7. What is the pH of a solution that contains 10.0 g of nitric acid (HNO_3) and 15.0 g of hydrochloric acid (HCl) dissolved in 1000 liters of water?

Lesson 6.10 hwk + review answers

Lesson 6.10: Neutralization and Solution Stoichiometry

What happens when an acid is mixed with a base? NEUTRALIZATION!

Products of Neutralization: The products of acid-base neutralization are always a metallic salt and H_2O.

The definition of a metallic salt is a class of compounds formed when the hydrogen ion of an acid is partly or wholly replaced by a metal. In order for neutralization to occur an equal number of moles of acid and base must combine, so that every 1 mole of H^+ combines with exactly 1 mole OH^- to form 1 mole of H_2O. Likewise, 1 mole of negative anion from the acid combines with 1 mole of positive cation from the base to form 1 mole of metallic salt.

It is a misconception that acid-base neutralization *always* yields a pH of 7. The resulting pH is 7 **only if** the number of moles of acid and base are exactly enough to react completely with each other. If there is any excess unreacted acid (or base) left over, it dictates the resulting pH of the solution. Excess unreacted hydronium ion, for example, will make the pH after neutralization be under 7. It is more correct to say that pH after neutralization will be *closer* to 7.

Examples of neutralizations:

$HCl + NaOH \rightarrow NaCl + H_2O$ $\quad$ $H_2SO_4 + Ca(OH)_2 \rightarrow CaSO_4 + 2\,H_2O$ $\quad$ $2\,HNO_3 + Mg(OH)_2 \rightarrow Mg(NO_3)_2 + 2\,H_2O$

All species are (aq) except $H_2O(l)$

Review: Write the balanced neutralization equation for each reaction. Assume that ionization is complete for all acids.

1. Hydrobromic acid and sodium hydroxide

   $HBr + NaOH \rightarrow NaBr + H_2O$

2. Hydroiodic acid and calcium hydroxide

   $2HI + Ca(OH)_2 \rightarrow CaI_2 + 2H_2O$

3. Sulfuric acid and rubidium hydroxide

   $H_2SO_4 + 2\,RbOH \rightarrow Rb_2SO_4 + 2H_2O$

In solution stoichiometry (including acid-base stoichiometry), if you know the molarity and volume of a reactant or product in aqueous solution, you can calculate how many moles of solute are contained in it.

 # moles = molarity (in M) x volume (in L)

From there you can calculate the # of moles of any other reactant or product in the chemical reaction. If precipitation occurs, you can calculate and predict the mass of precipitate that should be recovered.

Example: 150. mL of 0.100 M calcium nitrate solution is mixed with excess sodium hydroxide solution.

a) Write a balanced chemical equation for this process. (Hint: use your solubility rules.)

b) How many moles of calcium nitrate will react?

c) How many moles of precipitate will form?

don't worry

d) What is the mass of the precipitate?

<u>Review</u>: If 250. mL of 0.300 M sodium iodide react with excess lead (II) nitrate, how many grams of precipitate will form?

In acid-base neutralization, which is really just another type of solution stoichiometry, we are most often concerned with knowing about the other reactant as opposed to the yield of some product (although you could calculate that as well, if you really wanted to). In the food industry, which handles acids and bases quite often, a common measure of quality control is to identify the concentration of acid in a product, such as ketchup or vinegar. Ingredient quantities must be consistent, so samples of the same food are tested, over and over, to ensure that acid concentrations are always the same.

For neutralization: moles of acid = moles of base

$M_aV_an_b = M_bV_bn_a$ where M = molarity, n = mols of H^+ or OH^- **V =** volume

Example: What volume of 2.00 M (M_a) sulfuric acid would be required to neutralize 54.1 mL (V_b) of 1.40 M (M_b) calcium hydroxide?

$$\underset{\uparrow n_a}{\_}\,H_2SO_4 + \underset{\uparrow n_b}{\_}\,Ca(OH)_2 \rightarrow \underline{2}\,H_2O + CaSO_4$$

$$M_aV_an_b = M_bV_bn_a$$

$$(2.00\,M)(V_a)(1) = (1.40\,M)(.0541\,L)(1)$$

$$V_a = \frac{(1.40)(.0541)(1)}{(2.00)(1)} = .0379L = 37.9\,mL$$

Ma? Va Vb Mb
Example: Calculate the molarity of 75.0 mL hydrochloric acid that could neutralize 49.1 mL of 1.40 M $Ca(OH)_2$.

$$\underline{2}\ HCl + \underline{\ }\ Ca(OH)_2 \longrightarrow 2\ H_2O + CaCl_2$$

n_a n_b

$$M_aV_an_b = M_bV_bn_a$$

$$M_a = \frac{(1.40\,M)(.0491\,L)(2)}{(.0750\,L)(1)} = 1.83\,M$$

Review:

Va↓ H3A Vb↓ Mb↓
1) It was found that 33.8 mL of a triprotic acid was required to neutralize 43.1 mL of 1.10 M NaOH.
What is the molarity of the acid?
Ma?

$$\underline{\ }\ H_3A + \underline{3}\ NaOH \longrightarrow 3\ H_2O + Na_3A$$

n_a n_b

$$M_a = \frac{M_bV_bn_a}{V_an_b} = \frac{(1.10\,M)(.0431\,L)(1)}{(.0338\,L)(3)} = .468\,M$$

Ma↓ Vb↓
2) Industrially produced nitric acid is 8.0 M. What volume of nitric acid is needed to exactly neutralize 37.9 mL of 3.50 M calcium hydroxide spill?
Va?
Mb↑

$$\underline{2}\ HNO_3 + \underline{\ }\ Ca(OH)_2 \longrightarrow Ca(NO_3)_2 + 2\ H_2O$$

n_a n_b

$$V_a = \frac{M_bV_bn_a}{M_an_b} = \frac{(3.50\,M)(.0379\,L)(2)}{(8.0\,M)(1)} = .033\,L = 33\,mL$$

Titration is a widely used quantitative method of analyzing an unknown solution to determine its molarity. Let's say we want to determine the molarity of an acid.

- A known volume of that acid is placed in an Erlenmeyer flask, along with a few drops of a chemical called an **indicator**. The indicator changes color when the reaction between the acid and base has completed. This is only needed if the acid does not change color on its own. Two common indicators are: **phenolphthalein,** which is colorless in acids but turns pink when the solution is basic; and **bromothymol blue,** which is yellow in acids and blue in bases.

- The point at which the indicator changes color is called the **endpoint**. The endpoint occurs very close to the **equivalence point**, which is the point when the reaction has completed.

- Prepare the buret. (See inset.)

- A known volume and concentration of a base are placed in a buret. Take note of the initial volume of liquid in the buret. Read one uncertain digit and read the bottom of the meniscus. You turn the stopcock to control the flow of buret solution into the Erlenmeyer flask.

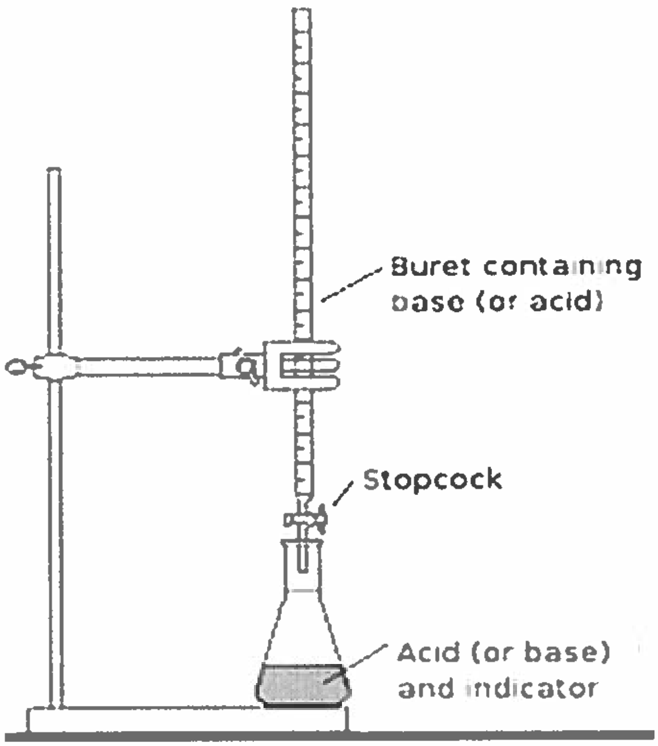
- When the indicator in the flask has just barely changed color, stop adding base. Any hanging drops on the inside walls of the Erlenmeyer flask and from the buret tip must be put in the Erlenmeyer flask. Take note of the final volume of liquid in the buret.

Technique: Preparing the Buret for Titration

1. Rinse a clean buret several times with 5 mL portions of distilled water and then with portions of base.

2. Allow the titrant to drain through the stopcock so that the tip gets rinsed with titrant as well.

3. Discard the rinse solution in a waste beaker. Clamp the buret into place, and fill it with the base. Remove air bubbles from the tip of the buret and the stopcock by draining several milliliters of titrant. Dispose of the drained titrant in a waste beaker.



Solving Titration Problems

1. Subtract the initial buret reading and the final buret reading.
2. You should know the buret solution molarity. Calculate the # moles of solute delivered to the flask.
3. Do solution stoichiometry to determine the # moles and molarity of the unknown solution in the flask.

M_b

Example: The initial volume reading on a buret filled with 0.10 M sodium hydroxide is 1.52 mL, and when the endpoint is reached, it reads 54.52 mL. What is the molarity of 125 mL HCl solution titrated by this solution? $V_b = 54.52 - 1.52 = 53.00\,mL$ V_a M_a?

$$\underset{n_a}{\_}\,HCl + \underset{n_b}{\_}\,NaOH \longrightarrow NaCl + H_2O$$

$$M_a = \frac{M_b V_b n_a}{V_a n_b} = \frac{(.10\,M)(.05300\,L)(1)}{(.125\,L)(1)} = .042\,M$$

M_b V_a

Example: 0.50 M calcium hydroxide solution is used to titrate 75.20 mL of perchloric acid. What is the molarity of the acid? M_a?

$$\underset{n_b}{\_}\,Ca(OH)_2 + \underset{n_a}{2}\,HClO_4 \longrightarrow Ca(ClO_4)_2 + 2H_2O$$

$Ca(OH)_2$	Volume (mL)
Initial volume	17.40 mL
Final volume	49.32 mL

$V_b = 31.92\,mL$

$$M_a = \frac{M_b V_b n_a}{V_a n_b} = \frac{(0.50\,M)(.03192\,L)(2)}{(.07520\,L)(1)} = .42\,M$$

Review

1. The concentration of a solution of potassium hydroxide is determined by titration with nitric acid. A 25.0 mL sample of potassium hydroxide reaches endpoint when 35.6 mL of 0.562 M nitric acid is added. What is the concentration of the potassium hydroxide solution? M_b?

$$\underset{n_b}{\_}\,KOH + \underset{n_a}{\_}\,HNO_3 \longrightarrow KNO_3 + H_2O$$

$$M_b = \frac{M_a V_a n_b}{V_b n_a} = \frac{(.562\,M)(.0356\,L)(1)}{(.0250\,L)(1)} = .800\,M$$

* Let's say your $M_b = 0.100\ M$

$NaOH + HC_2H_3O_2 \longrightarrow NaC_2H_3O_2 + H_2O$

2. In a titration of a sample of vinegar with sodium hydroxide, your initial buret reading is 17.05 mL and your final buret reading is 28.27 mL. A 10.0 mL sample of vinegar (acetic acid [CH_3COOH]) was titrated. (What is the molarity of the acid?) $V_b = 28.27 - 17.05 = 11.22\ mL$ M_a?

$$M_a = \frac{M_b V_b n_a}{V_a n_b} = \frac{(0.100\ M)(.01122\ L)(1)}{(.0100\ L)(1)} = .112\ M$$

3. What is the concentration of acid in rainwater when 100.0 mL is titrated with 25.12 mL of 0.00105 M NaOH? Since acid rain contains several acids, assume the acid is monoprotic and use the symbol HA to represent the acids that are present. M_a?

$\underset{\uparrow n_a}{\_}HA + \underset{\uparrow n_b}{\_}NaOH \longrightarrow H_2O + NaA$

$$M_a = \frac{M_b V_b n_a}{V_a n_b} = \frac{(0.00105\ M)(.02512\ L)(1)}{(.100\ L)(1)} = 2.638 \times 10^{-4}\ M$$

4. A 0.250 M solution of barium hydroxide neutralizes 25.0 mL stomach acid. What is the molarity of the HCl in the solution? Assume stomach acid is the HCl solution. M_a?

$\underset{\uparrow n_b}{\_}Ba(OH)_2 + \underset{\uparrow n_a}{2}HCl \longrightarrow BaCl_2 + H_2O$

$Ba(OH)_2$	Volume (mL)
Initial	3.20 mL
Final	47.02 mL

$V_b = 47.02 - 3.20 = 43.82\ mL$

$$M_a = \frac{M_b V_b n_a}{V_a n_b} = \frac{(0.250\ M)(.04382)(2)}{(.0250\ L)(1)} = .876\ M$$

Unit 6 Homework

Lesson 6.1

1. The boiling points of the following liquids increase in the order in which they are listed below:

$CH_4 < H_2S < NH_3$

Discuss the theoretical considerations involved and use them to account for this order.

2.

Substance	Melting Point, ºC
H_2	-259
C_3H_8	-190
HF	-92
CsI	621

Discuss how the trend in the melting points of the substances tabulated above can be explained in terms of the types of attractive forces and/or bonds in these substances.

3. Alcohol dissolves in water to give a solution that boils at a lower temperature than pure water. Salt dissolves in water to give a solution that boils at a higher temperature than pure water. Explain these

Lesson 6.10

1. Write the **balanced** equation for these neutralization reactions: (remember to balance charges for ionic compounds!)

   a) HCl + NaOH → $NaCl + H_2O$

   b) HNO_3 + KOH → $H_2O + KNO_3$

   c) $Ca(OH)_2 + H_2SO_4$ → $CaSO_4 + 2H_2O$

2. Solve the following solution stoichiometry problems.

   a) How many grams of precipitate form after the complete reaction of 10.0 mL of 0.500 M barium nitrate and 20.0 mL of 0.250 M sodium sulfate? (Warning: limiting/excess)

   don't worry

   M_a ? V_a↓ V_b↓ M_b↓

   b) Calculate the molarity of phosphoric acid when 31.1 mL of it is neutralized by 21.3 mL of 0.357 M potassium hydroxide.

   $$\underset{\uparrow n_a}{\_}H_3PO_4 + \underset{\uparrow n_b}{3}KOH \longrightarrow H_2O + K_3PO_4$$

   $$M_a = \frac{M_b V_b n_a}{V_a n_b} = \frac{(0.357M)(.0213L)(1)}{(0.0311L)(3)}$$

   $$= .0815\ M$$

   V_b ?

   c) What volume of 0.137 M rubidium hydroxide would be used to neutralize 15.3 mL of 0.0592 M sulfuric acid? (Assume complete ionization of the acid.)

   $$\underset{\uparrow n_b}{2}RbOH + \underset{\uparrow n_a}{\_}H_2SO_4 \longrightarrow Rb_2SO_4 + 2H_2O$$

   $$V_b = \frac{M_a V_a n_b}{M_b n_a} = \frac{(0.0592M)(.0153L)(2)}{(0.137M)(1)} = 0.0132\ L = 13.2\ mL$$

   d) 14.3 mL of calcium hydroxide solution is placed in an Erlenmeyer flask, along with a few drops of phenolphthalein indicator. The buret is filled with 0.400 M nitric acid. Figure 1 is the initial buret reading, and Figure 2 is the final buret reading after the flask solution has turned pale pink. Calculate the molarity of calcium hydroxide in the flask.

   $V_b = 14.3mL = .0143L$

   $M_a = 0.400\ M$

   $$\underset{\uparrow n_b}{\_}Ca(OH)_2 + \underset{\uparrow n_a}{2}HNO_3 \longrightarrow Ca(NO_3)_2 + 2H_2O$$

   $$M_b = \frac{M_a V_a n_b}{V_b n_a} = \frac{(0.400M)(.02278L)(1)}{(.0143L)(2)}$$

   $$= .319\ M$$

You can read from the top or from the bottom — just be consistent

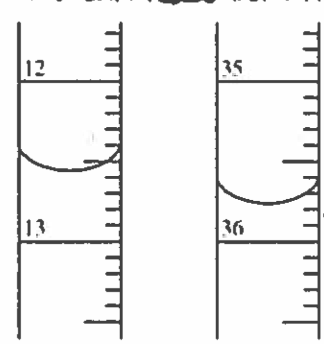


Figure 1 Figure 2

13.45mL 36.23mL

$V_a = 36.23 - 13.45 = 22.78mL$