01/21/14

Solution to the Assigned Problems of Workshop 1

Chem. 103, Spring 2014

Chapter 1: Chemistry - the Science of Change

Home-assigned problems: 1.49, 1.67, 1.75, 1.87, 1.91, 1.93, 1.99, 1.105, 1.107, 1.113

1.49 **Strategy:** The difference between the masses of the empty and filled bulbs is the mass of the gas in the bulb. The density is then determined using the density equation, $d = \frac{m}{V}$.

Solution:

243.22 g – 243.07 g = 0.15 g gas

$$d = \frac{m}{V} = \frac{0.15\text{ g}}{111.5\text{ mL}} = 1.3\times10^{-3}\text{ g/mL}$$

$$\frac{1.3\times10^{-3}\text{ g}}{1\cancel{\text{mL}}} \times \frac{1\times10^{3}\cancel{\text{mL}}}{\text{L}} = \mathbf{1.3\ g/L}$$

Because the density of gases are generally low, the density is typically expressed in g/L.

1.67 a. **Upper ruler: 2.5 cm** b. **Lower ruler: 2.55 cm**

1.75 **Strategy:** The difference between the masses of the empty and filled flasks is the mass of the water in the flask. The volume of the water (and the flask) can be found using the density equation.

Setup: Solve the density equation for V:

$$V = \frac{m}{d}$$

Solution:

87.39 – 56.12 = 31.27 g water

$$V = \frac{m}{d} = \frac{31.27\text{ g}}{0.9976\text{ g/cm}^3} = \mathbf{31.35\ cm^3}$$

1.87 **Strategy:** The volume of seawater is given. The strategy is to use the given conversion factors to convert L seawater → g seawater → g NaCl. This result can then be converted to kg NaCl and to tons NaCl. Note that 3.1% NaCl by weight means 100 g seawater = 3.1 g NaCl.

Setup: Use the conversion factors:

$$\frac{1000 \text{ mL seawater}}{1 \text{ L seawater}}, \frac{1.03 \text{ g seawater}}{1 \text{ mL seawater}}, \text{ and } \frac{3.1 \text{ g NaCl}}{100 \text{ g seawater}}$$

Solution:

$$1.5\times10^{21} \text{ L seawater} \times \frac{1000 \text{ m}\cancel{\text{L seawater}}}{1 \text{ L seawater}} \times \frac{1.03 \text{ g seawater}}{1 \text{ m}\cancel{\text{L seawater}}} \times \frac{3.1 \text{ g NaCl}}{100 \text{ g seawater}} = 4.8\times10^{22} \text{ g NaCl}$$

$$\textbf{mass NaCl (kg)} = 4.8\times10^{22} \cancel{\text{g NaCl}} \times \frac{1 \text{ kg}}{100\cancel{0 \text{ g}}} = \mathbf{4.8\times10^{19} \text{ kg NaCl}}$$

$$\textbf{mass NaCl (tons)} = 4.8\times10^{22} \cancel{\text{g NaCl}} \times \frac{\cancel{1 \text{ lb}}}{453.6 \cancel{\text{g}}} \times \frac{1 \text{ ton}}{2000 \cancel{\text{lb}}} = \mathbf{5.3\times10^{16} \text{ tons NaCl}}$$

1.91 a. **Strategy:** Use the percent error equation.

Setup: The percent error of a measurement is given by:

$$\frac{|\text{true value} - \text{experimental value}|}{\text{true value}} \times 100\%$$

Solution:

$$\frac{|0.798 \text{ g/mL} - 0.802 \text{ g/mL}|}{0.798 \text{ g/mL}} \times 100\% = \mathbf{0.5\%}$$

b. **Strategy:** Use the percent error equation.

Setup: The percent error of a measurement is given by:

$$\frac{|\text{true value} - \text{experimental value}|}{\text{true value}} \times 100\%$$

Solution:

$$\frac{|0.864 \text{ g} - 0.837 \text{ g}|}{0.864 \text{ g}} \times 100\% = \mathbf{3.1\%}$$

1.93 **Strategy:** Use the percent composition measurement to convert kg ore → g Cu. Note that 34.63% Cu by mass means 100 g ore = 34.63 g Cu.

Setup: Use the conversion factors:

$$\frac{34.63 \text{ g Cu}}{100 \text{ g ore}} \text{ and } \frac{1000 \text{ g}}{1 \text{ kg}}$$

Solution:

$$(5.11 \times 10^3 \text{ kg ore}) \times \frac{34.63 \text{ g Cu}}{100 \text{ g ore}} \times \frac{1000 \text{ g}}{1 \text{ kg}} = \mathbf{1.77 \times 10^6 \text{ g Cu}}$$

1.99 **Strategy:** To calculate the density of the pheromone, you need the mass of the pheromone, and the volume that it occupies. The mass is given in the problem.

Setup:

$$\text{volume of a cylinder} = \text{area} \times \text{height} = \pi r^2 \times h$$

Converting the radius and height to cm gives:

$$0.50 \cancel{\text{mi}} \times \frac{1609 \cancel{\text{m}}}{1 \cancel{\text{mi}}} \times \frac{1 \text{ cm}}{0.01 \cancel{\text{m}}} = 8.05 \times 10^4 \text{ cm}$$

$$5\ 40 \cancel{\text{ft}} \times \frac{12 \cancel{\text{in}}}{1 \cancel{\text{ft}}} \times \frac{2.54 \text{ cm}}{1 \cancel{\text{in}}} = 1.22 \times 10^3 \text{ cm}$$

Solution:

$$\text{volume} = \pi(8.05 \times 10^4 \text{ cm})^2 \times (1.22 \times 10^3 \text{ cm}) = 2.48 \times 10^{13} \text{ cm}^3$$

Density of gases is usually expressed in g/L. Let's convert the volume to liters.

$$(2.48 \times 10^{13} \cancel{\text{cm}^3}) \times \frac{1 \cancel{\text{mL}}}{1 \cancel{\text{cm}^3}} \times \frac{1 \text{ L}}{1000 \cancel{\text{mL}}} = 2.48 \times 10^{10} \text{ L}$$

$$\textbf{density} = \frac{\text{mass}}{\text{volume}} = \frac{1.0 \times 10^{-8} \text{ g}}{2.48 \times 10^{10} \text{ L}} = \mathbf{4.0 \times 10^{-19} \text{ g/L}}$$

1.105 **Gently heat the liquid to see if any solid remains after the liquid evaporates. Also, collect the vapor and then compare the densities of the condensed liquid with the original liquid. The composition of a mixed liquid frequently changes with evaporation along with its density.**

1.107 **Strategy:** As water freezes, it expands. First, calculate the mass of the water at 20°C. Then, determine the volume that this mass of water would occupy at –5°C.

Solution:

$$\text{Mass of water} = 242 \text{ mL} \times \frac{0.998 \text{ g}}{1 \text{ mL}} = 241.5 \text{ g}$$

$$\text{Volume of ice at } -5°C = 241.5 \text{ g} \times \frac{1 \text{ mL}}{0.916 \text{ g}} = 264 \text{ mL}$$

The volume occupied by the ice is larger than the volume of the glass bottle. The glass bottle would break.

1.113 a. **Strategy:** We are asked to determine when °C = °F. Since both °C and °F are used in the conversion factor, we can replace both °C and °F with a common variable, such as °C. Solving the algebraic equation for °C will yield the temperature at which the values are numerically equal.

Setup: Conversion from Celsius to Fahrenheit:

$$°F = \left(°C \times \frac{9°F}{5°C}\right) + 32°F$$

Solution: °C = °F

Replacing °F in the equation with °C yields:

$$°C = \left(°C \times \frac{9}{5}\right) + 32 \text{ or } °C = \frac{9}{5}°C + 32$$

Combine like terms to yield:

$$-\frac{4}{5}°C = 32$$

Solving for °C gives **–40°**

b. **Strategy:** We are asked to determine when °F = K. If °F = K, we can set the conversion factors equal to one another. Then, we solve for °C which is the variable common to both equations. This value, converted to both °F and K, will yield the value at which both scales are numerically equivalent.

Setup: Conversion from Celsius to Fahrenheit:

$$°F = \left(°C \times \frac{9°F}{5°C}\right) + 32°F.$$

Conversion from Celsius to Kelvin:

$$K = °C + 273.15.$$

Solution: If we set °F = K, then:

$$\left(°C \times \frac{9}{5}\right) + 32 = °C + 273.15$$

Solve the equation for °C:

$$1.8\ °C + 32 = °C + 273.15$$

$$0.8\ °C = 241.15$$

$$°C = 301.44$$

Use this value of Celsius to solve for Kelvin or Fahrenheit:

$$K = 301.44°C + 273.15 = \mathbf{574.59K}$$

$$°F = \left(301.44°C \times \frac{9°F}{5°C}\right) + 32°F = \mathbf{574.59°F}$$

c. **Strategy:** Use the equation which shows the conversion between Celsius and Kelvin.

Setup: Conversion from Celsius to Kelvin:

$$K = °C + 273.15$$

Solution: **No. Since the value of Kelvin is always equal to a value that is 273.15 greater than the Celsius value, there is no temperature at which the values can be numerically equal.**