R1 · Review of B1 to B5 (optional)

Completed notes

Goals: Use each method from B1 to B5 on a new question. Balance a redox equation in acidic and in basic solution, and check it. Choose the method a question needs.

Slide 1 of 7 · B1 review: oxidation, reduction and the two agents

A particle that loses electrons is **oxidized**. Its charge goes up.

A particle that gains electrons is **reduced**. Its charge goes down.

The **reducing agent** is what is oxidized. The **oxidizing agent** is what is reduced.

An oxidation half shows electrons on the **product** side. A reduction half shows them on the **reactant** side.

Example. Given: $Zn(s) + Pb^{2+}(aq) \rightarrow Zn^{2+}(aq) + Pb(s)$. Name what is oxidized, what is reduced and both agents. Write the two half-reactions.

1. Zn goes 0 → +2. It loses 2 electrons, so Zn is oxidized.

2. Pb^{2+} goes +2 → 0. It gains 2 electrons, so Pb^{2+} is reduced.

3. Reducing agent: Zn. Oxidizing agent: Pb^{2+}.

4. Oxidation: $Zn \rightarrow Zn^{2+} + 2e^{-}$. Reduction: $Pb^{2+} + 2e^{-} \rightarrow Pb$.

5. Check: 2 electrons lost and 2 gained.

Try it. Given: $Mg(s) + Ni^{2+}(aq) \rightarrow Mg^{2+}(aq) + Ni(s)$. Name what is oxidized, what is reduced and both agents. Write the two half-reactions.

1. Mg goes 0 → +2. It loses 2 electrons, so Mg is oxidized.

2. Ni^{2+} goes +2 → 0. It gains 2 electrons, so Ni^{2+} is reduced.

3. Reducing agent: Mg. Oxidizing agent: Ni^{2+}.

4. Oxidation: $Mg \rightarrow Mg^{2+} + 2e^{-}$. Reduction: $Ni^{2+} + 2e^{-} \rightarrow Ni$.

Slide 2 of 7 · B2 review: find an oxidation number

The oxidation numbers in a formula add to its **overall charge**, not always to zero.

Oxygen is usually **−2** and hydrogen is usually **+1**.

In a **peroxide**, each oxygen is −1. In a metal **hydride**, hydrogen is −1.

A Group 1 metal is **+1** in a compound. Set it first.

Example. Find the oxidation number of sulfur in $S_2O_7^{2-}$. Then find the oxidation number of oxygen in Na_2O_2.

1. $S_2O_7^{2-}$ has an overall charge of −2. Oxygen is usually −2. Let x be sulfur.

2. $2x + 7(-2) = -2$, so $2x = +12$ and $x = +6$.

3. Check: $2(+6) - 14 = -2$.

4. Na_2O_2: sodium is +1. $2(+1) + 2x = 0$, so $x = -1$.

5. Oxygen is −1, so Na_2O_2 is a peroxide.

Try it. Find the oxidation number of phosphorus in $P_2O_7^{4-}$. Then find the oxidation number of hydrogen in KH.

1. $P_2O_7^{4-}$ has an overall charge of −4. Oxygen is usually −2. Let x be phosphorus.

2. $2x + 7(-2) = -4$, so $2x = +10$ and $x = +5$.

3. Check: $2(+5) - 14 = -4$.

4. KH: potassium is a Group 1 metal, so it is +1. $+1 + x = 0$, so $x = -1$.

5. Hydrogen is −1, so KH is a metal hydride.

Slide 3 of 7 · B2 review: what was oxidized and what was reduced

Number every atom on both sides. Then **compare** each element before and after.

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

The oxidizing agent is the whole reactant that is **reduced**. The reducing agent is the whole reactant that is oxidized.

When one reactant is both oxidized and reduced, the reaction is a **disproportionation**.

Example. Given: $SnO_2 + 2C \rightarrow Sn + 2CO$. Name what is oxidized, what is reduced and both agents. Count the electrons lost and gained.

1. Sn is +4 in SnO_2 and 0 in Sn. Down, so SnO_2 is reduced. It is the oxidizing agent.

2. C is 0 in C and +2 in CO. Up, so C is oxidized. It is the reducing agent.

3. O stays −2.

4. One Sn gains 4 electrons. Two C atoms lose 2 each, so 4 are lost.

Try it. Given: $Fe_2O_3 + 2Al \rightarrow Al_2O_3 + 2Fe$. Name what is oxidized, what is reduced and both agents. Count the electrons lost and gained.

1. Fe is +3 in Fe_2O_3 and 0 in Fe. Down, so Fe_2O_3 is reduced. It is the oxidizing agent.

2. Al is 0 in Al and +3 in Al_2O_3. Up, so Al is oxidized. It is the reducing agent.

3. O stays −2.

4. Two Fe gain 3 electrons each, so 6 are gained. Two Al lose 3 each, so 6 are lost.

Slide 4 of 7 · B3 review: balance a redox equation in acid

Find the changes with oxidation numbers. Then balance each half: other atoms, O with **H_2O**, H with **H^+**, charge with **e^-**.

Make the electrons **equal**. Multiply every term in a half by the same whole number.

Add the halves. Then **cancel** e^-, H^+ and H_2O that appear on both sides.

Check atoms and **charge** on both sides. No electrons may be left.

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

1. Cr goes +6 → +3: reduced. N goes +4 → +5: oxidized.

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

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

4. Electrons are 6 and 1, so multiply the oxidation by 6: $6NO_2 + 6H_2O \rightarrow 6NO_3^- + 12H^+ + 6e^-$.

5. Add, then cancel $6e^-$, $12H^+$ and $6H_2O$.

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

7. Check: Cr 2, N 6, O 19, H 2 on each side. Charge: 0 and 0.

Try it. Given acidic solution: $MnO_2 + NO_2 \rightarrow Mn^{2+} + NO_3^-$. No half-reactions are given. Write the balanced net equation and check it.

1. Mn goes +4 → +2: reduced. N goes +4 → +5: oxidized.

2. Reduction: $MnO_2 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$.

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

4. Electrons are 2 and 1, so multiply the oxidation by 2: $2NO_2 + 2H_2O \rightarrow 2NO_3^- + 4H^+ + 2e^-$.

5. Add, then cancel $2e^-$, $4H^+$ and $2H_2O$. Here all the H^+ and all the H_2O cancel.

6. $MnO_2 + 2NO_2 \rightarrow Mn^{2+} + 2NO_3^-$.

7. Check: Mn 1, N 2, O 6 on each side. Charge: 0 and $+2 - 2 = 0$.

Slide 5 of 7 · B5 review: neutral and basic solutions

First read the **condition**: acidic, neutral or basic.

An acidic equation may contain H_2O and $\mathbf{H^+}$. A basic equation may contain H_2O and $\mathbf{OH^-}$ and no H^+.

In neutral water, H^+ or OH^- may stay only if the reaction **really makes** it. A counting H^+ must cancel.

An equation can balance and still be wrong if it does not fit the **condition**.

Example. Given basic solution: reduction $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$ and oxidation $Al + 4OH^- \rightarrow Al(OH)_4^- + 3e^-$. Write the lowest whole-number net equation and check it.

1. The smallest number that 4 and 3 both reach is 12. Multiply the reduction by 3 and the oxidation by 4.

2. $3O_2 + 6H_2O + 12e^- \rightarrow 12OH^-$ and $4Al + 16OH^- \rightarrow 4Al(OH)_4^- + 12e^-$.

3. Add, then cancel $12e^-$ and $12OH^-$.

4. $4Al + 3O_2 + 6H_2O + 4OH^- \rightarrow 4Al(OH)_4^-$.

5. Check: Al 4, O 16, H 16 on each side. Charge: −4 and −4. There is no H^+.

Try it. Given basic solution: reduction $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$ and oxidation $Zn + 2OH^- \rightarrow Zn(OH)_2 + 2e^-$. Write the lowest whole-number net equation and check it.

1. The reduction takes 4 electrons and the oxidation gives 2. Multiply the oxidation by 2.

2. $2Zn + 4OH^- \rightarrow 2Zn(OH)_2 + 4e^-$.

3. Add, then cancel $4e^-$ and $4OH^-$.

4. $2Zn + O_2 + 2H_2O \rightarrow 2Zn(OH)_2$.

5. Check: Zn 2, O 4, H 4 on each side. Charge: 0 and 0. There is no H^+.

Slide 6 of 7 · B4 review: rank from tests and predict with the table

A metal that gives electrons to another metal’s ion is the stronger **reducing agent**.

The ions run the other way: the weakest reducing agent has the strongest **oxidizing agent** ion.

E°reaction = E°(oxidizing agent) − E°(reducing agent). Use only what is in the **beaker**.

A positive E°reaction means the reaction is **spontaneous**. Do not multiply E° by coefficients.

Example. Given a strip of iron in $Cu^{2+}(aq)$ under standard conditions. From the Data Booklet table, page 7: E°(Cu^{2+}/Cu) = +0.34 V and E°(Fe^{2+}/Fe) = −0.45 V. Find E°reaction and say whether the reaction is spontaneous.

1. The beaker has Fe(s) and $Cu^{2+}(aq)$.

2. Cu^{2+} is the oxidizing agent. Fe is the reducing agent. In the table, Cu^{2+} is above Fe.

3. Halves: $Cu^{2+} + 2e^- \rightarrow Cu$ and $Fe \rightarrow Fe^{2+} + 2e^-$.

4. E°reaction = +0.34 − (−0.45) = +0.79 V.

5. Positive, so $Fe + Cu^{2+} \rightarrow Fe^{2+} + Cu$ is spontaneous.

Try it. Given a strip of aluminium in $Zn^{2+}(aq)$ under standard conditions. From the Data Booklet table, page 7: E°(Zn^{2+}/Zn) = −0.76 V and E°(Al^{3+}/Al) = −1.66 V. Find E°reaction, write the balanced equation and say whether the reaction is spontaneous.

1. The beaker has Al(s) and $Zn^{2+}(aq)$.

2. Zn^{2+} is the oxidizing agent. Al is the reducing agent. In the table, Zn^{2+} is above Al.

3. E°reaction = −0.76 − (−1.66) = +0.90 V.

4. Halves: $Zn^{2+} + 2e^- \rightarrow Zn$ and $Al \rightarrow Al^{3+} + 3e^-$. Six electrons: $2Al + 3Zn^{2+} \rightarrow 2Al^{3+} + 3Zn$.

5. Positive, so the reaction is spontaneous. E° is not multiplied by 2 or by 3.

Slide 7 of 7 · Which method does this question need?

A question about agents, or who lost and gained electrons, uses the method from **B1**.

A question about one atom in a formula, or what was oxidized, uses oxidation numbers from **B2**.

A skeleton equation in acid uses **B3**. Halves given for a basic solution use **B5**.

A question that asks if a reaction is spontaneous uses the table and E°reaction from **B4**.

Example. Name the lesson whose method each question needs. Do not solve. (a) Name the agents in $Mg + Cu^{2+} \rightarrow Mg^{2+} + Cu$. (b) Find the oxidation number of Cl in ClO_2^-. (c) Is $2H_2 + O_2 \rightarrow 2H_2O$ a redox reaction? (d) Does tin react with Ag^+?

1. (a) B1: name the electron transfer.

2. (b) B2: find an oxidation number.

3. (c) B2: number every atom and compare both sides.

4. (d) B4: find E°reaction from the table and read its sign.

Try it. Name the lesson whose method each question needs. Do not solve. (a) Find the oxidation number of N in NH_3. (b) In acidic solution, balance the change of MnO_4^- to Mn^{2+} and Cl^- to Cl_2. (c) Is $Ag + Cu^{2+}$ a spontaneous reaction? (d) Combine two halves given for a basic solution.

1. (a) B2: find an oxidation number.

2. (b) B3: balance a redox equation in acid.

3. (c) B4: find E°reaction from the table and read its sign.

4. (d) B5: keep OH^- and H_2O, match the electrons, add and cancel.

R1 · Worked practice

Review of B1 to B5 (optional) | Questions 1-2 of 15

1. B1. Given: $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$. Is chlorine oxidized or reduced? Explain using the change in charge.

Reduced. The electrons are on the reactant side, so chlorine gains them. Each chlorine atom goes from 0 to −1. Gaining negative electrons makes the charge go down.

2. B1. Given: $2Al(s) + 3Pb^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Pb(s)$. Name what is oxidized and what is reduced. Name both agents. Then write the two half-reactions.

Use Method: how to name an electron transfer (B1). Al goes 0 → +3, so each Al loses 3 electrons. Al is oxidized and is the reducing agent. Pb^{2+} goes +2 → 0, so each Pb^{2+} gains 2 electrons. Pb^{2+} is reduced and is the oxidizing agent. Oxidation: $Al \rightarrow Al^{3+} + 3e^-$. Reduction: $Pb^{2+} + 2e^- \rightarrow Pb$. Check: 2 Al lose 6 electrons and 3 Pb^{2+} gain 6.

R1 · Worked practice

Review of B1 to B5 (optional) | Questions 3-4 of 15

3. B2. Find the oxidation number of iodine in IO_4^-. Then find the oxidation number of oxygen in Li_2O_2 and say which usual value changed.

Use Method: how to find an oxidation number (B2). IO_4^-: oxygen is usually −2. x + 4(−2) = −1, so x = +7. Check: +7 − 8 = −1. Li_2O_2: lithium is a Group 1 metal, so it is +1. 2(+1) + 2x = 0, so x = −1. Oxygen is −1, not its usual −2. Li_2O_2 is a peroxide.

4. B2. Given: $WO_3 + 3H_2 \rightarrow W + 3H_2O$. Use oxidation numbers to name what is oxidized, what is reduced and both agents. Then count the electrons lost and gained.

Use Method: how to find what was oxidized and reduced (B2). W is +6 in WO_3 and 0 in W. The number goes down, so WO_3 is reduced and is the oxidizing agent. H is 0 in H_2 and +1 in H_2O. The number goes up, so H_2 is oxidized and is the reducing agent. O stays −2. One W gains 6 electrons. Six H atoms lose 1 each, so 6 are lost. They match.

R1 · Worked practice

Review of B1 to B5 (optional) | Questions 5-6 of 15

5. B2. Given: $3MnO_4^{2-} + 4H^+ \rightarrow 2MnO_4^- + MnO_2 + 2H_2O$. Find the oxidation number of manganese in MnO_4^{2-}, in MnO_4^- and in MnO_2. Is this disproportionation? Explain.

MnO_4^{2-}: $x + 4(-2) = -2$, so Mn is +6. MnO_4^-: $x + 4(-2) = -1$, so Mn is +7. MnO_2: $x + 2(-2) = 0$, so Mn is +4. Some manganese goes up from +6 to +7, so it is oxidized. Some goes down from +6 to +4, so it is reduced. One reactant is both oxidized and reduced, so this is disproportionation. Two Mn lose 1 electron each and one Mn gains 2.

6. B3. In acidic solution, IO_3^- changes to I_2. Write the balanced half-reaction. Then say whether it is an oxidation or a reduction.

Use the B3 order. Other atoms: $2IO_3^- \rightarrow I_2$. O with H_2O: $2IO_3^- \rightarrow I_2 + 6H_2O$. H with H^+: $2IO_3^- + 12H^+ \rightarrow I_2 + 6H_2O$. Charge: the left is $-2 + 12 = +10$ and the right is 0, so add $10e^-$ to the left. $2IO_3^- + 12H^+ + 10e^- \rightarrow I_2 + 6H_2O$. The electrons are on the reactant side, so it is a reduction.

R1 · Worked practice

Review of B1 to B5 (optional) | Questions 7-8 of 15

7. B3. Given acidic solution: reduction $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$ and oxidation $Fe^{2+} \rightarrow Fe^{3+} + e^-$. Write the lowest whole-number net equation and check it.

Use Method: how to balance a redox equation in acid (B3). The reduction takes 4 electrons and the oxidation gives 1. Multiply the oxidation by 4: $4Fe^{2+} \rightarrow 4Fe^{3+} + 4e^-$. Add and cancel $4e^-$: $O_2 + 4H^+ + 4Fe^{2+} \rightarrow 2H_2O + 4Fe^{3+}$. Check: O 2 = 2; H 4 = 4; Fe 4 = 4. Charge: +4 + 8 = +12 and +12.

8. B3. A student adds two acidic halves and stops here: $2NO_3^- + 8H^+ + 3H_2S \rightarrow 2NO + 4H_2O + 3S + 6H^+$. The electrons are already cancelled. Finish the equation and check it.

H^+ is on both sides. $8H^+$ on the left and $6H^+$ on the right leave $2H^+$ on the left. $2NO_3^- + 2H^+ + 3H_2S \rightarrow 2NO + 4H_2O + 3S$. Check: N 2 = 2; O 6 = 2 + 4; H 2 + 6 = 8 and 8; S 3 = 3. Charge: −2 + 2 = 0 and 0.

R1 · Worked practice

Review of B1 to B5 (optional) | Questions 9-10 of 15

9. B5. Given basic solution: reduction $MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_2 + 4OH^-$ and oxidation $Zn + 2OH^- \rightarrow Zn(OH)_2 + 2e^-$. Write the lowest whole-number net equation and check it.

Use Method: how to balance a redox equation in basic solution (B5). The smallest number that 3 and 2 both reach is 6. Multiply the reduction by 2 and the oxidation by 3. $2MnO_4^- + 4H_2O + 6e^- \rightarrow 2MnO_2 + 8OH^-$ and $3Zn + 6OH^- \rightarrow 3Zn(OH)_2 + 6e^-$. Add and cancel $6e^-$ and $6OH^-$: $2MnO_4^- + 4H_2O + 3Zn \rightarrow 2MnO_2 + 3Zn(OH)_2 + 2OH^-$. Check: Mn 2; Zn 3; O 8 + 4 = 12 and 4 + 6 + 2 = 12; H 8 and 6 + 2 = 8. Charge: −2 on each side. There is no H^+.

10. B5. A question says the solution is basic. A student answers $Zn + 2MnO_2 + 2H^+ \rightarrow Zn^{2+} + Mn_2O_3 + H_2O$. The atoms and the charge balance. Is this an acceptable final answer? Explain.

No. The equation has H^+ in it. A final equation for a basic solution may have H_2O and OH^- and must have no H^+. The equation balances, but it does not fit the condition.

R1 · Worked practice

Review of B1 to B5 (optional) | Questions 11-12 of 15

11. B4. Given three tests under the same conditions. Magnesium in Fe^{2+} solution gets a coating of iron. Iron in Ni^{2+} solution gets a coating of nickel. Nickel in Fe^{2+} solution: nothing happens. Rank Mg, Fe and Ni as reducing agents. Then rank Mg^{2+}, Fe^{2+} and Ni^{2+} as oxidizing agents.

Magnesium gave electrons to Fe^{2+}, so Mg is a stronger reducing agent than Fe. Iron gave electrons to Ni^{2+}, so Fe is a stronger reducing agent than Ni. Nickel did nothing in Fe^{2+}, which fits. Reducing agents, strongest first: Mg > Fe > Ni. The ions run the other way. Oxidizing agents, strongest first: Ni^{2+} > Fe^{2+} > Mg^{2+}.

12. B4. Given a strip of lead in $Ni^{2+}(aq)$ under standard conditions. From the Data Booklet table, page 7: $E°(Pb^{2+}/Pb) = -0.13$ V and $E°(Ni^{2+}/Ni) = -0.26$ V. Find $E°_{reaction}$. Then say whether $Pb + Ni^{2+} \rightarrow Pb^{2+} + Ni$ is spontaneous.

Use Method: how to predict a redox reaction with the table (B4). The beaker has Pb(s) and $Ni^{2+}(aq)$. Ni^{2+} is the oxidizing agent and Pb is the reducing agent. $E°_{reaction}$ = $E°$(oxidizing agent) − $E°$(reducing agent) = $-0.26 - (-0.13) = -0.13$ V. The sign is negative, so the reaction is not spontaneous. Nothing happens.

R1 · Worked practice

Review of B1 to B5 (optional) | Questions 13-14 of 15

13. B3. Given acidic solution: $Cr_2O_7^{2-} + H_2O_2 \rightarrow Cr^{3+} + O_2$. 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.

Cr goes +6 → +3: reduced. O is −1 in H_2O_2 (a peroxide) and 0 in O_2: up, so it is oxidized. Reduction: $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$. Oxidation: $H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$. Multiply the oxidation by 3. Add and cancel $6e^-$ and $6H^+$. $Cr_2O_7^{2-} + 8H^+ + 3H_2O_2 \rightarrow 2Cr^{3+} + 7H_2O + 3O_2$. Check: Cr 2; O 7 + 6 = 13 and 7 + 6 = 13; H 8 + 6 = 14 and 14. Charge: −2 + 8 = +6 and +6.

14. B5. Given basic solution: $MnO_4^- + S^{2-} \rightarrow MnO_2 + S$. 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. S goes −2 → 0: oxidized. As in acid: reduction $MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O$ and oxidation $S^{2-} \rightarrow S + 2e^-$. Multiply the reduction by 2 and the oxidation by 3. Add and cancel $6e^-$: $2MnO_4^- + 8H^+ + 3S^{2-} \rightarrow 2MnO_2 + 4H_2O + 3S$. Change to basic: add $8OH^-$ to each side. $8H^+ + 8OH^-$ make $8H_2O$. Cancel $4H_2O$ from each side. $2MnO_4^- + 4H_2O + 3S^{2-} \rightarrow 2MnO_2 + 3S + 8OH^-$. Check: Mn 2; S 3; O 8 + 4 = 12 and 4 + 8 = 12; H 8. Charge: −2 − 6 = −8 and −8. There is no H^+.

R1 · Worked practice

Review of B1 to B5 (optional) | Questions 15-15 of 15

15. Challenge: Given acidic solution. A strip of nickel is placed in a solution of MnO_4^-. From the Data Booklet table, page 7: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ has E° = +1.51 V, and $Ni^{2+} + 2e^- \rightarrow Ni$ has E° = −0.26 V. Find the oxidation number of Mn in MnO_4^-. Write the lowest whole-number net equation. Then find E°reaction and say whether the reaction is spontaneous.

B2: x + 4(−2) = −1, so Mn is +7. B4: MnO_4^- has the higher E°, so it is the oxidizing agent and is reduced. Nickel is the reducing agent: $Ni \rightarrow Ni^{2+} + 2e^-$. B3: the smallest number that 5 and 2 both reach is 10. Multiply the MnO_4^- half by 2 and the nickel half by 5. Add and cancel $10e^-$: $2MnO_4^- + 16H^+ + 5Ni \rightarrow 2Mn^{2+} + 8H_2O + 5Ni^{2+}$. Check: Mn 2; O 8; H 16; Ni 5. Charge: −2 + 16 = +14 and +4 + 10 = +14. E°reaction = +1.51 − (−0.26) = +1.77 V. Do not multiply E° by 2 or by 5. The sign is positive, so the reaction is spontaneous.