Chapter 19: Acids, Bases, and Salts

19.1 Acid-Base Theories

Properties of Acids and Bases

Acids

Acids taste sour, will change the color of an acid-base indicator, and can be strong or weak electrolytes in aqueous solution.

Properties of Acids and Bases

Bases

Bases taste bitter, feel slippery, will change the color of an acid-base indicator, and can be strong or weak electrolytes in aqueous solution.

Arrhenius Acids and Bases

Arrhenius said that acids are hydrogen-containing compounds that ionize to yield hydrogen ions (H^+) in aqueous solution. He also said that bases are compounds that ionize to yield hydroxide ions (OH^-) in aqueous solution.

Arrhenius Acids and Bases

Arrhenius Acids

- Acids that contain one ionizable hydrogen, (nitric - HNO_3), are called **monoprotic acids**.
- Acids that contain two, (sulfuric - H_2SO_4), are called **diprotic acids**.
- Acids that contain three, (phosphoric - H_3PO_4) are called **triprotic acids**.

Arrhenius Bases

Hydroxide ions are one of the products of the dissolution of an alkali metal in water.

$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$

Sodium metal — Water — Sodium hydroxide — Hydrogen

$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

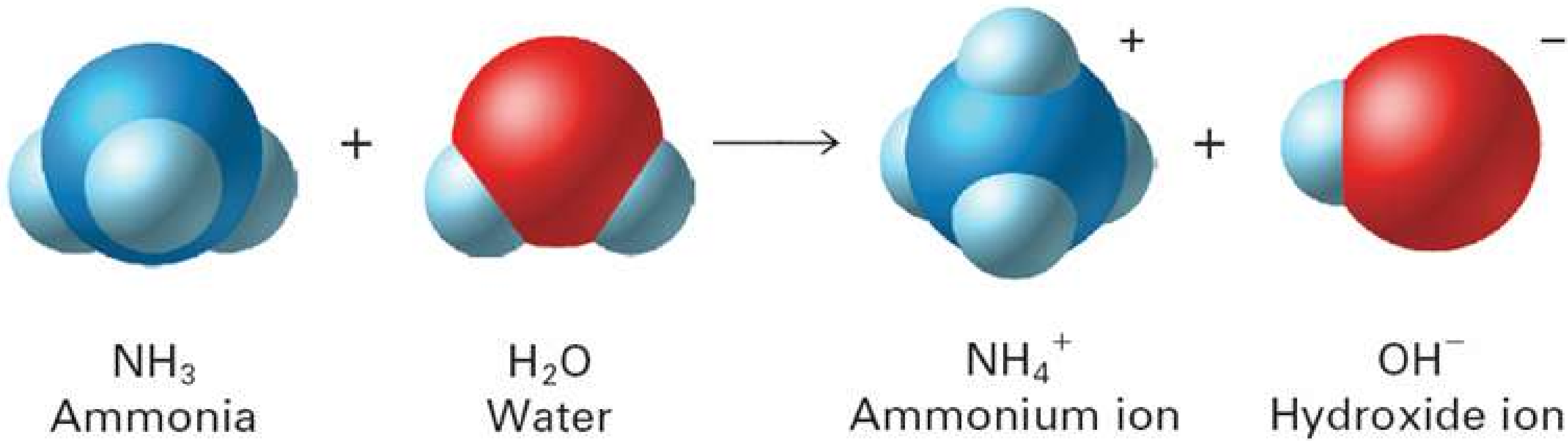
Sodium hydroxide — Sodium ion — Hydroxide ion

Brønsted-Lowry Acids and Bases

The Brønsted-Lowry theory defines an acid as a hydrogen-ion donor, and a base as a hydrogen-ion acceptor.

Brønsted-Lowry Acids and Bases

Why Ammonia is a Base



$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Ammonia (hydrogen-ion acceptor, Brønsted-Lowry base)	Water (hydrogen-ion donor, Brønsted-Lowry acid)	Ammonium ion	Hydroxide ion (makes the solution basic)

Conjugate Acids and Bases

A **conjugate acid** is the particle formed when a base gains a hydrogen ion.

A **conjugate base** is the particle that remains when an acid has donated a hydrogen ion.

$$\underset{\text{Base}}{NH_3(aq)} + \underset{\text{Acid}}{H_2O(l)} \rightleftharpoons \underset{\text{Conjugate acid}}{NH_4^+(aq)} + \underset{\text{Conjugate base}}{OH^-(aq)}$$

$$\underset{\text{Acid}}{HCl(g)} + \underset{\text{Base}}{H_2O(l)} \rightleftharpoons \underset{\text{Conjugate acid}}{H_3O^+(aq)} + \underset{\text{Conjugate base}}{Cl^-(aq)}$$

A **conjugate acid-base pair** consists of two substances related by the loss or gain of a single hydrogen ion.

$$\underset{\text{Base}}{NH_3(aq)} + \underset{\text{Acid}}{H_2O(l)} \rightleftharpoons \underset{\text{Conjugate acid}}{NH_4^+(aq)} + \underset{\text{Conjugate base}}{OH^-(aq)}$$

A substance that can act as both an acid and a base is said to be **amphoteric.**

Which are conj acid-base pairs?

1) HF, F^-
2) NH_4^+, NH_3
3) HCl, H_2O
4) $HClO_4$, ClO_4^-
5) HCl, ClO^-
6) HNO_3, NO_3^-

Write Conj Base for:

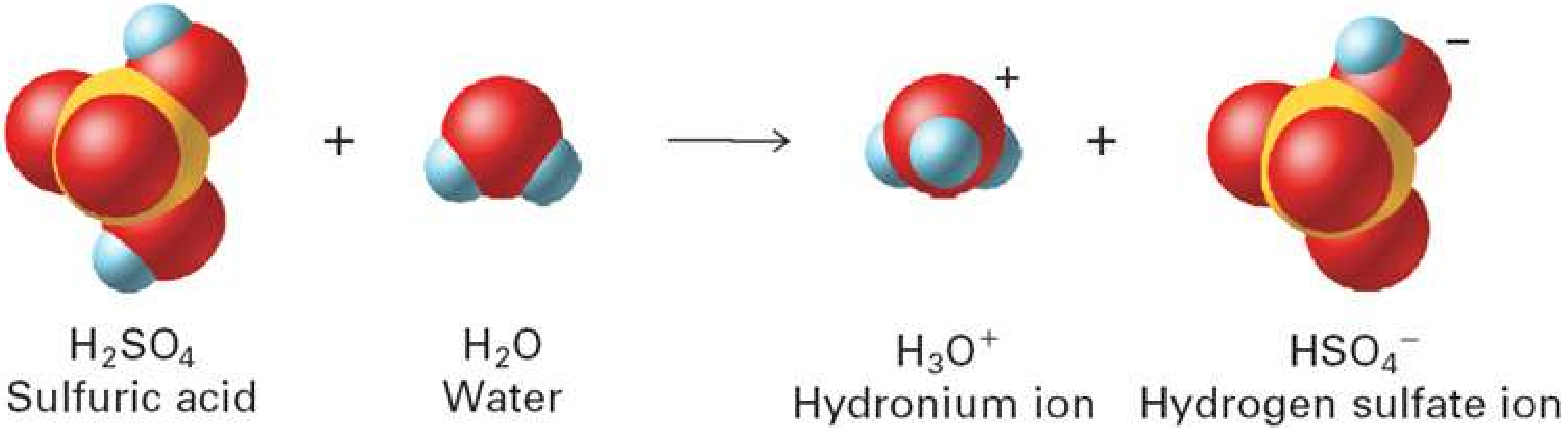
1) $HClO_4$
2) H_3PO_4
3) $CH_3NH_3^+$
4) H_2S
5) HS^-
6) H_2SO_3

Write Conj Acid of:

1) $H_2PO_4^-$
2) Cl^-
3) ClO_3^-
4) HS^-
5) HSO_4^-
6) NO_2^-

Brønsted-Lowry Acids and Bases

A water molecule that gains a hydrogen ion becomes a positively charged hydronium ion (H_3O^+).



Lewis Acids and Bases

Lewis proposed that an acid accepts a pair of electrons during a reaction, while a base donates a pair of electrons.

A **Lewis acid** is a substance that can accept a pair of electrons to form a covalent bond.

A **Lewis base** is a substance that can donate a pair of electrons to form a covalent bond.

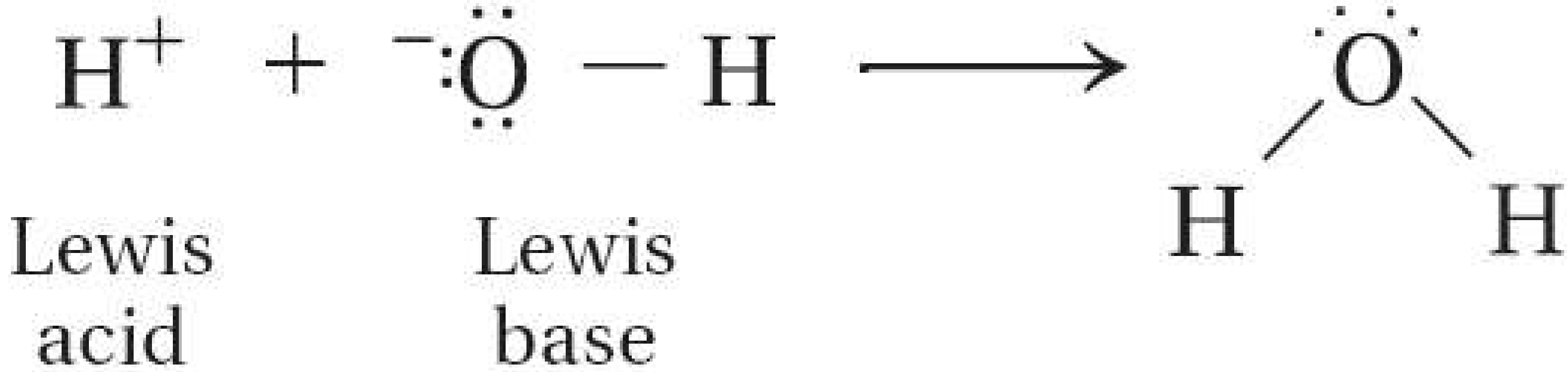


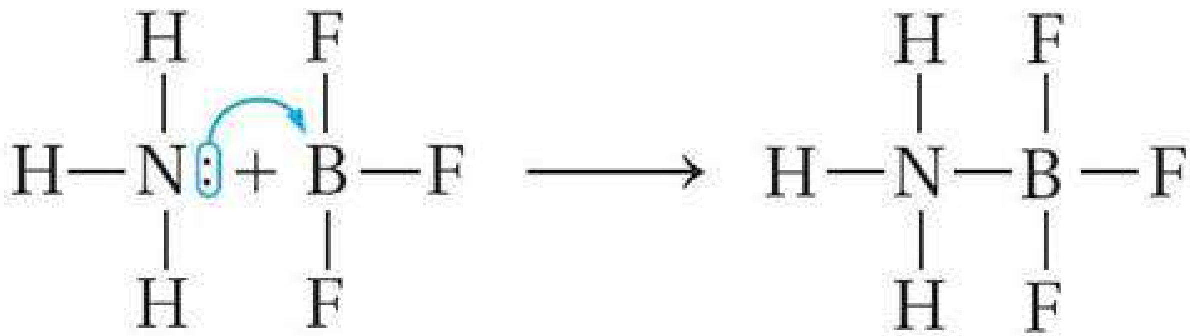
Table 19.4

Acid-Base Definitions		
Type	**Acid**	**Base**
Arrhenius	H^+ producer	OH^- producer
Brønsted-Lowry	H^+ donor	H^+ acceptor
Lewis	electron-pair acceptor	electron-pair donor

CONCEPTUAL PROBLEM 19.1

Identifying Lewis Acids and Bases

Ammonia is widely used in fertilizers, plastics, and explosives. Identify the Lewis acid and the Lewis base in this reaction involving ammonia.



for Conceptual Problem 19.1

1. Identify the Lewis acid and Lewis base in each reaction.

a. $H^+ +$  $\longrightarrow H_3O^+$

b. $AlCl_3 + Cl^- \longrightarrow AlCl_4^-$

19.1 Section Quiz.

1. Which of the following is NOT a characteristic of acids?
   a) taste sour
   b) are electrolytes
   c) feel slippery
   d) affect the color of indicators

19.1 Section Quiz.

2. Which compound is most likely to act as an Arrhenius acid?

   a) H_2O

   b) NH_3.

   c) NaOH.

   d) H_2SO_4.

19.1 Section Quiz.

3. A Lewis acid is any substance that can accept

   a) a hydronium ion.

   b) a proton.

   c) hydrogen.

   d) a pair of electrons.

19.2 Hydrogen Ions and Acidity

Hydrogen Ions from Water

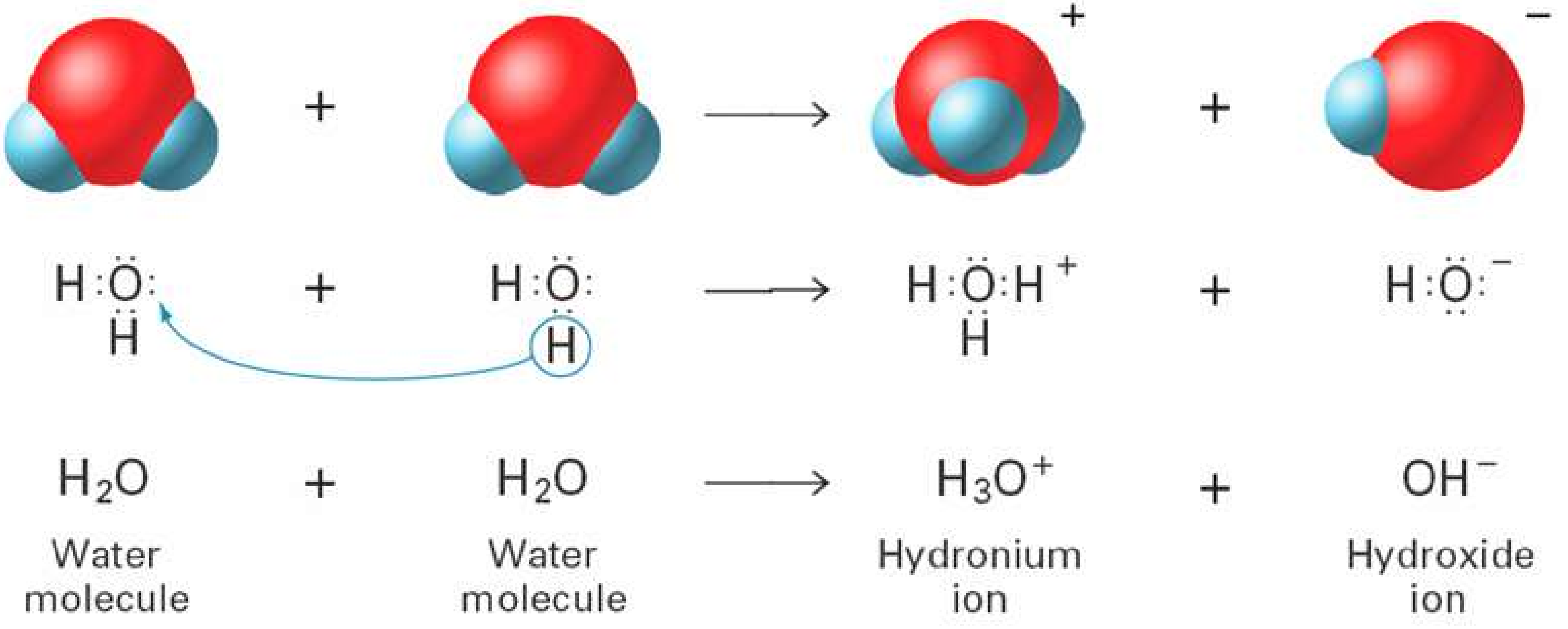
Hydrogen Ions from Water

The reaction in which water molecules produce ions is called the **self-ionization** of water.

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

Hydrogen ion — Hydroxide ion

In the self-ionization of water, a proton (hydrogen ion) transfers from one water molecule to another water molecule.



Ion Product Constant for Water

For aqueous solutions, the product of the hydrogen-ion concentration and the hydroxide-ion concentration equals 1.0×10^{-14}.

$$H^+(aq) + OH^-(aq) \rightleftharpoons H_2O(l)$$

Any aqueous solution in which $[H^+]$ and $[OH^-]$ are equal is described as a neutral solution.

Ion Product Constant for Water

The product of the concentrations of the hydrogen ions and hydroxide ions in water is called the **ion-product constant for water (K_w).**

$$K_w = [H^+] \times [OH^-] = 1.0 \times 10^{-14}$$

Ion Product Constant for Water

An **acidic solution** is one in which $[H^+]$ is greater than $[OH^-]$.

$$HCl(g) \xrightarrow{H_2O} H^+(aq) + Cl^-(aq)$$

Ion Product Constant for Water

A **basic solution** is one in which $[H^+]$ is less than $[OH^-]$. Basic solutions are also known as **alkaline solutions.**

$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

Sample Problem 19.1

Finding the $[OH^-]$ of a Solution

Colas are slightly acidic. If the $[H^+]$ in a solution is $1.0 \times 10^{-5}M$, is the solution acidic, basic, or neutral? What is the $[OH^-]$ of this solution?

for Sample Problem 19.1

10. If the hydroxide-ion concentration of an aqueous solution is $1 \times 10^{-3} M$, what is the $[H^+]$ in the solution? Is the solution acidic, basic, or neutral?

Calculate $[H^+]$ or $[OH^-]$ and tell if acid, base, or neutral

1) 1×10^{-5} M OH^-
2) 1×10^{-7} M OH^-
3) 10.0 M H^+
4) 3.4×10^{-4} M H^+
5) 2.6×10^{-8} M OH^-
6) 6.2×10^{-9} M H^+

The pH Concept

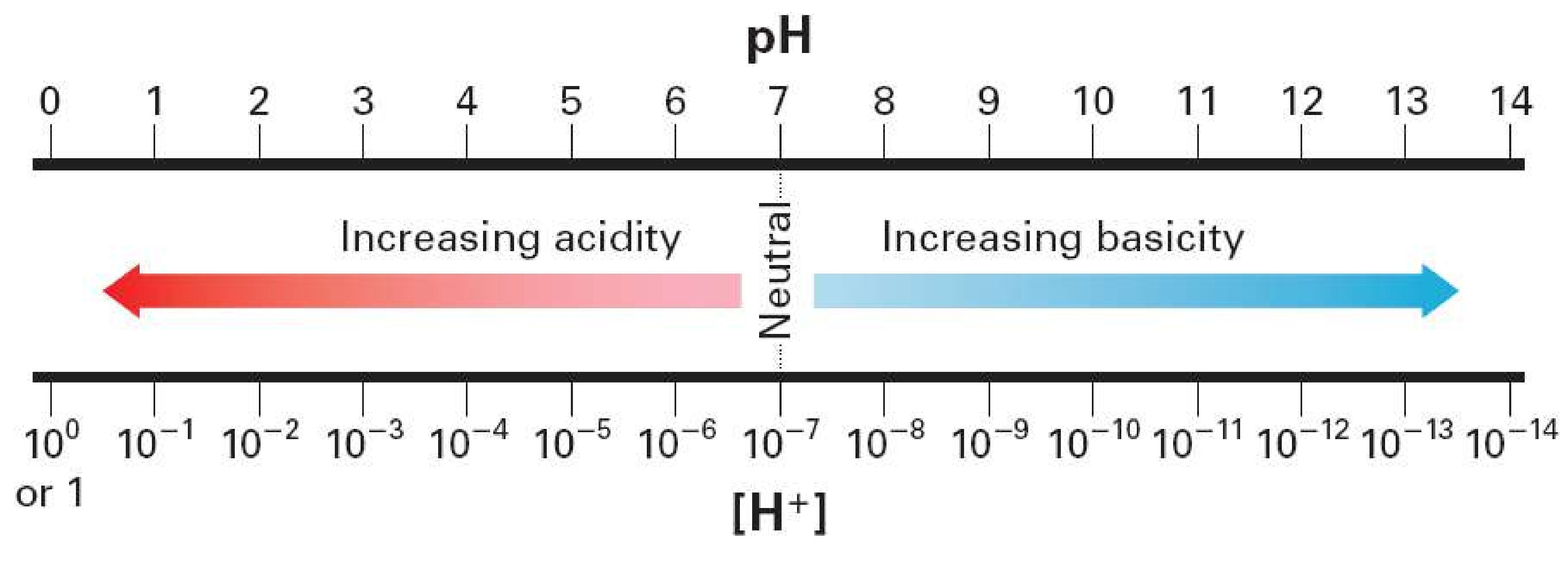
The **pH** of a solution is the negative logarithm of the hydrogen-ion concentration.

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

The pH Concept

A solution in which $[H^+]$ is greater than 1×10^{-7} *M* has a pH less than 7.0 and is acidic. The pH of pure water or a neutral aqueous solution is 7.0. A solution with a pH greater than 7 is basic and has a $[H^+]$ of less than 1×10^{-7} *M*.

The pH Concept



Sample Problem 19.2

Calculating pH from [H^+]

What is the pH of a solution with a hydrogen-ion concentration of $4.2 \times 10^{-10} M$?

for Sample Problem 19.2

12. What are the pH values of the following solutions, based on their hydrogen-ion concentrations?

a. $[H^+] = 1.0 \times 10^{-12} M$

b. $[H^+] = 0.045 M$

Calculating pOH

$$\mathrm{pOH} = -\log[\mathrm{OH}^-]$$

$$\mathrm{pH} + \mathrm{pOH} = 14$$

$$\mathrm{pH} = 14 - \mathrm{pOH}$$

$$\mathrm{pOH} = 14 - \mathrm{pH}$$

Sample Problem 19.4

Calculating pH from [OH^-]

What is the pH of a solution if [OH^-] $= 4.0 \times 10^{-11} M$?

for Sample Problem 19.4

15. Calculate the pH of each solution.

a. $[OH^-] = 4.3 \times 10^{-5} M$

b. $[OH^-] = 4.5 \times 10^{-11} M$

Calculate pH: acid, base or neutral?

1) $[H^+] = 1 \times 10^{-9}$ M
2) $[OH^-] = 1 \times 10^{-6}$ M
3) $[H^+] = 2.1 \times 10^{-5}$ M
4) $[OH^-] = 5.9 \times 10^{-8}$ M
5) $[H^+] = 1 \times 10^{-3}$ M
6) $[OH^-] = 5 \times 10^{-5}$ M

Sample Problem 19.3

Using pH to Find $[H^+]$

The pH of an unknown solution is 6.35. What is its hydrogen-ion concentration?

for Sample Problem 19.3

14. What are the hydrogen-ion concentrations for solutions with the following pH values?

a. 4.00

b. 11.55

Calculate $[H^+]$

1) pH = 7.41
2) pH = 3.50
3) pH = 3.14
4) pH = 8.53

Measuring pH

An indicator is a valuable tool for measuring pH because its acid form and base form have different colors in solution.

19.2 Section Quiz.

1. If the $[OH^-]$ in a solution is $7.65 \times 10^{-3} M$, what is the $[H^+]$ of this solution?
   a) $7.65 \times 10^{-17} M$
   b) $1.31 \times 10^{-12} M$
   c) $2.12 M$
   d) $11.88 M$

19.2 Section Quiz.

2. The $[OH^-]$ for four solutions is given below. Which one of the solution is basic?

a) $1.0 \times 10^{-6} M$

b) $1.0 \times 10^{-8} M$

c) $1.0 \times 10^{-7} M$

d) $1.0 \times 10^{-14} M$

19.2 Section Quiz.

3. What is the pH of a solution with a hydrogen-ion concentration of $8.5 \times 10^{-2} M$?

   a) 12.93

   b) 8.50

   c) 5.50

   d) 1.07

19.3 Strengths of Acids and Bases

Strong and Weak Acids and Bases

An **acid dissociation constant** (***K*****a**) is the ratio of the concentration of the dissociated (or ionized) form of an acid to the concentration of the undissociated (nonionized) form.

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

Strong and Weak Acids and Bases

Weak acids have small K_a values. The stronger an acid is, the larger is its K_a value.

Strong and Weak Acids and Bases

Strong acids are completely ionized in aqueous solution.

$$HCl(g) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq) \text{ (100\% ionized)}$$

Weak acids ionize only slightly in aqueous solution.

$$\underset{\text{Ethanoic acid}}{CH_3COOH(aq)} + \underset{\text{Water}}{H_2O(l)} \rightleftharpoons \underset{\text{Hydronium ion}}{H_3O^+(aq)} + \underset{\text{Ethanoate ion}}{CH_3COO^-(aq)}$$

Strong and Weak Acids and Bases

In general, the **base dissociation constant** (K_b) is the ratio of the concentration of the conjugate acid times the concentration of the hydroxide ion to the concentration of the base.

Strong and Weak Acids and Bases

Strong bases dissociate completely into metal ions and hydroxide ions in aqueous solution.

Weak bases react with water to form the hydroxide ion and the conjugate acid of the base.

$$\underset{\text{Ammonia}}{NH_3(aq)} + \underset{\text{Water}}{H_2O(l)} \rightleftharpoons \underset{\text{Ammonium ion}}{NH_4^+(aq)} + \underset{\text{Hydroxide ion}}{OH^-(aq)}$$

Dissociation Constants

To find the K_a of a weak acid or the K_b of a weak base, substitute the measured concentrations of all the substances present at equilibrium into the expression for K_a or K_b.

Calculating Dissociation Constants

Acid Dissociation Constant

The dissociation constant, K_a, of ethanoic acid is calculated from the equilibrium concentrations of all of the molecules and ions in the solution.

$$K_{eq} = \frac{[H_3O^+] \times [CH_3COO^-]}{[CH_3COOH] \times [H_2O]}$$

$$K_{eq} \times [H_2O] = K_a = \frac{[H_3O^+] \times [CH_3COO^-]}{[CH_3COOH]}$$

Calculating Dissociation Constants

Base Dissociation Constant

The dissociation constant, K_b, of ammonia is calculated from the equilibrium concentrations of all of the molecules and ions in the solution.

$$K_{eq} = \frac{[NH_4^+] \times [OH^-]}{[NH_3] \times [H_2O]}$$

$$K_{eq} \times [H_2O] = K_b = \frac{[NH_4^+] \times [OH^-]}{[NH_3]}$$

Sample Problem 19.5

Calculating a Dissociation Constant

A 0.1000M solution of ethanoic acid is only partially ionized. From measurements of the pH of the solution, $[H^+]$ is determined to be $1.34 \times 10^{-3} M$. What is the acid dissociation constant (K_a) of ethanoic acid?

for Sample Problem 19.5

23. In an exactly 0.2M solution of a monoprotic weak acid, $[H^+] = 9.86 \times 10^{-4} M$. What is the K_a for this acid?

Practice

Formic acid is used to process latex tapped from rubber trees into natural rubber. The pH of a 0.100 M solution of formic acid (HCOOH) is 2.38. What is the K_a for HCOOH?

Practice

- Calculate the K_a of a 0.220 M solution of H_3AsO_4 with a pH = 1.50.
- Calculate the K_a of a 0.00330 M solution of benzoic acid (C_6H_5COOH) with a pOH = 10.70.

19.3 Section Quiz.

1. H_2S is considered to be a weak acid because it

   is insoluble in water.

   ionizes only slightly.

   is completely ionized.

   is dilute.

19.3 Section Quiz.

2. Calcium hydroxide, $Ca(OH)_2$, is a strong base because it

- has a large K_b.
- has a small K_b.
- forms concentrated solutions.
- is highly soluble in water.

19.3 Section Quiz.

3. If the $[H^+]$ of a 0.205*M* solution of phenol (C_6H_5OH) at 25ºC is 2.340×10^{-6}, what is the K_a for phenol? Phenol is monoprotic.

$K_a = 2.67 \times 10^{-11}$

$K_a = 1.14 \times 10^{-5}$

$K_a = 5.48 \times 10^{-12}$

$K_a = 1.53 \times 10^{-3}$

19.3 Section Quiz.

4. The K_a of three acids is given below.
   (1) 5.1×10^{-3}
   (2) 4.8×10^{-11}
   (3) 6.3×10^{-5}
   Put the acids in order from the strongest acid to the weakest acid.
   - 1, 3, 2
   - 2, 3, 1
   - 3, 1, 2
   - 2, 1, 3

19.3 Section Quiz.

5. The K_b of four bases is given below.

(1) 7.41×10^{-5}

(2) 1.78×10^{-5}

(3) 4.27×10^{-4}

(4) 4.79×10^{-4}

Put the bases in order from the strongest base to the weakest base.

2, 3, 4, 1

19.4 Neutralization Reactions

Acid-Base Reactions

- In general, the reaction of an acid with a base produces water and one of a class of compounds called salts.

$$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

$$H_2SO_4(aq) + 2KOH(aq) \longrightarrow K_2SO_4(aq) + 2H_2O(l)$$

Acid-Base Reactions

Reactions in which an acid and a base react in an aqueous solution to produce a salt and water are generally called **neutralization reactions.**

Titration

The process of adding a known amount of solution of known concentration to determine the concentration of another solution is called **titration.**

The point of neutralization is the end point of the titration.

When an acid and base are mixed, the **equivalence point** is when the number of moles of hydrogen ions equals the number of moles of hydroxide ions.

$$\underset{1\text{ mol}}{HCl(aq)} + \underset{1\text{ mol}}{NaOH(aq)} \longrightarrow \underset{1\text{ mol}}{NaCl(aq)} + \underset{1\text{ mol}}{H_2O(l)}$$

$$\underset{2\text{ mol}}{2HCl(aq)} + \underset{1\text{ mol}}{Ca(OH)_2(aq)} \longrightarrow \underset{1\text{ mol}}{CaCl_2(aq)} + \underset{2\text{ mol}}{2H_2O(l)}$$

$$\underset{1\text{ mol}}{H_2SO_4(aq)} + \underset{2\text{ mol}}{2NaOH(aq)} \longrightarrow \underset{1\text{ mol}}{Na_2SO_4(aq)} + \underset{2\text{ mol}}{2H_2O(l)}$$

Sample Problem 19.6

Finding the Number of Moles of an Acid in Neutralization

How many moles of sulfuric acid are required to neutralize 0.50 mol of sodium hydroxide?

for Sample Problem 19.6

30. How many moles of potassium hydroxide are needed to completely neutralize 1.56 mol of phosphoric acid?

Titration

The solution of known concentration is called the **standard solution.**

Indicators are often used to determine when enough of the standard solution has been added to neutralize the acid or base.

The point at which the indicator changes color is the **end point** of the titration.

Sample Problem 19.7

Determining the Concentration of an Acid by Titration

A 25-mL solution of H_2SO_4 is completely neutralized by 18 mL of 1.0*M* NaOH. What is the concentration of the H_2SO_4 solution?

for Sample Problem 19.7

33. What is the molarity of H_3PO_4 if 15.0 mL is completely neutralized by 38.5 mL of 0.150 *M* NaOH?

Practice

How many mililiters of 0.500 M NaOH would neutralize 25.00 mL of 0.100 H_3PO_4?

19.4 Section Quiz

1. When a neutralization takes place, one of the products is always

   a) carbon dioxide.

   b) a salt.

   c) sodium chloride.

   d) a precipitate.

19.4 Section Quiz

2. In a titration, 45.0 mL of KOH is neutralized by 75.0 mL of 0.30*M* HBr. What is the concentration of the KOH solution?

   a) 0.18*M*

   b) 0.60*M*

   c) 0.25*M*

   d) 0.50*M*

19.4 Section Quiz

3. How many moles of HCl are required to neutralize an aqueous solution of 2.0 mol $Ca(OH)_2$?

    a) 0.5 mol

    b) 1.0 mol

    c) 2.0 mol

    d) 4.0 mol

19.5 Salts in Solution

Salt Hydrolysis

In general, salts that produce acidic solutions contain positive ions that release protons to water. Salts that produce basic solutions contain negative ions that attract protons from water.

Salt Hydrolysis

In **salt hydrolysis,** the cations or anions of a dissociated salt remove hydrogen ions from or donate hydrogen ions to water.

Salt Hydrolysis

To determine whether a salt solution is acidic or basic, remember the following rules:

Strong acid + Strong base $\longrightarrow$ Neutral solution

Strong acid + Weak base $\longrightarrow$ Acidic solution

Weak acid + Strong base $\longrightarrow$ Basic solution

Practice

Write equations for the salt hydrolysis reaction of the following. Classify each as acidic, basic, or neutral.

a) ammonium nitrate

b) potassium sulfate

c) rubidium acetate

d) calcium carbonate

Buffers

A buffer is a solution of a weak acid and one of its salts, or a solution of a weak base and one of its salts.

The pH of a **buffer** remains relatively constant when small amounts of acid or base are added.

The **buffer capacity** is the amount of acid or base that can be added to a buffer solution before a significant change in pH occurs.

CONCEPTUAL PROBLEM 19.2

Using Equations to Illustrate the Action of a Buffer

Show how the carbonic acid–hydrogen carbonate buffer can “mop up” added hydrogen ions and hydroxide ions.

for Conceptual Problem 19.2

39. Write an equation that shows what happens when acid is added to the ethanoic acid–ethanoate buffer.

19.5 Section Quiz.

1. Which of the following reactions would most likely yield a basic salt solution?

strong acid + weak base

weak acid + weak base

strong acid + strong base

weak acid + strong base

19.5 Section Quiz.

2. Choose the correct words for the spaces. A buffer can be a solution of a ____________ and its ____________.

   weak acid, salt

   strong acid, salt

   weak acid, conjugate base

   weak base, conjugate acid