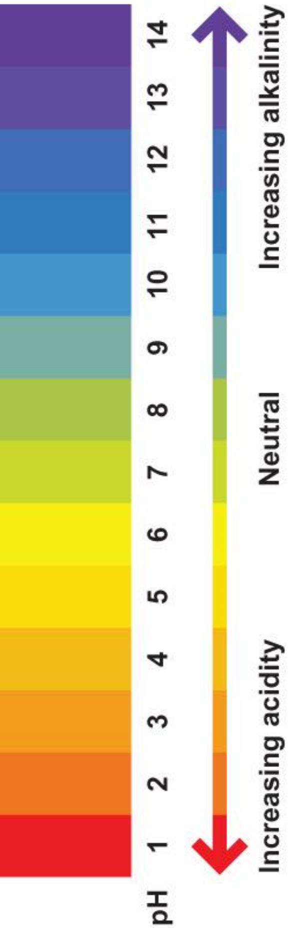

Lecture 2 part 1

Acids and Bases

Chemistry for Teacher II,

Semester 1, Academic Year 2026

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Introduction

Table 18.1 Some Common Acids and Bases and Their Household Uses

Substance	Use
Acids	
Acetic acid, CH_3COOH	Flavoring, preservative
Citric acid, $H_3C_6H_5O_7$	Flavoring
Phosphoric acid, H_3PO_4	Rust remover
Boric acid, H_3BO_3	Mild antiseptic, insecticide
Aluminum salts, $NaAl(SO_4)_2 \cdot 12H_2O$	In baking powder, with sodium hydrogen carbonate
Hydrochloric acid (muriatic acid), HCl	Brick and ceramic tile cleaner
Bases	
Sodium hydroxide (lye), NaOH	Oven and drain cleaners
Ammonia, NH_3	Household cleaner
Sodium carbonate, Na_2CO_3	Water softener, grease remover
Sodium hydrogen carbonate, $NaHCO_3$	Fire extinguisher, rising agent in cake mixes (baking soda), mild antacid
Sodium phosphate, Na_3PO_4	Cleaner for surfaces before painting or wallpapering

Outline

Acid-base: Part I

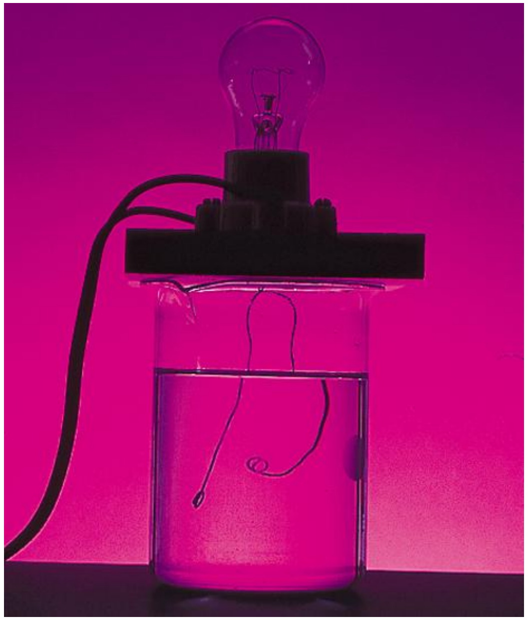
1) Electrolyte Vs nonelectrolytes
2) Strength of Acids and Bases
3) Acid-Base definitions
4) Types of Acids
5) Autoionization of water and Kw
6) pH—A Measure of Acidity
7) Strength of Acids and Bases
8)Calculation of weak acid?/weak bases
9) Percent Ionization

Acid-base: Part II

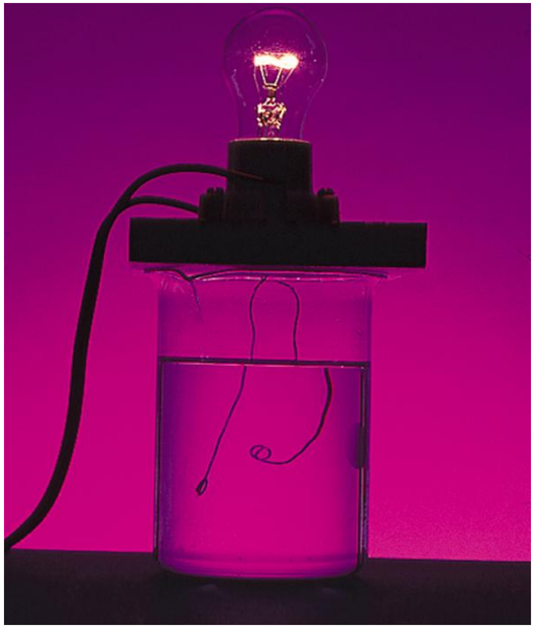
1) neutralization reaction
2) Hydrolysis of Acidic Salts
3) Acid-Base Properties of Salts
4) *Buffer solution*
5) Titration
6) acid-base indicators

An ***electrolyte*** is a substance that, when dissolved in water, results in a solution that can conduct electricity.

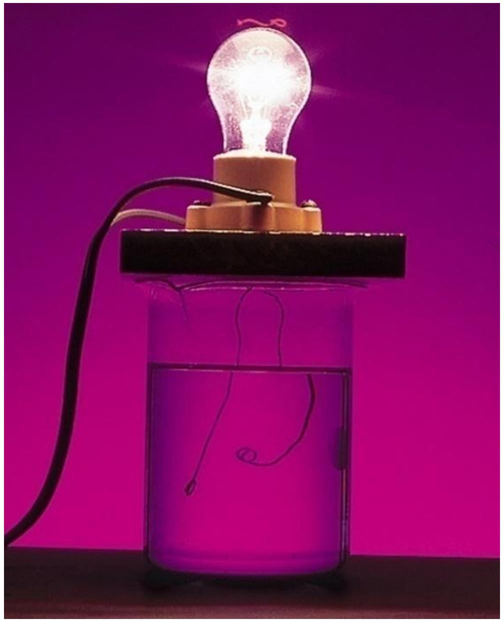
A ***nonelectrolyte*** is a substance that, when dissolved, results in a solution that does not conduct electricity.



nonelectrolyte



weak electrolyte



strong electrolyte

Conduct electricity in solution?

Cations (+) and Anions (-)

Strong Electrolyte – 100% dissociation

$$NaCl\ (s) \xrightarrow{H_2O} Na^+\ (aq) + Cl^-\ (aq)$$

Weak Electrolyte – not completely dissociated

$$CH_3COOH \rightleftharpoons CH_3COO^-\ (aq) + H^+\ (aq)$$

Electrolytes *Versus* Non electrolytes

Table 4.1 Classification of Solutes in Aqueous Solution

Strong Electrolyte	Weak Electrolyte	Nonelectrolyte
HCl	CH_3COOH	$(NH_2)_2CO$ (urea)
HNO_3	HF	CH_3OH (methanol)
$HClO_4$	HNO_2	C_2H_5OH (ethanol)
H_2SO_4*	NH_3	$C_6H_{12}O_6$ (glucose)
$NaOH$	$H_2O^{\dagger}$	$C_{12}H_{22}O_{11}$ (sucrose)
$Ba(OH)_2$		
Ionic compounds		

Strength of Acids and Bases

- Strong/weak acids
- Strong/weak bases

Strength of Acids and Bases

Strong acids

Strong acids are strong electrolytes which, for practical purposes, are assumed to ***ionize completely in water***

Weak acids

Weak acid dissociate very slightly into ions in water.

At equilibrium, aqueous solutions of weak acids contain a mixture of nonionized acid molecules, H^+ ions, and the conjugate base.

Table 4.2 Strong and Weak Acids and Bases

Acids

Strong	
Hydrochloric acid, HCl	
Hydrobromic acid, HBr	
Hydriodic acid, HI	
Nitric acid, HNO_3	
Sulfuric acid, H_2SO_4	
Perchloric acid, $HClO_4$	
Weak	
Hydrofluoric acid, HF	
Phosphoric acid, H_3PO_4	
Acetic acid, CH_3COOH (or $HC_2H_3O_2$)	

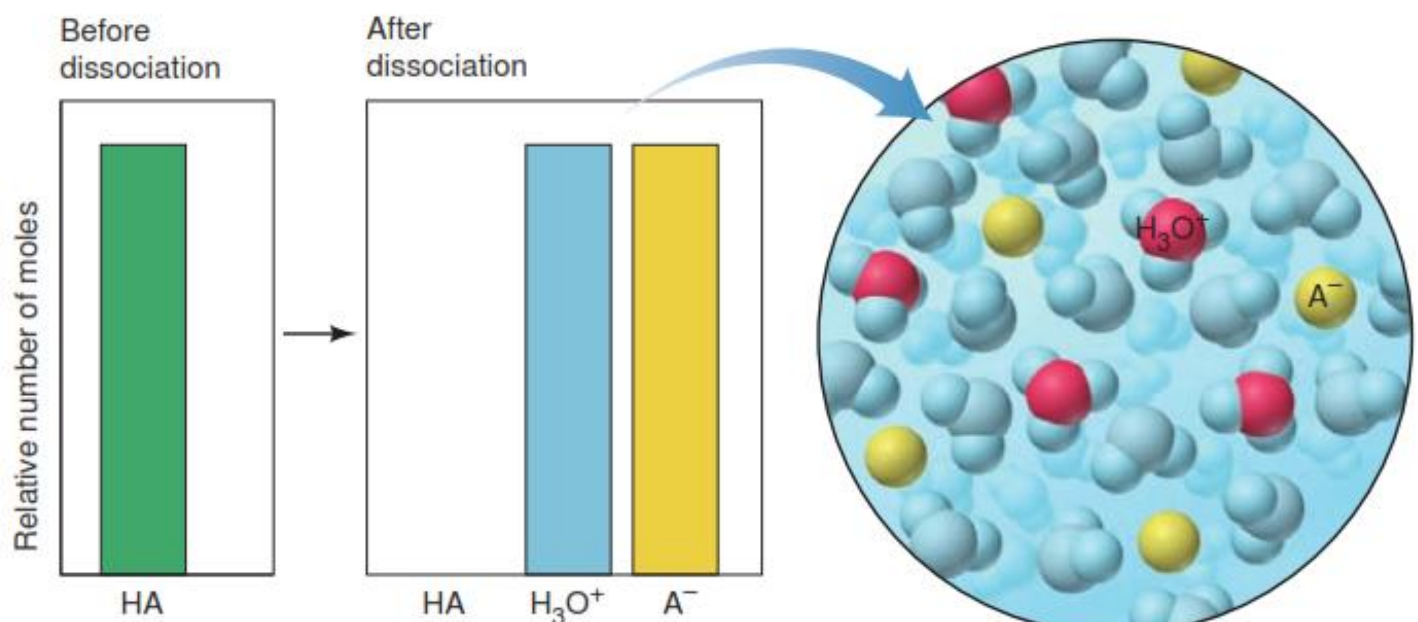
Note that H_2SO_4 is a diprotic acid; we show only the first stage of ionization here.

Strength of Acids and Bases

Strong acids dissociate completely into ions in water

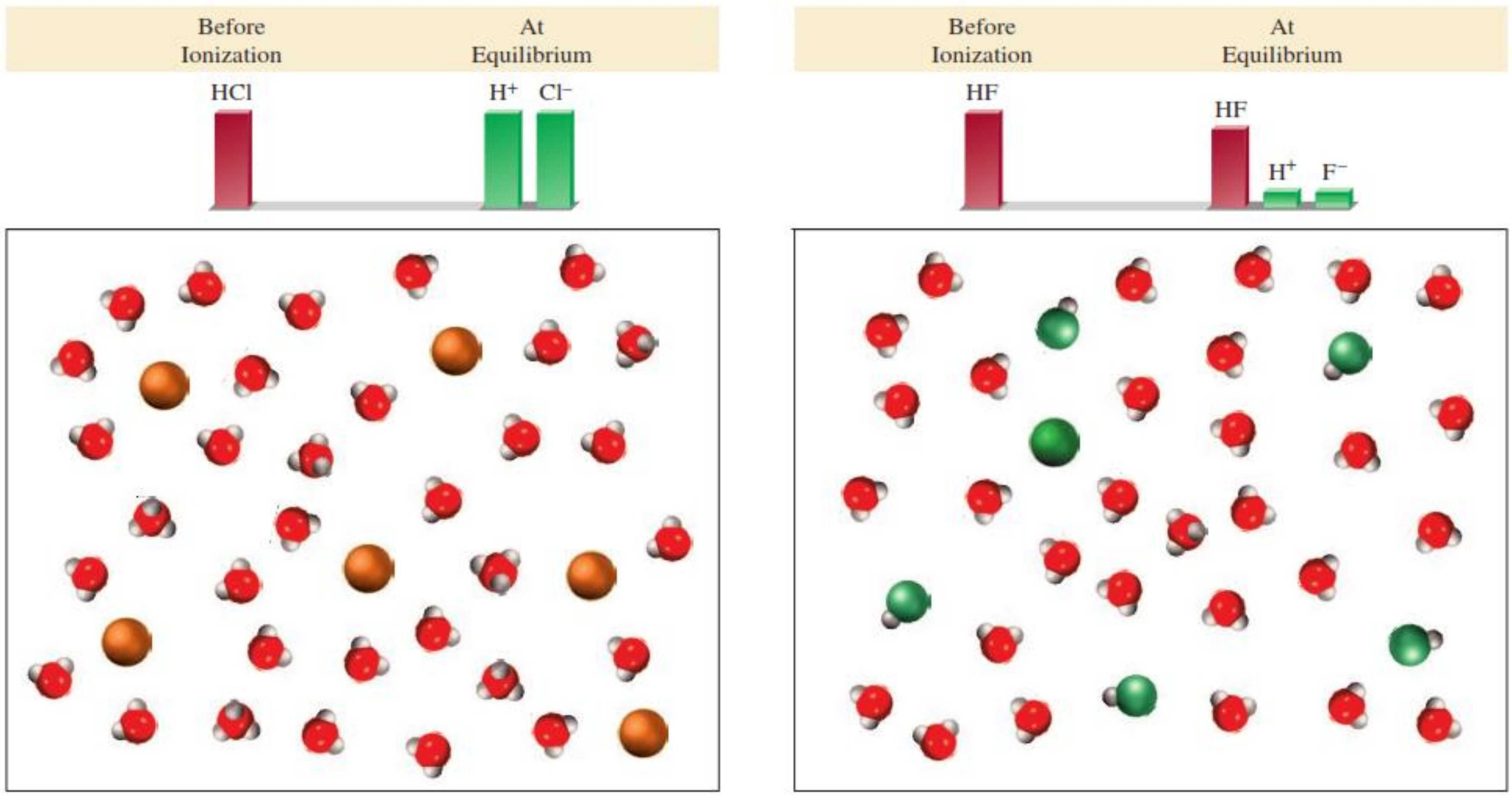
In a dilute solution of a strong acid, virtually no **HA** molecules are present; that is, $[H_3O^+] \approx [HA]_{initial}$.

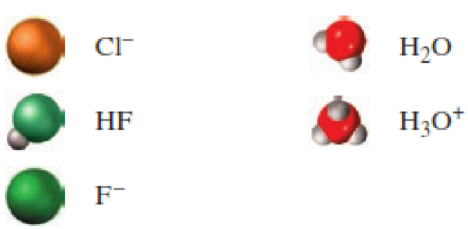
Because the reaction is essentially complete, it is not very useful to express it as an equilibrium process.



Strong acid: $HA(g \text{ or } l) + H_2O(l) \longrightarrow H_3O^+(aq) + A^-(aq)$

The extent of ionization of a strong acid such as HCl (left) and a weak acid





The extent of ionization of a strong acid such as HCl (left) and a weak acid such as HF (right). Initially, there were 6 HCl and 6 HF molecules present.
The strong acid is assumed to be completely ionized in solution.
The proton exists in solution as the hydronium ion (H_3O+)

Strength of Acids and Bases

Strong bases

Strong bases are all strong electrolytes that ionize completely in water

Weak bases

like weak acids, *are weak electrolytes. Ammonia ionizes in water as* follows:

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

Table 4.2 Strong and Weak Acids and Bases

Bases
Strong
Sodium hydroxide, $NaOH$
Potassium hydroxide, KOH
Calcium hydroxide, $Ca(OH)_2$
Strontium hydroxide, $Sr(OH)_2$
Barium hydroxide, $Ba(OH)_2$
Weak
Ammonia, NH_3

Strong/weak QUIZ

Summary : Identify acid/bases – strong and weak

Strong acid/Strong base dissociates/dissolves completely (100%); weak acid/base only a little

Remember

Strong acids: **HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_4$, $HClO_3$**

Strong bases: Group IA hydroxides **[LiOH, NaOH, KOH, RbOH, CsOH]**

Group IIA hydroxides **[$Ca(OH)_2$, $Sr(OH)_2$, and $Ba(OH)_2$]**

Weak acids: ***Any acid that isn't strong***

(common: **HF, HCN, CH_3COOH, H_3PO_4, NH_4^+**, *etc)*

Weak base: NH_3 *(sometime written* NH_4OH*)*

Types of Acids

***1 monoprotic acids:** each unit of the acid yields **one** hydrogen ion upon ionization:*

$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$
$$HNO_3(aq) \longrightarrow H^+(aq) + NO_3^-(aq)$$
$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

***2 diprotic acid:** each unit of the acid gives up ions, in two separate steps:*

$$H_2SO_4(aq) \longrightarrow H^+(aq) + HSO_4^-(aq)$$
$$HSO_4^-(aq) \rightleftharpoons H^+(aq) + SO_4^{2-}(aq)$$

***3 Triprotic acids,** which yield **three H^+ ions**, are relatively few in number.*

$$H_3PO_4(aq) \rightleftharpoons H^+(aq) + H_2PO_4^-(aq)$$
$$H_2PO_4^-(aq) \rightleftharpoons H^+(aq) + HPO_4^{2-}(aq)$$
$$HPO_4^{2-}(aq) \rightleftharpoons H^+(aq) + PO_4^{3-}(aq)$$

Acid–Base

Definition

- Arrhenius
- Brønsted–Lowry
  - Conjugate Acid-Base Pairs
- Lewis

1) Acid-Base definition

1.1) **Arrhenius:** Acids produce H^+ ions in solution,

bases produce OH^- ions.

1.2) **Brønsted–Lowry:** Acids are proton (H^+) donors,

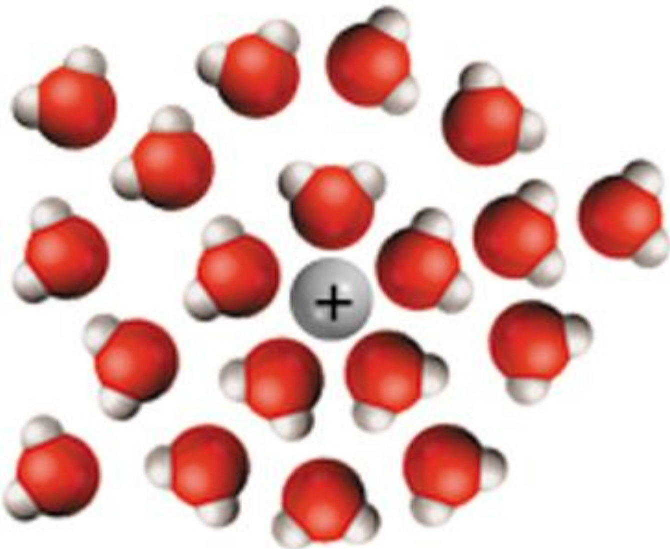
Bases are proton acceptors.

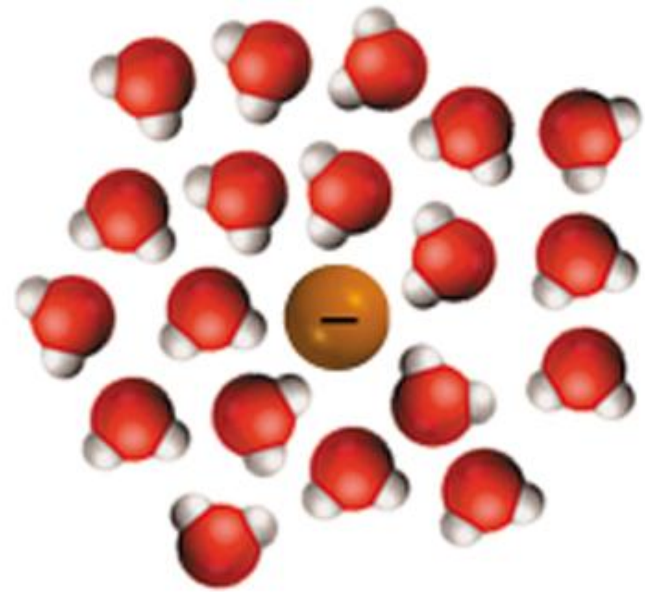
1.2A Conjugate acid-bases pair

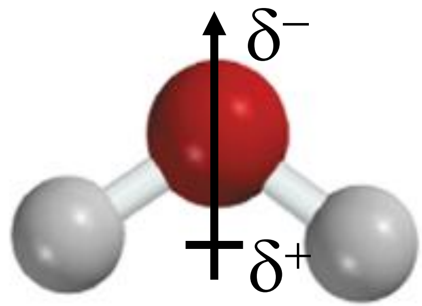
1.3) A ***Lewis acid*** is a substance that can accept a pair of electrons

A ***Lewis base*** is a substance that can donate a pair of electrons

Hydration is the process in which an ion is surrounded by water molecules arranged in a specific manner.







H_2O

The proton bonds covalently to one of the lone electron pairs of a water molecule’s O atom to form a hydronium ion, $\mathbf{H_3O^+}$

1.1) Arrhenius acid-base definition

In the ***Arrhenius acid-base definition***, acids and bases are classified in terms of their formulas and their behavior in water:

Acid

An acid is a substance that has ***H*** in its formula and dissociates in water to yield H_3O^+

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

base

A base is a substance that has ***OH*** in its formula and dissociates in water to yield OH^-

$$NaOH + H_2O \rightleftharpoons Na^+ + OH^-$$

The proton bonds covalently to one of the lone electron pairs of a water molecule's O atom to form a hydronium ion, $\mathbf{H_3O^+}$

1.2) Brønsted Acids and Bases

Arrhenius's definitions of acids and bases are limited in that they apply only to aqueous solutions.

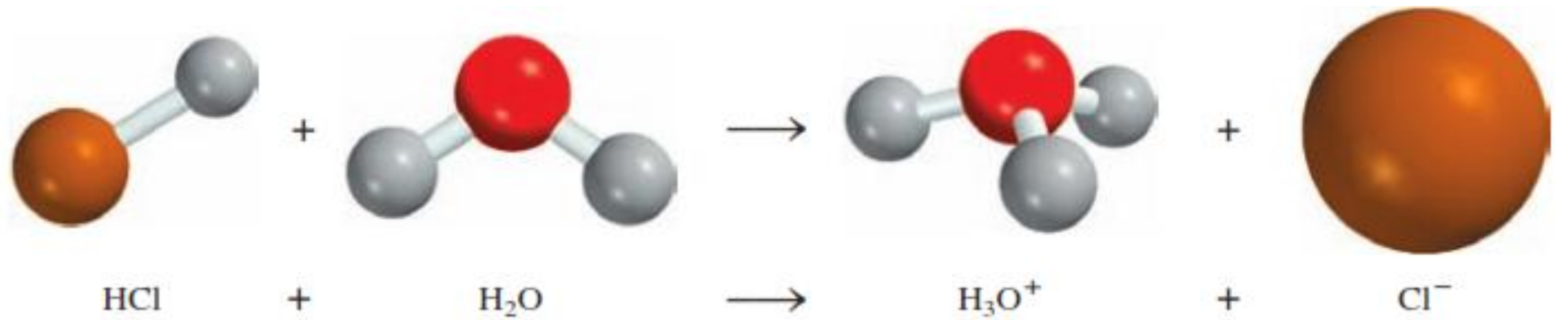
a Brønsted acid is a proton donor

a Brønsted base is a proton acceptor

Note that *Brønsted's definitions do not require acids and bases to be in* aqueous solution

1.2) Brønsted Acids

Example: Hydrochloric acid is a Brønsted acid because it **donates a proton** in water:

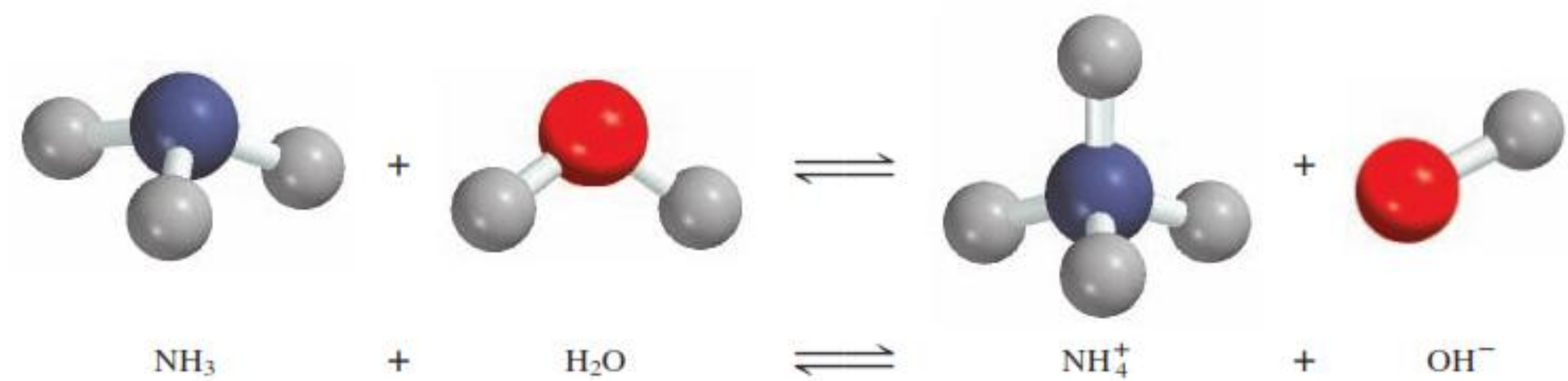


The proton that exists in the hydrated form is called *hydronium ion*

we will generally use H^+(*aq) to represent the hydrated proton. This notation is for convenience, but* H_3O^+ is closer to reality.

1.2) Brønsted Bases

Ammonia (NH_3) is classified as a Brønsted base because it can <u>accept a H^+ ion</u>



Question

Classify each of the following species in aqueous solution as a Brønsted acid or base: (a) HBr, (b) NO_2^- (c) HCO_3^-.

Extension of the Brønsted definition

1.2A) Conjugate Acid-Base Pairs

Conjugate acid-base pair can be defined as an acid and its conjugate base or a base and its conjugate acid.

*A **conjugate base** of a Brønsted acid is the species* that remains when one proton has been removed from the acid

A **conjugate acid** results from the addition of a proton to a Brønsted base

$$HA(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + A^-(aq)$$

acid	base	conjugate acid	conjugate base

Note that the atom in the Brønsted base that accepts a H^+ ion must have a lone pair.

Extension of the Brønsted definition

1.2A) Conjugate Acid-Base Pairs

$$\underset{\text{acid}}{CH_3COOH} + \underset{\text{base}}{H_2O} \rightleftharpoons \underset{\text{conjugate base}}{CH_3COO^-} + \underset{\text{conjugate acid}}{H_3O^+}$$

- the acetate ion (CH_3COO^-) is the *conjugate base* of the acid CH_3COOH
- the hydronium ion (H_3O^+) is *conjugate acid* of base H_2O

$$\underset{\text{base}}{NH_3} + \underset{\text{acid}}{H_2O} \rightleftharpoons \underset{\text{conjugate acid}}{NH_4^+} + \underset{\text{conjugate base}}{OH^-}$$

- the ammonium ion(NH_4^+) is *conjugate acid* of base NH_3
- the OH^- is the *conjugate base* of the acid H_2O

Exercise 1

1. Identify the conjugate acid-base pairs in the reaction between ammonia and hydrofluoric acid in aqueous solution

$$NH_3(aq) + HF(aq) \rightleftharpoons NH_4^+(aq) + F^-(aq)$$

2. Identify the conjugate acid-base pairs for the reaction

$$CN^- + H_2O \rightleftharpoons HCN + OH^-$$

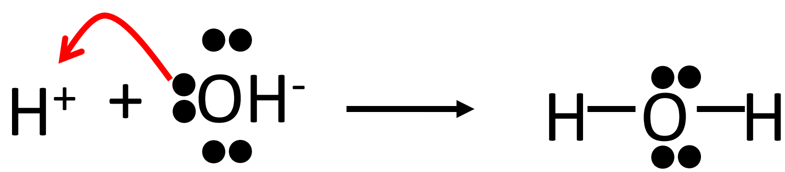
Exercise 2

1. What is the conjugate base of OH^-
2. What is the conjugate acid of OH^-
3. What is the conjugate base of NH_3
4. What is the conjugate acid of NH_3
5. What is the conjugate base of HCOOH

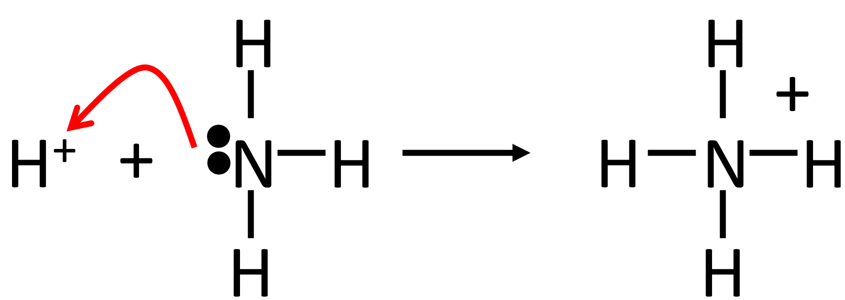
1.3) Lewis Acid and Bases

A ***Lewis acid*** is a substance that can accept a pair of electrons

A ***Lewis base*** is a substance that can donate a pair of electrons

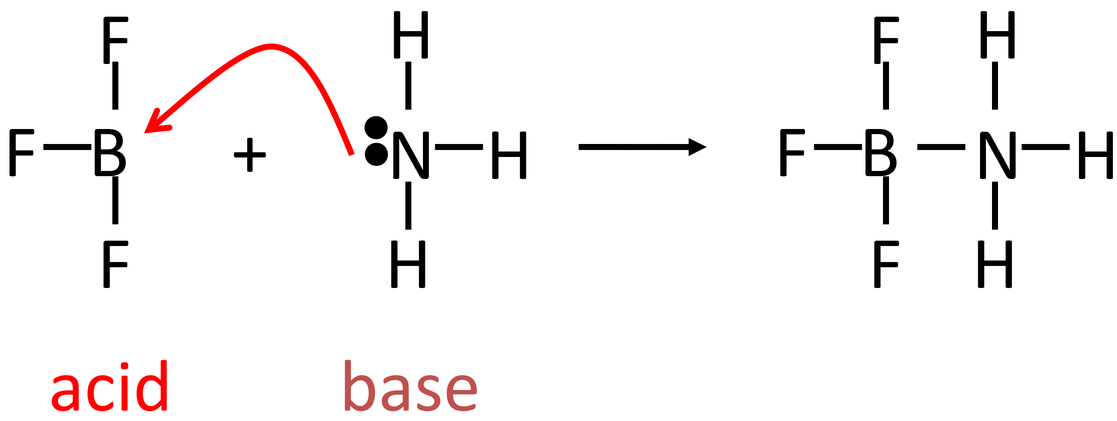


acid base



acid base

1.3) Lewis Acid and Bases

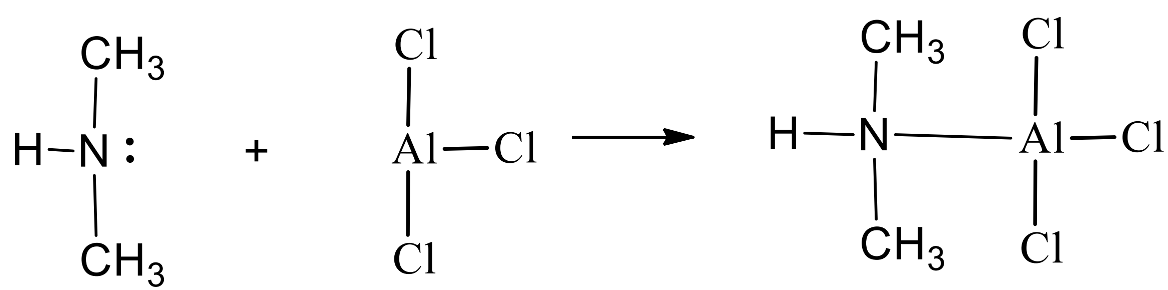


Note that, No protons ***donated*** or ***accepted***!

Example: Identify the Lewis acid and Lewis base in each of the following reactions:

(a) $C_2H_5-\ddot{\underset{\cdot\cdot}{O}}-C_2H_5 \;+\; AlCl_3 \rightleftharpoons (C_2H_5)_2OAlCl_3$

(b)



Practice Exercise: Which of the following cannot behave as a Lewis base?

(a) OF_2

(b) NH_3

(c) CH_4

(d) OH^-

Autoionization of water

Comparing ionic and covalent compounds

Ionic

Ionic compounds have high melting and boiling points

Ionic compounds are usually soluble in water

Ionic compounds conduct electricity, when melted or dissolved in water.

Covalent

Molecular covalent compounds have low melting and boiling points

Covalent compounds tend to be insoluble in water

Covalent compounds do not conduct electricity.

Water is the covalent compound.

Why does water conduct electricity?

4) The Acid-Base Properties of Water

- Water functions as a base in reactions with acids such as HCl and CH_3COOH, and it functions as an acid in reactions with bases such as NH_3.

- Water is a very weak electrolyte and therefore a poor conductor of electricity, but it does undergo ionization to a small extent. This reaction is sometimes called the ***autoionization of water***.

Autoionization of water

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

5) The Ion Product of Water (K_w)

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

The equilibrium constant for the autoionization of water, according to above Equation is;

$$K_w = [H_3O^+][OH^-]$$

where K_w is called the ***ionization constant of water.***
$[H^+]$ is concentration of H^+ ion
$[OH^-]$ is concentration of OH^-

5) The Ion Product of Water (K_w)

In pure water at 25°C, the concentrations of H^+ and OH- are equal and found to be $[H^+] = 1.0 \times 10^{-7}$ *M and* $[OH^-]$ = 1.0×10^{-7} *M .Thus, from Equation* (16.3), at 25°C

$$K_w = [H_3O^+][OH^-]$$

$$K_w = (1.0 \times 10^{-7})(1.0 \times 10^{-7})$$

$$K_w = 1.0 \times 10^{-14}$$

we have pure water or an aqueous solution of dissolved species, the following relation *always* holds at 25°C:

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

5) The Ion Product of Water (K_w)

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

Whenever $[H^+] = [OH^-]$ the aqueous solution is said to be **neutral**.

In an **acidic** solution, there is an excess of H+ ions and

[H+] > [OH-]

In a **basic** solution, there is an excess of hydroxide ions, so

[H+]< [OH-]

6) The Ion Product of Water (K_w)

Example of K_w calculation

The concentration of OH^- ions in a certain household ammonia cleaning solution is 0.0030 *M. Calculate the concentration of* H^+ *ions.*

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

Rearrange the above equation, we write;

$$[H^+] = \frac{K_w}{[OH^-]}$$

$$[H^+] = \frac{1.0 \times 10^{-14}}{0.003}$$

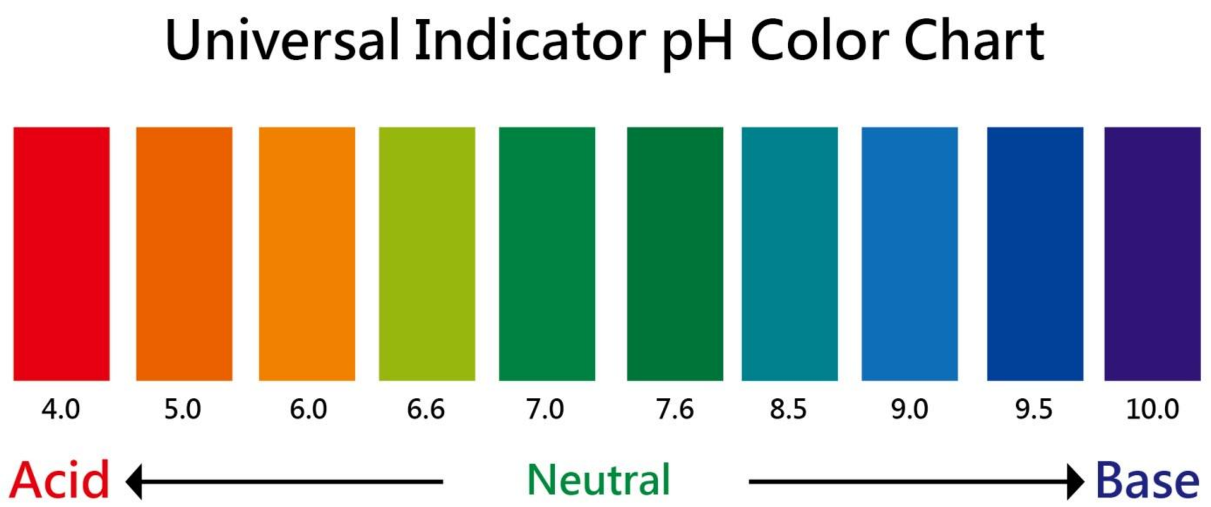
$$[H^+] = 3.3 \times 10^{-12}\,M$$

Question

Practice Exercise 1: Calculate the concentration of OH^- ions in a HCl solution whose hydrogen ion concentration is 1.3 *M.*

Practice exercise 2: Calculate the concentration of $[H^+]$ in a solution with a $[OH^-] = 1.0 \times 10^{-4}$ M.

pH—A Measure of Acidity



pH—A Measure of Acidity

The ***pH of a*** solution is defined as *the negative logarithm of the hydrogen ion concentration (in mol/L):*

$$pH = -\log[H^+]$$ ....Equation (16.5)

Acidic solutions:	$[H^+] > 1.0 \times 10^{-7}\, M$, pH < 7.00
Basic solutions:	$[H^+] < 1.0 \times 10^{-7}\, M$, pH > 7.00
Neutral solutions:	$[H^+] = 1.0 \times 10^{-7}\, M$, pH $= 7.00$

Notice that pH increases as $[H^+]$ decreases.

Sometimes we may be asked to calculate the H^+ ion concentration. In that case, we need to take the antilog of Equation (16.5) as follows:

$$[H^+] = 10^{-pH}$$

pH—A Measure of Acidity

In the laboratory, the pH of a solution is measured with a **pH meter**

Figure 16.2
A pH meter is commonly used in the laboratory to determine the pH of a solution. Although many pH meters have scales marked with values from 1 to 14, pH values can, in fact, be less than 1 and greater than 14.

The pHs of Some Common Fluids

Sample	pH Value
Gastric juice in the stomach	1.0–2.0
Lemon juice	2.4
Vinegar	3.0
Grapefruit juice	3.2
Orange juice	3.5
Urine	4.8–7.5
Water exposed to air*	5.5
Saliva	6.4–6.9
Milk	6.5
Pure water	7.0
Blood	7.35–7.45
Tears	7.4
Milk of magnesia	10.6
Household ammonia	11.5

pOH—A Measure of Acidity

A pOH scale analogous to the pH scale can be devised using the negative logarithm of the hydroxide ion concentration of a solution. Thus, we define pOH as;

$$pOH = -\log[OH^-]$$

We can calculate $[OH^-]$ from equation;

$$[OH^-] = 10^{-pOH}$$

pH and pOH

From the equation: $$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

Taking the negative logarithn of both sides, we obtain

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

$$(-\log[H^+]) + (-\log[OH^-]) = 14.00$$

We obtain;

$$pH + pOH = 14.00$$

Summary: *Equations*

$$K_w = [H^+][OH^-] = 1 \times 10^{-14}$$

$$pH + pOH = 14$$

$$pH = -\log[H^+] \ \& \ pOH = -\log[OH^-]$$

Solution Is

neutral	$[H^+] = [OH^-]$	$pH = 7$
acidic	$[H^+] > [OH^-]$	$pH < 7$
basic	$[H^+] < [OH^-]$	$pH > 7$

$$\begin{tikzcd}[column sep=huge, row sep=huge]
{[H^+]} \arrow[r, shift left, "-\log[H^+]"] \arrow[d, leftrightarrow, "{[H^+][OH^-] = 1 \times 10^{-14}}"'] & pH \arrow[l, shift left, "10^{-pH}"] \arrow[d, leftrightarrow, "pH + pOH = 14"] \\
{[OH^-]} \arrow[r, shift left, "-\log[OH^-]"] & pOH \arrow[l, shift left, "10^{-pOH}"]
\end{tikzcd}$$

$[H^+]\uparrow$ *or* $pH\downarrow \ \Rightarrow$ *acid strength* $\uparrow$

$[OH^-]\uparrow$ *or* $pOH\downarrow \ \Rightarrow$ *base strength* $\uparrow$

Calculating example of pH and pOH relationship

Example: The concentration of H+ ions in a bottle of table wine was 4.1 x 10^{-4} *M right after* the cork was removed. Only half of the wine was consumed. The other half, after it had been standing open to the air for a month, was found to have a hydrogen ion concentration equal to 2.3 x 10^{-3} *M. Calculate the pH of the wine on these two* occasions.

Strategy We are given the H^+ ion concentration and are asked to calculate the pH of the solution. What is the definition of pH?

Solution According to Equation (16.5), pH = $-\log[H^+]$. When the bottle was first opened, $[H^+] = 4.1 \times 10^{-4}$ *M*, which we substitute in Equation (16.5):

$$\begin{aligned} pH &= -\log[H^+] \\ &= -\log(4.1 \times 10^{-4}) = 3.39 \end{aligned}$$

On the second occasion, $[H^+] = 2.3 \times 10^{-3}$ *M*, so that

$$pH = -\log(2.3 \times 10^{-3}) = 2.64$$

Question

Practice Exercisec 1: Nitric acid (HNO_3) is used in the production of fertilizer, dyes, drugs, and explosives. Calculate the pH of a HNO3 solution having a hydrogen ion concentration [H^+] of 0.76 *M.*

Question

Practice Exercise 2 : The pH of rainwater collected in a certain region of the northeastern United States on a particular day was 5.23. **Calculate the H^+ ion concentration** of the rainwater.

Question

Practice Exercise 3 : In a NaOH solution $[OH^-]$ is 7.2×10^{-5} M. Calculate pH of the solution.

Calculation

Strong acid and Strong base Ionization

Summary : Identify acid/bases – strong and weak

Strong acid/Strong base dissociates/dissolves completely (100%); weak acid/base only a little

Remember

Strong acids: **HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_4$, $HClO_3$**

Strong bases: Group IA hydroxides **[LiOH, NaOH, KOH, RbOH, CsOH]**

Group IIA hydroxides **[$Ca(OH)_2$, $Sr(OH)_2$, and $Ba(OH)_2$]**

Weak acids: ***Any acid that isn't strong***

(common: **HF, HCN, CH_3COOH, H_3PO_4, NH^{4+}**, *etc)*

Weak base: NH_3 *(sometime written* NH_4OH*)*

Consider this example:

Calculate the pH of (a) a 2.0×10^{-3} *M* ***HNO_3*** *solution and (b) a 0.045 M* ***$Ba(OH)_2$*** solution.

Strategy Keep in mind that HCl is a strong acid and Ba(OH) is a strong base. Thus, these species are completely ionized and no HCl or Ba(OH)2 will be left in solution.

$$HNO_3\ (aq) + H_2O\ (l) \longrightarrow H_3O^+\ (aq) + NO_3^-\ (aq)$$

	HNO_3	H_3O^+	NO_3^-
Start	0.002 *M*	0.0 *M*	0.0 *M*
End	0.0 *M*	0.002 *M*	0.002 *M*

$$pH = -\log [H^+] = -\log [H_3O^+] = -\log(0.002) = 2.7$$

Consider this example:

Calculate the pH of (a) a 2.0×10^{-3} *M* ***HNO_3*** *solution and*

(b) a 1.8×10^{-2} *M* ***$Ba(OH)_2$*** solution.

Strategy Keep in mind that HCl is a strong acid and Ba(OH) is a strong base. Thus, these species are completely ionized and no HCl or $Ba(OH)_2$ will be left in solution.

(b)

$$Ba(OH)_2\,(s) \longrightarrow Ba^{2+}\,(aq) + 2OH^-\,(aq)$$

	$Ba(OH)_2\,(s)$	$Ba^{2+}\,(aq)$	$2OH^-\,(aq)$
Start	0.018 *M*	0.0 *M*	0.0 *M*
End	0.0 *M*	0.018 *M*	0.036 *M*

$$pH = 14.00 - pOH = 14.00 + \log(0.036) = 12.6$$

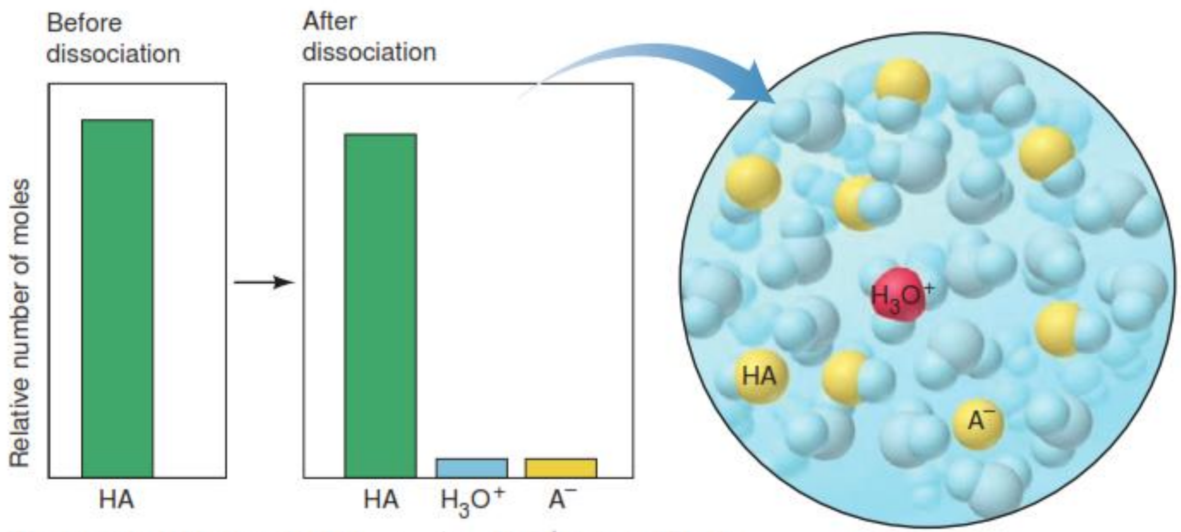
Sample Problem 1

Problem: Calculate the pH of each of the following solutions: (a) 0.0010 *M* HCl, (b) 0.76 *M* KOH.

Sample Problem 2

Problem: Calculate the number of moles of KOH in 5.50 mL of a 0.360 *M* KOH solution. What is the pOH of the solution?

How to measure the acid strength of weak acid?/weak bases



Weak acid: $HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$

How to measure the acid strength of weak acid?/weak bases

1) Weak Acids (HA)
2) Acid Ionization Constants (K_a and K_b)
3) Percent Ionization
4) **Solving weak acid ionization problems**

The Meaning of K_a

The equilibrium expression for the dissociation of a general weak acid, HA, in water is

$$HA\,(aq) + H_2O\,(l) \rightleftharpoons H_3O^+\,(aq) + A^-\,(aq)$$

$$K_c = \frac{[H_3O^+][A^-]}{[HA][H_2O]}$$

we simplify the equilibrium expression by multiplying $[H_2O]$ by K_c to define a new equilibrium constant, the **acid-dissociation constant** (or acid-ionization constant), $\boldsymbol{K_a}$**:**

:

$$K_c[H_2O] = K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

Weak Acids (HA) and Acid Ionization Constants

$$HA\ (aq) + H_2O\ (l) \rightleftharpoons H_3O^+\ (aq) + A^-\ (aq)$$

$$HA\ (aq) \rightleftharpoons H^+\ (aq) + A^-\ (aq)$$

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

K_a is the ***acid ionization constant***

K_a $\uparrow$ weak acid strength $\uparrow$

Weak Bases and Base Ionization Constants

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

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

K_b is the ***base ionization constant***

$K_b \uparrow$ weak base strength $\uparrow$

Solve weak base problems like weak acids ***except*** solve **for [OH⁻] instead of [H⁺].**

The Meaning of K_a

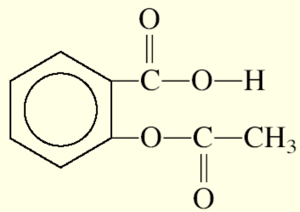
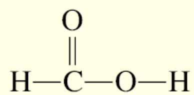
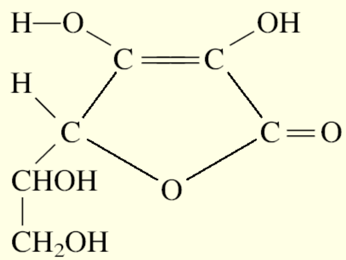
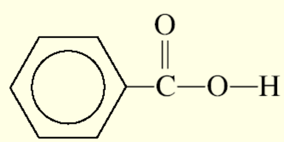
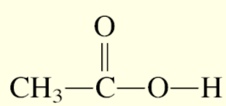
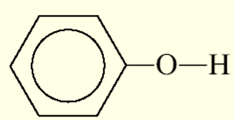
The stronger the acid, the higher the $[H_3O^+]$ at equilibrium, and the larger the Ka

:

Stronger acid $\Longrightarrow$ **higher $[H_3O^+]$** $\Longrightarrow$ **larger K_a**

Stronger base $\Longrightarrow$ **higher $[OH^-]$** $\Longrightarrow$ **larger K_b**

TABLE 15.3 Ionization Constants of Some Weak Acids and Their Conjugate Bases at 25°C

Name of Acid	Formula	Structure	K_a	Conjugate Base	K_b†
Hydrofluoric acid	HF	H—F	7.1×10^{-4}	F^-	1.4×10^{-11}
Nitrous acid	HNO_2	O=N—O—H	4.5×10^{-4}	NO_2^-	2.2×10^{-11}
Acetylsalicylic acid (aspirin)	$C_9H_8O_4$		3.0×10^{-4}	$C_9H_7O_4^-$	3.3×10^{-11}
Formic acid	$HCOOH$		1.7×10^{-4}	$HCOO^-$	5.9×10^{-11}
Ascorbic acid*	$C_6H_8O_6$		8.0×10^{-5}	$C_6H_7O_6^-$	1.3×10^{-10}
Benzoic acid	C_6H_5COOH		6.5×10^{-5}	$C_6H_5COO^-$	1.5×10^{-10}
Acetic acid	CH_3COOH		1.8×10^{-5}	CH_3COO^-	5.6×10^{-10}
Hydrocyanic acid	HCN	H—C≡N	4.9×10^{-10}	CN^-	2.0×10^{-5}
Phenol	C_6H_5OH		1.3×10^{-10}	$C_6H_5O^-$	7.7×10^{-5}

*For ascorbic acid it is the upper left hydroxyl group that is associated with this ionization constant.
†The base ionization constant K_b is discussed in Section 15.6.

Ionization Constants of Conjugate Acid-Base Pairs

Ionization Constants of Conjugate Acid-Base Pairs

$$\cancel{HA}\,(aq) \rightleftharpoons H^+\,(aq) + \cancel{A^-}\,(aq) \qquad K_a$$

$$\cancel{A^-}\,(aq) + H_2O\,(l) \rightleftharpoons OH^-\,(aq) + \cancel{HA}\,(aq) \qquad K_b$$

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

$$\boxed{K_aK_b = K_w}$$

<u>Weak Acid and Its Conjugate Base</u>

$$K_a = \frac{K_w}{K_b} \qquad\qquad K_b = \frac{K_w}{K_a}$$

Summary; Key Equation

EQUILIBRIUM EXPRESSIONS

WEAK ACIDS – K_a (acid dissociation) for weak acids

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

WEAK BASES – K_b (base) for weak bases; *when writing the reaction of a weak base always add water*

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^- \qquad K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

$$\text{Percent ionization} = \frac{[H^+]}{[HA]_0} \times 100\%$$

$$\boxed{K_aK_b = K_w}$$

$$K_w = [H_3O^+][OH^-] = 1.0\times10^{-14}$$

%

Percent

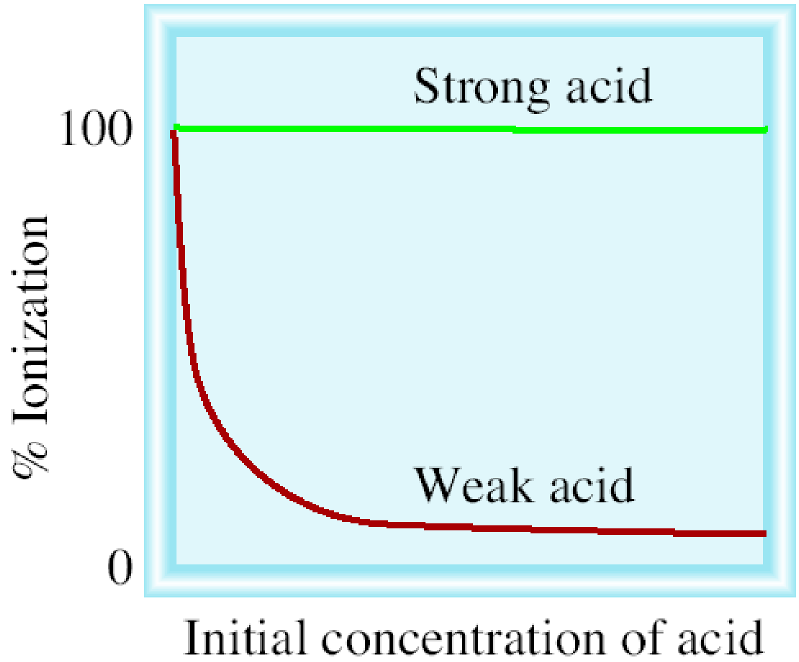
Ionization

$$\textbf{\textit{percent ionization}} = \frac{\text{Ionized acid concentration at equilibrium}}{\text{Initial concentration of acid}} \times 100\%$$

For a monoprotic acid HA

$$\text{Percent ionization} = \frac{[H^+]}{[HA]_0} \times 100\%$$

$[HA]_0$ = initial concentration



$$HF + H_2O \rightleftharpoons H_3O^+ + F^-$$

0.5 M HF

$$\% = \frac{0.019M}{0.5M} \times 100 = 3.8\%$$

0.05 M HF

$$\% = \frac{0.0056M}{0.05M} \times 100 = 11\%$$

How to measure the acid strength of weak acid?/weak bases

1) Weak Acids (HA)
2) Acid Ionization Constants (K_a and K_b)
3) Percent Ionization
4) **Solving weak acid ionization problems**

Chang, R. Chemistry; McGraw-Hill, 2010.

Example: What is the pH of a 0.5 *M* HF solution (at 25^0C)?

$$HF\ (aq) \rightleftharpoons H^+\ (aq) + F^-\ (aq) \qquad K_a = \frac{[H^+][F^-]}{[HF]} = 7.1 \times 10^{-4}$$

	$HF\ (aq)$	$\rightleftharpoons$ $H^+\ (aq)$	$+\ F^-\ (aq)$
Initial (*M*)	0.50	0.00	0.00
Change (*M*)	$-x$	$+x$	$+x$
Equilibrium (*M*)	$0.50 - x$	x	x

$$K_a = \frac{x^2}{0.50 - x} = 7.1 \times 10^{-4}$$

Example: What is the pH of a 0.5 *M* HF solution (at 25^0C)?

$$HF\ (aq) \rightleftharpoons H^+\ (aq) + F^-\ (aq) \qquad K_a = \frac{[H^+][F^-]}{[HF]} = 7.1 \times 10^{-4}$$

$$HF\ (aq) \rightleftharpoons H^+\ (aq) + F^-\ (aq)$$

	HF	H^+	F^-
Initial (*M*)	0.50	0.00	0.00
Change (*M*)	$-x$	$+x$	$+x$
Equilibrium (*M*)	$0.50 - x$	x	x

$$K_a = \frac{x^2}{0.50 - x} = 7.1 \times 10^{-4} \qquad K_a << 1 \quad 0.50 - x \approx 0.50$$

$$K_a \approx \frac{x^2}{0.50} = 7.1 \times 10^{-4} \qquad x^2 = 3.55 \times 10^{-4} \quad x = 0.019\ M$$

$[H^+] = [F^-] = 0.019\ M$ $\qquad$ $pH = -\log [H^+] = 1.72$

$[HF] = 0.50 - x = 0.48\ M$

Exercise 1: What is the pH of 0.2 M benzoic acid solution. (K_a = 6.5 x 10^{-5})

$$C_6H_5COOH \rightleftharpoons C_6H_5COO^- + H^+$$

Exercise 2 What is the pH of a 0.122 *M weak monoprotic acid whose* K_a is 5.7×10^{-4}

Shortcut formulas: For Admission exam only

$$K_a = \frac{[H^+]^2}{M}$$

$$[H^+] = \sqrt{K_a \cdot M}$$

$$\%ionization = \frac{[H^+]}{M} \times 100$$

$$K_b = \frac{[OH^-]^2}{M}$$

$$[OH^-] = \sqrt{K_b \cdot M}$$

$$\%ionization = \frac{[OH^-]}{M} \times 100$$

Exercise 3: What is the pH of a 0.40 *M ammonia solution? (*K_b *of ammonia is* 1.8×10^{-5}*)*

Summary; Solving weak acid/base problem

Use ICE to express the equilibrium concentrations in terms of single unknown *x*

Write K_a in terms of equilibrium concentrations.

Solve for *x* by the approximation method. If approximation is not valid, solve for *x* exactly.

Calculate concentrations of all species and/or pH of the solution.

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

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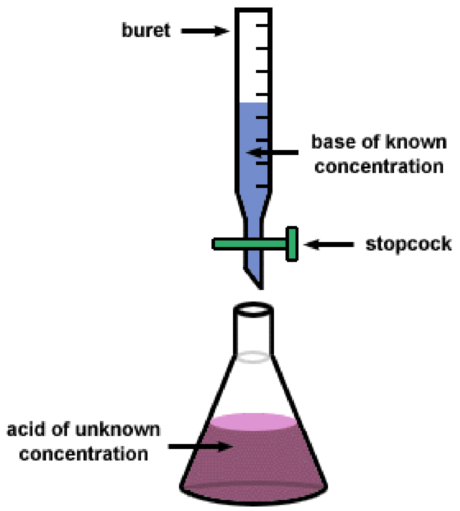
Lecture 2 part 2

Acids and Bases



Basic Chemistry II

Semester 1, Academic year 2026

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Outline

Acid-base: Part II

1) neutralization reaction

2) Hydrolysis of Acidic Salts

3) Acid-Base Properties of Salts

4) Buffer solution

5) Titration

Acid-Base Neutralization (ปฏิกิริยาสะเทิน)

- A ***neutralization reaction*** *is a reaction between an **acid** and **a base.***
- Generally, aqueous acid-base reactions produce water and a **salt.**

Acid-base reaction scheme

$$\text{Acid} + \text{Base} \longrightarrow \text{Salt} + \text{Water}$$

Example

$$HCl + NaOH \longrightarrow NaCl + H_2O$$

Neutralization Reaction

acid + **base** $\longrightarrow$ **salt** + water

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

$H^+ + \cancel{Cl^-} + \cancel{Na^+} + OH^- \longrightarrow \cancel{Na^+} + \cancel{Cl^-} + H_2O$

$H^+ + OH^- \longrightarrow H_2O$

Neutralization Reaction Involving a Weak Electrolyte

weak acid + base $\longrightarrow$ salt + water

$HCN\ (aq) + NaOH\ (aq) \longrightarrow$ **NaCN** $(aq) + H_2O$

$HCN + \cancel{Na^+} + OH^- \longrightarrow \cancel{Na^+} + CN^- + H_2O$

$HCN + OH^- \longrightarrow CN^- + H_2O$

Neutralization Reaction Producing a Gas

acid + base $\longrightarrow$ salt + water + CO_2

$2HCl\ (aq) + Na_2CO_3\ (aq) \longrightarrow 2NaCl\ (aq) + H_2O + CO_2$

$2H^+ + \cancel{2Cl^-} + \cancel{2Na^+} + CO_3^{2-} \longrightarrow \cancel{2Na^+} + \cancel{2Cl^-} + H_2O + CO_2$

$2H^+ + CO_3^{2-} \longrightarrow H_2O + CO_2$

Excercies1: Write the equation to show the reaction between acids and bases

a. $2HCl + Ca(OH)_2 \longrightarrow$

b. $H_2CO_3 + 2KOH \longrightarrow$

c. $2HCl + Ba(OH)_2 \longrightarrow$

d. $HNO_3 + NaOH \longrightarrow$

e. $2HNO_3 + Ca(OH)_2 \longrightarrow$

f. $BaO + H_2SO_4 \longrightarrow$

g. $Zn + 2HCl \longrightarrow$

h. $Na_2CO_3 + HCl \longrightarrow$

i. $Fe_2O_3 + HCl \longrightarrow$

j. $Al(OH)_3 + 3HBr \longrightarrow$

Hydrolysis of Salts

The term **salt hydrolysis** describes the reaction of an anion or a cation of a salt, or both, with water.

Acid-Base properties of salts (Hydrolysis of salts)

- Salts That Produce Acidic Solutions (Acidic Salts)
- Salts That Produce Basic Solutions (Basic Salts)
- Salts That Produce Neutral Solutions (Neutral Salts)

Summary : Identify acid/bases – strong and weak

Strong acid/Strong base dissociates/dissolves completely (100%); weak acid/base only a little

Remember

Strong acids: **HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_4$, $HClO_3$**

Strong bases: Group IA hydroxides **[LiOH, NaOH, KOH, RbOH, CsOH]**

Group IIA hydroxides **[$Ca(OH)_2$, $Sr(OH)_2$, and $Ba(OH)_2$]**

Weak acids: ***Any acid that isn't strong***

(common: **HF, HCN, CH_3COOH, H_3PO_4, NH^{4+}*, etc)***

Weak base: **NH_3 *(sometime written* NH_4OH*)***

Hydrolysis of Acidic Salts

A salt formed between a strong acid and a weak base is an **acidic salt**.

Ammonia is a weak base, and its salt with any strong acid gives a solution with a pH lower than 7. For example, let us consider the reaction:

$$\underset{\text{Strong acid}}{HCl} \quad + \quad \underset{\text{Weak base}}{NH_4OH} \longrightarrow NH_4Cl \quad + \quad H_2O$$

In the solution, the NH_4^+ ion reacts with water (called **hydrolysis**) according to the equation:

$$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$$

$$Cl^- + H_2O \nrightarrow$$

Hydrolysis of Basic Salts

A **basic salt** is formed between a weak acid and a strong base.

The basicity is due to the hydrolysis of the conjugate base of the (weak) acid used in the neutralization reaction.
For example, sodium acetate formed between the weak acetic acid and the strong base NaOH is a basic salt.

$$\underset{\text{weak acid}}{CH_3COOH} + \underset{\text{strong base}}{NaOH} \rightleftharpoons \underset{\text{basic salt}}{CH_3COO\mathbf{Na}} + H_2O$$

When the salt is dissolved, ionization takes place:

$$CH_3COONa \longrightarrow CH_3COO^- + Na^+$$

$$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$$

Acid-Base Properties of Salts

Neutral Salt: **Strong base + strong acid**

Salts containing an alkali metal or alkaline earth metal ion (except Be^{2+}) **and** the conjugate base of a **strong** acid (*e.g.* Cl^-, Br^-, and NO_3^-).

$$NaCl\ (s) \xrightarrow{H_2O} Na^+\ (aq) + Cl^-\ (aq)$$

$$Na^+ + H_2O \not\longrightarrow$$

$$Cl^- + H_2O \not\longrightarrow$$

Salts of weak acids and weak bases

A salt formed between a weak acid and a weak base can be neutral, acidic, or basic depending on the relative strengths of the acid and base.

$$NH_4CN \longrightarrow NH_4^+ + CN^-$$

$$NH_4^+ + H_2O \rightleftharpoons NH_3 + \mathbf{H_3O^+}$$

$$CN^- + H_2O \rightleftharpoons HCN + \mathbf{OH^-}$$

- K_b for the anion > K_a for the cation, solution will be basic
- K_b for the anion < K_a for the cation, solution will be acidic
- K_b for the anion ≈ K_a for the cation, solution will be neutral

TABLE 15.7 Acid-Base Properties of Salts

Type of Salt	Examples	Ions That Undergo Hydrolysis	pH of Solution
Cation from strong base; anion from strong acid	$NaCl$, KI, KNO_3, $RbBr$, $BaCl_2$	None	≈ 7
Cation from strong base; anion from weak acid	CH_3COONa, KNO_2	Anion	> 7
Cation from weak base; anion from strong acid	NH_4Cl, NH_4NO_3	Cation	< 7
Cation from weak base; anion from weak acid	NH_4NO_2, CH_3COONH_4, NH_4CN	Anion and cation	< 7 if $K_b < K_a$
			≈ 7 if $K_b \approx K_a$
			> 7 if $K_b > K_a$
Small, highly charged cation; anion from strong acid	$AlCl_3$, $Fe(NO_3)_3$	Hydrated cation	< 7

<u>Summary:</u> acid base properties of salts

<u>Neutral Salt =</u> **Strong base + strong acid**

acid salt = strong acid + weak base

basic salt = weak acid + strong base.

Salts that contain both WA & WB can produce an acid or base. Which one?
Compare K: **larger K wins!**

if K*a> Kb acidic salt;*

if Kb> Ka basic salt;

Exercise: Predict whether the following solutions will be acidic, basic, or nearly neutral and write **equations to** show how each of the following salts can be prepared from appropriate acid and base;

1) $NaNO_3 \longrightarrow$

2) $NaCN \longrightarrow$

3) $NH_4NO_3 \longrightarrow$

4) $CH_3COONH_4 \longrightarrow$

5) $KBr \longrightarrow$

6) $MgSO_4 \longrightarrow$

7) $CaCO_3 \longrightarrow$

8) $NH_4CN \longrightarrow$

9) $(NH_4)_3PO_4 \longrightarrow$

Exercise 1: Identify the following salts as neutral, acidic, or basic.

NO.	Salts	Neutral salt	Acidic salt	Basic salt
1	$NaCl$			
2	KI			
3	KCN			
4*	CH_3COONH_4			
5	$NaHCO_3$			
6	NH_4NO_3			
7*	NH_4CN			
8	NaF			
9	NH_4ClO_4			
10	$LiNO_3$			

K_a of HCN $= 4.9 \times 10^{-10}$ K_a of CH_3COOH $= 1.8 \times 10^{-5}$

K_b of NH_3 $= 1.8 \times 10^{-5}$ K_a of $HClO_4$ $>> 1.00$

Buffer solution

Buffer solution

A ***buffer solution*** is a solution of:

1. A weak acid or a weak base **and**
2. The salt of the weak acid or weak base

Both must be present!

A buffer solution has the ability to resist changes in pH upon the addition of small amounts of either acid or base.

Buffer solution

Consider an equal molar mixture of CH_3COOH and CH_3COONa

Add strong acid

$$H^+ (aq) + CH_3COO^- (aq) \longrightarrow CH_3COOH (aq)$$

Add strong base

$$OH^- (aq) + CH_3COOH (aq) \longrightarrow CH_3COO^- (aq) + H_2O (l)$$

Buffer solution

In general, a buffer system can be represented as salt-acid or conjugate base–acid.

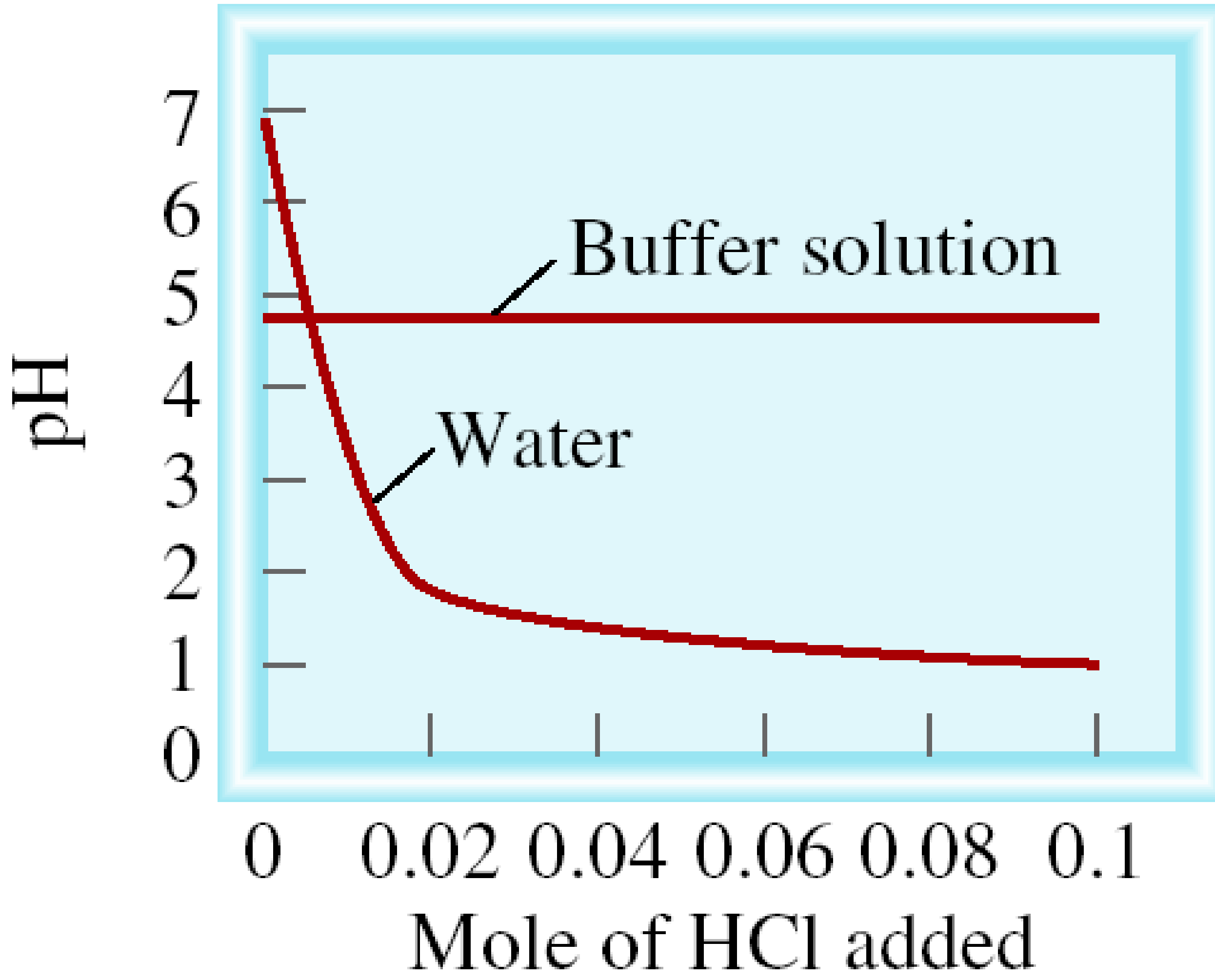
Thus, the sodium acetate–acetic acid buffer system discussed above can be written as;

$$CH_3COONa/CH_3COOH \text{ or simply } CH_3COO^-/CH_3COOH.$$

Strong acid/strong base cannot form buffer solution because it is strong electrolyte (100% dissociation)

$$HCl \longrightarrow H^+ + Cl^-$$

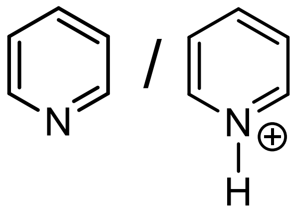
$$HCl + CH_3COO^- \longrightarrow CH_3COOH + Cl^-$$



Question 1

Which of the following are buffer systems?

(a) KF/HF (b) KBr/HBr, (c) NH_4OH/NH_4Cl

(d) 

pH calculation of Buffer solution

Let us consider the pH of a solution containing a weak acid, HA, and a soluble salt of the weak acid, such as NaA. We start by writing;

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$

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

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

$$\boxed{[H^+] = K_a \frac{[acid]}{[salt]}} \qquad \boxed{[OH^-] = K_b \frac{[base]}{[salt]}}$$

Example 1: What is the pH of a solution containing both 0.20 *M* CH_3COOH and 0.30 *M* CH_3COONa? The *K*a of CH_3COOH is 1.8×10^{-5}

Sodium acetate is a strong electrolyte, so it dissociates completely in solution:

$$\underset{}{CH_3COONa(aq)} \longrightarrow \underset{0.30\ M}{Na^+(aq)} + \underset{0.30\ M}{CH_3COO^-(aq)}$$

The initial concentrations, changes, and final concentrations of the species involved in the equilibrium are

	$CH_3COOH(aq) \rightleftharpoons$	$H^+(aq)$ +	$CH_3COO^-(aq)$
Initial (*M*):	0.20	0	0.30
Change (*M*):	$-x$	$+x$	$+x$
Equilibrium (*M*):	$0.20 - x$	x	$0.30 + x$

$$K_a = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$$

$$1.8 \times 10^{-5} = \frac{(x)(0.30 + x)}{0.20 - x}$$

Example 1: What is the pH of a solution containing both 0.20 *M* CH_3COOH and 0.30 *M* CH_3COONa? The *Ka* of CH_3COOH is 1.8 x 10^{-5}

$$K_a = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$$

$$1.8 \times 10^{-5} = \frac{(x)(0.30 + x)}{0.20 - x}$$

Assuming that $0.30 + x \approx 0.30$ and $0.20 - x \approx 0.20$, we obtain

$$1.8 \times 10^{-5} = \frac{(x)(0.30 + x)}{0.20 - x} \approx \frac{(x)(0.30)}{0.20}$$

or

$$x = [H^+] = 1.2 \times 10^{-5}\,M$$

Thus,

$$\begin{aligned} pH &= -\log [H^+] \\ &= -\log (1.2 \times 10^{-5}) = 4.92 \end{aligned}$$

Example 2: What is the pH of a solution containing both 0.30 *M* HCOOH and 0.52 *M* HCOOK? The K_a of HCOOH is 1.7 x 10^{-4}

Example 3: The pH of a bicarbonate-carbonic acid buffer is 8.00. Calculate the ratio of the concentration of carbonic acid (H_2CO_3) of carbonate ion (HCO_3^-.) [K_a of $H_2CO_3 = 4.2 \times 10^{-7}$]

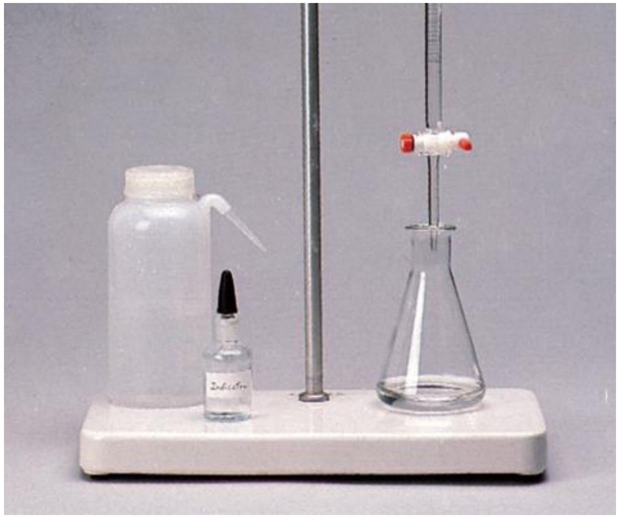
Titrations

Titrations

In a ***titration*** a solution of accurately known concentration is gradually added to another solution of unknown concentration until the chemical reaction between the two solutions is complete.

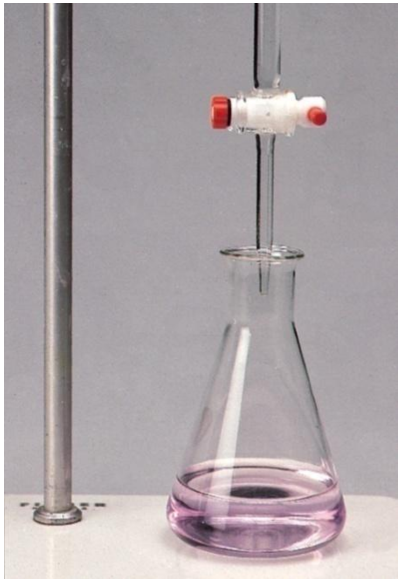
Equivalence point – the point at which the reaction is complete

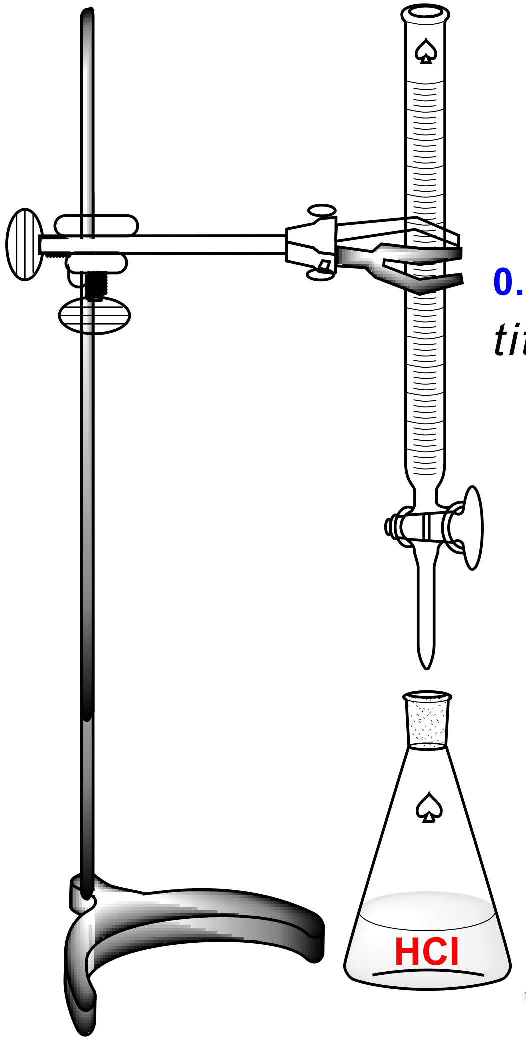
Indicator – substance that changes color at (or near) the equivalence point



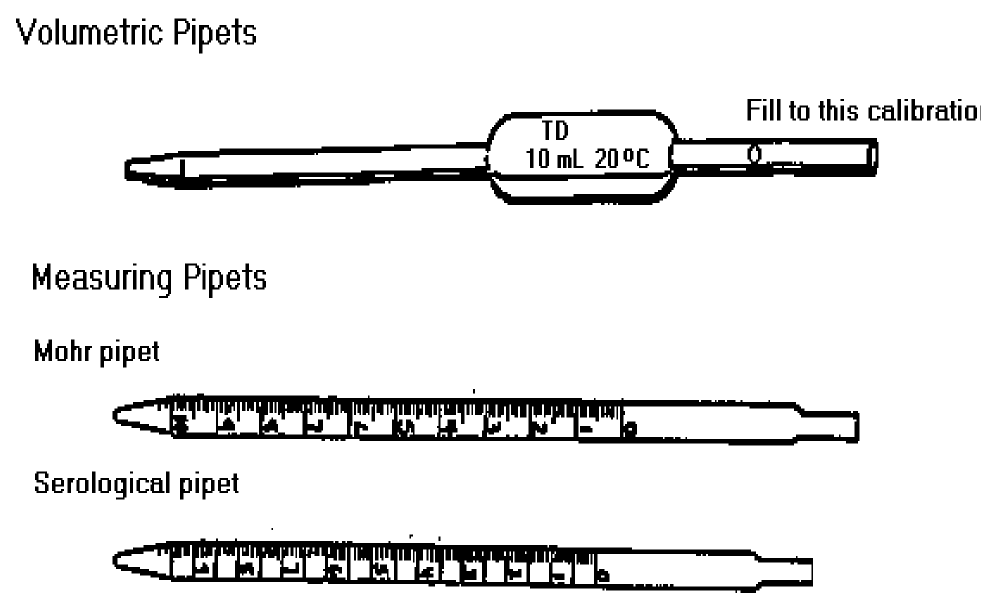
Slowly add base to unknown acid UNTIL

the indicator changes color

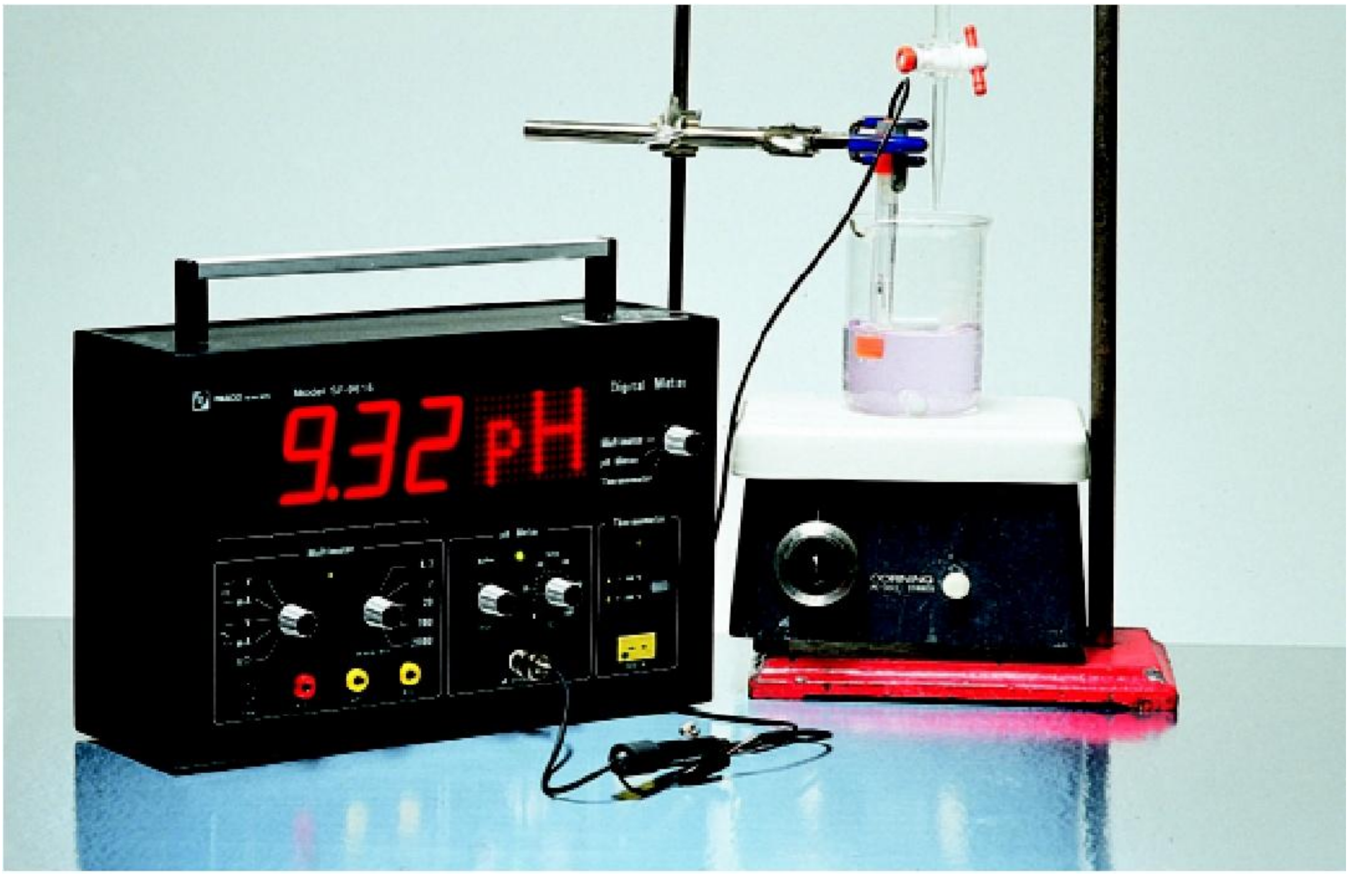




Unknown concentration



Alternative Method of Equivalence Point Detection



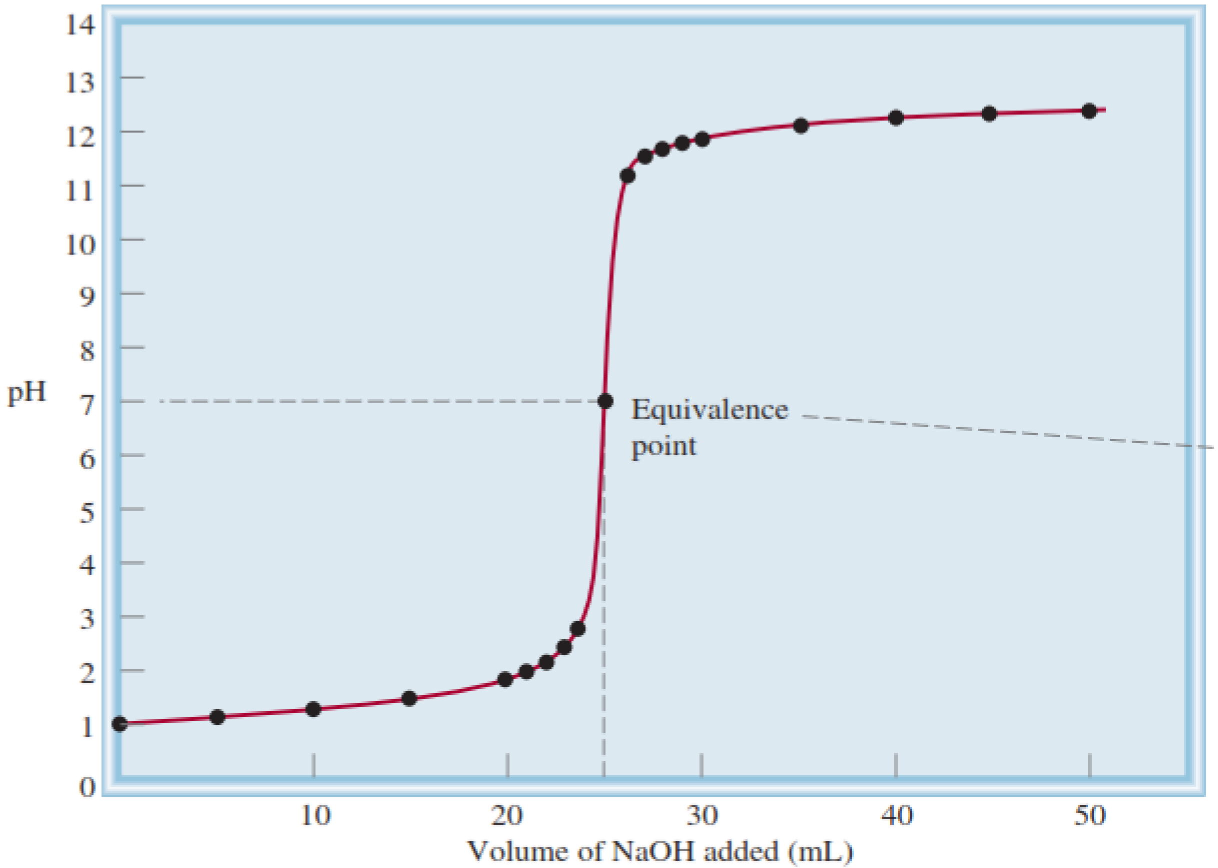
monitor pH

Strong Acid–Strong Base Titrations

Strong Acid–Strong Base Titrations

The reaction between a strong acid (say, HCl) and a strong base (say, NaOH) can be represented by

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



Volume NaOH added (mL)	pH
0.0	1.00
5.0	1.18
10.0	1.37
15.0	1.60
20.0	1.95
22.0	2.20
24.0	2.69
25.0	7.00
26.0	11.29
28.0	11.75
30.0	11.96
35.0	12.22
40.0	12.36
45.0	12.46
50.0	12.52

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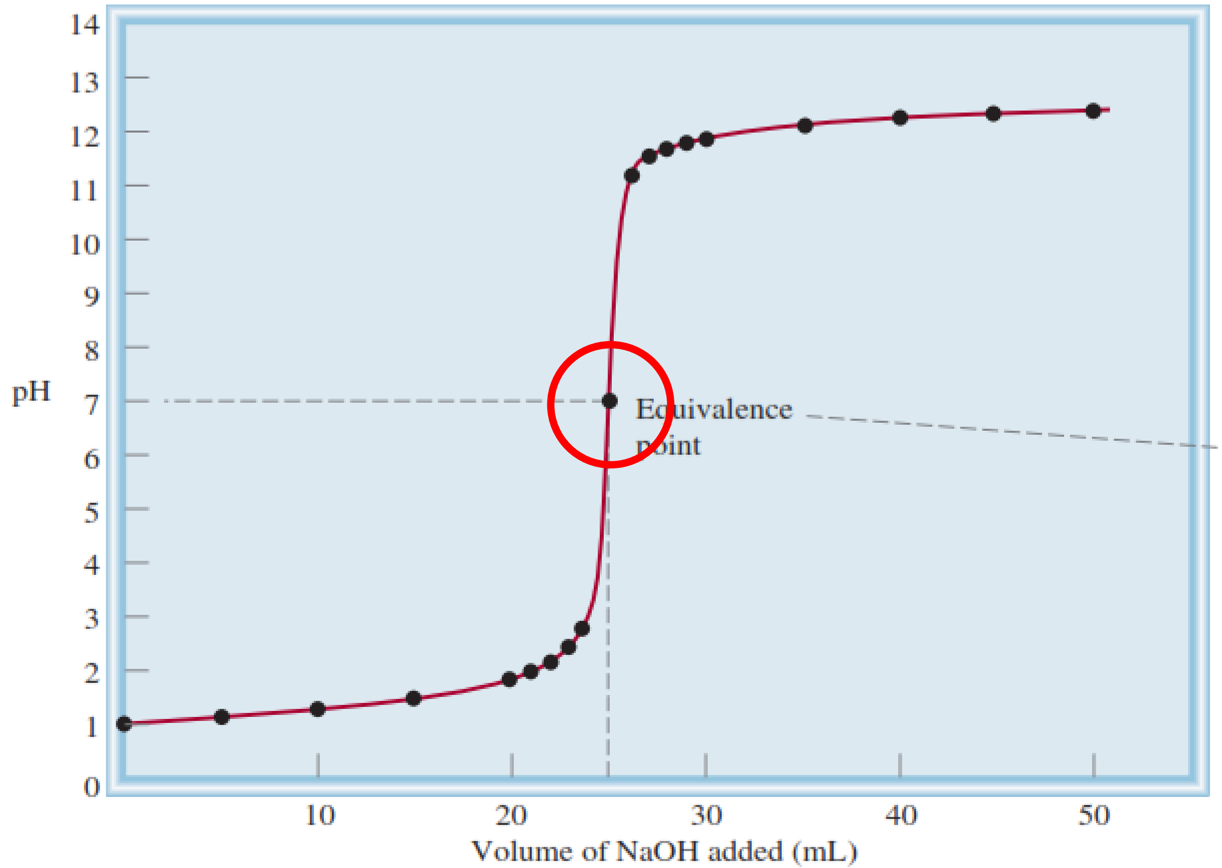

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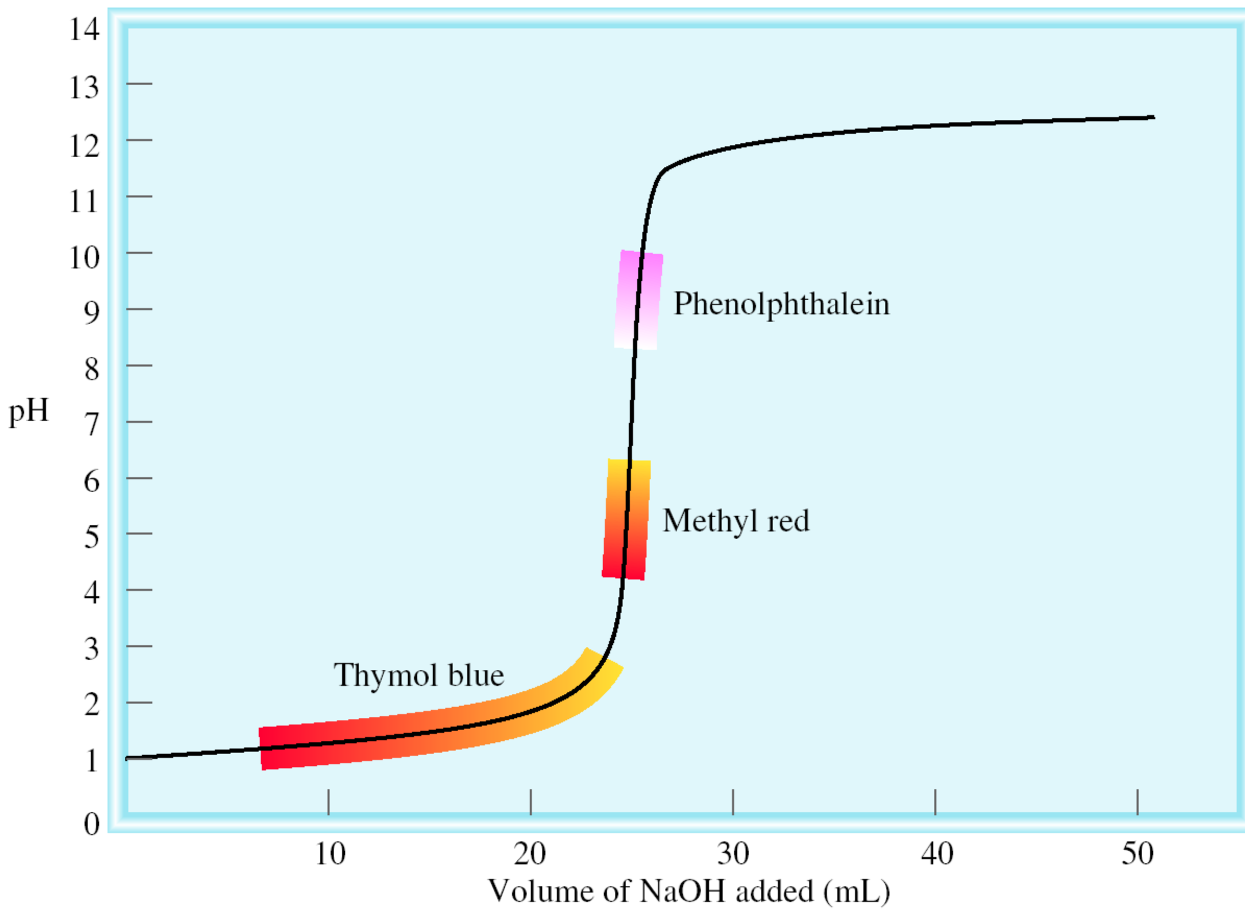
The titration curve of a strong acid with a strong base.



Because the regions over which the indicators methyl red and phenolphthalein change color along the steep portion of the curve, they can be used to monitor the equivalence point of the titration.

Thymol blue cannot be used for the same purpose because the color change does not match the steep portion of the titration curve

Example 1: What volume of a 1.420 *M* NaOH solution is required to titrate 25.00 mL of a 4.50 *M* H_2SO_4 solution?

WRITE THE CHEMICAL EQUATION!

$$H_2SO_4 + 2NaOH \longrightarrow 2H_2O + Na_2SO_4$$

$$\frac{mol_{H_2SO_4}}{mol_{NaOH}} = \frac{1}{2}$$

$$2 \times mol_{H_2SO_4} = mol_{NaOH}$$

$$2 \times MV_{H_2SO_4} = MV_{NaOH}$$

Answer is 158 mL

Acid-Base Titration Calculation

Student can be calculate using following formulas

$$mol = \frac{MV_{mL}}{1000}$$

$$aM_1V_1 = bM_2V_2$$

$$M_{all}V_{all} = aM_1V_1 + bM_2V_2 + cM_3V_3 + \ldots$$

M = molar concentration

V = volume of solution (mL)

a, b, c = number of $[H^+]$ or $[OH^-]$ that is ionized.

Example 2: In a titration experiment, 12.5 mL of 0.500 *M* H_2SO_4 neutralize 50.0 mL of NaOH solution. What is the concentration of the NaOH solution?

$$H_2SO_4 + 2NaOH \longrightarrow 2H_2O + Na_2SO_4$$

ตัวอย่าง 3

เมื่อนำสารละลาย H_2SO_4 เข้มข้น 0.5 mol/dm^3 ปริมาตร 20 cm^3 มาไทเทรตด้วยสารละลาย NaOH เข้มข้น X mol/dm^3 โดยใช้ฟีนอล์ฟทาลีนเป็นอินดิเคเตอร์ พบว่า เมื่อใช้สารละลาย NaOH ไป 10 cm^3 อินดิเคเตอร์เปลี่ยนเป็นสีชมพูอ่อน ข้อใดเป็นค่าของ X (**7 วิชา มค. 57**)

1. 0.33
2. 0.67
3. 0.50
4. 1.0
5. 2.0

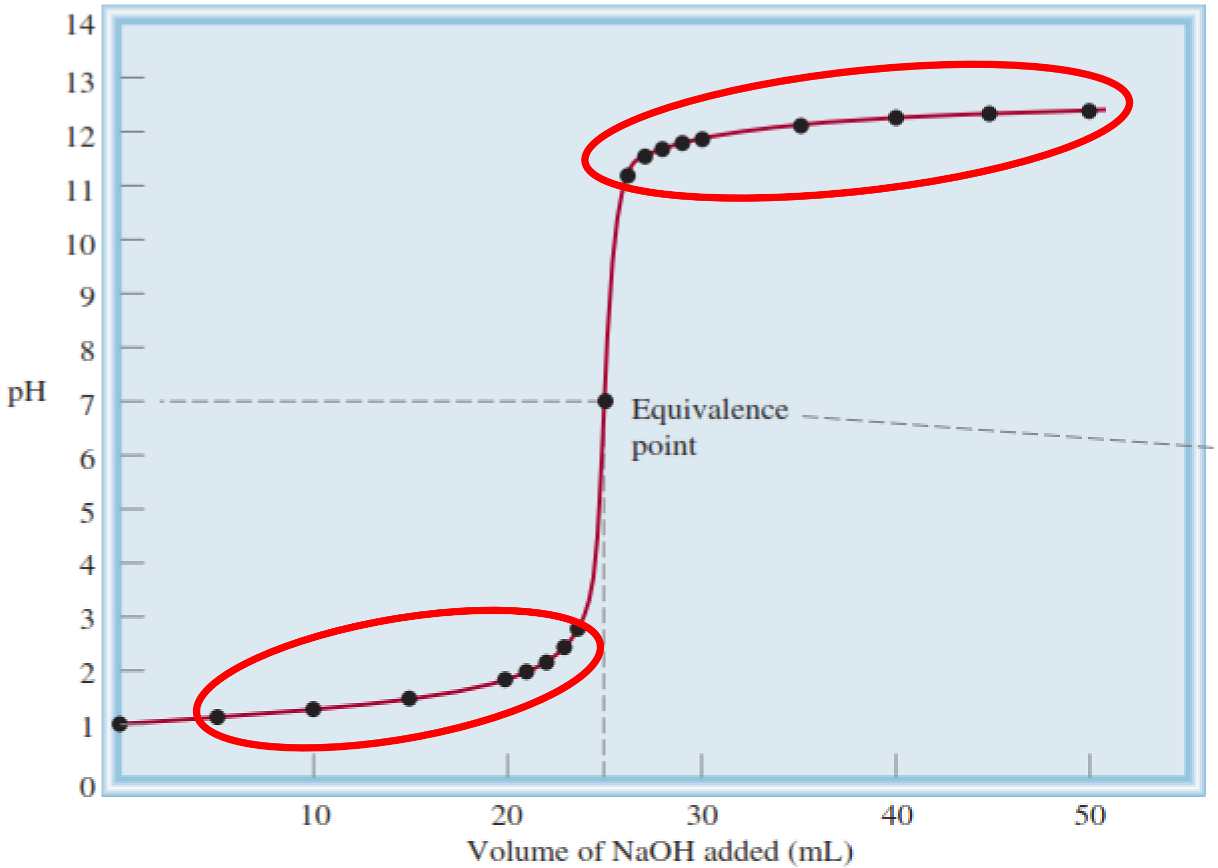
Example 4: A 0.2688-g sample of a monoprotic acid neutralizes 16.4 mL of 0.08133 *M KOH solution. Calculate the* molecular weight of this acid.

Strong Acid–Strong Base Titrations

Strong Acid–Strong Base Titrations

The reaction between a strong acid (say, HCl) and a strong base (say, NaOH) can be represented by

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



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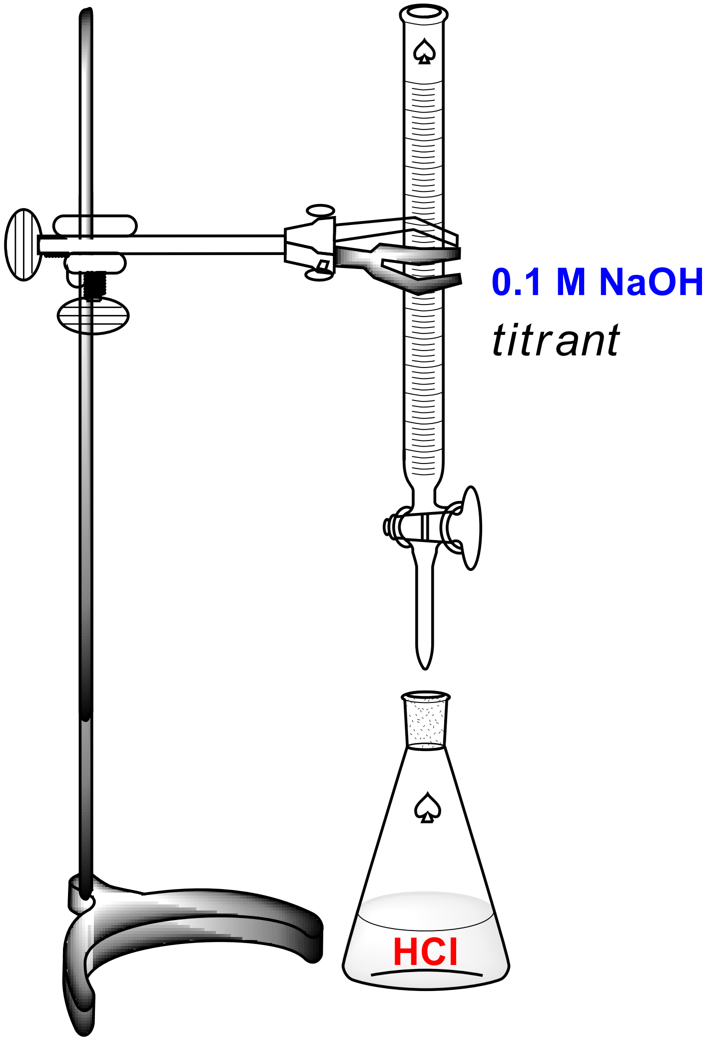

OH^- OH^- OH^- OH^- OH^-

H^+ H^+ H^+ H^+ H^+

Calculation of pH at every stage of Titrations

It is possible to calculate the pH of the solution at every stage of titration. Here are three sample calculations.

<u>Example 1:</u> 10 mL of 0.1 M NaOH solution are titrated with 25 mL of 0.1 M HCl solution. The total volume of the solution is 35 mL. What is the the pH of this solution?

	$NaOH(aq)$	+ $HCl(aq)$	$\longrightarrow$ $NaCl(aq)$	+ $H_2O(l)$
initial (mol)	0.001 mol	0.0025 mol	0	0
change (mol)	0.001 mol	0.001 mol	0.001 mol	
Final (mol)	0	0.0015 mol	0.001 mol	

Therefore, concentration of [HCl] can be find by $mol = \frac{MV_{mL}}{1000}$

Calculation of pH at every stage of Titrations

Example 1: 10 mL of 0.1 M NaOH solution are titrated with 25 mL of 0.1 M HCl solution. The total volume of the solution is 35 mL. What is the the pH of this solution?

	$NaOH(aq)$	+ $HCl(aq)$	$\longrightarrow$ $NaCl(aq)$	+ $H_2O(l)$
initial (mol)	0.001 mol	0.0025 mol	0	0
change (mol)	0.001 mol	0.001 mol	0.001 mol	
Final (mol)	0	0.0015 mol	0.001 mol	

[HCl] = 0.043 M

HCl is strong acid, $[H^+] = 0.043$ M

$$pH = -\log(0.043) = 1.37$$

Example 2: *Calculate the pH of the solution after the addition of 35.0 mL of 0.100 M NaOH to 25.0 mL of 0.100 M HCl. The* total volume of the solution is now 60.0 mL.

Calculation of pH at every stage of Titrations

Student can be calculate using following formulas

$$[H^+]or[OH^-]_{remained} = \frac{aM_1V_1 - bM_2V_2}{V_1 + V_2}$$

M = molar concentration (mol/L)
V = volume of solution (mL)
a, b, = number of $[H^+]$ or $[OH^-]$ that is ionized.

Note that, this equation can only be applied when strong acid is mix with strong base

ตัวอย่างที่ 3

เมื่อนำสารละลาย HCl เข้มข้น 0.2 mol/dm^3 ปริมาตร 10 cm^3 ผสมกับสารละลาย NaOH เข้มข้น 0.1 mol/dm^3 ปริมาตร 30 cm^3 สารละลายที่ได้มีpH เท่าใด

1. 1.6
2. 3.0
3. 7.0
4. 11.0
5. 12.4

R e f e r e n c e s

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