Acids and Bases

PreAP Chemistry

Chap. 14

Introduction to Acids and Bases

- Acids and bases are both aqueous solutions. A substance (solid or liquid) may be identified as an “acid” or “base” BUT, does not have acidic or basic properties until it is dissolved in water.
- Acids produce hydrogen ions (H^+) in solution while bases produce hydroxide ions (OH^-).
- Most acids are ionic compounds with hydrogen as the positive ion. Bases usually have hydroxide as the negative ion.

Introduction to Acids and Bases

- The self ionization of water is what enables acids and bases to work. Water is the only substance we know of that will react with itself.

$$H_2O + H_2O \rightarrow H_30^+ + OH^-$$

hydronium ***hydroxide***

Pure water has as many hydronium ions in solution as it does hydroxide ions. This produces a "neutral" solution.

Introduction to Acids and Bases

- Acids and bases both produce electrolytic solutions because they have ions in solution.
- Acids have a sour taste (like vinegar), turn litmus paper red, and react with active metals, carbonates, and bases.
- Bases have a bitter taste (pure cocoa), a slippery feel, turn litmus blue, and react with acids.

Acids and Base Theories

Arrhenius Acids and Bases

"An acid is a substance that contains hydrogen and ionizes to produce hydrogen ions in aqueous solutions. A base is a substance that contains hydroxide in the formula and produces a hydroxide ion in solution."

$$HCl_{(g)} \rightarrow H^{+}_{(aq)} + Cl^{-}_{(aq)}$$

$$NaOH_{(s)} \rightarrow Na^{+}_{(aq)} + OH^{-}_{(aq)}$$

This model works well but there are exceptions. Ammonia (NH_3) is a base but does not have hydroxide in the formula.

Acids and Base Theories

Bronsted-Lowry Model

This model focuses on what happens to hydrogen ions (protons) in reactions.

Acids are defined as proton donors and bases are proton acceptors. Consider the following general equation:

$$\underset{\text{acid}}{HX_{(aq)}} + \underset{\text{base}}{H_2O_{(l)}} \rightarrow H_3O^+_{(aq)} + X^-_{(aq)}$$

In this case, the water accepts a proton to become a hydronium ion. By the Bronsted-Lowry definition, it is a base. The HX donated the proton, so it must be an acid.

Acids and Base Theories

Bronsted-Lowry Model

$$HX_{(aq)} + H_2O_{(l)} \leftrightarrow H_3O^+_{(aq)} + X^-_{(aq)}$$

HX (aq)	H2O (l)	H3O+ (aq)	X- (aq)
acid	base	conjugate acid	conjugate base

Most acid/base reactions are reversible. In the reverse direction, the hydronium ion would donate its extra proton (H+) to X. In this case, it would behave like an acid and X would be acting like a base. These are called conjugate acids and bases and form pairs with their similar counterpart on the reactant side of the equation.

So, H_2O and H_3O^+ are a conjugate acid-base pair.

Acids and Base Theories

Bronsted-Lowry Model (examples)

HF + H_2O ←→ H_3O^+ + F^-

acid | base | conjugate acid | conjugate base

$NH_{3\ (aq)}$ + H_2O ←→ $NH_4^+{}_{(aq)}$ + $OH^-{}_{(aq)}$

base | acid | conjugate acid | conjugate base

Ionization of Polyprotic Acids

Diprotic acids require two steps to totally ionize

$H_2SO_{4\,(aq)} \rightarrow H^+_{(aq)} + HSO_4^-{}_{(aq)}$

$HSO_4^-{}_{(aq)} \rightarrow H^+_{(aq)} + SO_4^{2-}{}_{(aq)}$

Triprotic acids require three steps . . .

$H_3PO_{4\,(aq)} \rightarrow H^+_{(aq)} + H_2PO_4^-{}_{(aq)}$

$H_2PO_4^-{}_{(aq)} \rightarrow H^+_{(aq)} + HPO_4^{2-}{}_{(aq)}$

$HPO_4^{2-}{}_{(aq)} \rightarrow H^+_{(aq)} + PO_4^{3-}{}_{(aq)}$

Anhydrides

Anhydrides are compounds that do not contain hydrogen yet act as an acid or base in solution.

- An example of an *acidic anhydride* is carbon dioxide. CO2 will react with water to produce carbonic acid.

$$CO_{2(g)} + H_2O_{(l)} \rightarrow H_2CO_{3(aq)}$$

- Most metal oxides are examples of *basic anhydrides*. They react with water to produce hydroxide compounds which then dissociate to form hydroxide ions in solution.

$$CaO_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2\,(aq)}$$

$$Ca(OH)_{2\,(aq)} \rightarrow Ca^{2+}_{(aq)} + 2\,OH^{-}_{(aq)}$$

Strength of Acids and Bases

"Strength" is a reference to the degree of dissociation of an acid or base. A strong acid or base completely dissociates in water (like HCl), leaving only H^+ and Cl^- ions.

A weak acid (like acetic) will partially dissociate in water, leaving a combination of dissociated ions and the original compound.

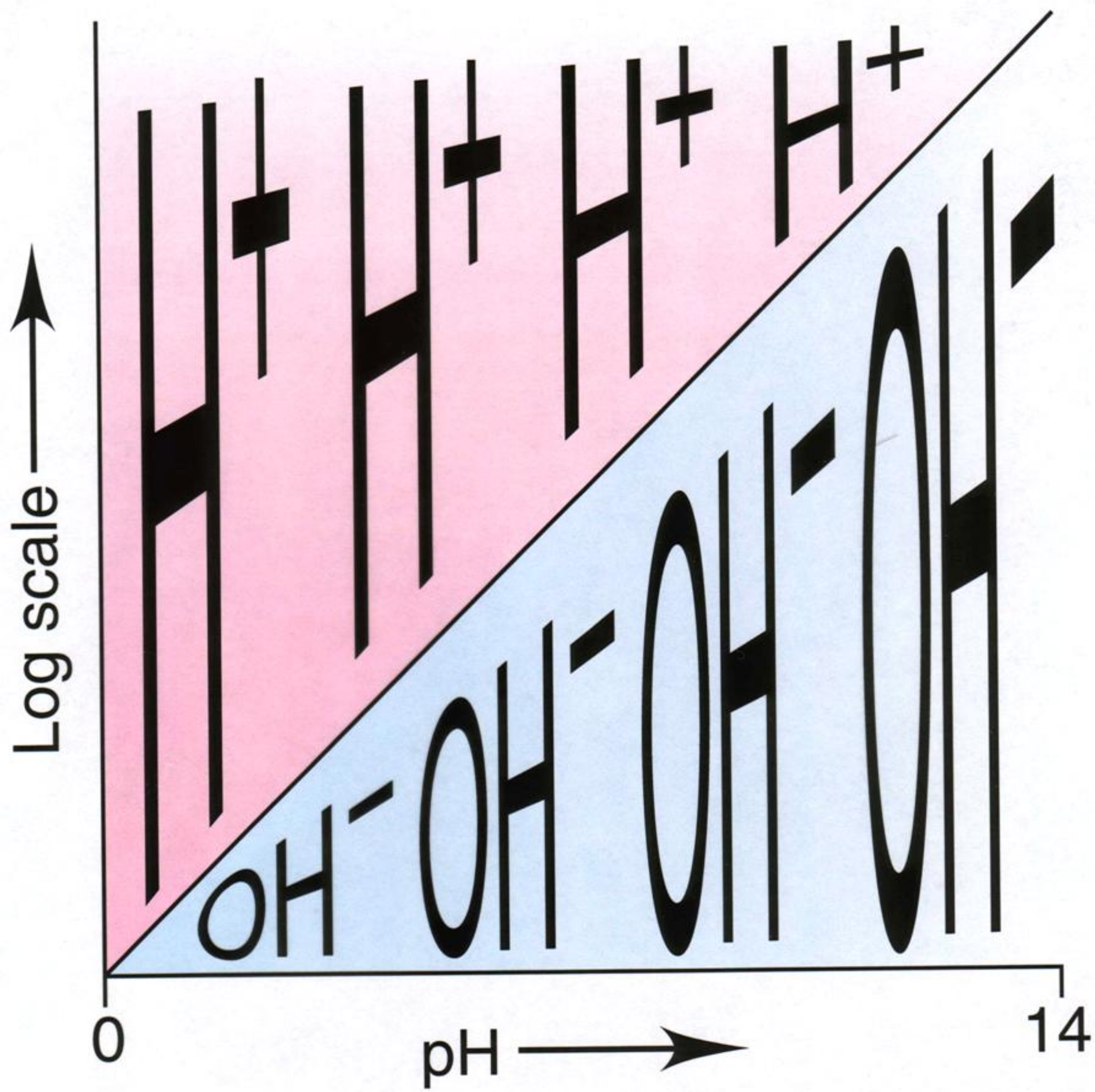
Strength has nothing to do with concentration! You can have a dilute, strong acid (muratic acid put in pools) or a concentrated, weak acid (grapefruit juice).

The pH Scale

- Remember that as water self-ionizes, it produces equal amounts of H_3O^+ ions and OH^- ions in solution.
- Pure water contains 0.0000001 moles of H_3O^+ ions in one liter of water, or 1×10^{-7} mol (and an equal number of OH^- ions).
- More acidic solutions contain more H_3O^+ ions and less OH^-. Lemon juice, which is a dilute solution of citric acid may contain 1×10^{-3} moles of H_3O^+ and only 1×10^{-11} moles of OH^-

pH

pH = -log [H^{1+}]



Kelter, Carr, Scott, Chemistry A World of Choices 1999, page 285

The pH Scale

- The pH scale was developed as a way of comparing the relative concentrations of H_3O^+ and OH^- ions in solution.
- The basic equation is pH = $-\log[H_3O^+]$

So, if pure water has a hydronium ion concentration of 1×10^{-7}, it has a pH of 7

Lemon juice, with a 1×10^{-3} concentration would have a pH of 3

pH of Common Substance

More basic

NaOH, 0.1 *M*
Household bleach
Household ammonia

Lime water
Milk of magnesia

Borax

Baking soda
Egg white, seawater
Human blood, tears
Milk
Saliva
Rain

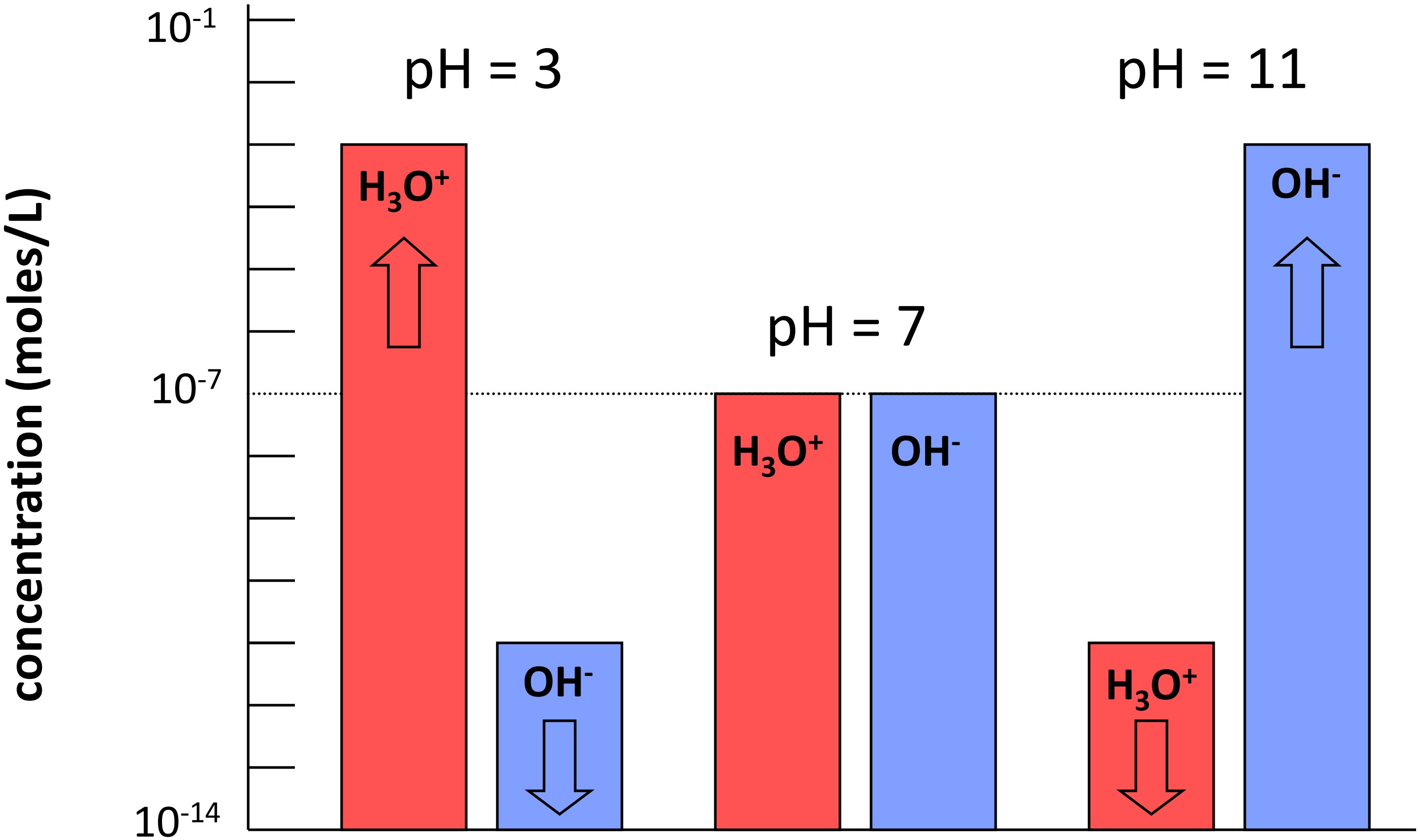
Black coffee
Banana
Tomatoes
Wine
Cola, vinegar
Lemon juice

Gastric juice

More acidic

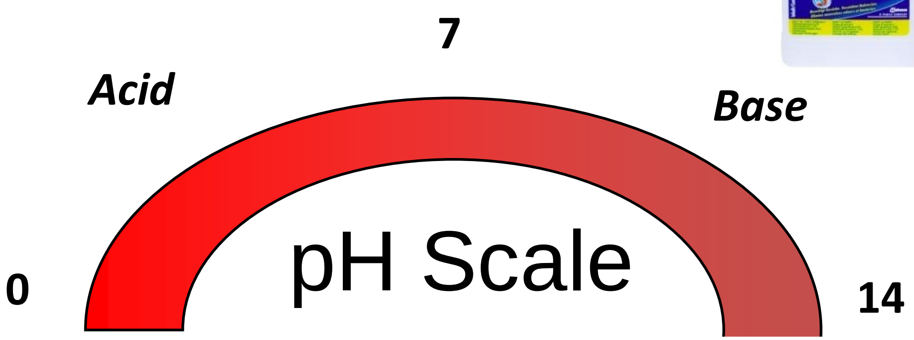
pH	$[H^{1+}]$	$[OH^{1-}]$	pOH
14	1×10^{-14}	1×10^{-0}	0
13	1×10^{-13}	1×10^{-1}	1
12	1×10^{-12}	1×10^{-2}	2
11	1×10^{-11}	1×10^{-3}	3
10	1×10^{-10}	1×10^{-4}	4
9	1×10^{-9}	1×10^{-5}	5
8	1×10^{-8}	1×10^{-6}	6
7	1×10^{-7}	1×10^{-7}	7
6	1×10^{-6}	1×10^{-8}	8
5	1×10^{-5}	1×10^{-9}	9
4	1×10^{-4}	1×10^{-10}	10
3	1×10^{-3}	1×10^{-11}	11
2	1×10^{-2}	1×10^{-12}	12
1	1×10^{-1}	1×10^{-13}	13
0	1×10^{0}	1×10^{-14}	14

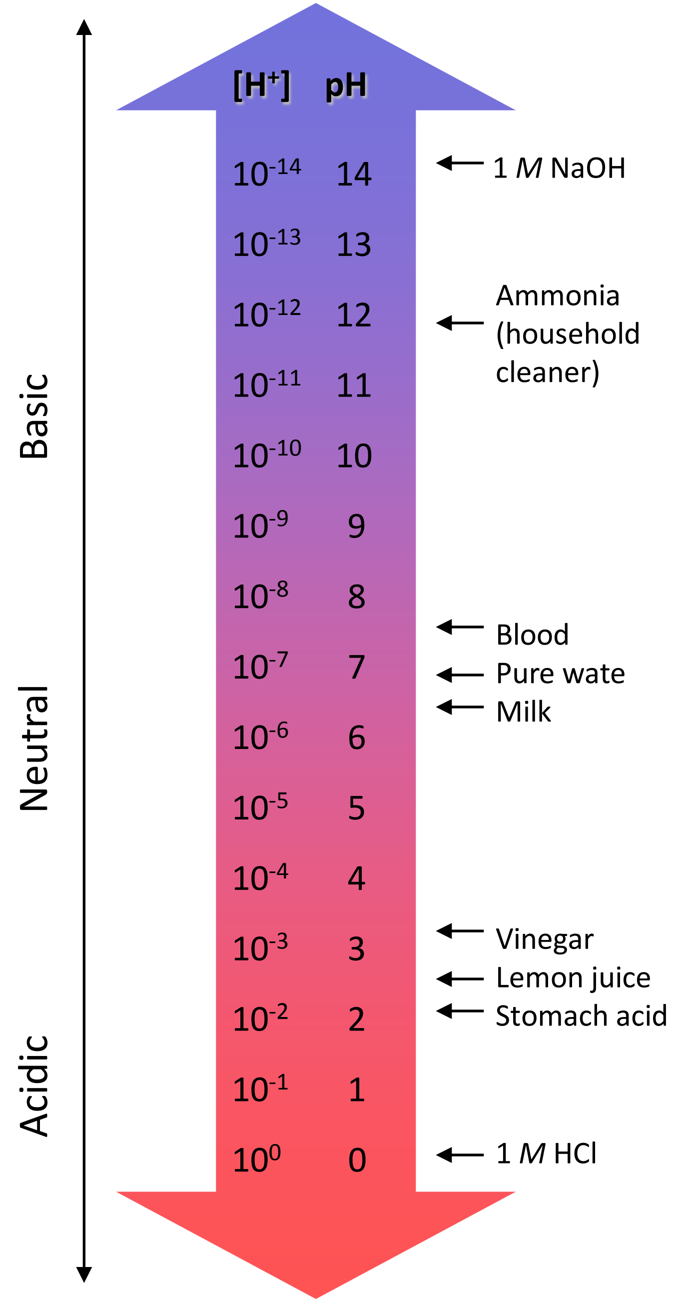
Acid – Base Concentrations



acidic solution	neutral solution	basic solution
$[H_3O^+] > [OH^-]$	$[H_3O^+] = [OH^-]$	$[H_3O^+] < [OH^-]$

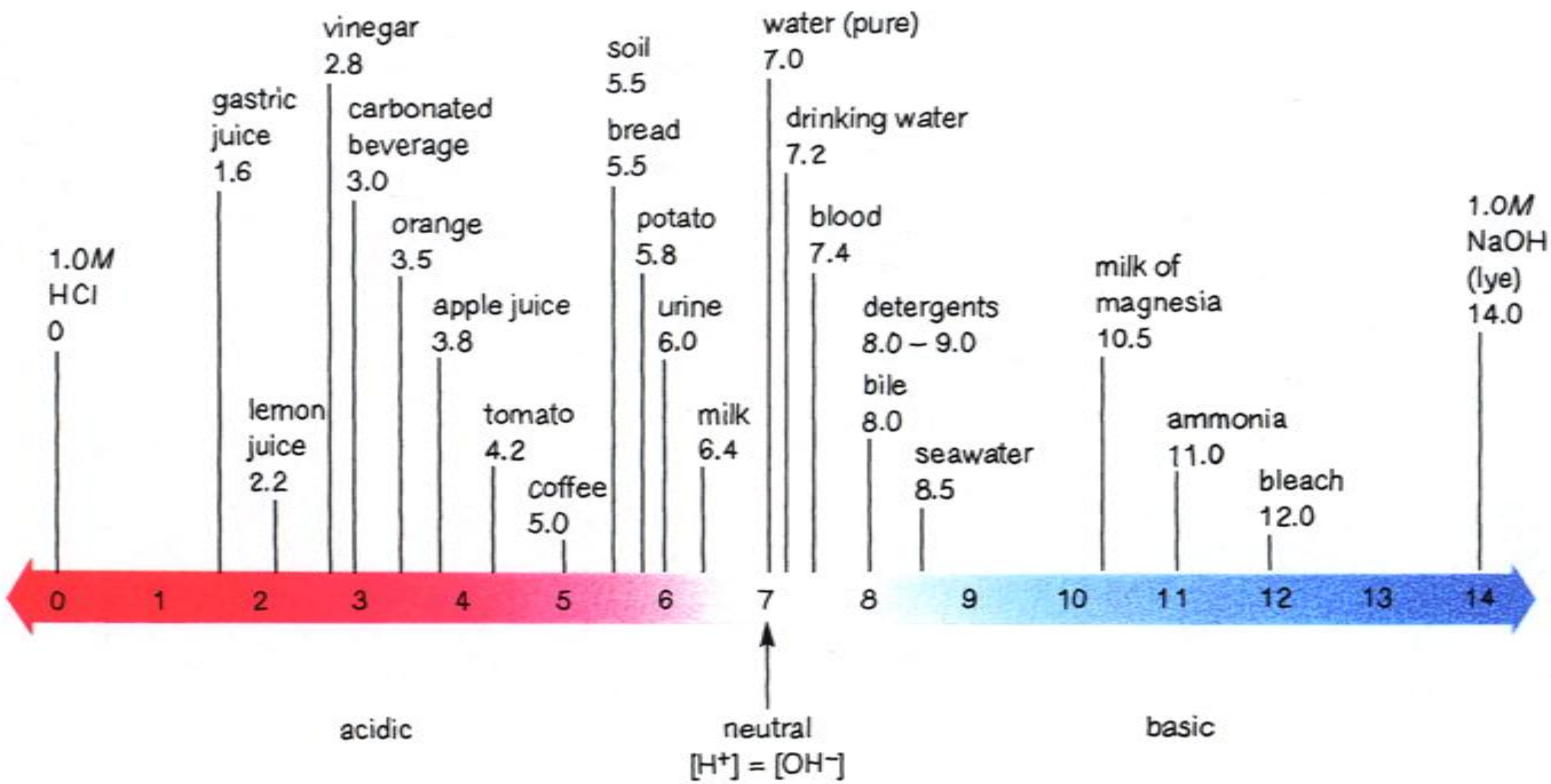
Timberlake, Chemistry 7th Edition, page 332





Zumdahl, Zumdahl, DeCoste, World of Chemistry 2002, page 515

pH of Common Substances



Timberlake, Chemistry 7th Edition, page 335