Calculate The PH Change That Results When 13 ML Of 2.4 M HCl Is Added To 600. ML Of Each Of The Following

Calculate The pH Change That Results When 13 ML Of 2.4 M HCl Is Added To 600. ML Of Each Of The Following is a common problem in acid-base chemistry, often encountered in laboratory settings and academic exercises. Understanding how to determine the change in pH when a strong acid like hydrochloric acid (HCl) is added to a solution is fundamental in various chemical applications, including titrations, buffer preparations, and environmental chemistry. This article will guide you through the process of calculating the pH change step-by-step, covering the essential concepts, formulas, and practical calculations involved.

Understanding the Basics of pH and Acid-Base Reactions

Before delving into calculations, it is crucial to review some foundational concepts.

What Is pH?

pH is a measure of the hydrogen ion concentration $[H^+]$ in a solution, expressed as:

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

A lower pH indicates a more acidic solution, while a higher pH indicates a more basic or alkaline solution.

Strong Acids and Bases

- Strong acids like HCl dissociate completely in water, releasing all their hydrogen ions.
- Strong bases dissociate completely to release hydroxide ions $[OH^-]$.

Since HCl is a strong acid, when it dissolves, it fully dissociates:

$$
\text{HCl} \rightarrow H^+ + Cl^-
$$

Step-by-Step Calculation Approach

To determine the pH change resulting from adding 13 mL of 2.4 M HCl to 600 mL of another solution, follow these steps:

1. Calculate the moles of HCl added.
2. Determine the initial moles of $[H^+]$ in the original solution, if any.
3. Calculate the total moles of $[H^+]$ after addition.
4. Determine the total volume of the solution after mixing.
5. Calculate the new $[H^+]$ concentration.
6. Find the new pH and compare it with the initial pH to determine the change.

Let's explore each step in detail.

Calculating the Moles of HCl Added

Given:
- Volume of HCl solution, $V_{HCl} = 13\, \text{mL} = 0.013\, \text{L}$
- Concentration of HCl, $C_{HCl} = 2.4\, \text{M}$

The number of moles of HCl added:

$$
n_{HCl} = C_{HCl} \times V_{HCl} = 2.4\, \text{mol/L} \times 0.013\, \text{L} = 0.0312\, \text{mol}
$$

This is the amount of $[H^+]$ introduced into the solution due to the added acid.

Assessing the Initial Conditions of the Solution

The problem states adding HCl to "600 mL of each of the following," but without specific initial solutions provided, we can assume two common scenarios:

Scenario A: The initial solution is pure water (neutral)
Scenario B: The initial solution has a known pH or $[H^+]$ concentration

For clarity, we will focus on Scenario A—adding acid to pure water—since no initial pH is specified. The process would be similar for other initial conditions, with adjustments based on initial $[H^+]$.

Initial moles of $[H^+]$ in the original solution:

$$
n_{initial} = 10^{-7}\, \text{mol/L} \times 0.6\, \text{L} \approx 3.0 \times 10^{-8}\, \text{mol}
$$

which is negligible compared to the moles of acid added.

Therefore, initial pH is approximately 7, and the solution is neutral.

Calculating the Total Moles of $[H^+]$ After Addition

Since HCl dissociates completely, the added moles of HCl contribute directly to the $[H^+]$ concentration.

Total moles of $[H^+]$ after addition:
$$
n_{total} = n_{initial} + n_{HCl} \approx 0 + 0.0312\, \text{mol} = 0.0312\, \text{mol}
$$

The total volume after mixing:
$$
V_{total} = 600\, \text{mL} + 13\, \text{mL} = 613\, \text{mL} = 0.613\, \text{L}
$$

The new $[H^+]$ concentration:
$$
[H^+]_{new} = \frac{n_{total}}{V_{total}} = \frac{0.0312\, \text{mol}}{0.613\, \text{L}} \approx 0.0509\, \text{M}
$$

Calculating the new pH:
$$
pH_{new} = -\log(0.0509) \approx 1.29
$$

Since the initial pH (assuming pure water) was about 7, the change in pH:
$$
\Delta pH = pH_{initial} - pH_{new} \approx 7 - 1.29 = 5.71
$$

This indicates a significant decrease in pH, making the solution highly acidic after the addition.

Considering Different Initial Conditions

If the initial solution is not pure water but an existing buffer or solution with known pH, the calculation becomes more complex. Here are some common scenarios:

1. Initial Solution with Known pH

Suppose the initial solution has a pH of 4:
- $[H^+]$ initially:

$$
[H^+]_{initial} = 10^{-4} = 0.0001\, \text{M}
$$
- Moles of initial $[H^+]$:
$$
n_{initial} = 0.0001\, \text{mol/L} \times 0.6\, \text{L} = 6 \times 10^{-5}\, \text{mol}
$$
- Total moles after addition:
$$
n_{total} = 6 \times 10^{-5} + 0.0312 \approx 0.0312\, \text{mol}
$$
- Total volume:
$$
V_{total} = 0.613\, \text{L}
$$
- New $[H^+]$:
$$
[H^+]_{new} = \frac{0.0312}{0.613} \approx 0.0509\, \text{M}
$$
- New pH:
$$
pH_{new} = -\log(0.0509) \approx 1.29
$$
- Initial pH:
$$
pH_{initial} = 4
$$
- pH change:
$$
\Delta pH = 4 - 1.29 = 2.71
$$

The pH drops significantly, although not as drastically as in the pure water case because of the initial buffering capacity.

Impact of Volume and Concentration on pH Change

The magnitude of pH change depends heavily on the volume and concentration of the added acid relative to the initial solution. Here are key points:

- Higher concentration or volume of acid results in a larger pH decrease.
- Larger initial volume of solution buffers the change, leading to a smaller pH shift.
- Buffer capacity of the solution influences how much the pH changes upon addition of acid.

Practical Applications and Considerations

Understanding and calculating pH changes are vital in various real-world applications:

- Titrations: Precise calculations help determine endpoint pH.
- Environmental Chemistry: Acid rain impacts pH of water bodies; calculations can predict effects.
- Pharmaceuticals and Food Industry: pH adjustments are crucial in formulation.

Important tips:
- Always convert volumes to liters for molarity calculations.
- Remember that strong acids dissociate completely; thus, moles of acid equate directly to moles of $[H^+]$.
- Consider initial solution conditions for accurate pH change assessment.
- Use logarithmic calculations carefully, especially with small or large concentrations.

Conclusion

Calculating the pH change when adding a strong acid like 2.4 M HCl involves understanding the dissociation behavior of the acid, the initial conditions of the solution, and the resulting concentration of hydrogen ions after mixing. Whether adding to pure water or an existing solution with a known pH, the fundamental principles remain the same: determine the moles of acid added, account for the total volume, and calculate the new hydrogen ion concentration to find the new pH. This process highlights the importance of stoichiometry and logarithmic functions in acid-base chemistry and provides a practical framework for analyzing and predicting pH changes in various chemical

Frequently Asked Questions

How do you calculate the pH change when 13 mL of 2.4 M HCl is added to 600 mL of solution?

First, determine the moles of HCl added (volume in liters × concentration), then find the initial moles of H+ in the solution, and finally calculate the new H+ concentration to find the pH before and after addition, allowing you to determine the change.

What is the initial pH of the solution before adding HCl if it contains 600 mL of water?

If the solution is pure water, the initial pH is approximately 7.0, since pure water has a neutral pH at 25°C. Otherwise, initial pH depends on the original solution's composition.

How many moles of HCl are added when 13 mL of 2.4 M HCl is introduced?

Number of moles = volume (L) × concentration (M) = 0.013 L × 2.4 mol/L = 0.0312 mol of HCl.

How does the addition of HCl affect the pH of a neutral solution?

Adding HCl introduces H+ ions, decreasing the pH and making the solution more acidic. The extent of the pH change depends on the total volume and buffering capacity of the solution.

If the solution initially contains a buffer, how does the addition of HCl influence the pH change?

The buffer will resist pH change by neutralizing some of the added H+ ions, resulting in a smaller pH decrease compared to a non-buffered solution.

How can you calculate the new pH after adding HCl to the solution?

Calculate the total moles of H+ ions after addition, divide by the total volume in liters, and then take the negative logarithm to find the new pH: pH = -log[H+].

What assumptions are made when calculating pH change in this scenario?

Assumptions include complete dissociation of HCl, no significant volume change due to addition, and that the solution behaves ideally without buffering effects unless specified.

How does the volume of the solution affect the magnitude of the pH change after adding HCl?

A larger volume dilutes the added H+ ions more, resulting in a smaller pH change. Conversely, a smaller volume leads to a more significant pH shift for the same amount of acid added.

Additional Resources

pH Calculation: Understanding the Impact of Adding 13 mL of 2.4 M HCl to 600 mL of Various Solutions

---

When delving into the realm of acid-base chemistry, one of the most fundamental and yet intriguing aspects is understanding how the addition of a strong acid like hydrochloric acid

(HCl) influences the pH of different solutions. This article offers an in-depth exploration of the calculation process involved in determining the pH change when 13 mL of 2.4 M HCl is added to 600 mL of various solutions. Whether you're a student, educator, or chemistry enthusiast, mastering this calculation enhances your grasp of solution chemistry and acid-base interactions.

Understanding the Core Concepts

Before diving into the calculations, it's essential to clarify some foundational concepts:

What is pH?

pH is a measure of the hydrogen ion concentration ([H^+]) in a solution. It is defined as:

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

A lower pH indicates higher acidity, while a higher pH suggests alkalinity.

Strong Acids and Bases

HCl is a strong acid, meaning it dissociates completely in aqueous solution:

$$ \text{HCl} \rightarrow \text{H}^+ + \text{Cl}^- $$

This complete dissociation simplifies the calculation of the number of moles of H^+ added to a solution.

Calculating Moles of HCl Added

Given the molarity (M) and volume (mL), the moles of HCl are:

$$ \text{moles} = \text{M} \times \text{L} $$

Step-by-Step Calculation Framework

To accurately determine the pH change, the process involves:

1. Calculating the moles of HCl added.
2. Determining the initial moles of H^+ or OH^- in the solution.
3. Considering the neutralization reaction between the acid and the solution's existing ions.
4. Calculating the remaining moles of H^+ after neutralization.
5. Computing the final concentration of H^+ in the total volume.
6. Deriving the new pH value.

Case Study: Adding 13 mL of 2.4 M HCl to 600 mL of Different Solutions

Let's apply this framework to various solutions, assuming different initial conditions.

Scenario 1: Pure Water (Neutral Solution)

Initial Conditions:
- Volume of water = 600 mL = 0.6 L
- Initial pH = 7 (neutral), so $[H^+] = 1\times10^{-7}$ M

Step 1: Moles of HCl added

$$ \text{Moles of HCl} = 2.4\, \text{M} \times 0.013\, \text{L} = 0.0312\, \text{mol} $$

Step 2: Initial moles of H^+ in water

$$ [H^+]_{initial} = 10^{-7}\, \text{M} $$

$$ \text{Moles} = 10^{-7}\, \text{mol/L} \times 0.6\, \text{L} = 6 \times 10^{-8}\, \text{mol} $$

This is negligible compared to the HCl moles added.

Step 3: Neutralization and final H^+ concentration
Since water is neutral, the added HCl fully dissociates, contributing 0.0312 mol of H^+ to the total volume.

Step 4: Total volume after addition

$$ 0.6\, \text{L} + 0.013\, \text{L} = 0.613\, \text{L} $$

Step 5: Final $[H^+]$

$$ [H^+] = \frac{0.0312\, \text{mol}}{0.613\, \text{L}} \approx 0.0509\, \text{M} $$

Step 6: Final pH

$$ pH = -\log(0.0509) \approx 1.29 $$

Result:
Adding 13 mL of 2.4 M HCl to pure water drops the pH from 7 to approximately 1.29, illustrating a significant increase in acidity.

Scenario 2: 600 mL of a Weak Basic Solution (e.g., 0.1 M NaOH)

Initial Conditions:
- Molarity of NaOH = 0.1 M
- Volume = 600 mL = 0.6 L

Initial moles of OH^-:
$$ 0.1\, \text{mol/L} \times 0.6\, \text{L} = 0.06\, \text{mol} $$

Step 1: Moles of HCl added (same as above)
$$ 0.0312\, \text{mol} $$

Step 2: Neutralization reaction
$$ \text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O} $$

Since HCl is a strong acid and NaOH is a strong base, the moles react in a 1:1 ratio:

- Moles of OH^- remaining after neutralization:
$$ 0.06\, \text{mol} - 0.0312\, \text{mol} = 0.0288\, \text{mol} $$

Step 3: Final solution volume
$$ 0.6\, \text{L} + 0.013\, \text{L} = 0.613\, \text{L} $$

Step 4: Final concentration of OH^-
$$ [OH^-] = \frac{0.0288\, \text{mol}}{0.613\, \text{L}} \approx 0.0469\, \text{M} $$

Step 5: Final pOH and pH
$$ pOH = -\log(0.0469) \approx 1.33 $$
$$ pH = 14 - pOH \approx 12.67 $$

Result:
The pH shifts from 13 (highly basic) to approximately 12.67, indicating a slight decrease in alkalinity due to the added acid but still remaining strongly basic.

Scenario 3: Acidic Solution (e.g., 0.05 M HNO_3)

Initial Conditions:
- Molarity of HNO_3 = 0.05 M
- Volume = 600 mL = 0.6 L

Initial moles of H^+:
$$ 0.05\, \text{mol/L} \times 0.6\, \text{L} = 0.03\, \text{mol} $$

Step 1: Moles of HCl added
$$ 0.0312\, \text{mol} $$

Step 2: Total H^+ after addition
$$ 0.03\, \text{mol} + 0.0312\, \text{mol} = 0.0612\, \text{mol} $$

Step 3: Total volume
$$ 0.613\, \text{L} $$

Step 4: Final $[H^+]$
$$ \frac{0.0612\, \text{mol}}{0.613\, \text{L}} \approx 0.1\, \text{M} $$

Step 5: Final pH
$$ pH = -\log(0.1) = 1 $$

Result:
The initial acidity increases slightly, lowering the pH from approximately 1.3 to 1.

Additional Factors and Considerations

While the above calculations provide a clear framework, real-world scenarios may involve additional factors that influence the pH change:

- Buffer Capacity: Solutions containing weak acids or bases can resist pH changes, requiring more acid or base to produce the same pH shift.
- Dilution Effects: Adding the acid increases the total volume, which dilutes ions and affects concentration calculations.
- Incomplete Dissociation: For weak acids or bases, dissociation is not complete, complicating the calculation.

Understanding these factors is crucial for accurate pH prediction and experimental planning.

Summary of Key Takeaways

- The amount of moles of HCl added can be precisely calculated from its molarity and volume.
- Neutralization reactions between strong acids and bases follow a 1:1 molar ratio, simplifying calculations.
- The pH change depends heavily on the initial solution's composition and buffering capacity.
- Final pH calculations involve considering total volume, remaining moles of ions, and their concentrations.

Practical Applications and Implications

Knowing how to calculate the pH change upon adding a specific volume of a strong acid to various solutions is invaluable in multiple fields:

- Laboratory Titrations: Precise pH measurements are essential for tit

Calculate The Ph Change That Results When 13 Ml Of 2 4 M Hcl Is Added To 600 Ml Of Each Of The Following

Calculate The Ph Change That Results When 13 Ml Of 2 4 M Hcl Is Added To 600 Ml Of Each Of The Following

Related Articles

- Are Visual Represenation Of Datat Or Information They Can Display Complex Information Quickly And Clearly
- Each Platform Varies In The Number Of Videos Or Images That Can Be Added For A Carousel Ad, But The Range
- A Logically Cohesive Module Is One, Where The Activities To Be Executed Are Chosen From Within The Module

Back to Home