B3 · Balance redox equations in acid

Completed notes

Goals: Balance a half-reaction for atoms and charge. Balance a half-reaction in acid with H_2O, H^+ and e^-. Join two half-reactions into one net ionic equation.

Slide 1 of 11 · A half-reaction must conserve atoms and charge

A half-reaction must keep atoms and **charge** balanced. A reaction cannot create atoms or charge.

In $Al \rightarrow Al^{3+}$, atoms balance (1 Al on each side) but charge does not. The left is 0 and the right is **+3**.

One electron has charge **−1**. Add **3** electrons to the right side: $+3 + 3(-1) = 0$.

Electrons on the product side mean **oxidation**. Electrons carry charge but no **atoms**.

Example. Given $Al(s) \rightarrow Al^{3+}(aq)$. Add electrons to make a balanced oxidation half-reaction. What must stay equal on both sides?

1. Atoms: 1 Al on each side, so atoms balance.

2. Charge: left 0, right +3, so they do not match yet.

3. Add 3 electrons to the right: $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$.

4. Check: right side $+3 + 3(-1) = 0$, which equals the left side.

Try it. Given $Sn^{2+}(aq) \rightarrow Sn^{4+}(aq)$. Add electrons to make a balanced half-reaction. Say whether it is oxidation or reduction.

1. Atoms: 1 Sn on each side.

2. Charge: left +2, right +4. Add 2 electrons to the right: $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$.

3. Check: right side $+4 + 2(-1) = +2$, which equals the left side.

4. Electrons are products, so this is oxidation.

Slide 2 of 11 · Match unequal electron counts before adding

$Al \rightarrow Al^{3+} + 3e^-$ gives away 3 electrons. $Cu^{2+} + 2e^- \rightarrow Cu$ takes **2**. Electrons given must equal electrons **taken**.

The lowest common multiple of 3 and 2 is **6**. So we need **2** Al and **3** Cu^{2+}.

Multiply every term of a half-reaction, including the **electrons**. A coefficient changes how many particles there are, not the charge on each one.

Add the halves and cancel the $6e^-$, because the same particle is on **opposite** sides. Check atoms and total charge: **+6** on each side.

Example. Given the halves $Al \rightarrow Al^{3+} + 3e^-$ and $Cu^{2+} + 2e^- \rightarrow Cu$. Build the lowest whole-number net ionic equation and verify it.

1. Lowest common multiple of 3 and 2 electrons is 6.

2. Multiply the Al half by 2: $2Al \rightarrow 2Al^{3+} + 6e^-$.

3. Multiply the Cu half by 3: $3Cu^{2+} + 6e^- \rightarrow 3Cu$.

4. Add and cancel $6e^-$: $2Al + 3Cu^{2+} \rightarrow 2Al^{3+} + 3Cu$.

5. Atoms: Al 2 = 2; Cu 3 = 3. Charge: 3(+2) = +6 and 2(+3) = +6.

6. Coefficients 2, 3, 2, 3 have no common factor, so this is the lowest whole-number equation.

Try it. Given the halves $Zn \rightarrow Zn^{2+} + 2e^-$ and $Fe^{3+} + 3e^- \rightarrow Fe$. Build the lowest whole-number net ionic equation and verify it.

1. Lowest common multiple of 2 and 3 electrons is 6.

2. Multiply the Zn half by 3: $3Zn \rightarrow 3Zn^{2+} + 6e^-$.

3. Multiply the Fe half by 2: $2Fe^{3+} + 6e^- \rightarrow 2Fe$.

4. Add and cancel $6e^-$: $3Zn + 2Fe^{3+} \rightarrow 3Zn^{2+} + 2Fe$.

5. Atoms: Zn 3 = 3; Fe 2 = 2. Charge: 2(+3) = +6 and 3(+2) = +6.

Slide 3 of 11 · Your turn: magnesium and silver

$Mg \rightarrow Mg^{2+}$ gives **2** electrons, so the electrons are on the **right**.

$Ag^{+} + e^{-} \rightarrow Ag$ takes only 1 electron, so we need **2** Ag^{+} to take 2 electrons: $2Ag^{+} + 2e^{-} \rightarrow 2Ag$.

Scale the halves first. Then add and cancel the **$2e^{-}$**.

Check the net equation: one Mg and two Ag on each side, and a total charge of **+2** on each side.

Example. Given $Mg \rightarrow Mg^{2+}$ and $Ag^{+} \rightarrow Ag$. Write both balanced halves, then the net ionic equation.

1. $Mg \rightarrow Mg^{2+} + 2e^{-}$ (loss, so electrons are products).

2. $Ag^{+} + e^{-} \rightarrow Ag$ takes 1 electron. Multiply by 2: $2Ag^{+} + 2e^{-} \rightarrow 2Ag$.

3. Add and cancel $2e^{-}$: $Mg + 2Ag^{+} \rightarrow Mg^{2+} + 2Ag$.

4. Check: Mg 1 = 1; Ag 2 = 2; charge 2(+1) = +2 and +2.

Try it. Given $Li \rightarrow Li^{+}$ and $Cu^{2+} \rightarrow Cu$. Write both balanced halves, then the net ionic equation.

1. $Li \rightarrow Li^{+} + e^{-}$ (1 electron lost).

2. $Cu^{2+} + 2e^{-} \rightarrow Cu$ (2 electrons gained).

3. Multiply the Li half by 2: $2Li \rightarrow 2Li^{+} + 2e^{-}$.

4. Add and cancel $2e^{-}$: $2Li + Cu^{2+} \rightarrow 2Li^{+} + Cu$.

5. Check: Li 2 = 2; Cu 1 = 1; charge +2 = 2(+1) = +2.

Slide 4 of 11 · Start the complete acidic reaction from its changes

Find oxidation numbers first. In MnO_4^-, four oxygens give −8 and the charge is 1−, so Mn is **+7**.

Mn goes +7 to +2, so it is **reduced**. Fe goes +2 to +3, so it is **oxidized**.

A skeleton is a half-reaction **before balancing**. Use the real species, so write MnO_4^-, not Mn^{7+}.

In acidic water we may add **H_2O** for oxygen and **H^+** for hydrogen. Electrons cannot fix a missing oxygen.

Example. Given acidic water and $MnO_4^- + Fe^{2+} \rightarrow Mn^{2+} + Fe^{3+}$. Find the changes in oxidation number, then write both skeleton halves.

1. MnO_4^-: x + 4(−2) = −1, so Mn is +7. Mn^{2+} is +2. Mn is reduced.

2. Fe^{2+} is +2 and Fe^{3+} is +3. Fe is oxidized.

3. Reduction skeleton: $MnO_4^- \rightarrow Mn^{2+}$.

4. Oxidation skeleton: $Fe^{2+} \rightarrow Fe^{3+}$.

Try it. Given acidic water and $ClO_3^- + Sn^{2+} \rightarrow Cl^- + Sn^{4+}$. Find the changes in oxidation number, then write both skeleton halves.

1. ClO_3^-: x + 3(−2) = −1, so Cl is +5. Cl^- is −1. Cl is reduced.

2. Sn^{2+} is +2 and Sn^{4+} is +4. Sn is oxidized.

3. Reduction skeleton: $ClO_3^- \rightarrow Cl^-$.

4. Oxidation skeleton: $Sn^{2+} \rightarrow Sn^{4+}$.

Slide 5 of 11 · Build and check the permanganate half

Balance atoms other than O and H first. Then balance O with **H_2O**, H with **H^+** and charge with **electrons**.

MnO_4^- → Mn^{2+} has 4 O on the left and none on the right. Add **4**H_2O on the right. They bring 8 H, so add 8H^+ on the left.

Use whole-particle charge. Left: −1 + 8 = **+7**. Right: **+2**. Add 5e^- on the left to make both sides +2.

Check: Mn drops from +7 to +2, which is **5**. That matches the 5 electrons gained.

Example. Given acidic solution: MnO_4^- → Mn^{2+}. Build the balanced half-reaction. Account for every atom before balancing charge.

1. Mn: 1 on each side. O: 4 on the left, 0 on the right.

2. Add 4H_2O on the right: MnO_4^- → Mn^{2+} + 4H_2O. That gives 8 H on the right.

3. Add 8H^+ on the left: MnO_4^- + 8H^+ → Mn^{2+} + 4H_2O.

4. Charge: left −1 + 8 = +7; right +2. Add 5e^- to the left.

5. MnO_4^- + 8H^+ + 5e^- → Mn^{2+} + 4H_2O. Charge: −1 + 8 − 5 = +2 and +2.

6. Check: Mn +7 → +2 is a drop of 5, which matches 5e^-.

Try it. Given acidic solution: ClO_3^- → Cl^-. Build the balanced half-reaction and check atoms, charge and the oxidation-number change.

1. Cl: 1 on each side. O: 3 on the left, 0 on the right.

2. Add 3H_2O on the right: ClO_3^- → Cl^- + 3H_2O. That gives 6 H on the right.

3. Add 6H^+ on the left: ClO_3^- + 6H^+ → Cl^- + 3H_2O.

4. Charge: left −1 + 6 = +5; right −1. Add 6e^- to the left.

5. ClO_3^- + 6H^+ + 6e^- → Cl^- + 3H_2O. Charge: −1 + 6 − 6 = −1 and −1.

6. Check: Cl +5 → −1 is a drop of 6, which matches 6e^-.

Slide 6 of 11 · Show the full sum before cancelling

The iron half is $Fe^{2+} \rightarrow Fe^{3+} + \mathbf{e^-}$. MnO_4^- takes 5 electrons, so use **5** Fe^{2+} for each MnO_4^-.

Multiply every term of the iron half by 5: $5Fe^{2+} \rightarrow 5Fe^{3+} + \mathbf{5e^-}$. Leave the Mn half alone.

Add the halves before cancelling. Cancel only the $5e^-$. H^+ is only on the left and H_2O only on the right, so **neither cancels**.

Net: $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$. Charge is **+17** on each side. Fe^{2+} is the **reducing** agent.

Example. Given the halves $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ and $Fe^{2+} \rightarrow Fe^{3+} + e^-$ in acidic solution. Write the overall net ionic equation and check it.

1. The Mn half takes $5e^-$, so multiply the Fe half by 5: $5Fe^{2+} \rightarrow 5Fe^{3+} + 5e^-$.

2. Add: $MnO_4^- + 8H^+ + 5e^- + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+} + 5e^-$.

3. Cancel $5e^-$: $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$.

4. Atoms: Mn 1, O 4, H 8, Fe 5 on each side.

5. Charge: left $-1 + 8 + 10 = +17$; right $+2 + 15 = +17$.

Try it. Given the halves $ClO_3^- + 6H^+ + 6e^- \rightarrow Cl^- + 3H_2O$ and $Fe^{2+} \rightarrow Fe^{3+} + e^-$ in acidic solution. Write the overall net ionic equation and check it.

1. The Cl half takes $6e^-$, so multiply the Fe half by 6: $6Fe^{2+} \rightarrow 6Fe^{3+} + 6e^-$.

2. Add and cancel $6e^-$: $ClO_3^- + 6H^+ + 6Fe^{2+} \rightarrow Cl^- + 3H_2O + 6Fe^{3+}$.

3. Atoms: Cl 1, O 3, H 6, Fe 6 on each side.

4. Charge: left $-1 + 6 + 12 = +17$; right $-1 + 18 = +17$.

Slide 7 of 11 · Water can be a reactant in an oxidation half

Nitrate has one more O than nitrite, so the left side is short one O. Add **H_2O** on the left.

That water brings 2 H to the left, so put **$2H^+$** on the right. H^+ does not always go on the reactant side.

Before electrons: left charge -1; right $-1 + 2 =$ **$+1$**. Add electrons to the more **positive** side.

So add **$2e^-$** on the right. Electrons on the product side mean oxidation. N goes from +3 to +5, up 2.

Example. Given acidic water: $NO_2^- \rightarrow NO_3^-$. Write the half-reaction and say which earlier decisions change.

1. N: 1 on each side. O: 2 on the left, 3 on the right, so the left is short one O.

2. Add H_2O on the left: $NO_2^- + H_2O \rightarrow NO_3^-$. That puts 2 H on the left, so add $2H^+$ on the right.

3. Charge: left -1; right $-1 + 2 = +1$. Add $2e^-$ to the right.

4. $NO_2^- + H_2O \rightarrow NO_3^- + 2H^+ + 2e^-$. Charge $-1 = -1 + 2 - 2$.

Try it. Given acidic water: $SO_3^{2-} \rightarrow SO_4^{2-}$. Write the half-reaction and check atoms, charge and the oxidation-number change.

1. S: 1 on each side. O: 3 on the left, 4 on the right, so add H_2O on the left.

2. $SO_3^{2-} + H_2O \rightarrow SO_4^{2-}$. That puts 2 H on the left, so add $2H^+$ on the right.

3. Charge: left -2; right $-2 + 2 = 0$. Add $2e^-$ to the right.

4. $SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2H^+ + 2e^-$. Charge $-2 = -2 + 2 - 2$. S goes +4 → +6, up 2.

Slide 8 of 11 · Cancel what is on both sides

The halves take 5 electrons and give 2. The first common total is **10**. Multiply the MnO_4^- half by **2** and the NO_2^- half by 5.

After adding, H^+ is on both sides: 16 on the left and 10 on the right. The net is **6** H^+ on the left.

H_2O is on both sides too: 5 on the left and 8 on the right. The net is **3** H_2O on the right.

Cancel only the **common amount** of the same particle. Net: $2MnO_4^- + 6H^+ + 5NO_2^- \rightarrow 2Mn^{2+} + 3H_2O + 5NO_3^-$, with charge -1 on each side.

Example. Given acidic halves $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ and $NO_2^- + H_2O \rightarrow NO_3^- + 2H^+ + 2e^-$. Write the lowest whole-number net equation.

1. The common electron total is 10. Multiply the Mn half by 2 and the nitrite half by 5.

2. $2MnO_4^- + 16H^+ + 10e^- \rightarrow 2Mn^{2+} + 8H_2O$ and $5NO_2^- + 5H_2O \rightarrow 5NO_3^- + 10H^+ + 10e^-$.

3. Add and cancel $10e^-$. Net H^+: $16 - 10 = 6$ on the left. Net H_2O: $8 - 5 = 3$ on the right.

4. $2MnO_4^- + 6H^+ + 5NO_2^- \rightarrow 2Mn^{2+} + 3H_2O + 5NO_3^-$.

5. Charge: $-2 + 6 - 5 = -1$ and $+4 - 5 = -1$. O 18 = 18; H 6 = 6.

Try it. Given acidic halves $ClO_3^- + 6H^+ + 6e^- \rightarrow Cl^- + 3H_2O$ and $NO_2^- + H_2O \rightarrow NO_3^- + 2H^+ + 2e^-$. Write the lowest whole-number net equation.

1. The Cl half takes $6e^-$ and the N half gives $2e^-$. Multiply the nitrite half by 3.

2. $3NO_2^- + 3H_2O \rightarrow 3NO_3^- + 6H^+ + 6e^-$.

3. Add and cancel $6e^-$. Both $6H^+$ cancel, and both $3H_2O$ cancel.

4. $ClO_3^- + 3NO_2^- \rightarrow Cl^- + 3NO_3^-$.

5. Check: Cl 1; N 3; O 3 + 6 = 9 and 9; charge $-1 - 3 = -4$ and $-1 - 3 = -4$.

Slide 9 of 11 · Method: how to balance a redox equation in acid

Step 1. Find the changes with oxidation numbers. Then write the skeleton **halves**. Show only the particles that change.

Step 2. Balance each half: atoms other than O and H, then O with **H_2O**, then H with **H^+**, then charge with **e^-**.

Step 3. Make the **electrons** equal. Multiply each half by the smallest whole number that works.

Step 4. **Add** the halves. All reactants go on the left and all products on the right.

Step 5. **Cancel** the e^-, H^+ and H_2O that appear on both sides.

Step 6. Check **atoms** and **charge** on both sides. No electrons are left.

Example. Given acidic solution: $MnO_4^- + NO_2^- \rightarrow Mn^{2+} + NO_3^-$. No half-reactions are given. Use the method to write the lowest whole-number net equation.

1. Step 1: Mn goes +7 → +2 (reduced). N goes +3 → +5 (oxidized). Skeleton halves: $MnO_4^- \rightarrow Mn^{2+}$ and $NO_2^- \rightarrow NO_3^-$.

2. Step 2: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ and $NO_2^- + H_2O \rightarrow NO_3^- + 2H^+ + 2e^-$.

3. Step 3: the smallest common number of electrons is 10. Multiply permanganate by 2 and nitrite by 5.

4. Steps 4 and 5: add, then cancel $10e^-$, $10H^+$ and $5H_2O$. This gives $2MnO_4^- + 5NO_2^- + 6H^+ \rightarrow 2Mn^{2+} + 5NO_3^- + 3H_2O$.

5. Step 6: Mn 2 and 2, N 5 and 5, O 18 and 18, H 6 and 6. Charge is −1 on each side.

Try it. Use the method on $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$ and $Fe^{2+} \rightarrow Fe^{3+} + e^-$. Write the lowest whole-number net equation.

1. The halves take $6e^-$ and give $1e^-$. Multiply the iron half by 6.

2. Add: $Cr_2O_7^{2-} + 14H^+ + 6Fe^{2+} \rightarrow 2Cr^{3+} + 7H_2O + 6Fe^{3+}$.

3. Check: charge −2 + 14 + 12 = +24 on the left and +6 + 18 = +24 on the right.

Slide 10 of 11 · Your turn: build the dichromate half

Count chromium first. $Cr_2O_7^{2-}$ has **2** Cr, so write $2Cr^{3+}$ on the right.

Seven O on the left need **7**H_2O on the right. These bring 14 H, so add **14**H^+ on the left.

Charge: left −2 + 14 = +12; right 2(+3) = **+6**. Add **6**e^- on the left to make both sides +6.

Check: each Cr drops from +6 to +3, so two Cr atoms give 2 × 3 = **6** electrons. Atoms: Cr 2, O 7, H 14.

Example. Given acidic water: $Cr_2O_7^{2-} \rightarrow Cr^{3+}$. Write the balanced half-reaction.

1. Cr: 2 on the left, so $Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$.

2. O: 7 on the left, so add $7H_2O$ on the right. That gives 14 H on the right.

3. Add $14H^+$ on the left: $Cr_2O_7^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$.

4. Charge: left −2 + 14 = +12; right 2(+3) = +6. Add $6e^-$ on the left.

5. $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$. Charge: +6 and +6.

6. Check: Cr +6 → +3 is a drop of 3 per atom; 2 × 3 = 6 electrons.

Try it. Given acidic water: $NO_3^- \rightarrow N_2O$. Write the balanced half-reaction. Check atoms, charge and the oxidation-number change.

1. N: 2 on the right, so $2NO_3^- \rightarrow N_2O$.

2. O: 6 on the left and 1 on the right. Add $5H_2O$ on the right. That gives 10 H.

3. Add $10H^+$ on the left: $2NO_3^- + 10H^+ \rightarrow N_2O + 5H_2O$.

4. Charge: left −2 + 10 = +8; right 0. Add $8e^-$ on the left.

5. $2NO_3^- + 10H^+ + 8e^- \rightarrow N_2O + 5H_2O$. Charge: −2 + 10 − 8 = 0 and 0.

6. Check: N +5 → +1 is a drop of 4 per atom; 2 × 4 = 8 electrons.

Slide 11 of 11 · Your turn: the whole method on dichromate and nitrite

When no half-reactions are given, start with **oxidation numbers**. They show what is reduced and what is oxidized.

If the number goes down, the element is **reduced**. If it goes up, it is **oxidized**.

Write a skeleton half for each change. Then balance each half: other atoms, O with **H_2O**, H with **H^+**, charge with **e^-**.

Make the electrons equal, add the halves, **cancel** what is on both sides, and check atoms and charge.

Example. Given acidic solution: $Cr_2O_7^{2-} + NO_2^- \rightarrow Cr^{3+} + NO_3^-$. No half-reactions are given. Find the changes with oxidation numbers. Then write the balanced net equation and check it.

1. Cr: $2x + 7(-2) = -2$, so Cr is +6. It goes +6 → +3: reduced.

2. N is +3 in NO_2^- and +5 in NO_3^-: oxidized.

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

4. Oxidation: $NO_2^- + H_2O \rightarrow NO_3^- + 2H^+ + 2e^-$. Multiply it by 3.

5. Add, then cancel $6e^-$, $6H^+$ and $3H_2O$.

6. $Cr_2O_7^{2-} + 8H^+ + 3NO_2^- \rightarrow 2Cr^{3+} + 4H_2O + 3NO_3^-$.

7. Check: Cr 2, N 3, O 13, H 8 on each side. Charge: +3 and +3.

Try it. Given acidic solution: $MnO_4^- + Sn^{2+} \rightarrow Mn^{2+} + Sn^{4+}$. No half-reactions are given. Find the changes with oxidation numbers. Then write the balanced net equation and check it.

1. Mn: $x + 4(-2) = -1$, so Mn is +7. It goes +7 → +2: reduced.

2. Sn goes +2 → +4: oxidized.

3. Reduction: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$.

4. Oxidation: $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$.

5. The smallest common number of electrons is 10. Multiply the reduction by 2 and the oxidation by 5.

6. $2MnO_4^- + 16H^+ + 5Sn^{2+} \rightarrow 2Mn^{2+} + 8H_2O + 5Sn^{4+}$.

7. Check: Mn 2, O 8, H 16, Sn 5 on each side. Charge: $-2 + 16 + 10 = +24$ and $+4 + 20 = +24$.

B3 · Worked practice

Balance redox equations in acid | Questions 1-2 of 9

1. Balance $Fe^{2+} \rightarrow Fe^{3+}$ as a half-reaction. Add the electrons and check the charge.

$Fe^{2+} \rightarrow Fe^{3+} + e^-$. Charge: +2 on the left, and +3 − 1 = +2 on the right. Electrons are lost, so this is an oxidation.

2. Given $Al \rightarrow Al^{3+}$ and $Ag^+ \rightarrow Ag$. Write the net ionic equation with the lowest whole numbers.

Use Method: how to balance a redox equation in acid. $Al \rightarrow Al^{3+} + 3e^-$ and $3Ag^+ + 3e^- \rightarrow 3Ag$. Net: $Al + 3Ag^+ \rightarrow Al^{3+} + 3Ag$. There is 1 Al and 3 Ag on each side. The charge is +3 on each side.

B3 · Worked practice

Balance redox equations in acid | Questions 3-4 of 9

3. In acidic solution, balance NO_3^- → NO as a half-reaction.

Add $2H_2O$ to the right for oxygen: $NO_3^- \rightarrow NO + 2H_2O$. Add $4H^+$ to the left for hydrogen. The charge is +3 on the left and 0 on the right, so add $3e^-$ to the left. $NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(l)$. Check: N 1, O 3, H 4, and a charge of 0 on each side.

4. Given, in acid: $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$ and $I^- \rightarrow \frac{1}{2}I_2 + e^-$. Write the overall equation with the lowest whole numbers.

Use Method: how to balance a redox equation in acid. Multiply the iodide half by 6. Net: $Cr_2O_7^{2-} + 14H^+ + 6I^- \rightarrow 2Cr^{3+} + 7H_2O + 3I_2$. Charge: $-2 + 14 - 6 = +6$ on the left, and +6 on the right. All the atoms match.

B3 · Worked practice

Balance redox equations in acid | Questions 5-6 of 9

5. A student writes $MnO_4^- + 8H^+ + 7e^- \rightarrow Mn^{2+} + 4H_2O$. Find the mistake and fix it.

The atoms match, but the charge does not. The left is $-1 + 8 - 7 = 0$ and the right is $+2$. Use $5e^-$ on the left. Then both sides are $+2$.

6. Given the balanced half-reaction $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$. Check the atoms and the charge. Is it an oxidation or a reduction? Explain.

Each side has 4 H and 2 O. The charge is -2 on each side. It is a reduction, because electrons are gained. Hydrogen goes from $+1$ in water to 0 in H_2.

B3 · Worked practice

Balance redox equations in acid | Questions 7-8 of 9

7. Given acidic solution: $MnO_2 + Cl^- \rightarrow Mn^{2+} + Cl_2$. 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.

Use Method: how to balance a redox equation in acid. Mn is +4 in MnO_2 and +2 in Mn^{2+}: down, so it is reduced. Cl is −1 in Cl^- and 0 in Cl_2: up, so it is oxidized. Reduction: $MnO_2 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$. Oxidation: $2Cl^- \rightarrow Cl_2 + 2e^-$. The electrons are 2 and 2, so add and cancel $2e^-$. $MnO_2 + 4H^+ + 2Cl^- \rightarrow Mn^{2+} + 2H_2O + Cl_2$. Check: Mn 1; O 2; H 4; Cl 2. Charge: +4 − 2 = +2 and +2.

8. Given acidic solution: $MnO_4^- + H_2S \rightarrow Mn^{2+} + S$. 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.

Use Method: how to balance a redox equation in acid. Mn is +7 in MnO_4^- and +2 in Mn^{2+}: down, so it is reduced. S is −2 in H_2S and 0 in S: up, so it is oxidized. Reduction: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$. Oxidation: $H_2S \rightarrow S + 2H^+ + 2e^-$. The smallest common number of electrons is 10. Multiply the reduction by 2 and the oxidation by 5. Add and cancel $10e^-$. $16H^+$ on the left and $10H^+$ on the right leave $6H^+$ on the left. $2MnO_4^- + 6H^+ + 5H_2S \rightarrow 2Mn^{2+} + 8H_2O + 5S$. Check: Mn 2; O 8; H 6 + 10 = 16 and 16; S 5. Charge: −2 + 6 = +4 and +4.

B3 · Worked practice

Balance redox equations in acid | Questions 9-9 of 9

9. Challenge: Given acidic solution: MnO_4^- changes to Mn^{2+} and SO_2 changes to SO_4^{2-}. Build both half-reactions. Then write the lowest whole-number net ionic equation and check atoms and charge.

Reduction: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$. Oxidation: $SO_2 + 2H_2O \rightarrow SO_4^{2-} + 4H^+ + 2e^-$. The lowest common multiple of 5 and 2 is 10. Multiply the reduction by 2 and the oxidation by 5. Add, cancel 10 e^-, then net the H_2O and H^+ that appear on both sides. Net: $2MnO_4^- + 5SO_2 + 2H_2O \rightarrow 2Mn^{2+} + 5SO_4^{2-} + 4H^+$. Check: 2 Mn, 5 S, 20 O and 4 H on each side. Charge is −2 on each side.