Section 10.3: Acid–Base Stoichiometry

Tutorial 1 Questions, page 481
1. It is necessary to keep the volume of indicator used to a minimum because many acid–base indicators are weak acids. Some of the base used in a titration reacts with the indicator.
2. It is necessary to rinse with titrant after rinsing with water so that the inner walls of the burette are lined with titrant rather than with water. If water were present, it could dilute the titrant.
3. The air bubble occupies a certain volume in the burette. The presence of an air bubble would make the volume of titrant used seem larger than it actually is.
4. Rinsing the sides of a titration flask near the end of the titration washes any unreacted titrant into the bottom of the flask.
5. The solution turned pink momentarily because the sodium hydroxide solution was briefly concentrated in one spot. Then, as the flask is swirled, the sodium hydroxide spreads throughout the flask where it reacts with excess acid, changing the indicator to colourless.

Tutorial 2 Practice, page 484
1. (a) Given: three volumes of acid
   Required: average volume of acid used, $V_{\text{HNO}_3}$
   Solution: Calculate the average of titrant (acid) used by subtracting the final from the initial burette volume for each trial.
   
   \[
   V_{\text{HNO}_3, \text{average}} = \frac{8.00 \text{ mL} + 8.06 \text{ mL} + 8.12 \text{ mL}}{3}
   \]
   \[
   V_{\text{HNO}_3, \text{average}} = 8.06 \text{ mL}
   \]
   Statement: The average volume of nitric acid used is 8.06 mL.
   (b) Given: concentration of acid, $c_{\text{HNO}_3} = 0.500 \text{ mol/L}$
   volume of base, $V_{\text{NaOH}} = 25.00 \text{ mL}$
   average volume of acid, $V_{\text{HNO}_3, \text{average}} = 8.06 \text{ mL}$
   Required: amount concentration of base, $c_{\text{NaOH}}$
   Analysis: $c = \frac{n}{V}$
   Solution:
   Step 1. Convert the volume of solution to litres.
   \[
   V_{\text{HNO}_3, \text{average}} = \frac{8.06 \text{ mL} \times 1 \text{ L}}{1000 \text{ mL}}
   \]
   \[
   V_{\text{HNO}_3, \text{average}} = 8.06 \times 10^{-3} \text{ L}
   \]
   \[
   V_{\text{NaOH}} = 25.00 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}}
   \]
   \[
   V_{\text{NaOH}} = 2.500 \times 10^{-2} \text{ L}
   \]
**Step 2.** Write a balanced equation for the reaction listing the given values.

\[ \text{NaOH(aq)} + \text{HNO}_3(\text{aq}) \rightarrow \text{NaNO}_3 + \text{H}_2\text{O} \]

\[ 2.5 \times 10^{-2} \text{ L} \quad 8.06 \times 10^{-3} \text{ L} \]

\[ c_{\text{NaOH}} = 0.500 \text{ mol/L} \]

**Step 3.** Use the concentration equation to determine the amount of acid.

\[ n_{\text{HNO}_3} = c_{\text{HNO}_3} V_{\text{HNO}_3} \text{(average)} \]

\[ = \frac{0.500 \text{ mol}}{1 \text{ L}} \times 8.06 \times 10^{-3} \text{ L} \]

\[ n_{\text{HNO}_3} = 4.03 \times 10^{-3} \text{ mol} \]

**Step 4.** Use the amount of acid and the mole ratio in the balanced equation to determine the amount of base.

\[ n_{\text{NaOH}} = 4.03 \times 10^{-3} \text{ mol}_{\text{HNO}_3} \times \frac{1 \text{ mol}_{\text{NaOH}}}{1 \text{ mol}_{\text{HNO}_3}} \]

\[ n_{\text{NaOH}} = 4.03 \times 10^{-3} \text{ mol}_{\text{NaOH}} \]

**Step 5.** Use the concentration equation to determine the concentration of base.

\[ c_{\text{NaOH}} = \frac{n_{\text{NaOH}}}{V_{\text{NaOH}}} \]

\[ = \frac{4.03 \times 10^{-3} \text{ mol}}{2.500 \times 10^{-2} \text{ L}} \]

\[ c_{\text{NaOH}} = 0.161 \text{ mol/L} \]

**Statement:** The amount concentration of the sodium hydroxide solution is 0.161 mol/L.

2. **Given:** mass of potassium carbonate, \( m_{\text{K}_2\text{CO}_3} = 1.00 \text{ g} \)

volume of acid, \( V_{\text{HNO}_3} = 24.20 \text{ mL} \)

**Required:** amount concentration of acid, \( c_{\text{HNO}_3} \)

**Analysis:** \( c = \frac{n}{V} \)

**Solution:**

**Step 1.** Convert the volume of titrant used to litres.

\[ V_{\text{HNO}_3} = 24.20 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \]

\[ V_{\text{HNO}_3} = 2.420 \times 10^{-2} \text{ L} \]

**Step 2.** Write the balanced equation listing the given values.

\[ \text{K}_2\text{CO}_3(\text{s}) + 2 \text{ HNO}_3(\text{aq}) \rightarrow 2 \text{ KNO}_3 + \text{H}_2\text{CO}_3 \]

\[ 1.00 \text{ g} \quad 2.420 \times 10^{-2} \text{ L} \]

\[ 138.21 \text{ g/mol} \quad c_{\text{HNO}_3} \]
Step 3. Convert the mass of potassium carbonate to the amount.

\[ n_{K_2CO_3} = 1.00 \text{ g} \times \frac{1 \text{ mol}}{138.21 \text{ g}} \]

\[ n_{K_2CO_3} = 7.2354 \times 10^{-3} \text{ mol} \] [2 extra digits carried]

Step 4. Use the amount and the mole ratio in the balanced equation to determine the amount of the acid.

\[ n_{HNO_3} = 7.2354 \times 10^{-3} \text{ mol}_{K_2CO_3} \times \frac{2 \text{ mol}_{HNO_3}}{1 \text{ mol}_{K_2CO_3}} \]

\[ n_{HNO_3} = 1.4471 \times 10^{-2} \text{ mol}_{HNO_3} \] [2 extra digits carried]

Step 5. Use the concentration equation to determine the concentration of acid.

\[ c_{HNO_3} = \frac{n_{HNO_3}}{V_{HNO_3}} \]

\[ c_{HNO_3} = \frac{1.4471 \times 10^{-2} \text{ mol}}{2.420 \times 10^{-2} \text{ L}} \]

\[ c_{HNO_3} = 0.598 \text{ mol/L} \]

Statement: The amount concentration of nitric acid is 0.598 mol/L.

3. Given: concentration of acid, \( c_{HCl} = 0.100 \text{ mol/L} \)

volume of acid, \( V_{HCl} = 15.52 \text{ mL} \)

volume of base, \( V_{Ba(OH)_2} = 25.00 \text{ mL} \)

Required: amount concentration of base, \( c_{Ba(OH)_2} \)

Analysis: \( c = \frac{n}{V} \)

Solution:

Step 1. Convert the volume of the solutions to litres.

\[ V_{HCl} = 15.52 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \]

\[ V_{HCl} = 1.552 \times 10^{-2} \text{ L} \]

\[ V_{Ba(OH)_2} = 25.00 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \]

\[ V_{Ba(OH)_2} = 2.500 \times 10^{-2} \text{ L} \]
Step 2. Write a balanced equation for the reaction listing the given values.

\[ 2 \text{HCl(aq)} + \text{Ba(OH)}_2(\text{aq}) \rightarrow \text{BaCl}_2 + 2 \text{H}_2\text{O} \]

\[ 1.552 \times 10^{-2} \text{ L} \quad 2.500 \times 10^{-2} \text{ L} \]

\[ 0.100 \text{ mol/L} \quad c_{\text{Ba(OH)}_2} \]

Step 3. Use the concentration equation to determine the amount of acid.

\[ n_{\text{HCl}} = c_{\text{HCl}} V_{\text{HCl}} \]

\[ = 0.100 \text{ mol} \times \frac{1}{1} \times 1.552 \times 10^{-2} \text{ L} \]

\[ n_{\text{HCl}} = 1.552 \times 10^{-3} \text{ mol} \]

Step 4. Use the amount of acid and the mole ratio in the balanced equation to determine the amount of base.

\[ n_{\text{Ba(OH)}_2} = 1.552 \times 10^{-3} \text{ mol} \times \frac{1 \text{ mol} \text{Ba(OH)}_2}{2 \text{ mol} \text{HCl}} \]

\[ n_{\text{Ba(OH)}_2} = 7.760 \times 10^{-4} \text{ mol} \text{Ba(OH)}_2 \]

Step 5. Use the concentration equation to determine the concentration of base.

\[ c_{\text{Ba(OH)}_2} = \frac{n_{\text{Ba(OH)}_2}}{V_{\text{Ba(OH)}_2}} \]

\[ = \frac{7.760 \times 10^{-4} \text{ mol}}{2.500 \times 10^{-2} \text{ L}} \]

\[ c_{\text{Ba(OH)}_2} = 3.10 \times 10^{-2} \text{ mol/L} \]

Statement: The amount concentration of the barium hydroxide solution is \( 3.10 \times 10^{-2} \text{ mol/L} \).

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1. The purpose of an acid–base indicator in a titration is to signify when the endpoint of the titration is reached.

2. The equivalence point is the point during a titration when neutralization is complete. The endpoint of a titration is a sudden change in a property of the solution such as a change in pH or conductivity. Ideally, the endpoint occurs at a pH very close to that of the equivalence point.

3. (a) Given: mass of sodium hydroxide, \( m_{\text{NaOH}} = 4.00 \text{ g} \)

volume of solution, \( V_{\text{NaOH}} = 100.0 \text{ mL} \)

Required: amount concentration of base, \( c_{\text{NaOH}} \)

Analysis: \( c_{\text{NaOH}} = \frac{n_{\text{NaOH}}}{V_{\text{NaOH}}} \)
Solution:

**Step 1.** Convert the volume of solution to litres.

\[ V_{\text{NaOH}} = \frac{100.0 \text{ mL}}{1000 \text{ mL}} \times \frac{1 \text{ L}}{1000 \text{ mL}} \]

\[ V_{\text{NaOH}} = 0.1000 \text{ L} \]

**Step 2.** Convert mass of the sodium hydroxide to the amount.

\[ n_{\text{NaOH}} = \frac{4.00 \text{ g}}{40.00 \text{ g/mol}} \times \frac{1 \text{ mol}}{40.00 \text{ g}} \]

\[ n_{\text{NaOH}} = 0.100 \text{ mol} \]

**Step 3.** Use the concentration equation to determine the concentration of the solution.

\[ c_{\text{NaOH}} = \frac{n_{\text{NaOH}}}{V_{\text{NaOH}}} \]

\[ c_{\text{NaOH}} = \frac{0.100 \text{ mol}}{0.100 \text{ L}} \]

\[ c_{\text{NaOH}} = 1.00 \text{ mol/L} \]

**Statement:** The amount concentration of sodium hydroxide solution is 1.00 mol/L.

(b) The actual concentration differs from the predicted concentration because, over time, contaminants may enter the solution and react with sodium hydroxide. Carbon dioxide, for example, is known to react with sodium hydroxide to form sodium carbonate. This decreases the amount of sodium hydroxide present. Assuming the volume of solution remains constant, carbon dioxide contamination decreases the concentration of the solution.

4. It is necessary to re-standardize a solution after it has been stored for some time because the concentration of the solution may vary over time. This may be caused by several factors including reactions with contaminants and changes in the volume of the solution.

5. Storage bottles of solutions are often filled right to the top to eliminate the amount of air that would otherwise occupy the space above the solution. Components in air, such as oxygen and carbon dioxide, may react with the solution.

6. Drying the samples in an oven before using them to standardize a base is necessary to remove any water that may have been absorbed by the solid. If the mass were not dried, the mass of a sample contaminated with water would be larger than a dried sample. As a result, the mass of potassium hydrogen phthalate recorded would be larger than it actually is, and the calculated amount and concentration of base would be incorrect.

7. A pH meter can be used to monitor the pH changes during a titration. Near the equivalence point, a sharp rise or drop should be observed. For an acid–base indicator to be suitable for a titration, the colour change should occur when the equivalence point is reached.

8. (a) The two acids have different initial pH because Acid 1 is stronger than Acid 2 and, thus, it has a lower pH.

(b) The volume that is required to neutralize these acids is 50 mL. The same volume is required for both acids because the acids are of the same concentration. This also assumes that both acids contain the same number of hydrogen atoms that ionize.

(c) An acid–base indicator that could be used for one of the acids, but not the other is methyl red.
9. (a) CaCO$_3$(s) + 2 HCl(aq) → CO$_2$(g) + 2 H$_2$O(l) + CaCl$_2$(aq)

(b) Given: volume of hydrochloric acid, $V_{\text{HCl}} = 25.20$ mL

amount concentration of hydrochloric acid, $c_{\text{HCl}} = 0.50$ mol/L

Required: mass of calcium carbonate, $m_{\text{CaCO}_3}$

Analysis: $c = \frac{n}{V}$

Solution:

Step 1. Convert the volume of solution to litres.

$$V_{\text{HCl}} = 25.20 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}}$$

$$V_{\text{HCl}} = 0.0252 \text{ L}$$

Step 2. Rearrange the concentration equation to determine the amount of hydrochloric acid.

$$n_{\text{HCl}} = c_{\text{HCl}} \times V_{\text{HCl}}$$

$$n_{\text{HCl}} = 0.50 \text{ mol/L} \times 0.0252 \text{ L}$$

$$n_{\text{HCl}} = 0.0126 \text{ mol}$$

Step 3. Use the amount of acid and the mole ratio in the balanced equation to determine the amount of calcium carbonate.

$$\text{CaCO}_3(s) + 2 \text{HCl}(aq) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(l) + \text{CaCl}_2(aq)$$

$$n_{\text{CaCO}_3} = 0.0126 \text{ mol}_{\text{HCl}} \times \frac{1 \text{ mol}_{\text{CaCO}_3}}{2 \text{ mol}_{\text{HCl}}}$$

$$n_{\text{CaCO}_3} = 0.0063 \text{ mol}_{\text{CaCO}_3}$$

Step 4. Calculate the mass of calcium carbonate.

$$m_{\text{CaCO}_3} = 0.0063 \text{ mol}_{\text{CaCO}_3} \times \frac{100.09 \text{ g}}{1 \text{ mol}}$$

$$m_{\text{CaCO}_3} = 0.63 \text{ g}$$

Statement: The mass of calcium carbonate required is 0.63 g.

10. (a) 2 NH$_4$OH(aq) + H$_2$SO$_4$(aq) → 2 H$_2$O(l) + (NH$_4$)$_2$SO$_4$(aq)

(b) Given: three volumes of acid

Required: average volume of acid used, $V_{\text{H}_2\text{SO}_4}$

Solution: Calculate the average of titrant (acid) used by subtracting the final from the initial burette volume for each trial.

$$V_{\text{H}_2\text{SO}_4(average)} = \frac{8.46 \text{ mL} + 8.50 \text{ mL} + 8.66 \text{ mL}}{3}$$

$$V_{\text{H}_2\text{SO}_4(average)} = 8.54 \text{ mL}$$

Statement: The average volume of sulfuric acid used is 8.54 mL.
(c) **Given:** concentration of acid, \( c_{H_2SO_4} = 0.25 \) mol/L

volume of base, \( V_{NH_4OH} = 10.00 \) mL

average volume of acid, \( V_{H_2SO_4 (average)} = 8.54 \) mL

**Required:** amount concentration of base, \( c_{NH_4OH} \)

**Analysis:** \( c = \frac{n}{V} \)

**Solution:**

**Step 1.** Convert the volume of solution to litres.

\[
V_{NH_4OH} = 10.00 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.010 \text{ 00 L}
\]

\[
V_{H_2SO_4 (average)} = 8.54 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.00854 \text{ L}
\]

**Step 2.** Write a balanced equation for the reaction listing the given values.

\[
2 \text{ NH}_4\text{OH(aq) } + \text{ H}_2\text{SO}_4(aq) \rightarrow 2 \text{ H}_2\text{O(l) } + \text{(NH}_4\text{)}_2\text{SO}_4(aq)
\]

\[
0.010 \text{ 00 L } \quad 0.00854 \text{ L}
\]

\[
c_{NH_4OH} = 0.25 \text{ mol/L}
\]

**Step 3.** Use the concentration equation to determine the amount of acid.

\[
n_{H_2SO_4} = c_{H_2SO_4} V_{H_2SO_4 (average)}
\]

\[
= \frac{0.25 \text{ mol}}{1 \text{ L}} \times 0.00854 \text{ L}
\]

\[
n_{H_2SO_4} = 0.002135 \text{ mol}
\]

**Step 4.** Use the amount of acid and the mole ratio in the balanced equation to determine the amount of base.

\[
n_{NH_4OH} = \frac{2 \text{ mol}_{NH_4OH}}{1 \text{ mol}_{H_2SO_4}} \times 0.002135 \text{ mol}_{H_2SO_4}
\]

\[
n_{NH_4OH} = 0.00427 \text{ mol}_{NH_4OH}
\]
**Step 5.** Use the concentration equation to determine the concentration of base.

\[
c_{\text{NH}_4\text{OH}_2} = \frac{n_{\text{NH}_4\text{OH}_2}}{V_{\text{NH}_4\text{OH}_2}}
\]

\[
= \frac{0.00427 \text{ mol}}{0.0100 \text{ L}}
\]

\[
c_{\text{NH}_4\text{OH}_2} = 0.43 \text{ mol/L}
\]

**Statement:** The amount concentration of ammonium hydroxide solution is 0.43 mol/L.

(d) It is necessary to conduct more than one trial when standardizing a solution to show that the data is reproducible and reliable.

11. **Given:** concentration of base, \(c_{\text{NaOH}} = 0.100 \text{ mol/L}\)

average volume of base, \(V_{\text{NaOH (average)}} = 1.60 \text{ mL}\)

volume of acid, \(V_{\text{milk}} = 10.00 \text{ mL}\)

**Required:** amount concentration of lactic acid in milk, \(c_{\text{HC}_3\text{H}_5\text{O}_2}\)

**Analysis:** \(c = \frac{n}{V}\)

**Solution:**

**Step 1.** Convert the volume of solutions to litres.

\[
V_{\text{NaOH (average)}} = 1.60 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}}
\]

\[
V_{\text{NaOH (average)}} = 0.00160 \text{ L}
\]

\[
V_{\text{milk}} = 10.00 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}}
\]

\[
V_{\text{milk}} = 0.0100 \text{ L}
\]

**Step 2.** Write the balanced equation for the reaction listing the given values.

\[
\text{HC}_3\text{H}_5\text{O}_2(\text{aq}) + \text{NaOH(}\text{aq}) \rightarrow \text{H}_2\text{O(l)} + \text{NaC}_3\text{H}_5\text{O}_2(\text{aq})
\]

\[
c_{\text{HC}_3\text{H}_5\text{O}_2} = 0.00160 \text{ L}
\]

\[
c_{\text{HC}_3\text{H}_5\text{O}_2} = 0.100 \text{ mol/L}
\]

**Step 3.** Use the concentration equation to determine the amount of base.

\[
n_{\text{NaOH}} = c_{\text{NaOH}}V_{\text{NaOH (average)}}
\]

\[
= \frac{0.100 \text{ mol}}{1 \text{ L}} \times 0.00160 \text{ L}
\]

\[
n_{\text{NaOH}} = 1.60 \times 10^{-4} \text{ mol}
\]
Step 4. Use the amount of acid and the mole ratio in the balanced equation to determine the amount of base.

\[ n_{\text{HC}_3\text{H}_5\text{O}_2} = 1.60 \times 10^{-4} \text{ mol NaOH} \times \frac{1 \text{ mol } \text{NaOH}}{1 \text{ mol } \text{HC}_3\text{H}_5\text{O}_2} \]

\[ n_{\text{HC}_3\text{H}_5\text{O}_2} = 1.60 \times 10^{-4} \text{ mol } \text{HC}_3\text{H}_5\text{O}_2 \]

Step 5. Use the concentration equation to determine the concentration of base.

\[ c_{\text{HC}_3\text{H}_5\text{O}_2} = \frac{n_{\text{HC}_3\text{H}_5\text{O}_2}}{V_{\text{HC}_3\text{H}_5\text{O}_2}} \]

\[ c_{\text{HC}_3\text{H}_5\text{O}_2} = \frac{1.60 \times 10^{-4}}{0.0100 \text{ L}} = 0.0160 \text{ mol/L} \]

Statement: The amount concentration of lactic acid in milk is 0.0160 mol/L.

12. (a) \[ \text{KH(}\text{IO}_3\text{)}_2(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + 2 \text{ KIO}_3(\text{aq}) \]

(b) Given: mass of potassium hydrogen iodate, \( m_{\text{KH(}\text{IO}_3\text{)}_2} = 1.50 \text{ g} \)

average volume of base, \( V_{\text{KOH}} = 25.42 \text{ mL} \)

Required: amount concentration of potassium hydroxide, \( c_{\text{KOH}} \)

Solution:

Step 1. Convert the volume of titrant used to litres.

\[ V_{\text{KOH}} = 25.42 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \]

\[ V_{\text{KOH}} = 2.542 \times 10^{-2} \text{ L} \]

Step 2. Write the balanced equation for the reaction listing the given values.

\[ \text{KH(}\text{IO}_3\text{)}_2(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + 2 \text{ KIO}_3(\text{aq}) \]

1.50 g \( \quad \) 2.420 \times 10^{-2} L

389.91 g/mol \( \quad \) \( c_{\text{KOH}} \)

Step 3. Convert the mass of potassium hydrogen iodate to the amount.

\[ n_{\text{KH(}\text{IO}_3\text{)}_2} = 1.50 \text{ g} \times \frac{1 \text{ mol}}{389.91 \text{ g}} \]

\[ n_{\text{KH(}\text{IO}_3\text{)}_2} = 3.847 \times 10^{-3} \text{ mol} \]

Step 4. Use the amount and the mole ratio in the balanced equation to determine the amount of the base.

\[ n_{\text{KOH}} = 3.847 \times 10^{-3} \text{ mol } \text{KH(}\text{IO}_3\text{)}_2 \times \frac{1 \text{ mol } \text{KOH}}{1 \text{ mol } \text{KH(}\text{IO}_3\text{)}_2} \]

\[ n_{\text{KOH}} = 3.847 \times 10^{-3} \text{ mol } \text{KOH} \]
Step 5. Use the concentration equation to determine the concentration of base.

\[ c_{\text{KOH}} = \frac{n_{\text{KOH}}}{V_{\text{KOH}}} \]

\[ = \frac{3.847 \times 10^{-3} \text{ mol}}{2.542 \times 10^{-2} \text{ L}} \]

\[ c_{\text{KOH}} = 0.151 \frac{\text{mol}}{\text{L}} \]

Statement: The amount concentration of potassium hydroxide is 0.151 mol/L.

13. Given: mass of potassium carbonate, \( m_{\text{K}_2\text{CO}_3} = 1.00 \text{ g} \)

average volume of acid, \( V_{\text{HCl(average)}} = 24.20 \text{ mL} \)

volume of base, \( V_{\text{K}_2\text{CO}_3} = 100 \text{ mL} \)

Required: amount concentration of acid, \( c_{\text{HCl}} \)

Analysis: \[ c = \frac{n}{V} \]

Solution:
Step 1. Convert the volume of solutions to litres.

\[ V_{\text{HCl(average)}} = 24.20 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \]

\[ V_{\text{HCl(average)}} = 2.420 \times 10^{-2} \text{ L} \]

\[ V_{\text{K}_2\text{CO}_3} = 100 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \]

\[ V_{\text{K}_2\text{CO}_3} = 0.100 \text{ L} \]

Step 2. Write a balanced equation for the reaction listing the values.
\( \text{K}_2\text{CO}_3(aq) + 2 \text{ HCl}(aq) \rightarrow 2 \text{ H}_2\text{O(l)} + \text{CO}_2(g) + 2 \text{ KCl(aq)} \)

1.00 g \( \quad 2.420 \times 10^{-2} \text{ L} \)

138.21 g/mol \( c_{\text{HCl}} \)

Step 3. Convert the mass of potassium carbonate to the amount.

\[ n_{\text{K}_2\text{CO}_3} = 1.00 \text{ g} \times \frac{1 \text{ mol}}{138.21 \text{ g}} \]

\[ n_{\text{K}_2\text{CO}_3} = 7.235 \times 10^{-3} \text{ mol} \]

Step 4. Use the amount of potassium carbonate and the mole ratio in the balanced equation to determine the amount of acid.

\[ n_{\text{HCl}} = 7.235 \times 10^{-3} \text{ mol}_{\text{K}_2\text{CO}_3} \times \frac{2 \text{ mol}_{\text{HCl}}}{1 \text{ mol}_{\text{K}_2\text{CO}_3}} \]

\[ n_{\text{HCl}} = 1.447 \times 10^{-2} \text{ mol}_{\text{HCl}} \]
**Step 5.** Use the concentration equation to determine the concentration of acid.

\[
c_{\text{HCl}} = \frac{n_{\text{HCl}}}{V_{\text{HCl}}}
\]

\[
= \frac{1.447 \times 10^{-2} \text{ mol}}{2.420 \times 10^{-2} \text{ L}}
\]

\[
c_{\text{HCl}} = 0.598 \ \text{mol/L}
\]

**Statement:** The amount concentration of the hydrochloric acid solution is 0.598 mol/L.