Chemistry
Carmel High School
Mr. Dooner/2018

ACIDS, BASES, AND SALTS
CH 19

- Acids, bases, and salts are three classes of compounds that **FORM IONS** in SOLUTION

ACIDS(some observable properties):

- acids taste **SOUR**
- acids change the color of **LITMUS PAPER** from **blue to red**
- acids INDICATE "**acidic values**" on UNIVERSAL INDICATOR PAPER
- acids react with certain metals to **PRODUCE HYDROGEN GAS**
- acids react with metal hydroxides(BASES) to **produce WATER AND SALT**

BASES(some observable properties):

- Bases taste **BITTER** or feel **SLIPPERY**
- Bases change the color of **LITMUS PAPER** from **red to blue**
- Bases indicate "**basic values**" on UNIVERSAL INDICATOR PAPER
- Bases react with many compounds containing hydrogen ions to **produce WATER AND SALT**

INDICATORS:

- INDICATORS are chemical compounds that **CHANGE COLOR** in response to acidic or basic conditions
- A simple indicator is **pigment from red cabbage**; basic solutions turn green while acidic solutions turn red
- **Universal indicator** can be used to test **a wide range** of substances
- **Electronic pH probes** can measure precise values for acidic and basic solutions

ACID-BASE DEFINITIONS

- there are multiple "definitions" of acids and bases; these definitions developed over time as chemists' understanding of acid-base chemistry evolved

- there is no "correct" definition; they all describe acids and bases but describe what is happening at the microscopic level from different perspectives or from relatively narrow or broader approaches

- three acid-base definitions are:

1) **ARRHENIUS**
2) **BRONSTED-LOWRY**
3) **LEWIS**

ARRHENIUS ACID-BASE DEFINITION

- this definition defines acids as hydrogen containing compounds that ionize to yield hydrogen ions(H+) in aqueous solution

- it also defines bases as compounds that ionize to yield hydroxide ions(OH-) in aqueous solution

- Arrhenius acids can be **MONOPROTIC**(one ionizable hydrogen- i.e nitric acid{H2SO4}, **DIPROTIC**(two ionizable hydrogens – i.e. sulfuric acid{H2SO4}, or **TRIPROTIC**(three ionizable hydrogens – i.e. phosphoric acid{H3PO4}

- ***Not all compounds that contain hydrogen are acids however***

- Not all the hydrogens in an acid are necessarily released—**ONLY** the hydrogens in **HIGHLY POLAR BONDS** are ionizable

- an Arrhenius base ***must contain hydroxide***, such as the compound KOH

- **HOWEVER**—a compound like **NH3** does not contain hydroxide and displays all the characteristics of a base—YET, would not be considered a base according to the Arrhenius definition

- The Bronsted-Lowry definition addresses this limitation

<u>BRONSTED-LOWRY ACID-BASE DEFINITION</u>

- according to this definition, ***acids*** **<u>donate HYDROGEN IONS</u>** and ***bases*** **<u>ACCEPT HYDROGEN IONS</u>**

- acids that are formed from the nonmetals on the right side of the periodic table easily DISSOCIATE to produce hydrogen ions because these nonmetals have a large electronegativity compared with that of hydrogen

- *all the acids and bases included in the Arrhenius theory are also acids and bases according to the Bronsted-Lowry theory*; BUT, some compounds not included in the Arrhenius theory are classified as bases in the Bronsted-Lowry theory

- Returning to the example of **NH3**—because NH3 accepts hydrogen ions, it would be **a BASE according to the Bronsted-Lowry definition**

- **Bronsted-Lowry leads to the concept of <u>CONJUGATE ACID-BASE PAIRS</u>**

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

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

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

- ***Be sure you can identify the conjugate acid-base pairs in the reaction of ammonia with water and the dissociation of hydrogen chloride in water***

- Rather than imagining single hydrogen ions(H+) in solution as we do in the Arrhenius theory, in the Bronsted-Lowry model we have the concept of the **HYDRONIUM ION(H3O+)**

- **<u>A water molecule that gains a hydrogen ion is called an hydronium ion</u>**

- In this model, a water molecule can sometimes accept a hydrogen ion, and in other cases donate a hydrogen atom—in other words ***sometimes it acts as an acid and sometimes as a base***—the term for this is **AMPHOTERIC**

LEWIS ACIDS AND BASES

- a **LEWIS ACID** is an **ELECTRON PAIR RECEPTOR**, and a **LEWIS BASE** is an **ELECTRON PAIR DONOR**

- ***using the Lewis definition extends the concept of acid-base reactions to non-aqueous systems***

- Lewis acids accept pairs of electrons to form a ***covalent bond*** and Lewis bases can donate a pair of electrons to form a ***covalent bond***

- This concept is more general than either Arrhenius or Bronsted-Lowry, and the Lewis definition includes some compounds not classified as Bronsted-Lowry acids or bases

- The compound **BF3**, for example, would not be an acid according to the Bronsted-Lowry definition because the hydrogen ion is not present

The pH Scale

- **the pH scale measures the concentration of hydrogen ions in solution**

- the scale is NOT linear, but LOGARITHMIC, meaning that at pH of 2 for example, the concentration of hydrogen ions is TEN TIMES greater than it is at pH 3.

- The ph scale ranges from **0(very ACIDIC)** to **14(very BASIC)**

- **a pH of 7 is considered to be a NEUTRAL solution**

- **mathematically, the pH scale is defined as -log[H+], where [H+] is the hydrogen-ion concentration in moles per liter of solution**

WEAK AND STRONG ACIDS AND BASES

- Acids dissociate by donating hydrogen ions

* bases ionize by dissociating to form hydroxide ions(from a hydroxide salt) OR by accepting hydrogen ions

- Some acids and bases either ***dissociate or ionize*** almost **COMPLETELY**; complete dissociation/ionization results in a **STRONG ACID** or a **STRONG BASE**

- **PARTIAL** DISSOCIATION/IONIZATION results in a **WEAK ACID** or **WEAK BASE**

- The strength of an acid or a base **can VARY**—depending on such conditions as temperature and concentration

BUFFERS:

- A **BUFFER** is a solution that stabilizes H+ concentration levels

- such a solution may *RELEASE HYDROGEN IONS as pH RISES* or *CONSUME HYDROGEN IONS AS pH DECREASES*

- an important but extremely complex example is the equilibria between carbon dioxide, carbonic acid, bicarbonate, and solid calcium carbonate that keeps the world's oceans at a nearly constant pH of about 8

pOH:

- ***pOH is the negative logarithm of the OH- concentration expressed in moles per liter of solution***

- **pOH = (-)log[OH-]**

- **The sum of pH and pOH is always 14.0 for a given solution at 25 C**