A. Complete And Balance The Following Redox Reaction In Acidic SolutionCr2O72-(aq)+Pb2+(aq)-->Cr3+(aq)+PbO2(s)b.

A. Complete And Balance The Following Redox Reaction In Acidic SolutionCr2O72-(aq)+Pb2+(aq)-->Cr3+(aq)+PbO2(s)b.

Understanding how to correctly balance redox reactions, especially in acidic solutions, is crucial in chemistry. This process involves identifying oxidation and reduction processes, writing separate half-reactions, balancing atoms and charge, and then combining these to obtain the overall balanced equation. In this article, we will analyze the given reaction involving dichromate ions and lead(II) ions, determine the oxidation states, write the half-reactions, balance them in acidic medium, and finally combine them to produce the balanced overall reaction.

Understanding the Reaction Components

Reaction Overview

The unbalanced chemical equation is:

$Cr_2O_7^{2-}(aq) + Pb^{2+}(aq) \rightarrow Cr^{3+}(aq) + PbO_2(s)$

This reaction involves the reduction of dichromate ions and the oxidation of lead(II) ions. To balance this reaction properly, we need to:

- Determine the oxidation states of all elements involved.
- Write the oxidation and reduction half-reactions.
- Balance each half-reaction in acidic solution.
- Combine the balanced half-reactions to get the overall equation.

Oxidation States and Electron Transfer

- In $Cr_2O_7^{2-}$: Chromium is in +6 oxidation state.
- In Cr^{3+}: Chromium is in +3 oxidation state.
- In Pb^{2+}: Lead is in +2 oxidation state.
- In PbO_2 (lead dioxide): Lead is in +4 oxidation state.

From this, we observe:

- Chromium is reduced from +6 to +3.
- Lead is oxidized from +2 to +4.

Writing the Half-Reactions

Reduction Half-Reaction (Chromium)

The reduction involves chromium in dichromate ions gaining electrons to form Cr^{3+}.

Unbalanced reduction half-reaction:

$Cr_2O_7^{2-}$ + electrons → Cr^{3+}

Steps to balance:

1. Balance Cr atoms: 2 Cr on left, so:

$Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+}$

2. Balance oxygen atoms by adding H_2O:

There are 7 oxygens in $Cr_2O_7^{2-}$, so:

$Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$ (initially, but since oxygens are on the left, we will balance after balancing charges)

3. Balance hydrogen by adding H^+ (since in acidic solution):

- Since no hydrogen appears on the left, and oxygens are balanced with water, proceed to balance charges to find electrons.

4. Balance charge by adding electrons:

- Total charge on left: -2 (from $Cr_2O_7^{2-}$), on right: $2\times(+3) = +6$.

- To balance charges, electrons are added to the side that needs to gain electrons:

Let’s write the half-reaction:

$Cr_2O_7^{2-} + 14\ H^+ + 6\ e^- \rightarrow 2\ Cr^{3+} + 7\ H_2O$

Balanced reduction half-reaction:

$Cr_2O_7^{2-} + 14\ H^+ + 6\ e^- \rightarrow 2\ Cr^{3+} + 7\ H_2O$

Oxidation Half-Reaction (Lead)

Lead goes from +2 to +4, indicating oxidation, and forms PbO_2.

Unbalanced oxidation half-reaction:

$Pb^{2+} \rightarrow PbO_2$

1. Balance Pb atoms: 1 on each side.

2. Balance oxygen: PbO_2 has 2 oxygens; to balance oxygens, include H_2O or H^+ as appropriate.

In acidic solution, the oxidation of Pb^{2+} to PbO_2 involves:

$Pb^{2+} + 2\ H_2O \rightarrow PbO_2 + 4\ H^+ + 2\ e^-$

Explanation:

- The oxidation of lead from +2 to +4 involves loss of 2 electrons.

- To balance oxygen, water molecules are added to the reactant side, and H^+ to the product side.

Balanced oxidation half-reaction:

$Pb^{2+} + 2\ H_2O \rightarrow PbO_2 + 4\ H^+ + 2\ e^-$

Balancing the Half-Reactions

Reduction Half-Reaction

- Already balanced in terms of atoms and charge.

Final reduction half-reaction:

$Cr_2O_7^{2-} + 14\ H^+ + 6\ e^- \rightarrow 2\ Cr^{3+} + 7\ H_2O$

Oxidation Half-Reaction

- Already balanced:

$Pb^{2+} + 2\ H_2O \rightarrow PbO_2 + 4\ H^+ + 2\ e^-$

Combining the Half-Reactions

To combine, we need the electrons to cancel out:

- Reduction half-reaction involves 6 e^-
- Oxidation half-reaction involves 2 e^-

Find least common multiple (LCM) of electrons: 6

- Multiply oxidation half-reaction by 3:

$3\ Pb^{2+} + 6\ H_2O \rightarrow 3\ PbO_2 + 12\ H^+ + 6\ e^-$

- Use reduction half-reaction as is:

$Cr_2O_7^{2-} + 14\ H^+ + 6\ e^- \rightarrow 2\ Cr^{3+} + 7\ H_2O$

Now, add the two:

$Cr_2O_7^{2-} + 14\ H^+ + 6\ e^- + 3\ Pb^{2+} + 6\ H_2O \rightarrow 2\ Cr^{3+} + 7\ H_2O + 3\ PbO_2 + 12\ H^+ + 6\ e^-$

Electrons cancel out:

$Cr_2O_7^{2-} + 14\ H^+ + 3\ Pb^{2+} + 6\ H_2O \rightarrow 2\ Cr^{3+} + 7\ H_2O + 3\ PbO_2 + 12\ H^+$

Simplify by subtracting common H^+ and H_2O:

- Net H^+: 14 H^+ - 12 H^+ = 2 H^+ remaining on the left.

- Net H_2O: 6 H_2O - 7 H_2O = -1 H_2O, so subtract 1 H_2O from both sides:

$Cr_2O_7^{2-} + 14\ H^+ + 3\ Pb^{2+} \rightarrow 2\ Cr^{3+} + 3\ PbO_2 + H_2O + 2\ H^+$

Now, combine H^+ ions:

14 H^+ - 2 H^+ = 12 H^+ on the left.

Final balanced equation:

$Cr_2O_7^{2-} + 12\ H^+ + 3\ Pb^{2+} \rightarrow 2\ Cr^{3+} + 3\ PbO_2 + H_2O$

Final Balanced Equation

In standard form:

$Cr_2O_7^{2-}(aq) + 12\ H^+(aq) + 3\ Pb^{2+}(aq) \rightarrow 2\ Cr^{3+}(aq) + 3\ PbO_2(s) + H_2O(l)$

This is the complete and balanced redox reaction in acidic solution.

Summary of the Balancing Process

- Identify oxidation states and electron transfer.
- Write separate half-reactions for reduction and oxidation.
- Balance each half-reaction for atoms and charge, adding H^+, H_2O, and electrons as needed.
- Adjust the number of electrons to be equal in both half-reactions.
- Combine the half-reactions, cancel electrons, and simplify.
- Write the final balanced equation with states of matter.

Conclusion

Balancing redox reactions in acidic solutions requires careful analysis of oxidation states, precise writing of half-reactions, and systematic balancing of atoms and charges. The reaction between dichromate ions and lead(II) ions exemplifies this process, illustrating how electrons are transferred and how the overall reaction can be accurately balanced. Mastery of this methodology is essential in understanding complex redox processes encountered in chemistry, industrial applications, and environmental chemistry.

Frequently Asked Questions

What is the first step to balance the given redox reaction in acidic solution?

The first step is to separate the reaction into its oxidation and reduction half-reactions and then balance each half-reaction for atoms and charge.

How do you determine the oxidation states of elements in the reaction $Cr_2O_7^{2-}$ and Pb^{2+}?

In $Cr_2O_7^{2-}$, chromium has an oxidation state of +6, and in Pb^{2+}, lead has an oxidation state of +2.

What is the oxidation process in the reaction?

The oxidation process involves Pb^{2+} being oxidized to PbO_2, where lead's oxidation state increases from +2 to +4.

What is the reduction process in the reaction?

The reduction process involves $Cr_2O_7^{2-}$ being reduced to Cr^{3+}, where chromium's oxidation state decreases from +6 to +3.

How do you balance the oxygen and hydrogen atoms in the acidic medium?

Balance oxygen atoms by adding H_2O molecules and hydrogen atoms by adding H^{+} ions, ensuring the number of atoms is equal on both sides.

What are the balanced half-reactions for this redox process?

The oxidation half-reaction: $2Pb^{2+} + 4H_2O \rightarrow 2PbO_2 + 4H^{+} + 4e^{-}$ and the reduction half-reaction: $Cr_2O_7^{2-} + 14H^{+} + 6e^{-} \rightarrow 2Cr^{3+} + 7H_2O$.

How do you combine the half-reactions to get the balanced overall reaction?

Multiply the half-reactions to equalize the electrons transferred, then add them together, canceling out common terms, to obtain the balanced overall equation.

What is the final balanced redox reaction in acidic

solution?

The balanced reaction is: $Cr_2O_7^{2-} + 3Pb^{2+} + 8H^{+} \rightarrow 2Cr^{3+} + 3PbO_2 + 4H_2O$.

Why is it important to balance redox reactions in acidic solution?

Balancing in acidic solution ensures the conservation of mass and charge, which is essential for accurately describing electron transfer processes in the reaction.

Additional Resources

Complete And Balance The Following Redox Reaction In Acidic Solution $Cr_2O_7^{2-}(aq) + Pb^{2+}(aq) \rightarrow Cr^{3+}(aq) + PbO_2(s)$

Balancing redox reactions is a fundamental skill in chemistry, especially when dealing with oxidation-reduction processes in acidic solutions. The reaction between dichromate ions ($\text{Cr}_2\text{O}_7^{2-}$) and lead(II) ions ($\text{Pb}^{2+}$) leading to the formation of chromium(III) ions (Cr^{3+}) and lead dioxide (PbO_2) exemplifies a classic redox process involving multiple electron transfers. This reaction serves as an excellent case study to understand the systematic approach to balancing complex redox reactions in acidic media, emphasizing the importance of identifying oxidation states, separating half-reactions, and ensuring mass and charge balance.

Understanding the Reaction Components and Oxidation States

Before delving into the balancing process, it is crucial to analyze the oxidation states of all elements involved.

Oxidation States of Reactants

- Dichromate ion ($\text{Cr}_2\text{O}_7^{2-}$):
Chromium in dichromate has an oxidation state of +6. Since oxygen generally has an oxidation state of -2, and there are 7 oxygens, the calculation is:
$2 \times \text{Cr} + 7 \times (-2) = -2$
$2 \times \text{Cr} - 14 = -2$
$2 \times \text{Cr} = 12$
$\text{Cr} = +6$

- Lead(II) ion (Pb^{2+}):
Lead has an oxidation state of +2.

Oxidation States of Products

- Chromium(III) ion (Cr^{3+}):
Chromium is in the +3 oxidation state.
- Lead dioxide (PbO_2):
Oxygen is at -2; thus,
$\text{Pb} + 2 \times (-2) = 0$
$\text{Pb} = +4$

Key observations:
- Chromium is reduced from +6 to +3.
- Lead is oxidized from +2 to +4.

Step-by-Step Approach to Balancing the Redox Reaction

Balancing redox reactions involves several systematic steps, especially in acidic solutions where H^+ and H_2O are used to balance hydrogen and oxygen.

1. Write the Unbalanced Skeleton Equation

$$
\text{Cr}_2\text{O}_7^{2-} + \text{Pb}^{2+} \rightarrow \text{Cr}^{3+} + \text{PbO}_2
$$

2. Separate into Half-Reactions

Identify the oxidation and reduction processes:

- Oxidation (Pb):
$\text{Pb}^{2+} \rightarrow \text{Pb}^{4+}$ (in PbO_2)
- Reduction (Cr):
$\text{Cr}_2\text{O}_7^{2-} \rightarrow \text{Cr}^{3+}$

Balancing the Oxidation Half-Reaction

Step 1: Write the unbalanced oxidation half-reaction:

$$
\text{Pb}^{2+} \rightarrow \text{PbO}_2
$$

Step 2: Balance lead atoms:

$$
\text{Pb}^{2+} \rightarrow \text{PbO}_2
$$

Step 3: Balance oxygen atoms by adding H_2O:

Since PbO_2 contains 2 oxygen atoms, add 2 waters to the side with Pb^{2+}:

$$
\text{Pb}^{2+} + 2 \text{H}_2\text{O} \rightarrow \text{PbO}_2
$$

Step 4: Balance hydrogen atoms by adding H^+:

To balance 4 hydrogen atoms on the left, add 4 protons:

$$
\text{Pb}^{2+} + 2 \text{H}_2\text{O} \rightarrow \text{PbO}_2 + 4 \text{H}^+
$$

Step 5: Balance charge by adding electrons:

Calculate total charge on each side:

- Left: $+2 + 0 = +2$
- Right: $0 + 4 \times (+1) = +4$

To balance charges, add electrons to the more positive side:

Add 2 electrons to the left:

$$
\text{Pb}^{2+} + 2 \text{H}_2\text{O} + 2 e^- \rightarrow \text{PbO}_2 + 4 \text{H}^+
$$

Rearranged as:

$$
\text{Pb}^{2+} + 2 \text{H}_2\text{O} + 2 e^- \rightarrow \text{PbO}_2 + 4 \text{H}^+
$$

Balancing the Reduction Half-Reaction

Step 1: Write the unbalanced reduction half-reaction:

$$
\text{Cr}_2\text{O}_7^{2-} \rightarrow \text{Cr}^{3+}
$$

Step 2: Balance chromium atoms:

$$
\text{Cr}_2\text{O}_7^{2-} \rightarrow 2 \text{Cr}^{3+}
$$

Step 3: Balance oxygen atoms with H_2O:

Oxygen atoms: 7 on the left. Add 7 water molecules on the right:

$$
\text{Cr}_2\text{O}_7^{2-} \rightarrow 2 \text{Cr}^{3+} + 7 \text{H}_2\text{O}
$$

Step 4: Balance hydrogen atoms with H^+:

Hydrogen atoms: 14 on the right (from 7 water molecules). Add 14 protons to the left:

$$
\text{Cr}_2\text{O}_7^{2-} + 14 \text{H}^+ \rightarrow 2 \text{Cr}^{3+} + 7 \text{H}_2\text{O}
$$

Step 5: Balance charge with electrons:

Calculate total charges:

- Left: $-2 + 14(+1) = +12$
- Right: $2 \times (+3) = +6$

To balance, add 6 electrons to the left:

$$
\text{Cr}_2\text{O}_7^{2-} + 14 \text{H}^+ + 6 e^- \rightarrow 2 \text{Cr}^{3+} + 7 \text{H}_2\text{O}
$$

Combine the Half-Reactions

Our goal is to multiply the half-reactions to cancel electrons and then sum them.

- Oxidation half-reaction involves 2 electrons.
- Reduction half-reaction involves 6 electrons.

Multiply the oxidation half-reaction by 3:

$$
3 \left(\text{Pb}^{2+} + 2 \text{H}_2\text{O} + 2 e^- \rightarrow \text{PbO}_2 + 4 \text{H}^+ \right)
$$

which becomes:

$$
3 \text{Pb}^{2+} + 6 \text{H}_2\text{O} + 6 e^- \rightarrow 3 \text{PbO}_2 + 12 \text{H}^+
$$

Now, add the reduction half-reaction:

$$
\text{Cr}_2\text{O}_7^{2-} + 14 \text{H}^+ + 6 e^- \rightarrow 2 \text{Cr}^{3+} + 7 \text{H}_2\text{O}
$$

to get the overall reaction:

$$
\text{Cr}_2\text{O}_7^{2-} + 14 \text{H}^+ + 3 \text{Pb}^{2+} + 6 \text{H}_2\text{O} \rightarrow 2 \text{Cr}^{3+} + 7 \text{H}_2\text{O} + 3 \text{PbO}_2 + 12 \text{H}
$$

A Complete And Balance The Following Redox Reaction In Acidic Solutioncr2o72 Aq Pb2 Aq Gt Cr3 Aq Pbo2 S B

A Complete And Balance The Following Redox Reaction In Acidic Solutioncr2o72 Aq Pb2 Aq Gt Cr3 Aq Pbo2 S B

Related Articles

- Mariel Is Buying A Home And Is Looking Forward To The Tax Deductions. She May Deduct Which Of The Following?Furnishings
- What Are The General Responsibilities Of The Sharecropper, Based On This Contract? What

Are The Responsibilities

- The Point(0,0)is An Equilibrium For The Following System. Determine Whether It Is Stable Or Unstable.dx1/dt=2x1+11x2+22x1x2

Back to Home