

Problem Set 2—Key

CENG 340—Introduction to Environmental Engineering

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1. Arsenic

$$[\text{MCL}]_{\text{arsenic}} = 10 \text{ ppb}_m; \text{MW}_{\text{arsenic}} = 75 \frac{\text{g}}{\text{mole}}$$

(a)

$$[As] = 10 \text{ ppb}_m \times \frac{1 \text{ ppm}_m}{10^3 \text{ ppb}_m} = 0.01 \text{ ppm}_m = 0.01 \frac{\text{mg}}{\text{L}}$$

$$[As] = 0.01 \frac{\text{mg}}{\text{L}} \times \frac{1 \text{ mole}}{75 \text{ g}} = 1.3 \times 10^{-4} \frac{\text{mmole}}{\text{L}}$$

$$[As] = 0.01 \frac{\text{mg}}{\text{L}} \times 10^3 \frac{\mu\text{g}}{\text{mg}} = 10 \frac{\mu\text{g}}{\text{L}}$$

$$[As] = 1.3 \times 10^{-4} \frac{\text{mmole}}{\text{L}} \times 10^6 \frac{\text{nmole}}{\text{mmole}} = 130 \frac{\text{nmole}}{\text{L}}$$

(b) Exposure to arsenic may lead to skin damage, problems with circulatory systems, and increased risk of cancer. (Source: water.epa.gov).

(c) Naturally occurring arsenic has been found in water worldwide in every continent (Table 10.5 in the text book).

Also, (not required for PSet) arsenic in drinking water sources comes from orchard runoff, electronics and glass production. (source: water.epa.gov).

2. Turbid water is opaque, hazy, and cloudy. Small particles in water cause it to be turbid. In other words, turbid water is dirty.

3. (a) $TS = 45 + 30 + 100 + 10 + 20 + 25 = 230 \text{ mg/L}$
 (b) $TDS = 45 + 30 + 10 + 20 = 105 \text{ mg/L}$
 (c) $TSS = 100 + 25 = 125 \text{ mg/L}$
 (d) $VSS = 25 \text{ mg/L}$
 (e) $FDS = 45 + 30 + 10 = 85 \text{ mg/L}$

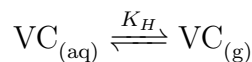
$$4. [VC]_{(aq)} = \frac{2 \text{ mg}}{60 \text{ mL}} = 0.05 \frac{\text{g}}{\text{L}}$$

$$MW_{VC} = 62.5 \frac{\text{g}}{\text{mole}}$$

$$[VC]_{(aq)} = 0.05 \frac{\text{g}}{\text{L}} \times \frac{1 \text{ mole}}{62.5 \text{ g}} = 8 \times 10^{-4} \frac{\text{mole}}{\text{L}}$$

$$K_H = 10^{0.04} \frac{\frac{\text{mole}_g}{\text{L}_g}}{\frac{\text{mole}_{aq}}{\text{L}_{aq}}} \text{ (definition of dimensionless } K_H \text{ found on p. 67 of the text book.)}$$

(a)



Write the equilibrium relationship of aqueous and gaseous VC—make sure you write the equation in such a way that the units of K_H work out.

$$K_H = \frac{VC_{(g)}}{VC_{(aq)}} = 10^{0.04} \frac{\frac{\text{mole}_g}{\text{L}_g}}{\frac{\text{mole}_{aq}}{\text{L}_{aq}}}$$

Calculate the concentration of VC (moles/L) in the air from Henry's constant:

$$VC_{(g)} = K_H \times VC_{(aq)} = 10^{0.04} \times 8 \times 10^{-4} = 8.8 \times 10^{-4} \frac{\text{mole}}{\text{L}} \text{ gas}$$

Volume of the air = 100 m^3 . Use this and calculate total moles of VC in the air:

$$\text{moles VC} = 8.8 \times 10^{-4} \frac{\text{mole}}{\text{L}} \times 100 \text{ m}^3 \times \frac{1000 \text{ L}}{\text{m}^3} = 88 \text{ moles VC in the air}$$

Use the ideal gas law and calculate the Volume of VC in the air:

$$V_{\text{VC}} = \frac{nRT}{P} = \frac{88 \text{ moles} \times 8.205 \times 10^{-5} \frac{\text{m}^3 \text{atm}}{\text{moleK}} \times 298.15 \text{ K}}{1 \text{ atm}} = 2.14 \text{ m}^3$$

Calculate the volume fraction and multiply by 10^6 to obtain the gaseous concentration of VC in units of ppm_v

$$[\text{VC}]_g = \frac{V_{\text{VC}}}{V_{\text{tot}}} \times 10^6 = \frac{2.1 \text{ m}^3}{100 \text{ m}^3} \times 10^6 = 21,500 \text{ ppm}_v$$

$$21,500 \text{ ppm}_v \gg 10 \text{ ppm}_v$$

BAD NEWS!

- (b) The students should have put the open (or unsealed bottle) in the *fume hood* as quickly as possible and told everyone to leave the lab immediately.

5. (16 pts) Assume atmospheric pressure of 1 atm, and convert atmospheric CO_2 concentration of 390 ppm_v to partial pressure of CO_2

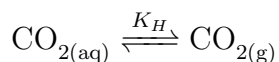
$$[\text{CO}_2]_g = 390 \text{ ppm}_v \times \frac{1 \text{ atm}}{10^6 \text{ ppm}_v} = 3.9 \times 10^{-4} \text{ atm}$$

$[\text{CO}_2]_{(g)}$ values between 310 and 400 ppm_v acceptable. Alternatively, $P_{\text{CO}_2} = 10^{-3.5} \text{ atm}$ is acceptable as well.

$$K_H = 29.4 \frac{\text{L} \times \text{atm}}{\text{mole}}$$

$$\text{MW}_{\text{CO}_2} = 44 \frac{\text{g}}{\text{mole}}$$

- (a) Step 1: Write equilibrium relationship for partitioning of CO_2 from the gas to the aqueous phase. Write the reaction in such a way that the units of K_H work out.



- (b) Calculate the dissolved aqueous concentration of CO₂ using Henry's Law constant and the atmospheric concentration of CO₂, which you can find online.

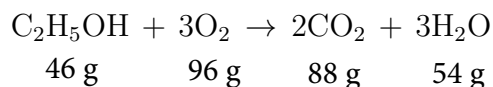
$$K_H = \frac{[\text{CO}_2]_{(g)}}{[\text{CO}_2]_{(aq)}}$$

$$[\text{CO}_2]_{(aq)} = \frac{[\text{CO}_2]_{(g)}}{K_H} = \frac{3.9 \times 10^{-4} \text{ atm}}{29.4 \frac{\text{L} \times \text{atm}}{\text{mole}}} = 1.33 \times 10^{-5} \frac{\text{moles}}{\text{L}}$$

$$[\text{CO}_2]_{(aq)} = 1.33 \frac{\text{moles}}{\text{L}} \times 44 \frac{\text{g}}{\text{mole}} \times \frac{1000 \text{ mg}}{\text{g}} = 0.58 \frac{\text{mg}}{\text{L}}$$

- (c) When the concentration of dissolved aqueous CO₂ increases, the concentration of H⁺ also increases and pH decreases.

6. Balance the reaction:



Convert lbs of ethanol to kg of ethanol:

$$500 \text{ lbs ethanol} \times \frac{1 \text{ kg}}{2.2 \text{ lbs}} = 227 \text{ kg ethanol}$$

- (a) Use the mass ratios noted above below the equation to calculate the mass of O₂ required for biodegradation:

$$\text{Mass of O}_2 = 227 \text{ kg eth} \times \frac{96 \text{ g O}_2}{46 \text{ g eth}} = 474 \text{ kg O}_2$$

(b) Calculate the mass of CO₂ produced:

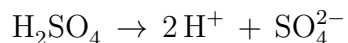
$$\text{Mass of CO}_2 = 227 \text{ kg eth} \times \frac{88 \text{ g O}_2}{46 \text{ g eth}} = 435 \text{ kg CO}_2$$

(c) Convert kg of CO₂ to moles. Then use the ideal gas law to convert to volume of CO₂.

$$\text{Moles of CO}_2 = 435 \text{ kg CO}_2 \times \frac{1000 \text{ g}}{\text{kg}} \times \frac{1 \text{ mole}}{44 \text{ g}} = 9881 \text{ mole}$$

$$V = \frac{nRT}{P} = \frac{9881 \times 8.205 \times 10^{-5} \frac{\text{m}^3 \text{atm}}{\text{moleK}} \times 303.15 \text{ K}}{1 \text{ atm}} = 245 \text{ m}^3$$

7. H₂SO₄ is a strong acid:



(a) To calculate moles of [H⁺], calculate number of moles of H₂SO₄ and multiply by 2:

$$10 \frac{\text{mg H}_2\text{SO}_4}{\text{L}} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ moles}}{98 \text{ g}} \times \frac{2 \text{ mole H}^+}{\text{mole H}_2\text{SO}_4} = 2 \times 10^{-4} \frac{\text{mole}}{\text{L}} \text{H}^+$$

$$\text{pH} = -\log[\text{H}^+] = -\log(2 \times 10^{-4}) = 3.7$$

(b) Normality of H₂SO₄ solution equal the moles per Liter of H₊, which equals $2 \times 10^{-4} \frac{\text{eq}}{\text{L}}$

(c) Convert [HCO₃⁻] from mg/L as CaCO₃ to eq/L:

$$[\text{HCO}_3^-] = 90 \frac{\text{mg CaCO}_3}{\text{L}} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ eq}}{50 \text{ g}} = 1.8 \times 10^{-3} \frac{\text{eq}}{\text{L}}$$

Use:

$$C_{\text{H}^+} \times V_{\text{H}^+ \text{ solution}} = C_{\text{bicarb}} \times V_{\text{bicarb solution}}$$

$$V_{\text{bicarb solution}} = \frac{C_{\text{H}^+} \times V_{\text{H}^+ \text{ solution}}}{C_{\text{bicarb}}} = \frac{2 \times 10^{-4} \frac{\text{eq}}{\text{L}} \times 100 \text{ mL}}{1.8 \times 10^{-3}} \frac{\text{eq}}{\text{L}} = 11 \text{ mL}$$

8. First Order Reaction:

$$A = A_0 \times e^{-kt}$$

(a) 90% of A destroyed, so 10% of initial concentration remains.

$$0.1A_0 = A_0 \times e^{-0.1t}$$

$$t = \frac{\ln(\frac{0.1}{1})}{-0.1 \text{ day}^{-1}} = 23 \text{ days}$$

(b) Same as part (a) except that 99% destroyed, so 1% remaining:

$$t = \frac{\ln(\frac{0.01}{1})}{-0.1 \text{ day}^{-1}} = 46 \text{ days}$$

(c) Same as part (a) except that 99.9% destroyed, so 0.1% remaining:

$$t = \frac{\ln(\frac{0.001}{1})}{-0.1 \text{ day}^{-1}} = 69 \text{ days}$$