B5 · Balance redox equations in neutral and basic solution

Completed notes | Notes 1/3

Decide whether H^+, OH^- and H_2O may stay, combine given half-reactions for a basic solution, balance a disproportionation and check every final equation.

1. What may stay in the final equation?

Read the condition first

The reaction happens in water, so first ask what kind of solution it is: acidic, basic or neutral. This is called the condition. Acidic water has extra **H^+**. Basic water has extra **OH^-** instead.

So a final equation for a basic solution should have OH^- and H_2O but no **H^+**. Basic water has almost no spare H^+ to use up.

Neutral water has no extra of either, but it is still water, so H_2O is there. A particle the reaction really makes, such as the H^+ made when Cl_2 reacts with water, **may stay**.

Writing the word “basic” beside $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ changes nothing. The half still takes in $8H^+$, so it is still **acidic**.

On the diploma: the words acidic, basic or neutral tell you what the correct option may contain, so check that first.

Watch: The question says acidic or basic (Tyler DeWitt, “How to Balance Redox Equations in Basic Solution”, 0:00)

H^+ used only for counting

Some H^+ are only counting tools. They show up on both sides and cancel. Given $H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$ and $H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$. Each half moves **two** electrons, so each is multiplied by **one**.

Add the halves. The $2H^+$ and $2e^-$ appear on **opposite sides**, so they cancel, because nothing real was used up or made. The net equation is $2H_2O_2 \rightarrow$ **$O_2 + 2H_2O$**.

Check it. Oxygen is **4** on each side, hydrogen is **4** on each side and charge is **0** on each side.

On the diploma: a correct neutral answer has no leftover counting H^+ or e^-, so scan the options for them.

Watch: H^+ is only a counting tool here (Tyler DeWitt, “How to Balance Redox Equations in Basic Solution”, 7:16)

Watch: Add OH^- to both sides (Tyler DeWitt, “How to Balance Redox Equations in Basic Solution”, 12:56)

Watch: H^+ and OH^- make water (Tyler DeWitt, “How to Balance Redox Equations in Basic Solution”, 13:58)

Watch: Subtract extra H_2O (Tyler DeWitt, “How to Balance Redox Equations in Basic Solution”, 14:28)

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Completed notes | Notes 2/3

2. Check an answer against its condition

Atoms, charge and the kind of solution

Here is how to check a half-reaction. Given the basic half $ClO^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$. Oxygen is **2** on each side and hydrogen is **2** on each side.

Charge on the left is $-1 + 0 - 2 = \mathbf{-3}$. Charge on the right is $-1 - 2 = \mathbf{-3}$. The two electrons on the left show **reduction**, because electrons are gained.

Last, look at what is in the equation. There is OH^- and H_2O and **no** H^+, so the half fits a basic solution.

Watch: Check atoms and charge at the end (Tyler DeWitt, “How to Balance Redox Equations in Basic Solution”, 16:02)

A balanced equation can still be a wrong answer

Given the acidic net equation $2MnO_4^- + 6H^+ + 5NO_2^- \rightarrow 2Mn^{2+} + 3H_2O + 5NO_3^-$. Atoms and charge **balance**.

But it contains **$6H^+$**. A question that says “basic solution” needs OH^- and H_2O and no H^+, so this equation is **the wrong answer**.

Last step of every balance: check what is left against the **condition**, as well as atoms and charge.

On the diploma: a wrong option is often balanced correctly but written for the wrong condition, so always compare it with the condition in the question.

Watch: Acidic answer versus basic answer (Tyler DeWitt, “How to Balance Redox Equations in Basic Solution”, 12:25)

> Basic half: $ClO^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$
>
> Oxygen: left 2; right 2. Hydrogen: left 2; right 2
>
> Charge: left -3; right -3
>
> Contains H^+? No, so it fits a basic solution

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Completed notes | Notes 3/3

3. Combine halves and disproportionation

Combine two given half-reactions for a basic solution

Electrons given must equal electrons taken, just like in B3. Given $MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_2 + 4OH^-$ and $2I^- \rightarrow I_2 + 2e^-$. The smallest number that both 3 and 2 divide into is **6**.

So multiply the first half by **2** ($2 \times 3 = 6$) and the second by **3** ($3 \times 2 = 6$). Multiply every term, including the H_2O and OH^-.

Add the halves and cancel the **6e⁻**. Nothing else is the same on both sides, so the net equation is $2MnO_4^- + 4H_2O + 6I^- \rightarrow 2MnO_2 + 3I_2 + \mathbf{8OH^-}$.

The finished equation has OH^- and H_2O and **no** H^+, so it fits a basic solution.

On the diploma: matching the electrons is the first step, and wrong options often come from a wrong multiplier.

Watch: Add the halves and cancel (Tyler DeWitt, “How to Balance Redox Equations in Basic Solution”, 10:53)

Disproportionation in acid

Look at $Cl_2 \rightarrow ClO_3^- + Cl^-$ in acid. This is a disproportionation, which you met in B2: one substance is oxidized and reduced at once. Cl goes 0 to +5 (**oxidation**) and 0 to −1 (**reduction**).

The oxidation half is $Cl_2 + 6H_2O \rightarrow 2ClO_3^- + 12H^+ + \mathbf{10e^-}$. The reduction half is $Cl_2 + 2e^- \rightarrow 2Cl^-$. The reduction half takes 2 electrons and the oxidation half gives 10, so multiply the reduction half by **five**. Add the halves and cancel the electrons.

That gives $6Cl_2 + 6H_2O \rightarrow 2ClO_3^- + 12H^+ + 10Cl^-$. Every coefficient divides by **two**. The lowest whole-number equation is $3Cl_2 + 3H_2O \rightarrow ClO_3^- + 6H^+ + 5Cl^-$. The charge is **0** on each side.

On the diploma: reduce to the lowest whole numbers at the end.

The same reactant in base

Now the same idea in base. Given $Cl_2 \rightarrow ClO^- + Cl^-$ in base. Chlorine goes from 0 to **+1** in ClO^- and from 0 to **−1** in Cl^-.

Both halves start with Cl_2: $Cl_2 + 4OH^- \rightarrow 2ClO^- + 2H_2O + 2e^-$ and $Cl_2 + 2e^- \rightarrow 2Cl^-$. Each half moves two electrons, so add them to get $2Cl_2 + 4OH^- \rightarrow 2ClO^- + 2Cl^- + 2H_2O$.

Every coefficient is even, so divide every term by **2**. An equation shows a ratio of particles, so this keeps it balanced. The lowest whole-number equation is $Cl_2 + 2OH^- \rightarrow ClO^- + Cl^- + \mathbf{H_2O}$.

On the diploma: a numerical-response question may ask for one coefficient, so simplify before you record it.

B5 · Worked practice

Balance redox equations in neutral and basic solution | Questions 1-2 of 12

State symbols: (s) solid, (aq) dissolved in water, (l) liquid, (g) gas. Use the supplied reaction and conditions. B4 values name the Alberta Chemistry 30 Data Booklet table where used.

B5-P01 · Easy · A reaction takes place in basic solution. Which of H^+ and OH^- may remain in the final equation, and which may not? Explain in one sentence.

1. OH^- may remain, together with H_2O. H^+ may not remain.

2. Basic water has extra OH^- and no extra H^+, so a final basic equation should not be left with H^+.

B5-P02 · Easy · Given neutral water: $Cl_2 + H_2O \rightarrow Cl^- + HOCl + H^+$. Check atoms and charge, and say why the H^+ may stay in the final equation.

1. Cl is 2 on each side, H is 2 and O is 1. Charge is 0 on the left and $-1 + 0 + 1 = 0$ on the right.

2. The H^+ is a real product of the reaction, not a counting H^+ that appears on both sides, so it may stay in neutral water.

B5 · Worked practice

Balance redox equations in neutral and basic solution | Questions 3-4 of 12

B5-P03 · Easy · A student copies $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ and writes “basic solution” above it. Explain why the equation is still acidic.

1. The equation still takes in $8H^+$. A basic solution has almost no H^+, so it describes an acidic solution.

2. The label does not change what is in the equation. A basic equation may contain OH^- and H_2O but should contain no H^+.

B5-P04 · Easy · Given $2H_2O_2 \rightarrow O_2 + 2H_2O$ in neutral water, state the oxidation number of oxygen in H_2O_2, in O_2 and in H_2O. Which change is oxidation and which is reduction?

1. Oxygen is -1 in H_2O_2 (peroxide), 0 in O_2 and -2 in H_2O.

2. The change from -1 to 0 is oxidation (to O_2). The change from -1 to -2 is reduction (to H_2O). The same reactant is both, so this is disproportionation.

B5 · Worked practice

Balance redox equations in neutral and basic solution | Questions 5-6 of 12

B5-P05 · Medium · Given basic solution: $ClO^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$ and $Fe(OH)_2 + OH^- \rightarrow Fe(OH)_3 + e^-$. Write the lowest whole-number net equation.

1. The reduction takes two electrons and each oxidation gives one, so multiply the oxidation by 2: $2Fe(OH)_2 + 2OH^- \rightarrow 2Fe(OH)_3 + 2e^-$.

2. Add and cancel $2e^-$ and $2OH^-$: $ClO^- + H_2O + 2Fe(OH)_2 \rightarrow Cl^- + 2Fe(OH)_3$.

3. Check: O 1 + 1 + 4 = 6 and 6; H 2 + 4 = 6 and 6; charge −1 on each side.

B5-P06 · Medium · Given basic solution: $BrO_3^- + 3H_2O + 6e^- \rightarrow Br^- + 6OH^-$ and $2I^- \rightarrow I_2 + 2e^-$. Write the lowest whole-number net equation.

1. The reduction takes 6 electrons and each oxidation gives 2, so multiply the oxidation by 3: $6I^- \rightarrow 3I_2 + 6e^-$.

2. Add and cancel $6e^-$: $BrO_3^- + 3H_2O + 6I^- \rightarrow Br^- + 3I_2 + 6OH^-$.

3. Check: Br 1, I 6, O 6 and H 6 on each side. Charge is −7 on the left and −7 on the right.

B5 · Worked practice

Balance redox equations in neutral and basic solution | Questions 7-8 of 12

B5-P07 · Medium · A question asks for the net equation in basic solution. A student gives $Cr_2O_7^{2-} + 14H^+ + 6I^- \rightarrow 2Cr^{3+} + 7H_2O + 3I_2$. It balances. Why is it not a correct answer?

1. The equation contains $14H^+$ and the product Cr^{3+}, both of which belong to an acidic solution.

2. A final equation for a basic solution should contain OH^- and H_2O but no H^+. Balanced atoms and charge are not enough; what is in the equation must fit the condition.

B5-P08 · Medium · Check $Zn + 2OH^- + 2H_2O \rightarrow Zn(OH)_4^{2-} + H_2$ for atoms and charge. State whether it fits a basic solution.

1. Zn is 1 on each side. O is 2 + 2 = 4 on the left and 4 on the right. H is 2 + 4 = 6 on the left and 4 + 2 = 6 on the right.

2. Charge is −2 on the left and −2 on the right. The equation contains OH^- and H_2O and no H^+, so it fits a basic solution.

B5 · Worked practice

Balance redox equations in neutral and basic solution | Questions 9-10 of 12

B5-P09 · Challenging · Given basic solution: $ClO^- \rightarrow ClO_3^- + Cl^-$. The oxidation half is $ClO^- + 4OH^- \rightarrow ClO_3^- + 2H_2O + 4e^-$ and the reduction half is $ClO^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$. Write the lowest whole-number net equation.

1. The oxidation gives 4 electrons and the reduction takes 2, so multiply the reduction by 2: $2ClO^- + 2H_2O + 4e^- \rightarrow 2Cl^- + 4OH^-$.

2. Add and cancel $4e^-$, $4OH^-$ and $2H_2O$: $3ClO^- \rightarrow ClO_3^- + 2Cl^-$.

3. Check: Cl 3, O 3; charge −3 on each side. Chlorine goes from +1 to +5 and from +1 to −1, so ClO^- is both oxidized and reduced.

B5-P10 · Challenging · A student gives $MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O$ as a half for a basic solution. Atoms and charge balance. Why is it not correct?

1. Oxygen is 4 on each side, hydrogen is 4 and charge is 0 on each side, so it does balance.

2. It contains $4H^+$, which a basic solution does not have. A half for a basic solution should contain OH^- and H_2O and no H^+.

B5 · Worked practice

Balance redox equations in neutral and basic solution | Questions 11-12 of 12

B5-P11 · Challenging · Given basic solution: reduction $ClO^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$ and oxidation $2I^- \rightarrow I_2 + 2e^-$. Write the lowest whole-number net equation.

1. Each half moves two electrons, so no multiplier is needed.

2. Add and cancel $2e^-$: $ClO^- + H_2O + 2I^- \rightarrow Cl^- + I_2 + 2OH^-$.

3. Check: Cl 1, O 2, H 2, I 2 on each side. Charge is -3 on the left and -3 on the right.

B5-P12 · Challenging · Given basic solution: $Br_2 \rightarrow BrO_3^- + Br^-$. The oxidation half is $Br_2 + 12OH^- \rightarrow 2BrO_3^- + 6H_2O + 10e^-$ and the reduction half is $Br_2 + 2e^- \rightarrow 2Br^-$. Write the lowest whole-number net equation.

1. The oxidation gives 10 electrons and the reduction takes 2, so multiply the reduction by 5: $5Br_2 + 10e^- \rightarrow 10Br^-$.

2. Add and cancel $10e^-$: $6Br_2 + 12OH^- \rightarrow 2BrO_3^- + 10Br^- + 6H_2O$. Divide every coefficient by 2: $3Br_2 + 6OH^- \rightarrow BrO_3^- + 5Br^- + 3H_2O$.

3. Check: Br 6, O 6, H 6 on each side; charge -6 on each side.