

## **Problem Set 2b—KEY**

CEEG 340—Introduction to Environmental Engineering

Instructor: Deborah Sills

Fall 2017

### **Questions**

1. (15 pts) Text: 2.14—NEXT PAGE
2. (15 pts) Text: 2.20—NEXT PAGE

**2.14** Total mercury concentrations in the San Francisco Bay Area are reported to be 1.25 ng/L in water, 8 mg/L in rain, 2.1 mg/m<sup>3</sup> in air, and 250 ng in one gram of dry sediment. (a) Report all these concentrations in ppt. Assume the air temperature is 20°C.

$$1.25 \frac{\text{ng}}{\text{L}} \text{Hg} \times \frac{1 \text{ g}}{10^9 \text{ ng}} \times \frac{1.0 \text{ L H}_2\text{O}}{1000 \text{ g H}_2\text{O}} \times \frac{1 \text{ g}}{1000 \text{ mg}} = \frac{1.25 \times 10^{-12} \text{ g}}{\text{g}} \times \frac{10^{12} \text{ ppt}}{\text{mass fraction}} = \mathbf{1.25 \text{ ppt Hg}}$$

$$8 \frac{\text{mg}}{\text{L}} \text{Hg} = 8 \text{ ppm Hg} \times \frac{10^6 \text{ ppt}}{1 \text{ ppm}} = \mathbf{8.0 \times 10^6 \text{ ppt Hg}}$$

Solving for ppm<sub>v</sub> in equation 2.9:

$$\text{ppm}_v = \frac{\mu\text{g}}{\text{m}^3} \times \frac{1}{\text{MW}} \times \frac{RT}{1000P}$$

$$\begin{aligned} \text{ppm}_v &= 2.1 \frac{\text{mg}}{\text{m}^3} \times \frac{10^3 \mu\text{g}}{\text{mg}} \times \frac{1 \text{ mole}}{200.6 \text{ g Hg}} \times \frac{0.08205 \text{ atm} \times (20^\circ\text{C} + 273)}{\text{mol K} \times 1000 (1 \text{ atm})} \\ &= 0.217 \text{ ppm}_v \times \frac{10^6 \text{ ppt}}{1 \text{ ppm}} = \mathbf{252 \times 10^3 \text{ ppt}_v \text{ Hg}} \end{aligned}$$

$$250 \frac{\text{ng}}{\text{g sediment}} = \mathbf{250 \text{ ppb} = 250,000 \text{ ppt}}$$

Solutions Manual prepared by: Colleen Naughton, Ziad Katirji, Heather E. Wright Wendel, and James Mihelcic

*Environmental Engineering: Fundamentals, Sustainability, Design, 2<sup>nd</sup> Edition*

James R. Mihelcic and Julie Beth Zimmerman, John Wiley & Sons, New York, 2014.

**2.20** Assume that the average concentration of chlordane – a chlorinated pesticide now banned in the United States – in the atmosphere above the Arctic Circle in Norway was found to be  $0.6 \text{ pg/m}^3$ . In this measurement, approximately 90 percent of this compound is present in the gas phase; the remainder is adsorbed to particles. For this problem, assume that all the compound occurs in the gas phase, the humidity is negligibly low, and the average barometric pressure is 1 atm. Calculate the partial pressure of chlordane. The molecular formula for chlordane is  $\text{C}_{10}\text{Cl}_8\text{H}_6$ . The average air temperature through the period of measurement was  $-5^\circ\text{C}$ .

Solution:

First, find concentration in  $\text{ppm}_v$ . The molecular weight equals 409.8.

$$\frac{0.6 \text{ pg}}{\text{m}^3} \times \frac{\mu\text{g}}{10^6 \text{ pg}} = 0.6 \times 10^{-6} \mu\text{g} / \text{m}^3$$

$$\text{ppm}_v = nRT/PV = Q*(RT)/MW*P; \text{MW}=\text{mass}/n; Q=\text{mass}/V$$

$$\text{ppm}_v = \frac{\left(0.6 \times 10^{-6} \frac{\mu\text{g}}{\text{m}^3}\right) \left(0.08205 \frac{\text{L-atm}}{\text{mole-K}}\right) \times 268 \text{ K}}{\left(409.8 \frac{\text{g}}{\text{mole}}\right) \times \left(1000 \frac{\text{L}}{\text{m}^3}\right) \times 1 \text{ atm}}$$

$$\text{ppm}_v = 3.2 \times 10^{-11}$$

Now find partial pressure,

$$3.2 \times 10^{-11} = \frac{P_i}{1 \text{ atm}} \times 10^6$$

$$\text{Solve for } \boxed{P_i = 3.2 \times 10^{-17} \text{ atm}}$$

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3. (25 pts) **Who loves hockey?** (modified from Mihelcic and Zimmerman)

Ice resurfacing machines (aka Zambonis) use internal combustion vehicles that give off exhaust containing carbon monoxide (CO) and nitrogen oxides (NO<sub>x</sub>). The outdoor air quality 1-h standard of CO is set at 35 mg/m<sup>3</sup>. Average CO concentrations measured at Lynah Rink (at Cornell University) have been reported to be as high as 115 ppm<sub>v</sub> and as low as 35 ppm<sub>v</sub>. (1) Should Prof. Sills be concerned about spending 1 h at Lynah Rink when she watches Cornell's Women's Hockey Team beat Harvard this season? Assume the temperature equals 20°C. (2) Calculate the partial pressure (in atm) of CO in the rink. Assume that the atmospheric pressure is 1 atm.

*Part One:*

- **Step 1: Convert [CO]<sub>std</sub> from  $\frac{\text{mg}}{\text{m}^3}$  to  $\frac{\text{mole}}{\text{m}^3}$**

Molecular weight of CO = 12 + 16 = 28  $\frac{\text{g}}{\text{mole}}$

$$[\text{CO}]_{\text{std}} = 35 \frac{\text{mg}}{\text{m}^3} \times \frac{1 \text{ mole}}{28 \text{ g}} \times \frac{1 \text{ g}}{1000 \text{ mg}} = 1.25 \times 10^{-3} \frac{\text{mole}}{\text{m}^3}$$

- **Step 2: Convert [CO]<sub>std</sub> from  $\frac{\text{mole}}{\text{m}^3}$  to  $\frac{\text{m}^3}{\text{m}^3}$**

Use the ideal gas law: PV = nRT, where (1) temperature in Kelvin (K) = temperature in Celsius (°C) + 273.15, (2) the ideal gas constant R = 8.205 × 10<sup>-5</sup>  $\frac{\text{m}^3 \times \text{atm}}{\text{mole} \times \text{Kelvin}}$ , and (3) P = 1 atm.

$$\begin{aligned} [\text{CO}]_{\text{std}} &= \frac{1.25 \times 10^{-3} \text{ moles CO} \times \frac{R \times T}{P}}{\text{m}^3 \text{ air}} = \\ &= \frac{1.25 \times 10^{-3} \text{ moles CO} \times \frac{8.205 \times 10^{-5} \frac{\text{m}^3 \times \text{atm}}{\text{mole} \times \text{K}} \times 293.15 \text{ K}}{1 \text{ atm}}}{\text{m}^3 \text{ air}} = 3.01 \times 10^{-5} \frac{\text{m}^3 \text{ CO}}{\text{m}^3 \text{ air}} \end{aligned}$$

- **Step 3: Convert [CO]<sub>std</sub> from  $\frac{\text{m}^3}{\text{m}^3}$  to ppm<sub>v</sub>:**

$$[\text{CO}]_{\text{std}} = 3.01 \times 10^{-5} \frac{\text{m}^3 \text{ CO}}{\text{m}^3 \text{ air}} \times 10^6 \frac{\text{ppm}_v}{\frac{\text{m}^3}{\text{m}^3}} = 30 \text{ ppm}_v$$

At Lynah [CO] ranges from 35 to 115 ppm<sub>v</sub>—violates the air quality standard all of the time. Prof. Sills might want to bring a gas mask.

*Part Two:*

$$[\text{CO}] \text{ in units of ppm}_v = \frac{P_i}{P_{\text{air}}} \times 10^6$$

$$P_i(\text{low}) = \frac{35 \text{ ppm}_v \times 1 \text{ atm}}{10^6} = 3.5 \times 10^{-5} \text{ atm}$$

$$P_i(\text{high}) = \frac{115 \text{ ppm}_v \times 1 \text{ atm}}{10^6} = 1.15 \times 10^{-4} \text{ atm}$$

The partial pressure of CO at Lynah Rink ranges from  $3.5 \times 10^{-5}$  atm to  $1.15 \times 10^{-4}$  atm.

4. (25 pts)[VC]

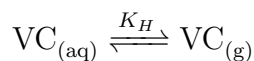
$$[\text{VC}]_{(\text{aq})} = \frac{3 \text{ mg}}{60 \text{ mL}} = 0.05 \frac{\text{g}}{\text{L}}$$

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

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

$$K_H = 26.8 \text{ L}\cdot\text{atm}/\text{mole}$$

(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.

$$K_H = \frac{\text{VC}_{(\text{g})}}{\text{VC}_{(\text{aq})}} = 26.8 \text{ L}\cdot\text{atm}/\text{mole}$$

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

$$\text{VC}_{(\text{g})} = K_H \times \text{VC}_{(\text{aq})} = 26.8 \times 8.0 \times 10^{-4} = 0.0214 \text{ atm}$$

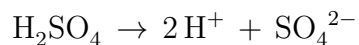
$$[\text{VC}]_g = \frac{P_{\text{VC}}}{P_{\text{tot}}} \times 10^6 = \frac{0.021 \text{ atm}}{1 \text{ atm}} \times 10^6 = 21,400 \text{ ppm}_v$$

$$21,400 \text{ ppm}_v \gg 10 \text{ ppm}_v$$

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. (20 pts)  $\text{H}_2\text{SO}_4$  is a strong acid:



- (a) To calculate moles of  $[\text{H}^+]$ , calculate number of moles of  $\text{H}_2\text{SO}_4$  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  $\text{H}_2\text{SO}_4$  solution equal the moles per Liter of  $\text{H}^+$ , which equals  $2 \times 10^{-4} \frac{\text{eq}}{\text{L}}$