Analytical Chemistry

Lecture 2 Acids & Bases

Strengths of Acids and Base

Acids are typically classified according to their ability to donate protons. Strong acids give up protons easily, whereas in a weak acid most of the molecules keep their protons and just a few give them up . Bases are classified in terms of their ability to accept protons. Strong bases have a strong attraction for protons, whereas weak bases have little attraction for protons.

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

A similar equation and equilibrium expression can be written for a weak base such as ammonia in water

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

Problem Using a pH meter a chemist finds that 42.50 ml of a 0.1 M NaOH solution is required to titrate (neutralize) 31 ml of a hydrochloric acid solution of unknown concentration . What is the molarity of the hydrochloric acid ?

Solution

Step 1 Calculate the moles of added OH^- from M of OH^- × V(in litters) = moles of OH^-

$$\frac{0.1 \text{ ml}}{1 \text{ L}} \times 0.0425 \text{ L} = .00425 \text{ mol of } OH^-$$

(0.1 M NaOH is 0.1 M in OH^- ; 42.5 ml × 1L/1000 ml = 0.0425 L)

Step 2 At the equivalence point , moles of OH^- = moles of H^+ Therefore, 0.00425 mol of H^+ was neutralized

Step 3 Calculate the acid molarity by dividing moles by litters of solution . There is 0.00425 mol of H^+ in o.oo425 mol of HCl .

$$31 \text{ ml of HCl} \times \frac{1\text{L}}{1000\text{ml}} = 0.031 \text{ L of HCl solution}$$

$$\text{M of HCl} = \frac{\text{Moles of HCl}}{\text{Litter of solution}}$$

$$\text{M of HCl} = \frac{0.00425 \text{ mol of HCl}}{0.031 \text{ L of HCl}} = 0.137 \text{ M}$$

Problem

A student neutralizes 29.1 ml of a potassium hydroxide solution with 15.3 ml of a 0.5 M H_2SO_4 solution . What is the molarity of the KOH solution ?

Solution

Note that in this case an acid of known concentration is used to titrate a base of unknown concentration . The principles are adjust the same as those of the reverse type of titration .

Step 1 As before, we want to calculate moles of added known, in this case, H^+ from M of $H^+ \times V$ (in litters)= moles H^+ . 0.5 M H_2SO_4 is 1 M in H^+ because there are 2 mol of H^+ per 1 mol of H_2SO_4 .

$$\frac{0.5 \text{ mol of } H_2SO_4}{\text{L}} \quad \frac{2 \text{ mol of } H^+}{1 \text{ mol of } H_2SO_4} \quad \frac{1 \text{ mol o f } H^+}{\text{L}}$$

$$15.3 \text{ ml of soln} \times \frac{1 \text{ L}}{1000 \text{ ml}} = 0.0153 \text{ L of soln}$$

$$\text{M} \times \text{V} = \frac{1 \text{ ml of } H^+}{\text{L}} \times 0.0153 \text{ L} = 0.0153 \text{ mol of } H^+$$

Step 2 At the equivalence point, moles of H^+ = moles of OH^- . Therefore, 0.0153 mol of OH^- was neutralized .

Step 3 Calculate the base molarity by dividing moles by litters of solution. There is 0.0153 mol of OH^- in 0.0153 mol of KOH

$$29.1 \text{ ml of KOH} \times \frac{1 \text{ L}}{1000 \text{ ml}} = 0.0291 \text{ L of KOH}$$

$$\text{M} = \frac{0.0153 \text{ mol KOH}}{0.0291 \text{ L KOH}} = 0.526 \text{ M KOH}$$

Antacids

Antacids are substances used to neutralize excess stomach acid (HCL). Some antacids are mixtures of aluminum hydroxide and magnesium hydroxide. These hydroxides are not very soluble in water, so the levels of available OH are not damaging to the intestinal tract, which may cause weakness and loss of appetite.

Magnesium hydroxide has a laxative effect. These side effects are less likely when a combination of antacids is used.

$$Al(OH)_3 + 3HCl \longrightarrow AlCl_3 + 3H_2O$$
$$Mg(OH)_2 + 2HCl \longrightarrow MgCl_2 + 2H_2O$$

Some antacids used calcium carbonate to neutralize excess stomach acid. About 10% of the calcium is absorbed into the bloodstream, where it elevates the levels of serum calcium . Calcium carbonate is not recommended for patients who have peptic ulcers or a tendency to form kidney stones.

$$CaCO_3 + 2HCl \longrightarrow H_2O + CO_2(g) + CaCl_2$$

Still other antacids contain sodium bicarbonate. This type of antacids has a tendency to increase blood pH and elevate sodium levels in the body fluids. It also is not recommended in the treatment of peptic ulcers.

$$NaHCO_3 + HCl \longrightarrow NaCl + CO_2(g) + H_2O$$

The pH Scale

Many Kinds of careers such as respiratory therapy, medicine, agriculture, spa cleaning, and soap manufacturing require personnel to measure the $[H_3O^+]$ and $[OH^-]$ of solutions. The proper levels of acidity are necessary for soil to support plant growth and prevent algae in swimming pool water. Measuring the acidity levels of blood and urine checks the function of kidneys.

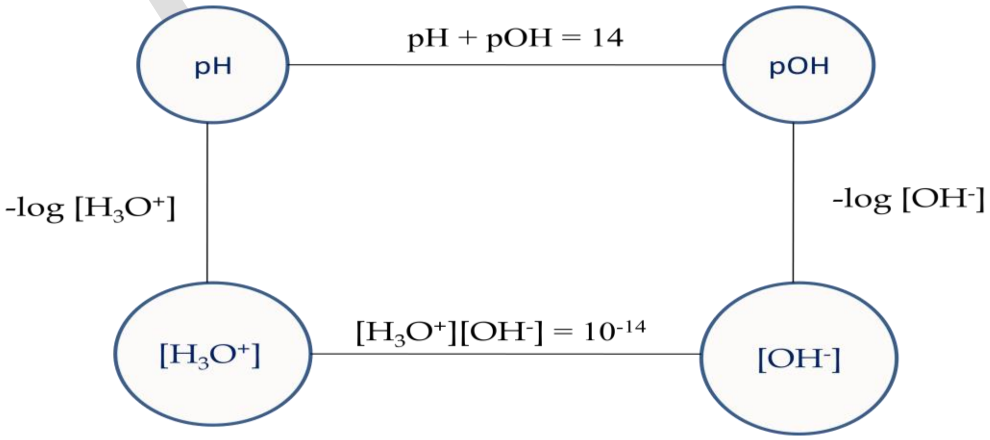
The simpler way of describing acidity is used: the pH scale. On the pH scale, a number between 0 and 14 represents the H_3O^+ concentration. A pH value less than 7 corresponds to an acidic solution; a pH greater than 7 indicates a basic solution.

Acidic solution	pH < 7	$[H_3O^+] > 1.0 \times 10^{-7}$ M
Neutral solution	pH = 7	$[H_3O^+] = 1.0 \times 10^{-7}$ M
Basic solution	pH > 7	$[H_3O^+] < 1.0 \times 10^{-7}$ M

Calculation the pH and pOH of solution

$$pH = -\log [H_3O^+]$$

$$pOH = -\log [OH^-]$$



Problem

Determine the pH for the following solutions:

a. $[H_3O^+] = 1\times10^{-5}$ M b. $[H_3O^+] = 2.5 \times 10^{-8}$

Solution

a. pH = -log $[1\times10^{-5}]$ = 5 b. pH = -log $[2.5 \times10^{-8}]$ = *?*

Problem

Determine the $[H_3O^+]$ for the following solutions:

a. pH = 3 b. pH= 12

Solution

a. $[H_3O^+] = 1\times10^{\overset{\text{pH}}{-3}}$ **M** **b.** $[H_3O^+] = 1\times10^{\overset{\text{pH}}{-12}}$ **M**

Calculation the pK_w

K_w = 1.0×10^{-14}

pK_w = - log $[K_w]$ = 14

= (-log $[H_3O^+]$) + (-log $[OH^-]$)

= pH + pOH = 14

Problem

a. What is the pOH and pH of seawater if the $[OH^-]$ is 1.0×10^{-6} M ?

b. What is the pOH and pH of sample if $[H_3O^+]$ is 1.5×10^{-3} M ?

Solution

a. pOH = -log $[OH^-]$ = - log $[1.0 \times 10^{-6}]$ = 6

From the pK_w we know that pH + pOH = 14 . Solving for pH gives

$pH = 14 – pOH = 14 – 6 = 8$

b. $pH = -\log [H_3O^+] = -\log [1.5 \times 10^{-3}] = 2.82$

From the pK_w we know that pH + pOH = 14 . Solving for pOH gives

$pOH = 14 – pH = 14 – 2.82 = 11.18$

Blood Buffers

We can discuss the body's mechanisms for protecting blood from dramatic pH changes. Three interconnected systems maintain the pH in blood: (1) the blood buffers, which actually serve to neutralize the added hydrogen and hydroxide ions which form from the body's metabolic reactions; (2) the lungs, which are involved in the excretion and inhalation of carbon dioxide, and thereby maintain the concentration of carbonic acid in blood; and (3) the kidneys, which excrete hydrogen ions and bicarbonate ions from the blood. In this lecture we will concentrate on the role of the blood buffers and witness the supporting role of the other two systems.

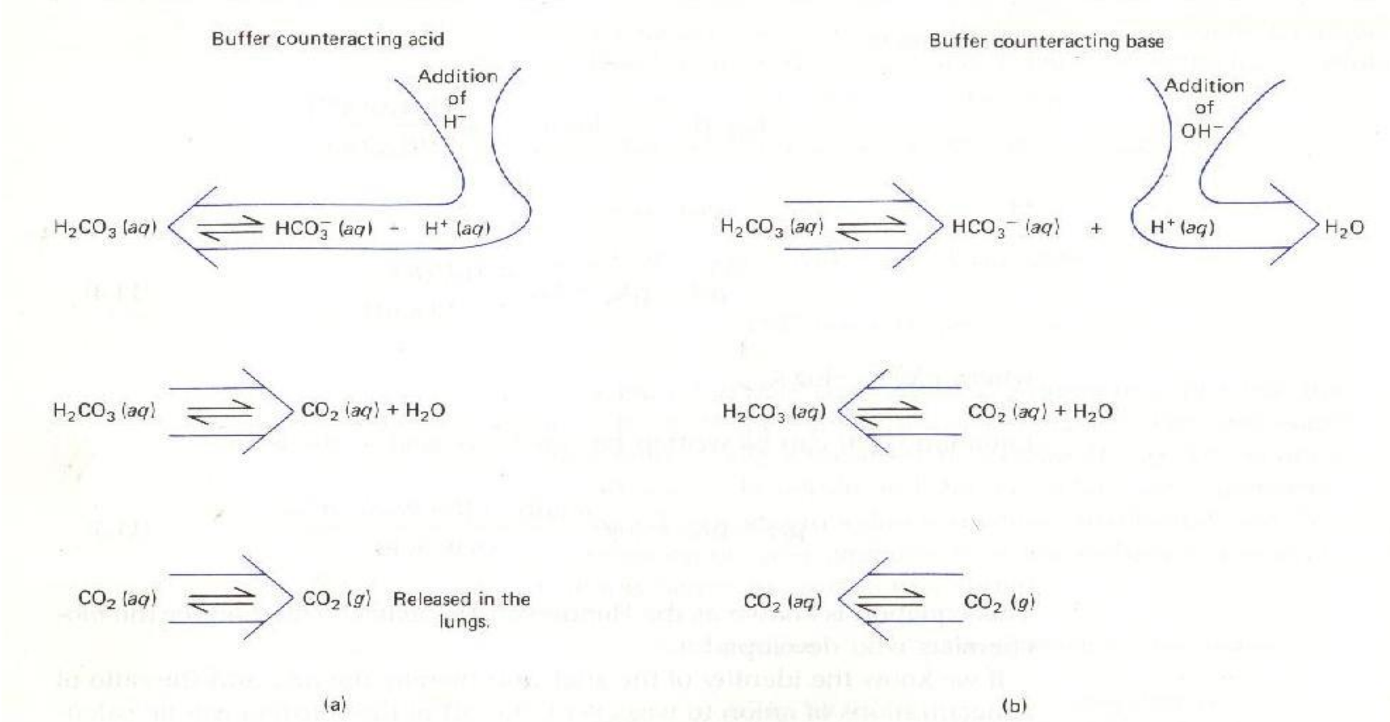
There are three major body buffers:(1)the H_2CO_3 / HCO_3^- buffer, (2) the $H_2PO_4^-$ / HPO_4^{2-} buffer, and (3) the protein buffers. The carbonic acid-bicarbonic acid is an unstable weak acid which in aqueous solution , in this case blood , is always in equilibrium with CO_2(aq).

$$\mathbf{H_2CO_3(aq) \rightleftharpoons CO_2(aq) + H_2O}$$

The position of this equilibrium is to the right. Dissolved CO_2(aq)is also in equilibrium with CO_2(g) in the lungs.

$$\underset{\text{Blood}}{\mathbf{CO_2(aq)}} \rightleftharpoons \underset{\text{lungs}}{\mathbf{CO_2(g)}}$$

A moment's though (and perhaps some help from Le Chatelier) should convince you that the concentration of H_2CO_3(aq) can be directly affected by that of CO_2(g).



Action of the carbonic acid-bicarbonate buffer upon addition of acid (a) or base (b) .

$$pH = pKa + \log \frac{[\text{anion of the }]}{[\text{weak acid}]}$$

The concentration of $HCO_3^-(aq)$ and $H_2CO_3(aq)$ in the blood of a healthy individual are 2.5×10^{-2}M and 1.25×10^{-3} M , respectively. The pKa of carbonic acid is 6.1 . Using these data and the above equation .

The concentration of $HCO_3^-(aq)$ and $H_2CO_3(aq)$ in the blood of a healthy individual are 2.5×10^{-2}M and 1.25×10^{-3} M , respectively. The pKa of carbonic acid is 6.1 . Using these data and the above equation .

$$pH = pKa + \log \frac{[HCO_3^-]}{[H_2CO_3]}$$

$$pH = 6.1 + \log \frac{2.5 \times 10^{-2}}{1.25 \times 10^{-3}}$$

$$pH = 6.1 + \log 20$$

$$pH = 6.1 + 1.3 = 7.4$$