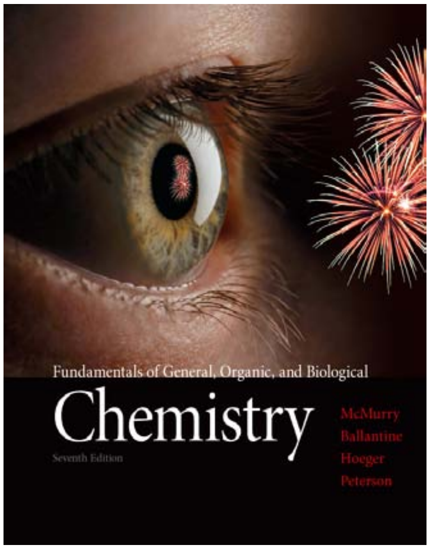
Chapter 10 Lecture



Fundamentals of General, Organic, and Biological Chemistry

7th Edition

McMurry, Ballantine, Hoeger, Peterson

Chapter Ten

Acids and Bases

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Outline

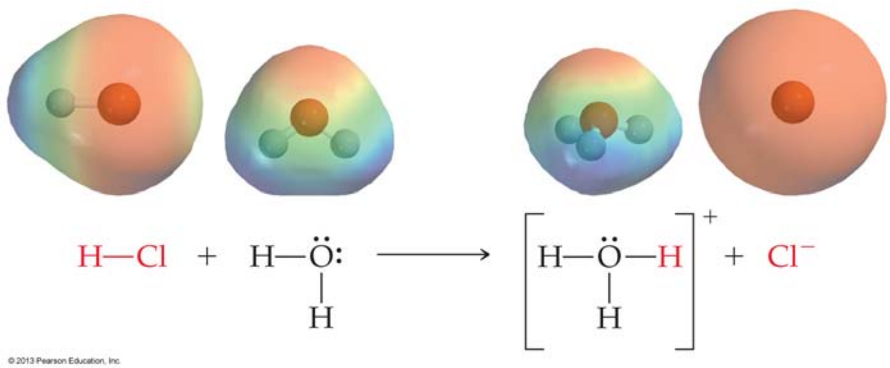
Goals

1. **What are acids and bases?**
   Be able to recognize acids and bases and write equations for common acid–base reactions.
2. **What effect does the strength of acids and bases have on their reactions?**
   Be able to interpret acid strength using acid dissociation constants K_a and predict the favored direction of acid–base equilibria.
3. **What is the ion-product constant for water?**
   Be able to write the equation for this constant and use it to find the concentration of H^+ or OH^-.
4. **What is the pH scale for measuring acidity?**
   Be able to explain the pH scale and find pH from the H_3O^+ concentration.
5. **What a buffer?**
   Be able to explain how a buffer maintains pH and how the bicarbonate buffer functions in the body.
6. **How is the acid or base concentration of a solution determined?**
   Be able to explain how a titration procedure works and use the results of a titration to calculate acid or base concentration in a solution.

10.1 Acids and Bases in Aqueous Solution

- An acid is a substance that produces hydrogen ions, H^+, when dissolved in water.
- A base is a substance that produces hydroxide ions, OH^-, when dissolved in water.
- The neutralization reaction of an acid with a base yields water plus a *salt*.
- These definitions are limited to reactions that take place in aqueous solution.
- The H^+ ion is so reactive it does not exist in water. It reacts with H_2O to give the **hydronium ion**, H_3O^+.

10.1 Acids and Bases in Aqueous Solution



- When gaseous HCl dissolves in water, H_3O^+ and Cl^- are formed.

10.1 Acids and Bases in Aqueous Solution

- In practice, the notations and H^+(*aq*) are often used interchangeably.
- OH^- ions can come from two sources.
- Metal hydroxides are ionic compounds that release OH^- ions when they dissolve in water.
- Some molecular compounds such as ammonia, can act as bases to produce OH^- ions in reactions with water.

10.2 Some Common Acids and Bases

- Acids generally have a sour taste:
- Lemons, oranges, and grapefruit contain citric acid, and sour milk contains lactic acid.
- Bases are present in many household cleaning agents, from bar soap to ammonia-based window cleaners, to drain openers.
- It is a good idea to learn the names and formulas of common acids and bases.

10.2 Some Common Acids and Bases

- **Sulfuric acid, H_2SO_4,** is manufactured in greater quantity worldwide than any other industrial chemical. It is the acid found in automobile batteries, is highly corrosive, and can cause painful burns.
- **Hydrochloric acid, HCl,** or *muriatic acid* has many industrial applications, including metal cleaning and the manufacture of high-fructose corn syrup. Aqueous HCl is stomach acid in the digestive systems of most mammals.

10.2 Some Common Acids and Bases

- **Phosphoric acid, H_3PO_4,** is used in phosphate fertilizers, and as an additive in foods and toothpastes. The tart taste of many soft drinks is due to the presence of phosphoric acid.
- **Nitric acid, HNO_3,** is a strong oxidizing agent that is used for the manufacture of ammonium nitrate fertilizer and military explosives. When spilled on the skin, it leaves a characteristic yellow coloration because of its reaction with skin proteins.
- **Acetic acid, CH_3CO_2H,** is the primary constituent of vinegar. It occurs in all living cells and is used in the preparation of solvents, lacquers, and coatings.

10.2 Some Common Acids and Bases

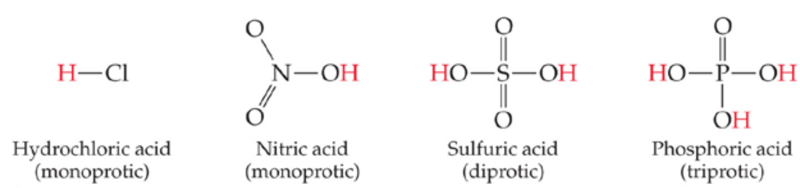
- **Sodium hydroxide, NaOH,** *caustic soda* or *lye*, is the most commonly used of all bases. It is used in the production of aluminum, glass, soap, and drain cleaners. Concentrated solutions can cause severe burns.
- **Calcium hydroxide, $Ca(OH)_2$,** or *slaked lime*, is made by treating lime (CaO) with water. It is used in mortars and cements. Aqueous $Ca(OH)_2$ is *limewater*.
- **Magnesium hydroxide, $Mg(OH)_2$,** or *milk of magnesia*, is an additive in foods, toothpaste, and over-the-counter medications. Antacids contain magnesium hydroxide.
- **Ammonia, NH_3,** is used primarily as a fertilizer, but has industrial applications including pharmaceuticals and explosives. A dilute solution of ammonia is frequently used as a glass cleaner.

10.3 The Brønsted–Lowry Definition of Acids and Bases

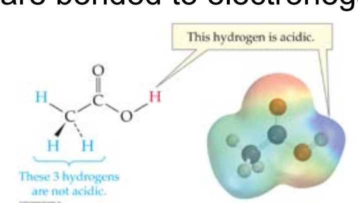
A more general definition was proposed in 1923 by the Danish chemist Johannes Brønsted and the English chemist Thomas Lowry.

- A **Brønsted–Lowry acid** is any substance that is able to give a hydrogen ion, H^+, to another molecule or ion.
- Acids are *proton donors*.
- The reaction need not occur in water, and a Brønsted–Lowry acid need not give H^+ ions in water.

10.3 The Brønsted–Lowry Definition of Acids and Bases

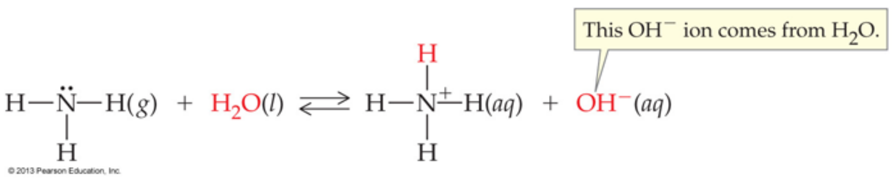


- Different acids can supply different numbers of H^+ ions.
- Acids with one proton (HCl) are *monoprotic acids*.
- H_2SO_4 is *diprotic* because it has two protons to donate.
- H_3PO_4 is *triprotic* because it has three protons to donate.
- Acidic H atoms are bonded to electronegative atoms.



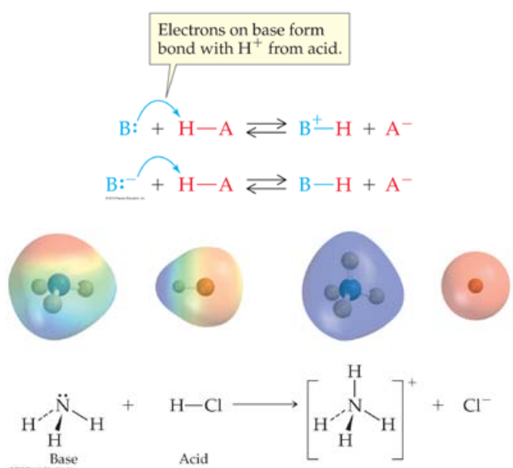
10.3 The Brønsted–Lowry Definition of Acids and Bases

- A **Brønsted–Lowry base** is a substance that *accepts* H^+ from an acid.
- Ammonia will react with water to produce OH^- ions.



10.3 The Brønsted–Lowry Definition of Acids and Bases

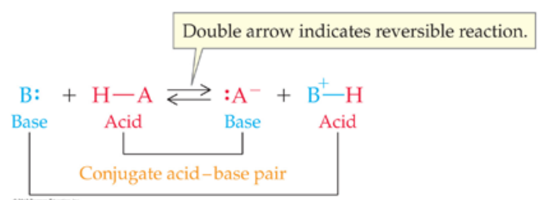
- An acid–base reaction is one in which a proton is transferred.



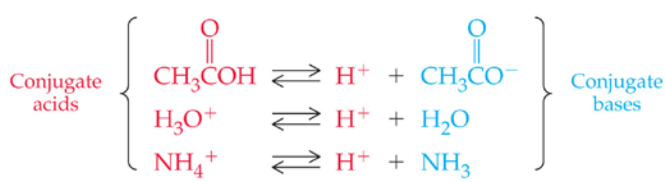
10.3 The Brønsted–Lowry Definition of Acids and Bases

- A Brønsted–Lowry base *must* have a lone pair of electrons to accept H^+ from an acid.
- A base can be either neutral (B:) or negatively charged ($B:^-$).
- If the base is neutral, the product has a positive charge after H^+ is added.
- The *products* of an acid–base reaction can also behave as acids and bases.
- Many acid–base reactions are reversible.

10.3 The Brønsted–Lowry Definition of Acids and Bases



- Pairs of chemical species such as B, BH^+ and HA, A^- are **conjugate acid–base pairs**.
- They are found on opposite sides of a chemical reaction, with formulas that differ by one H^+.



10.4 Acid and Base Strength

- **Strong acids** give up a proton easily and are essentially 100% **dissociated** in water.
- **Weak acids** give up a proton with difficulty and are substantially less than 100% dissociated in water.
- **Weak bases** have little affinity for a proton.
- **Strong bases** grab and hold a proton tightly.

10.4 Acid and Base Strength

TABLE 10.1 Relative Strengths of Acids and Conjugate Bases

Increasing acid strength (upward) — Increasing base strength (downward)

	ACID		CONJUGATE BASE		
Strong acids: 100% dissociated	Perchloric acid	$HClO_4$	ClO_4^-	Perchlorate ion	Little or no reaction as bases
	Sulfuric acid	H_2SO_4	HSO_4^-	Hydrogen sulfate ion	
	Hydriodic acid	HI	I^-	Iodide ion	
	Hydrobromic acid	HBr	Br^-	Bromide ion	
	Hydrochloric acid	HCl	Cl^-	Chloride ion	
	Nitric acid	HNO_3	NO_3^-	Nitrate ion	
	Hydronium ion	H_3O^+	H_2O	**Water**	
Weak acids	Hydrogen sulfate ion	HSO_4^-	SO_4^{2-}	Sulfate ion	Very weak bases
	Phosphoric acid	H_3PO_4	$H_2PO_4^-$	Dihydrogen phosphate ion	
	Nitrous acid	HNO_2	NO_2^-	Nitrite ion	
	Hydrofluoric acid	HF	F^-	Fluoride ion	
	Acetic acid	CH_3COOH	CH_3COO^-	Acetate ion	
Very weak acids	Carbonic acid	H_2CO_3	HCO_3^-	Bicarbonate ion	Weak bases
	Dihydrogen phosphate ion	$H_2PO_4^-$	HPO_4^{2-}	Hydrogen phosphate ion	
	Ammonium ion	NH_4^+	NH_3	Ammonia	
	Hydrocyanic acid	HCN	CN^-	Cyanide ion	
	Bicarbonate ion	HCO_3^-	CO_3^{2-}	Carbonate ion	
	Hydrogen phosphate ion	HPO_4^{2-}	PO_4^{3-}	Phosphate ion	
	Water	H_2O	OH^-	**Hydroxide ion**	Strong base

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10.4 Acid and Base Strength

- Polyprotic acids undergo stepwise dissociations in water.
- The first dissociation occurs to the extent of nearly 100%.
- The second dissociation takes place to a much lesser extent because separation of a positively charged H+ from a negatively charged anion is difficult.

10.4 Acid and Base Strength

- There is an inverse relationship between acid strength and base strength.
- The stronger the acid, the weaker its conjugate base; the weaker the acid, the stronger its conjugate base.
- Knowing the relative strengths of acids makes it possible to predict proton-transfer reactions.
- An acid–base proton-transfer equilibrium always favors reaction of the stronger acid with the stronger base, and formation of the weaker acid and base.

10.4 Acid and Base Strength

GERD—Too Much Acid or Not Enough?

- The major component of gastric juices is hydrochloric acid.
- Stomach acid is essential for the digestion of proteins and absorption of certain micronutrients. It also creates a sterile environment by killing yeast and bacteria that may be ingested.
- If these gastric juices leak up into the esophagus, they can cause heartburn or acid indigestion. Persistent irritation of the esophagus is known as gastro-esophageal reflux disease (GERD).
- Those who suffer from acid indigestion can obtain relief using over-the-counter antacids.
- Proton-pump inhibitors (PPI), such as Prevacid and Prilosec, prevent the production the H^+ ions in the parietal cells, while H2-receptor blockers (Tagamet, Zantac, and Pepcid) prevent the release of stomach acid into the lumen.
- The valve that controls the release of stomach contents to the small intestine is triggered by acidity. If the stomach is not acidic enough, the contents of the stomach can be churned up into the esophagus.

10.5 Acid Dissociation Constants

- The reaction of a weak acid with water, like any chemical equilibrium, can be described by an equilibrium equation.

$$K = \frac{[H_3O^+][A^-]}{[HA]}$$

- The K_a values for strong acids are so large that it is difficult and not very useful to measure them.
- For a weak acid, K_a is small.

10.5 Acid Dissociation Constants

TABLE 10.2 Some Acid Dissociation Constants, K_a, at 25 °C

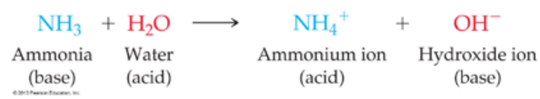
Acid	K_a	Acid	K_a
Hydrofluoric acid (HF)	3.5×10^{-4}	*Polyprotic acids*	
Hydrocyanic acid (HCN)	4.9×10^{-10}	Sulfuric acid	
Ammonium ion (NH_4^+)	5.6×10^{-10}	H_2SO_4	Large
		HSO_4^-	1.2×10^{-2}
Organic acids		Phosphoric acid	
Formic acid (HCOOH)	1.8×10^{-4}	H_3PO_4	7.5×10^{-3}
Acetic acid (CH_3COOH)	1.8×10^{-5}	$H_2PO_4^-$	6.2×10^{-8}
Propanoic acid (CH_3CH_2COOH)	1.3×10^{-5}	HPO_4^{2-}	2.2×10^{-13}
		Carbonic acid	
Ascorbic acid (vitamin C)	7.9×10^{-5}	H_2CO_3	4.3×10^{-7}
		HCO_3^-	5.6×10^{-11}

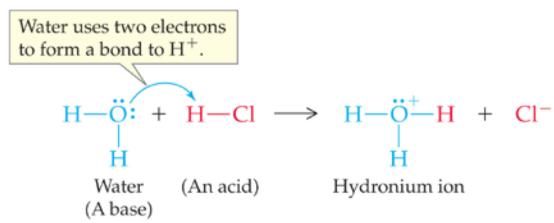
10.5 Acid Dissociation Constants

- Strong acids have K_a values much greater than 1 because dissociation is favored.
- Weak acids have K_a values much less than 1 because dissociation is not favored.
- Donation of each successive H^+ from a polyprotic acid is more difficult than the one before it, so K_a values become successively lower.
- Most organic acids, which contain the —COOH group, have K_a values near 10^{-5}.

10.6 Water as Both an Acid and a Base

- Water can act as *both* an acid and a base.
- When in contact with a base, water *donates* a proton.
- When in contact with an acid, water *accepts* a proton.
- Substances which can react as either an acid or a base are **amphoteric**.





10.6 Water as Both an Acid and a Base

- Water can dissociate to form hydronium and hydroxide ions.
- The **ion-product constant for water (*K_w*)**, is the H^+ concentration times the OH^- concentration.
- At 25 °C, $K_w = 1.00 \times 10^{-14}$.
- This equation applies to all aqueous solutions, not just pure water.
- As a result, we can determine the concentration of one species, if we know the concentration of the other.

10.6 Water as Both an Acid and a Base

- Solutions are identified as acidic, neutral, or basic (*alkaline*) according to the value of their H_3O^+ and OH^- concentrations.

Acidic solution: $[H_3O^+] > 10^{-7}$ and $[OH^-] < 10^{-7}$

Basic solution: $[H_3O^+] < 10^{-7}$ and $[OH^-] > 10^{-7}$

Neutral solution: $[H_3O^+] = 10^{-7}$ and $[OH^-] = 10^{-7}$

10.7 Measuring Acidity in Aqueous Solution: pH

- The pH of an aqueous solution is a number, usually between 0 and 14, that indicates the H_3O^+ concentration of the solution.

$$pH = -\log[H_3O^+]$$

- A pH smaller than 7 corresponds to an acidic solution.
- A pH larger than 7 corresponds to a basic solution.
- A pH of exactly 7 corresponds to a neutral solution.

10.7 Measuring Acidity in Aqueous Solution: pH

Acidic solution: pH < 7, $[H_3O^+] > 10^{-7}$

Basic solution: pH > 7, $[H_3O^+] < 10^{-7}$

Neutral solution: pH = 7, $[H_3O^+] = 10^{-7}$

- The pH scale covers an enormous range of acidities because it is *logarithmic*.
- A change of 1 pH unit means a tenfold change in H_3O^+.
- Calculations can sometimes be simplified by realizing that pH + pOH = 14.

10.8 Working with pH

- Converting from pH to $[H_3O^+]$ requires finding the *antilogarithm* of the negative pH.
- This is done on many calculators with an "INV" key and a "log" key.
- Consult your calculator instructions if you are not sure how to use these keys.
- Remember that the sign of the number given by the calculator must be changed to get the pH.

10.9 Laboratory Determination of pH

- The simplest but least accurate method of pH determination is **acid–base indicators**. These change color depending on the pH of the solution.
- Test kits contain a mixture of indicators known as *universal indicator* that give approximate pH measurements in the range 2–10.
- pH paper makes it possible to determine pH simply by putting a drop of solution on the paper and comparing the color that appears to the color on a calibration chart.
- A much more accurate way to determine pH is an electronic pH meter. Electrodes are dipped into the solution, and the pH is read from the meter.

10.10 Buffer Solutions

- Maintaining the pH of blood and other fluids within narrow limits is accomplished through the use of **buffers**.
- Most buffers are mixtures of a weak acid and its conjugate base.
- If OH^- is added to a buffer solution, the pH increases only slightly; the acid component of the buffer neutralizes the added OH^-.
- If H^+ is added to a buffer solution, the pH decreases only slightly; the base component of the buffer neutralizes the added H^+.

10.10 Buffer Solutions

- Rearranging the K_a equation to its logarithmic form gives the **Henderson–Hasselbalch equation:**

$$pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right)$$

- Where HA is the weak acid and A^- is the conjugate base.

10.10 Buffer Solutions

- The effective pH range of a buffer will depend on the pK_a of the acid and the ratio of HA and conjugate base.
- The pK_a for the weak acid should be close to the desired pH of the buffer solution.
- The ratio of [HA] to A^- should be close to 1.
- The molar amounts of HA and A^- in the buffer should be approximately 10 times greater than the molar amounts of acid or base that may be added.

10.10 Buffer Solutions

- The pH of body fluids is maintained by three major buffer systems.
- The carbonic acid–bicarbonate system and the dihydrogen phosphate–hydrogen phosphate system depend on weak acid–conjugate base interactions.
- The third buffer system depends on the ability of proteins to act as either proton acceptors or proton donors at different pH values.

10.10 Buffer Solutions

Buffers in the Body: Acidosis and Alkalosis

- Each of the fluids in our bodies has a pH range suited to its function.
- Blood plasma and the interstitial fluid surrounding cells have a slightly basic pH with a normal range of 7.35–7.45.
- The reactions and equilibria that take place throughout the body are very sensitive to pH—variations of even a few tenths of a pH unit can produce severe symptoms.
- Maintaining the pH of blood serum is accomplished by the carbonic acid-bicarbonate buffer system.
- The bicarbonate buffer system is intimately related to the elimination of CO_2.
- *Respiratory acidosis* can be caused by a decrease in respiration, which leads to a buildup of excess CO_2 in the blood and a corresponding decrease in pH.
- *Metabolic acidosis* results from an excess of other acids in the blood that reduce the bicarbonate concentration.
- Increased breathing can remove too much CO_2 from the blood, causing *respiratory alkalosis*.

10.11 Acid and Base Equivalents

- For acids and bases, the property of interest is the number of H^+ ions (for an acid) or the number of OH^- ions (for a base) per formula unit.
- *One equivalent of any acid neutralizes one equivalent of any base*.
- The **normality (N)** of an acid or base solution is the number of equivalents of acid or base per liter of solution.

10.11 Acid and Base Equivalents

- The values of molarity (M) and normality (N) are the same for monoprotic acids, but are not the same for diprotic or triprotic acids.

$$1\ N\ HCl = 1\ M\ HCl$$

$$1\ N\ H_2SO_4 = 0.5\ M\ H_2SO_4$$

- For any acid or base, normality is always equal to molarity times the number of H^+ or OH^- ions produced per formula unit.

10.12 Some Common Acid-Base Reactions

Reaction of Acids with Hydroxide Ion

- One equivalent of an acid reacts with 1 Eq of a metal hydroxide to yield water and a salt in a neutralization reaction.
- Such reactions are written with a single arrow; their equilibria lie far to the right.
- The net ionic equation for all such reactions makes clear why acid-base equivalents are useful.
- The equivalent ions for the acid H^+ and the base OH^- are used up in the formation of water.

10.12 Some Common Acid-Base Reactions

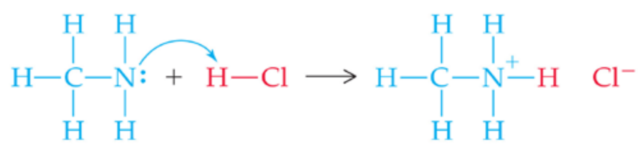
Reaction of Acids with Bicarbonate and Carbonate Ion

- Bicarbonate ion reacts with acid by accepting H^+ to yield carbonic acid, H_2CO_3. Carbonic acid decomposes to carbon dioxide gas and water.
- Most metal carbonates are insoluble in water but react easily with aqueous acid.
- Geologists test for carbonate-bearing rocks by putting a few drops of aqueous HCl on the rock and watching to see if bubbles of CO_2 form.
- Antacids that contain carbonates neutralize excess stomach acid.

10.12 Some Common Acid-Base Reactions

Reaction of Acids with Ammonia

- Acids react with ammonia to yield water-soluble ammonium salts.
- Living organisms contain *amines*, which contain nitrogen atoms bonded to carbon.
- Amines react with acids just as ammonia does, yielding water-soluble salts.



Methylamine — Methylammonium chloride

10.13 Titration

- pH gives H^+ concentration but not necessarily total acid concentration.
- The H^+ concentration gives the amount of acid that has dissociated into ions.
- Total acid concentration gives the sum of dissociated plus undissociated acid.
- Strong acids are 100% dissociated, weak acids typically dissociate 1% or less.

10.13 Titration

- The total acid or base concentration of a solution is found by carrying out a **titration.**

  Titration of an acid solution of unknown concentration with a base solution of known concentration:
- A measured volume of acid solution is placed in the flask along with an indicator.
- The base of known concentration is added from a buret until the color change of the indicator shows that neutralization is complete (the *end point*).

10.13 Titration

1. Reading from the buret gives the volume of the NaOH solution that has reacted with the known volume of HCl.
2. Knowing the concentration and volume of the NaOH solution allows calculation of the molar amount of NaOH.
3. The coefficients in the balanced equation allow us to find the molar amount of HCl that has been neutralized.
4. Dividing the molar amount of HCl by the volume of the HCl solution gives the concentration.

10.13 Titration

- When the titration involves a neutralization reaction in which one mole of acid reacts with one mole of base, the moles of acid and base needed for complete reaction can be represented as:

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

10.13 Titration

Acid Rain

- Under normal conditions, rain is slightly acidic, with a pH close to 5.6, because of atmospheric CO_2 that dissolves to form carbonic acid.
- In recent decades, however, the acidity of rainwater in many industrialized areas of the world has increased by a factor of over 100, to a pH between 3 and 3.5.
- The primary cause of this so-called *acid rain* is industrial and automotive pollution.
- Each year, large power plants and smelters pour millions of tons of sulfur dioxide gas into the atmosphere. Sulfur oxides then dissolve in rain to form dilute sulfurous and sulfuric acids.
- Nitrogen oxides produced by coal-burning plants and automobile engines further contribute to the problem. Nitrogen dioxide (NO_2) dissolves in water to form dilute nitric acid.
- Many processes in nature require such a fine pH balance that they are dramatically upset by the shift that has occurred in the pH of rain.
- Industrial emissions of sulfur and nitrogen oxides decreased by over 40% from 1990 to 2007, resulting in a decrease in acid rain depositions.

10.14 Acidity and Basicity of Salt Solutions

- Salt solutions can be neutral, acidic, or basic, depending on the ions present.
- Some ions react with water to produce H^+ and some ions react with water to produce OH^-.

TABLE 10.3 Acidity and Basicity of Salt Solutions

Anion Derived from Acid That Is:	Cation Derived from Base That Is:	Solution	Example
Strong	Weak	Acidic	NH_4Cl, NH_4NO_3
Weak	Strong	Basic	$NaHCO_3$, KCH_3CO_2
Strong	Strong	Neutral	NaCl, KBr, $Ca(NO_3)_2$
Weak	Weak	More information needed	

10.14 Acidity and Basicity of Salt Solutions

Salt of Strong Acid + Weak Base = Acidic

- A salt such as NH_4Cl which can be formed by reaction of a strong acid (HCl) with a weak base NH_3 yields an acidic solution.
- The Cl^- ion does not react with water, but the NH_4^+ ion is a weak acid that gives H^+ ions.

$$NH_4^+\ (aq) + H_2O\ (l) \leftrightarrows NH_3\ (aq) + H_3O^+\ (aq)$$

10.14 Acidity and Basicity of Salt Solutions

Salt of Weak Acid + Strong Base = Basic

- A salt such as sodium bicarbonate, which can be formed by reaction of a weak acid, H_2CO_3, with a strong base (NaOH), yields a basic solution.
- The Na^+ ion does not react with water, but the HCO_3^- ion is a weak base that gives OH ions.

$$HCO_3^- (aq) + H_2O (l) \leftrightharpoons H_2CO_3 (aq) + OH^-(aq)$$

10.14 Acidity and Basicity of Salt Solutions

Salt of Strong Acid + Strong Base = Neutral

- A salt such as NaCl, which can be formed by reaction of a strong acid (HCl) with a strong base (NaOH), yields a neutral solution.

Salt of Weak Acid + Weak Base = ?

- Both cation and anion in this type of salt react with water. The ion that reacts to the greater extent will govern the pH.

Chapter Summary

1. What are acids and bases?

- According to the *Brønsted–Lowry definition*, an acid is a substance that donates a hydrogen ion (a proton, H^+) and a base is a substance that accepts a hydrogen ion.
- Thus, the generalized reaction of an acid with a base involves the reversible transfer of a proton:

$$B: + H—A \leftrightarrows A:^- + H—B^+$$

- In aqueous solution, water acts as a base and accepts a proton from an acid to yield a *hydronium ion*, H_3O^+.
- Reaction of an acid with a metal hydroxide, such as KOH, yields water and a salt; reaction with bicarbonate ion HCO_3^- or carbonate ion CO_3^{-2} yields water, a salt, and CO_2 gas; and reaction with ammonia yields an ammonium salt.

Chapter Summary, *Continued*

2. What effect does the strength of acids and bases have on their reactions?

- Different acids and bases differ in their ability to give up or accept a proton.
  - A *strong acid* gives up a proton easily and is 100% *dissociated* in aqueous solution;
  - A *weak acid* gives up a proton with difficulty, is only slightly dissociated in water, and establishes an equilibrium between dissociated and undissociated forms.
  - Similarly, a *strong base* accepts and holds a proton readily, whereas a *weak base* has a low affinity for a proton and establishes an equilibrium in aqueous solution.
- The two substances that are related by the gain or loss of a proton are called a *conjugate acid–base pair*. The exact strength of an acid is defined by an *acid dissociation constant*, K_a.
- A proton-transfer reaction always takes place in the direction that favors formation of the weaker acid.

Chapter Summary, *Continued*

3. What is the ion-product constant for water?

- Water is *amphoteric*; that is, it can act as either an acid or a base. Water also dissociates slightly into H_3O^+ ions and OH^- ions; the product of these concentrations in any aqueous solution is the *ion-product constant for water,* $K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$ at 25 °C.

4. What is the pH scale for measuring acidity?

- The acidity or basicity of an aqueous solution is given by its *pH*, defined as the negative logarithm of the hydronium ion concentration, $[H_3O^+]$. A pH below 7 means an acidic solution; a pH equal to 7 means a neutral solution; and a pH above 7 means a basic solution.

Chapter Summary, *Continued*

5. What is a buffer?

- The pH of a solution can be controlled through the use of a *buffer* that acts to remove either added H_3O^+ ions or added OH^- ions. Most buffer solutions consist of roughly equal amounts of a weak acid and its conjugate base. The bicarbonate buffer present in blood and the hydrogen phosphate buffer present in cells are particularly important examples.

6. How is the acid or base concentration of a solution determined?

- Acid (or base) concentrations are determined in the laboratory by *titration* of a solution of unknown concentration with a base (or acid) solution of known strength until an indicator signals that neutralization is complete.