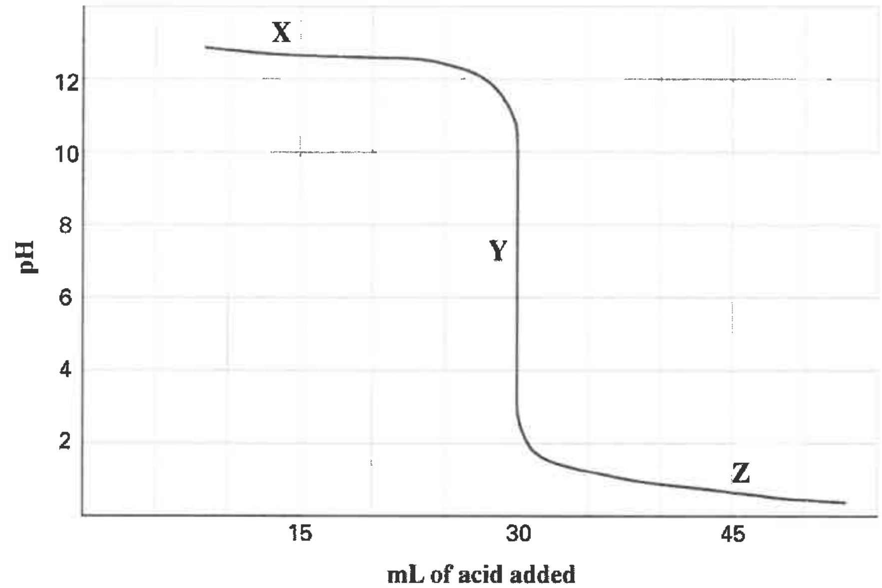
# Acid-Base Titration and pH

## SECTION 2

**SHORT ANSWER Answer the following questions in the space provided.**

1. Below is a pH curve from an acid-base titration. On it are labeled three points: X, Y, and Z.

## *Acid-Base Titration Curve*




a. Which point represents the equivalence point?

b. At which point is there excess acid in the system?

c. At which point is there excess base in the system?

d. If the base solution is 0.250 M and there is one equivalent of OH– ions for each mole of base, how many moles of OH– ions are consumed at the end point of the titration?

**PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.**

2. A standardized solution of 0.065 M HCl is titrated with a saturated solution of calcium hydroxide to determine its molarity and its solubility. It takes 25.0 mL of the base to neutralize 10.0 mL of the acid.

a. Write the balanced molecular equation for this neutralization reaction.

SECTION 2 **continued**

b.Determine the molarity of the $Ca(OH)_2$ solution.

c.Based on your answer to part b, calculate the solubility of the base in grams per liter of solution. (Hint: What is the concentration of $Ca(OH)_2$ in the saturated solution?)

3.It is possible to carry out a titration without any indicator. Instead, a pH probe is immersed in a beaker containing the solution of unknown molarity, and a solution of known molarity is slowly added from a buret. Use the titration data below to answer the following questions.

Volume of KOH(*aq*) in the beaker = 30.0 mL

Molarity of HCl(*aq*) in the buret = 0.50 M

At the instant pH falls from 10 to 4, the volume of acid added to KOH = 27.8 mL.

a.What is the mole ratio of chemical equivalents in this system?

b.Calculate the molarity of the KOH solution, based on the above data.