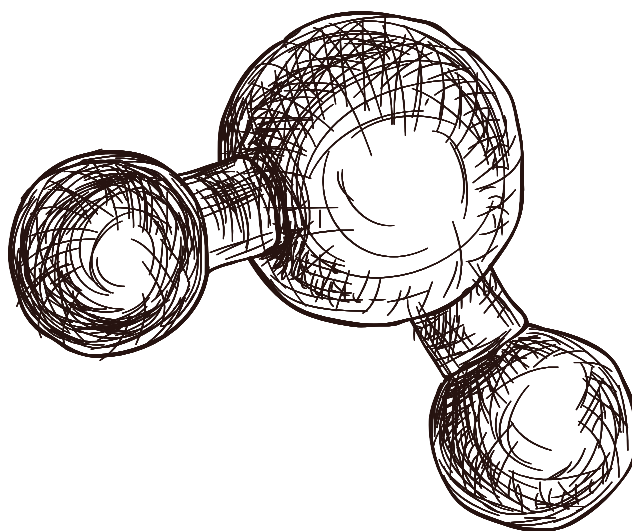


CHAPTER 3

ACIDS AND BASES



CHAPTER SUMMARY

by Abd Alrahman Basim

SECTION 1:
INTRODUCTION TO ACIDS AND BASES

SECTION 2:
STRENGTHS OF ACIDS AND BASES

SECTION 3:
HYDROGEN IONS AND PH

SECTION 4:
NEUTRALIZATION



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INTRODUCTION TO ACIDS AND BASES

Properties of Acids and Bases

Acids and bases are two of the most common classifications of substances. You can recognize them by the tart taste of some of your favourite beverages and by the pungent odor of ammonia in some household cleaners.

Physical Properties

Acidic solutions taste **sour** (حامض). Carbonic and phosphoric acids give many carbonated beverages their sharp taste; citric and ascorbic acids give lemons and grapefruit their tartness; and acetic acid makes vinegar taste sour.

Basic solutions taste **bitter** (مر) and feel **slippery**. A bar of soap becomes slippery when it gets wet.

Electrical Conductivity

Pure water is a non-conductor of electricity, but the addition of an **acid** or **base** produces ions that cause the resulting solution to become a **conductor**.

Chemical Properties

- **Reactions with Litmus**

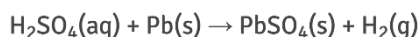
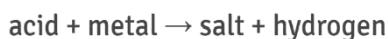
Litmus (عباد الشمس) is one of the dyes commonly used to distinguish solutions of acids and bases.

Aqueous solutions of **acids** cause blue litmus paper to **turn red**.

Aqueous solutions of **bases** cause red litmus paper to **turn blue**.

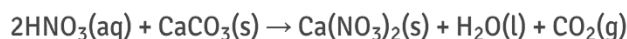
- **Reactions with Metals**

Unlike bases, acids react with most metals to yield **hydrogen gas** and a **salt**. Most metals replace the hydrogen atoms from an acid in a **single-replacement reaction**.



- **Reactions with Metal Carbonates**

Acids also react with **metal carbonates** and **hydrogen carbonates** to produce carbon dioxide (CO₂) gas, along with the corresponding metal salts and water.

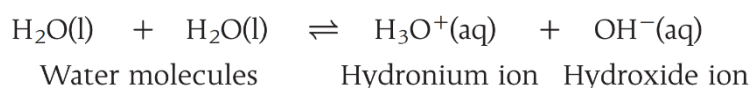


INTRODUCTION TO ACIDS AND BASES

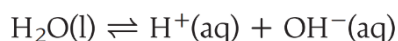
The concentrations of the ions in an aqueous solution determine the nature of the solution.

- **Acidic** solution: $[H^+] > [OH^-]$
- **Neutral** solution: $[H^+] = [OH^-]$
- **Basic** solution: $[H^+] < [OH^-]$

Pure water produces equal numbers of **H⁺** ions and **OH⁻** ions in a process called **self-ionization**, in which water molecules react to form a **hydronium ion (H₃O⁺)** and a **hydroxide ion**.

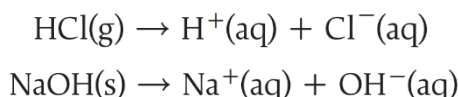


The hydronium ion is a hydrogen ion that has a water molecule attached to it by a covalent bond. The symbols **H⁺** and **H₃O⁺** can be used interchangeably, as this simplified self-ionization equation shows.



The Arrhenius Model

The **Arrhenius model** states that an **acid** is a substance that contains **hydrogen** and ionizes to produce **hydrogen ions** in aqueous solution. A **base** is a substance that contains a **hydroxide group** and dissociates to produce a **hydroxide ion** in aqueous solution.



Although the Arrhenius model is useful in explaining many acidic and basic solutions, it has some shortcomings. For example, **ammonia (NH₃)** and **sodium carbonate (Na₂CO₃)** do not contain a hydroxide group, yet both substances produce hydroxide ions in solution and are well-known bases.

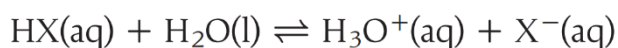
The Brønsted-Lowry Model

In the **Brønsted-Lowry model** of acids and bases,

An acid is a hydrogen-ion donor.

A base is a hydrogen-ion acceptor.

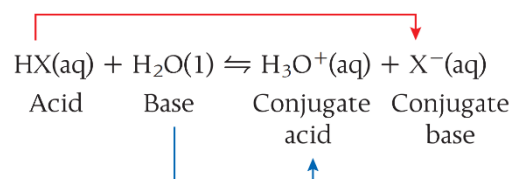
The symbols **X** and **Y** represent **non-metallic** elements or **negative polyatomic** ions. Thus, the general formula for an acid can be represented as **HX** or **HY**.



INTRODUCTION TO ACIDS AND BASES

Conjugate Acids and Bases

The forward reaction is the reaction of an acid and a base. The reverse reaction is also the reaction of an acid and a base. The acid and base that react in the reverse reaction are identified under the equation as a conjugate acid and a conjugate base.

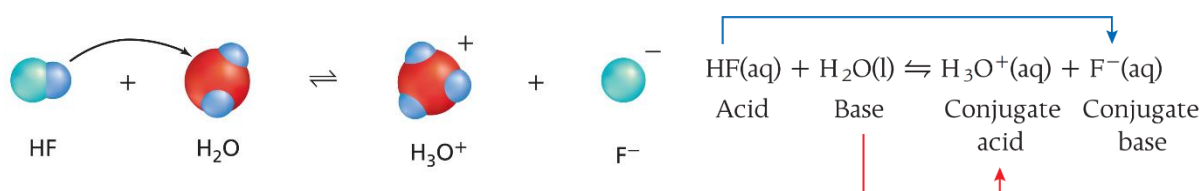


A **conjugate acid** is the species produced when a base accepts a hydrogen ion. The base H_2O accepts a hydrogen ion from the acid HX and becomes the conjugate acid H_3O^+ .

A **conjugate base** is the species that results when an acid donates a hydrogen ion. The acid HX donates its hydrogen ion and becomes the conjugate base X^- .

A **conjugate acid-base pair** consists of two substances related to each other by the donating and accepting of a single hydrogen ion. The X^- ion is the conjugate base of the acid HX .

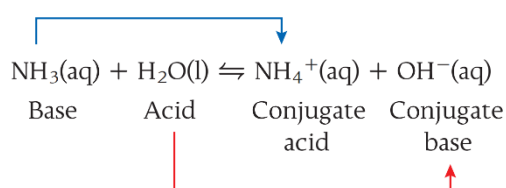
Hydrogen Fluoride – A Brønsted-Lowry Acid



Hydrogen fluoride donates a hydrogen ion to a water molecule, so hydrogen fluoride is an acid.

Ammonia – A Brønsted-Lowry Base

Some substances that lack a hydroxide group and cannot be considered bases according to the Arrhenius definition but can be classified as bases according to the Brønsted-Lowry model. One example is ammonia (NH_3).



Because the NH_3 molecule accepts a H^+ ion to form the ammonium ion (NH_4^+), ammonia is a Brønsted-Lowry base in the forward reaction. The ammonium ion is the conjugate acid of the base ammonia.

Water – A Brønsted-Lowry Acid and Base

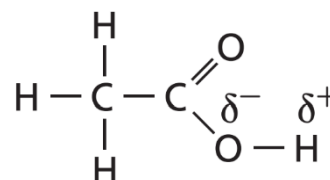
When HF dissolves in water, water acts as a base; when NH_3 dissolves in water, water acts as an acid. Depending on what other substances are in the solution, water can act as either an acid or a base. Water and other substances that can act as both acids and bases are said to be **amphoteric**.

INTRODUCTION TO ACIDS AND BASES

Monoprotic and Polyprotic Acids

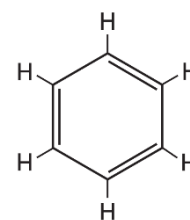
An acid that can donate only one hydrogen ion is called a **monoprotic acid**. Because acetic acid is a monoprotic acid, its formula is often written $\text{HC}_2\text{H}_3\text{O}_2$ to emphasize the fact that only one of the four hydrogen atoms in the molecule is ionizable.

The difference between acetic acid's ionizable hydrogen atom and the other three hydrogen atoms is that the ionizable atom is bonded to the element oxygen, which is more **electronegative** than hydrogen. The difference in electronegativity makes the bond between oxygen and hydrogen **polar**.



The hydrogen atom in **hydrogen fluoride** is bonded to the highly electronegative fluorine atom, so the hydrogen-fluorine bond is **polar** and the fluorine atom is ionizable to a certain extent.

However, the hydrogen atoms in **benzene** are each bonded to a carbon atom. Carbon atoms have about the same electronegativity as hydrogen. These bonds are **nonpolar**, so benzene is not an acid.



Acids that contain **two** ionizable hydrogen atoms per molecule are called **diprotic acids**. Acids with **three** hydrogen ions to donate are called **triprotic acids**. The term **polyprotic acid** can be used for any acid that has more than one ionizable hydrogen atom.

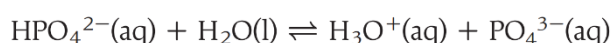
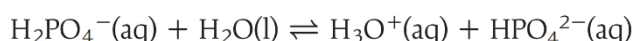
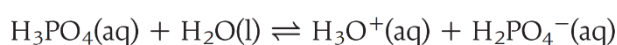
Monoprotic acids:

- HCl – Hydrochloric acid
- HF – Hydrofluoric acid
- HClO_4 – Perchlorate
- HNO_3 – Nitrate
- HBr – Hydrobromic acid
- CH_3COOH – Acetic acid

Polyprotic acids:

- H_2SO_4 – Sulfuric acid
- H_2CO_3 – Carbonic acid
- H_3PO_4 – Phosphoric acid
- H_3BO_3 – Boric acid

All polyprotic acids ionize **in steps**. The three ionizations of phosphoric acid are described by the following equations.



INTRODUCTION TO ACIDS AND BASES

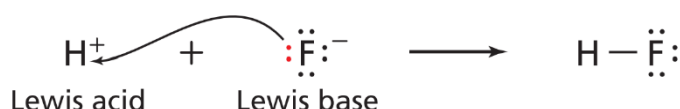
The Lewis Model

Lewis developed the **electron-pair** theory of chemical bonding and introduced Lewis structures to keep track of the electrons in atoms and molecules.

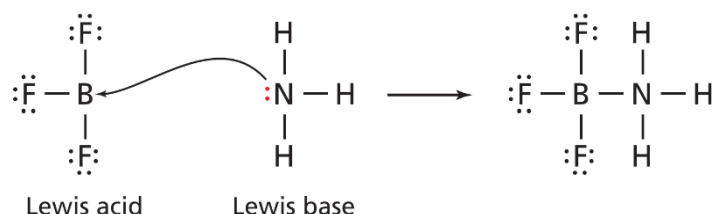
Lewis proposed that an **acid** is an ion or molecule with a **vacant atomic orbital** that can **accept** (share) an electron pair. A **base** is an ion or molecule with a **lone electron pair** that it can **donate** (share).

According to the **Lewis model**, a Lewis acid is an electron-pair **acceptor** and a Lewis base is an electron-pair **donor**. Note that the Lewis model includes all the substances classified as Brønsted-Lowry acids and bases and many more.

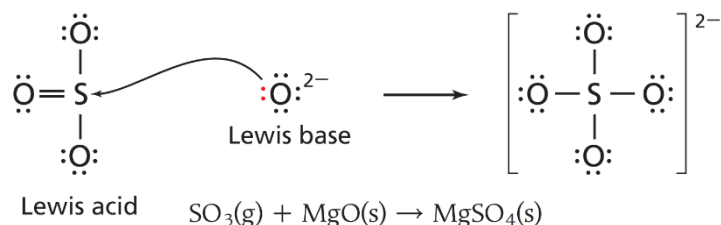
Electron Pair Donors and Acceptors



In this reaction, the **H⁺** ion is the Lewis acid. Its vacant 1s orbital accepts an electron pair from the **F⁻** ion. The fluoride ion is the Lewis base. It donates a lone electron pair to form the hydrogen-fluorine bond in **HF**.



The reaction of gaseous boron trifluoride (**BF₃**) with gaseous ammonia (**NH₃**) to form **BF₃NH₃** is a Lewis acid-base reaction.



The **SO₃** molecule, a Lewis acid, accepts an electron pair from the **O²⁻** ion, a Lewis base.

The reaction of SO₃ and MgO is important because it produces magnesium sulfate, a salt that forms the heptahydrate known as **Epsom salt** (MgSO₄·7H₂O). Epsom salt has many uses, including soothing sore muscles and acting as a plant nutrient.

The reaction to form magnesium sulfate also has environmental applications. When MgO is injected into the flue gases of coal-fired power plants, it reacts with and removes SO₃.

INTRODUCTION TO ACIDS AND BASES

Table 2 Three Models for Acids and Bases

Model	Acid Definition	Base Definition
Arrhenius	H^+ producer	OH^- producer
Brønsted-Lowry	H^+ donor	H^+ acceptor
Lewis	electron-pair acceptor	electron-pair donor

Anhydrides

An **acid anhydride** is an **oxide** that can combine with water to form an acid. Other oxides combine with water to form bases. For example, calcium oxide (**CaO**, lime) forms the base calcium hydroxide **Ca(OH)₂**. In general,

- oxides of **metallic** elements form **bases**;
- oxides of **nonmetals** form **acids**.

STRENGTHS OF ACIDS AND BASES

Strengths of Acids

In solution, strong acids and bases ionize completely, but weak acids and bases ionize only partially.

Strong Acids

HCl solution conducts electricity better than $\text{HC}_2\text{H}_3\text{O}_2$ solution. This is because ions carry electric current through the solution and all the HCl molecules contained in the solution are ionized completely into hydronium ions and chloride ions.

Strong acids are acids that ionize completely. Because strong acids produce the maximum number of ions, they are good conductors of electricity.

The ionization of **hydrochloric** acid in water can be represented by the following equation, which has a **single arrow** pointing to the right. Recall that a single arrow means that a reaction goes to **completion**.

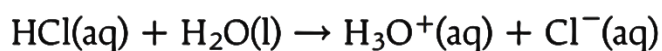


Table 3 Ionization Equations			
Strong Acids		Weak Acids	
Name	Ionization Equation	Name	Ionization Equations
Hydrochloric	$\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$	Hydrofluoric	$\text{HF} \rightleftharpoons \text{H}^+ + \text{F}^-$
Hydroiodic	$\text{HI} \rightarrow \text{H}^+ + \text{I}^-$	Acetic	$\text{HC}_2\text{H}_3\text{O}_2 \rightleftharpoons \text{H}^+ + \text{C}_2\text{H}_3\text{O}_2^-$
Perchloric	$\text{HClO}_4 \rightarrow \text{H}^+ + \text{ClO}_4^-$	Hydrosulfuric	$\text{H}_2\text{S} \rightleftharpoons \text{H}^+ + \text{HS}^-$
Nitric	$\text{HNO}_3 \rightarrow \text{H}^+ + \text{NO}_3^-$	Carbonic	$\text{H}_2\text{CO}_3 \rightleftharpoons \text{H}^+ + \text{HCO}_3^-$
Sulfuric	$\text{H}_2\text{SO}_4 \rightarrow \text{H}^+ + \text{HSO}_4^-$	Hypochlorous	$\text{HClO} \rightleftharpoons \text{H}^+ + \text{ClO}^-$

Weak Acids

Weak acid is an acid that ionizes only partially in dilute aqueous solution. Acetic acid does not ionize completely. Weak acids produce fewer ions and thus cannot conduct electricity as well as strong acids.

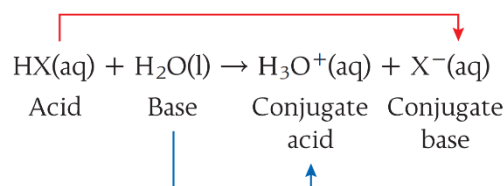
Acid Strength and the Brønsted-Lowry Model

Consider the ionization of any strong acid, **HX**. Remember that the **acid** on the reactant side of the equation produces a **conjugate base** on the product side. Similarly, the **base** on the reactant side produces a **conjugate acid**.

STRENGTHS OF ACIDS AND BASES

Ionization Equilibrium for a Strong Acid

HX represents a **strong acid** and its conjugate base is weak. That is, HX is nearly 100% ionized because H₂O is a stronger base (in the forward reaction) than is the conjugate base X⁻ (in the reverse reaction).

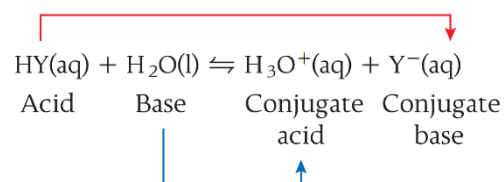


In other words, the ionization equilibrium lies almost completely **to the right** because the base H₂O has a much greater attraction for the H⁺ ion than does the base X⁻.

Think of this as the **battle** of the bases: Which of the two (H₂O or X⁻) has a greater attraction for the hydrogen ion? **In the case of all strong acids, water is the stronger base.**

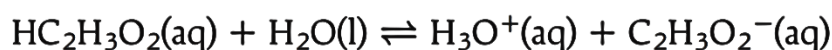
Ionization Equilibrium for a Weak Acid

The ionization equilibrium for a weak acid lies far **to the left** because the conjugate base Y⁻ has a greater attraction for the H⁺ ion than does the base H₂O.



In the battle of the bases, the conjugate base Y⁻ (in the reverse reaction) is stronger than the base H₂O (in the forward reaction) and manages to capture the H⁺ ion.

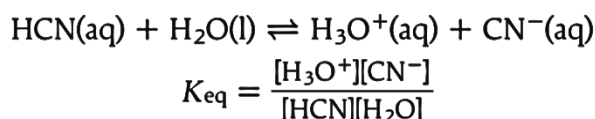
In the case of **acetic acid**, the conjugate base **C₂H₃O₂⁻** (in the reverse reaction) has a stronger attraction for hydrogen ions than does the base H₂O (in the forward reaction).



Acid Ionization Constants

Although the **Brønsted-Lowry model** helps explain acid strength, the model does not provide a **quantitative** way to express the strength of an acid or to compare the strengths of various acids. The **equilibrium constant** expression provides the quantitative measure of acid strength.

Consider **hydrocyanic acid (HCN)**, also known as **prussic acid** which is used in dyeing, engraving, and tempering steel. The ionization equation and equilibrium constant expression for hydrocyanic acid are as follows.



The concentration of liquid H₂O in the denominator of the expression is **constant** in dilute aqueous solutions, so it can be combined with **K_{eq}** to give a new equilibrium constant, **K_a**.

$$K_{\text{eq}} [\text{H}_2\text{O}] = K_{\text{a}} = \frac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}]} = 6.2 \times 10^{-10}$$

STRENGTHS OF ACIDS AND BASES

K_a is called the acid ionization constant. The **acid ionization constant** is the value of the equilibrium constant expression for the ionization of a weak acid.

The value of K_a indicates whether reactants or products are **favored** at equilibrium. For weak acids, the concentrations of the ions (products) in the numerator tend to be small compared to the concentration of un-ionized molecules (reactant) in the denominator.

The weakest acids have the **smallest** K_a values because their solutions have the lowest concentrations of ions and the highest concentrations of un-ionized acid molecules.

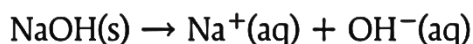
Note that polyprotic acids are not necessarily strong acids for any of their ionizations. Each ionization of a polyprotic acid has a K_a value, and the values decrease for each successive ionization.

Strengths of Bases

What you have read about acids can be applied to bases, except that OH^- ions, rather than H^+ ions, are involved. The conductivity of a base depends on the extent to which the base produces OH^- ions in aqueous solution.

Strong Bases

Strong base is a base that dissociates entirely into metal ions and hydroxide ions. Therefore, metallic hydroxides, such as sodium hydroxide (NaOH), are strong bases.



Some metallic hydroxides, such as calcium hydroxide (Ca(OH)_2) have low solubility and thus are poor sources of OH^- ions. Note that the solubility product constant, K_{sp} , for calcium hydroxide (Ca(OH)_2) is small, indicating that few OH^- ions are present in a saturated solution.



WATCH [How to Memorize Strong Acids and Bases](#)

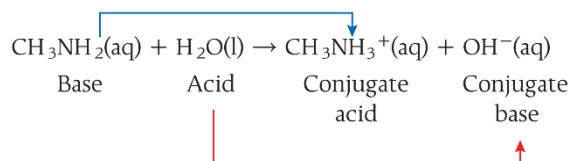
Weak Bases

Weak base is a base that ionizes only partially in dilute aqueous solution.

Methylamine (CH_3NH_2) reacts with water to produce an equilibrium mixture of CH_3NH_2 molecules, CH_3NH_3^+ ions, and OH^- ions.

STRENGTHS OF ACIDS AND BASES

This equilibrium lies far **to the left** because the base, CH_3NH_2 , is weak and the conjugate base, OH^- , is strong. The hydroxide ion has a greater attraction for a hydrogen ion than has a molecule of methylamine.



Base Ionization Constants

The equilibrium constant provides a measure of the extent of the base's ionization. The equilibrium constant for the ionization of methylamine in water is defined by the following equilibrium constant expression.

$$K_b = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]}$$

The **base ionization constant, K_b** , is the value of the equilibrium constant expression for the ionization of a base. The smaller the value of K_b , the weaker the base.

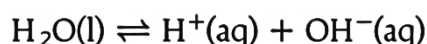
HYDROGEN IONS AND PH

Ion Product Constant for Water

pH and pOH are logarithmic scales that express the concentrations of hydrogen ions and hydroxide ions in aqueous solutions.

Writing K_w

Pure water contains **equal** concentrations of H^+ and OH^- ions produced by self-ionization. The **double arrow** indicates that this is an **equilibrium**.



The Ion Product of Water

$$K_w = [H^+][OH^-]$$

In dilute aqueous solutions, the product of the concentrations of the hydrogen ion and the hydroxide ion equals K_w .

The expression for K_w is a special equilibrium constant expression that applies **only** to water. The **ion product constant for water** K_w is the value of the equilibrium constant expression for the self-ionization of water.

Experiments show that in pure water at **298 K**, $[H^+]$ and $[OH^-]$ are both equal to **$1.0 \times 10^{-7} M$** . Therefore, at 298 K, the value of K_w is **1.0×10^{-14}** .

$$\begin{aligned} K_w &= [H^+][OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) \\ K_w &= 1.0 \times 10^{-14} \end{aligned}$$

K_w and Le Châtelier's Principle

Adding extra hydrogen ions to water at equilibrium is a **stress** on the system. The system reacts in a way to **relieve** the stress. The added H^+ ions **react** with OH^- ions to form more water molecules. Thus, the concentration of OH^- ions **decreases**.

pH and pOH

Chemists express the concentration of hydrogen ions using a pH scale based on common logarithms.

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

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

HYDROGEN IONS AND PH

At 298 K,

- **acidic solutions** have pH values **below 7**.
- **Basic solutions** have pH values **above 7**.

The logarithmic nature of the pH scale means that a change of **one pH** unit represents a **tenfold** change in ion concentration. A solution having a pH of **3.0** has **ten times** the hydrogen ion concentration of a solution with a pH of **4.0**.

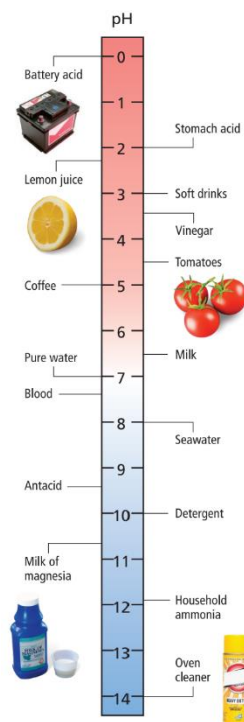
$$\text{pOH} = -\log[\text{OH}^-]$$

*The **pOH** of a solution equals the **negative logarithm** of the **hydroxide ion concentration**.*

At 298 K,

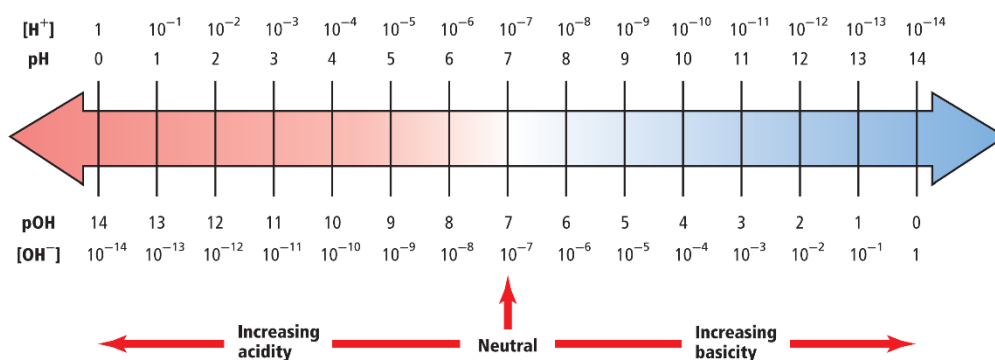
- **Basic solutions** have pOH values **below 7**.
- **acidic solutions** have pOH values **above 7**.

As with the pH scale, a change of one pOH unit expresses a tenfold change in ion concentration.



$$\text{pH} + \text{pOH} = 14.00$$

The sum of pH and pOH is 14.00.



Molarity and the pH of Strong Acids

For all **strong monoprotic acids**, the concentration of the acid is the concentration of **H⁺** ions. Thus, you can use the molarity of the acid to calculate pH.

Every HCl molecule produces one **H⁺** ion. A bottle labeled **0.1 M HCl** contains **0.1 mol of H⁺** ions per liter and 0.1 mol of **Cl⁻** ions per liter.

HYDROGEN IONS AND PH

Molarity and the pH of Strong Bases

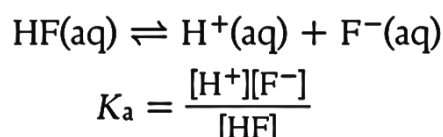
Similarly, the **0.1 M** solution of the **strong base NaOH** is fully ionized.

Some strong bases, such as calcium hydroxide, **Ca(OH)₂**, contain two or more **OH⁻** ions in each formula unit. The concentration of OH⁻ ion in a solution of Ca(OH)₂ is **twice** the molarity of the ionic compound.

Although **strong acids** and **strong bases** are **completely ionized** in dilute aqueous solutions, remember that **weak acids** and **weak bases** are only **partially ionized**. Therefore, you must use **K_a** and **K_b** values to determine the concentrations of H⁺ and OH⁻ ions in solutions of weak acids and bases.

Calculating K_a from pH

Suppose you measured the **pH** of a **0.100 M** solution of the weak acid **HF** and found it to be **3.20**.



From the pH, you could calculate [H⁺]. Then, for every mole per liter of **H⁺ ion** there must be an equal concentration of **F⁻ ion**. That means you know two of the variables in the **K_a** expression.

The concentration of **HF** at equilibrium is equal to the **initial concentration** of the acid (**0.100M**) minus the moles per liter of **HF** that dissociated, which is equal to [H⁺].

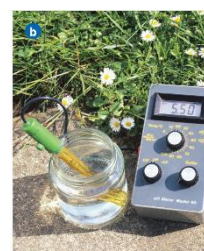
Measuring pH

All pH paper is treated with one or more substances called **indicators** that **change color depending on the concentration of hydrogen ions in a solution**. While litmus paper only indicates if a substance is an acid or a base, pH papers treated with several indicators can identify a range of pH values. **Phenolphthalein** is another example of an indicator.



When a strip of pH paper is dipped into an acidic solution or a basic solution, the color of the paper changes. To determine the pH, the new color of the paper is compared with standard pH colors on a chart.

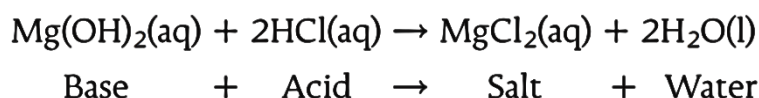
The **pH meter** provides a more accurate measure of pH. When electrodes are placed in a solution, the meter gives a direct analog or digital readout of pH.



NEUTRALIZATION

Reactions Between Acids and Bases

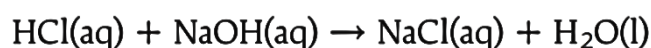
In a **neutralization reaction**, an acid reacts with a base to produce a salt and water. When $\text{Mg}(\text{OH})_2$ and HCl react, a neutralization reaction occurs.



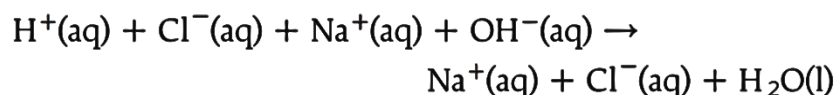
A **salt** is an ionic compound made up of a cation from a base and an anion from an acid. Neutralization is a **double-replacement** reaction.

Writing Neutralization Equations

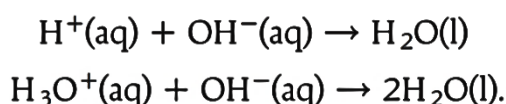
When writing neutralization equations, you must know whether all of the reactants and products in the solution exist as **molecules** or as **formula units**.



Because HCl is a strong acid, NaOH a strong base, and NaCl a soluble salt, all three compounds exist as ions in an aqueous solution.



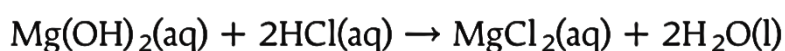
The **chloride ion** and the **sodium ion** appear on both sides of the equation, so they are **spectator ions**. They can be eliminated to obtain the **net ionic equation** for the neutralization of a strong acid by a strong base.



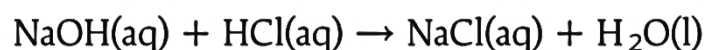
In an aqueous solution, a H^+ ion exists as a H_3O^+ ion.

Acid-base Titration

In the **antacid** reaction below, 1 mol of $\text{Mg}(\text{OH})_2$ neutralizes 2 mol of HCl .



In the reaction of **sodium hydroxide** and **hydrogen chloride**, 1 mol of NaOH neutralizes 1 mol of HCl .



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Stoichiometry provides the basis for a procedure called **titration** (المعايرة), which is used to determine the concentrations of acidic and basic solutions. **Titration is a method for determining the concentration of a solution by reacting a known volume of that solution with a solution of known concentration.**

To find the concentration of an acid solution, you would titrate the acid solution with a solution of a base of known concentration.

Molarity Equation

$$\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2}$$

Titration Procedure:

1. A measured volume of an acidic or basic solution of unknown concentration is placed in a beaker. The electrodes of a pH meter are immersed in this solution, and the initial pH of the solution is read and recorded.
2. A **buret** (سحاحة) is filled with the titrating solution of known concentration. This is called the **standard solution**, or **titrant** (محلول المعايرة).
3. Measured volumes of the standard solution are added slowly and mixed into the solution in the beaker. The pH is read and recorded after each addition.



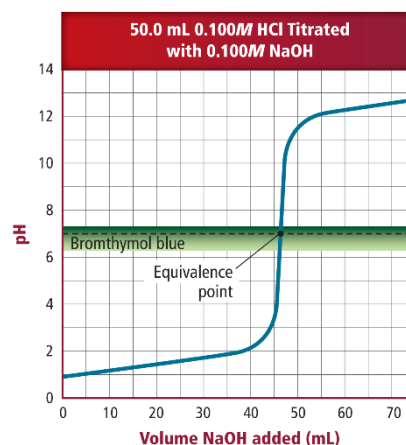
This process continues until the reaction reaches the **equivalence point**, which is the point at which moles of H^+ ion from the acid **equal** moles of OH^- ion from the base.

Strong-Strong Titration

The initial pH of the 0.100M HCl is 1.00. As NaOH is added, the acid is neutralized and the solution's pH increases gradually.

However, when nearly all of the H^+ ions from the acid have been used up, the pH increases dramatically with the addition of an exceedingly small volume of NaOH.

This abrupt increase in pH occurs at the equivalence point of the titration. Beyond the equivalence point, the addition of more NaOH again results in a gradual increase in pH.

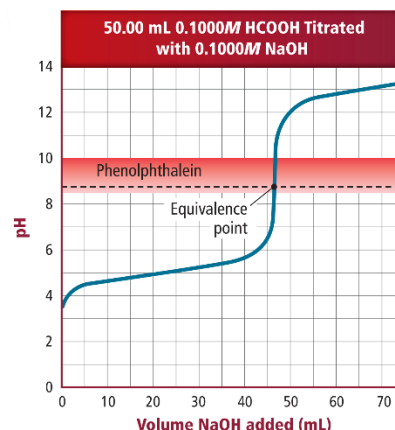


NEUTRALIZATION

Strong-Weak Titration

Some titrations have equivalence points at pH values **less than 7**, and some have equivalence points at pH values **greater than 7**. These differences occur because of reactions between the newly formed salts and water.

The figure shows that the equivalence point for the titration of methanoic acid (a weak acid) with sodium hydroxide (a strong base) lies between pH 8 and pH 9.

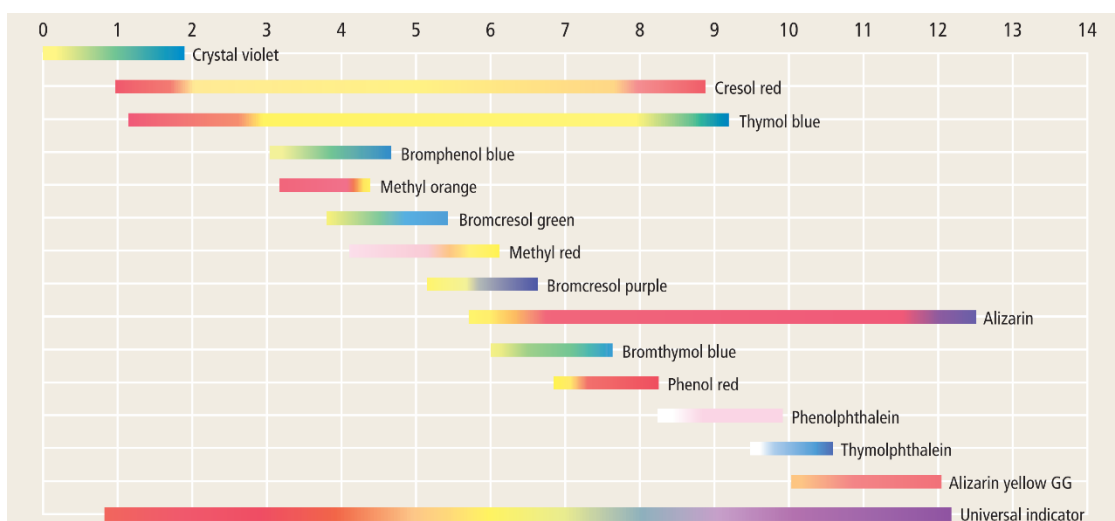


Acid-base Indicators

Acid-base indicators are chemical dyes whose colors are affected by acidic and basic solutions. Chemists often use a chemical dye rather than a pH meter to detect the equivalence point of an acid-base titration.

The familiar dark color of tea becomes lighter when lemon juice is added. A substance contained in tea is an indicator.

Most indicators are large molecules that act as **weak acids**. Slight differences in bonding patterns when an indicator molecule is ionized or un-ionized account for the color changes.



Many of the indicators used by chemists are shown in the figure above. **Bromthymol blue** is a good choice for a titration of a strong acid with a strong base, and **phenolphthalein** changes color at the equivalence point of a titration of a weak acid with a strong base.

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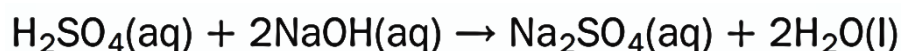
Indicators and Titration End Point

Many indicators used for titration are **weak acids**. Each has its own particular pH or pH ranges over which it changes color.

The **end point** of the titration is the point at which the indicator used in a titration changes color. It is important to choose an indicator for a titration that will change color at the equivalence point of the titration.

Calculating Molarity

The balanced equation for a titration reaction is the key to calculating the unknown molarity. For example, sulfuric acid is titrated with sodium hydroxide according to this equation.



From the equation, you know that the mole ratio of NaOH to H_2SO_4 is **2:1**. Two moles of NaOH are required to neutralize 1 mol of H_2SO_4 . Use the **molarity equation** to find the unknown.

Molarity Equation

$$\frac{M_1 V_1}{n_1} = \frac{M_2 V_2}{n_2}$$

Salt Hydrolysis

Many salts react with water in a process known as salt hydrolysis. In **salt hydrolysis**,

- the anions of the dissociated salt accept hydrogen ions from water
- or the cations of the dissociated salt donate hydrogen ions to water.

The indicator **bromthymol blue** provides surprising results when added to three solutions of ionic salts.

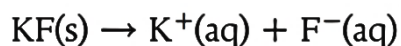
- NaNO_3 turns the indicator green, the solution is neutral.
- KF turns the indicator blue, the solution is basic.
- NH_4Cl turns the indicator yellow, the solution is acidic.



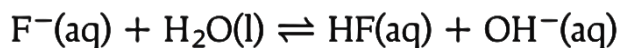
Salts that Produce Basic Solutions

Potassium fluoride is the salt of a strong base (KOH) and a weak acid (HF).

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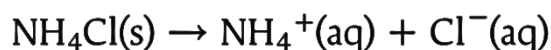
The **K⁺ ions** do not react with water, but the **F⁻ ion** is a weak Brønsted-Lowry base. Some fluoride ions establish this equilibrium with water.



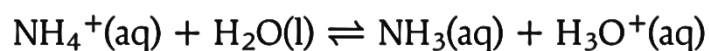
Hydrogen fluoride molecules and **OH⁻ ions** are produced. The production of the **OH⁻ ions** makes the solution **basic**.

Salts that Produce Acidic Solutions

NH₄Cl is the salt of a weak base (NH₃) and a strong acid (HCl).



The **Cl⁻ ions** do not react with water, but the **NH₄⁺ ion** is a weak Brønsted-Lowry acid. Ammonium ions react with water molecules to establish this equilibrium.



Ammonia molecules and hydronium ions are produced. The presence of **hydronium ions** makes the solution **acidic**.

Salts that Produce Neutral Solutions

Sodium nitrate (NaNO₃) is the salt of a strong acid (HNO₃) and a strong base (NaOH). Little or no salt hydrolysis occurs because neither Na⁺ nor NO₃⁻ react with water. Therefore, a solution of sodium nitrate is neutral.

Buffer solutions and their applications are not required for the EOT2 exam this year.

CHAPTER 3

ACIDS AND BASES

CHAPTER SUMMARY

Resources

- *Acids and Bases, from Glencoe Chemistry: Matter and Change ©2017*

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