**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.

**Names and formulas of common acids and bases**

|  |  |  |
| --- | --- | --- |
| **Binary Acids** | **Ternary (oxo) Acids** | **Hydroxy Bases** |
| Hydrofluoric Acid - HFHydrochloric Acid - HClHydrobromic Acid - HBrHydroiodic Acid - HI | Nitric Acid - HNO3Sulfuric Acid - H2SO4Phosphoric Acid - H3PO4Carbonic Acid - H2CO3 Hypochlorous Acid - HClOChlorous Acid - HClO2Chloric Acid - HClO3Perchloric Acid - HClO4 | Sodium Hydroxide - NaOHPotassium Hydroxide - KOHAmmonium Hydroxide - NH4OHCalcium Hydroxide - Ca(OH)2Magnesium Hydroxide - Mg(OH)2Barium Hydroxide - Ba(OH)2Aluminum Hydroxide - Al(OH)3Zinc Hydroxide - Zn(OH)2Lithium Hydroxide - LiOH  |
|  **Organic Acids** |
| Acetic Acid - HC2H3O2Benzoic Acid - C6H5CO2H Oxalic Acid –H2C2O4 |

**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+ (proton) 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.** Arrhenius Acid- HCl (acid), Arrhenius Base-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 NH3CH3NH3+ and CH3NH2HC2H3O2 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)**Error! Bookmark not defined.** 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-Lowery 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**

Acidity of a compound is mainly created by certain types of hydrogen present in the acid molecule. Not all hydrogen in an acid molecule break off to produce H+ ions: **the acidic protons**. Most common family acids are known as **monoprotic acids** have only one acidic proton. Each of these acids has a single H+ ion, or proton, it can donate when it acts as a Brnsted acid. Hydrochloric acid (HCl), acetic acid (CH3CO2H or HOAc), nitric acid (HNO3), and benzoic acid (C6H5CO2H) are all monoprotic acids.

**Polyprotic acids**

Several important acids can be classified as **polyprotic acids**, which can lose more than one H+ ion when they act as Brnsted acids.

 **Diprotic acids**: Have two acidic protons as in the sulfuric acid (H2SO4), carbonic acid (H2CO3), hydrogen sulfide (H2S), chromic acid (H2CrO4), and oxalic acid (H2C2O4).

**Triprotic acids**: Have three acidic protons as in such as phosphoric acid (H3PO4) and citric acid (C6H8O7).

**10.4 Strengths of Acids and Bases**
**Strong acid and bases**

In strong acids the first acidic hydrogen is completely separated or ionized (100%) exists as an hydronium ions, H3O+ in aqueous solutions.

 HCl + H2O  H3O+ (aq) + Cl- (aq) (100 % ions)

Strong bases also show complete separation (100%) into ions in aqueous solutions.

 NaOH(s) + H2O  Na+ (aq) + OH-(aq) (100 % ions)

Most bases in Groups 1A and 2A are strong bases.

|  |  |
| --- | --- |
| **Strong acid** | **Strong bases** |
| **Name- formula** | **Name- formula** |
| Hydrioidic HI Hydrobromic HBrPerchloric HClO4Hyrdrochloric HCl Chloric HClO3Sulfuric H2SO4Nitric HNO3  | Lithium hydroxide LiOHSodium hydroxide NaOHPotassium hydroxide KOH Rubidium hydroxide RbOHCesium hydroxideCsOHCalcium hydroxide Ca(OH)2 |

In weak acids the first acidic hydrogen in some molecules are not separated or ionized (< 1-10%) only exists as an hydronium ions, H3O+ in aqueous solutions.Acetic acid, CH3COOH in water contains mostly molecules and a few CH3COO¯ ions and it called **partial dissociation**.

 CH3COO**H** + H2O  H3O+ (aq) + CH3COO¯ (aq) (< 1 % ions)

Weak does not bases also show complete dissociation (< 1%) into ions in aqueous solutions.

 NH4OH + H2O  NH4+ (aq) + OH-(aq) (< 1 % ions)

|  |  |
| --- | --- |
| **Weak acid** | **Weak bases** |
| **Name- formula** | **Name- formula** |
| Hydrofluoric acid HF Formic acid HCOOHAcetic acid CH3COOH Nitrous acid HNO2 Hydrocyanic acid HCN  | Magnesium hydroxide Mg(OH)2Calcium hydroxide Ca(OH)2 Strontium hydroxide Sr(OH)2 Barium hydroxide Ba(OH)2  |

**10.5 Ionization Constants for Acids and Bases**The equilibrium constants equilibrium expression for the ionization of strong and weak acids is useful in predicting the strength of acids and bases.

**For a strong acid** such as HCl in water:

 HCl(aq) + H2O(l)  H3O+ (aq) + Cl- (aq) (100 % ions)

Equilibrium constant expression for the above equilibrium:

             [H3O+]  [Cl-]
**Ka** =   ----------------- =    1 x 107 **Ka large for strong acids**
                [HCl]

**For a strong** bases such as NaoH which show complete separation (100%) into ions in aqueous solutions.

 NaOH(s) + H2O  Na+ (aq) + OH-(aq) (100 % ions)

             [Na+]  [OH-]
**Kb** =   ----------------- =    1 x 102 **Kb very large for strong bases**
                [NaOH]

**For a weak acid** that does not show complete dissociation (< 1%) into ions in aqueous solutions.

CH3COO**H** + H2O  H3O+ (aq) + CH3COO¯ (aq) (< 1 % ions)

             [H3O+]  [CH3COO¯]
**Ka** =   ----------------------- =    1 x 10-8 **Ka very small for weal acids**
                [CH3COO**H**]

**For a weak acid** that does not show complete dissociation (< 1%) into ions in aqueous solutions.

 NH4OH + H2O  NH4+ (aq) + OH-(aq) (< 1 % ions)

             [NH4+]  [OH-]
**Kb** =   ----------------- =    1 x 10-8 **Kb very small for weak bases**
                [NH4OH]

The definition of Ka is: The pKa value is defined from Ka and can be calculated from the Ka value and from the equation

 **pKa = −Log (Ka)**

**Comparing Ka & Kb**

* Any compound with a larger Ka value is the stronger acid.
* Any compound with a larger Kb value is the stronger base.

|  |  |
| --- | --- |
| **Strong acid** | **Strong bases** |
| **Name- formula** | **Ka** | **pKa** | **Name- formula** | **Kb** | **pKb** |
|  HI HBrHClO4HCl HClO3H2SO4HNO3  | ~ 1011~ 109~ 107~ 107~ 103~ 102~ 20 | -11-9-7-7-3-2-1.3  | LiOHNaOHKOH RbOHCsOHCa(OH)2 | 102-103 | 2-3 |

|  |  |
| --- | --- |
| **Strong acid** | **Strong bases** |
| **Name- formula** | **Ka** | **pKa** | **Name- formula** | **Kb** | **pKb** |
| HF HCOOH CH3COOH HNO2 HCN  | 6.6x10-4 1.77x10-4 1.76x10-5 4.6x10-4 6.17x10-10  | 3.183.754.75 3.349.21 | Mg(OH)2Ca(OH)2  Sr(OH)2  Ba(OH)2  | 0.01-0.1 | -2 **to** -1 |

**Which acid is weaker? Answer: HCN**

HNO2    ;  Ka= 4.0 x 10-4.

HOCl2 ;Ka= 1.2 x 10-2.

HOCl     ;  Ka= 3.5 x 10-8.

HCN      ;  Ka= 4.9 x 10-10. **(Any compound with a smaller Ka value is the weaker acid)**

**Polyprotic Acids have more than one Ka**H3PO4 is triportic therefore has three **Ka1**, **Ka2**, and **Ka3.**

H3PO4 + H2O  H3O+ (aq) + H2PO4¯ (aq); **Ka1 =**7.1 x 10-3

H2PO4¯ + H2O  H3O+ (aq) + HPO42¯ (aq); **Ka2 =**6.3 x 10-8

HPO42¯ + H2O  H3O+ (aq) + PO43¯ (aq); **Ka3 =**4.2 x 10-13

**Acid-Dissociation Equilibrium Constants for Common Polyprotic Acids**

|  |  |  |  |
| --- | --- | --- | --- |
| **Acid** | **Ka1** | **Ka2** | **Ka3** |
| sulfuric acid (H2SO4) | 1.0 x 103  | 1.2 x 10-2 |  |
| chromic acid (H2CrO4) | 9.6 | 3.2 x 10-7 |  |
| oxalic acid (H2C2O4) | 5.4 x 10-2  | 5.4 x 10-5 |  |
| sulfurous acid (H2SO3) | 1.7 x 10-2  | 6.4 x 10-8 |  |
| phosphoric acid (H3PO4) | 7.1 x 10-3  | 6.3 x 10-8  | 4.2 x 10-13 |
| glycine (C2H6NO2)  | 4.5 x 10-3  | 2.5 x 10-10 |  |
| citric acid (C6H8O7)  | 7.5 x 10-4  | 1.7 x 10-5 | 4.0 x 10-7 |
| carbonic acid (H2CO3) | 4.5 x 10-7  | 4.7 x 10-11 |  |
| hydrogen sulfide (H2S) | 1.0 x 10-7  | 1.3 x 10-13 |  |

**Citric acid (C6H8O7 or H3C6H5O7) is triprotic. Write three dissociation equilibriums for it:**

H3C6H5O + H2O  H3O+ (aq) + H2C6H5O ¯ (aq); **Ka1 =** 7.5 x 10-4

H2C6H5O¯ + H2O  H3O+ (aq) + HC6H5O2¯ (aq); **Ka2 =** 1.7 x 10-5

HC6H5O2¯ + H2O  H3O+ (aq) + C6H5O3¯ (aq); **Ka3 =** 4.0 x 10-7

**10.6 Salts**

Salts are formed when an acid reacts with a base to from water.

 **E.g. HCl(acid) + NaOH(base) --> NaCl (salt) + H2O.**

Salts normally don't have acidic protons or OH- ions. One can get the parent acid and base from a salt by breaking the compound into metal (cation) and anion (non-metal or polyatomic anions) and adding H+ to anion and OH- to metal.

**Identify acids bases and salts among the following list:**

**AgNO3, NaCl, H3PO4, NaOH, HCl, NaNO3, H2SO4, HC2H3O2, KNO3, HNO3, Mg(OH)2, KOH, H2SO3, NH4OH.**

**Salts**: AgNO3(Salt), NaCl(Salt), NaNO3(salt), KNO3(salt),

**Acids**: H3PO4(acid), HCl(acid), H2SO4(acid), HC2H3O2(acid), HNO3(acid), H2SO3(acid)

**Bases**: Mg(OH)2(base), NaOH(base), KOHbase), NH4OH(base).

**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. 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:

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

 **Acid + Base  Salt + H2O(l)**

 **Acid + NaOH  Sodium Salt + H2O(l)**

|  |  |  |
| --- | --- | --- |
| **Acid Name** | **Formula Name** | **Sodium salt** |
| **Name** | **Formula** |
| Acetic acid | HC2H3O2 | Sodium acetate | NaC2H3O2 |
| Hydrogen chloride | HCl | Sodium chloride |   |
| Nitric acid  | HNO3 | Sodium nitrate | NaNO3  |
| Phosphoric acid | H3PO4 | Sodium phosphate | Na3PO4 |
| Sulfuric acid | H2SO4 | Sodium sulfate | Na2SO4 |

 **HCl + Base  Chloride Salt + H2O(l)**

|  |  |  |
| --- | --- | --- |
| **Base Name** | **Formula Name** | **Chloride salt** |
| **Name**  | **Formula** |
| Sodium hydroxide | NaOH | Sodium chloride | NaCl |
| Barium oxide | BaO | Barium chloride | BaCl2 |
| Calcium oxide | CaO | Calcium chloride | CaCl2 |

**10.8 Self-Ionization of Water**

A water molecule can react with another molecule of water to produce ions and the process is called autoionization.

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

This equilibrium is called **autoionization** or **self-ionization** of water. The equilibrium constant for this reaction is called **Kw**.

**Kw  = [H3O+]  [OH¯]**; **Kw =**1 x 10-14 **is very small** and water there are only **1 x 10-7 moles of H3O+** and **1 x 10-7 moles of OH¯**in one liter (1L) of water.

**10.9 The pH Concept**

In aqueous solutions 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=) alter the autoinization equilibrium of the water and change the concentration of hydrogen [ions](http://www.visionlearning.com/library/pop_glossary_term.php?oid=853&l=).  Acids increase the concentration of hydrogen ions, while bases decrease the concentration of hydrogen ions (by converting them to water).  The acidity or basicity of an aqueous solution, therefore, can be measured by its hydrogen ion concentration.

In 1909, the Danish biochemist 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=). |

**Neutral pH**

For example, pure water [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+ or H3O+ concentration (written [H+] or [H3O+]), ignoring the minus sign). The pH scale ranges from 0 to 14.

**Acidic pH**

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+]).

**Basic pH**

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



|  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| **pH** | 0 | 2 | 4 | 6 | 7 | 8 | 10 | 12 | 14 |
|  | **1 X 100** | **1 x 10-2** | **1 x 10-4** | **1 x 10-6** | **1 x 10-7** | **1 x 10-8** | **1 x 10-10** | **1 x 10-12** | **1 x 10-14** |

 **What is the pH of a solution that has a hydrogen ion concentration of 1.0 x 10-3 M?**
****

**What is the hydrogen ion concentration of a solution that has a pH of 5?
**

**10.10 The pKa Method for Expressing Acid Strength**

**Ka and pKa Values for acid dissociations**

The Ka value is a value used to describe the tendency of compounds or ions to dissociate

and is also called the dissociation constant, the ionization constant and the acid constant.

The pKa value is a value related to the Ka value in a logic manner. pKa values are easier

to remember than Ka values and pKa values are in many cases easier to use for fast

approximations of concentrations of compounds and ions in equilibriums than Ka values.

**Definition of pKa**

The definition of Ka is: The pKa value is defined from Ka and can be calculated from the Ka value and from the equation

 **pKa = −Log (Ka)**

**Example on how pKa is calculated**

Given Ka = 1.8 x 10-5 what is pKa? Answer 4.74

pKa = - log Ka means pKa =4.74 do this in steps take log 1.8 x 10-5 = -4.74 and change sign (+/-) give 4.74

Any time you see a "**p**" in terms of acid-base chemistry, it automatically signifies

 **-log** (whatever is after p).

**Comparing pKa & pKb**

* Any compound with a smaller pKa value is the stronger acid.
* Any compound with a smaller pKb value is the stronger base.

**Which acid is Stronger? Answer: propanoic acid**

|  |  |  |
| --- | --- | --- |
| Acid | p*K*a |  |
| HCOOH (formic acid) | 3.75 |  |
| CH3COOH (acetic acid) | 4.76 |  |
| CH3CH2COOH (propanoic acid) | 4.86  | **Any compound with a smaller pKa value is the stronger acid.**  |

**10.11 The pH of Aqueous Salt Solutions**

Some salts dissolved in water make the solution acidic or basic depending on the type of salt. For example, when ammonium chloride, NH4Cl, is dissolved in water the solution become acidic. Since NH4Cl is produced by the reaction of HCl (a strong acid) with NH4OH (a weak base) which is combination to produce acidic solutions.

1) Salt of **strong acid** and **strong base** forms ions which are **neutral**.

2) Salt of **weak acid** and **strong base** forms ions which are **basic**.

3) Salt of **strong acid** and **weak base** forms ions which are **acidic**.

4) Salt of **weak acid** and **weak base** forms ions which are either **neutral/ basic/acidic** depending on the relative strengths of the acid and the base.

**Which of the following salt solutions in water would produce ions which are acidic, basic or neutral ions?**

a) NaCl b) NaC2H3O2 c) NaHSO4 d) NH4Cl

 **Answer**: a) NaCl (**neutral**) b) NaC2H3O2 (**basic**) c) Na2SO4 (**neutral**) d) NH4Cl (**acidic**)

**Chemistry at a Glance: Acids and Acidic Solutions**

**10.12 Buffers**

A buffer solution are usually combinations of **weak acid/soluble-salt** or **weak base/ soluble-salt** combinations Buffers resists changes in pH when small quantities of an acid or an alkali are added to it.

**Acidic buffer solutions**

An acidic buffer (**weak base/soluble-salt**: **NH3/NH4Cl**) solution is simply one which has a pH less than 7. Acidic buffer solutions are commonly made from a weak acid and one of its salts - often a sodium salt.

Acidic buffer: weak acid/soluble-salt of weak base; NH3/NH4Cl

The main hydrolysis equilibrium that creates the acidity is:

 **H2O + NH4+  H3O+ + NH3**

 common ion with NH3

The common ion NH4+ will increase the H+ or [H3O+] decreasing pH of the solution. Similar to [HC2H3O2/NaC2H3O2] this equilibrium can buffer changes in pH.

**Basic buffer solutions**

Basic buffer solutions buffer (**weak acid/soluble-salt**: **HC2H3O/NaC2H3O**) solution is simply one which has a pH more than 7. Baisc buffer solutions are commonly made from a weak acid and one of its salts - often a sodium salt.

Regardless of acidic buffer or basic buffers, you can change the pH of the buffer solution by changing the ratio of acid to salt, or by choosing a different acid and one of its salts.

Basic buffer: weak acid/soluble-salt of weak acid; HC2H3O2/NaC2H3O2.

The main hydrolysis equilibrium that creates the basicity is:

H2O + C2H3O2-  HO- (basicty) + HC2H3O2 (acid dissociation )

 common ion with HC2H3O2

 The common ion C2H3O- will decrease the [HO-] increasing pH of the solution. This solution can act as a buffer for changes in pH. If the pH of the solution is increased (adding more HO-) this equilibrium will shift to left to maintain the original pH. If the pH of solution is decreased (adding H+) the equilibrium will shift to right to remove extra H+.

**10.13 The Henderson-Hasselbalch Equation**

The Henderson-Hasselbalch equation can be used to prepare buffer solutions and to estimate charges on ionizable species in solution, such as amino acid side chains in proteins. Caution must be exercised in using this equation because pH is sensitive to changes in temperature and salt concentration in the solution being prepared. This equation can then be rearranged to give the **Henderson-Hasselbalch** equation:



**Calculate the pH of a solution in which concentrations of NH4Cl is 0.9346 M and NH3 is 1.9072 M. (Ka(NH4Cl); 5.56 x 10-10)**

 [Base] = [NH3] = 1.9072 M and [Acid] = [NH4Cl] = 0.9346 M.

**pKa** = -**log** **Ka = -[**log (**5.56 x 10-10)] =-(-9.26) =9.26**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **pH** | = | **pKa** | **+ log** | **[Base]** |
| **[Acid]** |
| **pH** | = | **9.26** | **+ log** | **1.9072** |
| **0.9346** |
|  **pH** | = | **9.26** | **+**  | **0.3098**  | **= 9.57** |  |

**Chemistry at a Glance: Buffer Systems**

**10.14 Electrolytes**

**Electrolytes**

Certain substances that are called **electrolytes** produce ions when they dissolve in solution. Because these ions are free to move in the solution, the **solution conducts electricity strongly**. Ions can be produced in solution in either of two ways. Electrolytes can be either ionic compounds (i.e. sodium hydroxide, potassium nitrate) that dissolve in water, giving solutions of ions, or they may be covalent compounds that react with water and form ions in solution as a result. Strong electrolytes fall into three categories: **strong acids**, **strong bases**, and **salts**. (Salts are sometimes also called ionic compounds, but really strong bases are ionic compounds as well.)

Strong Electrolytes are strong acids such as HCl, HBr, HI, HNO3, HClO3, HClO4, and H2SO4

Strong bases NaOH, KOH, LiOH, Ba(OH)2, and Ca(OH)2

Soluble salts NaCl, KBr, and MgCl2, nitrate and acetate salts.

**Weak Electrolytes**

Weak electrolytes are compounds that partially ionize when they dissolve in water. Because there are fewer ions that are free to move in the solution, the solution **does not conduct electricity**. The weak electrolytes include weak acids and weak bases. HF, HC2H3O2 (acetic acid), H2CO3 (carbonic acid), H3PO4 (phosphoric acid), and weak bases such as NH3 (ammonia), C5H5N (pyridine)

 **Non-Electrolytes**

Non-electrolytes are compounds that don't ionize when they dissolve in water. Because there are no ions that are free to move in the solution, the solution **conducts electricity weakly**. Non-electrolytes are limited to **covalent compounds**. Many molecular compounds of carbon such as mathane CH4, benzene C6H6, ethanol C2H5OH, ether (C2H5)2O, and formaldehyde CH2O, are non-electrolytes.

**10.15 Acid-Base Titrations**

A **acid/base titration** is a procedure used in analytical chemistry to determine the amount or unknown concentration of a acid or base. In a titration one reagent, the **titrant**, is added to another slowly using a **burette** to a second solution in a conical flask. As it is added a chemical stoichiometric reaction occurs until one of the reagents is exhausted, and some process or device signals that this has occurred. The purpose of a titration is generally to determine the quantity or concentration of one of the reagents, that of the other being known beforehand. In any titration at the **end point** there must be a rapid quantitative reaction taking place as the titrant is added, and in acid-base titrations this is a stoichiometric **neutralization**.

**Acid-Base Indicators**

Many substances, including **litmus**, the one dye almost everyone associates with acids and bases, change color in response to acid or base.  The pigment in **red cabbage** is another natural substance very commonly used to show color change. **Phenolphthalein** is one of the most common indicators used for beginning chemistry, because its color change is very obvious which makes it easy to use.  There are many other indicators that change colors at different pH's, and so are useful for different purposes.  pH paper commonly contains a mixture of different indicators that change colors at different pH's.  The mixture is applied to paper, and then compared to a color chart to see what the pH of a solution is, approximately.

**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**