**10. Acids, Bases, and Salts**

**Acids and bases**

For centuries people have known acids are in vinegar, lemon juice and many other foods taste sour. Only few hundred years ago that it was discovered that acids taste sour and change litmus **red**, and they could be identified as **acids**. In Latin term *acere*, which means "**sour**." While **bases** feel slippery, change litmus **blue**. There are three different definitions of acids and bases, in this section we will introduce the fundamentals of acid/base chemistry.

* **Acids** taste sour, make metals corrode, change litmus (a dye from plants) red, and get neutralized when bases are added.
* **Bases** feel slippery, change litmus blue, and get neutralized when mixed with acids.

**10.1 Arrhenius Acid-Base Theory**

**Arrhenius Definition**:

Arrhenius's theory **explains** why all acids have similar properties to each other (and, conversely, why all bases are similar): because all acids release H+ (prtoton) or H3O+ (**hydronium ions**) into solution (and all bases release **hydroxide ions**, OH-).

**Arrhenius Acid**:

A substance that produces H +, or (protons) H+3O, (hydronium ion) in an aqueous solution.

**Arrhenius Base**:

A substance that produces OH-, or hydroxide ion in an aqueous solution.

This is the first acid/base concept to be developed to describe typical acid/base reactions.

**E.g.** HCl (acid), NaOH (base).

**10.2 Bronsted-Lowry Acid-Base Theory**

Brønsted (Denmark) and Lowry (England) came up with an alternative acids and base bases definition to Arrhenius. There solved the problems associated with **non hydroxide bases**, especially ammonia which Arrhenius' definition could not include as a base limiting the acid base reactions to few reactions.

According to **Brønsted-Lowry**: Acids and bases are substances that are capable of donating and accepting protons (hydrogen ions, H+), respectively. An acid-base reaction consists of the transfer of a proton from an acid to a base. Acid and bases are considered as proton transfer agents.

* An acid is a "**proton donor**."
* A base is a "**proton acceptor**."

Lowry proposed the use of hydronium ion H3O+ in the place of H+ that is commonly used today. He pointed out acidity is a relative thing comparing proton donor ability of two pure compounds. Even hydrogen chloride only becomes an acid when mixed with water.

HCl + H2O  H3O+ + Cl¯

This reaction proceeds to **right** to a large extent:

**HCl** - this is a Brønsted-Lowry acid, because it has a proton available to be transferred.

**H2O** - this is a Brønsted-Lowrybase, since it gets the proton that the Brønsted-Lowryacid lost.

**Conjugate Acid-Base Pairs**

Since this an equilibrium reaction the reverse reaction could also be considered as an acid/base reaction:

H3O+ + Cl¯  HCl + H2O

Acid and bases involved in the reverse reaction is call **conjugate acid** and **conjugate base**

**Conjugate acid**: **H3O+** - this is a Brønsted-Lowry **conjugate acid**, because it can give a proton.

**Conjugate base: Cl¯** - this is a Brønsted-Lowry **conjugate base**, since it has the capacity to receive a proton.

A conjugate pair is an acid-base pair that differs by one proton in their formulas (remember: proton, hydrogen ion, etc.).

A conjugate pair is always one acid and one base.

HCl + H2O  H3O+ + Cl¯

Here is the one conjugate pair (acid/conjugate base) from the first example reaction:

**HCl and Cl¯**

The other conjugate pair is:

**H2O and H3O+**

**Some more conjugate acid-base pairs to look for**:

|  |
| --- |
| H2O and OH¯  HCO3¯ and CO32¯  H2PO4¯ and HPO42¯  HSO4¯ and SO42¯  NH4+ and NH3  CH3NH3+ and CH3NH2  HC2H3O2 and C2H3O2¯ |

**Identify the Bronsted-Lowery acid/conjugate base and base/conjugate acid pairs in the equilibrium reactions given below**

a) HCl(aq) + H 2O(l)  H 3+O(aq) + Cl¯(aq)

b) H2SO4(aq) + H2O(l)  H 3+O(aq) + HSO4¯(aq)

c) H2O(l) + H 2O(l)  H 3+O(aq) + OH¯(aq)

d) NH3 (aq) +H 2O(l)  NH 4+ + OH ¯(aq)

The concept of acid\conjugate base pair and base\conjugate acid pair came out of Bronsted definition describing proton transfer reactions.

a) HCl/Cl¯ is an acid/conjugate base pair

H2O/H3+O is a base/conjugate acid pair in this equilibrium.

b) H2SO4/HSO4¯ is an acid/conjugate base pair and H2O/H3+O is a base/conjugate acid pair in this equilibrium.

HSO4¯/SO42¯ is an acid/conjugate base pair, if the second dissociation of HSO4¯ (aq) took place.

c) H2O/OH¯ is an acid/conjugate base pair in this equilibrium.

H2O/ H3+O is a base/conjugate acid pair and H2O/ OH ¯ is an acid/conjugate base pair in this equilibrium.

H2O/ H3+O is a base/conjugate acid pair and HC2H3O2/C2H3O2¯ an acid/conjugate base pair in this equilibrium.

d) NH3 / NH4+ is a base/conjugate acid pair and H2O/OH¯ is an acid/conjugate base pair in this equilibrium.

**10.3 Mono-, Di-, and Triprotic Acids**

**Polyprotic Acids**  
In contrast to a simple monoprotic acid like acetic acid, with only one equilibrium between the acid and conjugate base, a polyprotic acid contains more than one acidic hydrogen. For a polyprotic acid, n acidic hydrogens will exist in solution in equilibrium with n conjugate base forms (for a total of n+1 species). For example, when phosphoric acid (n = 3, a triprotic acid) is dissolved in solution, the following equilibria are established among the four species H3PO4 (phosphoric acid itself),  
10.4 Strengths of Acids and Bases  
10.5 Ionization Constants for Acids and Bases  
10.6 Salts  
**10.7 Acid-Base Neutralization Reactions**

**Neutralization**: This idea, that a base (or acid) can make an acid (or base) weaker by converting them to water is called neutralization.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| H+(aq) | + | OH-(aq) | http://www.howe.k12.ok.us/~jimaskew/arrow.jpg | H2O(l) |

As you can see from the equations, acids release H+ into solution and bases release OH-. If we were to mix an acid and base together, the H+ [ion](http://www.visionlearning.com/library/pop_glossary_term.php?oid=853&l=) would combine with the OH- ion to make the molecule H2O, or plain water:

The neutralization reaction of an acid with a base will always produce water and a [salt](http://www.visionlearning.com/library/pop_glossary_term.php?oid=1575&l=), as shown below:

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| **Acid** |  | **Base** |  | **Water** |  | **Salt** |
| HCl | + | NaOH | http://www.howe.k12.ok.us/~jimaskew/arrow.jpg | H2O | + | NaCl |
| HBr | + | KOH | http://www.howe.k12.ok.us/~jimaskew/arrow.jpg | H2O | + | KBr |

10.8 Self-Ionization of Water

H2O(l) + H 2O(l)  H 3+O(aq) + OH¯(aq)

This equilibrium is called **autoionization** or **self-ionization** of water.

10.9 The pH Concept

Under the Brønsted-Lowry definition, both [acids](http://www.visionlearning.com/library/pop_glossary_term.php?oid=1573&l=) and [bases](http://www.visionlearning.com/library/pop_glossary_term.php?oid=1574&l=) are related to the concentration of hydrogen [ions](http://www.visionlearning.com/library/pop_glossary_term.php?oid=853&l=) present.  Acids increase the concentration of hydrogen ions, while bases decrease the concentration of hydrogen ions (by accepting them).  The acidity or basicity of something therefore can be measured by its hydrogen ion concentration.

In 1909, the Danish biochemist Sören Sörensen invented the [pH](http://www.visionlearning.com/library/pop_glossary_term.php?oid=1577&l=) scale for measuring acidity.  The pH scale is described by the formula:

|  |  |  |
| --- | --- | --- |
| |  |  | | --- | --- | | **pH = -log [H+]** | Note: concentration is commonly abbreviated by using square brackets, thus [H+] = hydrogen [ion](http://www.visionlearning.com/library/pop_glossary_term.php?oid=853&l=) concentration.  When measuring [pH](http://www.visionlearning.com/library/pop_glossary_term.php?oid=1577&l=), [H+] is in [units](http://www.visionlearning.com/library/pop_glossary_term.php?oid=848&l=) of [moles](http://www.visionlearning.com/library/pop_glossary_term.php?oid=1515&l=) of H+ per liter of [solution](http://www.visionlearning.com/library/pop_glossary_term.php?oid=1571&l=). | |

For example, a [solution](http://www.visionlearning.com/library/pop_glossary_term.php?oid=1571&l=) with [H+] = 1 x 10-7 moles/liter has a [pH](http://www.visionlearning.com/library/pop_glossary_term.php?oid=1577&l=) equal to 7 (a simpler way to think about pH is that it equals the exponent on the H+ concentration, ignoring the minus sign). The pH scale ranges from 0 to 14. Substances with a pH between 0 and less than 7 are [acids](http://www.visionlearning.com/library/pop_glossary_term.php?oid=1573&l=) (pH and [H+] are inversely related - lower pH [means](http://www.visionlearning.com/library/pop_glossary_term.php?oid=4221&l=) higher [H+]). Substances with a pH greater than 7 and up to 14 are [bases](http://www.visionlearning.com/library/pop_glossary_term.php?oid=1574&l=) (higher pH means lower [H+]). Right in the middle, at pH = 7, are [neutral](http://www.visionlearning.com/library/pop_glossary_term.php?oid=855&l=) substances, for example, pure water. The relationship between [H+] and pH is shown in the table below alongside some common examples of acids and bases in everyday life.

|  |  |  |  |
| --- | --- | --- | --- |
|  | **[H+]** | **pH** | **Example** |
| Acids | 1 X 100 | 0 | HCl |
| 1 x 10-1 | 1 | Stomach acid |
| 1 x 10-2 | 2 | Lemon juice |
| 1 x 10-3 | 3 | Vinegar |
| 1 x 10-4 | 4 | Soda |
| 1 x 10-5 | 5 | Rainwater |
| 1 x 10-6 | 6 | Milk |
| Neutral | 1 x 10-7 | 7 | Pure water |
| Bases | 1 x 10-8 | 8 | Egg whites |
| 1 x 10-9 | 9 | Baking soda |
| 1 x 10-10 | 10 | Tums® antacid |
| 1 x 10-11 | 11 | Ammonia |
| 1 x 10-12 | 12 | Mineral lime - Ca(OH)2 |
| 1 x 10-13 | 13 | Drano® |
| 1 x 10-14 | 14 | NaOH |

10.10 The pKa Method for Expressing Acid Strength  
10.11 The pH of Aqueous Salt Solutions  
Chemistry at a Glance: Acids and Acidic Solutions  
10.12 Buffers  
10.13 The Henderson-Hasselbalch Equation  
Chemistry at a Glance: Buffer Systems  
10.14 Electrolytes  
10.15 Acid-Base Titrations  
Chemical Connections: Excessive Acidity Within the Stomach: Antacids and Acid Inhibitors; Acid Rain: Excess Acidity; Blood Plasma pH and Hydrolysis; Buffering Action in Human Blood; Electrolytes and Body Fluids