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)

B5 · Balance redox equations in neutral and basic solution

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

B5 · Balance redox equations in neutral and basic solution

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.

RT · Redox titrations

Completed notes | Notes 1/3

Use titre, mole ratio and volume to find an unknown concentration or mass, and judge how an error changes the result.

1. The titration and its volumes

Two solutions, two jobs

A titration finds an unknown concentration. The **titrant** is the solution of known concentration in the buret, the tall tube with a tap. The **analyte** is the solution of unknown concentration in the flask.

We add titrant until the reaction is just complete. A **pipette** delivers a fixed volume of analyte into the flask more precisely than a graduated cylinder. A **buret** delivers the titrant and measures how much was used.

The titre is the volume of titrant delivered: titre = **final reading − initial reading**. A buret counts up from zero as liquid leaves. Read the meniscus, the curved surface, at eye level.

With initial 0.40 mL and final 18.90 mL, the titre is **18.50 mL**. Never assume the start is 0.00.

On the diploma: the readings are given, and a wrong option often uses the final reading without subtracting.

Watch: A titration has one known and one unknown (Alberta Chemistry Teacher, “Chemistry 30 - Redox titration example. Acidified chromium (II) ions with dichromate.”, 0:24)

Choose the trials

One trial may hold a mistake, so we repeat it and average only the trials that agree within the limit the question gives. Given 18.50 mL, 18.60 mL and 19.60 mL with a 0.20 mL limit, the agreeing trials are **18.50 and 18.60**.

Their average is **18.55 mL**. The 19.60 mL trial is left out because it **disagrees by more than the limit**, probably because too much titrant was added past the endpoint, which is the colour change that tells us to stop.

Watch: Average the volumes (Alberta Chemistry Teacher, “Chemistry 30 - Redox titration example. Acidified chromium (II) ions with dichromate.”, 8:31)

Watch: Leave out any trial more than 0.2 mL away (Alberta Chemistry Teacher, “Chemistry 30 - Redox titration example. Acidified chromium (II) ions with dichromate.”, 9:17)

Initial 0.40 mL; final 18.90 mL. Titre = 18.50 mL

Trials 18.50, 18.60, 19.60 mL (limit 0.20 mL). Average of agreeing trials = 18.55 mL

Chain: n(titrant) → mole ratio → n(analyte) → c(analyte)

Divide moles of analyte by the analyte (flask) volume.

RT · Redox titrations

Completed notes | Notes 2/3

2. The equation and the endpoint

The equation gives the ratio

The balanced equation counts particles, so its coefficients also give mole ratios. Given $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$. One MnO_4^- reacts with **5** Fe^{2+}, so $n(Fe^{2+}) = \mathbf{5} \times n(MnO_4^-)$.

The reason is electrons, as in B3. Each MnO_4^- takes **5e⁻** (Mn +7 → +2) and each Fe^{2+} gives only **1e⁻** (+2 → +3). So five Fe^{2+} give the electrons for one MnO_4^-.

Reduction values from the data booklet: MnO_4^-/Mn^{2+} +1.51 V and Fe^{3+}/Fe^{2+} +0.77 V. The higher value is the oxidizing agent, so E° = 1.51 − 0.77 = **+0.74 V**. Acidified MnO_4^- or $Cr_2O_7^{2-}$ is usually the strongest oxidizing agent in a titration.

A positive value means the direction is spontaneous. It does not say the reaction is fast.

On the diploma: the mole ratio comes from the equation in the question, and wrong options often use a 1 : 1 ratio.

Watch: Start with a balanced redox equation (Alberta Chemistry Teacher, “Chemistry 30 - Redox titration example. Acidified chromium (II) ions with dichromate.”, 0:57)

Watch: Match the electrons, then write the equation (Alberta Chemistry Teacher, “Chemistry 30 - Redox titration example. Acidified chromium (II) ions with dichromate.”, 5:01)

Watch: Required over given: the coefficient ratio (Alberta Chemistry Teacher, “Chemistry 30 - Redox titration example. Acidified chromium (II) ions with dichromate.”, 13:42)

Endpoint and equivalence

MnO_4^- is **purple** and Mn^{2+} in dilute solution is almost **colourless**. While Fe^{2+} remains, each drop of purple turns colourless.

After the last Fe^{2+} reacts, one extra drop leaves a faint **pink** that stays. Permanganate is its own indicator, so no extra indicator is needed.

Equivalence is the exact point where moles are in the **equation ratio**, and we calculate it. The endpoint is the colour change we **see**. A deep pink means too much titrant was added, so the titre is too **large**.

On the diploma: describe the endpoint as the first faint colour that stays, and keep equivalence (calculated) apart from endpoint (seen).

Watch: A redox titration: titrant and sample (Alberta Chemistry Teacher, “Chemistry 30 - Example: predicting redox reaction and observations of reactants and products”, 10:26)

Watch: Colours of ions on page 11 of the booklet (Alberta Chemistry Teacher, “Chemistry 30 - Example: predicting redox reaction and observations of reactants and products”, 10:47)

Watch: Orange fades to blue-green at the equivalence point (Alberta Chemistry Teacher, “Chemistry 30 - Example: predicting redox reaction and observations of reactants and products”, 11:45)

Watch: No indicator needed: use the colours of the reactants (Alberta Chemistry Teacher, “Chemistry 30 - Redox titration example. Acidified chromium (II) ions with dichromate.”, 17:01)

Watch: Permanganate: the purple tells you when to stop (Alberta Chemistry Teacher, “Chemistry 30 - Redox titration example. Acidified chromium (II) ions with dichromate.”, 18:21)

RT · Redox titrations

Completed notes | Notes 3/3

3. The calculation chain and errors

Titrant moles to concentration

Concentration is moles per litre, so volumes go in litres. Given 25.00 mL of Fe^{2+} and an average titre of 18.50 mL of 0.0200 mol/L MnO_4^-. $n(MnO_4^-) = 0.0200 \times 0.01850 =$ **3.70×10^{-4} mol**.

Use the ratio, 5 Fe^{2+} for each MnO_4^-: $n(Fe^{2+}) = 5 \times 3.70 \times 10^{-4} =$ **1.85×10^{-3} mol**. Divide by the **flask volume** 0.02500 L to get **0.0740 mol/L**.

We do not divide by the **titre**. It is the volume of titrant, not the volume of the solution that held the iron(II).

On the diploma: forgetting to convert mL to L is a common wrong option, so convert first and write the unit beside each number.

Watch: The same three steps every time (Alberta Chemistry Teacher, “Chemistry 30 - Redox titration example. Acidified chromium (II) ions with dichromate.”, 10:45)

Watch: Step 1: moles of the known (Alberta Chemistry Teacher, “Chemistry 30 - Redox titration example. Acidified chromium (II) ions with dichromate.”, 12:02)

Watch: Step 3: concentration of the unknown (Alberta Chemistry Teacher, “Chemistry 30 - Redox titration example. Acidified chromium (II) ions with dichromate.”, 14:55)

Aliquots and errors

An aliquot is a measured part of a larger solution. If 20.00 mL is taken from 100.0 mL, the whole is 100.0 ÷ 20.00 = 5.000 times bigger, so multiply the aliquot moles by **5.000**.

Mass of iron = moles × **55.85 g/mol**. Mass percent = mass of iron ÷ mass of **sample** × 100.

The titre sets the moles of titrant, those set the moles of analyte, and those set the concentration. Adding too much titrant makes the titre too large, so all three come out too **high**.

On the diploma: to judge an error, follow the chain and mark each step larger, smaller or unchanged before you choose.

Watch: Buret and pipette: what is in each (Alberta Chemistry Teacher, “Chemistry 30 - Redox titration example. Acidified chromium (II) ions with dichromate.”, 8:11)

Watch: Pipette precision and significant figures (Alberta Chemistry Teacher, “Chemistry 30 - Redox titration example. Acidified chromium (II) ions with dichromate.”, 15:45)

RT · Worked practice

Redox titrations | 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.

RT-P01 · Easy · Buret readings: initial 3.05 mL, final 24.55 mL. Find the titre.

1. Titre = final − initial = 24.55 − 3.05 = 21.50 mL.

2. The readings increase as liquid leaves the buret, so the final reading is larger. Do not assume the initial reading is 0.00.

RT-P02 · Easy · Trials: 20.10 mL, 20.20 mL, 21.30 mL. The limit is 0.20 mL. Which trials are averaged, and what is the average?

1. The first two trials differ by 0.10 mL, within the limit. The third differs by more than 1 mL, so it is left out.

2. The average is (20.10 + 20.20) ÷ 2 = 20.15 mL.

RT · Worked practice

RT-P03 · Easy · Given $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$. How many moles of Fe^{2+} react with 2.0×10^{-4} mol of MnO_4^-? Explain the ratio using electrons.

1. $n(Fe^{2+}) = 5 \times 2.0 \times 10^{-4} = 1.0 \times 10^{-3}$ mol.

2. Each MnO_4^- takes 5 electrons (Mn +7 → +2) and each Fe^{2+} gives 1 electron, so five Fe^{2+} are needed for each MnO_4^-.

RT-P04 · Easy · Use +1.51 V for MnO_4^-/Mn^{2+} and +0.77 V for Fe^{3+}/Fe^{2+}. Which reactant is the oxidizing agent, and what is E°cell?

1. MnO_4^- is the oxidizing agent because +1.51 V is the higher reduction value. Fe^{2+} is the reducing agent.

2. E°cell = 1.51 − 0.77 = +0.74 V. The positive value says the direction is spontaneous; it does not say how fast the reaction is.

RT · Worked practice

Redox titrations | Questions 5-6 of 12

RT-P05 · Medium · Given $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$. A 25.00 mL sample of $Fe^{2+}(aq)$ needs 22.40 mL of 0.0150 mol/L $MnO_4^-(aq)$. Find $c(Fe^{2+})$ in mol/L.

1. $n(MnO_4^-) = 0.0150 \times 0.02240 = 3.36 \times 10^{-4}$ mol.

2. $n(Fe^{2+}) = 5 \times 3.36 \times 10^{-4} = 1.68 \times 10^{-3}$ mol.

3. $c(Fe^{2+}) = 1.68 \times 10^{-3} \div 0.02500 = 0.0672$ mol/L.

RT-P06 · Medium · Explain why a permanganate titration of iron(II) needs no added indicator. Describe the colour before and at the endpoint.

1. MnO_4^- is purple and is reduced to almost colourless Mn^{2+} while Fe^{2+} remains, so each drop turns colourless.

2. After the last Fe^{2+} reacts, one extra drop of MnO_4^- leaves a faint pink-purple that stays. The titrant is its own indicator.

RT · Worked practice

Redox titrations | Questions 7-8 of 12

RT-P07 · Medium · Given $5H_2C_2O_4 + 2MnO_4^- + 6H^+ \rightarrow 10CO_2 + 2Mn^{2+} + 8H_2O$. A 10.00 mL sample of $H_2C_2O_4(aq)$ needs 12.60 mL of 0.0200 mol/L $MnO_4^-(aq)$. Find $c(H_2C_2O_4)$ in mol/L.

1. $n(MnO_4^-) = 0.0200 \times 0.01260 = 2.52 \times 10^{-4}$ mol.

2. The ratio from this equation is 5 $H_2C_2O_4$ for every 2 MnO_4^-: $n(H_2C_2O_4) = 2.52 \times 10^{-4} \times 5/2 = 6.30 \times 10^{-4}$ mol.

3. $c(H_2C_2O_4) = 6.30 \times 10^{-4} \div 0.01000 = 0.0630$ mol/L.

RT-P08 · Medium · A student divides the moles of Fe^{2+} by the titre instead of by the 25.00 mL flask volume. Explain the error and name the volume to divide by.

1. The concentration of the analyte is its moles divided by the volume of the analyte solution that held them.

2. The titre is the volume of titrant. Divide by the flask (pipetted) volume, 25.00 mL = 0.02500 L.

RT · Worked practice

Redox titrations | Questions 9-10 of 12

RT-P09 · Challenging · Given $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$. A 10.00 mL aliquot from a 250.0 mL flask of $Fe^{2+}(aq)$ needs 15.60 mL of 0.0100 mol/L $MnO_4^-(aq)$. Find the mass of iron in the whole flask (Fe: 55.85 g/mol, Alberta Chemistry 30 Data Booklet).

1. $n(MnO_4^-) = 0.0100 \times 0.01560 = 1.56 \times 10^{-4}$ mol. $n(Fe^{2+})$ in the aliquot = $5 \times 1.56 \times 10^{-4} = 7.80 \times 10^{-4}$ mol.

2. The whole flask holds $250.0 \div 10.00 = 25.00$ times as much: 1.95×10^{-2} mol.

3. Mass = $1.95 \times 10^{-2} \times 55.85 = 1.09$ g.

RT-P10 · Challenging · A student rinses the buret with distilled water, not with the permanganate solution, before filling it. Is the calculated $c(Fe^{2+})$ too high, too low or unchanged? Explain with the chain of calculation.

1. The water makes the titrant weaker, so more volume is needed to reach the endpoint. The titre is larger than it should be.

2. The calculation still uses the given concentration, so $n(MnO_4^-)$, $n(Fe^{2+})$ and $c(Fe^{2+})$ are all too high.

RT · Worked practice

Redox titrations | Questions 11-12 of 12

RT-P11 · Challenging · An iron tablet of mass 0.800 g is dissolved and made up to 100.0 mL. A 20.00 mL aliquot needs 14.10 mL of 0.0200 mol/L MnO_4^-(aq). Find the mass percent of iron in the tablet ($MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$; Fe: 55.85 g/mol).

1. $n(MnO_4^-) = 0.0200 \times 0.01410 = 2.82 \times 10^{-4}$ mol. $n(Fe^{2+})$ in the aliquot = $5 \times 2.82 \times 10^{-4} = 1.41 \times 10^{-3}$ mol.

2. The whole is 100.0 ÷ 20.00 = 5.000 times the aliquot: 7.05×10^{-3} mol. Mass of iron = $7.05 \times 10^{-3} \times 55.85 = 0.3937$ g.

3. Mass percent = 0.3937 ÷ 0.800 × 100 = 49.2 %.

RT-P12 · Challenging · A bubble sits in the buret tip at the start and leaves during a titration. The recorded titre is larger than the liquid delivered to the flask. Is the calculated analyte concentration too high, too low or unchanged? Explain.

1. The bubble volume is counted as titrant delivered although it did not enter the flask, so the titre is too large.

2. Following the chain, n(titrant), n(analyte) and the calculated concentration are all too high.

B6 · Voltaic cells

Completed notes | Notes 1/3

Explain how a voltaic cell works, label its parts, write its notation and half-reactions, predict observations and calculate E°cell.

1. Separate the halves

From direct contact to a cell

Zinc in copper(II) solution passes electrons **directly** at the metal surface, and the energy appears as **heat**. To use that energy, we make the electrons travel along a path we choose.

In a voltaic cell (also called a galvanic cell), each metal sits in its own solution and a **wire** joins them. A metal strip is an electrode. One electrode plus its solution is a half-cell, and the solution is the electrolyte, because its ions conduct charge.

Use the reduction table from B4. The higher value, Cu^{2+}/Cu **+0.34 V**, is the oxidizing agent. The lower value, Zn^{2+}/Zn **−0.76 V**, is the reducing agent.

The anode is where **oxidation** happens and the cathode is where **reduction** happens. Here the zinc electrode is the **anode** and the copper electrode is the **cathode**. Memory aid: AN OX (anode, oxidation) and RED CAT (reduction, cathode).

On the diploma: label the electrodes by their job, because an option with the anode and cathode swapped is a classic wrong choice.

Watch: What makes electricity in a zinc–copper cell (Tyler DeWitt, “Introduction to Electrochemistry”, 4:33)

Watch: Set up the two half-cells (Tyler DeWitt, “Galvanic Cells (Voltaic Cells)”, 1:34)

Where charge moves

Electrons leave the anode, travel through the **wire** and arrive at the cathode. Electrons do not travel through the **solutions**. Moving ions carry the charge there.

Zinc releases electrons, so in a voltaic cell the anode is **negative** and the cathode is **positive**.

On the diploma: describe electron flow as going from the anode through the wire to the cathode.

Watch: Why electrons flow through the wire (Tyler DeWitt, “Galvanic Cells (Voltaic Cells)”, 4:41)

Watch: Anode and cathode (Tyler DeWitt, “Galvanic Cells (Voltaic Cells)”, 10:04)

Watch: A way to remember anode and cathode (Tyler DeWitt, “Introduction to Electrochemistry”, 10:44)

Zn^{2+}/Zn −0.76 V; Cu^{2+}/Cu +0.34 V. Higher value is reduced.

Anode (oxidation): Zn Cathode (reduction): Cu

Electron path: Zn electrode → wire → Cu electrode

Notation: $Zn(s) \mid Zn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$

B6 · Voltaic cells

Completed notes | Notes 2/3

2. The salt bridge and cell notation

Keep the charge balanced

At the anode, $Zn \rightarrow Zn^{2+} + 2e^-$ makes the solution gain **positive** charge. At the cathode, Cu^{2+} leaves the solution, which leaves extra **negative** charge on anions, the negative ions such as sulfate.

A salt bridge, a tube of unreactive ions, or a porous barrier lets **ions** move between the solutions. Anions move toward the **anode** beaker and cations, the positive ions, move toward the **cathode** beaker.

Without a bridge, charge builds up and the electron flow **stops**. The bridge does not carry **electrons**.

On the diploma: to explain why a cell stops, say charge builds up in each beaker. Match the ion type to the beaker with the opposite extra charge.

Watch: What the salt bridge does (Tyler DeWitt, “Galvanic Cells (Voltaic Cells)”, 16:48)

Watch: How ions move through the bridge (Tyler DeWitt, “Galvanic Cells (Voltaic Cells)”, 20:00)

Write the cell in one line

Notation: $Zn(s) \mid Zn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$. The **anode** is on the far left and the **cathode** is on the far right, because electrons start at the anode.

A single line | separates **phases**, such as an electrode and its solution. The double line || stands for the **salt bridge**. Always write state symbols.

Read from the left: the half on the left of the double line is **oxidized**. Halves for this cell: $Zn \rightarrow Zn^{2+} + \mathbf{2e^-}$ and $Cu^{2+} + 2e^- \rightarrow \mathbf{Cu}$.

On the diploma: a complete notation has the anode on the left, the double line in the middle and state symbols on every substance.

Watch: Cell notation (Tyler DeWitt, “Galvanic Cells (Voltaic Cells)”, 14:37)

B6 · Voltaic cells

Completed notes | Notes 3/3

3. Predict, then calculate

Observations from the half-reactions

Every prediction comes from a half-reaction. Given Cu(s) | Cu^{2+}(aq) || Ag^{+}(aq) | Ag(s). The copper anode **loses** mass as Cu → Cu^{2+} + $2e^{-}$. The copper solution is already blue, and more Cu^{2+} makes it **deeper** blue.

The silver cathode **gains** mass because $2Ag^{+}$ + $2e^{-}$ → **2Ag**. The Ag^{+} concentration **decreases**.

The overall reaction is Cu + $2Ag^{+}$ → **Cu^{2+} + 2Ag**. Copper gives two electrons and each Ag^{+} takes one, so the silver half is multiplied by **2**.

On the diploma: state the change in mass, concentration and colour from the half-reaction, and do not mix up the two electrodes.

Watch: What you would see in each half-cell (Tyler DeWitt, “Galvanic Cells (Voltaic Cells)”, 9:18)

Cell potential

E°cell = E°cathode − E°**anode**. This is the B4 rule, oxidizing agent minus reducing agent, with new names. Zn/Cu: +0.34 − (−0.76) = **+1.10 V**. Cu/Ag: +0.80 − (+0.34) = **+0.46 V**.

Do not multiply a table value when you multiply a half-reaction. A negative E°cell warns that the **electrodes were swapped**.

The notation always puts the **anode** on the left, even when a drawing shows it on the right. Zn/Ag: E°cell = +0.80 − (−0.76) = **+1.56 V**.

On the diploma: keep the table signs, subtract cathode minus anode, and label electrodes from the table values, not from left and right.

Watch: Which metal pulls electrons harder (Tyler DeWitt, “Introduction to Electrochemistry”, 6:06)

B6 · Worked practice

Voltaic cells | 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.

B6-P01 · Easy · Given Zn^{2+}/Zn −0.76 V and Cu^{2+}/Cu +0.34 V. Name the anode and the cathode, and write the half-reaction at each.

1. Cu^{2+}/Cu has the higher table value, so copper(II) is reduced at the cathode. Zinc is oxidized at the anode.

2. Anode: $Zn \rightarrow Zn^{2+} + 2e^-$. Cathode: $Cu^{2+} + 2e^- \rightarrow Cu$.

B6-P02 · Easy · Describe the path of the electrons and the path of the charge in a Zn / Cu cell with a salt bridge.

1. Electrons leave the zinc anode, travel through the wire and arrive at the copper cathode.

2. In the solutions and the salt bridge, the charge is carried by ions. Electrons do not travel through the solutions.

B6 · Worked practice

Voltaic cells | Questions 3-4 of 12

B6-P03 · Easy · Write the cell notation for a cell in which Mg(s) is oxidized and $Cu^{2+}(aq)$ is reduced, with a salt bridge.

1. Mg(s) | $Mg^{2+}(aq)$ || $Cu^{2+}(aq)$ | Cu(s).

2. The anode half is on the far left, the double line is the salt bridge, and the cathode half is on the far right.

B6-P04 · Easy · Explain why a Zn / Cu cell with wires but no salt bridge or porous barrier stops after a moment.

1. Zn^{2+} builds up in the anode beaker, so positive charge builds up there. The cathode beaker is left with extra negative charge.

2. Without a bridge the ions cannot balance the charge, so the electron flow stops.

B6 · Worked practice

Voltaic cells | Questions 5-6 of 12

B6-P05 · Medium · In a cell with a salt bridge containing $NaNO_3(aq)$, toward which beaker do the NO_3^- ions move in the cell Cu(s) | $Cu^{2+}(aq)$ || $Ag^+(aq)$ | Ag(s)? Explain.

1. The NO_3^- ions move toward the copper (anode) beaker.

2. Cu^{2+} ions are made at the anode, so positive charge builds up there. The anions move in to balance it, and the Na^+ ions move toward the silver beaker.

B6-P06 · Medium · Given Cu^{2+}/Cu +0.34 V and Ag^+/Ag +0.80 V. Write the half-reactions and the overall equation for the Cu / Ag cell.

1. Anode: $Cu \rightarrow Cu^{2+} + 2e^-$. Cathode: $2Ag^+ + 2e^- \rightarrow 2Ag$, because each silver ion takes one electron and copper gives two.

2. Overall: $Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$. Copper is 1 on each side, silver is 2 and charge is +2 on each side.

B6 · Worked practice

Voltaic cells | Questions 7-8 of 12

B6-P07 · Medium · Predict what happens to the mass of each electrode and to the colour of the copper solution while the Cu / Ag cell runs.

1. The copper electrode loses mass because $Cu \rightarrow Cu^{2+} + 2e^{-}$ at the anode. The silver electrode gains mass because Ag^{+} becomes Ag at the cathode.

2. Cu^{2+} increases, so the blue colour of the copper solution gets deeper.

B6-P08 · Medium · Calculate E°cell for the Zn / Cu cell and for the Cu / Ag cell. Use Zn^{2+}/Zn −0.76 V, Cu^{2+}/Cu +0.34 V and Ag^{+}/Ag +0.80 V.

1. Zn / Cu: E°cell = E°cathode − E°anode = 0.34 − (−0.76) = +1.10 V.

2. Cu / Ag: E°cell = 0.80 − 0.34 = +0.46 V.

B6 · Worked practice

Voltaic cells | Questions 9-10 of 12

B6-P09 · Challenging · A student calculates E°cell = 0.34 − 0.80 = −0.46 V for the Cu / Ag cell. What was done wrong, and what is the correct value?

1. The values were subtracted the wrong way round. Silver is the cathode, so its value goes first.

2. The correct value is 0.80 − 0.34 = +0.46 V. A negative answer warns that the electrodes were swapped.

B6-P10 · Challenging · Given Zn^{2+}/Zn −0.76 V and Ag^{+}/Ag +0.80 V. A drawing shows the silver half on the left. Name the anode, write the cell notation and calculate E°cell.

1. Silver has the higher table value, so it is the cathode. Zinc is the anode, even though it is drawn on the right.

2. Notation always puts the anode on the left: $Zn(s) \mid Zn^{2+}(aq) \parallel Ag^{+}(aq) \mid Ag(s)$.

3. E°cell = 0.80 − (−0.76) = +1.56 V.

B6 · Worked practice

Voltaic cells | Questions 11-12 of 12

B6-P11 · Challenging · Predict the change in mass of each electrode in the Zn / Ag cell, Zn(s) | Zn^{2+}(aq) || Ag^{+}(aq) | Ag(s), and write the half-reaction that explains each change.

1. The zinc electrode loses mass: $Zn \rightarrow Zn^{2+} + 2e^{-}$ at the anode.

2. The silver electrode gains mass: $Ag^{+} + e^{-} \rightarrow Ag$ at the cathode.

B6-P12 · Challenging · Explain why the double line in a cell notation is needed and what two half-cells it joins.

1. The double line stands for the salt bridge or porous barrier. Without it, the circuit through the solutions is not complete.

2. It separates the anode half on its left from the cathode half on its right. Each single line separates a metal electrode from its solution.