Acids and Bases

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CHAPTER **1**

Acids and Bases

CHAPTER OUTLINE

Aqua regia, a mixture of nitric acid and hydrochloric acid, is one of the few materials that will dissolve gold. This material, translated to "royal water" in English, was named so because it could dissolve the royal metal gold. First noted in the fourteenth century, aqua regia could be used to help ascertain whether a particular material was actually gold or was some trickery of the alchemist. Nitric acid by itself will not dissolve gold, but will in combination with hydrochloric acid. The chemistry of the process is rather complex, with both acids reacting with the metal to form soluble gold compounds. The gold can be recovered from the solution, making the process useful for purification purposes.

Acids and bases have had many uses throughout history. In this chapter we want to explore the properties of acids and bases and the reactions in which they take part.

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1.1 Acid-Base Definitions

Lesson Objectives

- List some common properties of acids and bases.
- Describe how acids and bases are defined based on the Arrhenius, Bronsted-Lowry, and Lewis definitions. Understand the differences between these three systems.

Lesson Vocabulary

- **Arrhenius acid**: A compound that releases H^+ ions when dissolved in water.
- **Arrhenius base**: A compound that generates hydroxide ions (OH^-) when dissolved in water.
- **Brønsted-Lowry acid**: A compound that can donate a proton (an H^+ ion) to an appropriate acceptor.
- **Brønsted-Lowry base**: A compound that can remove a proton from a relatively strong Brønsted-Lowry acid.
- **Lewis acid**: A chemical species that accepts a pair of electrons.
- **Lewis base**: A chemical species that donates a pair of electrons.
- **monoprotic acid**: A compound with only one acidic hydrogen atom that can be transferred to a strong base.
- **polyprotic acid**: A compound with two or more acidic hydrogen atoms that can be transferred to a strong base.

Check Your Understanding

- What comes to mind when you think of acids? How does a base differ from an acid?

Introduction

When you think of acids, "sour" is likely to come to mind. For the ancient Greeks, "sour-tasting" is in fact what defined a category of substances which later became known as acids. Bases were first categorized by their uses in soaps, as well as their ability to counteract properties of acids. In this lesson, you will learn about a few definitions for acids and bases as well as some of their properties.

Properties of Acids and Bases

Acids and bases are versatile and useful materials in many chemical reactions. Some properties that are common to aqueous solutions of acids and bases are listed in **Table** 1.1.

TABLE 1.1: Properties of acids and bases

Acids	Bases
conduct electricity in solution	conduct electricity in solution
turn blue litmus paper red	turn red litmus paper blue
have a sour taste	have a slippery feeling
react with bases to create a neutral solution	react with acids to create a neutral solution
react with active metals to produce hydrogen gas	

Note: Litmus paper is a type of treated paper that changes color based on the acidity of the solution it comes in contact with.
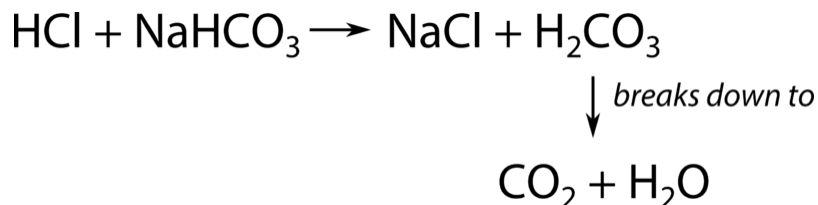
Defining Acids and Bases

An early way of classifying acids and bases was proposed by Svante Arrhenius, a Swedish chemist, in 1894. An **Arrhenius acid** is any compound that releases H^+ ions when dissolved in water. An **Arrhenius base** is a compound that generates hydroxide ions (OH^-) when dissolved in water. Some representative examples are given in the **Table 1.2**.

TABLE 1.2: Arrhenius acids and bases

Acids	Bases
hydrochloric acid: $HCl \rightarrow H^+ + Cl^-$	sodium hydroxide: $NaOH \rightarrow Na^+ + OH^-$
nitric acid: $HNO^3 \rightarrow^H + + NO_3^-$	potassium hydroxide: $KOH \rightarrow K^+ + OH^-$
hydrobromic acid: $HBr \rightarrow H^+ + Br^-$	calcium hydroxide: $Ca(OH)_2 \rightarrow Ca^{2+} + 2\ OH^-$

Many strong acids and bases can be identified based on the Arrhenius model. However, there are many compounds that share a number of common characteristics with acids and bases but do not fit the Arrhenius definitions. In the early 1920s, the Danish scientist Johannes Brønsted and the English researcher Thomas Lowry each published ideas that expanded the Arrhenius concept. According to this newer definition, a **Brønsted-Lowry acid** is any compound that can donate a proton (an H^+ ion) to an appropriate acceptor. A **Brønsted-Lowry base** is a compound that can remove (or accept) a proton from a relatively Brønsted-Lowry acid.

Overall, the Brønsted-Lowry model suggested that any acid-base reaction could be reduced to the transfer of a proton from an acid to a base. For example, the reaction between hydrochloric acid (HCl) and sodium bicarbonate ($NaHCO_3$) involves the transfer of an H^+ ion from the acid (HCl) to the base (the bicarbonate ion). The resulting carbonic acid (H_2CO_3) is unstable and breaks down to form carbon dioxide and water:

$$HCl + NaHCO_3 \longrightarrow NaCl + H_2CO_3$$

$$\downarrow \textit{breaks down to}$$

$$CO_2 + H_2O$$

Another more general definition of acids and bases was offered by the American chemist G.N. Lewis. A **Lewis acid** is any chemical species that accepts a pair of electrons, and a **Lewis base** is a chemical species that donates a pair of electrons. This is the broadest most commonly used definition, and all compounds that qualify as an acid or base under the previous definitions are also Lewis acids and bases.
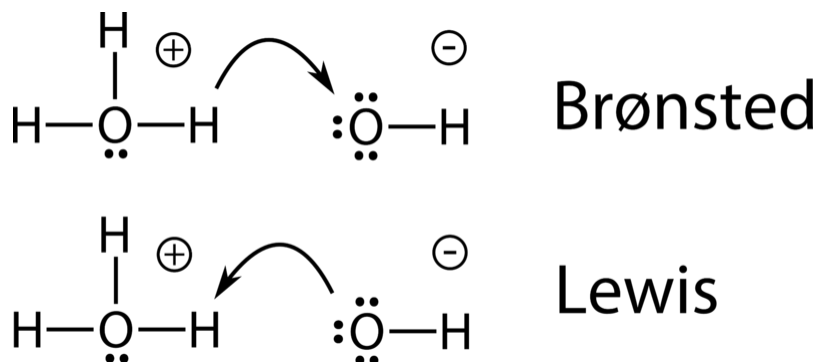
In the diagram above, we see the same process illustrated multiple times, highlighting the ways in which the compounds are acting as an acid or a base according to each definition. According to the Brønsted-Lowry model, the protonated water molecule (the acid) is donating a proton to the OH^- ion (the base). According to the Lewis model, the hydroxide ion has a pair of electrons (indicated by the black bar) that it donates to the protonated water molecule. "Pronated" refers to the extra hydrogen atom in the molecule, which increases its charge from neutral to +1. In both instances, the hydroxide ion serves as a base and the protonated water molecule is the acid.

An example of a Lewis acid-base reaction that would not fit the other definitions of acid and base is the formation of an adduct between boron trifluoride and ammonia:
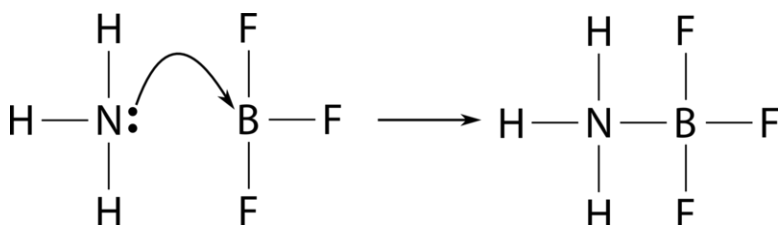

FIGURE 1.1

The ammonia serves as a Lewis base by donating its lone pair of electrons to make a new bond with boron (the Lewis acid), which has an empty orbital (indicated by ○) that can accept two electrons.

Unless otherwise indicated, we will be using the Brønsted-Lowry model of acids and bases for the remainder of the chapter.

Monoprotic and Polyprotic Acids

Acids can further be categorized based on how many acidic hydrogen atoms they contain. Acidic hydrogen atoms are those which will be transferred to a base. A **monoprotic acid** has only one acidic hydrogen that would be transferred to a strong base, whereas a **polyprotic acid** has two or more. Common monoprotic acids include HCl, HBr, and HNO_3. A common diprotic acid is sulfuric acid (H_2SO_4), and phosphoric acid (H_3PO_4) provides an example of a triprotic acid. In each case, all hydrogens are available to participate in acid-base reactions. However, that is not the case for all acidic molecules. For example, in acetic acid (CH_3COOH), only the hydrogen bonded to the oxygen atom is acidic. The other three hydrogens are covalently bonded to carbon and cannot be removed by any of the bases that we will consider in this chapter.

TABLE 1.3: Acids

Name	Structure
hydrobromic acid	HBr

TABLE 1.3: (continued)

Name	Structure
hydrochloric acid	HCl
hydrofluoric acid	HF
hydroiodic acid	HI
nitric acid	HNO_3
perchloric acid	$HClO_4$
acetic acid	CH_3COOH

TABLE 1.4:

Name	Structure
carbonic acid	H_2CO_3
sulfuric acid	H_2SO_4
sulfurous acid	H_2SO_3

TABLE 1.5:

Name	Structure
phosphoric acid	H_3PO_4

Lesson Summary

- Acids and bases were originally identified based on their physical and chemical properties.
- According to the Arrhenius model, acids and bases are compounds that release H^+ or OH^- ions when dissolved in water.
- In the Brønsted-Lowry model, an acid is a compound that can donate a proton (H^+ ion), and a base is a compound that can accept a proton.
- A Lewis base is an electron-pair donor, whereas a Lewis acid is an electron-pair acceptor. This is the broadest of the three definitions.
- Acids can be classified as monoprotic or polyprotic based on the number of acidic hydrogens they contain.

Lesson Review Questions

1. List three characteristics exhibited by acids and three exhibited by bases.
2. Based on the given acid and base characteristics, can you think of a few substances that might be classified as acids and a few that might be classified as bases?
3. Based on Arrhenius' definitions, how can you tell the difference between an acidic substance and a basic substance?
4. Why were Arrhenius' definitions expanded upon?
5. Based on Brønsted-Lowry definitions, how can you tell the difference between an acidic substance and a basic substance?
6. What definition is most widely applicable in defining acids and bases?
7. Describe the difference between a monoprotic acid and a polyprotic acid.

8. The formula for propanoic acid is CH_3CH_2COOH. Is this acid monoprotic or polyprotic? Explain your answer.

Further Reading/Supplementary Links

- Summary of acid and base properties: http://www.chemtutor.com/acid.htm
- Strong and weak acids and bases: http://www.highlands.edu/academics/divisions/scipe/chemistry/Site/GH andouts_files/Strong%20and%20Weak%20Acids%20and%20Bases.pdf

Points to Consider

- How can we tell how acidic or basic a solution is?

1.2 Acid-Base Neutralization Reactions and Titrations

Lesson Objectives

- Define and give examples of acid-base neutralization reactions. Be able to write both molecular and net ionic equations for these reactions.
- Describe how antacids deal with stomach acidity.
- Describe the titration process when using an indicator.
- Describe the titration process when using a pH meter.
- Describe titration curves for both monoprotic and polyprotic acids.

Lesson Vocabulary

- **neutralization**: The reaction between an acid and a base in which neutral water is produced.
- **titration**: A laboratory technique that very accurately measures the concentration of a solution of an acid or a base.
- **titrant**: The solution of unknown concentration used in a titration.
- **equivalence point**: The point at which the moles of acid that were initially present in the flask are equal to the moles of base that were added via the buret.
- **titration curve**: A plot of pH vs. the amount of titrant added.

Check Your Understanding

- What information can we obtain from an acid equilibrium constant?

Introduction

If you have ever been stung by a bee, you've likely experienced the painful burning sensation left by the stinger. This burning or stinging is caused by an acid in the bee's stinger. One way to relieve this discomfort is to apply a paste made of baking soda to the sting. Baking soda contains a weak base called sodium hydrogen carbonate which reacts with the acid and helps relieve the sting. In this lesson you will learn about various acid-base reactions and how they work.

Neutralization of Acids and Bases
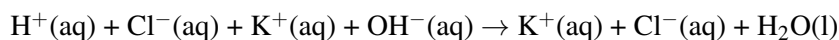
The Neutralization Reaction

So far, we have been looking at the interaction between acids and water (in which water acts as a very weak base) or bases and water (in which water acts as a very weak acid). However, acids and bases can react directly with each other as well. A typical acid-base reaction would have the following form:

acid + base → salt + water
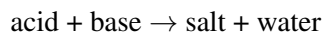
where the term salt can refer to essentially any ionic compound. An example would be the reaction between hydrochloric acid and potassium hydroxide:

$HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(l)$
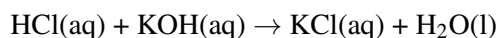
Recall from the chapter *Chemical Reactions* that there are several different ways to write reactions between ionic substances that take place in water. The equation above is the molecular equation for this reaction. We can also write the ionic equation as follows:

$H^+(aq) + Cl^-(aq) + K^+(aq) + OH^-(aq) \rightarrow K^+(aq) + Cl^-(aq) + H_2O(l)$

Removing the spectator ions, which do not change over the course of the reaction, we get the following net ionic equation:

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
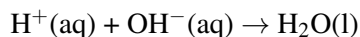
The reaction between an acid and a base is referred to as a **neutralization** reaction, since the more reactive H^+ and OH^- ions combine to produce neutral water. When equal amounts of a strong acid and a strong base are combined, we are left with a neutral solution (pH = 7).

How Antacids Work

Have you ever experienced heartburn? Heartburn is actually a condition caused by stomach acid, rather something in your heart. The stomach produces HCl, which is a necessary component of digestion. However, excessive amounts of stomach acid can cause a great deal of discomfort. Various factors can cause the overproduction of acid, mostly related to the types (and amounts) of food and beverages that are being consumed. Caffeine and other oils in coffee (even decaf) will increase HCl production in the stomach. Smoking cigarettes and drinking alcohol also increase the formation of stomach acid.

To relieve the discomfort associated with excess stomach acid, many people turn to antacids. Antacids are essentially just bases that react with some of the unwanted HCl via a neutralization reaction. Strong bases like NaOH are far too harsh to ingest, but some less soluble hydroxide bases can be consumed as suspensions. For example, milk of magnesia is a suspension of $Mg(OH)_2$. $Al(OH)_3$ is also available as a safe formulation. Both of these compounds will remain largely undissolved until they reach the stomach, where they will react with the excess acid without subjecting other parts of the body to strongly basic conditions. However, these can lead to other types of issues with continued use, such as constipation, kidney damage, and possibly even bone depletion.

These days, the most common antacids use weak bases, such as carbonate or bicarbonate ions, to neutralize the HCl. The resulting carbonic acid can then decompose into water and carbon dioxide. For example, drinking a solution of sodium bicarbonate (baking soda) will result in the following reactions:

FIGURE 1.2
Many antacid tablets use calcium carbonate as an active ingredient.

HCl(aq) + NaHCO$_3$(aq) $\rightarrow$ NaCl(aq) + H$_2$CO$_3$(aq)
H$_2$CO$_3$(aq) $\rightarrow$ H$_2$O(l) + CO$_2$(g)
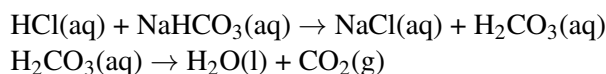
It should be noted that a newer class of antacids does not work by directly neutralizing stomach acid. Instead, they act to inhibit the proton pumps that generate the acid in the first place. Proton pump inhibitors decrease the activity of these pumps and lower the amount of acid produced by the stomach, but they do not neutralize the acid already present.

Acid-Base Titrations

A **titration** is a laboratory technique that very accurately measures the concentration of a solution of acid or base. It makes use of a neutralization reaction and the fact that pH changes very rapidly for neutral (and nearly neutral) solutions.

The following is a typical procedure for titration of a strong acid with a strong base: Let's say you have a solution of HCl, but the concentration is not known. First, a known volume of the acid is placed in a flask, and a few drops of an acid-base indicator, such as phenolphthalein, are added. Because this solution is acidic, and phenolphthalein is colorless at low pH values, the solution remains clear.

Next, a solution of a strong base, such as NaOH is placed into a piece of glassware called a buret (**Figure** 1.3). This solution, which has a known concentration, is referred to as the **titrant**. Then, the titrant is added drop by drop to the acid while swirling the flask. Because phenolphthalein turns pink when exposed to base, the drops of the titrant may initially appear pink when they hit the acid, but after swirling, the base is neutralized by the HCl, and the pink color fades. Once the pink color no longer fades after swirling, the solution in the flask is basic (pH >7), which means all of the HCl has been neutralized by the NaOH titrant. The volume of the titrant that was added to the acid can be measured by looking at the initial and final levels of the solution in the buret.

At this point, called the **equivalence point**, the moles of acid that were initially present in the flask are equal to the number of moles of base that were added via the buret. Remember that to calculate the number of moles of solute, we multiply the concentration of the solution (in mol/L) by the volume of the solution (in L). At the equivalence point:

moles acid = moles base

FIGURE 1.3

$$M_{acid} \times V_{acid} = M_{base} \times V_{base}$$

where M and V are the concentration and volume, respectively, of the two solutions. After performing the titration above, we know the volume of the initial acid in the flask, the concentration of the base, and the volume of the titrant that was required to reach the equivalence point. We can therefore solve for the concentration of the HCl solution.

Example 21.6

A 25.0 mL sample of an HCl solution is titrated with 0.050 M NaOH. The equivalence point is reached after the addition of 46.7 mL of the NaOH solution. What is the concentration of the HCl solution?

Answer:

Simply plug these values into the equation above:

$$M_{acid} \times V_{acid} = M_{base} \times V_{base}$$
$$M_{acid} \times 25.0 \text{ mL} = 0.050 \text{ M} \times 46.7 \text{ mL}$$
$$M_{acid} = 0.093 \text{ M}$$

Titration Curves

If the pH of a solution is tracked with a pH meter over the course of the titration, we can get even more useful information about what is occurring in solution. A **titration curve** is a plot of pH vs. the amount of titrant added. The shape and position of the curve can be used not just to determine the equivalence point, but also to determine whether the acid is strong or weak and the K_a value (for weak acids).

Figure 1.4 shows two titration curves. The curve on the left represents the experiment we just described, in which a strong base (NaOH) is added to a strong acid (HCl). The volume of base added is plotted on the x-axis, and the pH obtained is plotted on the y-axis. We see a very sharp rise in pH as the equivalence point is reached. The dotted line shows us the volume of base that must be added to get to this point.

The graph on the right shows a titration of a strong base with a strong acid. Although the pH is decreasing rather than increasing, the principle is the same. At the equivalence point, the moles of added acid are equal to the moles of base that were initially present, so the same equation can be used to determine the unknown concentration.

This approach can also be used to titrate a weak acid. In **Figure** 1.5, we see a comparison between the titration curves obtained for HCl and for acetic acid (a weak acid). There are a couple of noteworthy differences. First, the
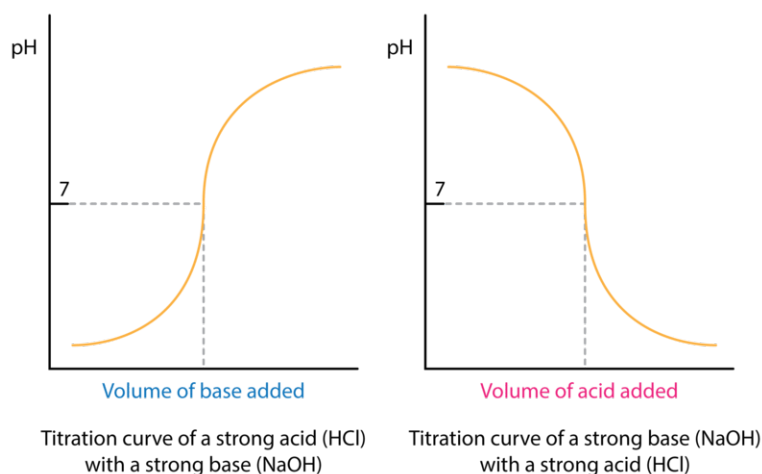
FIGURE 1.4

Titration curve of a strong acid (HCl) with a strong base (NaOH)

Titration curve of a strong base (NaOH) with a strong acid (HCl)

curve for the weak acid has a small plateau about halfway to the equivalence point. We will not go into the details, but as it turns out, the pH at the flattest point of this plateau can be used to determine the K_a of the acid according to the following equation:

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

Additionally, the pH at the equivalence point is higher than 7 when titrating a weak acid. This is because the anion from the weak acid (acetate) is slightly basic, whereas the anion from the strong acid (chloride) is an extremely poor base.
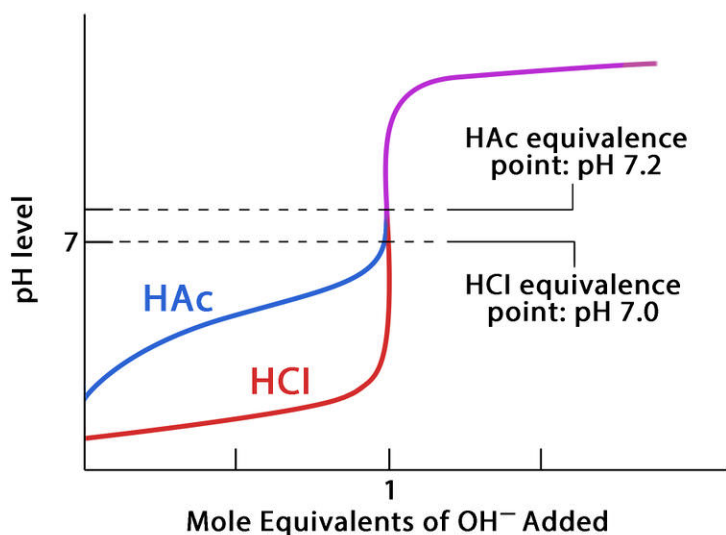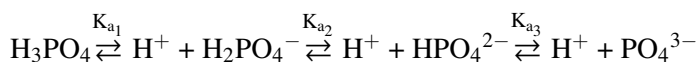

FIGURE 1.5

Polyprotic acids have distinctive titration curves. These acids have multiple dissociation constants, one for each acidic proton. For example, phosphoric acid has three acidic hydrogens:

$$H_3PO_4 \overset{K_{a_1}}{\rightleftarrows} H^+ + H_2PO_4^- \overset{K_{a_2}}{\rightleftarrows} H^+ + HPO_4^{2-} \overset{K_{a_3}}{\rightleftarrows} H^+ + PO_4^{3-}$$

$$K_{a_1} = 7.5 \times 10^{-3} \quad K_{a_2} = 6.2 \times 10^{-8} \quad K_{a_3} = 4.8 \times 10^{-13}$$

11

On the titration curve for phosphoric acid, we can see three vertical inflection points corresponding to the equivalence points for each acidic proton. (The third one is not very clear, because HPO_4^{2-} is not much more acidic than water.) As in the case of a monoprotic weak acid, the plateaus in the curve above can be used to determine the K_a values for each acidic hydrogen in phosphoric acid.
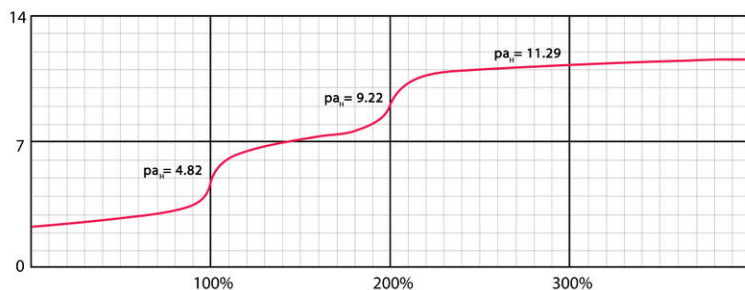

FIGURE 1.6

Lesson Summary

- Combining an acid and a base results in a neutralization reaction. For strong acids and bases, the net ionic equation is simply $H^+ + OH^- \rightarrow H_2O$.
- Weak bases or poorly soluble strong bases can be ingested to neutralize excess stomach acid.
- The concentration of an acid or base can be determined by titration.
- A pH meter allows the construction of a titration curve. The data from a titration curve can be used to determine the equivalence point, the type of acid or base (strong vs. weak), and the K_a or K_b value of the acid or base being titrated.

Lesson Review Questions

1. What are the products of a typical acid-base reaction?
2. Why are acid-base reactions referred to as neutralization reactions?
3. How do antacids relieve heartburn?
4. Read the ingredients on a container of antacids. Write equations for the neutralization of HCl based on those ingredients.
5. What property of near-neutral solutions do titration procedures rely on?
6. Define "titrant" and "indicator" as they pertain to titration procedures.
7. Estimate the pH at the equivalence point for the following acid-base titrations (i.e. =7, >7, <7):

    a. strong acid-strong base
    b. weak acid-strong base
    c. strong acid-weak base

8. 30. mL of an HBr solution is titrated with 0.075 M KOH. If 39.1 mL of KOH is required to reach the equivalence point, what is the concentration of the acid?
9. 45.0 mL of an NaOH solution is titrated with 0.12 M HCl. If 29.6 mL is required for neutralization, what is the concentration of the NaOH solution?
10. You have two weak acids in unlabeled bottles (a very bad practice, by the way). Explain how you could determine which one was acetic acid and which one was phosphoric acid.

Further Reading/Supplementary Links

- Acid and base titrations: http://www.sparknotes.com/chemistry/acidsbases/titrations/section1.html
- Neutralization of acids and bases: http://www.iun.edu/~cpanhd/C101webnotes/chemical%20reactions/acidbase.html
- Titration video: http://www.youtube.com/watch?v=g8jdCWC10vQ

Points to Consider

- Many biochemical processes require strict pH control. How can we take advantage of the properties of weak acids and bases to stabilize a solution at a specific pH?

1.3 References

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