13-62)

The partial pressure is found by ignoring the other gases in the mixture; that which would be the particular gas's pressure if only its molecules were present. Second, oxygen is 20.95% of the atmosphere not by mass but by the number of moles. Since PV = nRT, and we assume that in a given volume all gases present have the same temperature, the partial pressures are in the same ratio as the number of moles.[2] So,

$$\frac{P_{O_2}}{P_{air}} = \frac{0.2095}{1.0000} = 0.2095 \ .$$

At sea level, where the pressure is 1.01×10^5 Pa, the partial pressure of oxygen is

$$P_{O_2} = 0.2097\ P_{air} = 0.2095 \times 1.01 \times 10^5 = 21{,}160\ \text{Pa.}$$

FYI, atop the mountain,[3] where the atmospheric pressure is only 3.3×10^4 Pa, the partial pressure would be

$$P_{O_2} = 0.2097\ P_{air} = 0.2095 \times 3.3 \times 10^4 = 6914\ \text{Pa.}$$

But what we want is the necessary percentage of oxygen (x) to give the same partial pressure as at the surface:

$$\frac{P_{O_2}}{P_{air}} = \frac{x}{1.0000} \quad \rightarrow \quad x = 1.0000\frac{P_{O_2}}{P_{air}} = 1.0000\frac{21{,}160}{3.3 \times 10^4} = 0.641 = 64.1\%$$

As for drying out, the concentration of water vapor in the air would also be reduced to about a third that of sea level.

[2] $n_{Total} = n_{N_2} + n_{O_2} + n_{Ar} = \frac{P_{N_2}V}{RT} + \frac{P_{O_2}V}{RT} + \frac{P_{Ar}V}{RT} \quad \rightarrow \quad \frac{n_{Total}RT}{V} = P_{N_2} + P_{O_2} + P_{Ar}$.

[3] To get a pressure this low, one would perhaps need to climb K2 or Everest.