#63251 Chemistry, General Chemistry

Chapter 13 (13.107)

At ordinary body temperature (37°C), the solubility of N_2 in water at ordinary atmospheric pressure (1.0 atm) is 0.015 g/L. Air is approximately 78 mol % N_2.

1) **Calculate the number of moles of N_2 dissolved per liter of blood, assuming blood is a simple aqueous solution.**
2) **At a depth of 100 ft in water, the external pressure is 4.0 atm. What is the solubility of N_2 from air in blood at this pressure?**
3) **If a scuba diver suddenly surfaces from this depth, how many milliliters of N_2 gas, in the form of tiny bubbles, are released into the bloodstream from each liter of blood?**

Answer:

a. 0.015 g/L · (1 mol N_2/28 g N_2) = about 0.5.36E-4 M

b. k = p/c
XN_2 = 0.78 atm from the problem.
$XN_2 = pN_2/P_{total}$
0.78 · 1 atm = pN_2 = 0.78 atm
k = p/c = 0.78/5.36E-4 = 1455.

At 100 ft the P_{total} = 4.0 atm.
Then pN_2 = XN_2 · Ptotal = 0.78·4.0 atm = 3.12 atm.
c = pN_2/k = 3.12/1455 = 0.002 M.

c. Take the difference in moles in a liter at the two parts of the problem.
2E-3- 5.36E-4 = 0.0015 moles N_2 for each liter.
PV = nRT
1 atm · V = 0.0015·0.082 L atm K^{-1} mol^{-1}·(37 + 273)
V= (0.0015·0.082·(37 + 273))/1 atm = 0.038 L N_2 per liter of blood = 38 ml N_2 per liter of blood.