# Acids, Bases, and Salts

Unit 8

## Acid-Base Theory

Electrolyte: a substance which, when dissolved in water, forms a solution capable of conducting an electric current

Electrolytes must:

- dissolve in water (check Table F for solubility or insolubility)
- form charged particles to carry current

Electrolytes include:

- ionic compounds that are soluble

  $NaCl_{(s)} \rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$

- many polar molecules (HCl, HBr, $H_2SO_4$)

  $HCl_{(g)} \rightarrow H^{+}_{(aq)} + Cl^{-}_{(aq)}$

Two commonly used acid-base theories:

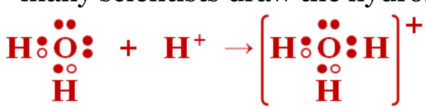
Arrhenius Theory:

- acids: electrolytes that produce $H^{+}_{(aq)}$ ions as the only positive ions in water solutions

  Example:

  $H_2CO_{3\,(aq)} \rightleftharpoons H^{+}_{(aq)} + HCO_3^{-}{}_{(aq)}$

  To show the bonding between the proton, $H^+$ ion, and the surrounding water molecules, many scientists draw the hydronium ion, $H_3O^+$ ion



- bases: electrolytes that produce $OH^{-}_{(aq)}$ ions as the only negative ions in water solutions

  Example:

  $KOH_{(s)} \rightleftharpoons K^{+}_{(aq)} + OH^{-}_{(aq)}$

Brønsted-Lowry Theory:

- acids: any $H^+$ donor (no $H_2O$ is required)
- bases: any $H^+$ acceptor (no $OH^-$ is required)

  Example:

  $NH_3 + NH_3 \rightleftharpoons NH_4^{+} + NH_2^{-}$

  Notice that one $NH_3$ donates an $H^+$ (the acid) and the other $NH_3$ accepts an $H^+$ (the base) therefore $NH_3$ is both an acid and a base according to the Brønsted-Lowry theory

  Substances that act as both an acid and a base are *amphoteric* or *amphiprotic*

  The NYS Regents always calls the Brønsted-Lowry theory *one alternative theory*

Description of acids:

- are electrolytes
- react with active metals to form $H_{2\,(g)}$: active metals are those above $H_2$ on Table J
- cause indicators to change color: indicators and color changes are on Table M
- react with bases – a neutralization reaction

  $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(\ell)}$

- react with $CO_3^{2-}$ to form $CO_{2\,(g)}$
- taste sour like lemon juice or vinegar *Note*: never taste anything in the lab

Table K: a list of common acids

Description of bases:

- are electrolytes
- cause indicators to change color: indicators and color changes are on Table M
- react with acids – a neutralization reaction

  $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(\ell)}$

- the only common weak base is ammonia, $NH_{3\,(aq)}$, which is sometimes written $NH_4OH$
- Group 1 metals form strong bases (NaOH and KOH, e.g.)
- feel slippery and taste bitter *Note*: never taste anything in the lab

Table L: a list of common bases

Strong—Weak Acid—Base Theory

Do not confuse the terms *strong* and *concentrated*:

- strong means the acid or base electrolyte ionizes 100%

  $HCl_{(g)} \rightarrow H^{+}_{(aq)} + Cl^{-}_{(aq)}$

- weak means the acid or base electrolyte ionizes less than 100% (usually less than 2%)

  $HC_2H_3O_{2\,(aq)} \rightleftharpoons H^{+}_{(aq)} + C_2H_3O_2^{-}{}_{(aq)}$

- concentrated means a lot of solute
- dilute means only a small amount of solute

It is possible to have dilute strong acids and bases and concentrated weak acids and bases

Acid-Base Neutralization

Possible reactions include:

- strong acid vs. strong base
- strong acid vs. weak base
- weak acid vs. strong base
- weak acid vs. weak base (not in the scope of a Regents class)

Examples of strong acids and bases:

Strong acids from Table K: $HCl_{(aq)}$, $HNO_{3\,(aq)}$, and $H_2SO_{4\,(aq)}$ ($HSO_4^{1-}$ is weak)

$H_2SO_{4\,(aq)} \rightarrow H^{+}_{(aq)} + HSO_4^{1-}{}_{(aq)}$

Strong bases from Table L: $NaOH_{(aq)}$, $KOH_{(aq)}$, and $Ca(OH)_{2\,(aq)}$

$Ca(OH)_{2\,(aq)} \rightarrow Ca^{2+}_{(aq)} + 2\ OH^{-}_{(aq)}$

Weak acids from Table L: $HNO_{2\,(aq)}$, $H_2SO_{3\,(aq)}$, $H_3PO_{4\,(aq)}$, $H_2CO_{3\,(aq)}$, and $CH_3COOH_{(aq)}$

$CH_3COOH_{(aq)} \rightleftharpoons H^{+}_{(aq)} + CH_3COO^{-}_{(aq)}$

Weak base from Table L: $NH_{3\,(aq)}$

$NH_{3\,(aq)} \rightarrow NH_4^{+}{}_{(aq)} + OH^{-}_{(aq)}$

Neutralization reactions always produce salt plus water

Strong acid – strong base

$2\ HCl_{(aq)} + Ca(OH)_{2\,(aq)} \rightarrow CaCl_{2\,(aq)} + 2\ H_2O_{(\ell)}$

acid + base → salt + water

Strong acid – weak base

$HCl_{(aq)} + NH_4OH_{(aq)} \rightarrow NH_4Cl_{(aq)} + 2\ H_2O_{(\ell)}$

acid + weak base → acidic salt + water

Weak acid – strong base

$HC_2H_3O_{2\,(aq)} + KOH_{(aq)} \rightarrow NH_4Cl_{(aq)} + 2\ H_2O_{(\ell)}$

weak acid + base → basic salt + water

Indicators and Endpoints

Titration: using a measured volume of a known concentration of one substance and adding a measured volume (typically, by using a buret, old school spelling is burette) of a second substance to determine the concentration of the second substance

Indicator: a substance that changes color when the equivalence point of a titration is reached

Equivalence point: occurs when the amount of acid is equal to the amount of base in a titration

The endpoint of a neutralization reaction occurs at the equivalence point, or when the amount of acid, $H^{+}_{(aq)}$, is equal to the amount of base, $OH^{-}_{(aq)}$, see the balanced equation below

$H^{+}_{(aq)} + OH^{-}_{(aq)} \rightarrow H_2O_{(\ell)}$

The endpoint of a strong acid – strong base titration (which produces a neutral salt) is pH 7

The indicator changes color at the endpoint which should occur at the equivalence point

For this to happen, we must choose the indicator properly

Table M shows a few commonly used indicators:

**Table M**
**Common Acid – Base Indicators**

| Indicator | Approximate pH Range for Color Change | Color Change |
|---|---|---|
| methyl orange | 3.1 – 4.4 | red to yellow |
| bromthymol blue | 6.0 – 7.6 | yellow to blue |
| phenolphthalein | 8 – 9 | colorless to pink |
| litmus | 4.5 – 8.3 | red to blue |
| bromcresol green | 3.8 – 5.4 | yellow to blue |
| thymol blue | 8.0 – 9.6 | yellow to blue |

Source: *The Merck Index*, 14th ed., 2006, Merck Publishing Group

**Reference Tables for Physical Settings / Chemistry – 2011 Edition**

According to Table M, litmus indicator will be:

- red when the pH is less than 4.5 (acidic)
- blue when the pH is greater than 8.3 (acidic)
- between 4.5 and 8.3, litmus will be purple

Notice that the colors in the color change column correspond to the pH range column

Note that the color change and change range for litmus is

**red** **blue**

**← 4.5** **purple** **8.3 →**

All the indicators on Table M will be the *left* color to the *left* of the pH range, they will be the *right* color to the *right* of the pH range, and the mix of the two colors *between* the listed range

The strong acid – strong base equivalence point is pH = 7.0, so what is the best indicator?

The only indicator color range that brackets pH = 7 is bromthymol blue at 6.0 – 7.6

If titrating a base into a beaker with an acidic solution, what colors will appear?

- yellow when the pH is less than 6.0 (acidic)
- green at the endpoint when the pH is between 6.0 and 7.6
- blue when the pH is more than 7.6 (basic)

Will the endpoint of a strong acid – weak base titration be acidic, neutral, or basic?
The salt of a strong acid and a weak base will be acidic

Name *one* indicator that would be appropriate for a strong acid – weak base titration.
Either of the following would be appropriate: methyl orange or bromcresol green

Will the endpoint of a weak acid – strong base titration be acidic, neutral, or basic?
The salt of a weak acid and a strong base will be basic

Name *one* indicator that would be appropriate for a weak acid – strong base titration.
Either of the following would be appropriate: phenolphthalein or thymol blue

Balancing Acid – Base Reactions

Acid – Base Reactions: (sometimes called neutralization reactions)

The reaction of an Arrhenius acid and an Arrhenius base will always be:
acid + base → salt + water

Recognizing an Arrhenius acid:

- produces $H^+$ as the only positive ion in water solution
- the formula will begin with $H^+$
- the formula will end with a negative ion (a nonmetal or a polyatomic ion)

Recognizing an Arrhenius base:

- produces $OH^-$ as the only positive ion in water solution
- the formula will end with $OH^-$ ($NH_3$ is one exception)
- the formula will begin with a metal ion (or $NH_4^+$)

Example: write the balanced equation for the reaction of calcium hydroxide with hydrochloric acid

Change words to formulas:
$Ca(OH)_{2\,(aq)} + HCl_{(aq)} \rightarrow$

Check reaction type to determine what products will form:
Double replacement (AX + BY → AY + BX)
$Ca(OH)_{2\,(aq)} + HCl_{(aq)} \rightarrow CaCl + HOH$

Make sure product formulas are correct
Ca is a metal and will form $Ca^{2+}$ ions, Cl is $Cl^-$ so easy criss-cross gives $CaCl_2$
Table F says $CaCl_2$ is soluble, so $CaCl_{2\,(aq)}$
HOH is just another way to write water, $H_2O_{(\ell)}$

$Ca(OH)_{2\,(aq)} + HCl_{(aq)} \rightarrow CaCl_{2\,(aq)} + H_2O_{(\ell)}$

Use conservation of atoms (mass) to balance the equation
$Ca(OH)_{2\,(aq)} + 2\ HCl_{(aq)} \rightarrow CaCl_{2\,(aq)} + H_2O_{(\ell)}$ use R → L, 2 Cl as products ($Cl_2$) so 2 HCl
$Ca(OH)_{2\,(aq)} + 2\ HCl_{(aq)} \rightarrow CaCl_{2\,(aq)} + H_2O_{(\ell)}$ use L → R, 2 O in $Ca(OH)_2$, so need 2 $H_2O$
This equation atom checks, so it is balanced

Example: here is a tough one – try sulfuric acid and iron(III) hydroxide
Check your answer online

Metals with Acids:

Example: write the balanced equation for the reaction of aluminum with hydrochloric acid

Change words to formulas:

$Al_{(s)} + HCl_{(aq)} \rightarrow$

Check Table J to see if Al can replace $H^+$:

Al is above $H_2$ on Table J, so the reaction will occur

Check reaction type to determine what products will form:

Single replacement ($A + BX \rightarrow AX + B$ or $Y + AX \rightarrow AY + X$)

Will Al replace $H^+$ or $Cl^-$?

Al is a metal and will form $Al^{3+}$ ions, so it would replace the $H^+$

$Al_{(s)} + HCl_{(aq)} \rightarrow AlCl + H$

Make sure product formulas are correct

Al is a metal and will form $Al^{3+}$ ions, Cl is $Cl^-$ so easy criss-cross gives $AlCl_3$

Table F shows that $AlCl_3$ is soluble in water, so $AlCl_{3\,(aq)}$

H is hydrogen and all the 'gens' are diatomic, so $H_{2\,(g)}$

$Al_{(s)} + HCl_{(aq)} \rightarrow AlCl_{3\,(aq)} + H_{2\,(g)}$

Use conservation of atoms (mass) to balance the equation

The product side has 3 Cl and 2 H, but they come from HCl, so the LCM is 6

$Al_{(s)} + HCl_{(aq)} \rightarrow 2\,AlCl_{3\,(aq)} + 3\,H_{2\,(g)}$ now we have 6 H and 6 Cl as products

$Al_{(s)} + 6\,HCl_{(aq)} \rightarrow 2\,AlCl_{3\,(aq)} + 3\,H_{2\,(g)}$ we used $R \rightarrow L$ to balance the HCl

$2\,Al_{(s)} + HCl_{(aq)} \rightarrow 2\,AlCl_{3\,(aq)} + 3\,H_{2\,(g)}$ Al was the only atom not balanced, so $R \rightarrow L$

Example: write the balanced equation for the reaction of copper with hydrochloric acid

Change words to formulas:

$Cu_{(s)} + HCl_{(aq)} \rightarrow$

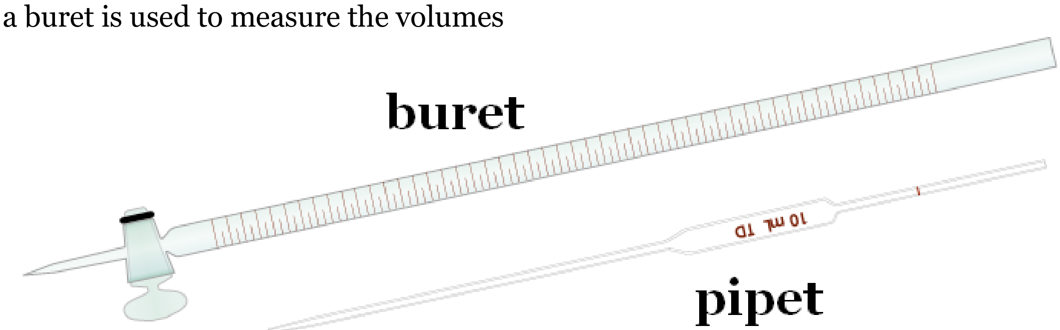
Check Table J to see if Cu can replace $H^+$:

$H_2$ is above Cu on Table J, so *no reaction* will occur

## Titration

Titration: a process in which the volume of a solution of known concentration is used to determine the concentration of a measured volume of another solution

- standard: a solution of known concentration
- neutralization is accomplished with an indicator
- a buret is used to measure the volumes




Titration equation: $M_A V_A = M_B V_B$

Where: $M_A$ = molarity of the $H^+$ $M_B$ = molarity of the $OH^-$

$V_A$ = volume of the $H^+$ $V_B$ = volume of the $OH^-$

Explaining the titration equation: $M_A V_A = M_B V_B$

- endpoint, by definition, means moles of $H^+$ = moles of $OH^-$
- $MV$ (molarity × volume) = $\frac{\text{mol}}{\text{L}} \times \text{L} = \text{mol}$

  therefore, $M_A V_A$ = moles acid = mole sof $H^+$ and $M_B V_B$ = moles of $OH^-$
- if a scientist measures any three variables, the fourth can be calculated
- volume (V) can be in any units, but the units of $V_A$ and $V_B$ *must* match
- units on molarity (M) are always mol/L

Example 1: What volume of 0.1000 M NaOH is required to neutralize 22.37 mL of 0.1095 M HCl?

$V_B$ = ?  $M_A V_A = M_B V_B$

$M_B$ = 0.1000 M  $V_B = M_A V_A / M_B$

$V_A$ = 22.37 mL  = ( 0.1095 M ) ( 22.37 mL ) / ( 0.1000 M )

$M_A$ = 0.1095 M  = 24.50 mL NaOH

Example 2: Find the concentration of a solution of NaOH for which a 25.00 mL sample requires 37.52 mL of 0.1108 M HCl

$M_B$ = ?  $M_A V_A = M_B V_B$

$V_B$ = 25.00 mL  $M_B = M_A V_A / V_B$

$V_A$ = 37.52 mL  = ( 0.1108 M ) ( 37.52 mL ) / ( 25.00 mL )

$M_A$ = 0.1108 M  = 0.1663 M NaOH

pH

pH is a logarithmic scale used to express acidity or alkalinity of a solution

Water is amphoteric (or amphiprotic), which means it produces both $H^+_{(aq)}$, making water an acid, and $OH^-_{(aq)}$, making water a base at the same time (see equation below)

$$H_2O_{(\ell)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$$

where:

$[H^+] = [OH^-] = 1 \times 10^{-7}$

$[H^+] = [OH^-]$ is read, “The molar concentration of the hydrogen ion is equal to the molar concentration of the hydroxide ion”

- $[H^+]$ must equal $[OH^-]$ because in the equation, for every $H^+$ there is one $OH^-$
- water is neutral because $[H^+] = [OH^-]$
- an increase in $[H^+]$ is acidic and $[H^+] = 1 \times 10^{-6}$ or higher

  note: $(1 \times 10^{-7})(10) = 1 \times 10^{-6}$
- a decrease in $[H^+]$ is basic and $[H^+] = 1 \times 10^{-8}$ or lower

  note: $(1 \times 10^{-7}) \div (10) = 1 \times 10^{-8}$

A logarithmic scale is used to avoid numbers like $1 \times 10^{-8}$ or $1 \times 10^{-11}$

$pH \equiv -\log[H^+]$  or, by definition, p means to take the negative $\log_{10}$

Example: Find the pH of a solution with $[H^+] = 1 \times 10^{-6}$ M

$pH = -\log[H^+] = -\log(1 \times 10^{-6}) = -(-6) = 6$

Example: Find the pH of a solution with $[H^+]$ 100 times more acid than $1 \times 10^{-6}$ M

$100\ (1 \times 10^{-6}\text{ M}) = 1 \times 10^{-4}$ M

$pH = -\log(1 \times 10^{-4}) = -(-4) = 4$

A pH 100 times more acid than a pH of 6 is a pH of 4 and neutral is pH = 7

Titrating a strong base with a strong acid

What is the pH of a neutral solution?

pH = 7

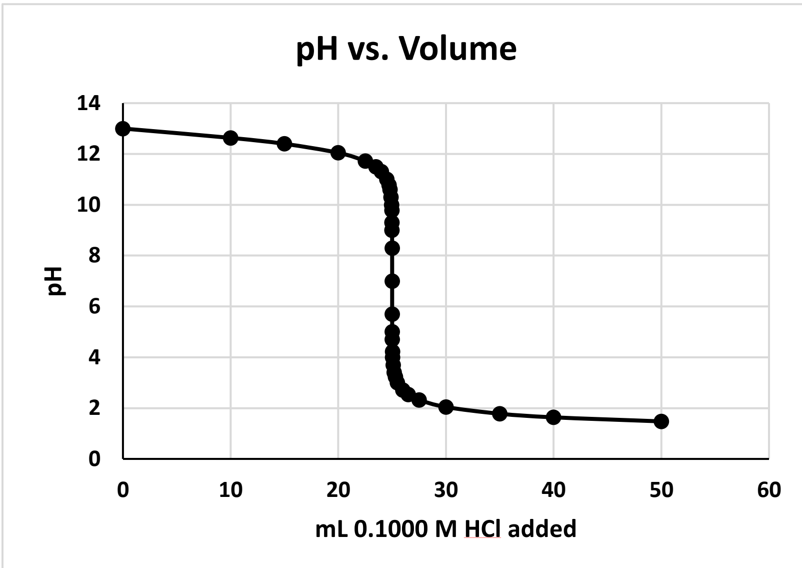
What is the pH of a basic solution?

pH > 7

What is the pH of an acidic solution?

pH < 7

A strong acid – strong base titration will have a neutral endpoint (see graph below)




For strong acids and bases, calculate pH from the molar concentration

What is the pH of a 0.00001 M HCl solution?

$[H^+] = 0.00001\ M = 1\times10^{-5}\ M$

$pH = -\log(1\times10^{-5}) = -(-5) = 5$

What is the pH of a 0.00001 M NaOH solution?

$[OH^-] = 0.00001\ M = 1\times10^{-5}\ M$

$pOH = -\log(1\times10^{-5}) = -(-5) = 5$

$pH + pOH = 14$

$pH = 14 - pOH = 14 - 5 = 9$

For weak acids and bases, the pH (or pOH) will be higher than the molar concentration

What is the pH of a 0.01 M $C_2H_3OOH$ solution?

$[HAc] = 0.01\ M = 1\times10^{-2}\ M$, but only some of the HAc ionizes, so

$[H^+] = 4.24\times10^{-4}\ M$

$pH = -\log(4.24\times10^{-4}) = 3.4$

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