Question #47454, Chemistry, Inorganic Chemistry

The next three (3) problems deal with the titration of 421 mL of 0.501 M carbonic acid (H2CO3) (Ka1 = 4.3 x 10-7, Ka2 = 5.6 x 10-11) with 2.1 M NaOH.

1. What is the pH of the solution at the 2nd equivalence point?

2. What will the pH of the solution be when 0.1316 L of 2.1 M NaOH are added to the 421 mL of 0.501 M carbonic acid?

3. How many mL of the 2.1 M NaOH are needed to raise the pH of the carbonic acid solution to a pH of 6.019?

ANSWER:

1) pH at the 2nd equal. point
   [H$^+$]= $Kw * Ka2 * 2/C$) =sqrt((10^(-14)*5.6*10^(-11)*2)/0.501)=1.495*10^(-12)
   pH=-lg[H+]=-lg(1.495*10^(-12))=**11.82**

2) If we added 0.1316 L 2.1M NaOH to the 421 mL 0.501 M carbonic acid
   [H$^+$]=$\sqrt{}(Kw * Ka1 * \frac{V1+V2}{C*V1})$=
   =sqrt((1*10^(-14)*4.3*10^(-7)*(0.1316+0.421))/(0.501*0.421)=1.06*10^(-10)
   pH= -lg(H+)=-lg(1.06*10^(-10))=**9.97**