D\_\_\_\_\_ of Water

The most important solvent to the chemical world is \_\_\_\_\_ due to its \_\_\_\_\_ … it is easy to forget the important role of water compared to the role of the solute. For example, milk, orange juice, coffee, tea, soft drinks, etc. etc. all have water as the \_\_\_\_\_.



* \_\_ eq of water’s dissociation. 
* This means that water only \_\_\_\_\_ dissociates into ions.
* We can use the same principle as Ksp = [\_\_+][\_\_−] .
* Notice the [H20] is omitted. For aqueous solutions: [H+][OH−] = 1.0 × 10−\_\_
* The product of the concentrations of the hydrogen ions and the hydroxide ions in water is called the ion-product constant for water (*Kw*).

K\_\_ = [H+][OH−] = 1.0 × 10−\_\_

* In pure water at 25 ⁰C, the concentration of hydrogen ions ([\_\_\_+]) is only 1 × 10−\_\_*M*.
* The concentration of hydroxide ions ([(\_\_)−]) is also 1 × 10−\_\_*M* because the numbers of H+ and (OH)− ions are \_\_\_\_\_ in \_\_\_\_\_ water.

The A\_\_\_\_\_ Concept

* Arrhenius \_\_\_\_\_: a substance that when dissolved in water increases \_\_\_+ (or H+) concentration.
* Arrhenius \_\_\_\_\_: a substance that when dissolved in water increases (\_\_\_)− concentration.

**A**\_\_\_\_\_  **Solutions**

* A substance that \_\_\_\_\_ the concentration of \_\_\_+ (H3O+) ions in solution.
* H+: \_\_\_\_\_ ion … H3O+: \_\_\_\_\_ ion

kw = [H+][(OH)-] = 10-14

* In acidic solutions, the hydrogen-ion concentration is \_\_\_\_\_ than the hydroxide-ion concentration. [\_\_+] > [(\_\_\_)-]
	+ The [\_\_+] is \_\_\_\_\_ than 1 × 10−\_\_\_*M*.
	+ The [(\_\_\_)-] is \_\_\_\_\_ than 1 × 10−\_\_\_*M*.
* Examples of Acids
	+ \_\_\_\_\_ acid (HCl) …stomach acid
	+ \_\_\_\_\_ acid (H2SO4) … battery acid
	+ \_\_\_\_\_ acid (HCH3COO) … vinegar
	+ Nitric acid (HNO3) … explosives, fertilizer
	+ Carbonic acid (H2CO3) … soda water
* Not all compounds that contain hydrogen are acids.
	+ Only a hydrogen that is bonded to a very \_\_\_\_\_ element can be released as an ion. Such bonds are highly \_\_\_\_\_.
* In an \_\_\_\_\_ solution, hydrogen ions are not present. Instead, the hydrogen ions are joined to water molecules as \_\_\_\_\_ ions.
* A hydronium ion (\_\_3\_\_\_+) is the ion that forms when a water molecule \_\_\_\_\_ a hydrogen ion [H+].
* Chemical and Physical Properties of Acids
	+ Have a \_\_\_\_\_ taste
	+ Turn \_\_\_\_\_ litmus paper \_\_\_\_\_
	+ Turn methyl orange solution red
	+ React with carbonates and bicarbonates
	+ Are \_\_\_\_\_ by bases
	+ React with metals

**B**\_\_\_\_\_ **Solutions**

* A substance that \_\_\_\_\_ the concentration of (\_\_\_)- ions in solution.
* In basic solutions, the \_\_\_\_\_ -ion concentration is greater than the hydrogen-ion concentration. [(\_\_\_)-] > [\_\_+]
	+ The [(\_\_\_)-] is \_\_\_\_\_ than 1 × 10−\_\_\_*M*.
	+ The [H+] is less than 1 × 10−7*M*.
* Basic solutions are also known as \_\_\_\_\_ solutions.
* Examples of Bases
	+ \_\_\_\_\_: Al(OH)3
	+ Drain Cleaner: Sodium hydroxide (NaOH)
	+ Fertilizer/soap: Potassium hydroxide (KOH)
	+ Limewater: Calcium Hydroxide Ca(OH)2
* Ionic compounds that \_\_\_\_\_ into (\_\_\_)- ions in an aqueous solution.
* Chemical and Physical Properties of Bases
	+ Taste \_\_\_\_\_
	+ Are \_\_\_\_\_
	+ Turn \_\_\_\_\_ litmus paper \_\_\_\_\_
	+ Turn colorless phenolphthalein \_\_\_\_\_
	+ Neutralize \_\_\_\_\_
	+ May react with some metal salts
	+ May react with carbon dioxide

Arrhenius Acids and Bases SUMMARY

Arrhenius’ theory defines acids, bases, and salts based on the use of \_\_\_\_\_. He used observable properties:

* \_\_\_\_\_ taste sour; \_\_\_\_\_ taste bitter
* acids/bases conduct electricity (\_\_\_\_\_)
* strong acids & bases are \_\_\_\_\_
* \_\_\_\_\_ turns red in an acid and blue in a base
* indicators change color based on \_\_\_

Arrhenius also used the \_\_\_\_-product of water concentrations to define acids [\_\_+], bases [\_\_\_-] and salts [\_\_+\_\_\_-].

**pH**

* Expressing hydrogen-ion concentration in molarity (Moles/liter) is often not practical or easy.
* So, rather than focus on [H+], Sorenson suggested that concentration values be expressed as a power of 10 that arises when hydrogen ion concentrations ([H+]) are put in scientific notation.
* pH measures [\_\_\_+] on a \_\_\_\_\_ scale.
	+ The pH scale works in powers of \_\_\_\_\_, so each jump in number is a multiple of 10 in concentration.
	+ For example, a pH of \_\_\_ is 10 times more acidic than a pH of \_\_\_.

pH measures [H+] on a logarithmic scale.

* Corresponds to the variation of [H+] related to k\_\_\_.
* [\_\_\_+] ranges from 10-1 to 10-14 M
* Avoids the use of small fractions and negative exponents
* Indicates “\_\_\_\_\_ strength” in a solution

pH = -\_\_\_10[\_\_+] or pH = -log10[H3O+]

\_\_\_\_\_\_\_

\_\_\_\_\_\_\_

\_\_\_\_\_\_\_



**The pH Scale**

\_\_\_\_\_: pH \_\_\_\_\_ 7 [H+] > 1 × 10−7*M* \_\_\_\_\_ pH indicates greater acidity

\_\_\_\_\_: pH \_\_\_\_\_ 7 [OH-] < 1 × 10−7*M* \_\_\_\_\_ pH indicates greater alkalinity

\_\_\_\_\_: pH = 7 [H+] = 10−\_\_\_*M* [OH-] = 10−\_\_\_*M* [\_\_+] = [\_\_\_-]

pOH … the “potential” of (OH)- in solution

pOH is a measure of the \_\_\_\_\_ ion concentration ([OH–]) in an aqueous solution

p\_\_\_\_ = –\_\_\_10[\_\_\_–]

* pH and pOH are \_\_\_\_\_ related
* As pH \_\_\_\_\_, pOH \_\_\_\_\_
* Acids have \_\_\_\_\_ pH, \_\_\_\_\_ pOH
* \_\_\_\_\_ have high pH, low pOH

pH and pOH pH + pOH = \_\_\_

* pH = 5 = -log (10-5) … pOH = 9 = -log (10-9)
* pH = 5 = -~~log (10~~-5) … pOH = 9 = -~~log (10~~-9)
* 5 + 9 = 14

\_\_\_\_\_ concentrations are described by: kw = [H+][OH-] = 10-14 p\_\_ + p\_\_\_ = 14

Measuring pH

* A pH \_\_\_\_\_ can be easier to use than liquid indicators or indicator strips.
* The pH reading is visible in a display window on the meter.
* The color and cloudiness of the solution do not affect the accuracy of the pH value obtained.

**B**\_\_\_\_\_ **-L**\_\_\_\_\_  **Acids and Bases**

* A more general definition of acids and bases
* Extends the Arrhenius theory
* Includes other solvents beside \_\_\_\_\_

An acid-base reaction in which there is a \_\_\_\_\_ (H+) transfer from one species to another.

* The species which \_\_\_\_\_ the proton is a B-L \_\_\_\_\_.
* The species which \_\_\_\_\_ the proton is a B-L \_\_\_\_\_.



For the f\_\_\_\_\_ reaction:

* Water donates the H+ … B-L \_\_\_\_\_
* Ammonia accepts H+ … B-L \_\_\_\_\_

Bronsted-Lowry proposed that the \_\_\_\_\_ reaction of an acid/base reaction is also an acid/base reaction.

For the r\_\_\_\_\_ reaction:

* + Ammonium donates the H+ … B-L \_\_\_\_\_
	+ Hydroxide accepts H+ … B-L \_\_\_\_\_

Brønsted-Lowry \_\_\_\_\_ Acid-Base Pairs

* According to the Bronsted-Lowry concept, in every reaction, an acid reacts with a base to form another acid and another base which are \_\_\_\_\_ of the \_\_\_\_\_ acid and base.
* \_\_\_\_\_ means that the acid-base pairs have the same derivation and are distinguished by the presence or absence of a \_\_\_\_\_ (\_\_+).



* NH4+ is the \_\_\_\_\_ acid of the base NH3.
* OH– is the \_\_\_\_\_ base of the acid H2O.

A\_\_\_\_\_ Substances

* A substance that can act as either an \_\_\_\_\_ or a \_\_\_\_\_ is said to be amphoteric.
* \_\_\_\_\_ is amphoteric. In the reaction with hydrochloric acid, water \_\_\_\_\_ a proton and is therefore a \_\_\_\_\_.
* In the reaction with ammonia, water \_\_\_\_\_ a proton and is therefore an \_\_\_\_\_.

Common amphoteric substances include: H2O, HSO4-, HCO3-, HS- [Reference Table L]

  HCO3- (aq) + OH- (aq) 🡪 H2O(aq) + CO3-2 (aq)

 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

  HF (aq) + HCO3- (aq) 🡪 H2CO3(aq) + F- (aq)

 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Notice that bicarbonate can act as an acid (\_\_\_\_\_ an H+) or a base (\_\_\_\_\_ the H+).

**L**\_\_\_\_\_  **Acids and Bases (BONUS)**

* The Arrhenius theory is the most restricted “acid-base” explanation since an acid must possess an H+ ion and a base must possess an OH- ion.
* The Bronsted-Lowry concept extended the acid-base understanding, but can only be applied to reactions involving a proton transfer [an ionizable Hydrogen atom]
* A **Lewis acid** is a substance that can \_\_\_\_\_ a pair of \_\_\_\_\_ to form a covalent bond.
* A **Lewis** \_\_\_\_\_ is a substance that can \_\_\_\_\_ a pair of electrons to form a covalent bond.
* Note that acids now “accept” rather than “donate” since electrons are opposite protons (H+).

|  |
| --- |
| Acid-Base Definitions |
| Type | Acid Description | Base Description |
| Arrhenius |  |  |
| Bronsted-Lowry |  |  |
| Lewis |  |  |

* C\_\_\_\_\_ Covalent Bonds
* Any reaction that leads to the formation of a coordinate covalent bond (both e- in the bond are furnished by the same species) is an acid-base reaction.

Lewis Bases

a. All Bronsted-Lowry bases are included, plus a few more.

b. All the bases must possess an unshared pair of \_\_\_\_\_ which can be \_\_\_\_\_ to an acid



* When a Lewis acid reacts with a Lewis base, the \_\_\_\_\_ donates a pair of electrons and the \_\_\_\_\_ accepts the donated pair.

Summary: Acid – Base Theories

Arrhenius

* \_\_\_\_\_ 🡪 H+ ions \_\_\_\_\_ scale
* \_\_\_\_\_ 🡪 (OH)- ions

Bronsted-Lowry

* Acid 🡪 \_\_\_\_\_ H+ ions \_\_\_\_\_ acid-base pairs
* Base 🡪 \_\_\_\_\_ H+ ions

Lewis

* Acid 🡪 \_\_\_\_\_ electrons most inclusive definition
* Base 🡪 \_\_\_\_\_ electrons

19B What makes one acid safer than another?

Lemon juice, which contains citric acid, has a pH of about 2.3 which is relatively strong. Yet, most people love lemonade! Battery acid is very dangerous (sulfuric acid: H2SO4 (aq)).

**I**\_\_\_\_\_

* The foundational concept of Acids and Bases is their ability to \_\_\_\_\_ in solution.
* \_\_\_\_\_: acids and bases can conduct \_\_\_\_\_ due to the presence of free moving \_\_\_\_\_ which exist because of ionization.
	+ i.e. HCl (aq) is a \_\_\_\_\_ electrolyte (acid)
	+ NaOH (aq) is a \_\_\_\_\_ electrolyte (base)
	+ CH3COOH (aq) is a \_\_\_\_\_ electrolyte (acid)
	+ Pure water does \_\_\_\_\_ conduct electricity

**Strong vs. Weak Acids**

* Acids and bases are classified as \_\_\_\_\_ or \_\_\_\_\_ based on the \_\_\_\_\_ to which they ionize in water.
* In general, a \_\_\_\_\_ acid \_\_\_\_\_ ionizes in aqueous solution. Hydrochloric and sulfuric acid are examples of strong acids.
	+ HCl(*g*) + H2O(*l*) → H3O+(*aq*) + Cl–(*aq*)
	+ H2SO4(s) + 2H2O(*l*) → 2H3O+(*aq*) + SO4–2(*aq*)
* A \_\_\_\_\_ acid ionizes only \_\_\_\_\_ in aqueous solution.
* The ionization of acetic acid (CH3COOH), a typical weak acid, is not complete.

D\_\_\_\_\_ of a \_\_\_\_\_ acid HX (aq) 🡪 H+(aq) + X- (aq)

* Ionize \_\_\_\_\_ (\_\_\_ %) … no reactant remains
* Goes to “\_\_\_\_\_” (FORWARD rxn favored)
* 100% dissociation, strong \_\_\_\_\_



Dissociation of a \_\_\_\_\_ acid HX (aq) H+(aq) + X- (aq)

* \_\_\_\_\_ ions produced, weak electrolyte
* Favors the \_\_\_\_\_ reaction
* \_\_\_\_\_ % dissociation

Acid Dissociation Constant

* The equilibrium-constant expression for any acid: HX (aq) ↔ H+(aq) + X- (aq)
* Ka = [H+][ X-] / [HX]
* The acid dissociation constant (*Ka*) indicates the \_\_\_\_\_ of \_\_\_\_\_ produced and which reaction (forward/reverse) is \_\_\_\_\_.
* Ka = 5.7 x 10-10 (very small) 🡪 \_\_\_\_\_ acid, \_\_\_\_\_ ions, \_\_\_\_\_ reaction
* Ka = 2.4 x 104 (large) 🡪 \_\_\_\_\_ acid, \_\_\_\_\_ ions, \_\_\_\_\_ rxn

\_\_\_\_\_ acid: is an acid which fully \_\_\_\_\_ (ionizes) in solution. \_\_\_\_\_ acids: do not fully \_\_\_\_\_ (dissociate) in solution.

List some strong acids:

Base Dissociation Constant

* The equilibrium-constant expression for any base:
* NH3 (aq) + H2O  NH4+(aq) + OH- (aq)
* Kb = [NH4+][ OH -] / [NH3]
* Kb = 1.8 x 10-4
* The Base dissociation constant (*Kb*) indicates the amount of ions produced and which reaction (forward/reverse) is favored.
	+ Kb = 1.8 x 10-4 (small) 🡪 \_\_\_\_\_ base, \_\_\_\_\_ ions, \_\_\_\_\_ reaction
	+ Kb = 7.2 x 106 (large) 🡪 \_\_\_\_\_ base, \_\_\_\_\_ ions, \_\_\_\_\_ rxn

Strong vs. Weak Bases

* Strong bases: like sodium hydroxide or potassium hydroxide which are fully ionized.
* A strong base is a compound that is \_\_\_\_\_ % split up into metal ions and \_\_\_\_\_ ions in solution.
* All \_\_\_\_\_ bases will contain a hydroxide group.
* \_\_\_\_\_ bases: do not \_\_\_\_\_ fully into hydroxide ions in solution. E.g. Ammonia NH3

**Strength vs. Concentration**

* The strength of an acid or base is related the degree of dissociation in water to produce ions.
* The \_\_\_\_\_ tells you about how much of the original acid (solute) is \_\_\_\_\_ in the solution.
* It is entirely possible to have a concentrated solution of a weak acid, or a dilute solution of a strong acid or base.
* Concentrated and dilute indicate how much of an acid or base is dissolved in solution and refer to the number of moles of the acid or base in a given volume (solute dissoved).
* The words strong and weak refer to the extent of \_\_\_\_\_ or dissociation of an acid or base (Ka or Kb)… \_\_\_\_\_ produced



**Acid-Base I**\_\_\_\_\_

* Either acid-base indicators or pH meters can be used to measure pH.
* An indicator (HIn) is an acid or a base that dissociates in a \_\_\_\_\_ pH range. Indicators work because their acid form and base form have different \_\_\_\_\_ in solution.



* The acid form of the indicator (HIn) is dominant at low pH and high [H+]. The base form (In−) is dominant at high pH and high [OH−].
* \_\_\_\_\_ turn blue \_\_\_\_\_ paper red. Bases turn \_\_\_\_\_ litmus paper \_\_\_\_\_.
* Indicator strips were developed to cover a wide range of pH values (e.g. pH or \_\_\_\_\_ paper).

**Neutralization**

* Neutralization is the complete reaction of a \_\_\_\_\_ acid with a \_\_\_\_\_ base to produce a \_\_\_\_\_ solution.
* A neutral solution has a pH of \_\_\_ because it contains \_\_\_\_\_ concentrations of \_\_\_\_\_ ions (H+) and \_\_\_\_\_ ions (OH-). [H+] = [OH-]
* The general equation for an acid base neutralization:

\_\_\_\_\_ + \_\_\_\_\_ → \_\_\_\_\_ + \_\_\_\_\_

Salt = the anion from an acid and the cation from a base

* A reaction between an acid and a base will go to completion when the solutions contain equal numbers of hydrogen ions and hydroxide ions. This can occur between \_\_\_\_\_ acids and \_\_\_\_\_ bases because both dissociate \_\_\_\_\_.
* The balanced equation provides the correct \_\_\_\_\_ of moles of acid to \_\_\_\_\_ of base to produce the neutralization.
* For hydrochloric acid and sodium hydroxide, the mole ratio is 1:1 so equal quantities of the acid and base neutralize. HCl(*aq*) + NaOH(*aq*) → NaCl(*aq*) + H2O(*l*)

Each of the following incomplete reactions are neutralization reactions. What is the formula of the salt formed?

H2SO4 + NH4OH →

NaOH + H2CO3 →

HNO3 + KOH →

H3PO4 + NaOH →

H\_\_\_\_\_ of a Salt

Show the acid and base used to form each salt in a neutralizing reaction.

K2SO4 →

NaBr →



MgCl2 →

(NH4)3PO4 →

Hydrolysis: \_\_\_\_\_ + \_\_\_\_\_ 🡪 \_\_\_\_\_ + \_\_\_\_\_

\_\_\_\_\_: Acid + Base 🡪 Salt + Water

**T**\_\_\_\_\_

* Chemists use a neutralization reaction to determine the \_\_\_\_\_ of an acid or base.
* The process of adding a measured amount of a solution of \_\_\_\_\_ concentration to a solution of \_\_\_\_\_ concentration is called a titration.
* A measured \_\_\_\_\_ of an acid solution of unknown concentration is added to a flask.
* Several drops of an \_\_\_\_\_ (phenolphthalein) are added to the solution.
* Measured volumes of a base of known concentration are mixed into the acid until the indicator shows a \_\_\_\_\_ \_\_\_\_\_ change.
* Used to determine the composition or concentration of acids and bases in solution.
* Used to \_\_\_\_\_ mixtures. Used in nearly every laboratory: university, power plants, breweries, etc.
* Titration is one of the most \_\_\_\_\_ methods of chemical analysis.

E\_\_\_\_\_ Point

* The point at which \_\_\_\_\_ occurs.
* The point at which moles of one reactant (acid) have been added in stoichiometric quantity to react completely with the moles of the other reactant (base).
* The indicator chosen for this titration changes color at or near the pH of the equivalence point.
* The point at which the indicator forms a \_\_\_\_\_ color change is the \_\_\_\_\_ point of the titration.

Titrations can be done by adding strong acid to strong base or strong base to strong acid.

 Adding \_\_\_\_\_ to \_\_\_\_\_ Adding \_\_\_\_\_ to \_\_\_\_\_



The \_\_\_ changes gradually until near the equivalence point where equal moles of acid and base react completely (pH = \_\_\_). At the equivalence point pH changes \_\_\_\_\_.

Titrations can also involve a \_\_\_\_\_ acid with a \_\_\_\_\_ base or a \_\_\_\_\_ acid with a \_\_\_\_\_ base.

 Adding \_\_\_\_\_\_ Acid to \_\_\_\_\_\_ Base Adding \_\_\_\_\_\_\_ Base to \_\_\_\_\_\_ Acid



 LEFT: \_\_\_\_\_ (pH \_\_\_\_\_) RIGHT: \_\_\_\_\_ (pH \_\_\_\_\_).

 Adding \_\_\_\_\_\_ Acid to \_\_\_\_\_\_ Base Adding \_\_\_\_\_\_\_ Base to \_\_\_\_\_\_ Acid



The equivalence point is \_\_\_ longer at a pH of 7 as with the \_\_\_\_\_ acids and \_\_\_\_\_ bases. At the equivalence point:

 LEFT: \_\_\_\_\_ (pH \_\_\_\_\_) RIGHT: \_\_\_\_\_ (pH \_\_\_\_\_).

Salt Hydrolysis

* Titration depends on neutralization: acid + base 🡪 SALT + water
* The strength of the acids & bases used will determine the type of salt solution produced: acidic, neutral, or basic.
* Ammonium chloride, NH4Cl(*aq*), is acidic (pH of about 5.3).
	+ NH4Cl(*aq*) 🡪 NH4(OH)(*aq*), + HCl(*aq*)
	+ \_\_\_\_\_ \_\_\_\_\_ base \_\_\_\_\_ acid
* Sodium chloride, NaCl(*aq*), is neutral (pH of 7).
	+ NaCl(*aq*) 🡪 HCl(*aq*), + NaOH(*aq*)
	+ \_\_\_\_\_ \_\_\_\_\_ acid \_\_\_\_\_ base
* Sodium acetate, NaCH2COOH(*aq*), is basic (pH of about 8.7).
	+ NaCH2COOH(*aq*) 🡪 CH3COOH(*aq*) + NaOH(*aq*)
	+ \_\_\_\_\_ \_\_\_\_\_ acid \_\_\_\_\_ base
* When a STRONG acid is titrated with a WEAK base, an \_\_\_\_\_ salt solution is produced. \_\_\_\_\_ acid + \_\_\_\_\_ base → Acidic solution
* When a STRONG acid is titrated with a STRONG base, a \_\_\_\_\_ salt solution is produced. \_\_\_\_\_ acid + \_\_\_\_\_ base → Neutral solution
* When a WEAK acid is titrated with a STRONG base, a \_\_\_\_\_ salt solution is produced. \_\_\_\_\_ acid + \_\_\_\_\_ base → Basic solution

Calculating in Titrations

Calculations involving titration is identical to “\_\_\_\_\_” in molar concentrations (M1V1 = M2V2): Titration is based on the \_\_\_\_\_ and \_\_\_\_\_ of the acids and bases involved.

MbVb = MaVa

This calculation can be applied to all concentrations and strengths of the acids and bases (include the \_\_\_\_\_ factor of a balanced chemical reaction).

**B**\_\_\_\_\_

A buffer is an [aqueous solution](http://chemistry.about.com/od/chemistryglossary/a/aqueoussoldef.htm) that has a highly \_\_\_\_\_ [pH](http://chemistry.about.com/od/chemistryglossary/a/phdef.htm).

* If you add acid or base to a buffered solution, its pH will not change significantly.
* Buffers resist \_\_\_\_\_ in \_\_\_.

A buffer is made by mixing a large volume of a \_\_\_\_\_ acid or weak base together with its \_\_\_\_\_.

* A weak acid and its conjugate base (salt) can remain in solution without neutralizing each other.
* A weak base and its conjugate acid (salt) do not neutralize.

Buffers cause \_\_\_\_\_ reactions that will not have much effect on the overall \_\_\_ of the buffer solution.

* When hydrogen ions are added to a buffer, they will be neutralized by the base in the buffer.
* Hydroxide ions will be neutralized by the acid.

How Buffers Work

* A \_\_\_\_\_ solution is better able to \_\_\_\_\_ drastic changes in pH than pure \_\_\_\_\_ since water dissociates so little (kw = 10-14).
* A buffer solution contains one component that can react with \_\_\_\_\_ ions (hydrogen-ion acceptor) and one that can react with \_\_\_\_\_ ions (hydrogen-ion donor).
* Buffer systems often include \_\_\_\_\_ substances.

pH too high?

* pH rises because of excess \_\_\_\_\_ in the blood which reduces the \_\_\_\_\_ ions.
* 2 H2O + CO2 ↔ H2CO3 + (OH)- → **H3O+** + HCO3-
* Therefore, the \_\_\_\_\_ generates H3O+ by having carbonic acid dissociate into hydronium ions to \_\_\_\_\_ the base, favoring the \_\_\_\_\_ reaction.

pH too low?

* pH drops because of excess \_\_\_\_\_ in the blood which \_\_\_\_\_ the hydronium ions.
* 2 H2O + CO2 ↔ H2CO3 + H2O ← H3O+ + HCO3-
* Therefore, the \_\_\_\_\_ reduces H3O+ by having bicarbonate ions accept hydronium ions to \_\_\_\_\_ the \_\_\_\_\_ (\_\_\_\_\_ rxn favored).

Buffer solutions have their limits.

* Adding too much acid or base will exceed the buffer capacity of a solution.
* The buffer \_\_\_\_\_ is the amount of acid or base that can be added to a buffer solution before a significant change in pH occurs.
* Buffers work best in environments with specific pH ranges.
	+ For \_\_\_\_\_ solutions, use a buffer with higher pH range (10.01).
	+ For \_\_\_\_\_ solutions, use a buffer with lower pH range (4.01).
	+ For \_\_\_\_\_ solutions, use a buffer with neutral pH range (7.00).