Assignment 2 Solutions

Problems: Gilbert, Chapter1: #1.62, 1.70, Chapter2: #2.57, Chapter 3: #3.12, 3.22, 3.27, 3.44, 3.56, 3.62.

1.62. **Collect and Organize**

There is a small amount of mercury in this lake per liter, but the volume of the lake will be quite large. In this problem we have to find the volume of the lake and use the concentration of mercury in one liter of the lake water to find the total amount of mercury in the lake.

Analyze

The volume of the lake can be calculated from the surface area and the average depth. However, we need this answer in liters since the concentration of mercury is expressed in micrograms per liter. We should first convert square miles to square meters and the depth of the lake into meters. Then we can calculate the volume of the lake. We need these conversions and the formula for volume from surface area and depth:

$$\frac{1\ \text{mi}}{1.609\ \text{km}},\quad \frac{1\ \text{km}}{1000\ \text{m}},\quad \frac{1\ \text{ft}}{0.3048\ \text{m}}$$

$$\text{Volume of lake} = \text{surface area (m}^2) \times \text{depth (m)}$$

We then need to convert cubic meters into liters through

$$\frac{1\ \text{m}^3}{1000\ \text{L}}$$

Next we can use the concentration of the mercury to find the total mass of mercury in the lake (which will be in µmicrograms) and convert that to kilograms:

$$\text{Mass of mercury in lake } (\mu\text{g}) = \text{volume of lake} \times \frac{\text{mass of mercury } (\mu\text{g})}{\text{volume (L)}}$$

$$\text{Mass of mercury (kg)} = \text{mass of mercury } (\mu\text{g}) \times \frac{1\ \text{g}}{1 \times 10^6\ \mu\text{g}} \times \frac{1\ \text{kg}}{1000\ \text{g}}$$

Solve

$$\text{Surface area of lake (m}^2) = 10.0\ \text{mi}^2 \times \left(\frac{1.609\ \text{km}}{1\ \text{mi}}\right)^2 \times \left(\frac{1000\ \text{m}}{1\ \text{km}}\right)^2 = 2.59 \times 10^7\ \text{m}^2$$

$$\text{Depth of lake (m)} = 45\ \text{ft} \times \frac{0.3048\ \text{m}}{1\ \text{ft}} = 13.7\ \text{m}$$

$$\text{Volume of lake (m}^3) = \left(2.59 \times 10^7\ \text{m}^2\right) \times \left(13.7\ \text{m}\right) = 3.55 \times 10^8\ \text{m}^3$$

$$\text{Volume of lake (L)} = 3.55 \times 10^8\ \text{m}^3 \times \frac{1000\ \text{L}}{1\ \text{m}^3} = 3.55 \times 10^{11}\ \text{L}$$

The amount of mercury in the lake then is computed as

$$3.55 \times 10^{11}\ \text{L} \times \frac{0.33\ \mu\text{g}}{\text{L}} \times \frac{1\ \text{g}}{1 \times 10^6\ \mu\text{g}} \times \frac{1\ \text{kg}}{1000\ \text{g}} = 120\ \text{kg}$$

Think About It

Although the concentration of the mercury is quite low, the entire lake contains a relatively large amount of mercury.

1.70. **Collect and Organize**

We are to express the result of each calculation to the correct number of significant figures.

Analyze

The rules regarding the significant figures that carry over in calculations are given in Section 1.8 in the textbook. Remember to operate on the weak-link principle.

Solve

(a) The least well-known value has two significant figures so the calculator result of 1.5506 $\times 10^{-1}$ is reported as 1.5×10^{-1}.
(b) The least well-known value has three significant figures so the calculator result of 146.3988 is reported as 146.
(c) The least well-known value has three significant figures so the calculator result of 2.25857×10^{-2} is reported as 2.26×10^{-2}.
(d) The least well-known value has three significant figures so the calculator result of 3.5700 $\times 10^{3}$ is reported as 3.57×10^{3}.

Think About It

Indicating the correct number of significant figures for a calculated value indicates the level of confidence we have in our calculated value. Reporting too many significant figures would indicate a higher level of precision in our number than we actually have.

2.57. **Collect and Organize**

All the compounds here are oxides of nitrogen. These are all molecular compounds composed of two nonmetallic elements. We name these using the rules for binary compounds.

Analyze

We will use prefixes (Table 2.2) to indicate the number of oxygen atoms in these compounds. The nitrogen atom is always first in the formula; so it is named first. If there is only one nitrogen atom in the formula, we do not need to use the prefix *mono-* for the nitrogen. If there is more than one nitrogen atom, however, we will indicate the number with the appropriate prefix. Also, since *oxide* begins with a vowel, it would be awkward to say *pentaoxide;* so we shorten the double vowel in this part of the chemical name to *pentoxide.*

Solve

(a) NO_3, nitrogen trioxide
(b) N_2O_5, dinitrogen pentoxide
(c) N_2O_4, dinitrogen tetroxide
(d) NO_2, nitrogen dioxide
(e) N_2O_3, dinitrogen trioxide
(f) NO, nitrogen monoxide
(g) N_2O, dinitrogen monoxide
(h) N_4O, tetranitrogen monoxide

Think About It

All of these many binary compounds of nitrogen and oxygen are uniquely named.

3.12. **Collect and Organize**

In this exercise, we convert the number of molecules of each gas found in the sample to moles.

Analyze

To convert the number of molecules to moles, we divide by Avogadro's number.

Solve

(a) $\frac{1.4 \times 10^{13} \text{ molecules of } H_2}{6.022 \times 10^{23} \text{ molecules/mol}} = 2.3 \times 10^{-11} \text{ mol } H_2$ (c)

$\frac{7.7 \times 10^{12} \text{ molecules of } N_2O}{6.022 \times 10^{23} \text{ molecules/mol}} = 1.3 \times 10^{-11} \text{ mol } N_2O$

(b) $\frac{1.5 \times 10^{14} \text{ atoms of He}}{6.022 \times 10^{23} \text{ atoms/mol}} = 2.5 \times 10^{-10} \text{ mol He}$ (d)

$\frac{3.0 \times 10^{12} \text{ molecules of CO}}{6.022 \times 10^{23} \text{ molecules/mol}} = 5.0 \times 10^{-12} \text{ mol CO}$

Think About It

The trace gas that has the highest number of atoms or molecules present also has the highest number of moles present. In this sample of air, the amount of the trace gases decreases in the order He > H_2 > N_2O > CO.

3.22. **Collect and Organize**

We are asked to convert a mass of gold in ounces to moles.

Analyze

We need the mass of 1 mol of gold to compute the number of moles of gold in the 2.00 oz sample. From the periodic table, we see that the molar mass of gold is 196.967 g/mol. In addition, we have to convert the mass of the gold in ounces to the mass in grams using 1 oz = 28.35 g.

Solve

$$2.00 \text{ oz Au} \times \frac{28.35 \text{ g}}{1 \text{ oz}} \times \frac{1 \text{ mol}}{196.967 \text{ g}} = 0.288 \text{ mol Au}$$

Think About It

Because gold's molar mass is relatively high at 197 g/mol, a few ounces do not contain many moles of gold atoms.

3.27. **Collect and Organize**

This exercise has us compute the molar mass of various molecular compounds of oxygen.

Analyze

The molar mass of each of the compounds can be found by adding the molar mass of each element from the periodic table, taking into account the number of moles of each atom present in 1 mol of the substance.

Solve

(a) SO_2: 32.065 + 2(15.999) = 64.063 g/mol

(b) O_3: 3(15.999) = 47.997 g/mol

(c) CO_2: 12.011 + 2(15.999) = 44.009 g/mol

(d) N_2O_5: 2(14.007) + 5(15.999) = 108.009 g/mol

Think About It

Notice that three compounds, SO_2, O_3, and CO_2, have three atoms in their chemical formula, but each has a different molar mass.

3.44. **Collect and Organize**

To balance the chemical equations, we use the five steps described in the textbook.

Analyze

First we start with the correct formula for the reactants and products. Here we are given the equation, so we may skip this step. Next, we assign one of the reactants or products a coefficient of 1 (one with the most different elements in it). Next we assign coefficients to the other reactants and products so that the numbers of each type of atom on both sides of the equation are equal. If there are any fractional coefficients, we multiply the entire equation through to eliminate all fractions. We can also eliminate any 1s as coefficients since the formula itself means 1. Finally, check the balanced equation for equal numbers of atoms on both sides.

Solve

(a) Step 2: ___$SO_2(g)$ + ___$O_2(g) \rightarrow$ 1 $SO_3(g)$

Step 3: 1 $SO_2(g)$ + 1/2 $O_2(g) \rightarrow$ 1 $SO_3(g)$

Step 4: 2 $SO_2(g)$ + $O_2(g) \rightarrow$ 2 $SO_3(g)$

Step 5: 2 S + 6 O $\rightarrow$ 2 S + 6 O Balanced!

(b) Step 2: ___$H_2S(g)$ + ___$O_2(g) \rightarrow$ 1 $SO_2(g)$ + ___$H_2O(g)$

Step 3: 1 $H_2S(g)$ + 3/2 $O_2(g) \rightarrow$ 1 $SO_2(g)$ + 1 $H_2O(g)$

Step 4: 2 $H_2S(g)$ + 3 $O_2(g) \rightarrow$ 2 $SO_2(g)$ + 2 $H_2O(g)$

Step 5: 4 H + 2 S + 6 O $\rightarrow$ 2 S + 6 O + 4 H Balanced!

(c) Step 2: $H_2S(g)$ + 1 $SO_2(g) \rightarrow$ ___$S_8(s)$ + ___$H_2O(g)$

Step 3: 2 $H_2S(g)$ + 1 $SO_2(g) \rightarrow$ 3/8 $S_8(s)$ + 2 $H_2O(g)$

Step 4: 16 $H_2S(g)$ + 8 $SO_2(g) \rightarrow$ 3 $S_8(s)$ + 16 $H_2O(g)$

Step 5: 32 H + 24 S + 16 O $\rightarrow$ 24 S + 32 H + 16 O Balanced!

Think About It

Part c has large coefficients due to the formation of S_8, the most stable elemental form of sulfur.

3.56. **Collect and Organize**

We need to convert the given mass of carbon dioxide generated to moles of CO_2, then to the mass of carbon present in that amount of carbon dioxide.

Analyze

The mass of carbon dioxide we are given is in metric tons. We first have to convert this into grams using conversions for mass [1 t (metric ton) = 1000 kg and 1000 g = 1 kg]. We can then find the number of moles of carbon dioxide in this mass by dividing by the molar mass of CO_2. Because 1 mol of carbon is contained in 1 mol of carbon dioxide, the moles of carbon are equal to the moles of carbon dioxide. To find the mass of carbon present in the CO_2 in grams, we need only to multiply the moles of carbon by the molar mass of carbon.

Solve

(a) $$27 \times 10^{9}\ \text{t CO}_2 \times \frac{1000\ \text{kg}}{1\ \text{t}} \times \frac{1000\ \text{g}}{1\ \text{kg}} \times \frac{1\ \text{mol CO}_2}{44.01\ \text{g}} = 6.1 \times 10^{14}\ \text{mol CO}_2$$

(b) $6.1\times10^{14}\text{ mol CO}_2\times\frac{1\text{ mol C}}{1\text{ mol CO}_2}\times\frac{12.011\text{ g}}{1\text{ mol}}=7.4\times10^{15}\text{ g C}$

Think About It
This 7.4×10^{15} g of carbon is 7.4 billion metric tons of carbon that are generated each year. This mass amount of carbon is less than the mass amount of carbon dioxide. In evaluating the environmental impact of carbon dioxide emissions, it is important to know which species is being compared. The carbon mass will always be less than that of carbon dioxide.

3.62. **Collect and Organize**
We are looking at the differences in the amount of CO_2 produced from burning charcoal versus propane.

Analyze
For each fuel we need to write a balanced chemical equation for the combustion (reaction with oxygen), then calculate the amount of CO_2 produced from burning 500.0 g of each fuel. For this we need the molar masses of charcoal (carbon, 12.011 g/mol), propane (C_3H_8, 44.097 g/mol), and carbon dioxide (CO_2, 44.01 g/mol).

Solve
(a) For the combustion of charcoal

Step 1: ___C(s) + ___O_2(g) $\longrightarrow$ ___CO_2(g)
Step 2: ___C(s) + ___O_2(g) $\longrightarrow$ 1 CO_2(g)
Step 3: 1 C(s) + 1 O_2(g) $\longrightarrow$ 1 CO_2(g)
Step 4: C(s) + O_2(g) $\longrightarrow$ CO_2(g)
Step 5: 1 C + 2 O $\rightarrow$ 1 C + 2 O Balanced!

For the combustion of propane

Step 1: ___C_3H_8(g) + ___O_2(g) $\longrightarrow$ ___CO_2(g) + ___H_2O(g)
Step 2: 1 C_3H_8(g) + ___O_2(g) $\longrightarrow$ ___CO_2(g) + ___H_2O(g)
Step 3: 1 C_3H_8(g) + 5 O_2(g) $\longrightarrow$ 3 CO_2(g) + 4 H_2O(g)
Step 4: C_3H_8(g) + 5 O_2(g) $\longrightarrow$ 3 CO_2(g) + 4 H_2O(g)
Step 5: 3 C + 8 O + 10 O $\longrightarrow$ 3 C + 10 O + 8 H Balanced!

(b) For the combustion of charcoal

$$500.0\text{ g C}\times\frac{1\text{ mol C}}{12.011\text{ g}}\times\frac{1\text{ mol CO}_2}{1\text{ mol C}}\times\frac{44.01\text{ g CO}_2}{1\text{ mol}}=1832\text{ g CO}_2$$

For the combustion of propane

$$500.0\text{ g C}_3\text{H}_8\times\frac{1\text{ mol C}_3\text{H}_8}{44.097\text{ g}}\times\frac{3\text{ mol CO}_2}{1\text{ mol C}_3\text{H}_8}\times\frac{44.01\text{ g CO}_2}{1\text{ mol}}=1497\text{ g CO}_2$$

Think About It
Burning charcoal produces more carbon dioxide than the same mass of propane. The propane burns “cleaner” by mass.