B3 · Balance redox equations in acid

Completed notes | Notes 1/3

Conserve atoms and charge, match electrons, build acidic halves and verify given basic halves.

1. Two checks and two matching halves

Atoms and charge both have to balance

A balanced equation keeps two things equal on both sides: the count of each **element's atom** and the **total signed charge**. Electrons carry charge but add **no atoms**. Equal charge does not have to mean zero charge.

Take $Al \rightarrow Al^{3+} + 3e^-$. The charge is 0 on the left and +3 - 3 = 0 on the right, so it balances. For $Cu^{2+} + 2e^- \rightarrow Cu$, the left is +2 - 2 = 0. Al gives 3 electrons and Cu takes 2. The smallest number both divide into is **six**. So multiply the whole Al half by **two** and the whole Cu half by **three**.

Now add the halves: $2Al + 3Cu^{2+} + 6e^- \rightarrow 2Al^{3+} + 6e^- + 3Cu$. Only identical terms on **opposite sides** cancel, so the six electrons cancel. The net charge is **+6** on each side. Never change a subscript or an ion charge just to make things balance.

On the diploma: check atoms, charge, cancelling and lowest whole-number coefficients, one at a time.

Watch: Why redox balancing checks atoms and charge (Tyler DeWitt, "How to Balance Redox Equations in Acidic Solution", 0:19)

Start from the full reaction you are given

Say the solution is acidic and you are given $MnO_4^- + Fe^{2+} \rightarrow Mn^{2+} + Fe^{3+}$. Mn goes from **+7** in permanganate to +2, which is reduction. Fe goes from +2 to +3, which is **oxidation**.

Write the real species in each skeleton: **$MnO_4^- \rightarrow Mn^{2+}$** and **$Fe^{2+} \rightarrow Fe^{3+}$**. A +7 oxidation number does not mean a free Mn^{7+} ion exists. It is only a counting tool.

Then balance in this order: other atoms first, then O with **H_2O**, then H with **H^+**, then total charge with **electrons**. The solution type and the products are given to you, so do not guess them.

On the diploma: write the species exactly as given. Do not swap in a different ion.

Watch: Give each atom an oxidation number (Tyler DeWitt, "How to Balance Redox Equations in Acidic Solution", 1:04)

Watch: Which atom is oxidized and which is reduced (Tyler DeWitt, "How to Balance Redox Equations in Acidic Solution", 4:15)

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2. Build one acidic half, then combine

Build the permanganate half one step at a time

Start with MnO_4^- → Mn^{2+}. Mn matches, but the left has four O and the right has none. So **four O are missing on the right**. Add **$4H_2O$** to the right. Four waters also bring in **eight H** on the right.

Add **$8H^+$** to the left to match those hydrogens. Now Mn is 1, O is 4 and H is 8 on both sides. Check the charge before adding electrons: the left is -1 + 8 = **+7**, and the right is **+2**.

Add **$5e^-$ on the left**: +7 - 5 = +2. The finished half is $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$. Making the left side zero would **fail to match the right charge**, and the charges must match.

As a second check, one Mn goes from +7 to +2. That is **five assigned electrons gained**. It supports the charge result, but it does not replace counting atoms.

On the diploma: write the electrons last, after atoms and charge are checked.

Watch: Balance oxygen with H_2O (Tyler DeWitt, “How to Balance Redox Equations in Acidic Solution”, 6:42)

Watch: Balance hydrogen with H^+ (Tyler DeWitt, “How to Balance Redox Equations in Acidic Solution”, 7:14)

Watch: Balance charge with electrons (Tyler DeWitt, “How to Balance Redox Equations in Acidic Solution”, 8:16)

Combine the two halves

$Fe^{2+} \rightarrow Fe^{3+} + e^-$ releases one electron, but the Mn half takes five. So multiply the iron half by **five**. The manganese half stays as it is.

Write every term before cancelling. The left has $MnO_4^- + 8H^+ + 5e^- + 5Fe^{2+}$, and the right has $Mn^{2+} + 4H_2O + 5Fe^{3+} + 5e^-$. Cancel **five electrons from each side**.

The net equation is $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$. The charge is **+17** on each side. H^+ and water stay in the equation because they have **no identical opposite-side terms**.

On the diploma: only cancel identical terms that sit on opposite sides.

Watch: Make the electrons equal (Tyler DeWitt, “How to Balance Redox Equations in Acidic Solution”, 10:46)

Watch: Add the halves and cancel (Tyler DeWitt, “How to Balance Redox Equations in Acidic Solution”, 11:51)

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

Before e^-: left +7; right +2

After e^-: left +2; right +2

Equal signed charge is required; neutral totals are not required.

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3. Changed sides, new examples and given basic halves

Water does not always go on the right

Try NO_2^- → NO_3^- in acid. The right has one more O, so add **H_2O on the left**. That puts two H on the left, so add **$2H^+$ on the right**. The charge is now -1 on the left and +1 on the right.

Add **$2e^-$ on the right**: $NO_2^- + H_2O \rightarrow NO_3^- + 2H^+ + 2e^-$. N goes from +3 to +5, which is **oxidation**. H^+ can appear as a product even in acid. Put it on whichever side the atom count needs.

On the diploma: do not assume H_2O is always a product. Check which side is missing oxygen.

Combine, cancel what repeats, then simplify

Combine $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ with $NO_2^- + H_2O \rightarrow NO_3^- + 2H^+ + 2e^-$. The smallest shared electron total is **ten**. So multiply the first half by **two** and the second by **five**.

Add the halves and cancel $10e^-$. The left has $16H^+$ and the right has $10H^+$, so **$6H^+$** is left over on the left. The left has $5H_2O$ and the right has $8H_2O$, so **$3H_2O$** is left over on the right.

The net equation is $2MnO_4^- + 6H^+ + 5NO_2^- \rightarrow 2Mn^{2+} + 3H_2O + 5NO_3^-$. The charge is **-1** on each side.

On the diploma: after cancelling electrons, also cancel repeated H^+ and H_2O.

Watch: Cancel what appears on both sides (Tyler DeWitt, “How to Balance Redox Equations in Acidic Solution”, 12:50)

Watch: Divide by a common factor (Tyler DeWitt, “Balance Redox Equations in Acid Example 2 (Advanced)”, 12:24)

Dichromate: a changed attempt

In acid, take $Cr_2O_7^{2-} \rightarrow Cr^{3+}$. First make **$2Cr^{3+}$** so the Cr atoms match. The left has seven O, so add **$7H_2O$** on the right. That brings 14 H, so add **$14H^+$** on the left. The charges are +12 on the left and +6 on the right, so add **$6e^-$ left**.

Final check: Cr is 2, O is 7, H is 14 and the charge is **+6 on both sides**. Each of the two Cr atoms drops by three. That accounts for **six electrons**.

Watch: Balance the chromium (Tyler DeWitt, “How to Balance Redox Equations in Acidic Solution Example 1”, 5:40)

Watch: Add 7 H_2O (Tyler DeWitt, “How to Balance Redox Equations in Acidic Solution Example 1”, 6:11)

Watch: Add 14 H^+ (Tyler DeWitt, “How to Balance Redox Equations in Acidic Solution Example 1”, 6:43)

Watch: Add 6 electrons (Tyler DeWitt, “How to Balance Redox Equations in Acidic Solution Example 1”, 8:18)

Combine given basic halves

A neutral solution is built the same way as an acidic one, with H_2O and H^+. Basic half-reactions are **supplied** to you, not made up by you.

Say you are given $OCl^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$ and $2I^- \rightarrow I_2 + 2e^-$. Both halves already use **two electrons**, so just cancel them. You get $OCl^- + H_2O + 2I^- \rightarrow Cl^- + 2OH^- + I_2$.

The charge is **-3** on each side. Use the basic form you were given. Do not turn it into an acidic one. Also, balancing does not tell you **whether the reaction goes in this direction**. That is the next lesson.

On the diploma: if basic halves are given, only match electrons and cancel.

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

B3 · Worked practice

Balance redox equations in acid | 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.

B3-P01 · Easy · Write the balanced oxidation half-reaction for $Mg(s) \rightarrow Mg^{2+}(aq)$. Add the electrons and check the charge.

1. $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^{-}$. The two electrons balance the +2 on the product.

2. There is one Mg atom on each side. The charge is 0 on the left and +2 - 2 = 0 on the right. The electrons are on the product side because Mg loses them.

B3-P02 · Easy · Write the balanced reduction half-reaction for $Fe^{3+}(aq) \rightarrow Fe^{2+}(aq)$. Show the charge on each side.

1. $Fe^{3+}(aq) + e^{-} \rightarrow Fe^{2+}(aq)$. One electron on the left gives +3 - 1 = +2.

2. There is one Fe atom on each side, and the charge is +2 on each side. The product does not have to be neutral.

B3 · Worked practice

Balance redox equations in acid | Questions 3-4 of 12

B3-P03 · Easy · Given $Al \rightarrow Al^{3+} + 3e^-$ and $Ag^+ + e^- \rightarrow Ag$. What is the lowest number of electrons that both halves can move? What do you multiply each half by?

1. The lowest number is 3. Multiply the aluminium half by 1 and the silver half by 3.

2. Then three silver ions take the three electrons that one Al atom gives. Do not multiply the charge on Ag^+.

B3-P04 · Easy · Check $Zn + 2Ag^+ \rightarrow Zn^{2+} + 2Ag$. Count each element and add the charge on each side. Is the charge balanced even though it is not zero?

1. Each side has one Zn and two Ag atoms. The charge is $2(+1) = +2$ on the left and $+2$ on the right.

2. Both checks pass. In a net ionic equation the two sides must be equal. They do not have to be zero.

B3 · Worked practice

Balance redox equations in acid | Questions 5-6 of 12

B3-P05 · Medium · Join $Mg \rightarrow Mg^{2+} + 2e^-$ and $Ag^+ + e^- \rightarrow Ag$ into one net ionic equation. Show the half you multiplied and what you cancelled.

1. Multiply every part of the silver half by two: $2Ag^+ + 2e^- \rightarrow 2Ag$. Leave the magnesium half as it is.

2. Add the halves and cancel the $2e^-$: $Mg + 2Ag^+ \rightarrow Mg^{2+} + 2Ag$. The atoms match, and the charge is +2 on each side.

B3-P06 · Medium · In acidic solution, balance $MnO_4^- \rightarrow Mn^{2+}$ with H_2O, H^+ and e^-. Explain why you add them in that order.

1. Mn is already balanced. Four O on the left need $4H_2O$ on the right. The water brings 8 H, so add $8H^+$ on the left.

2. Before the electrons, the charge is -1 + 8 = +7 on the left and +2 on the right. Add $5e^-$ to the left to bring +7 down to +2. Charge comes last, because electrons have no atoms.

3. $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$. Check: Mn 1, O 4, H 8, and a charge of +2 on each side.

B3 · Worked practice

Balance redox equations in acid | Questions 7-8 of 12

B3-P07 · Medium · In acidic solution, balance $Cr_2O_7^{2-} \rightarrow Cr^{3+}$ as a reduction half-reaction. Check the atoms and the charge.

1. Balance Cr first with $2Cr^{3+}$. Add $7H_2O$ on the right for the seven O. Then add $14H^+$ on the left for the fourteen H.

2. Before the electrons, the charge is -2 + 14 = +12 on the left and 2(+3) = +6 on the right. Add $6e^-$ to the left.

3. $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$. Check: Cr 2, O 7, H 14, and a charge of +6 on each side.

B3-P08 · Medium · A student writes $MnO_4^- + 8H^+ + 7e^- \rightarrow Mn^{2+} + 4H_2O$. Which check fails? Fix only the number of electrons. Explain why the student may have chosen seven.

1. The atoms are balanced. But the charge is -1 + 8 - 7 = 0 on the left and +2 on the right. The charge check fails.

2. Use $5e^-$. Then the left is +2. Seven electrons make the left side zero, but the left must match the right. It does not have to be zero.

B3 · Worked practice

Balance redox equations in acid | Questions 9-10 of 12

B3-P09 · Challenging · In acidic solution, join $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$ with $I^- \rightarrow 1/2\ I_2 + e^-$. Write the net equation with the lowest whole numbers. Check the atoms and the charge.

1. Six iodide ions each give one electron: $6I^- \rightarrow 3I_2 + 6e^-$. Add this to the dichromate half and cancel the six electrons.

2. $Cr_2O_7^{2-} + 14H^+ + 6I^- \rightarrow 2Cr^{3+} + 7H_2O + 3I_2$.

3. Cr 2, O 7, H 14 and I 6 match. The charge is -2 + 14 - 6 = +6 on the left and 2(+3) = +6 on the right.

B3-P10 · Challenging · A question gives three reactants: MnO_4^-, Fe^{2+} and H^+. The products are Mn^{2+}, Fe^{3+} and H_2O. Which two reactants change oxidation number? What does the third one do?

1. Check each atom. Mn goes from +7 to +2 and Fe goes from +2 to +3. Hydrogen is +1 on both sides and oxygen is -2 on both sides.

2. Only MnO_4^- and Fe^{2+} lose or gain electrons, so each one gets a half-reaction. H^+ only helps with balancing. You add H^+ to balance hydrogen and H_2O to balance oxygen. On the diploma: find the two particles that change first.

3. Net: $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$. If one particle has two atoms that both change, such as Cu_2S, keep them in the same half-reaction and count all their electrons together.

B3 · Worked practice

Balance redox equations in acid | Questions 11-12 of 12

B3-P11 · Challenging · Given this half-reaction in basic solution: $OCl^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$. Check the atoms and the charge. Find the oxidation number of chlorine before and after. Is it an oxidation or a reduction?

1. Each side has Cl 1, O 2 and H 2. The charge is -1 - 2 = -3 on the left and -1 + 2(-1) = -3 on the right.

2. In OCl^-, x - 2 = -1, so Cl = +1. In Cl^- it is -1. The number goes down from +1 to -1, and the electrons are on the left. Both show a reduction.

B3-P12 · Challenging · Given two half-reactions for hydrogen peroxide: $H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$ and $H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$. Join them and cancel what is on both sides. Explain why the net reaction is a disproportionation.

1. Both halves move 2 electrons, so add them as they are. Then cancel the $2H^+$ and the $2e^-$ that are on both sides.

2. The net equation is $2H_2O_2 \rightarrow O_2 + 2H_2O$. There are 4 H and 4 O on each side, and the charge is 0 on each side.

3. O starts at -1 in one reactant. It ends at 0 in O_2 and at -2 in H_2O. So one reactant is both oxidized and reduced. That is disproportionation.