



AIR POLLUTION PROBLEMS (I) SOLUTIONS

Problem 1.-

A part per trillion by volume (ppt_v) is:

$$\text{ppt}_v = \frac{1 \text{ volume pollutant}}{10^{12} \text{ volume of air}} = \frac{1 \text{ L pollutant}}{10^{12} \text{ L of air}} = \frac{1 \text{ nL pollutant}}{\text{m}^3 \text{ of air}}$$

Assuming that Ideal Gas Law ($P V = n R T$) is accurate under these conditions (298 K and 1 atmosphere), one mole of a gas would occupy:

$$V_m = \frac{R T}{P}$$

Where: P=pressure
V= volume
n= mole
R= gas constant
T= temperature

$$V_m = \frac{0.082 \frac{\text{atm L}}{\text{mol K}} \cdot 298 \text{ K}}{1 \text{ atm}} = 24.44 \text{ L}$$

The molecular weight of dimethyl sulfide (DMS) is $62 \text{ g}\cdot\text{mol}^{-1}$. Thus, atmospheric concentrations of DMS between 250 and $500 \text{ ng}\cdot\text{m}^{-3}$ equal to:

$$250 \frac{\text{ng DMS}}{\text{m}^3} \cdot \frac{1 \text{ g}}{10^9 \text{ ng}} \cdot \frac{1 \text{ mol DMS}}{62 \text{ g}} \cdot \frac{24.44 \text{ L}}{1 \text{ mol}} \cdot \frac{10^9 \text{ nL}}{1 \text{ L}} = \frac{98.5 \text{ nL DMS}}{\text{m}^3} = \mathbf{98.5 \text{ ppt}}$$

$$500 \frac{\text{ng DMS}}{\text{m}^3} \cdot \frac{1 \text{ g}}{10^9 \text{ ng}} \cdot \frac{1 \text{ mol DMS}}{62 \text{ g}} \cdot \frac{24.44 \text{ L}}{1 \text{ mol}} \cdot \frac{10^9 \text{ nL}}{1 \text{ L}} = \frac{98.5 \text{ nL DMS}}{\text{m}^3} = \mathbf{197.5 \text{ ppt}}$$

Problem 2.-

Considering that there are 2.5 million vehicles on the move and that these vehicles travel a distance of 12,000 km per year and the NO_x emission factor of the fleet of this city is 2.1 grams of NO_x per kilometer, the total annual and daily NO_x emissions are:





$$2.5 \cdot 10^6 \text{ vehicles} \cdot \frac{12 \text{ km vehicle}}{\text{year}} \cdot \frac{2.1 \text{ g NO}_x}{\text{km}} = \frac{6.3 \cdot 10^{10} \text{ g NO}_x}{\text{year}}$$

$$\frac{6.3 \cdot 10^{10} \text{ g NO}_x}{\text{year}} \cdot \frac{\text{year}}{365 \text{ days}} = \frac{1.73 \cdot 10^8 \text{ g NO}_x}{\text{day}}$$

The city road network totals 1,600 km. If we assume that the traffic is homogeneously distributed over the city, the daily NO_x emissions per square meter are:

$$\frac{1.73 \cdot 10^8 \frac{\text{g NO}_x}{\text{day}}}{400 \text{ km}^2 \cdot \frac{10^6 \text{ m}^2}{1 \text{ km}^2}} = \frac{0.432 \text{ g NO}_x}{\text{day} \cdot \text{m}^2}$$

Problem 3.-

There are 92 people in a public establishment, of these, half are smokers and, on average they smoke 2 cigarettes per hour and that each cigarette, when smoked, releases 1.4 milligrams of formaldehyde, we get:

$$46 \text{ smokers} \cdot \frac{2 \text{ cigarettes}}{\text{h}} \cdot \frac{1.4 \text{ mg HCHO}}{\text{cigarette}} \cdot \frac{1 \text{ mg}}{10^3 \text{ mg}} \cdot \frac{90 \text{ g}}{100 \text{ g}} \cdot \frac{1 \text{ mol}}{30 \text{ g HCHO}} = \frac{0.039 \text{ mol HCHO}}{\text{h}}$$

The ambient temperature and the pressure in the establishment are 25°C and 100.66 kPa, respectively. Assuming that Ideal Gas Law ($PV = nRT$) is accurate under these conditions, 0.0039 mol of formaldehyde at 298.15 K at 100.66 kPa would occupy:

$$V = \frac{0.0039 \text{ mol HCHO} \cdot \frac{0.082 \text{ atm L}}{\text{mol K}} \cdot 298.15 \text{ K}}{100.66 \text{ kPa} \cdot \frac{1 \text{ atm}}{101.3 \text{ kPa}}} = 0.09452 \text{ L HCHO}$$

The volume of the 16 x 12 x 2.5 meters establishment is 480 m³.

A part per million by volume (ppm_v) is:

$$\text{ppm}_v = \frac{1 \text{ volume pollutant}}{10^6 \text{ volume of air}} = \frac{1 \text{ L pollutant}}{10^6 \text{ L of air}} = \frac{1 \text{ mL pollutant}}{\text{m}^3 \text{ of air}}$$

Therefore, the concentration of formaldehyde in units of parts per million is:

$$\frac{0.094 \text{ L HCHO}}{480 \text{ m}^3} \cdot \frac{1000 \text{ mL}}{1 \text{ L}} = 0.197 \text{ ppm}_v \text{ HCHO}$$

When this toxic compound is present in the air at levels at or above 0.05 ppm_v, it can lead to acute physical symptoms such as watery eyes, burning sensations in the eyes and other irritating effects. Assuming that 10% (volume) of the formaldehyde escapes through opened windows and doors and that 60% (volume) of the formaldehyde retained inside the building is



oxidized to carbon dioxide, the real concentration of formaldehyde is 0.0787 ppm_v, which is above 0.05 ppm_v.

$$0.197 \text{ ppm}_v \text{ HCHO} \cdot \frac{40 \text{ ppm}_v \text{ HCHO not oxidated}}{100 \text{ ppm}_v \text{ HCHO}} = 0.0787 \text{ ppm}_v \text{ HCHO}$$

And, the concentration of CO resulting from the oxidation of the formaldehyde in the air is:

$$0.197 \text{ ppm}_v \text{ HCHO} \cdot \frac{40 \text{ ppm}_v \text{ oxidated to CO}_2}{100 \text{ ppm}_v \text{ HCHO}} = \mathbf{0.1182 \text{ ppm}_v \text{ CO}_2}$$

Problem 4.-

According to the box model, if we assume steady state emissions and atmospheric conditions and upwind background concentrations of the pollutant are insignificant, the concentration of is given by:

$$C_{\text{pollutant}} = \frac{\Delta x \cdot Q}{z \cdot u}$$

Where: Δx =distance over which the emissions take place

Q = area emission rate

z = mixing height

u = average wind speed through the vertical extent of the box

If the city emits 0.012 mg·m⁻²·s⁻¹ of sulfur dioxide, it extends 15 km in the wind direction and 10 km perpendicularly the wind speed is 3 m·s⁻¹ and the pollutants are mixed up to 600 m altitude, the concentration of SO₂ is:

$$[\text{SO}_2] = \frac{15 \text{ km} \cdot \frac{10^3 \text{ m}}{1 \text{ km}} \cdot \frac{0.012 \text{ mg SO}_2}{\text{m}^2 \text{ s}} \cdot \frac{10^3 \mu\text{g}}{1 \text{ mg}}}{600 \text{ m} \cdot \frac{3 \text{ m}}{\text{s}}} = \frac{\mathbf{100 \mu\text{g SO}_2}}{\mathbf{m}^3}$$

Problem 5.-

A power plant with no sulfur removal equipment burns 30 tons of coal per hour to generate electricity. The analyses of the samples of the exhaust gases show that sulfur dioxide concentration is 2,000 ppm_v and that they are emitted at a temperature of 460°C and an overpressure of 0.05 atmospheres.

Applying the Ideal Gas Law ($PV = nRT$), we get that a mol of a gas at the conditions of this problem (1.05 atmosphere and 460°C) would occupy:

$$V_m = \frac{\frac{0.082 \text{ atm L}}{\text{mol K}} \cdot (460 + 273.15) \text{ K}}{1.05 \text{ atm}} = 57.26 \text{ L}$$



Thus, the SO₂ emitted per cubic meter is:

$$2000 \text{ ppm}_v \text{ SO}_2 \cdot \frac{64 \text{ g} \cdot \text{mol}^{-1}}{57.26 \text{ L}} = \frac{2235.6 \text{ mg SO}_2}{\text{m}^3} \cdot \frac{1 \text{ g}}{10^3 \text{ mg}} = \frac{2.2356 \text{ g SO}_2}{\text{m}^3}$$

As the exhaust gas flow is 25,000 m³·min⁻¹, the mass of SO₂ released per second is:

$$\frac{2.2356 \text{ g SO}_2}{\text{m}^3} \cdot \frac{25000 \text{ m}^3}{\text{min}} \cdot \frac{1 \text{ min}}{60 \text{ s}} = \frac{931.5 \text{ g SO}_2}{\text{s}}$$

To calculate the percentage of sulfur in coal, first, we calculate the kilograms of SO₂ emitted per hour. It is set by:

$$\frac{2.2356 \text{ g SO}_2}{\text{m}^3} \cdot \frac{3600 \text{ s}}{1 \text{ h}} \cdot \frac{1 \text{ kg}}{10^3 \text{ g}} = \frac{3353.38 \text{ kg SO}_2}{\text{h}}$$

Assuming that the reaction between one mole of S and molecular oxygen produces one mole of SO₂ and considering that the molecular weight of SO₂ (64 g·mol⁻¹) doubles the one of the sulfur (32g·mol⁻¹), the mass of S 'burned' per hour is:

$$\frac{3353.38 \text{ kg SO}_2}{\text{h}} \cdot \frac{1 \text{ kg S}}{2 \text{ kg SO}_2} = \frac{1676.69 \text{ kg S}}{\text{h}}$$

Thus, the percentage of sulfur in coal by mass is:

$$\frac{1676.69 \text{ kg S}}{\text{h}} \cdot \frac{1 \text{ h}}{30000 \text{ kg coal}} \cdot 100 = \mathbf{5.59\% \text{ S}}$$