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1. Perchloric Acid (HClO₄) comes in a 70% w/w solution that has a density of 1.664 g/mL

A. What is the molarity of this solution?

1 liter would weight 1000 mL x 1.664 g/mL = 1664 g

But of the 1664 g, only 70%, .7 is actually HClO₄ so:

\[ \text{g HClO}_4 = 1664 \times .7 = 1164.8 \text{ g MClO}_4 \]

Mole HClO₄ = 1164.8g / 100.458 = 11.59 mole = 11.59M

B. How many mLs of this concentrated solution to I need to put in a 1000 ml volumetric flask to make 1.0 M solution?

\[ \frac{M_1V_1}{M_2V_2} \]

\[ 11.59 \times V = 1000 \times 1.00M \]

\[ V = \frac{1000 \times 1.00M}{11.59} \]

\[ = 86.24 \text{ mL} \]

C. How many mLs of the concentrated stock solution do I need to make 5000L of a 500 ppb solution?

500 ppb = 500 ng/mL = 500 µg/L

5000L x 500 µg/L = 2500000 µg = 2500 mg = 2.5 g

If we need 2.5 g of HClO₄ and the solution is 70% or .7 HClO₄

\[ \frac{2.5 \text{ g HClO}_4}{X \text{ g total}} = .7 \]

\[ X = \frac{2.5}{.7} = 3.57 \text{ g total solution} \]

\[ 3.57 \text{ g} \times \frac{1 \text{ mL}}{1.664 \text{ g}} = 2.146 \text{ mL} \]

2. Define or give equations for the following terms

- % e relative error = absolute error/actual amount
- median The middle value in an ordered set of results
- coefficient of variation Standard deviation/mean \( X \) 100%
- Confidence interval \[ \text{Confidence interval} = x \pm \frac{ts}{\sqrt{n}} \]

\( K_{sp} \) for the reaction \( XY_2(s) \rightarrow X^{2+} + 2Y^{-} \); \( K_{sp} = [X^{2+}][Y]^{2} \)
3. If I have the following 5 values
   1.59  1.61  1.61  1.58  1.57

Determine the largest possible value for that 6th value that will not be rejected by the Grubbs test using a 95% confidence interval.

To have a 6th value rejected $G >= 1.822$

Since adding a 6th value changes both the s and the mean this problem is harder than I anticipated. The simplest attack is to just start guessing values and calculating G

<table>
<thead>
<tr>
<th>Value</th>
<th>1.63</th>
<th>1.64</th>
<th>1.65</th>
<th>1.66</th>
<th>1.67</th>
</tr>
</thead>
<tbody>
<tr>
<td>G</td>
<td>1.42</td>
<td>1.58</td>
<td>1.69</td>
<td>1.76</td>
<td>1.82</td>
</tr>
</tbody>
</table>

You can see that 1.67 would be rejected but 1.66 would be accepted.

Since this approach takes some time (or excel) I didn’t expect anybody to try in on the test. If you at least calculated the average and the standard deviation for the set of 5, and guessed what the sixth value would be to be rejected, I accepted this for full points, but it is only the first step in a set of successive approximations.

4. A salt with the formula AB₃ has a solubility of .05 g/L What is the $K_{sp}$ of this salt? (Assume the salt has a molar mass of 150 g/mole)

If it has a solubility of .05 g/L this would be .05/150g/mol = 3.33x10⁻⁴ M

Since this would make 1 mole of A and 3 moles of B the concentrations would be:

$[A] = 3.33x10^{-4}M$

$[B] = 3(3.33x10^{-4}M) = 1x10^{-3}M$

$K_{sp} = [A][B]³ = 3.33x10^{-4}M(1x10^{-3})³$

$= 3.33x10^{-13}$

5. I am going to dilute 1.00 ml of 0.5190 M HCl and place it in a 1 liter volumetric flask and fill to the line with water. I am then going to take 1.00 mL of that solution and put in a second 1L flask and fill that to the mark with water. Finally I am then going to take 1.00 mL of that solution and place it in a 500 mL flask and add water until my final volume is 500 mL. What is the pH of the final solution?

$1^{st}$ dilution 1(.5190) = 1000(X) ; $X = 5.19x10^{-4}$

$2^{nd}$ dilution 1(5.19x10⁻⁴) = 1000(X) ; $X = 5.19x10^{-7}$

$3^{rd}$ dilution 1(5.19x10⁻⁷)=500(X) ; $X = 1.038x10^{-9}$

If you simply take $-log(1.038x10^{-9})$ you get 8.98 which is a basic pH so you forgot to include the H⁺ from water.

Doing the quick and dirty correction

$H^+ = 1.038x10^{-9} + 1x10^{-7} = 1.01x10^{-7}$

$pH = 6.99$
6. What is the pH of 0.05 M Acetic acid ($K_a = 1.75 \times 10^{-5}$)

$K_a$ is small enough to do quick and dirty

$[H^+] = \sqrt{0.05 \times 1.75 \times 10^{-5}}$

$= \sqrt{8.75 \times 10^{-7}}$

$= 9.35 \times 10^{-4}$

$pH = -\log(9.35 \times 10^{-4}) = 3.03$

0.05 Sodium Acetate

So this is the conjugate base of acetic acid

$K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{1.75 \times 10^{-5}} = 5.71 \times 10^{-10}$

$K_b$ is small enough to do the Q&D

$[OH^-] = \sqrt{0.05 \times 5.71 \times 10^{-10}}$

$= \sqrt{2.855 \times 10^{-11}}$

$= 5.34 \times 10^{-6}$

$pOH = 5.27$

$pH = 8.73$

7. Describe the two different ways that were discussed in class to find the endpoint of a titration with $Ag^+$ and $Cl^-$.  

I was looking for a verbal description of the Fajans and Volhard methods for doing Ag titrations.
Take home questions - take this page with you. You may use your book, and you may ask me for hints, but you may not ask other students, and you may not use any Web-related resources to help answer the problems. You may use Excel.

8. What is the pH of $\text{Na}_3\text{PO}_4$ While this is a triprotic species, you may ignore two of the Ka's and concentrate on the single Ka that is appropriate for this problem.

$\text{Na}_3\text{PO}_4$ would ionize to $\text{PO}_4^{3-}$
Which is the conjugate base of the last ionization so :
$K_b = K_{w}/K_{a3} = \frac{1\times10^{-14}}{4.22\times10^{-13}} = 0.0237$

This $K$ is $>10^{-4}$ so you can’t do the Q&D
$.0237 = X/(.1-X)$
using the solver function
$X = [\text{OH}^-] = .03825$
$p\text{OH} = 1.42$
$p\text{H} = 12.58$

9. Remember # 5 “I am going to dilute 1.00 ml of 0.5190 M HCl and place it in a 1liter volumetric flask and fill to the line with water. I am then going to take 1.00 mL of that solution and put in a second 1L flask ad fill that to the mark with water. Finally I am then going to take 1.00 mL of that solution and place it in a 500 mL flask and add water until my final volume is 500 mL. What is the pH of the final solution?”

Making the following assumptions: 1.) All volumetric flasks are class A volumetrics. 2.) The 1.00 mL amounts were measured with a 1000 $\mu$L adjustable micropipet. 3.) The original HCl solution had an uncertainty of .0003M in its molarity. What is the concentration of HCl in the final solution, and what is the uncertainty in that molarity?

From my book the error in
the 1000 mL volumetric is .3 ml in 1000 mL = .0003 relative error
the 500 mL volumetric is .2 mL in 500 mL = .0004 relative error
the 1 ml adjustable pipet is .3% = .003 relative error
The error in the original molarity is .0003/.519 = .000578

When you look at the calculation of the concentration you see that it is a series of divisions so you use square root of the squares of the relative errors. Since this could actually be done in a single equation with several divisions, there is no need to do each error analysis separately. You can lump them all into a single large calculation line this:

$$relative\ error = \sqrt{.003^2 + .003^2 + .003^2 + .0003^2 + .0003^2 + .0004^2 + .000578^2}$$
$$= .0052$$

Done like this it is easier to see the individual errors and you make less mistakes in doing one calculation following another. In fact in making up this key I spotted that I forgot one of the errors for the pipet. If you got the above number and I took a point off, let me know and I will give you that point back

$$Molarity = 1.038\times10^{-9} \text{ absolute error} = 1.038\times10^{-9} \times .0052 = 5.4\times10^{-12}$$
$$= 1.038 \pm .005 \times 10^{-9}$$
10. I want to determine the pH of a .1M weak acid solution. I know if $K_a$ is $10^{-5}$ or smaller I can use the approximation that $0.1 - x = 0.1$.
I know that if $K_a$ is $10^{-3}$ or larger, I cannot use this approximation. And must solve exactly.

If my professor says my answer cannot be off by more than 5% or I will get it marked wrong.

For what values of $K$ can I use the Q&D approximation $0.1 - x \approx 1$, and for what values of $K$ must I solve the equation exactly.

I did this one by assuming the $K = 1 \times 10^{-4}$, $2 \times 10^{-4}$ etc, then calculating $X$ both with the Q&D and with the solver.

<table>
<thead>
<tr>
<th>$K_a$</th>
<th>$X$ (Q&amp;D)</th>
<th>$X$ Exact using solver</th>
<th>% error</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.00E-04</td>
<td>0.003162278</td>
<td>0.00311</td>
<td>-0.01681</td>
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<td>2.00E-04</td>
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<tr>
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<tr>
<td>9.00E-04</td>
<td>0.009486833</td>
<td>0.00905</td>
<td>-0.04827</td>
</tr>
</tbody>
</table>

So it looks like any value below $9 \times 10^{-4}$ will give <5% error.