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chapter

Answers to Practice Problems in Chap. 2, Preparation of Reagents and Buffers

Catrin Tyl (✉) • Baraem P. Ismail
Department of Food Science and Nutrition, University of Minnesota,
St. Paul, MN, USA
e-mail: tylxx001@umn.edu; bismailm@umn.edu

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1. (a) This problem can be solved by using Eq. (2.5): The molecular weight of NaH_2PO_4 is 120 g/mol. Make sure to use the same units throughout. Molecular weights are stated in mol/g; the unit of concentration is in mol/L; thus, the 500 mL should also be converted into L:

$$m[\text{g}] = M\left[\frac{\text{mol}}{\text{L}}\right] \times v[\text{L}] \times \text{MW}\left[\frac{\text{g}}{\text{mol}}\right] \quad (2.5)$$

$$m[\text{g}] = 0.1\left[\frac{\cancel{\text{mol}}}{\cancel{\text{L}}}\right] \times 0.5[\cancel{\text{L}}] \times 120\left[\frac{\text{g}}{\cancel{\text{mol}}}\right] = 6\,\text{g}$$

   (b) Just like for Example A2, the only change in the calculation is the use of 156 instead of 120, thus:

$$m[\text{g}] = 0.1\left[\frac{\cancel{\text{mol}}}{\cancel{\text{L}}}\right] \times 0.5[\cancel{\text{L}}] \times 156\left[\frac{\text{g}}{\cancel{\text{mol}}}\right] = 7.8\,\text{g}$$

2. According to the definition of % wt/vol, as found in Table 2.1:

$$\%\frac{w}{v} = \frac{\text{weight solute}\,[\text{g}] \times 100}{\text{total volume}\,[\text{mL}]}$$

$$\text{weight of solute} = \%\frac{\text{wt}}{v} \times \frac{1}{100} \times \text{volume}$$

$$= 10\left[\frac{\text{g}}{\cancel{\text{mL}}}\right] \times \frac{1}{100} \times 150[\cancel{\text{mL}}] = 15\,\text{g}$$

3. Determine NaOH's molarity in a 40% wt/vol solution: Due to NaOH's equivalence of 1, molarity equals normality. Use the definition of wt % to obtain the mass of NaOH in 1 L and Eq. (2.9) to calculate normality:

$$\%\frac{w}{v} = \frac{\text{weight solute}\,[\text{g}] \times 100}{\text{total volume}\,[\text{mL}]}$$

$$\text{weight of solute}\,[\text{g}] = \%\frac{w}{v} \times \frac{1}{100} \times \text{volume}$$

$$= 40\left[\frac{\text{g}}{\text{mL}}\right] \times \frac{1}{100} \times 1000[\text{mL}]$$

$$= 400[\text{g}]$$

$$n[\text{mol}] = \frac{m[\text{g}]}{\text{MW}\left[\frac{\text{g}}{\text{mol}}\right]} = \frac{400}{40} = 10[\text{mol}](\text{in}\,1\,\text{L})$$

   Thus, the molarity and the normality of this solution are 10.

4. The number of equivalents for H_2SO_4 is 2, because it can donate 2 H^+, and so the normality is two times the molarity. The mL of NaOH can be found through inserting into Eq. (2.19):

$$\text{mL of NaOH} = \frac{\text{mL of sulfuric acid} \times N \text{ of sulfuric acid}}{N \text{ of NaOH}}$$

$$\text{mL of NaOH} = \frac{200 \times 4}{10} = 80\,\text{mL}$$

5. For HCl, normality and molarity are equal, because 1 H^+ is released per molecule HCl. The problem can be solved like Example A3 by using Eq. (2.13) to calculate M of concentrated HCl, followed by Eq. (2.15):

$$M\left[\frac{\text{mol}}{\text{L}}\right] = \frac{d \times 1000\left[\frac{\text{g}}{\text{L}}\right]}{\text{MW}\left[\frac{\text{g}}{\text{mol}}\right]} \times \%\frac{\text{wt}}{\text{wt}} \quad (2.13)$$

$$M\left[\frac{\text{mol}}{\text{L}}\right] = \frac{1.2 \times 1000\left[\frac{\text{g}}{\text{L}}\right]}{36.5\left[\frac{\text{g}}{\text{mol}}\right]} \times 0.37 = 12.16\left[\frac{\text{mol}}{\text{L}}\right]$$

$$M_1 \times v_1 = M_2 \times v_2 \quad (2.15)$$

$$v \text{ of concentrated HCl}[\text{L}] = \frac{\text{vol of diluted HCl}\,[\text{L}] \times \text{M of diluted HCl}\left[\frac{\text{mol}}{\text{L}}\right]}{M \text{ of concentrated HCl}\left[\frac{\text{mol}}{\text{L}}\right]}$$

$$v \text{ of concentrated HCl}\,[\text{L}] = \frac{0.25[\text{L}] \times 2\left[\frac{\cancel{\text{mol}}}{\cancel{\text{L}}}\right]}{12.16\left[\frac{\cancel{\text{mol}}}{\cancel{\text{L}}}\right]}$$

$$= 0.041\,\text{L or } 41\,\text{mL}$$

6. Just like for Example A3 and A5, determine the molarity of the concentrated acetic acid with Eq. (2.13) (ignore %wt/wt):

$$M \text{ acetic acid} = \frac{1.05 \times 1000}{60.06}\left[\frac{\text{g} \times \text{mol}}{\text{g} \times \text{L}}\right] = 17.5\left[\frac{\text{mol}}{\text{L}}\right]$$

   The desired amount is 0.04 mol; thus, take the amount calculated below with Eq. (2.16) and dilute to 1 L:

$$v[L] = \frac{n}{M}\left[\frac{\text{mol} \times \text{L}}{\text{mol}}\right] = \frac{0.04}{17.5} = 0.0023\,\text{L}$$

7. The weight of acetic acid in 1 L of solution can be found analogously to Problem 2:

$$\%\frac{\text{w}}{v} = \frac{\text{weight solute}\,[\text{g}] \times 100}{\text{total volume}\,[\text{mL}]}$$

$$\text{weight of acetic acid} = \%\frac{\text{wt}}{v} \times \frac{1}{100} \times \text{volume}$$

$$= 1\left[\frac{\text{g}}{\cancel{\text{mL}}}\right] \times \frac{1}{100} \times 1000[\cancel{\text{mL}}] = 10\,\text{g}$$

The corresponding number of moles of 10 g can be found using Eq. (2.9):

$$n[\text{mol}] = \frac{m[\text{g}]}{\text{MW}\left[\frac{\text{g}}{\text{mol}}\right]} = \frac{10}{60.02} = 0.167\,[\text{mol}](\text{in}\,1\,\text{L})$$

This already answers the question: A 1 % acetic acid solution contains 0.167 mol per L, not 0.1 mol per L.

8. This problem can be solved analogously to Problem 14:

$$\%\frac{w}{v} = \frac{\text{weight solute}\,[\text{g}]\times 100}{\text{total volume}\,[\text{mL}]}$$

$$\begin{aligned}\text{weight of sodium hydroxide} &= \%\frac{\text{wt}}{v}\times\frac{1}{100}\times\text{volume}\\ &= 10\left[\frac{\text{g}}{\cancel{\text{mL}}}\right]\times\frac{1}{100}\times 1000[\cancel{\text{mL}}]\\ &= 100\,\text{g}\end{aligned}$$

The corresponding number of moles of 100 g can be found using Eq. (2.9):

$$n[\text{mol}] = \frac{m[\text{g}]}{\text{MW}\left[\frac{\text{g}}{\text{mol}}\right]} = \frac{10}{40} = 0.25\,[\text{mol}](\text{in}\,1\,\text{L})$$

This already answers the question: A 10% sodium hydroxide solution contains 0.25 mol per L, not 1 mol per L.

9. The normality of $K_2Cr_2O_7$ is 6 times the molarity. Use (Eqs. 2.1 and 2.9) to calculate the molarity and then multiply with the number of equivalents to obtain normality:

$$\text{Molarity}\,(M)\left[\frac{\text{mol}}{\text{L}}\right] = \frac{\text{number of moles}\,(n)[\text{mol}]}{\text{vol}\,(v)[\text{L}]} \tag{2.1}$$

$$n[\text{mol}] = \frac{m}{\text{MW}}\left[\frac{\text{g}\times\text{mol}}{\text{g}}\right] \tag{2.9}$$

$$M\left[\frac{\text{mol}}{\text{L}}\right] = \frac{\frac{0.2\,[\text{g}]}{294.187\left[\frac{\text{g}}{\text{mol}}\right]}}{0.1\,[\text{L}]} = 0.0068\left[\frac{\text{mol}}{\text{L}}\right]$$

$$\begin{aligned}N\left[\frac{\text{equivalents}}{\text{L}}\right] &= M\times\text{number of equivalents}\\ &= 0.0068\times 6 = 0.04\left[\frac{\text{equivalents}}{\text{L}}\right]\end{aligned}$$

10. The molecular weight of KHP is 204.22 g/mol, and as it only contains one unionized carboxyl group, its number of equivalents is 1 and its molarity equals its normality. According to Eq. (2.5), 100 mL would contain:

$$m[\text{g}] = M\left[\frac{\text{mol}}{\text{L}}\right]\times v[\text{L}]\times\text{MW}\left[\frac{\text{g}}{\text{mol}}\right] \tag{2.5}$$

$$m[\text{g}] = 0.1\times 0.1\times 204.22 = 2.0422\,\text{g}$$

11. The first step is to find the desired amount of Ca in the 1000 mL. As listed in Table 2.1, ppm corresponds to:

$$\text{ppm} = \frac{\text{mg solute}}{\text{kg solution}}$$

$$\text{Thus, }1000\text{ ppm} = \frac{1000\text{ mg Ca}}{\text{kg standard solution}}$$

For our example, the solution's density can be assumed to be 1, and thus there would be 1000 mg Ca/L or 1 g Ca/L. If 110.98 g of $CaCl_2$ contain 40.078 g Ca, then 1 g Ca is supplied by:

$$m[\text{g}]\text{ of }CaCl_2 = \frac{1\times 110.98}{40.078} = 2.7691\,\text{g}$$

12. First, the correct ratio of sodium acetate/acetic acid needs to be determined; similar to Example problem C3, $[A^-]$ is expressed through [AH] to have only one unknown quantity in the equation and then inserted into Eq. (2.2):

$$\left[A^-\right] = 0.1 - [\text{AH}]$$

$$5.5 = 4.76 + \log\frac{0.1-[\text{AH}]}{[\text{AH}]}$$

$$10^{(5.5-4.76)} = \frac{0.1-[\text{AH}]}{[\text{AH}]}$$

$$5.5\times[\text{AH}] = 0.1-[\text{AH}]$$

$$[\text{AH}] = \frac{0.1}{(5.5+1)} = 0.0154 \quad \left[A^-\right] = 0.0846$$

Preparing the buffer by approach 1 (Chap 2, Sect, 2.3) (1 L of 0.1 M solutions of sodium acetate and acetic acid):

$$m\text{ of sodium acetate }[\text{g}] = \text{MW}\times M = 82\times 0.1 = 8.2[\text{g}]$$

$$v\text{ of acetic acid }[\text{mL}] = \frac{\text{MW}\times M}{d} = \frac{60.02\times 0.1}{1.05} = 5.7[\text{mL}]$$

Mix them so that the resulting concentrations correspond to 0.0846 M of sodium acetate and 0.0154 M of acetic acid using Eq. (2.15):

$$M_1\times v_1 = M_2\times v_2 \tag{2.15}$$

$$v \text{ of sodium acetate solution } [\mathrm{mL}] = \frac{M \text{ of } \left[\mathrm{A}^{-}\right] \text{ in buffer} \times v \text{ of buffer}}{M \text{ of stock}} = \frac{0.0846 \times 0.25}{0.1} = 211.5[\mathrm{mL}]$$

$$v \text{ of acetic acid solution } [\mathrm{mL}] = \frac{M \text{ of } [\mathrm{AH}] \text{ in buffer} \times v \text{ of buffer}}{M \text{ of stock}} = \frac{0.0154 \times 0.25}{0.1} = 38.5[\mathrm{mL}]$$

Preparing the buffer via approach 2 (Chap 2, Sect, 2.3) (directly dissolve appropriate amounts of sodium acetate and acetic acid in 250 mL):

$$m \text{ of sodium acetate}[\mathrm{g}] = M \text{ of } \left[\mathrm{A}^{-}\right] \text{ in buffer}\left[\frac{\mathrm{mol}}{\mathrm{L}}\right] \times v \text{ of buffer}[\mathrm{L}] \times \mathrm{MW}\left[\frac{\mathrm{g}}{\mathrm{mol}}\right] \tag{2.5}$$

$$m \text{ of sodium acetate}[\mathrm{g}] = 0.25[\mathrm{L}] \times 0.0846\left[\frac{\mathrm{mol}}{\mathrm{L}}\right] \times 82\left[\frac{\mathrm{g}}{\mathrm{mol}}\right] = 1.73[\mathrm{g}]$$

$$\text{vol of acetic acid }[\mathrm{mL}] = \frac{\left\{\text{vol of buffer}[\mathrm{L}] \times M \text{ of } [\mathrm{AH}] \text{ in buffer}\left[\frac{\mathrm{mol}}{\mathrm{L}}\right] \times \mathrm{MW}\left[\frac{\mathrm{g}}{\mathrm{mol}}\right]\right\}}{\left[\frac{\mathrm{g}}{\mathrm{mL}}\right]}$$

$$v \text{ of acetic acid }[\mathrm{mL}] = \frac{0.25[\mathrm{L}] \times 0.0154\left[\frac{\cancel{\mathrm{mol}}}{\mathrm{L}}\right] \times 60.02\left[\frac{\mathrm{g}}{\cancel{\mathrm{mol}}}\right]}{1.05\left[\frac{\mathrm{g}}{\mathrm{mL}}\right]} = 2.2[\mathrm{mL}]$$

These would be dissolved in, e.g., 200 mL, and then the pH adjusted and the volume brought up to 250 mL.

Preparing the buffer via approach 3 (Chap 2, Sect, 2.3) (dissolve the appropriate amount of acetic acid to yield 250 mL of a 0.1 M solution in < 250 mL, and then adjust the pH with NaOH of a high molarity, e.g., 6 or 10 M to pH 5.5):

$$v \text{ of acetic acid }[\mathrm{mL}] = \frac{v \text{ of buffer}[\mathrm{L}] \times M \text{ of buffer}\left[\frac{\mathrm{mol}}{\mathrm{L}}\right] \times \mathrm{MW}\left[\frac{\mathrm{g}}{\mathrm{mol}}\right]}{\left[\frac{\mathrm{g}}{\mathrm{mL}}\right]}$$

$$v \text{ of acetic acid }[\mathrm{mL}] = \frac{0.25[\mathrm{L}] \times 0.1\left[\frac{\mathrm{mol}}{\mathrm{L}}\right] \times 60.02\left[\frac{\mathrm{g}}{\mathrm{mol}}\right]}{1.05\left[\frac{\mathrm{g}}{\mathrm{mL}}\right]} = 1.43[\mathrm{mL}]$$

13. The molarities of Na_2EDTA and $MgSO_4$ are found by rearranging Eq. (2.5):

$$m[\mathrm{g}] = M\left[\frac{\mathrm{mol}}{\mathrm{L}}\right] \times v[\mathrm{L}] \times \mathrm{MW}\left[\frac{\mathrm{g}}{\mathrm{mol}}\right] \tag{2.5}$$

$$M\left[\frac{\mathrm{mol}}{\mathrm{L}}\right] = \frac{m[\mathrm{g}]}{v[\mathrm{L}] \times \mathrm{MW}\left[\frac{\mathrm{g}}{\mathrm{mol}}\right]}$$

$$M\left[\frac{\mathrm{mol}}{\mathrm{L}}\right] \text{ of } \mathrm{Na_2\,EDTA} = \frac{1.179\,[\mathrm{g}]}{0.25\,[\mathrm{L}] \times 372.24\left[\frac{\mathrm{g}}{\mathrm{mol}}\right]} = 0.0127\left[\frac{\mathrm{mol}}{\mathrm{L}}\right]$$

$$M\left[\frac{\mathrm{mol}}{\mathrm{L}}\right] \text{ of } \mathrm{MgSO_4} = \frac{0.78\,[\mathrm{g}]}{0.25\,[\mathrm{L}] \times 246.47\left[\frac{\mathrm{g}}{\mathrm{mol}}\right]} = 0.0127\left[\frac{\mathrm{mol}}{\mathrm{L}}\right]$$

To calculate the buffer pH, determine molarity of NH_4Cl using Eqs. (2.1 and 2.3) and NH_3 with Eq. (2.27):

$$M \text{ of } \mathrm{NH_4Cl} \text{ in buffer}\left[\frac{\mathrm{mol}}{\mathrm{L}}\right] = \frac{m[\mathrm{g}]}{\mathrm{MW}\left[\frac{\mathrm{g}}{\mathrm{mol}}\right] \times v[\mathrm{L}]} = \frac{16.9}{53.49 \times 0.25} = 1.26\left[\frac{\mathrm{mol}}{\mathrm{L}}\right]$$

$$M \text{ of conc.} \mathrm{NH_3}\left[\frac{\mathrm{mol}}{\mathrm{L}}\right] = \frac{\left[\frac{\mathrm{g}}{\mathrm{mL}}\right] \times 1000}{\mathrm{MW}\left[\frac{\mathrm{g}}{\mathrm{mol}}\right]} \times \%\mathrm{wt} = \frac{0.88 \times 1000}{17} \times 0.28 = 14.5\left[\frac{\mathrm{mol}}{L}\right]$$

$$M \text{ in buffer}\left[\frac{\text{mol}}{\text{L}}\right] = \frac{M \text{ of stock solutions}\left[\frac{\text{mol}}{\text{L}}\right]\times v \text{ of stock solutions }[\text{L}]}{v \text{ of buffer }[\text{L}]} \tag{2.27}$$

$$M \text{ of } \mathrm{NH_3} \text{ in buffer}\left[\frac{\text{mol}}{\text{L}}\right] = \frac{M \text{ of conc.}\mathrm{NH_3}\left[\frac{\text{mol}}{\text{L}}\right]\times v \text{ of conc.}\mathrm{NH_3}[\text{L}]}{v \text{ of buffer}[\text{L}]}$$

$$M \text{ of } \mathrm{NH_3} \text{ in buffer}\left[\frac{\text{mol}}{\text{L}}\right] = \frac{14.5\left[\frac{\text{mol}}{\text{L}}\right]\times 0.143[\text{L}]}{0.25[\text{L}]} = 8.29\left[\frac{\text{mol}}{\text{L}}\right]$$

The normal form of the Henderson-Hasselbalch equation may be used after calculating the pK$_a$ of NH$_4^+$. NH$_4$Cl acts as the acid, NH$_4$OH is the base (NH$_4$OH is just another way of writing NH$_3$ in water). Note: The actual pK$_a$ and pK$_b$ may be slightly different because of the added salts affecting the ionic strength.

$$\mathrm{pK_a} \text{ of } \mathrm{NH_4^+} = 14 - \mathrm{pK_b} = 14 - 4.74 = 9.26$$

$$\mathrm{pH} = 9.26 + \log\frac{8.29}{1.26} = 9.26 + 0.82 = 10.08$$

14. (a) Use Eq. (2.5) to obtain the masses:

$$m[\text{g}] = M\left[\frac{\text{mol}}{\text{L}}\right]\times v[\text{L}]\times \text{MW}\left[\frac{\text{g}}{\text{mol}}\right] \tag{2.5}$$

$$m[\text{g}] \text{ of } \mathrm{NaH_2PO_4.H_2O} = 0.2\times 0.5\times 138 = 13.8[\text{g}]$$

$$m[\text{g}] \text{ of } \mathrm{Na_2HPO_4.7H_2O} = 0.2\times 0.5\times 268 = 26.8[\text{g}]$$

(b) Use Eq. (2.38) to express [A$^-$] through [AH] to substitute into Eq. (2.25); find the pK$_a$ in Table 2.2:

$$0.1 = \left[\mathrm{A^-}\right] + [\mathrm{AH}]$$
$$\left[\mathrm{A^-}\right] = 0.1 - [\mathrm{AH}]$$

$$6.2 = 6.71 + \log\frac{0.1 - \mathrm{AH}}{\mathrm{AH}}$$

$$-0.51 = \log\frac{0.1 - [\mathrm{AH}]}{[\mathrm{AH}]}$$

$$0.309 = \frac{0.1 - [\mathrm{AH}]}{[\mathrm{AH}]}$$

$$[\mathrm{AH}]\times(0.309 + 1) = 0.1$$

$$[\mathrm{AH}]\left[\frac{\text{mol}}{\text{L}}\right] = \frac{0.1}{1.219} = 0.0764\left[\frac{\text{mol}}{\text{L}}\right]\left[\mathrm{A^-}\right]\left[\frac{\text{mol}}{\text{L}}\right] = 1 - 0.0764 = 0.0236\left[\frac{\text{mol}}{\text{L}}\right]$$

The molarities of the stock solutions are 0.2 $\left[\frac{\text{mol}}{\text{L}}\right]$. To find the volumes to mix, use Eq. (2.27):

$$M_1 \times v_1 = M_2 \times v_2 \tag{2.27}$$

$$v \text{ of } \mathrm{NaH_2PO_4} \text{ stock solution }[\text{L}] = \frac{M \text{ of } \mathrm{NaH_2PO_4} \text{ in buffer} \times v \text{ of buffer}}{M \text{ of stock solution}}$$

$$v \text{ of } \mathrm{NaH_2PO_4} \text{ stock solution }[\text{L}] = \frac{0.0764\times 0.2}{0.2} = 0.0764\,[\text{L}] \text{ or } 76 \text{ mL}$$

$$v \text{ of } \mathrm{Na_2HPO_4} \text{ stock solution }[\text{L}] = \frac{M \text{ of } \mathrm{Na_2HPO_4} \text{ in buffer} \times v \text{ of buffer}}{M \text{ of stock solution}}$$

$$v \text{ of } \mathrm{NaH_2PO_4} \text{ stock solution}[\text{L}] = \frac{0.0236\times 0.2}{0.2} = 0.024\,[\text{L}] \text{ or } 24\,\text{mL}$$

(c) This problem is similar to Example C3. 1 mL of 6*M* NaOH supplies Eq. (2.2):

$$n \text{ of NaOH }[\text{mol}] = M\times v = 6\times 0.001 = 0.006[\text{mol}]$$

The amounts of NaH$_2$PO$_4$ and Na$_2$HPO$_4$ are found through Eq. (2.2) (use either the buffer molarity or values from the stock solutions calculated for Problem12b):

$$n \text{ of } \mathrm{NaH_2PO_4}[\text{mol}] = M\times v = 0.2\times 0.076 = 0.015[\text{mol}]$$

$$n \text{ of } \mathrm{NaH_2PO_4}\,[\text{mol}] = M\times v = 0.2\times 0.024 = 0.0048\,[\text{mol}]$$

Addition of NaOH changes the ratio by increasing the amount of Na_2HPO_4 and decreasing NaH_2PO_4:

$$n \text{ of } NaH_2PO_4 \text{ after HCl}[\text{mol}] = 0.015 - 0.006 = 0.009[\text{mol}]$$

$$n \text{ of } Na_2HPO_4 \text{ after HCl}[\text{mol}] = 0.0048 + 0.006 = 0.0108\,[\text{mol}]$$

Substitute these values into Eq. (2.25) to find the new pH. (Note: You may insert molar ratios, and you do not need to convert to concentrations, because the ratio would stay the same):

$$\text{pH} = 6.71 + \log\frac{0.0108}{0.009} = 6.77$$

15. To find the pH at 25°C, the acid/base ratio needs to be substituted into Eq. (2.25):

$$\text{pH} = 8.06 + \log\frac{4}{1}$$

$$\text{pH} = 8.06 - 0.6 = 7.46$$

Calculate the pK_a at 60 °C with Eq. (2.61), and then insert into Eq. (2.25):

$$\text{pK}_a = 8.06 - \left(0.023 \times (60 - 25)\right) = 7.26$$

$$\text{pH} = 7.26 - 0.6 = 6.65$$

16. This problem is analogous to Example C3. Find the ratio of acid to base through Eqs. (2.25 and 2.38), and then use Eq. (2.52) to find the volume of formic acid and Eq. (2.5) to find the mass of ammonium formate:

$$0.01 = [A^-] + [AH]$$

$$[A^-] = 0.01 - [AH]$$

$$3.5 = 3.75 + \log\frac{0.01 - [AH]}{[AH]}$$

$$-0.25 = \log\frac{0.01 - [AH]}{[AH]}$$

$$0.562 = \frac{0.01 - [AH]}{[AH]}$$

$$[AH] = \frac{0.01}{1.562} = 0.0064\,[A^-] = 0.0036$$

$$v[\text{mL}] = \frac{M \times v \times \text{MW}}{d} \qquad (2.52)$$

$$v \text{ of formic acid }[\text{mL}] = \frac{M \times v \text{ of buffer} \times \text{MW}}{d} = \frac{0.0064\left[\frac{\cancel{\text{mol}}}{\cancel{\text{L}}}\right] \times 46\left[\frac{\cancel{\text{g}}}{\cancel{\text{mol}}}\right] \times 1[\cancel{\text{L}}]}{1.22\left[\frac{\cancel{\text{g}}}{\text{mL}}\right]} = 0.24[\text{mL}]$$

$$m \text{ of ammonium formate}[\text{g}] = M \times v \times \text{MW of buffer} = 0.0036\left[\frac{\cancel{\text{mol}}}{\cancel{\text{L}}}\right] \times 63.06\left[\frac{\text{g}}{\cancel{\text{mol}}}\right] \times 1[\cancel{\text{L}}] = 0.227[\text{g}]$$