

# Chem101: General Chemistry

## Lecture 9 – Acids and Bases

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### I. Introduction

- A. In chemistry, and particularly biochemistry, water is the most common solvent
1. In studying acids and bases we are going to see that water can also participate in chemical reactions
  2. It is usually not the water itself, but its ionization products  $H^+$  (hydrogen ion) and  $OH^-$  (hydroxyl ion)
- B. The O–H bond is a very polar bond and some times behaves like an ionic bond:
1.  $H-O-H \rightarrow H^+ + OH^-$ 
    - a. This reaction is called an ionization reaction
  2. In pure water only about 1 water molecule in 10,000,000 is ionized
    - a. This is why in chemical reactions we write water in its molecular form ( $H_2O$ ).
- C. However, the hydrogen and hydroxyl ions are highly reactive, so at even very low concentrations their presence is important.
1. Acids are compounds that when added to water increase the  $H^+$  concentration.
  2. Bases are compounds that when added to water increase the  $OH^-$  concentration.
    - a. In the process they lower the hydrogen ion concentration
- D. One measure that is used to indicate the hydrogen ion concentration is the *pH* value.
1. Pure water has a *pH* of 7.
  2. **Acids** are compounds that when added to water cause the *pH* value to become less than 7.
  3. **Bases** are compounds that when added to water cause the *pH* value to become greater than 7.
  4. These are what I call *operational definitions* of acids and bases
    - a. We are also going to discuss a couple of other definitions of acids and bases.
- E. When acids and bases are mixed they react to counteract (neutralize) each other and produce salts.

### II. Arrhenius Theory

- A. The Arrhenius theory provides us with another definition for acids and bases.
1. The theory was introduced in 1887 by the Danish chemist Svante Arrhenius.
- B. *The Arrhenius definition of acids and bases:*
1. **Acids** are electrolytes (ionic substances) that when dissolved in water release hydrogen ions ( $H^+$ ).
  2. **Bases** are electrolytes (ionic substances) that when dissolved in water release hydroxyl ions ( $OH^-$ ).
- C. Hydrogen chloride is an example of an Arrhenius acid:
1. Hydrogen chloride ( $HCl$ ) is a gas that when dissolved in water ionizes (dissociates) to produce hydrogen ions ( $H^+$ ) and chloride ions ( $Cl^-$ ).

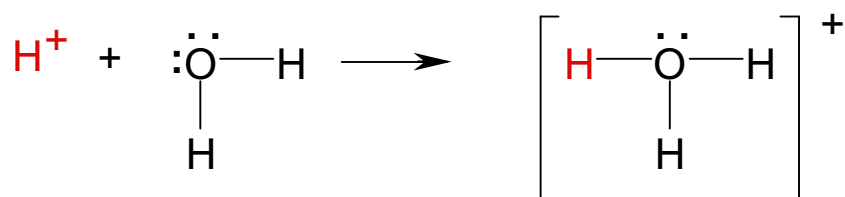


- D. Sodium hydroxide is an example of an Arrhenius base:
1. When sodium hydroxide dissolves in water it ionizes (dissociates) to produce sodium ions ( $\text{Na}^+$ ) and hydroxyl ions ( $\text{OH}^-$ ).



### III. The Brønsted Theory

- A. A hydrogen ion is a hydrogen atom that is missing its one electron.
1. This leaves only a proton.
    - a.  $\text{H}^+ = {}_1^1\text{p}$
  2. Protons do not exist free in water, but instead attach themselves to one of the water molecules to form a hydronium ion ( $\text{H}_3\text{O}^+$ )



*hydronium ion*

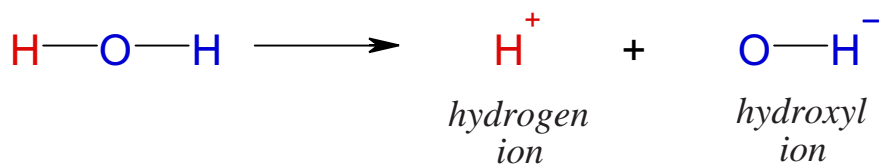
- B. Some substances that are operationally bases are not accounted for by the Arrhenius definition of a bases.
- $\text{NH}_3$  in water
- C. Johannes Brønsted of Denmark and Thomas Lowry of England redefined acid
1. **Acid** is a proton donor
  2. **Base** is a proton acceptor
- D. Brønsted/Lowry acid/base pairs
1. Acid becomes the **conjugate base**
  2. Base becomes the **conjugate acid**.

### IV. Naming Acids

- A. Binary acids
1. Hydrogen chloride - hydrochloric acid
- B. Poly atomic acids
1. Trihydrogen phosphate - phosphoric acid
  2. Dihydrogen sulphate - sulfuric acid
  3. Dihydrogen sulphite - sulfurous acid

### V. The Self-Ionization of Water

- A. Water is both an acid and a base
1. Arrhenius definition:



- B. Equilibrium constant for the ionization of water.
- C. Ionization product for water ( $K_w$ )
- D. Neutral solution  $[\text{H}_3\text{O}^+] = [\text{OH}^-]$ .
- E. Acidic solution  $[\text{H}_3\text{O}^+] > [\text{OH}^-]$ .
- F. Basic or alkaline solution  $[\text{H}_3\text{O}^+] < [\text{OH}^-]$ .

### VI. The *pH* concept

- A.  $[\text{H}^+]$  can have a wide range,  $10\text{ M}$  to  $1 \times 10^{-14}\text{ M}$ .
- B. Sørensen notation
  1.  $pH = -\log([\text{H}_3\text{O}^+])$  or  $pH = -\log([\text{H}^+])$ .
  2.  $[\text{H}_3\text{O}^+] = [\text{H}^+] = 10^{-pH}$  (antilog)

**Table 9.1** - Relationships between  $[\text{H}^+]$ ,  $[\text{OH}^-]$ , and *pH*.

**Table 9.2** - Calculating *pH* from molarity with a calculator.

**Table 9.3** - Calculating molarity from *pH* with a calculator

**Table 9.4** - Common laboratory acids and bases.

### VII. Properties of Acids

- A. All acids taste sour.
- B. All acids produce  $\text{H}_3\text{O}^+$  ions.
- C. Undergo characteristic double-replacement reactions with solid oxides, hydroxides, carbonates and bicarbonates.
  1. Reactions with  $\text{Cu}_2\text{O}$ ,  $\text{Ca}(\text{OH})_2$  and  $\text{CaCO}_3$
- D. React with certain metals to produce hydrogen gas
  1. Zn and HCl
  2. K and  $\text{H}_2\text{O}$

**Table 9.5** - Activity series of the metals.

### VIII. Properties of Bases

- A. Feel soapy
- B. Neutralize acids

### IX. Salts

- A. Solids at room temperature.

- B. Product of an acid/base neutralization reaction
  1. Cation comes from the base
  2. Anion comes from the acid.
- C. Acid base reactions that produce salts
  1. Acid + metal  $\rightarrow$  salt +  $H_2$
  2. Acid + metal oxide  $\rightarrow$  salt + water.
  3. Acid + metal hydroxide  $\rightarrow$  salt + water.
  4. Acid + metal carbonate  $\rightarrow$  salt + water +  $CO_2$ .
  5. Acid + metal bicarbonate  $\rightarrow$  salt + water +  $CO_2$ .
- D. Some salts exist as **hydrates** in their solid form
  1. In lab you determined the number of waters of hydration for the alum
  2. The waters of hydration combine with the salt in a very specific mole ratio.
    - a. For alum, you found that  $n=12$ .

<b>Table 9.6</b> - Some useful and common hydrates.
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### X. Strengths of Acids and Bases

- A. When salts dissolve in water they dissociate completely, not so with acids and bases.
  1. H- bond is a polar covalent bond.
  2. The more polar the bond is, the greater its ionic character, the more likely it is to dissociate into ions.
  3. Not all the hydrogens on molecules are necessarily acidic
    - a. For example, only one of the 4 hydrogens of acetic acid is acidic.
- B. The equilibrium constant for an acid dissolved in water.
  1.  $HB + H_2O \rightarrow H_3O^+ + B^-$
- C. The acid dissociation constant,  $K_a$ .
  1. For strong acids  $K_a$  is greater than 1.
  2. For weak acids  $K_a$  is much less than 1.

<b>Table 9.7</b> - Some common strong and weak acids
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- D. Mono-, Di- and Polyprotic acids.
  1. Each proton that the acid transfers has its own acid dissociation constant
  2. Species that serve as both acids and bases are called *amphiprotic* or *amphoteric*.
  3. The number of ionizable hydrogens cannot always be determined by looking at the molecular formula.
    - a. You need to look at the structural formula to determine which hydrogens are bonded to electronegative atoms such as oxygen, sulfur, and the halogens.

**XI. Analysis of Acids and Bases**

- A. Determining the concentrations of acids and bases.
1. The quantity of an acid can be determined by reacting them with a known amount of base.
  2. The quantity of a base can be determined by reacting them with a known amount of acid.
- B. For weak acids, all of the acid will react, not just the portion that has ionized.
1. This can be applied by applying *LeChatlier's Principal*.
- C. The procedure of neutralizing an acid with a base or a base with an acid for the purpose of determining the concentration of the acid or base, is called an acid/base **titration**.
1. When titrating an acid of unknown concentration with a base of known concentration, the basic solution is added to the acid solution until the **equivalence point** is reached.
    - a. This is the point where the number of moles of base added is equal to (equivalent to) the number of moles of acid present at the beginning of the reaction.
  2. During the titration, the acid reacts with the added base by donating its proton to the base, and in the process becomes its conjugate base.
    - a. When a solution contains a mixture of an acid and its conjugate base, it is called a **buffer**.
    - b. Buffers resist changes in *pH*.
      - i. Consequently, as the base is added the *pH* of the solution changes slowly at first.
      - ii. When the equivalence point is reached, there is no longer any acid left to neutralize it, so the *pH* will suddenly increase very rapidly.
      - iii. This is how the equivalence point is detected.
        1. In lab you used the *pH* indicator phenolphthalein to determine the equivalence points.
        2. Phenolphthalein is a *pH* indicator which changes from colorless to pink when the *pH* of a solution is above *pH* 9.
        3. When it turns pink during a titration, it is an indication that the equivalence point has been reached.

**XII. Titration Calculations**

- A. Acid/base titrations can be used to determine the concentration of either an acid or a base solution.
1. For an acid solution of unknown concentration, a strong base of known concentration is added to a known volume of the acid until the equivalence point is reached.
    - a. At the equivalence point we know:
      - i. The initial volume of the acid
      - ii. The volume and the concentration of the base
    - b. We also know the chemical equation for the acid/base reaction.

- c. From the volume and concentration of the bases we can calculate the number of moles of base that reacted.
- d. From the chemical equation we can determine the number of moles of the acid that reacted with each mole of the base added.
- e. Dividing the number moles of acid by its initial volume gives its initial concentration.

### **XIII. Hydrolysis Reactions of Salts**

- A. Consider salts as the product of an acid base reaction
  1. Compare the relative strengths of the acid and the base that would go make the salt.

### **XIV. Buffers**

- A. A mixture of a weak acid and its conjugate base
  1. The weak acid will neutralize added base.
  2. The weak conjugate base will neutralize added acid.
  3. In this way buffers help to resist changes in  $pH$ .
- B. Henderson-Hasselbalch equation