1) Solid aluminum phosphate, AlPO₄, is in equilibrium with its ions in solution:

\[ \text{AlPO}_4 \rightleftharpoons \text{Al}^{3+} + \text{PO}_4^{3-} \]  with \( K_{sp} = 10^{-22} \)

Find the equilibrium concentration of phosphate ion (in mg/L)

2) Water is frequently disinfected with chlorine gas, forming Hypochlorous acid (HOCl), which partially ionizes to hypochlorite and hydrogen ions as follows:

\[ \text{HOCl} \rightleftharpoons \text{H}^+ + \text{OCl}^- \] \( K_a = 2.9 \times 10^{-8} \)

The amount of [HOCl], which is desired disinfectant, depends on the pH. Find the fraction that is hypochlorous acid – that is, \([\text{HOCl}]/([\text{HOCl}]+[\text{OCl}^-])\) – as a function of pH. What would the hypochlorous fraction be for pH = 6, 8, and 10?

3) A solution of sodium bicarbonate is prepared by adding 45.00 g of sodium bicarbonate to a 1.00 L volumetric flask and adding distilled water until it reaches the 1.00 L mark. What is the concentration of sodium bicarbonate in units of (a) milligram per liter, (b) molarity, (c) normality and (d) milligram per liter as CaCO₃?

4) A magnesium hydroxide solution is prepared by adding 10.00 g of magnesium hydroxide to a volumetric flask and bringing the final volume to 1.00 L by adding water buffered such that the pH will remain at 7.0. What is the concentration of magnesium in this solution? (Assume standard pressure and temperature and the ionic strength is negligible)

5) The ground water in Pherric, New Mexico contains 1.80 mg/L of iron as Fe³⁺. What pH is required to precipitate all but 0.30 mg/L of the iron, as Fe(OH)₃ at 25°C? Hint: You will need to use the solubility table handed out in class.

6) Calculate the molarity and normality of the following:
   a. 80 ug/L HNO₃
   b. 135 ug/L CaCO₃
   c. 10 ug/L Cr(OH)₃
   d. 1000 ug/L Ca(OH)₂

7) At a pH of 4.75 what acetic acid (CH₃CO₂H) specie would be predominant? How about at a pH of 6? And finally at a pH of 3?
8) Convert the following from milligrams per liter as the ion or compound to milligrams per liter as CaCO₃.
   a. 200.00 mg/L NH₄⁺
   b. 280.00 mg/L K⁺
   c. 123.45 mg/L SO₄²⁻
   d. 85.05 mg/L Ca²⁺
   e. 19.90 mg/L Na⁺

9) You are in an environmental engineering laboratory:
   a. You have been asked to prepare a solution sodium hydroxide (NaOH). You add 5 grams of pure NaOH pellets and then fill the beaker to 1.0 liter with deionized water. What is the pH of the solution after equilibrium has been reached?
   b. Now you are asked to predict the pH of distilled water after adding a specific amount of acid. The type of acid used is sulfuric acid (H₂SO₄). Using a 0.1 M solution of H₂SO₄ and a 0.5 cubic centimeter syringe you are told to add 15 full injections to a beaker and then fill the beaker to the 2 liter mark….what is the pH of the solution?

10) a. Suppose the gas above the soda in a bottle of soft drink is pure CO₂ at a pressure of 2 atm. Calculate the [CO₂] at 25 °C.
    Henry’s Law = [CO₂] = K_H ’P_gas
    Where K_H = 0.0333947 mol/(L·atm) @ 25 °C
    b. Calculate the pH of the soft drink. Start with the following chemical reaction and realize the solution will be somewhat acidic (negligible carbonate):
    CO₂ + H₂O ⇌ H⁺ + HCO₃⁻
    (Hint: Assume the only cations and anions in the system are H⁺, OH⁻ and HCO₃⁻)
1. **Problem Statement:** Solid aluminium phosphate is in equilibrium with its ions in solution.

**Known:**

\[ \text{AlPO}_4 \rightleftharpoons \text{Al}^{3+} + \text{PO}_4^{3-} \]

\[ K_{sp} = 10^{-22} \]

**Find:**

The equilibrium concentration of phosphate ion in mg/L

**Assumptions:**

\[ K_{sp} = \frac{[\text{Al}^{3+}][\text{PO}_4^{3-}]}{[\text{AlPO}_4]} \]

standard temperature and pressure

\[ [\text{Al}^{3+}] = [\text{PO}_4^{3-}] = x \]

**Solve:**

\[ K_{sp} = \frac{[\text{Al}^{3+}][\text{PO}_4^{3-}]}{[\text{AlPO}_4]} = x^2 \]

\[ 10^{-22} = x^2 \]

\[ x = 1.0 \times 10^{-11} \text{ mole/L} \]

\[ \text{MW} (\text{PO}_4^{3-}) = 95.91 \text{ g/mole} \]

\[ 1.0 \times 10^{-11} \text{ mole} \times \frac{95.91 \text{ g}}{1 \text{ mole}} \times \frac{1000 \text{ mg}}{1 \text{ g}} = 9.59 \times 10^{-7} \text{ mg/L} \]

**Equilibrium concentration** \[ \text{PO}_4^{3-} = 9.59 \times 10^{-7} \text{ mg/L} \]

2. **Problem Statement:** Water is frequently disinfected with chlorine gas, forming hypochlorous acid, which partially ionizes to hypochlorite and hydrogen ions.
2. (Continued)

Known:
\[ \text{HOCI} \rightleftharpoons \text{H}^+ + \text{OCI}^- \quad \text{\( K_a = 2.9 \times 10^{-8} \)}

The amount of \([\text{HOCI}]\), which is desired disinfectant, depends on the pH.

Find:

The fraction that is hypochlorous acid, \([\text{HOCI}]\), as a function of pH. What would the hypochlorous fraction be for a pH = 6, pH = 8 and pH = 10?

Assumptions:

\[ K_a = \frac{[\text{H}^+][\text{OCI}^-]}{[\text{HOCI}]} \quad \text{pH} = -\log [\text{H}^+] \]

Solve:

pH = 6, HOCI is the dominant species.

\[ [\text{OCI}^-] = K_a [\text{HOCI}] \quad \text{pH} = -\log [\text{H}^+] \]

\[ [\text{HOCI}] = \frac{[\text{HOCI}]}{[\text{HOCI}] + [\text{H}^+] + K_a [\text{HOCI}]} \]

\[ [\text{HOCI}] = \frac{1}{1 + \frac{K_a [\text{HOCI}]}{[\text{H}^+]}} \]

\[ \text{HOCI} = \frac{1}{1 + \frac{2.9 \times 10^{-8}}{1 \times 10^{-6} \text{M}}} = 0.972 = 97.2\% \]
2. (continued)

\[
\text{PH} = 8 \quad \Rightarrow \quad \frac{1}{1 + 2.9 \times 10^{-8}} = 0.250 = 25.0\%
\]

\[
\text{HCl} = \frac{1}{1 + 2.9 \times 10^{-8}} = 0.0034 = 0.34\%
\]

\[
\text{PH} = 6 \quad \text{HCl Fraction} = 0.972 = 97.2\%
\]

\[
\text{PH} = 8 \quad \text{HCl Fraction} = 0.250 = 25.0\%
\]

\[
\text{PH} = 10 \quad \text{HCl Fraction} = 0.0034 = 0.34\%
\]

3. Problem statement (known): A solution of sodium bicarbonate is prepared by adding 45.0 g of NaHCO₃ to a 1.0 L volumetric flask and adding DI water until it fills to 1.0 L. Find:

What is the concentration of sodium bicarbonate in units of mg/L, molarity, normality, and mg/L as CaCO₃?

Assumptions:
Standard temperature and pressure.

Solve:

a. \[45.0 \text{ g NaHCO}_3 \times \frac{1 \text{ mol}}{84 \text{ g}} = \frac{45.000 \text{ mg/L NaHCO}_3}{1000} \]

b. Molarity = \[\frac{\text{conc.}}{\text{Mw}} = \frac{45.0 \text{ mg/L}}{84 \text{ g/mol}}\]

\[
\text{Mw (NaHCO}_3) = 23 + 1 + 12 + 3(16) = 84 \text{ g/mol}
\]
3. (continued)

\[ \text{Normality} = \frac{\text{conc.}}{\text{EW}} \]

\[ \text{EW} = \frac{\text{MW}}{\text{charge}} \]

\[ N = \frac{45.0 \text{ g L}^{-1}}{84.91 \text{ g mol}^{-1}} = 0.536 \text{ eq L}^{-1} \text{ NaHCO}_3 \]

a. \[ \text{MW (CaCO}_3) = 40 + 12 + 3(16) = 100 \text{ g mol}^{-1} \]
\[ \left( \frac{100 \text{ g mol}^{-1}}{2} \right) = 50 \text{ g mol}^{-1} \]

\[ (0.536 \text{ eq L}^{-1})(50 \text{ g mol}^{-1}) = 26.78 \text{ g L}^{-1} = 2.68 \times 10^4 \text{ mg L}^{-1} \text{ as CaCO}_3 \]

4. Problem statement known: A magnesium hydroxide solution is prepared by adding 10.0 g of magnesium hydroxide to a volumetric flask and bringing the final volume to 1.0 L by adding water to buffered such that the pH will remain at 7.0.

Find:

What is the concentration of magnesium in this solution?

Assumptions:

Standard temperature and pressure

\[ \text{Mg(OH)}_2 \rightarrow \text{Mg}^{2+} + 2\text{OH}^- \]

Ionic strength is negligible

\[ pK_{sp} = 11.25 \]

\[ K_{sp} = \left[ \frac{\text{Mg}^{2+}}{[\text{Mg(OH)}_2]} \right] \]
4) Problem 2-7 Davis and Masten

Concentration of magnesium hydroxide
Given: 10.00 g Mg(OH)_2, 1 L water, pH = 7, assuming temperature is 25º C and ionic strength is negligible

Solution:
Note that the molecular weight of Mg(OH)_2 = 58.526 g · mol⁻¹.

A 10.00 g sample is: 10 g Mg(OH)_2 X (58.526 g · mol⁻¹)⁻¹ = 0.17 mol

Since there is 1 mol Mg per 1 mol Mg(OH)_2 and this is dissolved in 1 L water, it follows that [Mg²⁺] = 0.17 M

Now calculate maximum solubility to prove max amount of Mg²⁺ that could be in solution

Using the solubility equilibrium reaction from Table 2-1
Mg(OH)_2 ⇌ Mg²⁺ + 2OH⁻

\[ pK_{sp} = 11.25 \]

\[ [Mg^{2+}][OH^-]^2 = K_s = 10^{-11.25} \]

Knowing that pH is buffered to a pH of 7, so pOH is also 7 then we know that [OH⁻] = 10⁻⁷ M

Substituting
\[ [Mg^{2+}](10^{-7})^2 = 10^{-11.25} \]
\[ [Mg^{2+}] = 562.34 \text{ M} = \text{Saturated Solution} - \text{Max Value of Mg}^{2+} \text{ that would be in solution at saturation} = \text{Even if there were enough Mg}^{2+} \text{ present this is an unrealistic value.} \]

**NOW!**
Since only added enough magnesium hydroxide to have 0.17 M, the concentration will be

**0.17 M**
5. Problem statement known: The groundwater in pherfect New Mexico contains 1.80 mg/l of iron as Fe³⁺.

Find:

What pH is required to precipitate all but 0.30 mg/l of the iron at 25°C?

Assumptions:

Standard temperature and pressure

Fe³⁺ + 30H⁻ → Fe(OH)₃

Ksp = \left(\frac{Fe^{3+}}{[Fe(OH)_3]}\right)^3

Ksp = 1.10 × 10^{-37}
5. (continued)

Solve:

\[ 0.30 \text{ mg L}^{-1} \times \frac{9}{1000 \text{ mg}} \times \frac{\text{mol}}{50 \text{ g}} = 5.3 \times 10^{-6} \text{ mol L}^{-1} \]

\[ K_{sp} = \left( \frac{\text{Fe}^{3+}}{[\text{OH}^-]^3} \right) \left( \frac{[\text{Fe}(\text{OH})_2]}{[\text{OH}^-]} \right) \]

\[ [\text{OH}^-] = 6.68 \times 10^{-12} \text{ mol L}^{-1} \]

\[ p\text{OH} = -\log [\text{OH}^-] = -\log (6.68 \times 10^{-12}) = 11.18 \]

\[ pK = pH + p\text{OH} = 14 \]

\[ pH = 14 - p\text{OH} = 14 - 11.18 = 2.82 \]

\[ pH = 2.82 \]

b. Problem statement known:
Calculate molarity and normality.

Solve:
Molarity: \( \frac{\text{Conc.}}{\text{MW}} \)

Normality: \( \frac{\text{Conc.}}{\text{EW}} \) \( \text{EW} = \frac{\text{MW}}{\text{charge}} \)

a. 80 \text{ mg L}^{-1} \text{ HNO}_3

\[ \text{MW} (\text{HNO}_3) = 1 + 14 + 3(16) = 63 \text{ g mol}^{-1} \]

\[ M = \left( 80 \text{ mg L}^{-1} \times \frac{g}{1 \times 10^2 \text{ mg}} \right) = 1.27 \times 10^{-6} \text{ mol L}^{-1} \]

\[ N = \left( 80 \text{ mg L}^{-1} \times \frac{g}{1 \times 10^2 \text{ mg}} \right) = 1.27 \times 10^{-6} \text{ eq L}^{-1} \]

\[ M = 1.27 \times 10^{-6} \text{ mol L}^{-1} \]

\[ N = 1.27 \times 10^{-6} \text{ eq L}^{-1} \]
b. 135 mg/L CaCO₃

\[ M_{\text{CaCO}_3} = 40 + 12 + 3(16) = 100 \text{ g/mol} \]

\[ M = \frac{135 \times 10^{-4}}{100 \text{ g/mol}} = 1.35 \times 10^{-4} \text{ mol/L} \]

\[ N = \frac{1.35 \times 10^{-4}}{\frac{1000 \text{ g/mol}}{2}} = 2.7 \times 10^{-5} \text{ eq/L} \]

\[ M = 1.35 \times 10^{-4} \text{ mol/L} \]

\[ N = 2.7 \times 10^{-5} \text{ eq/L} \]

C. 10 mg/L Cr(OH)₃

\[ M_{\text{Cr(OH)}_3} = 52 + 3(1) + 3(16) = 103 \text{ g/mol} \]

\[ M = \frac{1.0 \times 10^{-5} \text{ g/L}}{103 \text{ g/mol}} = 9.709 \times 10^{-8} \text{ mol/L} \]

\[ N = \frac{1.0 \times 10^{-5} \text{ g/L}}{103 \text{ g/mol}} = 2.913 \times 10^{-7} \text{ eq/L} \]

\[ M = 9.709 \times 10^{-8} \text{ mol/L} \]

\[ N = 2.913 \times 10^{-7} \text{ eq/L} \]

D. 1000 mg/L Ca(OH)₂

\[ M_{\text{Ca(OH)}_2} = 40 + 2(16) + 2(1) = 74 \text{ g/mol} \]

\[ M = \frac{0.001 \text{ g/L}}{74 \text{ g/mol}} = 1.35 \times 10^{-5} \text{ mol/L} \]

\[ N = \frac{0.001 \text{ g/L}}{74 \text{ g/mol}} = 2.703 \times 10^{-5} \text{ eq/L} \]

\[ M = 1.35 \times 10^{-5} \text{ mol/L} \]

\[ N = 2.703 \times 10^{-5} \text{ eq/L} \]
7. Problem Statement: Known: You have acetic acid at a pH of 4.75.

Find:

What CH₃CO₂H species would predominate? At pH = 6? At pH = 3?

Assumptions:

Standard temperature and pressure

\[ \text{CH}_3\text{CO}_2\text{H} \rightleftharpoons \text{CH}_3\text{CO}_2^- + \text{H}^+ \]

\[ \text{pK}_a = 4.75 \]

\[ \text{Ka} = \frac{[\text{CH}_3\text{CO}_2^-][\text{H}^+]}{[\text{CH}_3\text{CO}_2\text{H}]} \]

\[ \text{pH} = \text{pK}_a + \log \left( \frac{[\text{HA}]}{[\text{H}_2\text{A}]} \right) \]

\[ \text{pH} = 4.75 = 4.75 + \log \left( \frac{[\text{CH}_3\text{CO}_2^-]}{[\text{CH}_3\text{CO}_2\text{H}]} \right) \]

\[ 1 = \frac{[\text{CH}_3\text{CO}_2^-]}{[\text{CH}_3\text{CO}_2\text{H}]} \]

\[ [\text{CH}_3\text{CO}_2\text{H}] = [\text{CH}_3\text{CO}_2^-] \]

\[ \text{pH} = 6 \]

\[ 6 = 4.75 + \log \left( \frac{[\text{CH}_3\text{CO}_2^-]}{[\text{CH}_3\text{CO}_2\text{H}]} \right) \]

\[ 1.25 = \log \left( \frac{[\text{CH}_3\text{CO}_2^-]}{[\text{CH}_3\text{CO}_2\text{H}]} \right) \]

\[ 10^{1.25} = \frac{[\text{CH}_3\text{CO}_2^-]}{[\text{CH}_3\text{CO}_2\text{H}]} = 17.75 \]

\[ \text{pH} = 3 \]

\[ 3 = 4.75 + \log \left( \frac{[\text{CH}_3\text{CO}_2^-]}{[\text{CH}_3\text{CO}_2\text{H}]} \right) \]

\[ 0.0174 = \frac{[\text{CH}_3\text{CO}_2^-]}{[\text{CH}_3\text{CO}_2\text{H}]} \]

\[ \text{pH} = 4.75 \text{ [CH}_3\text{CO}_2^- \rightleftharpoons [CH}_3\text{CO}_2\text{H}] \]

\[ \text{pH} = 6 \text{ CH}_3\text{CO}_2^- \text{ predominates} \]

\[ \text{pH} = 3 \text{ CH}_3\text{CO}_2\text{H} \text{ predominates} \]
8. **Problem Statement:** Convert from mg/l as the anion or compound to mg/l as CaCO₃.

**Solution:**

\[ MW_{(CaCO_3)} = 40 + 12 + 3(16) = 100 \ \text{g/mol} \]

\[ \left( \frac{100 \ \text{g/mol}}{2} \right) = 50 \ \text{g/mol} \]

\[ N = \frac{\text{Conc.}}{\text{EW}} \]

\[ N = \frac{200.00 \ \text{mg/l}}{18} \]

\[ MW_{(NH_4^+)} = 14 + 4(1) = 18 \ \text{g/mol} \]

\[ N = \frac{200.0 \ \text{mg/l}}{18} = 10.01 \ \text{eq/L}(50 \ \text{g/mol}) = \]

\[ 555.56 \ \text{mg/l as CaCO}_3 \]

b. 280.0 mg/l K⁺

\[ MW_{(K^+)} = 39 \ \text{g/mol} \]

\[ N = \frac{280.0 \ \text{mg/l}}{39} = (0.007 \ \text{eq/L})(50 \ \text{g/mol}) = \]

\[ 358.97 \ \text{mg/l as CaCO}_3 \]

c. 123.45 mg/l SO₄²⁻

\[ MW_{(SO_4^{2-})} = 32 + 4(16) = 96 \ \text{g/mol} \]

\[ N = \frac{123.45 \ \text{mg/l}}{96} = (10.0027 \ \text{eq/L})(50 \ \text{g/mol}) = \]

\[ 128.59 \ \text{mg/l as CaCO}_3 \]
8. (continued)

a. $85.05 \text{mg/L CA}_2^+$

$molar\ ( \text{CA}_2^+) = \frac{40 \text{g/mol}}{}$

$N = \frac{85.05\ \text{mg/L}}{40 \text{g/mol}} = 212.53 \text{mg/L as CaCO}_3$

b. $19.90 \text{mg/L Na}^+$

$molar\ ( \text{Na}^+) = \frac{23 \text{g/mol}}{}$

$N = \frac{19.90\ \text{mg/L}}{23 \text{g/mol}} = 43.26 \text{mg/L as CaCO}_3$

9. a. Problem statement: You have been asked to prepare a 1.0 L solution of NaOH. You add 5g of pure NaOH pellets to the 1.0 L beaker and fill with DI water.

Find:

What is the pH of the solution after equilibrium has been reached?

Assumptions:

Standard temperature and pressure

NaOH $\rightarrow$ Na$^+$ + OH$^-$ strong base, dissociates completely

Solve:

$Mw\ (\text{NaOH}) = 23 + 1 \times 1 = 40 \text{g/mol}$

$5g \times \frac{1\ \text{mol}}{40g} = 0.125\ \text{mol/L}$ $pKw = 14$
a. (continued)

Hydroxide ions are insignificant.

\[ \text{pOH} = -\log([\text{OH}^-]) \]
\[ \text{pOH} = -\log(0.125 \text{ mol/L}) = 1.92 \]
\[ \text{pH} = 14 - 0.903 = 13.01 \]
\[ \text{pH} = 13.01 \]

b. Empty statement known: you are asked to predict the pH of a 2.0 L bottle of DI water after adding a specific amount of acid, sulfuric acid, using a 1.0 M solution of H\(_2\)SO\(_4\) and a 0.5 cm\(^3\) syringe. You are told to add 15 full injections.

Find:

What is the pH of the solution?

Assumptions:

Strong acids dissociate completely. STP

\[
\text{H}_2\text{SO}_4 \rightarrow 2\text{H}^+ + \text{SO}_4^{2-}
\]

1 cm\(^3\) = 1 mL

0.5 cm\(^3\) = 0.5 mL \times 15 injections = 7.5 mL = 0.0075 L

0.0075 L \times 0.1 \text{ mol/L} = 0.00075 \text{ mol/L}

0.00075 \text{ mol/L} \times 3.75 \times 10^{-4} \text{ mol/L} \text{ H}_2\text{SO}_4

3.75 \times 10^{-4} \text{ mol/L} = 7.5 \times 10^{-4} \text{ mol/L H}^+

\[
\text{pH} = -\log([\text{H}^+])
\]
\[
\text{pH} = -\log(7.5 \times 10^{-4}) = 3.12
\]
9) You are in an environmental engineering laboratory:

a. You have been asked to prepare 1.0 liter of a sodium hydroxide (NaOH) solution. You fill a beaker with 1.0 liter of deionized water (DI) and add 5 grams of pure NaOH pellets to the beaker. What is the pH of the solution after equilibrium has been reached?

\[
\text{NaOH} \rightarrow \text{Na}^+ + \text{OH}^- \quad \text{Strong Base, therefore dissociates completely}
\]

**DISCOUNT** OH\(^-\) ions in DI water because the concentration will be very small compared to the amount added.

\[
\frac{5 \text{g}}{1 \text{L}} = \frac{5 \text{g}}{1 \text{L}} \rightarrow \frac{\frac{5 \text{g}}{40 \text{g/mole}}}{\text{mole}} = 0.125 \frac{\text{mole NaOH}}{\text{L}}
\]

Therefore \(\text{pOH} = -\log[\text{OH}^-] = -\log[0.125 \text{ mole/L}] = 0.903\)

\[
\text{pH} = 14 - \text{pOH} = 14 - 0.903 = 13.01
\]

b. Now you are asked to predict the pH of a 2.0 liter beaker of distilled water after adding a specific amount of acid. The type of acid used is sulfuric acid (H\(_2\)SO\(_4\)). Using a 0.1 M solution of H\(_2\)SO\(_4\) and a 0.5 cubic centimeter syringe you are told to add 15 full injections….what is the pH of the solution?

\[
\text{H}_2\text{SO}_4 \rightarrow 2\text{H}^+ + \text{SO}_4^{2-} \quad \text{Strong Acid, therefore dissociates completely}
\]

1 cubic centimeter (cc) = 1 milliliter (mL)

\[(0.5\text{mL/injection}) \times (15 \text{ injections}) = 7.5 \text{ mL of 0.1 M H}_2\text{SO}_4\]

So \(0.1 \frac{\text{mole H}_2\text{SO}_4}{\text{L}} \left(\frac{1 \text{ L}}{1000 \text{ mL}}\right) (7.5 \text{ mL}) = 7.5 \times 10^{-4} \text{ mole H}_2\text{SO}_4\)

Therefore the molar conc. in 2 liters is \(\frac{7.5 \times 10^{-4} \text{ mole H}_2\text{SO}_4}{2 \text{ L}} = 3.75 \times 10^{-4} \frac{\text{mole H}_2\text{SO}_4}{\text{L}}\)

And there are 2 moles of H\(^+\) for every mole of H\(_2\)SO\(_4\) so \(\frac{2 \text{ mole H}^+}{1 \text{ mole H}_2\text{SO}_4}\)\(\times\left(\frac{3.75 \times 10^{-4} \text{ mole H}_2\text{SO}_4}{\text{L}}\right) = 7.5 \times 10^{-4} \frac{\text{mole H}^+}{\text{L}}\)

Therefore \(\text{pH} = -\log[\text{H}^+] = -\log [7.5 \times 10^{-4} \text{ mole H}^+/\text{L}] = 3.12\)
Problem Statement: Suppose the gas above the soda in a bottle of soft drink is pure CO₂ at a pressure of 2 atm.

Find:
The CO₂ at 25°C.

Solve:

Henry's Law: \[ [\text{CO}_2] = K_H \cdot P_{\text{gas}} \]

\[ K_H = 0.0333 \text{ mol/(L atm)} \] at 25°C

\[ [\text{CO}_2] = 10.033 \text{ mol/L} \] (2 atm)

\[ [\text{CO}_2] = 0.0478 \text{ mol/L} \]

\[ b. \quad \text{Problem statement: known:} \]

\[ \text{CO}_2 + \text{H}_2\text{O} \rightarrow \text{H}^+ + \text{HCO}_3^- \]

Find:
Calculate the pH of the soft drink.

Assumptions:
Only cations and anions in the system are H⁺, OH⁻, and HCO₃⁻.

Solution will be somewhat acidic (negligible carbonates).

Solve:

\[ K_a = \frac{[\text{H}^+][\text{HCO}_3^-]}{[\text{CO}_2][\text{H}_2\text{O}]} \]

\[ pK_a (\text{H}_2\text{CO}_3) = 6.35 \]

\[ 10^{-6.35} = 4.4 \times 10^{-7} \]
10 (continued)

\[ K_a = \frac{[H^+] [HCO_3^-]}{[CO_2]} \]

\[ [H^+] = [HCO_3^-] = x \quad \text{(more than 0)} \]

\[ K_a = \frac{x^2}{[CO_2]} \]

\[ 4.47 \times 10^{-7} = \frac{x^2}{0.00378 \text{ mol/L}} \]

\[ x = 1.73 \times 10^{-4} \text{ mol/L} = [H^+] \]

\[
\begin{align*}
\text{pH} &= -\log [H^+] \\
\text{pH} &= -10 \log (1.73 \times 10^{-4}) \\
\text{pH} &= 3.70
\end{align*}
\]