Answer on the question #63495, Chemistry / General Chemistry

Question:

16.117

How many milliliters of concentrated hydrochloric acid solution (36.0% HCl by mass, density = 1.18 g/mL) are required to produce 16.0 L of a solution that has a pH of 1.90?

Solution:

pH of the solution is the logarithm of the concentration of hydrogen ion, multiplied by -1:

$$pH = -\log[H^+].$$

Thus, we can calculate the concentration of hydrogen ion:

$$[H^+] = 10^{-1.90} = 0.012589\, mol\, L^{-1}.$$

Hydrochloric acid is a strong acid, so we can assume that the concentration of hydrogen ions is equal to the overall concentration of hydrochloric acid:

$$c = [H^+] = 0.012589\, mol\, L^{-1}$$

In order to produce 16 L of solution, we should take the following number of the moles of hydrochloric acid:

$$n(HCl) = cV = 0.012589(mol\, L^{-1}) \cdot 16(L) = 0.20143\, mol.$$

It is easy to calculate the mass of hydrochloric acid (pure) we should add:

$$m(HCl) = n(HCl) \cdot M(HCl) = 0.20143(mol) \cdot 36.461(g\, mol^{-1}) = 7.344\, g$$

The mass of concentrated acidic solution is:

$$m = \frac{m(HCl)}{0.36} = \frac{7.344(g)}{0.36} = 20.40\, g.$$

As the density of this solution is 1.18 g/mL, the volume of the concentrated solution to be used is:

$$V = \frac{m}{d} = \frac{20.40(g)}{1.18(g\, mL^{-1})} = 17.29\, mL$$

Answer: 17.29 mL