

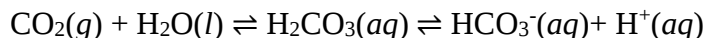
Chapter 10

Problems

Problem 1

(Wikipedia)

The **bicarbonate buffer system** is an acid-base homeostatic mechanism involving the balance of carbonic acid (H_2CO_3), bicarbonate ion (HCO_3^-), and carbon dioxide (CO_2) in order to maintain pH in the blood and duodenum, among other tissues, to support proper metabolic function. Catalyzed by carbonic anhydrase, carbon dioxide (CO_2) reacts with water (H_2O) to form carbonic acid (H_2CO_3), which in turn rapidly dissociates to form a bicarbonate ion (HCO_3^-) and a hydrogen ion (H^+) as shown in the following reaction:



(i). What is the conjugate acid of HCO_3^- ?

- (a) HCO_3^-
- * (b) H_2CO_3
- (c) H_2CO_3^+
- (d) CO_2
- (e) CO_3^{2-}

Answer

The conjugate ACID of a species A^- is HA . Thus, adding a proton (H^+) to HCO_3^- gives H_2CO_3 .

(ii). What is the conjugate base of HCO_3^- ?

- (a) HCO_3^-
- (b) H_2CO_3
- (c) H_2CO_3^+
- (d) CO_2
- * (e) CO_3^{2-}

Answer

The conjugate BASE of a species HA is A^- . Thus, REMOVING a proton (H^+) from HCO_3^- gives CO_3^{2-} .

Problem 2

<https://sfyl.ifas.ufl.edu/media/sfylifasufledu/miami-dade/documents/sea-grant/Temperature,-Salinity-and-pH.pdf>

(Adapted from the link above)

The pH and temperature are critical to the survival of most aquatic plants and animals and are therefore important parameters for water monitoring. The average pH for sea water is 8.2 but can range between 7.5 and 8.5 depending on the local conditions. Human activities such as sewage overflows or runoff can cause significant short-term fluctuations in pH and long-term impacts can be extremely harmful to plants and animals.

(i). Which solution below has the **highest** concentration of hydroxide ions?

- (a) pH = 7.50
- (b) pH = 7.60
- (c) pH = 8.00
- (d) pH = 8.20
- *(e) pH = 8.50

Answer

The HIGHER the pH, the HIGHER the concentration of hydroxide ions in an aqueous solution.

(ii). Which solution below has the **highest** concentration of hydronium ions?

- *(a) pH = 7.50
- (b) pH = 7.60
- (c) pH = 8.00
- (d) pH = 8.20
- (e) pH = 8.50

Answer

The LOWER the pH, the HIGHER the concentration of hydronium ions in an aqueous solution.

(iii) The concentration of hydronium ions in a solution whose pH is 7.50 at 25°C is _____?

Answer

$$[H_3O^+] = 10^{-pH} = 10^{-7.50} = 3.16 \times 10^{-8}$$

This will be reported with 2 sf as $3.2 \times 10^{-8} M$.

(iv) The concentration of hydroxide ions in a solution whose pH is 7.50 at 25°C is _____?

Answer

$$[H_3O^+] = 10^{-pH} = 10^{-7.50} = 3.16 \times 10^{-8}$$

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

$$[OH^-] = \frac{1.0 \times 10^{-14}}{3.16 \times 10^{-8}} = 3.16 \times 10^{-7}$$

This will be reported with 2 sf as $3.2 \times 10^{-7} M$.

Alternately, at 25°C

$$pH + pOH = 14.00$$

$$\text{Thus, } pOH = 14.00 - 7.50 = 6.50$$

$$[OH^-] = 10^{-pOH} = 10^{-6.50} = 3.16 \times 10^{-7}$$

Problem 3

(i). What is the pH of a 0.10 M solution of nitric acid (HNO₃) at 25°C?

(a) 1.0×10^{-1}

*(b) 1.00

(c) 7.00

(d) 13.00

(e) 14.00

Answer

HNO₃ is a strong acid, and will completely dissociate, giving rise to $[H_3O^+] = 0.10 M$

$$\text{Thus, } pH = -\text{Log}(0.10) = 1.00$$

(ii). What is the pH of a 0.10 M solution of barium hydroxide at 25°C?

(a) 1.0×10^{-1}

(b) 1.00

(c) 7.00

(d) 13.00

*(e) 13.30

Answer

Barium hydroxide, Ba(OH)₂, a strong base, will completely dissociate, giving rise to $[OH^-] = 0.20 M$.

$$pOH = -\text{Log}(0.20) = 0.70$$

$$pH = 14.00 - 0.70 = 13.30$$

Problem 4

(i). A 0.10 M aqueous solution of ammonium nitrate has a pH _____ at 25°C.

- *(a) less than 7.00
- (b) equal to 7.00
- (c) greater than 7.00

Answer

Ammonium ions are acidic, and nitrate ions are neutral. Thus, the solution is acidic.

(ii). Saline, a 0.15 M solution of salt (sodium chloride) in water is used to clean wounds and remove contact lenses. A 0.15 M aqueous solution of sodium chloride has a pH _____ at 25°C.

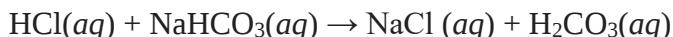
- (a) less than 7.00
- *(b) equal to 7.00
- (c) greater than 7.00

Answer

Both sodium and chloride ions are neutral. Thus, the solution is neutral.

Blurb on gastric juice, mostly from Wikipedia

While gastric juice is secreted with a pH as low as 0.8, it is immediately diluted in the stomach lumen and its pH rises to between 1.5 and 3.5. In the [duodenum](#), gastric acid is [neutralized](#) by [sodium bicarbonate](#). This also blocks gastric enzymes that have their optima in the acid range of pH. The secretion of sodium bicarbonate from the [pancreas](#) is stimulated by [secretin](#). This [polypeptide](#) hormone gets activated and secreted from so-called [S cells](#) in the mucosa of the duodenum and [jejunum](#) when the pH in the duodenum falls below 4.5 to 5.0. The neutralization is described by the equation:



The [carbonic acid](#) rapidly equilibrates with [carbon dioxide](#) and [water](#) through catalysis by carbonic anhydrase enzymes bound to the gut epithelial lining, leading to a net release of carbon dioxide gas within the lumen associated with neutralization.

Qualitative problems on titrations

Problem 5

(i). A 25.0 mL solution of 0.010 M hydrochloric acid (to mimic gastric juice) is titrated with 0.010 M solution of sodium hydroxide at 25°C. At the equivalence point, the pH is _____.

- (a) less than 7.00
- * (b) equal to 7.00
- (c) greater than 7.00

Answer

Hydrochloric acid is a strong acid and sodium hydroxide is a strong base. At the equivalence point, the salt sodium nitrate (which is present as neutral Na^+ ions and neutral nitrate ions) is neutral.

(ii). A 25.0 mL solution of 0.010 M acetic acid (approximately 80-fold diluted vinegar) is titrated with 0.010 M solution of sodium hydroxide at 25°C. At the equivalence point, the pH is _____.

- (a) less than 7.00
- (b) equal to 7.00
- * (c) greater than 7.00

Answer

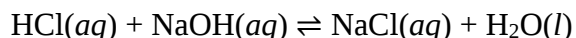
Acetic acid is a weak acid and sodium hydroxide is a strong base. At the equivalence point, the salt sodium acetate (which is present as neutral Na^+ ions and basic acetate ions) is basic.

Problem 6

To model gastric juice, a student prepared 100.0 mL of a 0.0100 M solution of HCl. What is the volume of 0.100 M NaOH needed to titrate the HCl sample to the endpoint?

Answer

We begin with a balanced equation:



$$\begin{aligned} \text{Volume of NaOH} &= (100.0 \text{ mL HCl}) \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{0.0100 \text{ mol HCl}}{1 \text{ L HCl}} \right) \left(\frac{1 \text{ mol NaOH}}{1 \text{ mol HCl}} \right) \left(\frac{1 \text{ L NaOH}}{0.100 \text{ mol NaOH}} \right) \\ &= 1.00 \times 10^{-2} \text{ L} \end{aligned}$$

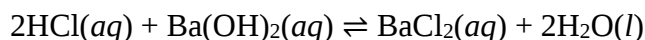
This is conveniently expressed in mL as 10.0 mL

Problem 7

To model gastric juice, a student prepared 100.0 mL of a 0.0100 M solution of HCl. What is the volume of 0.100 M Ba(OH)₂ needed to titrate the HCl sample to the endpoint?

Answer

We begin with a balanced equation:



$$\begin{aligned} \text{Volume of Ba}(\text{OH})_2 &= (100.0 \text{ mL HCl}) \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{0.0100 \text{ mol HCl}}{1 \text{ L HCl}} \right) \left(\frac{1 \text{ mol Ba}(\text{OH})_2}{2 \text{ mol HCl}} \right) \left(\frac{1 \text{ L Ba}(\text{OH})_2}{0.100 \text{ mol Ba}(\text{OH})_2} \right) \\ &= 5.00 \times 10^{-3} \text{ L} \end{aligned}$$

This is conveniently expressed in mL as 5.00 mL.

Problem 8

To model gastric juice, a student prepared 100.0 mL of a 0.0100 M solution of HCl. What is the mass of CaCO_3 (the major ingredient in Tums) needed to react completely with the HCl?

Answer

We begin with a balanced equation:



$$\begin{aligned} \text{Mass of CaCO}_3 &= (100.0 \text{ mL HCl}) \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{0.0100 \text{ mol HCl}}{1 \text{ L HCl}} \right) \left(\frac{1 \text{ mol CaCO}_3}{2 \text{ mol HCl}} \right) \left(\frac{100.09 \text{ g CaCO}_3}{1 \text{ mol CaCO}_3} \right) \\ &= 0.0500 \text{ g CaCO}_3 \end{aligned}$$

This is the mass of one-tenth of a regular-sized tablet of Tums!

Problem 9

A student was carrying a 1.00 L glass bottle with 1.00 M sulfuric acid in it. He dropped the bottle, which broke, spilling the contents on the laboratory floor. While waiting for the safety personnel to arrive, he sprinkled baking soda (NaHCO_3) to neutralize the acid. What is the mass of baking soda needed to react with all the acid?

Answer

We begin with a balanced equation:



$$\begin{aligned} \text{Mass of NaHCO}_3 &= (1.00 \text{ L H}_2\text{SO}_4) \left(\frac{1.00 \text{ mol H}_2\text{SO}_4}{1 \text{ L H}_2\text{SO}_4} \right) \left(\frac{2 \text{ mol NaHCO}_3}{1 \text{ mol H}_2\text{SO}_4} \right) \left(\frac{84.008 \text{ g NaHCO}_3}{1 \text{ mol NaHCO}_3} \right) \\ &= 168 \text{ g NaHCO}_3 \end{aligned}$$

From Wikipedia

Lactic acidosis is a medical condition characterized by the buildup of [lactate](#) (especially [L-lactate](#)) in the body, with formation of an excessively low [pH](#) in the bloodstream. It is a form of [metabolic acidosis](#), in which excessive acid accumulates due to a problem with the body's oxidative [metabolism](#).

Lactic acidosis is typically the result of an underlying acute or chronic medical condition, medication, or poisoning. The symptoms are generally attributable to these underlying causes, but may include [nausea](#), [vomiting](#), [Kussmaul breathing](#) (laboured and deep), and generalized weakness.

Problem 10

An aqueous buffer solution is prepared from equal volumes of 1.00 M sodium lactate and 1.50 M lactic acid. Calculate the pH of the resulting buffer. (For lactic acid, $K_a = 1.4 \times 10^{-4}$)

Answer

Use the Henderson-Hasselbalch equation

$$\begin{aligned} \text{pH} &= \text{p}K_a + \text{Log}\left(\frac{[\text{sodium lactate}]}{[\text{lactic acid}]}\right) \\ &= 3.854 + \text{Log}\left(\frac{1.0 \text{ M}}{1.5 \text{ M}}\right) = 3.68 \end{aligned}$$

Please note that in the equation above, we have used the ratios of concentrations of the [original solutions](#) since [equal volumes](#) of the two solutions were used. We have rounded the answer to two decimal places in the calculation of pH since there are two significant figures for the value of K_a .

Problem 11

True or false.

For the buffer in Problem 10, 1.0 L of water was added. The pH of the buffer drops.

Answer

False. The individual concentrations of sodium lactate and lactic acid are both reduced; however, the ratio of the concentrations of sodium lactate and lactic acid remains unchanged.

Problem 12

True or false

A buffer is prepared by adding equal number of moles of a weak monoprotic acid and its sodium salt at 25°C. The pH of this buffer varies, depending on the total volume of the buffer, since concentrations are used in the Henderson-Hasselbalch equation.

Answer

False. The individual concentrations of the weak acid (HA) and its sodium salt (NaA) do change depending on the volume of the solution. However, the ratio of the concentrations remains unchanged at 1 since equal numbers of moles of HA and NaA are used.

$$\begin{aligned} \text{pH} &= \text{p}K_a + \text{Log}\left(\frac{[\text{NaA}]}{[\text{HA}]}\right) \\ &= \text{p}K_a \end{aligned}$$

We note that $\text{Log}(1) = 0$

Problem 13

An aqueous buffer solution is prepared from equal volumes of 1.00 M sodium acetate and 2.00 M acetic acid. Calculate the pH of the resulting buffer. (For acetic acid, $K_a = 1.77 \times 10^{-5}$)

Answer

Use the Henderson-Hasselbalch equation

$$\begin{aligned} \text{pH} &= \text{p}K_a + \text{Log}\left(\frac{[\text{sodium acetate}]}{[\text{acetic acid}]}\right) \\ &= 4.752 + \text{Log}\left(\frac{1.00 \text{ M}}{2.00 \text{ M}}\right) = 4.451 \end{aligned}$$

Please note that in the equation above, we have used the ratios of concentrations of the original solutions since equal volumes of the two solutions were used. We have rounded the answer to three decimal places in the calculation of pH since there are three significant figures for K_a .

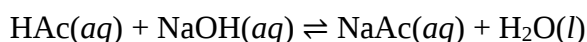
Question 14

True or false

For the buffer in Problem 13, solid NaOH is added. The pH increases a little.

Answer

True. The reaction that occurs is:



Ac is used as the abbreviation for acetate. The concentration of sodium acetate will increase and that of acetic acid will decrease. Thus, the pH increases. The increase is quite small, since this is a buffer.