B5 · Balance redox equations in neutral and basic solution

Completed notes

Goals: Decide whether H^+, OH^- and H_2O may stay in an acidic, neutral or basic equation. Balance a redox equation for a basic solution, with the half-reactions given or not given. Check every final equation against the kind of solution in the question.

Slide 1 of 10 · The solution decides what may stay

First read the **condition**: acidic, neutral or basic. It decides whether H^+, OH^- or H_2O may stay in the final equation.

An acidic equation may contain H_2O and **H^+**. A basic equation may contain H_2O and **OH^-** and should contain no H^+.

The half $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ takes in $8H^+$, so it describes an **acidic** solution. Writing "basic" beside it changes nothing.

In neutral water, H^+ or OH^- may stay only if the reaction **really makes** it, as in $Cl_2 + H_2O \rightarrow Cl^- + HOCl + H^+$.

Example. Given $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$. Could this half-reaction be used for a basic solution? Explain using what is in the equation.

1. No. The half takes in $8H^+$ on the left.

2. A basic solution has almost no H^+, so the half describes an acidic solution.

3. Writing the word basic beside it does not remove the H^+.

Try it. Given neutral water: $Cl_2(g) + H_2O(l) \rightarrow Cl^-(aq) + HOCl(aq) + H^+(aq)$. Check atoms and charge, and give the oxidation number of chlorine in Cl_2, Cl^- and HOCl.

1. Atoms: Cl 2 = 1 + 1; H 2 = 1 + 1; O 1 = 1.

2. Charge: left 0; right $(-1) + 0 + (+1) = 0$.

3. Cl_2 is an element by itself, so Cl is 0. In Cl^- it is -1. In HOCl, H is $+1$ and O is -2, so Cl is $+1$.

4. The H^+ is a real product, so it may stay.

Slide 2 of 10 · Neutral water: hydrogen peroxide breaks down

In H_2O_2 oxygen is **−1**. In H_2O it is −2 and in O_2 it is 0. One reactant that is both oxidized and reduced shows **disproportionation**.

Oxidation: $H_2O_2 \rightarrow O_2 + 2H^+ + \mathbf{2e^-}$. Reduction: $H_2O_2 + 2H^+ + 2e^- \rightarrow \mathbf{2H_2O}$.

Each half moves 2 electrons, so multiply each half by **1** and add them.

The $2H^+$ and $2e^-$ are on both sides and cancel, so they were only **counting** H^+. Final equation: $2H_2O_2 \rightarrow O_2 + 2H_2O$.

Example. Given neutral water: H_2O_2 breaks down into H_2O and O_2. Write the two halves, the balanced overall equation, and explain why the H^+ cancel.

1. Oxygen: −1 in H_2O_2; −2 in H_2O (reduction); 0 in O_2 (oxidation).

2. Oxidation: $H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$. Reduction: $H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$.

3. Each half moves 2 electrons, so add them as they are: $2H_2O_2 + 2H^+ + 2e^- \rightarrow O_2 + 2H^+ + 2e^- + 2H_2O$.

4. Cancel $2H^+$ and $2e^-$: $2H_2O_2 \rightarrow O_2 + 2H_2O$.

5. The H^+ cancel because they were only counting H^+: none is really used up or made.

6. Check: O 4 = 2 + 2; H 4 = 4; charge 0 = 0.

Try it. Given neutral water: $Cl_2 \rightarrow HOCl + Cl^-$. Use the halves $Cl_2 + 2H_2O \rightarrow 2HOCl + 2H^+ + 2e^-$ and $Cl_2 + 2e^- \rightarrow 2Cl^-$. Write the lowest whole-number equation and say whether H^+ cancels.

1. Each half moves 2 electrons, so add them directly and cancel $2e^-$.

2. $2Cl_2 + 2H_2O \rightarrow 2HOCl + 2H^+ + 2Cl^-$.

3. Every coefficient is even. Divide by 2: $Cl_2 + H_2O \rightarrow HOCl + H^+ + Cl^-$.

4. Check: Cl 2 = 1 + 1; H 2 = 1 + 1; O 1 = 1; charge 0 = 0 + 1 − 1.

5. H^+ is on one side only, so it does not cancel. It is a real product.

Slide 3 of 10 · Combine two given basic half-reactions

Reduction takes 3 electrons and oxidation gives 2. The smallest number both reach is **6**.

Multiply the reduction by **2** and the oxidation by **3**. Multiply every term, including H_2O and OH^-.

Add and cancel the **6e**$^-$. H_2O is only on the left and OH^- only on the right, so neither cancels. Never cancel H_2O against OH^-.

Check: Mn 2, I 6, O 12, H 8 and charge −8 on each side. A basic equation has OH^- and H_2O and no **H**$^+$.

Example. Given basic solution: reduction $MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_2 + 4OH^-$ and oxidation $2I^- \rightarrow I_2 + 2e^-$. Write the lowest whole-number net equation and check it.

1. The smallest number that both 3 and 2 reach is 6. Multiply the reduction by 2 and the oxidation by 3.

2. $2MnO_4^- + 4H_2O + 6e^- \rightarrow 2MnO_2 + 8OH^-$ and $6I^- \rightarrow 3I_2 + 6e^-$.

3. Add and cancel $6e^-$: $2MnO_4^- + 4H_2O + 6I^- \rightarrow 2MnO_2 + 3I_2 + 8OH^-$.

4. Check: Mn 2; I 6; O 8 + 4 = 12 and 4 + 8 = 12; H 8 = 8.

5. Charge: −2 − 6 = −8 and −8. Coefficients 2, 4, 6, 2, 3, 8 share no factor above 1.

Try it. Given basic solution: reduction $MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_2 + 4OH^-$ and oxidation $Fe(OH)_2 + OH^- \rightarrow Fe(OH)_3 + e^-$. Write the lowest whole-number net equation and check it.

1. The reduction takes $3e^-$ and the oxidation gives $1e^-$. Multiply the oxidation by 3.

2. $3Fe(OH)_2 + 3OH^- \rightarrow 3Fe(OH)_3 + 3e^-$.

3. Add and cancel $3e^-$. Then cancel $3OH^-$ on the left against $4OH^-$ on the right, leaving $1OH^-$ on the right.

4. $MnO_4^- + 2H_2O + 3Fe(OH)_2 \rightarrow MnO_2 + 3Fe(OH)_3 + OH^-$.

5. Check atoms: Mn 1; Fe 3; O 4 + 2 + 6 = 12 and 2 + 9 + 1 = 12.

6. H: 4 + 6 = 10 and 9 + 1 = 10. Charge: −1 and −1.

Slide 4 of 10 · Method: how to balance a redox equation in basic solution

Step 1. Read the **condition**. The solution is basic.

Step 2. Keep the halves as written with **OH^-** and H_2O. Do not add **H^+**.

Step 3. Make the **electrons** equal.

Step 4. **Add** the halves and cancel any e^-, H_2O or OH^- on both sides.

Step 5. Check: no H^+, equal **atoms** and equal **charge**.

Example. Use the method on $MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_2 + 4OH^-$ and $2I^- \rightarrow I_2 + 2e^-$. Write the lowest whole-number net equation.

1. Steps 1 and 2: the solution is basic and the halves already use OH^- and H_2O.

2. Step 3: the smallest common number of electrons is 6. Multiply the first half by 2 and the second by 3.

3. Step 4: add and cancel the $6e^-$.

4. Step 5: $2MnO_4^- + 4H_2O + 6I^- \rightarrow 2MnO_2 + 3I_2 + 8OH^-$. No H^+. Charge is -8 on each side.

Try it. Use the method on $ClO^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$ and $Zn + 2OH^- \rightarrow Zn(OH)_2 + 2e^-$. Write the net equation.

1. Both halves have $2e^-$, so no multiplying is needed.

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

3. Check: no H^+. Charge is -1 on each side.

Slide 5 of 10 · Your turn: hypochlorite and chromium(III) hydroxide

Reduction takes 2 electrons and oxidation gives 3. The smallest number both reach is 6. Multiply the reduction by **3** and the oxidation by **2**.

In $Cr(OH)_3$ the 2 in front multiplies Cr and everything inside the **brackets**.

After adding and cancelling $6e^-$, OH^- is 10 on the left and 6 on the right, so **4**OH^- stays on the left. H_2O is 3 left and 8 right, so **5**H_2O stays on the right.

Net: $2Cr(OH)_3 + 3ClO^- + 4OH^- \rightarrow 2CrO_4^{2-} + 3Cl^- + 5H_2O$. Check O 13, H 10 and charge **−7** on each side.

Example. Given basic solution: reduction $ClO^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$ and oxidation $Cr(OH)_3 + 5OH^- \rightarrow CrO_4^{2-} + 4H_2O + 3e^-$. Write your multipliers, then the lowest whole-number net equation.

1. The smallest number that both 2 and 3 reach is 6. Multiply the reduction by 3 and the oxidation by 2.

2. $3ClO^- + 3H_2O + 6e^- \rightarrow 3Cl^- + 6OH^-$ and $2Cr(OH)_3 + 10OH^- \rightarrow 2CrO_4^{2-} + 8H_2O + 6e^-$.

3. Add and cancel $6e^-$, $6OH^-$ and $3H_2O$.

4. $2Cr(OH)_3 + 3ClO^- + 4OH^- \rightarrow 2CrO_4^{2-} + 3Cl^- + 5H_2O$.

5. Check: O 13 = 13; H 10 = 10; charge −7 = −7.

Try it. Given basic solution: reduction $BrO^- + H_2O + 2e^- \rightarrow Br^- + 2OH^-$ and oxidation $Cr(OH)_3 + 5OH^- \rightarrow CrO_4^{2-} + 4H_2O + 3e^-$. Write your multipliers, then the lowest whole-number net equation.

1. The smallest number that both 2 and 3 reach is 6. Multiply the reduction by 3 and the oxidation by 2.

2. $3BrO^- + 3H_2O + 6e^- \rightarrow 3Br^- + 6OH^-$ and $2Cr(OH)_3 + 10OH^- \rightarrow 2CrO_4^{2-} + 8H_2O + 6e^-$.

3. Add and cancel $6e^-$, $6OH^-$ and $3H_2O$.

4. $2Cr(OH)_3 + 3BrO^- + 4OH^- \rightarrow 2CrO_4^{2-} + 3Br^- + 5H_2O$.

5. Check: O 13 = 13; H 10 = 10; charge −7 = −7.

Slide 6 of 10 · No half-reactions given: balance as in acid, then change to basic

When no half-reactions are given for a basic solution, first balance the equation as in **acid** (B3).

Then add **OH^-** to both sides, one for each H^+.

Each H^+ joins an OH^- to make **H_2O**. Cancel the H_2O that is on both sides.

Check atoms and charge. A basic equation has no **H^+** in it.

Example. Given basic solution: $ClO^- + I^- \rightarrow Cl^- + I_2$. No half-reactions are given. Write the balanced net equation and check it.

1. Cl goes +1 → −1: reduced. I goes −1 → 0: oxidized.

2. Reduction: $ClO^- + 2H^+ + 2e^- \rightarrow Cl^- + H_2O$. Oxidation: $2I^- \rightarrow I_2 + 2e^-$.

3. As in acid: $ClO^- + 2H^+ + 2I^- \rightarrow Cl^- + H_2O + I_2$.

4. Add $2OH^-$ to each side. $2H^+ + 2OH^-$ make $2H_2O$.

5. $ClO^- + 2H_2O + 2I^- \rightarrow Cl^- + H_2O + I_2 + 2OH^-$. Cancel $1H_2O$ from each side.

6. $ClO^- + H_2O + 2I^- \rightarrow Cl^- + I_2 + 2OH^-$.

7. Check: Cl 1, O 2, H 2, I 2 on each side. Charge: −3 and −3. There is no H^+.

Try it. Given basic solution: $ClO^- + Br^- \rightarrow Cl^- + Br_2$. No half-reactions are given. Write the balanced net equation and check it.

1. Cl goes +1 → −1: reduced. Br goes −1 → 0: oxidized.

2. Reduction: $ClO^- + 2H^+ + 2e^- \rightarrow Cl^- + H_2O$. Oxidation: $2Br^- \rightarrow Br_2 + 2e^-$.

3. As in acid: $ClO^- + 2H^+ + 2Br^- \rightarrow Cl^- + H_2O + Br_2$.

4. Add $2OH^-$ to each side. $2H^+ + 2OH^-$ make $2H_2O$. Cancel $1H_2O$ from each side.

5. $ClO^- + H_2O + 2Br^- \rightarrow Cl^- + Br_2 + 2OH^-$.

6. Check: Cl 1, O 2, H 2, Br 2 on each side. Charge: −3 and −3. There is no H^+.

Slide 7 of 10 · Your turn: permanganate and nitrite in basic solution

Find the changes with **oxidation numbers** first.

Balance each half as in acid, make the **electrons** equal, add and cancel.

Count the H^+ that are left. Add that many **OH^-** to both sides.

Turn each H^+ and OH^- pair into **H_2O**, cancel H_2O on both sides, and check.

Example. Given basic solution: $MnO_4^- + NO_2^- \rightarrow MnO_2 + NO_3^-$. No half-reactions are given. Write the balanced net equation and check it.

1. Mn goes +7 → +4: reduced. N goes +3 → +5: oxidized.

2. Reduction: $MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O$. Oxidation: $NO_2^- + H_2O \rightarrow NO_3^- + 2H^+ + 2e^-$.

3. Multiply the reduction by 2 and the oxidation by 3. Add and cancel.

4. As in acid: $2MnO_4^- + 2H^+ + 3NO_2^- \rightarrow 2MnO_2 + H_2O + 3NO_3^-$.

5. Add $2OH^-$ to each side. $2H^+ + 2OH^-$ make $2H_2O$. Cancel $1H_2O$ from each side.

6. $2MnO_4^- + H_2O + 3NO_2^- \rightarrow 2MnO_2 + 3NO_3^- + 2OH^-$.

7. Check: Mn 2, N 3, O 15, H 2 on each side. Charge: −5 and −5. There is no H^+.

Try it. Given basic solution: $MnO_4^- + I^- \rightarrow MnO_2 + I_2$. No half-reactions are given. Write the balanced net equation. Then compare it with the equation on slide 3, where the half-reactions were given.

1. Mn goes +7 → +4: reduced. I goes −1 → 0: oxidized.

2. Reduction: $MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O$. Oxidation: $2I^- \rightarrow I_2 + 2e^-$.

3. Multiply the reduction by 2 and the oxidation by 3. Add and cancel $6e^-$.

4. As in acid: $2MnO_4^- + 8H^+ + 6I^- \rightarrow 2MnO_2 + 4H_2O + 3I_2$.

5. Add $8OH^-$ to each side. $8H^+ + 8OH^-$ make $8H_2O$. Cancel $4H_2O$ from each side.

6. $2MnO_4^- + 4H_2O + 6I^- \rightarrow 2MnO_2 + 3I_2 + 8OH^-$.

7. Check: charge −8 and −8. There is no H^+. It is the same equation as on slide 3.

Slide 8 of 10 · Balance a disproportionation

In $Cl_2 \rightarrow ClO_3^- + Cl^-$, chlorine goes from 0 to **+5** in ClO_3^- and from 0 to **−1** in Cl^-. One species both oxidized and reduced is **disproportionation**.

Oxidation half: $Cl_2 + 6H_2O \rightarrow 2ClO_3^- + 12H^+ + \mathbf{10}e^-$. Reduction half: $Cl_2 + 2e^- \rightarrow 2Cl^-$.

Multiply the reduction half by **5** to match 10 electrons. Adding gives $6Cl_2 + 6H_2O \rightarrow 2ClO_3^- + 12H^+ + 10Cl^-$.

Every coefficient divides by 2, so the lowest whole-number equation is $3Cl_2 + 3H_2O \rightarrow ClO_3^- + 6H^+ + 5Cl^-$. Divide **every** term.

Example. Given acidic water: $Cl_2 \rightarrow ClO_3^- + Cl^-$. Chlorine is both oxidized and reduced. Construct both halves, combine them and write the lowest whole-number equation.

1. Oxidation half: $Cl_2 + 6H_2O \rightarrow 2ClO_3^- + 12H^+ + 10e^-$. Reduction half: $Cl_2 + 2e^- \rightarrow 2Cl^-$.

2. Multiply the reduction half by 5: $5Cl_2 + 10e^- \rightarrow 10Cl^-$.

3. Add and cancel $10e^-$: $6Cl_2 + 6H_2O \rightarrow 2ClO_3^- + 12H^+ + 10Cl^-$.

4. Divide every term by 2: $3Cl_2 + 3H_2O \rightarrow ClO_3^- + 6H^+ + 5Cl^-$.

5. Check: Cl 6 = 6; O 3 = 3; H 6 = 6; charge $0 = -1 + 6 - 5 = 0$.

Try it. Given acidic water: $Br_2 \rightarrow BrO_3^- + Br^-$. Bromine is both oxidized and reduced. Use the halves $Br_2 + 6H_2O \rightarrow 2BrO_3^- + 12H^+ + 10e^-$ and $Br_2 + 2e^- \rightarrow 2Br^-$. Write the lowest whole-number equation.

1. The halves give 10 and take 2 electrons. Multiply the reduction half by 5: $5Br_2 + 10e^- \rightarrow 10Br^-$.

2. Add and cancel $10e^-$: $6Br_2 + 6H_2O \rightarrow 2BrO_3^- + 12H^+ + 10Br^-$.

3. Divide every term by 2: $3Br_2 + 3H_2O \rightarrow BrO_3^- + 6H^+ + 5Br^-$.

4. Check: Br 6 = 1 + 5; O 3 = 3; H 6 = 6; charge $0 = -1 + 6 - 5 = 0$.

Slide 9 of 10 · A disproportionation in basic water

In base, $Cl_2 \rightarrow ClO^- + Cl^-$ is disproportionation. Chlorine goes from 0 to **+1** in ClO^- and from 0 to **−1** in Cl^-.

Oxidation: $Cl_2 + 4OH^- \rightarrow 2ClO^- + 2H_2O + \mathbf{2e^-}$. Reduction: $Cl_2 + 2e^- \rightarrow 2Cl^-$.

Both halves move 2 electrons. Add and cancel them: $2Cl_2 + 4OH^- \rightarrow 2ClO^- + 2Cl^- + 2H_2O$.

Every coefficient is even, so divide every term by **2**: $Cl_2 + 2OH^- \rightarrow ClO^- + Cl^- + H_2O$. Charge is **−2** on each side.

Example. Given basic solution: $Cl_2 \rightarrow ClO^- + Cl^-$. Use the halves $Cl_2 + 4OH^- \rightarrow 2ClO^- + 2H_2O + 2e^-$ and $Cl_2 + 2e^- \rightarrow 2Cl^-$. Write the lowest whole-number equation and check it.

1. Both halves move 2 electrons, so add them and cancel $2e^-$.

2. $2Cl_2 + 4OH^- \rightarrow 2ClO^- + 2Cl^- + 2H_2O$.

3. Every coefficient is even. Divide every term by 2: $Cl_2 + 2OH^- \rightarrow ClO^- + Cl^- + H_2O$.

4. Check: Cl 2 = 1 + 1; O 2 = 1 + 1; H 2 = 2; charge $-2 = -1 - 1$.

Try it. Given basic solution: $Br_2 \rightarrow BrO^- + Br^-$. Use the halves $Br_2 + 4OH^- \rightarrow 2BrO^- + 2H_2O + 2e^-$ and $Br_2 + 2e^- \rightarrow 2Br^-$. Write the lowest whole-number equation and check it.

1. Both halves move 2 electrons, so add them and cancel $2e^-$.

2. $2Br_2 + 4OH^- \rightarrow 2BrO^- + 2Br^- + 2H_2O$.

3. Divide every term by 2: $Br_2 + 2OH^- \rightarrow BrO^- + Br^- + H_2O$.

4. Check: Br 2 = 1 + 1; O 2 = 1 + 1; H 2 = 2; charge $-2 = -1 - 1$.

Slide 10 of 10 · Same method, new conditions

Step 1: read the **condition**. Step 2: combine halves and match electrons. Step 3: cancel only what is the **same** on both sides. Step 4: check what is left against the condition.

Acidic allows H_2O and **H^+**. Basic allows H_2O and **OH^-**.

Neutral allows H_2O, plus H^+ or OH^- only if the reaction **really makes** them.

An equation can balance in atoms and charge and still be the wrong answer if it does not match the **condition**.

Example. Given $2MnO_4^- + 6H^+ + 5NO_2^- \rightarrow 2Mn^{2+} + 3H_2O + 5NO_3^-$. A question asks for a basic solution. Does the equation balance, and is it a correct final answer?

1. Atoms: Mn 2; N 5; O 8 + 10 = 18 and 3 + 15 = 18; H 6 = 6.

2. Charge: $-2 + 6 - 5 = -1$ and $+4 - 5 = -1$. It balances.

3. Step 4: the equation contains H^+, but basic water allows only OH^- and H_2O.

4. So it is not a correct final answer for a basic solution.

Try it. A question asks for neutral water. Is $2H_2O_2 \rightarrow O_2 + 2H_2O$ a correct final answer? Check atoms, charge and what is left.

1. Atoms: O 4 = 2 + 2; H 4 = 4. Charge: 0 = 0.

2. Besides H_2O_2 and O_2, the equation has only H_2O. No H^+, OH^- or e^- is left over, so it is a correct final answer.

B5 · Worked practice

Balance redox equations in neutral and basic solution | Questions 1-2 of 9

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

Use Method: how to balance a redox equation in basic solution. The reduction takes two electrons, so multiply the oxidation by 2: $2Fe(OH)_2 + 2OH^- \rightarrow 2Fe(OH)_3 + 2e^-$. Add and cancel $2e^-$ and $2OH^-$: $ClO^- + H_2O + 2Fe(OH)_2 \rightarrow Cl^- + 2Fe(OH)_3$. O 6 = 6; H 6 = 6; charge $-1 = -1$.

2. Given basic solution: reduction $BrO_3^- + 3H_2O + 6e^- \rightarrow Br^- + 6OH^-$ and oxidation $Fe(OH)_2 + OH^- \rightarrow Fe(OH)_3 + e^-$. Write the lowest whole-number net equation.

Use Method: how to balance a redox equation in basic solution. Multiply the oxidation by 6: $6Fe(OH)_2 + 6OH^- \rightarrow 6Fe(OH)_3 + 6e^-$. Add and cancel $6e^-$ and $6OH^-$: $BrO_3^- + 3H_2O + 6Fe(OH)_2 \rightarrow Br^- + 6Fe(OH)_3$. O 18 = 18; H 18 = 18; charge $-1 = -1$.

B5 · Worked practice

Balance redox equations in neutral and basic solution | Questions 3-4 of 9

3. 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.

Multiply the reduction by 2 to match 4 electrons. Add and cancel $4e^-$, $4OH^-$ and $2H_2O$: $3ClO^- \rightarrow ClO_3^- + 2Cl^-$. Cl 3 = 3; O 3 = 3; charge $-3 = -3$.

4. Given neutral water: $Br_2(l) + H_2O(l) \rightarrow Br^-(aq) + HOBr(aq) + H^+(aq)$. Check atoms and charge. Give the oxidation number of bromine in Br_2, in Br^- and in HOBr. Say why H^+ may stay in the equation.

Br 2 = 2; H 2 = 2; O 1 = 1; charge $0 = -1 + 0 + 1$. Bromine goes $0 \rightarrow -1$ in Br^- and $0 \rightarrow +1$ in HOBr. The H^+ is a real product, not a counting H^+, so it may stay.

B5 · Worked practice

Balance redox equations in neutral and basic solution | Questions 5-6 of 9

5. Given basic solution: $I_2 \rightarrow IO_3^- + I^-$. The oxidation half is $I_2 + 12OH^- \rightarrow 2IO_3^- + 6H_2O + 10e^-$ and the reduction half is $I_2 + 2e^- \rightarrow 2I^-$. Write the lowest whole-number equation.

Multiply the reduction by 5. Add: $6I_2 + 12OH^- \rightarrow 2IO_3^- + 10I^- + 6H_2O$. Divide by 2: $3I_2 + 6OH^- \rightarrow IO_3^- + 5I^- + 3H_2O$. I 6 = 6; O 6 = 6; H 6 = 6; charge $-6 = -6$.

6. Explain why the half-reaction $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ cannot be used in a basic solution. Then name what a final equation for a basic solution may contain.

It takes in $8H^+$, which basic water does not supply. A final equation for a basic solution may contain OH^- and H_2O and should contain no H^+.

B5 · Worked practice

Balance redox equations in neutral and basic solution | Questions 7-8 of 9

7. Given acidic solution: $BrO_3^- + Fe^{2+} \rightarrow Br^- + Fe^{3+}$. No half-reactions are given. Use oxidation numbers to find what is oxidized and what is reduced. Write both half-reactions. Then write the lowest whole-number net equation and check it against the condition.

Use the B3 method. Br is +5 in BrO_3^- and −1 in Br^-: down, so it is reduced. Fe goes +2 → +3: up, so it is oxidized. Reduction: $BrO_3^- + 6H^+ + 6e^- \rightarrow Br^- + 3H_2O$. Oxidation: $Fe^{2+} \rightarrow Fe^{3+} + e^-$. Multiply the oxidation by 6. Add and cancel $6e^-$. $BrO_3^- + 6H^+ + 6Fe^{2+} \rightarrow Br^- + 3H_2O + 6Fe^{3+}$. Check: Br 1; O 3; H 6; Fe 6. Charge: −1 + 6 + 12 = +17 and −1 + 18 = +17. The solution is acidic, so the H^+ may stay.

8. Given basic solution: $MnO_4^- + Br^- \rightarrow MnO_2 + BrO_3^-$. No half-reactions are given. Use oxidation numbers to find what is oxidized and what is reduced. Balance as in acid. Then change the equation to basic and check it.

Mn goes +7 → +4: reduced. Br goes −1 → +5: oxidized. As in acid: reduction $MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O$ and oxidation $Br^- + 3H_2O \rightarrow BrO_3^- + 6H^+ + 6e^-$. Multiply the reduction by 2. Add and cancel $6e^-$, $6H^+$ and $3H_2O$: $2MnO_4^- + 2H^+ + Br^- \rightarrow 2MnO_2 + H_2O + BrO_3^-$. Change to basic: add $2OH^-$ to each side. $2H^+ + 2OH^-$ make $2H_2O$. Cancel $1H_2O$ from each side. $2MnO_4^- + H_2O + Br^- \rightarrow 2MnO_2 + BrO_3^- + 2OH^-$. Check: Mn 2; Br 1; O 8 + 1 = 9 and 4 + 3 + 2 = 9; H 2. Charge: −3 and −3. There is no H^+.

B5 · Worked practice

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

9. Challenge: Given basic solution: reduction $MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_2 + 4OH^-$ and oxidation $SO_3^{2-} + 2OH^- \rightarrow SO_4^{2-} + H_2O + 2e^-$. Write the lowest whole-number net equation. Then check atoms, charge and the condition.

The lowest common multiple of 3 and 2 is 6. Multiply the reduction by 2 and the oxidation by 3. Add and cancel 6 e^-. Then cancel the H_2O and OH^- that appear on both sides: 4 H_2O and 3 H_2O leave 1 H_2O on the left, and 8 OH^- and 6 OH^- leave 2 OH^- on the right. Net: $2MnO_4^- + 3SO_3^{2-} + H_2O \rightarrow 2MnO_2 + 3SO_4^{2-} + 2OH^-$. Check: 2 Mn, 3 S, 18 O and 2 H on each side. Charge is −8 on each side. Only H_2O and OH^- appear, with no H^+, so it fits a basic solution.