SOLUTIONS -ACIDS AND BASES

TOPICS COVERED :

- ✓ARRHENIUS ACID-BASE THEORY
- ✓BRONSTED-LOWRY ACID-BASE THEORY
- ✓CONJUGATE ACIDS AND BASES
- ✓TYPES OF ACIDS- MONO-DI AND TRIPROTIC ACIDS
- ✓STRENGTHS OF ACIDS AND BASES
- ✓REACTIONS OF ACIDS, BASES, SELF IONIZATION OF WATER
- ✓THE pH SCALE, CALCULATION OF pH OF SOLUTION
- ✓BUFFERS
- ✓ACID-BASE TITRATIONS

ACID PROPERTIES

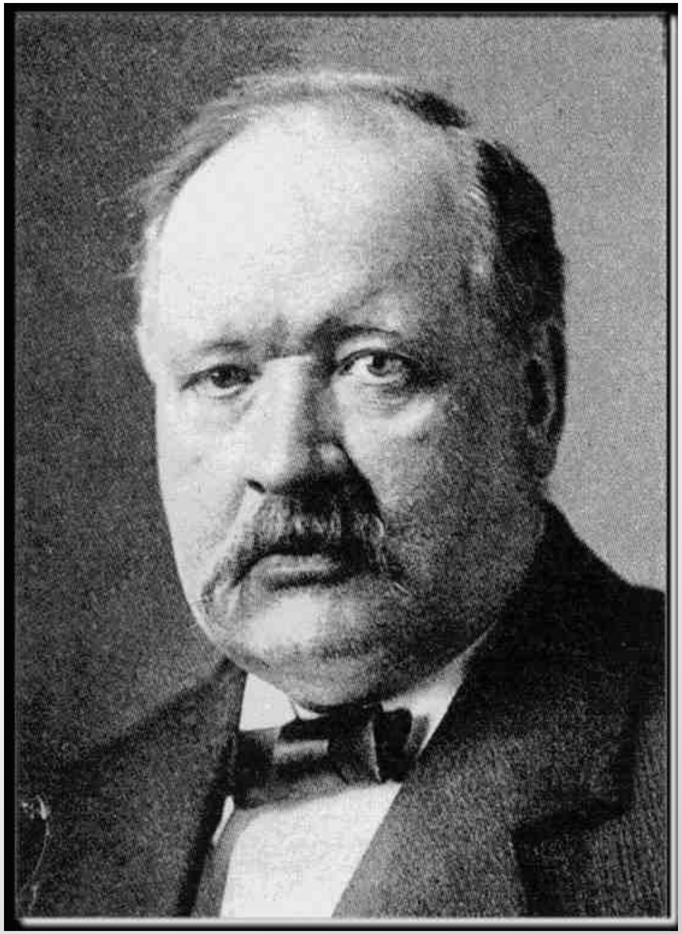
- SOUR TASTE
- CHANGE THE COLOR OF LITMUS FROM BLUE TO RED.
- REACT WITH
  - METALS SUCH AS ZINC AND MAGNESIUM TO PRODUCE HYDROGEN GAS
  - HYDROXIDE BAS[illegible] COMPOUND (A SALT)
  - CARBONATES TO PRODUCE CARBON DIOXIDE.

These properties are due to the release of hydrogen ions, H^+, in water solution.

BASE PROPERTIES

- BITTER OR CAUSTIC TASTE (CHOCOLATE)
- THE ABILITY TO CHANGE LITMUS RED TO BLUE
- DISSOLVE FATS (DRAIN CLEANER)
- HAVE A SLIPPERY, SOAPY FEELING (BUT SOAP IS A FATTY ACID SALT – NOT A BASE).
- THE ABILITY TO REACT WITH ACIDS

- SVANTE AUGUST ARRHENIUS WAS A SWEDISH SCIENTIST WHO LIVED FROM 1859-1927.
- IN 1884 HE ADVANCED A THEORY OF ACIDS AND BASES.



Svante August Arrhenius

Arrhenius Acid-Base Theory

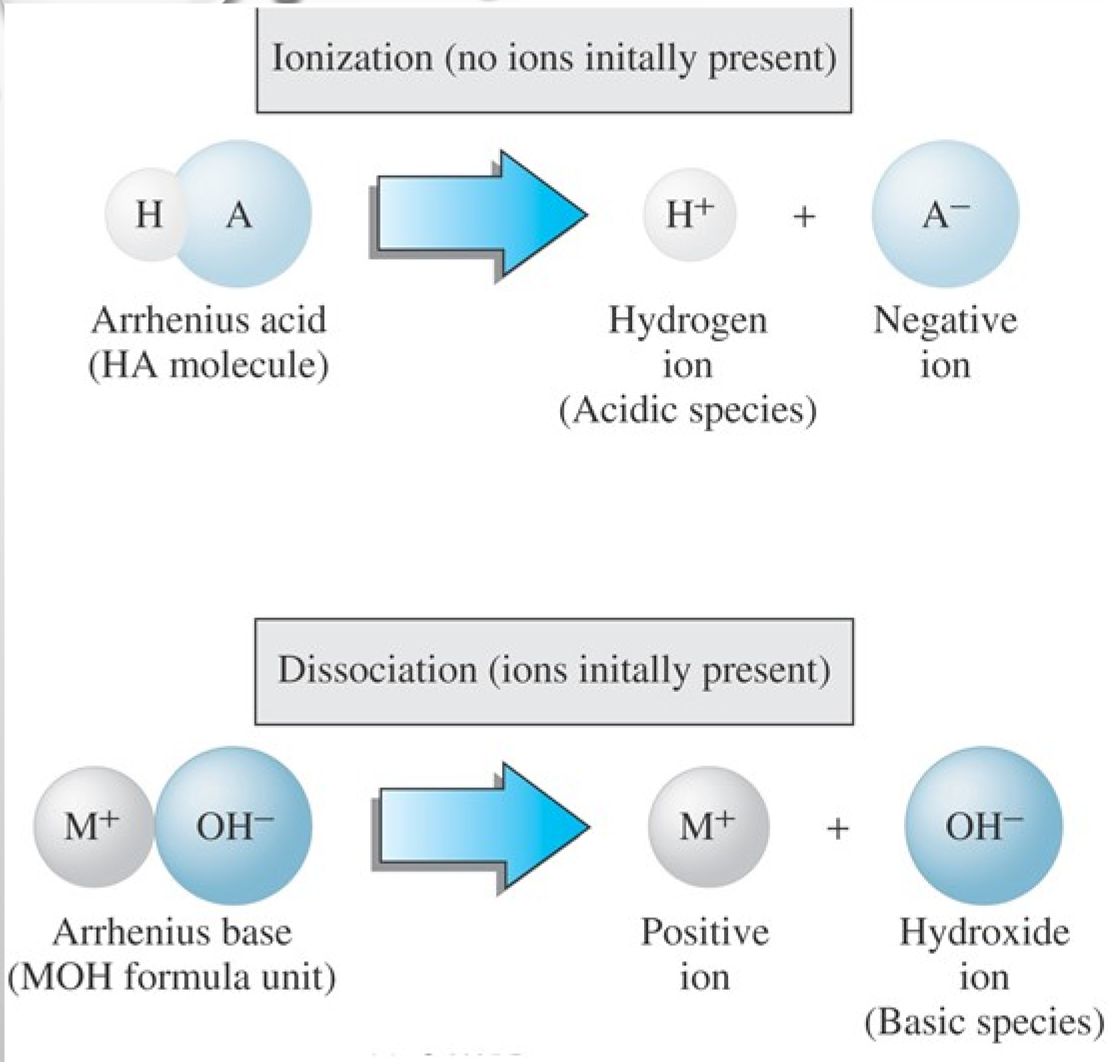
AN **ARRHENIUS ACID** IS A HYDROGEN-CONTAINING SUBSTANCE THAT UNDERGOES IONIZATION TO PRODUCE HYDROGEN IONS IN AQUEOUS SOLUTION.

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

acid

AN **ARRHENIUS BASE** IS A HYDROXIDE-CONTAINING SUBSTANCE THAT DISSOCIATES TO PRODUCE HYDROXIDE IONS IN AQUEOUS SOLUTION.

$$\underset{\text{base}}{MOH} \rightarrow M^{+}(aq) + OH^{-}(aq)$$



The difference between the aqueous solution process of **ionization** (Arrhenius acids) and **dissociation** (Arrhenius base). **Ionization** is the production of ions from a molecular compound that has been dissolved in a solvent. **Dissociation** is the production of ions from an ionic compound that has been dissolved in solvent.

Bronsted and Lowry Acid-Base Theory

J.N. BRONSTED (1897-1947) WAS A DANISH CHEMIST AND T. M. LOWRY (1847-1936) WAS AN ENGLISH CHEMIST.

- In 1923 they advanced their theory of acids and bases.

A Bronsted-Lowry acid is a proton (H^+) donor.

A Bronsted-Lowry base is a proton (H^+) acceptor.

- **You cannot have a B-L acid without a B-L base. The proton donor (acid) needs the proton acceptor (base).**
- **All B-L acids are also Arrhenius acids but not all B-L bases are Arrhenius bases**
- **The B-L acidic species in *aqueous* solution is not hydrogen ion, H+ but hydronium ion, H_3O+**

B-L reactions can occur in the gas phase – no water required depending on the B-L acid-base involved.

Bronsted-Lowry Acid

ry Base

$$HCl + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

Bronsted-Lowry Acid

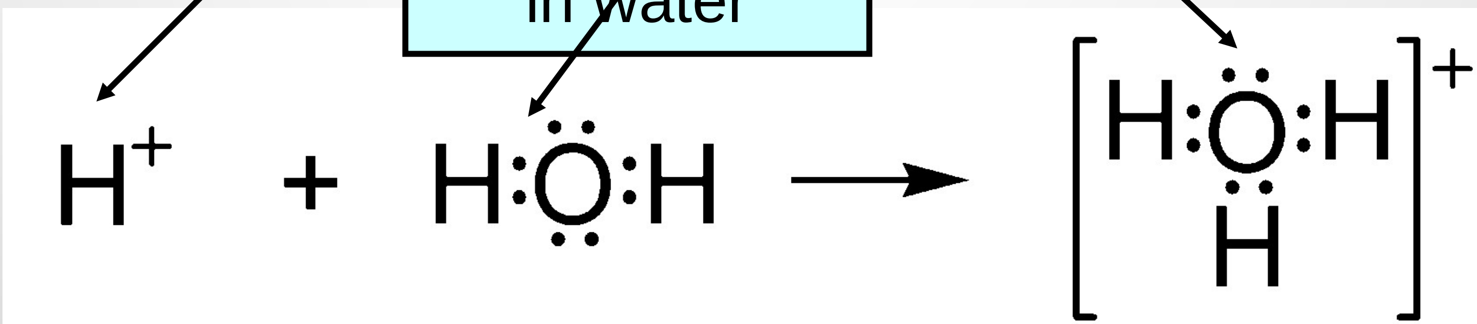
ry Base

$$HCl + NH_3(l) \rightarrow NH_4^+(aq) + Cl^-(aq)$$

a hydronium ion is formed

ydrogen ion
ines with water

in water



When an acid donates a proton the species that remains is called the **conjugate base**.

$$\underset{\text{acid}}{HCl} \xrightarrow{-\,H^+} \underset{\text{conjugate base}}{Cl^-}$$

When a base accepts a proton, the species it becomes is called the **conjugate acid**.

$$H_2O \xrightarrow{+\ H^+} H_3O^+$$

base → conjugate acid

Conjugate pair

$$HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$$

Acid Base Acid Base

Conjugate pair

In a B-L acid-base pair:

-the acid has one more acidic H atom and one fewer negative charge than the base

-the base always has one fewer acidic H atom and one more negative charge than the acid.

HF/F-

The two acids and two bases involved in a Brønsted-Lowry equilibrium situation can be grouped into two <u>conjugate acid-base</u> pairs.

$$HCl(g) + H_2O(l) \rightarrow Cl^{-}(aq) + H_3O^{+}(aq)$$

acid base Conjugate base Conjugate acid

Bronsted-Lowry acid – base pair

What is an *amphiprotic* substance?

A substance that can be either a B-L acid or B-L base.

Example: H_2O

$$HNO_{3(l))} + H_2O_{(l)} \longleftrightarrow H_3O^+_{(aq)} + NO_3^-{}_{(aq)}$$

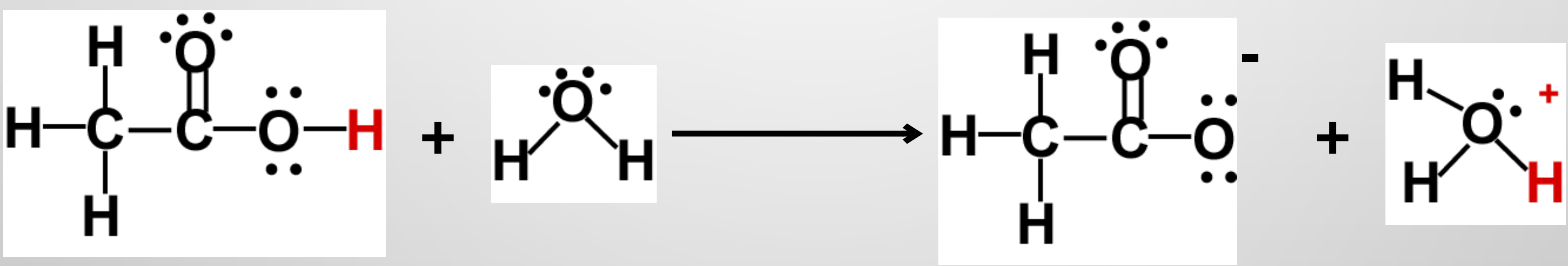
**bas
e**

$$NH_{3(l)} + H_2O_{(l)} \longleftrightarrow NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$$

Acid

ACIDIC HYDROGEN: the hydrogen atom in an acid molecule that is transferred as a proton in an acid-base reaction .

- The acidic hydrogen is a formula is always written at the beginning and separate from non-acidic hydrogens. example: $HC_2H_3O_2$
- In an oxyacid the acidic hydrogen is bonded to oxygen.



An acid can have more than one acidic hydrogen.

Types of Acid-

- **Monoprotic acid** only transfers 1 H^+(Ex: HNO_3, HCl)
- **Polyprotic acid**: acids that can transfer more than one proton in an acid/base reaction.
  - **Diprotic acid**: can transfer 2 H^+. (Ex: $\mathbf{H_2SO_4}$)
  - **Triprotic acid**: can transfer 3 H^+. (Ex: $\mathbf{H_3PO_4}$)

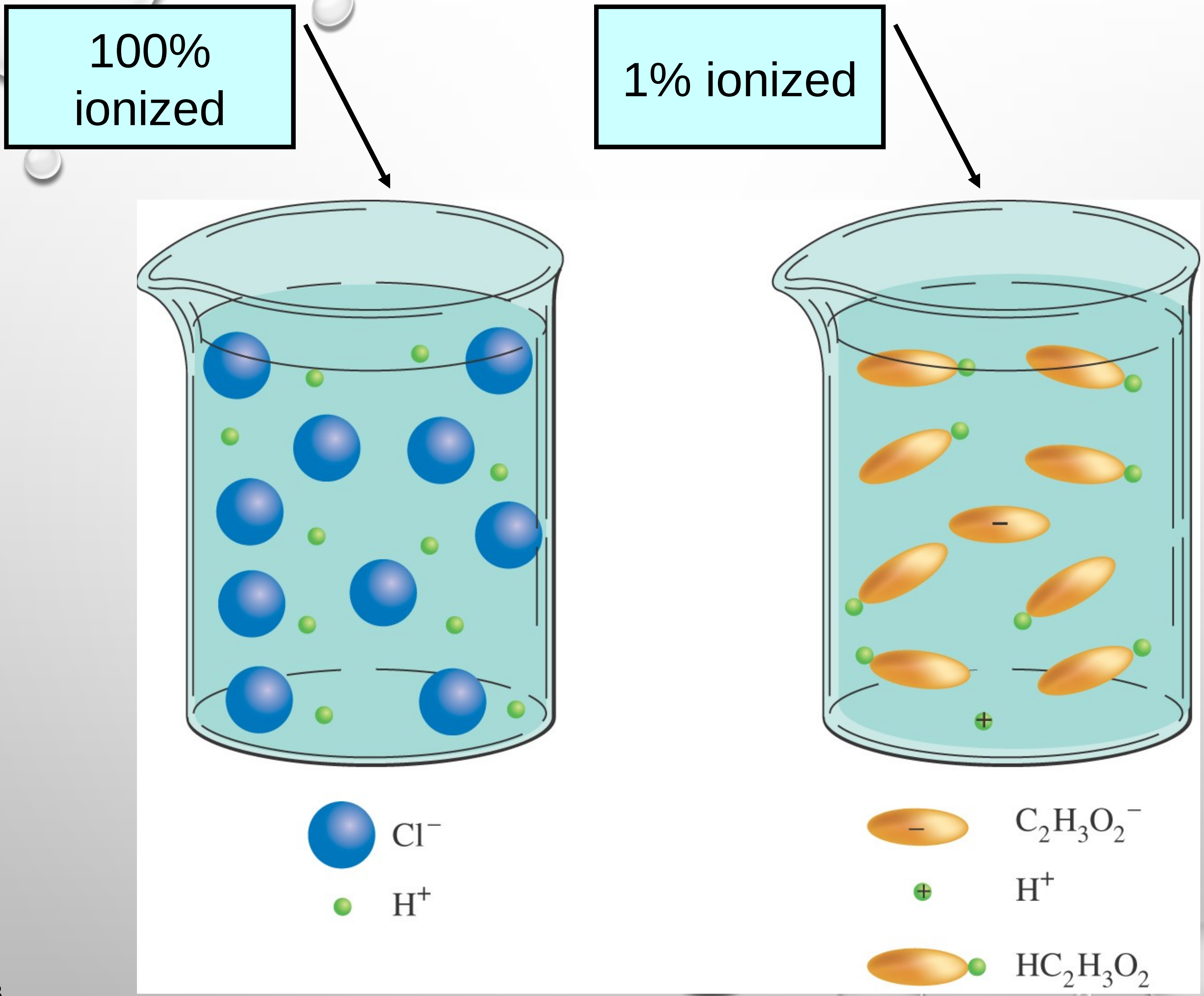
$$H_3PO_4(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + H_2PO_4^-(aq)$$
$$H_2PO_4^-(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + HPO_4^{2-}(aq)$$
$$HPO_4^{2-}(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + PO_4^{3-}(aq)$$

STRENGTHS OF ACIDS AND BASES

SOME IONIC COMPOUNDS (SALTS) DISSOCIATE COMPLETELY IN AQUEOUS SOLUTION. THEY ARE CALLED **STRONG ELECTROLYTES**. AT LEAST ONE COMPONENT OF THE SALT WILL BE FROM A STRONG ACID OR BASE.

A **STRONG ACID** IS AN ACID THAT, IN AQUEOUS SOLUTION TRANSFERS 100% (OR NEARLY 100%) OF ITS ACIDIC PROTONS TO WATER.

A **WEAK ACID** IS AN ACID THAT, IN AQUEOUS SOLUTION, ONLY TRANSFERS A SMALL PERCENTAGE OF ITS ACID PROTONS TO WATER.


100% ionized
1% ionized
Cl^-
H^+
$C_2H_3O_2^-$
H^+
$HC_2H_3O_2$

Both the ionized and un-ionized forms of a weak acid (or any weak electrolyte) are present in aqueous solution.

$$HC_2H_3O_2(l) \leftrightarrow H^+(aq) + C_2H_3O_2^-(aq)$$

un-ionized

ionized

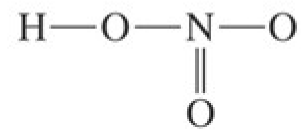
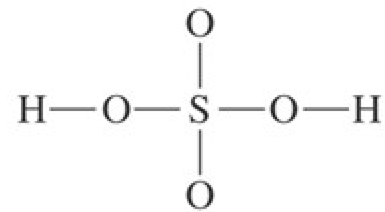
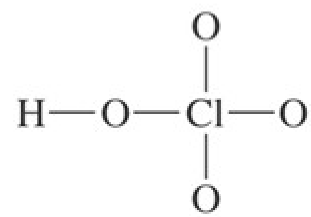
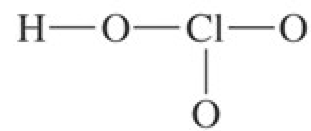
HNO_3, a strong acid, is 100 % ionized.

$$HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$$

HNO_2, a weak acid, is only slightly ionized.

$$HNO_2(l) \Leftrightarrow H^+(aq) + NO_2^-(aq)$$

Commonly Encountered Strong Acids

Name*	Molecular Formula	Molecular Structure
Nitric acid	HNO_3	
Sulfuric acid	H_2SO_4	
Perchloric acid	$HClO_4$	
Chloric acid	$HClO_3$	
Hydrochloric acid	HCl	H—Cl
Hydrobromic acid	HBr	H—Br
Hydroiodic acid	HI	H—I

Note- *Acid nomenclature discussed in lecture on chapter #4

STRONG BASES

Group 1A
Hydroxides

LiOH

NaOH

KOH

RbOH

CsOH

Group 2A
Hydroxides

$Ca(OH)_2$

$Sr(OH)_2$

$Ba(OH)_2$

REACTIONS OF ACIDS

IN AQUEOUS SOLUTION, THE H^+ OR H_3O^+ IONS ARE RESPONSIBLE FOR THE CHARACTERISTIC REACTIONS OF ACIDS.

Acids react with:

- Metals
- Bases
- Carbonates and bicarbonates

- **Reaction with Metals** Acids react with metals that lie above hydrogen in the activity series of elements to produce hydrogen and an ionic compound (salt):

acid + metal → hydrogen + ionic compound

$$2HCl(aq) + Ca(s) \rightarrow H_2(g) + CaCl_2(aq)$$

$$H_2SO_4(aq) + Mg(s) \rightarrow H_2(g) + MgSO_4(aq)$$

- **Reaction with Bases** The reaction of an acid with a base is called a *neutralization* reaction. When it is a hydroxide base, in aqueous solution, the products are a salt and water:

$$\underset{\text{acid}}{HBr(aq)} + \underset{\text{base}}{KOH(aq)} \rightarrow \underset{\text{salt}}{KBr(aq)} + H_2O(l)$$

$$\underset{\text{acid}}{2HNO_3(aq)} + \underset{\text{base}}{Ca(OH)_2(aq)} \rightarrow \underset{\text{salt}}{Ca(NO_3)_2(aq)} + 2H_2O(l)$$

- **Reaction with Carbonates and Bicarbonates**
  Most acids react with carbonates to produce an ionic compound and carbonic acid (double displacement reaction). Carbonic acid formed then decomposes to water and carbon dioxide.

$$2HCl(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$$

acid **carbonate** **salt**

$$H_2SO_4(aq) + MgCO_3(s) \rightarrow MgSO_4(aq) + H_2O(l) + CO_2(g)$$

acid **carbonate** **salt**

$$HCl(aq) + NaHCO_3(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

acid **carbonate** **salt**

Self –Ionization of Water

Even though water is a molecular (covalently-bonded) substance, a very small percentage of water molecules interact with one another to form ions. We refer to this phenomenon as the self-ionization of water.

Water only ionizes slightly.

hydronium ion

hydroxide ion

$$H_2O + H_2O \Leftrightarrow H_3O^+ + OH^-$$

acid base acid base

Water ionization can be expressed more simply as:

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

$[H_3O^+] = 1.0 \times 10^{-7}$ mol/L

$[OH^-] = 1.0 \times 10^{-7}$ mol/L

Brackets mean molar concentration

The pH scale of Acidity and Basicity

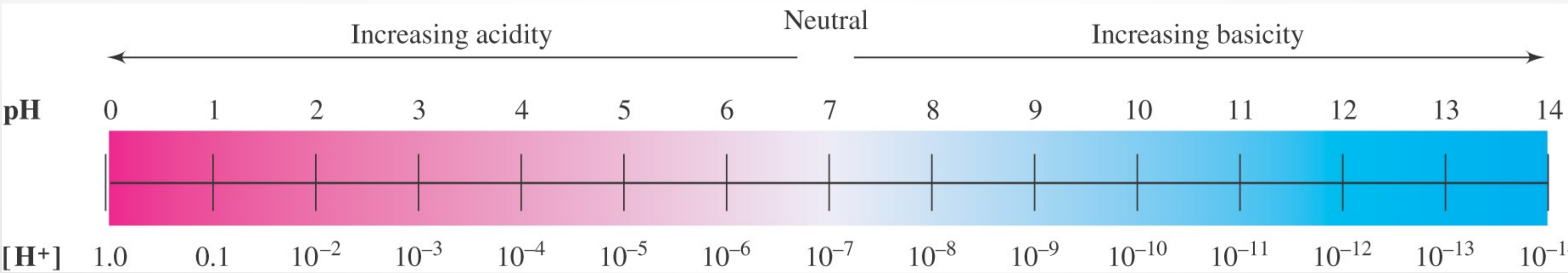


Table 14.7 Logarithm Values for Selected Numbers

Number	Number Expressed as a Power of 10	Common Logarithm
10,000	1×10^{4}	4.0
1,000	1×10^{3}	3.0
100	1×10^{2}	2.0
10	1×10^{1}	1.0
1	1×10^{0}	0.0
0.1	1×10^{-1}	−1.0
0.01	1×10^{-2}	−2.0
0.001	1×10^{-3}	−3.0
0.0001	1×10^{-4}	−4.0

INTRODUCTION TO PH

pH is the negative logarithm of the hydronium ion concentration.

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

Brackets mean molar concentration

1). What is the pH of a solution with an $[H_3O^+]$ of 1.0×10^{-11} M?

2 Significant figures

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

$$pH = -\log[1.0 \times 10^{-11}] = -(-11.00)$$

$$pH = 11.00$$

2 decimal places

2). What is the pH of a solution with an $[H_3O^+]$ of 6.0×10^{-4} M?

2 Significant Figures

$$pH = log[H_3O^+] = -log\ [6.0 \times 10^{-4}] = -(-3.22) = 3.22$$

2 decimal places

3). What is the pH of a solution with an $[H_3O^+]$ of 5.47×10^{-8} M?

3 Significant Figures

$$pH = -\log[H^+] = -\log [\, 5.47 \times 10^{-8}]$$

$$= -(-7.262) = 7.262$$

3 decimal places

BUFFERS

- A **BUFFER** IS A SOLUTION THAT RESISTS MAJOR CHANGES IN pH WHEN SMALL AMOUNTS OF ACID OR BASE ARE ADDED TO IT.
- A **BUFFER** SOLUTION IS MADE OF A **WEAK ACID** WHICH CAN REACT AND REMOVE ADDED BASE AND ITS **CONJUGATE BASE** WHICH CAN REACT WITH AND REMOVE ACID.

RECOGNIZING PAIRS THAT FUNCTION AS BUFFERS IN AQUEOUS SOLUTIONS

WHICH OF THE FOLLOWING PAIRS COULD FUNCTION AS A BUFFER?

HCl	**NaCl**	**HCl**	**NaOH**
HCN	**KCN**	**$HC_2H_3O_2$**	**$KC_2H_3O_2$**
HCl	**HCN**	**NaCl**	**NaCN**
NaCN	**KCN**	**HCN**	**$HC_2H_3O_2$**

HCN KCN **$HC_2H_3O_2$ $KC_2H_3O_2$**

These are a weak acid and the salt of that weak acid so the conjugate pair HCN/CN^- is present.

Chemical Equations for Buffer Action

Acid with a buffer:

$$H_3O^+ + C_2H_3O_2^- \rightleftharpoons HC_2H_3O_2 + H_2O$$

Base with a buffer:

$$OH^- + HC_2H_3O_2 \rightleftharpoons C_2H_3O_2^- + H_2O$$

Summary:

$$C_2H_3O_2^- \underset{OH^-}{\overset{H_3O^+}{\longleftrightarrow}} HC_2H_3O_2$$

Buffers:

- **Buffers do not hold the pH constant.**
- **pH change is less with a buffered solution than an unbuffered solution .**
- **Buffer solutions are not all pH 7.**

Unbuffered solution	pH
1L water	7.0
1 L water + 0.01 mole strong base	12.0
1 L water + 0.01 mole strong acid	2.0
Buffered solution	
1L Buffer 0.1MHPO_4^{2-}: 0.1M $H_2PO_4^-$	7.2
1 L water + 0.01 mole strong base	7.3
1 L water + 0.01 mole strong acid	7.1

ACID-BASE TITRATIONS

Acid-base titrations is a procedure where a measured volume of acid or base of known concentration is exactly reacted with a measured volume of a base or acid of unknown concentration.

Acid-base titrations function based on neutralization of the acid/base.

NEUTRALIZATION: THE REACTION OF AN ACID AND A BASE TO FORM A SALT AND WATER.

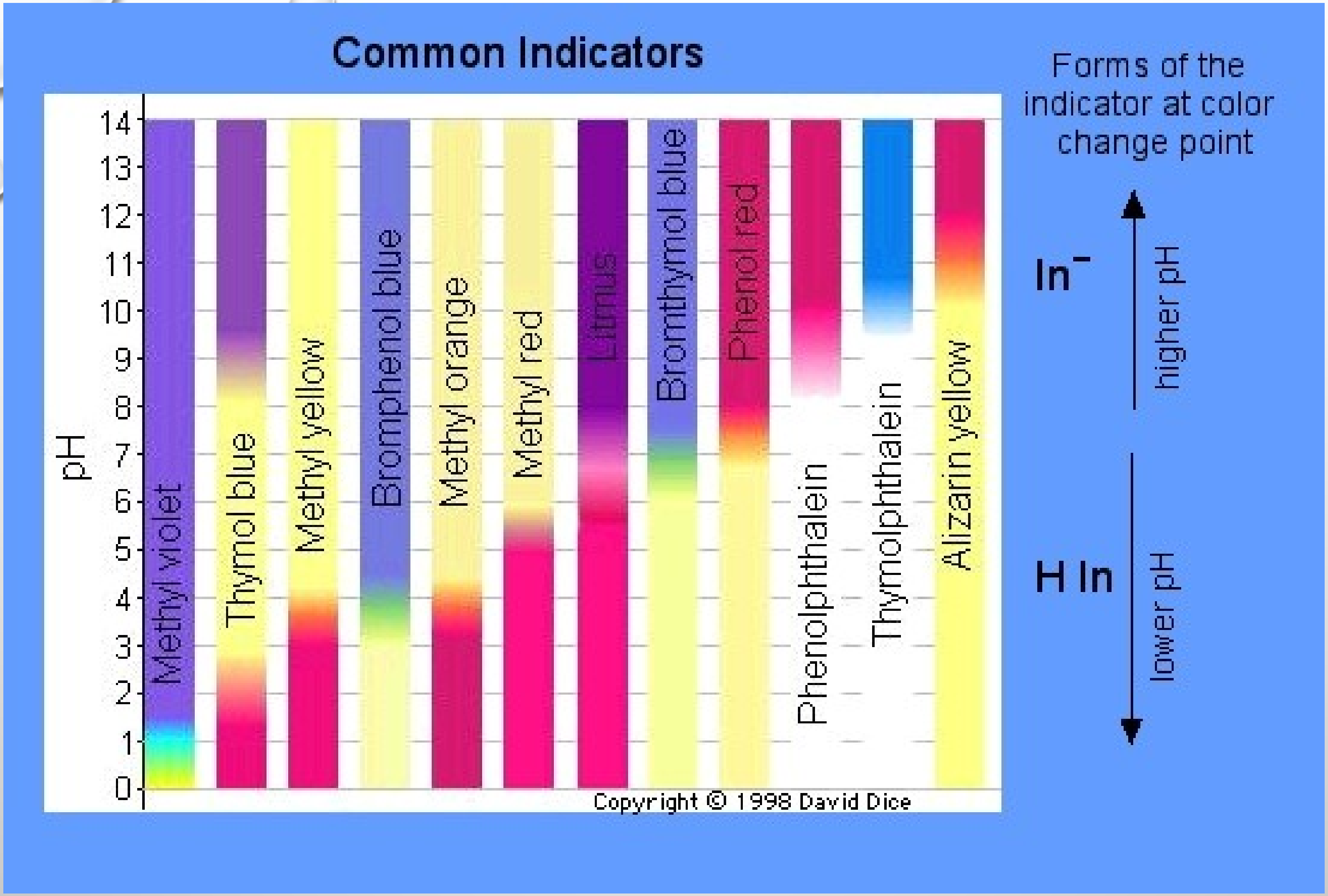
$$HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(l)$$

acid base salt

Titration: The process of measuring the volume of one reagent required to react with a measured mass or volume of another reagent.

Indicator: a compound that exhibits different colors depending on the pH of its surroundings.

Phenolphthalein: pH 8.2 = colorless; pH 10 = pink



http://www.elmhurst.edu/~chm/vchembook/images2/186indicators.jpg