Answer on Question #50329, Chemistry, Inorganic Chemistry

How many grams of nitrogen and oxygen are dissolved in 125 g of water at 20^{0}C when the water is saturated with air, in which $P_{nitrogen}$ equals 593 torr and P_{oxygen} equals 159 torr? At 1.00 atm pressure, the solubility of pure oxygen in water is 0.00430 g O_2/100.0 g H_2O, and the solubility of pure nitrogen in water is 0.00190 g N_2/100.0 g H_2O?

Solution:

Henry's law can be put into mathematical terms (at constant temperature) as

$$p = k_H \times c$$

where **p** is the partial pressure of the gaseous solute above the solution,
c is the concentration of the dissolved gas and
k_H is a constant with the dimensions of pressure divided by concentration. The constant, known as the Henry's law constant, depends on the solute, the solvent and the temperature.

$$k_H = \frac{p_0}{c_0} = \frac{1\ atm \times 100\ g}{m_{gas0}}$$

$$k_H = \frac{p_{partial}}{c} = \frac{p_{partial}}{\frac{m_{gas}}{m_{water}}} = \frac{p_{partial} \times m_{water}}{m_{gas}}$$

$$\frac{1\ atm \times 100\ g}{m_{gas0}} = \frac{p_{partial} \times m_{water}}{m_{gas}}$$

$$m_{gas} = \frac{p_{partial} \times m_{water} \times m_{gas0}}{1\ atm \times 100\ g}$$

Convertion of torr into atm:

$$p_{atm} = \frac{p_{torr} \times 1\ atm}{760\ torr}$$

So:

$$m_{gas} = \frac{\frac{p_{torr} \times 1\ atm}{760\ torr} \times m_{water} \times m_{gas0}}{1\ atm \times 100\ g} = \frac{p_{torr} \times m_{water} \times m_{gas0}}{760\ torr \times 100\ g}$$

Oxygene:

$$m_{O_2} = \frac{159\ torr \times 125\ g \times 0.00430\ g}{760\ torr \times 100\ g} = \frac{85.4625\ g}{76000} = 0.0011245\ g$$

Nitrogene:

$$m_{O_2} = \frac{593 torr \times 125\ g \times 0.00190\ g}{760\ torr \times 100\ g} = \frac{140.8375\ g}{76000} = 0.001853125\ g$$

Answer:

0.0011245 g of Oxygene (O_2)

0.001853125 g of Nitrogene (N_2)