B1 · Electron transfer and redox

Completed notes | Notes 1/3

Follow the electrons, explain charge changes, write half-reactions and name what each agent makes possible.

1. One transfer and what it does to charge

Start with the moving electron

Redox is all about electrons moving from one particle to another. Here, one electron moves from A to B. A **loses** it and B **gains** it. A reaction that moves electrons like this is called an oxidation–reduction reaction, or redox for short.

Oxidation means loss of electrons, so A is oxidized. **Reduction** means gain of electrons, so B is reduced. The electron B gains is the same electron A lost, so the two changes always happen together.

On the diploma: decide loss or gain of electrons first. Then choose "oxidized" or "reduced". The letters A and B just label particles. They are not element symbols.

Watch: Reduction is gain, oxidation is loss (Tyler DeWitt, “Introduction to Oxidation Reduction (Redox) Reactions”, 0:41)

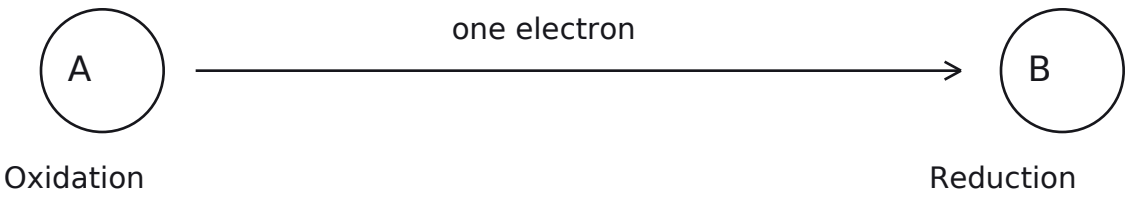
Why the charge changes

An electron carries **negative** charge. Adding a negative amount makes the charge value **decrease**. That is why gaining electrons is called reduction: the charge is reduced.

For example, 0 + (-1) = **-1**, and +3 + (-1) = **+2**. The +2 ion is still positive, but its charge went down. So a positive product can still be a reduction. Removing a negative electron makes the charge **increase**.

A neutral atom has equal positive and negative charge. The nucleus stays the same in these changes, so only the electrons move. When you are given particle counts, charge number = **proton count - electron count**.

Watch: Why gaining electrons is called reduction (Tyler DeWitt, “Introduction to Oxidation Reduction (Redox) Reactions”, 2:02)



A loses; B gains that same electron.

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2. Read the particles and the reaction

Particle words and notation

First, the words. An **atom** is a nucleus with electrons around it. An **ion** is a particle with a net charge, and it can be one atom or a bonded group. A **molecule** is a small bonded group of atoms. NaCl(s) is different: it builds a big ionic lattice, not separate molecules.

Now the symbols. A coefficient is the big number in front, and it counts **copies of a formula**. A subscript is the small lower number, and it counts **atoms within that formula**. A superscript is the small raised number, and it gives the **charge**. So $2Cl^-$ means **two chloride ions**, but Cl_2 is one neutral molecule with **two chlorine atoms**.

The letters in brackets tell you the state. (s) means **solid**, (l) means **liquid**, (g) means **gas**, and (aq) means **dissolved in water**.

On the diploma: read every number in a formula before you count anything. A coefficient, a subscript and a charge look alike but do three different jobs.

A simple shell model

Valence electrons are the ones in the **outer occupied shell**. We are given Na as 2,8,1 and Cl as 2,8,7. When Na gives one electron to Cl, Na^+ becomes **2,8** and Cl^- becomes **2,8,8**. The proton counts stay **unchanged**, because only electrons move.

Full outer shells help explain why Na^+ and Cl^- are common ions. But a full shell does not prove a reaction will happen. What matters is the energy change of the **complete reaction**.

Keep one real reaction in view

Take $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$. Sodium metal and chlorine gas start out neutral. Each Na gives away **one electron**. The Cl_2 molecule has two Cl atoms, and each takes one electron, so together they take in **two electrons in total**.

The Na^+ and Cl^- ions that form attract each other and build an ionic **lattice**, a repeating pattern of ions. Two +1 and two -1 balance, so the solid is neutral overall. The formula NaCl shows the **one-to-one ion ratio**. It does not mean a separate NaCl molecule exists.

On the diploma: always compare the particles before and after the reaction.

Watch: Sodium and chlorine make sodium chloride (Tyler DeWitt, “Introduction to Oxidation Reduction (Redox) Reactions”, 3:45)

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3. Half-reactions, agents and changed examples

Two halves of the same sodium/chlorine story

A half-reaction shows only one half of a redox reaction. Reduction: Cl_2 + **$2e^-$** → $2Cl^-$. The electrons go on the left because chlorine **takes them in**. Check it: two Cl on each side, and the total charge is **-2** on each side.

Oxidation: 2Na → $2Na^+$ + **$2e^-$**. The electrons go on the right because sodium **releases them**. The right side is +2 - 2 = **0**, which matches the neutral left side.

Now add the two halves. The same $2e^-$ sits on **opposite sides**, so it cancels. The electrons are not destroyed. They move from sodium to chlorine. What is left is 2Na + Cl_2 → $2Na^+$ + $2Cl^-$, and those ions form NaCl(s).

On the diploma: in a reduction half-reaction the electrons are reactants. In an oxidation half-reaction they are products.

Watch: Writing a half-reaction (Tyler DeWitt, “Introduction to Oxidation Reduction (Redox) Reactions”, 7:41)

Watch: Two half-reactions make the whole reaction (Tyler DeWitt, “Introduction to Oxidation Reduction (Redox) Reactions”, 10:45)

Watch: Oxidation and reduction always happen together (Tyler DeWitt, “Introduction to Oxidation Reduction (Redox) Reactions”, 11:14)

Agent names describe the partner’s change

An agent is a substance that makes a change happen in another substance. Na supplies the electrons, so chlorine can be **reduced**. That makes Na the **reducing agent**, even though Na itself is oxidized.

Cl_2 accepts the electrons, so sodium can be **oxidized**. That makes Cl_2 the **oxidizing agent**, even though Cl_2 itself is reduced. The two kinds of name answer two questions: what happens to this substance, and what does it do to its partner?

On the diploma: name the agents from the starting reactants, such as Cl_2, not from the product ions such as Cl^-. The reducing agent is the substance that is oxidized.

Watch: What an agent is (Tyler DeWitt, “Oxidizing Agents and Reducing Agents”, 0:57)

Watch: Which agent is which (Tyler DeWitt, “Oxidizing Agents and Reducing Agents”, 4:54)

Watch: Name the whole compound, not just one atom (Tyler DeWitt, “Oxidizing Agents and Reducing Agents”, 11:01)

Change one feature, keep the reasoning

Now start with NaCl that already exists. Solid NaCl already holds Na^+ and Cl^- ions. Dissolving it only separates ions that already have charges. Na stays +1 and Cl stays

-1, so no electrons move and this is **not redox**.

Seeing charge symbols in an equation does not mean redox. Compare each particle before and after, not just the charge marks.

Now try $Mg + Cu^{2+} \rightarrow Mg^{2+} + Cu$. Mg gives **two electrons** to **Cu^{2+}**. So Mg is the **reducing agent** and Cu^{2+} is the **oxidizing agent**. The oxidizing agent starts positive here, but the same loss and gain reasoning still works.

On the diploma: draw the electron arrow first. Then name the agents from the reactants, Mg and Cu^{2+}, not from the products.

Watch: Charges and ions do not make a reaction redox (Tyler DeWitt, “The Oxidation Reduction Question that Tricks Everyone!”, 0:21)

Watch: The charges never change, so no electrons move (Tyler DeWitt, “The Oxidation Reduction Question that Tricks Everyone!”, 1:22)

Watch: The same reaction starting neutral is redox (Tyler DeWitt, “The Oxidation Reduction Question that Tricks Everyone!”, 3:07)

Watch: Finding the agents in an equation (Tyler DeWitt, “Oxidizing Agents and Reducing Agents”, 6:21)

B1 · Worked practice

Electron transfer and redox | 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.

B1-P01 · Easy · A neutral magnesium atom has 12 protons and 12 electrons. It loses two electrons. How many electrons does it have now? What is its charge? Was it oxidized or reduced?

1. The protons do not change, so there are still 12. The electrons drop to 12 - 2 = 10. The charge is 12 - 10 = +2, so Mg^{2+} forms.

2. Losing electrons is oxidation. The charge is positive because negative electrons were removed, not because protons were added.

B1-P02 · Easy · A Cu^{2+} ion has 29 protons and 27 electrons. It gains two electrons. How many electrons does it have now? What is its charge? Was it oxidized or reduced?

1. Add the electrons: 27 + 2 = 29 electrons. With 29 protons, the net charge is 29 - 29 = 0.

2. Gaining electrons is reduction. The product is neutral Cu metal, not a negative copper ion.

B1 · Worked practice

Electron transfer and redox | Questions 3-4 of 12

B1-P03 · Easy · Complete $Fe^{2+} \rightarrow Fe^{3+} + ___$ with electrons. Is this an oxidation or a reduction? Check the charge on each side.

1. $Fe^{2+} \rightarrow Fe^{3+} + e^-$. Check the charge: the right side is +3 - 1 = +2, which matches the left side.

2. The electron is given off, so this is an oxidation. Do not change the charge written on Fe^{3+} to make it balance. Add the electron.

B1-P04 · Easy · Complete $Ag^+ + ___ \rightarrow Ag$ with electrons. Name the oxidizing agent and check the charge on each side.

1. $Ag^+ + e^- \rightarrow Ag$. The charge on the left is +1 - 1 = 0. The charge on the right is 0. They match.

2. Ag^+ takes the electron, so Ag^+ is reduced. Ag^+ is the oxidizing agent. Ag metal is the product, not the agent.

B1 · Worked practice

B1-P05 · Medium · Given $Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$, name what is oxidized, what is reduced, the reducing agent and the oxidizing agent. Explain the agent names.

1. Mg goes from 0 to +2 by losing two electrons, so Mg is oxidized. Cu^{2+} becomes Cu by gaining two electrons, so Cu^{2+} is reduced.

2. Mg is the reducing agent, because its electrons let the copper ions be reduced. Cu^{2+} is the oxidizing agent, because it takes those electrons and lets magnesium be oxidized.

B1-P06 · Medium · Given $Zn(s) + 2Ag^{+}(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$. Write the oxidation half-reaction and the reduction half-reaction for the amounts shown. Check that the electrons match.

1. $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^{-}$. Zinc is oxidized, so its two electrons are on the product side.

2. $2Ag^{+}(aq) + 2e^{-} \rightarrow 2Ag(s)$. Two silver ions each take one electron. Two electrons are lost and two are gained, so they match.

B1 · Worked practice

B1-P07 · Medium · Given $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$. A student says this cannot be a reduction because the product is positive. Correct the student. Use the electron and the charge.

1. Reduction means gaining electrons. It does not mean the product is negative. Here Fe^{3+} gains one electron.

2. Its charge goes down from +3 to +2. Both are positive, but the charge went down and the electron is on the reactant side. So it is a reduction.

B1-P08 · Medium · Compare $KCl(s) \rightarrow K^+(aq) + Cl^-(aq)$ with $2K(s) + Cl_2(g) \rightarrow 2KCl(s)$. Which one is redox? Explain what happens to the charges in each.

1. Dissolving KCl only separates ions that were already K^+ and Cl^- in the solid. No ion changes its charge and no electrons move. So dissolving is not redox.

2. When KCl forms from K and Cl_2, each K atom (charge 0) becomes K^+ (+1). Each Cl atom (charge 0) becomes Cl^- (-1). Potassium loses electrons and chlorine gains them. So this reaction is redox.

B1 · Worked practice

B1-P09 · Challenging · A question says: three Cu^{2+} ions each gain two electrons. Aluminium atoms each lose three electrons to become Al^{3+}. How many Al atoms are needed? Write the balanced net equation and name both agents.

1. Three copper ions need 3 × 2 = 6 electrons. Each Al atom gives 3 electrons, so two Al atoms give 2 × 3 = 6 electrons.

2. $2Al(s) + 3Cu^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Cu(s)$. The charge is +6 on each side, and the atoms match too.

3. Al is the reducing agent and Cu^{2+} is the oxidizing agent. Coefficients count particles. They never change the charge of an ion.

B1-P10 · Challenging · A student writes the half-reaction $Zn \rightarrow Zn^{2+} + e^{-}$. Which check finds the mistake? Fix it. Explain why changing Zn^{2+} to Zn^{+} is not a correct fix.

1. The atoms match. But the charge is 0 on the left and +2 - 1 = +1 on the right. The charge check finds the mistake.

2. Use $Zn \rightarrow Zn^{2+} + 2e^{-}$. Now the charge is 0 on the left and +2 - 2 = 0 on the right. Zn^{2+} is the product that was given. Zn^{+} would be a different ion.

B1 · Worked practice

Electron transfer and redox | Questions 11-12 of 12

B1-P11 · Challenging · A beaker holds Fe^{2+} and Ag^{+}. The reaction $Fe^{2+} + Ag^{+} \rightarrow Fe^{3+} + Ag$ happens. Draw or describe the arrow that shows the electron moving. A student says “both are positive, so neither can give an electron”. Explain why that is wrong.

1. $Fe^{2+} \rightarrow Fe^{3+} + e^{-}$ gives one electron, and $Ag^{+} + e^{-} \rightarrow Ag$ takes it. Draw the arrow from Fe^{2+} to Ag^{+} and write “1 electron” on it.

2. A positive ion can still lose another electron. What matters is the change that happens, not whether the ion starts positive.

B1-P12 · Challenging · Zinc is put in blue copper(II) sulfate solution. A reddish solid forms on the zinc and the blue colour fades. Give one observation and one explanation in terms of particles. Use $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$ to explain both changes. What happens to the sulfate ion?

1. Observation: a reddish solid appears, or the blue colour fades. These are things we can see. They do not show electrons directly.

2. The equation tells us the solid is Cu. It forms when Cu^{2+} gains electrons. Zinc gives the electrons and becomes Zn^{2+} in the solution. Blue Cu^{2+} is used up, so the blue colour fades.

3. The sulfate ion does not change. It is a spectator ion, as in lesson NI. It is in the solution but takes no part in the reaction. It is not the oxidizing agent.

B2 · Oxidation numbers and redox changes

Completed notes | Notes 1/3

Find the oxidation number of each atom with the rules and the sum rule. Know the exceptions. Use the numbers to find what is oxidized and reduced.

1. Rules and formula counting

An assigned value, not always a real ion charge

An oxidation number is an **electron-counting assignment**. It is the charge an atom would have if we gave all the bonding electrons to the more electronegative atom, the one that pulls on bonding electrons harder. Identical atoms share the electrons equally.

In H_2 the two atoms are the same, so each H is 0. In HCl, chlorine pulls harder, so Cl is -1 and H is +1. This is a counting method, not a measurement. HCl gas does not contain separate H^+ and Cl^- ions.

An element by itself has oxidation number **0**, including O_2 and F_2. A one-atom ion, called a monatomic ion, has oxidation number equal to **its ion charge**. For Na^+: 11 protons - 10 electrons = +1.

On the diploma: Cl_2 is 0, not -1, because it is an element by itself. Giving it the usual compound value is a common wrong option.

Watch: The oxidation number rules (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 0:26)

Watch: An element by itself is 0 (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 0:58)

Watch: A one-atom ion equals its charge (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 1:47)

Keep the rule set together

Check the free-element and one-atom-ion cases first. Then use the compound rules. In compounds, Group 1 is **+1**, Group 2 is **+2**, and F is **-1**. Other halogens are usually **-1**, but Cl, Br and I can be positive when combined with O or F.

H is usually **+1**, but it is **-1** in metal hydrides such as NaH. O is usually **-2**, but it is **-1** in peroxides and can be positive with fluorine. These are usual values, so check each formula for an exception.

The oxidation numbers of all atoms must add up to the charge of the whole formula. A neutral compound has sum **0**. A polyatomic ion, a charged group of bonded atoms, has sum equal to **the whole ion’s signed charge**. A subscript multiplies an atom’s value, but a coefficient in front does not change it.

On the diploma: write oxidation numbers with the sign first, such as +6. Numerical-response answers record only the digits asked for, so read how to enter the sign.

Watch: Group 1A, group 2A and the halogens (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 2:49)

Watch: Hydrogen is −1 with a metal (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 9:45)

Watch: Oxygen is −1 in a peroxide (Tyler DeWitt, “How to Calculate Oxidation Number Practice Problems”, 9:49)

Watch: Halogens are positive with oxygen (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 10:36)

Watch: Fluorine is always −1 (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 11:50)

Build the sum

CO: oxygen is -2 and CO is neutral, so x + (-2) = 0. That makes carbon **+2**. We find carbon’s value from its own formula, not from another compound.

NH_3: each H is +1, and three H contribute 3 × (+1) = **+3**. The total must be 0, so nitrogen is **-3**. The subscript 3 counts atoms. It does not make each H +3.

For the sulfate diagram, label three things: which number is an atom count, which is the whole-ion charge, and which is the per-oxygen value. Then complete the sulfur calculation.

Watch: A neutral compound adds to 0 (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 3:31)

Watch: Find the unknown: carbon in CO (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 4:53)

Watch: An ion adds to its charge (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 7:45)

Watch: Diploma example: sulfur in H_2SO_3 (Alberta Chemistry Teacher, “Electrochemistry Diploma Prep - Chemistry 30 - Part 1 of 4”, 12:57)

Given SO_4^{2-}; let sulfur = x.

O atom count: 4

Whole-ion charge: -2

Each O: -2

All O total: -8

Sum and sulfur: x + 4(-2) = -2 → x = +6

B2 · Oxidation numbers and redox changes

Completed notes | Notes 2/3

2. Read the formula, then solve step by step

Read the whole formula first

Before you assign anything, count what is in the formula. In SO_4^{2-} there is **one** S and **four** O. The whole ion has a charge of **-2**. In $2SO_4^{2-}$ you have two copies, but each ion still has a charge of **-2**.

Brackets keep a group together. In $Ca(ClO_4)_2$ the 2 outside multiplies **both Cl and O inside the parentheses**. So the full count is **one Ca, two Cl and eight O**. No charge is written, so the atoms must add up to **0**.

On the diploma: count the atoms inside brackets first. Most wrong answers come from forgetting that the outside number multiplies everything inside.

Watch: Multiply by the number of atoms (Tyler DeWitt, “How to Calculate Oxidation Number Practice Problems”, 2:28)

Two signed values, written two ways

An ion charge is written with the number first, then the sign: Ca^{2+} is **2+** and SO_4^{2-} is **2−**. An oxidation number is written with the sign first, then the number: calcium is **+2** and sulfur in sulfate is **+6**.

The charge belongs to the **whole ion**. The oxidation number is a value we assign to **one atom**. Try NH_4^+. Hydrogen is +1, so x + 4(+1) = +1. That makes N equal to **-3**. Check: -3 + 4 = **+1**, which matches the ion charge.

Watch: Write the sign first for oxidation numbers (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 2:18)

Watch: Charge and oxidation number: number and sign swap places (Alberta Chemistry Teacher, “Electrochemistry Diploma Prep - Chemistry 30 - Part 1 of 4”, 22:56)

When the unknown element appears twice

Take $H_2P_2O_7^{2-}$. Hydrogen is +1 and oxygen is -2. The known atoms add to 2(+1) + 7(-2) = **-12**. The ion must total -2, so the two phosphorus atoms together must supply **+10**.

Write it as an equation, with x as the value for one P: 2(+1) + 2x + 7(-2) = -2. Solve it: 2x = **+10**, so x = **+5**. Check: +2 + 2(+5) - 14 = **-2**.

On the diploma: if the unknown element appears twice, divide your total by two at the end. Then check the sum against the ion charge.

Watch: Two carbons in CaC_2 (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 6:58)

Watch: Two phosphorus atoms (Tyler DeWitt, “How to Calculate Oxidation Number Practice Problems”, 6:11)

A full worked example, then your turn

Let’s do $Ca(ClO_4)_2$ with Ca as +2 and O as -2. The equation is +2 + 2x + 8(-2) = 0. Solve: 2x = **14**, so each Cl is **+7**. Check: +2 + 14 - 16 = **0**.

Look at one ClO_4^- group on its own: +7 + 4(-2) = **-1**. Two of those groups balance one Ca^{2+}, so the Cl value stays the same in each group.

Your turn with $Al_2(SO_4)_3$. Take Al as +3 and O as -2. Count first: two Al, **three S** and **twelve O**. Then 2(+3) + 3x + 12(-2) = 0, so each S is **+6**. Check: +6 + 18 - 24 = **0**.

Watch: A formula with brackets: $Ca(ClO_4)_2$ (Tyler DeWitt, “How to Calculate Oxidation Number Practice Problems”, 12:55)

B2 · Oxidation numbers and redox changes

Completed notes | Notes 3/3

3. Exceptions, averages and reaction changes

Model and check

In SO_4^{2-} the target is the charge printed on the ion, -2. So x + 4(-2) = -2, and sulfur is **+6**. Check: +6 - 8 = **-2**. The +6 belongs to sulfur alone. Sulfate does not contain a free S^{6+} ion.

H_2O_2: with H at +1, 2(+1) + 2x = 0 gives O = **-1**. NaH: +1 + x = 0 gives H = **-1**. OF_2: F is -1, so x + 2(-1) = 0 gives O = **+2**. Give fluorine its value first. The sum rule still works when a usual value changes.

In C_3H_8 the eight H at +1 give +8, so 3x + 8(+1) = 0. Average carbon is **-8/3**. This is an average, not a real fractional charge, so do not round it away.

On the diploma: wrong options are often built from a wrong target. Use the printed ion charge, never zero, for an ion such as sulfate.

Watch: An ion: CO_3^{2-} (Tyler DeWitt, “How to Calculate Oxidation Number Practice Problems”, 4:37)

Compare reactants with products

An increase in oxidation number means **oxidation**. A decrease means **reduction**. This works even when both numbers are positive, such as +5 to +2. In $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$, carbon changes **-4 to +4** and oxygen changes **0 to -2**.

Check that the counting balances. Carbon rises by 8. Four oxygen atoms each fall by 2, and 4 × 2 = 8. So the electrons lost equal the electrons gained.

The reducing agent is the whole starting substance **CH_4**. The oxidizing agent is **O_2**. Find the changing element, then go back to its reactant formula. Naming the element or a product is a common wrong choice.

In $2H_2O_2 \rightarrow 2H_2O + O_2$, oxygen starts at -1 and ends at **-2 and 0**. The same reactant is both oxidized and reduced. This is **disproportionation**. Some peroxide oxygen goes down to water and some goes up to O_2.

On the diploma: to call a reaction "not redox", say that no oxidation number changes.

Watch: Number goes up is oxidation, down is reduction (Tyler DeWitt, “Oxidizing Agents and Reducing Agents”, 6:13)

Watch: Diploma example: what is oxidized in aluminium and chlorine (Alberta Chemistry Teacher, “Electrochemistry Diploma Prep - Chemistry 30 - Part 1 of 4”, 2:47)

Watch: Diploma example: oxidation numbers in a combustion reaction (Alberta Chemistry Teacher, “Electrochemistry Diploma Prep - Chemistry 30 - Part 1 of 4”, 8:32)

B2 · Worked practice

Oxidation numbers and redox changes | 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.

B2-P01 · Easy · Find the oxidation numbers of Mg(s), O_2(g), Fe^{3+}(aq) and Cl^-(aq). Say which rule you used for each one.

1. Mg and O_2 are elements by themselves, so each is 0. The subscript in O_2 does not make oxygen -2.

2. Fe^{3+} is +3 and Cl^- is -1. For an ion made of one atom, the oxidation number is the charge.

B2-P02 · Easy · Find the oxidation numbers of K and Cl in KCl. Show that they add up to the charge of the formula.

1. K is +1 because it is a Group 1 metal. Cl is -1 here, the usual value for a halogen.

2. One of each gives +1 + (-1) = 0. KCl is neutral, so the total must be 0.

B2 · Worked practice

Oxidation numbers and redox changes | Questions 3-4 of 12

B2-P03 · Easy · Find the oxidation number of carbon in CO_2. Show how the subscript on oxygen is used in the sum.

1. Each oxygen is -2, so two oxygens give 2(-2) = -4. Let carbon be x.

2. x - 4 = 0, so x = +4. The total is 0 because CO_2 is neutral. Each O is still -2.

B2-P04 · Easy · Find the oxidation number of nitrogen in NH_3. Give the value for one H, and the total for all the H atoms.

1. Each H is +1. Three hydrogens give 3(+1) = +3.

2. x + 3 = 0, so N = -3. Do not give +3 to each H. The 3 counts the atoms.

B2 · Worked practice

Oxidation numbers and redox changes | Questions 5-6 of 12

B2-P05 · Medium · Find the oxidation number of sulfur in SO_3^{2-}. Write the total the numbers must add to before you solve. Then check your answer.

1. The charge of the whole ion is -2, so the numbers must add to -2. Each O is -2, so x + 3(-2) = -2.

2. x - 6 = -2, so S = +4. Check by putting it back: +4 - 6 = -2.

B2-P06 · Medium · Find the oxidation number of nitrogen in NO_3^- and then in NO_2^-. If nitrate changes to nitrite, is nitrogen oxidized or reduced?

1. Nitrate: x + 3(-2) = -1, so x = +5. Nitrite: x + 2(-2) = -1, so x = +3.

2. The number goes down from +5 to +3, so nitrogen is reduced. The charge of the ion stays -1, but the oxidation number of nitrogen changed.

B2 · Worked practice

Oxidation numbers and redox changes | Questions 7-8 of 12

B2-P07 · Medium · Find the oxidation number of oxygen in H_2O_2 and in OF_2. Say why the usual value for oxygen does not work in each one.

1. H_2O_2 is a peroxide. With H at +1, 2 + 2x = 0, so O = -1.

2. In OF_2, F is always -1. So x - 2 = 0, and O = +2. The usual -2 for oxygen does not work in either one.

B2-P08 · Medium · Find the oxidation number of hydrogen in CaH_2. Then find the oxidation number of bromine in BrO_3^-. Show both sums.

1. Ca is a Group 2 metal, so +2 + 2x = 0 and each H = -1. This is a metal hydride, one of the exceptions.

2. In BrO_3^-, oxygen is -2, so x - 6 = -1 and Br = +5. A halogen can be positive when it is joined to oxygen.

B2 · Worked practice

Oxidation numbers and redox changes | Questions 9-10 of 12

B2-P09 · Challenging · For $C_2H_6 + 7/2\ O_2 \rightarrow 2CO_2 + 3H_2O$, find the average oxidation number of carbon on each side. What is oxidized and what is reduced? Name both agents. How many electrons does each C_2H_6 lose?

1. In C_2H_6, $2x + 6 = 0$, so the average for C is -3. In CO_2, C is +4. Each carbon goes up by 7, so the two carbon atoms lose 14 electrons in total.

2. O goes down from 0 to -2, so it is reduced. C_2H_6 is the reducing agent and O_2 is the oxidizing agent. Name the whole reactants, not a single carbon or oxygen atom.

3. Seven oxygen atoms each gain two electrons, so 14 are gained. That matches the 14 lost. The 7/2 in front of O_2 is only a ratio. Doubling every coefficient gives whole numbers.

B2-P10 · Challenging · A student solves NO_3^- with $x + 3(-2) = 0$ and gets +6. Find the mistake and fix the sum. Explain why checking the arithmetic alone would not find the mistake.

1. The student set the total to 0, but the nitrate ion has a charge of -1. The arithmetic is right, but the sum is the wrong one.

2. Use $x - 6 = -1$, so N = +5. Always check the total against the charge of the ion. Not every formula is neutral.

B2 · Worked practice

Oxidation numbers and redox changes | Questions 11-12 of 12

B2-P11 · Challenging · For $2H_2O_2 \rightarrow 2H_2O + O_2$, find the oxidation number of oxygen in all three substances. Explain why H_2O_2 is both agents. Explain why having two products does not prove disproportionation.

1. O is -1 in H_2O_2, -2 in H_2O and 0 in O_2. Some oxygen goes down from -1 and some goes up from -1.

2. H_2O_2 is both the oxidizing agent and the reducing agent, because the same reactant is reduced and oxidized. Two products alone would not show those two changes.

B2-P12 · Challenging · Find the average oxidation number of carbon in C_2H_6O. A student says the answer is the real charge on each carbon atom. Explain what the calculation shows and what it does not show.

1. Use H = +1 and O = -2: 2x + 6(+1) - 2 = 0. So 2x = -4, and the average for C is -2.

2. This is a count for the whole formula, averaged over two carbon atoms. It does not show that the two carbons are the same. It does not show real C^{2-} ions in the molecule. Looking at each carbon by itself is not part of this lesson.

NI · Net ionic equations

Completed notes | Notes 1/2

Write a dissolved ionic compound as its ions. Write a net ionic equation, name the spectator ions, and check atoms and charge.

1. Which substances are written as ions

A dissolved ionic compound is separate ions

An ionic compound that is dissolved in water breaks into its ions. This is called **dissociation**.

$Na_2SO_4(aq)$ is really $2Na^+(aq)$ and $SO_4^{2-}(aq)$. The subscript 2 on Na becomes the **coefficient** 2.

An ion with several atoms, like SO_4^{2-} or NO_3^-, stays **together**. Every ion keeps its charge and the state (aq).

Check the charges: $2(+1) + (-2) = 0$. The ions of a compound always add to **zero**.

Read the state before you split

Split only an ionic compound or a strong acid marked **(aq)**.

A solid, a liquid or a gas stays **whole**: Cu(s), $H_2O(l)$, $Cl_2(g)$.

A solid ionic compound stays whole too. AgCl(s) is **not** split.

A strong acid in water is written as ions. HCl(aq) is $\mathbf{H^+(aq)}$ and $Cl^-(aq)$.

On the diploma: read the state of every substance first. A common wrong choice splits a solid product into ions.

NI · Net ionic equations

Completed notes | Notes 2/2

2. From a full equation to a net ionic equation

Split, cancel, write what is left

Start with the balanced equation, with states: $Mg(s) + 2AgNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + 2Ag(s)$.

Split every dissolved ionic compound: $Mg(s) + 2Ag^+(aq) + 2NO_3^-(aq) \rightarrow Mg^{2+}(aq) + 2NO_3^-(aq) + 2Ag(s)$. This is the **total ionic equation**.

$2NO_3^-(aq)$ is the same on both sides, so it did not change. An ion that does not change is a **spectator ion**.

Cancel the spectator ions. What is left is the **net ionic equation**: $Mg(s) + 2Ag^+(aq) \rightarrow Mg^{2+}(aq) + 2Ag(s)$.

To count an ion, multiply the coefficient by the subscript. $3CuCl_2(aq)$ gives $3Cu^{2+}(aq)$ and $\mathbf{6}Cl^-(aq)$.

On the diploma: a question may ask for the net ionic equation or for the spectator ion. Read which one it wants.

Check it, then ask if it is redox

Check atoms: 1 Mg and 2 Ag on each side. Check **charge**: 2(+1) = +2 on the left and +2 on the right.

Now the electron transfer is easy to see. Mg goes 0 → +2, so it is **oxidized**. Ag^+ goes +1 → 0, so it is **reduced**.

Not every net ionic equation is redox. In $Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$, lead stays +2 and iodine stays −1.

No **oxidation number** changes there, so no electrons move. The ions only join to make a solid.

On the diploma: do not call a reaction redox only because it has ions. Check the oxidation numbers.

NI · Worked practice

Net ionic equations | 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.

NI-P01 · Easy · Given: $MgCl_2(aq)$. Write the ions it forms in water, with the number of each ion.

1. $MgCl_2(aq)$ is a dissolved ionic compound, so it is written as separate ions.

2. $MgCl_2(aq) \rightarrow Mg^{2+}(aq) + 2Cl^-(aq)$. The subscript 2 on Cl becomes the coefficient 2.

NI-P02 · Easy · Given: $(NH_4)_2SO_4(aq)$. Write the ions it forms in water, with the number of each ion.

1. NH_4^+ and SO_4^{2-} are ions with several atoms. Each one stays together.

2. $(NH_4)_2SO_4(aq) \rightarrow 2NH_4^+(aq) + SO_4^{2-}(aq)$. Check the charge: $2(+1) + (-2) = 0$.

NI · Worked practice

Net ionic equations | Questions 3-4 of 12

NI-P03 · Easy · Which of these is written as ions in a total ionic equation: Fe(s), NaOH(aq), $Br_2(l)$, $BaSO_4(s)$? Explain.

1. Only NaOH(aq) is split. It is an ionic compound dissolved in water, so it is $Na^+(aq)$ and $OH^-(aq)$.

2. Fe(s) and $BaSO_4(s)$ are solids and $Br_2(l)$ is a liquid, so those three stay whole.

NI-P04 · Easy · Given the total ionic equation $Zn(s) + Fe^{2+}(aq) + SO_4^{2-}(aq) \rightarrow Zn^{2+}(aq) + SO_4^{2-}(aq) + Fe(s)$. Name the spectator ion and write the net ionic equation.

1. $SO_4^{2-}(aq)$ is the same on both sides, so it is the spectator ion. Cancel it.

2. Net ionic: $Zn(s) + Fe^{2+}(aq) \rightarrow Zn^{2+}(aq) + Fe(s)$. The charge is +2 on each side.

NI · Worked practice

Net ionic equations | Questions 5-6 of 12

NI-P05 · Medium · Given: $Mg(s) + ZnCl_2(aq) \rightarrow MgCl_2(aq) + Zn(s)$. Write the total ionic equation and the net ionic equation. Name the spectator ion.

1. Total ionic: $Mg(s) + Zn^{2+}(aq) + 2Cl^-(aq) \rightarrow Mg^{2+}(aq) + 2Cl^-(aq) + Zn(s)$. Cancel $2Cl^-(aq)$, so Cl^- is the spectator ion.

2. Net ionic: $Mg(s) + Zn^{2+}(aq) \rightarrow Mg^{2+}(aq) + Zn(s)$. Check: Mg 1, Zn 1 and a charge of +2 on each side.

NI-P06 · Medium · Given: $Pb(s) + 2AgNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + 2Ag(s)$. Write the net ionic equation and check its charge.

1. Total ionic: $Pb(s) + 2Ag^+(aq) + 2NO_3^-(aq) \rightarrow Pb^{2+}(aq) + 2NO_3^-(aq) + 2Ag(s)$. Cancel $2NO_3^-(aq)$.

2. Net ionic: $Pb(s) + 2Ag^+(aq) \rightarrow Pb^{2+}(aq) + 2Ag(s)$. Charge: $2(+1) = +2$ on the left and +2 on the right.

NI · Worked practice

Net ionic equations | Questions 7-8 of 12

NI-P07 · Medium · Given: $Mg(s) + 2HBr(aq) \rightarrow MgBr_2(aq) + H_2(g)$. HBr(aq) is a strong acid. Write the net ionic equation and name the spectator ion.

1. Total ionic: $Mg(s) + 2H^+(aq) + 2Br^-(aq) \rightarrow Mg^{2+}(aq) + 2Br^-(aq) + H_2(g)$. Cancel $2Br^-(aq)$, so Br^- is the spectator ion.

2. Net ionic: $Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$. Check: Mg 1, H 2 and a charge of +2 on each side.

NI-P08 · Medium · Given: $Br_2(l) + 2NaI(aq) \rightarrow 2NaBr(aq) + I_2(s)$. Write the net ionic equation. Then say what is oxidized and what is reduced.

1. Total ionic: $Br_2(l) + 2Na^+(aq) + 2I^-(aq) \rightarrow 2Na^+(aq) + 2Br^-(aq) + I_2(s)$. Cancel $2Na^+(aq)$. Net ionic: $Br_2(l) + 2I^-(aq) \rightarrow 2Br^-(aq) + I_2(s)$.

2. I^- goes $-1 \rightarrow 0$, so it is oxidized. Br goes $0 \rightarrow -1$, so Br_2 is reduced. The charge is -2 on each side.

NI · Worked practice

Net ionic equations | Questions 9-10 of 12

NI-P09 · Challenging · Given: $2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$. HCl(aq) is a strong acid. Write the total ionic equation and the net ionic equation. Check the charge.

1. Total ionic: $2Al(s) + 6H^+(aq) + 6Cl^-(aq) \rightarrow 2Al^{3+}(aq) + 6Cl^-(aq) + 3H_2(g)$. Two $AlCl_3$ give $2 \times 3 = 6$ Cl^-. Cancel $6Cl^-(aq)$.

2. Net ionic: $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$. Charge: $6(+1) = +6$ on the left and $2(+3) = +6$ on the right.

NI-P10 · Challenging · Given: $3CaCl_2(aq) + 2Na_3PO_4(aq) \rightarrow Ca_3(PO_4)_2(s) + 6NaCl(aq)$. Write the net ionic equation. Is it a redox reaction? Explain.

1. Split the three dissolved compounds. $Ca_3(PO_4)_2(s)$ is a solid and stays whole. $6Na^+(aq)$ and $6Cl^-(aq)$ are on both sides, so cancel them. Net ionic: $3Ca^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$.

2. It is not a redox reaction. Calcium stays +2, phosphorus stays +5 and oxygen stays −2, so no oxidation number changes. The charge is $3(+2) + 2(-3) = 0$ on the left and 0 on the right.

NI · Worked practice

Net ionic equations | Questions 11-12 of 12

NI-P11 · Challenging · For $Zn(s) + 2AgNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + 2Ag(s)$, a student writes the net ionic equation $Zn(s) + Ag^+(aq) \rightarrow Zn^{2+}(aq) + Ag(s)$. Which check finds the mistake? Write the correct net ionic equation.

1. The charge check finds it. The left side is +1 and the right side is +2, so the equation is not balanced for charge.

2. The student dropped the coefficient 2. Correct net ionic: $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$. Now the charge is +2 on each side.

NI-P12 · Challenging · Given: $Cu(s) + 2Fe(NO_3)_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Fe(NO_3)_2(aq)$. Write the total ionic equation and the net ionic equation. How many nitrate ions cancel from each side?

1. Total ionic: $Cu(s) + 2Fe^{3+}(aq) + 6NO_3^-(aq) \rightarrow Cu^{2+}(aq) + 2NO_3^-(aq) + 2Fe^{2+}(aq) + 4NO_3^-(aq)$. The right side has 2 + 4 = 6 nitrate ions, so $6NO_3^-(aq)$ cancel from each side.

2. Net ionic: $Cu(s) + 2Fe^{3+}(aq) \rightarrow Cu^{2+}(aq) + 2Fe^{2+}(aq)$. Charge: 2(+3) = +6 on the left and +2 + 2(+2) = +6 on the right.

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)

B3 · Balance redox equations in acid

Completed notes | Notes 2/3

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.

B3 · Balance redox equations in acid

Completed notes | Notes 3/3

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.

B4 · Predict reactions and investigate

Completed notes | Notes 1/3

Use test results and the Data Booklet reduction table to predict direction, then judge an investigation’s conclusions.

1. Evidence and strength

Separate observations from their explanation

Balancing cannot tell us which way a reaction goes. $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$ balances, and so does its reverse. A balanced equation does not by itself prove a **spontaneous reaction in the written direction**. Spontaneous means it goes on its own. It does not mean fast.

Here are two example test results. Your class did not measure them. A strip of zinc in a Cu^{2+} solution gets a copper coating. A strip of copper in an Ag^{+} solution gets a silver coating. A new metal coating is evidence that dissolved ions **gain electrons** to become metal. The electrons came from the metal that dissolved.

So a metal that gives electrons to another metal’s ions is the stronger reducing agent. The two comparisons link through copper. Reducing agents rank **Zn > Cu > Ag**. The matching ions run the other way: oxidizing agents rank **Ag^{+} > Cu^{2+} > Zn^{2+}**. A metal and its ion have different roles.

On the diploma: build the order from the pairs, and say which metal was the reducing agent in each pair. Write the exact formula, such as Ag^{+}, not just silver.

Watch: What you would see: bubbles of hydrogen gas (Alberta Chemistry Teacher, “Chemistry 30 - Predicting redox reaction from species in the reduction table”, 11:01)

Watch: Observations to report: bubbles, pop test, litmus (Alberta Chemistry Teacher, “Chemistry 30 - Example: predicting redox reaction and observations of reactants and products”, 12:03)

Watch: What a falling H^{+} does to the pH (Alberta Chemistry Teacher, “Chemistry 30 - Example: predicting redox reaction and observations of reactants and products”, 12:22)

Read a reduction row

Every row in the Alberta Chemistry 30 Data Booklet table of standard reduction potentials is written as a reduction. The electrons are on the **left**. The oxidizing agents are on the left, and the strongest is at the top. The reducing agents are on the right, and the strongest is at the bottom. A metal on the right side of a row can act as a reducing agent, but only when you read that row **reversed**.

An oxidizing agent reacts spontaneously with a reducing agent that is below it in the table. Ag^{+} is above Cu, so they react. Zn^{2+} is below Cu, so they do not. That is why a copper container can hold a Zn^{2+} solution.

You can only choose substances that are really in the beaker. List every ion, any metal, and water. An acidic or “acidified” solution also has H^{+}, and a basic solution has OH^{-}. A

strong oxidizing agent elsewhere in the booklet is not in your beaker. It is like picking a team from the people who showed up. Choose the strongest **oxidizing agent** and **reducing agent** from what is in the beaker.

Label the diagram with the order of reducing agents and the order of oxidizing agents. These are example results, not measurements from a class experiment.

On the diploma: use the substances given in the question, not every row of the table.

Watch: The forward arrow in the table is the reduction (Alberta Chemistry Teacher, “Chemistry 30 - Predicting redox reaction from species in the reduction table”, 2:44)

Watch: Flip the row to write the oxidation (Alberta Chemistry Teacher, “Chemistry 30 - Predicting redox reaction from species in the reduction table”, 3:10)

Zn in Cu^{2+}: copper coating. Cu in Ag^{+}: silver coating.

Metal donors, strongest first: Zn > Cu > Ag

Ion acceptors, strongest first: Ag^{+} > Cu^{2+} > Zn^{2+}

Example results. Your class did not measure them.

B4 · Predict reactions and investigate

Completed notes | Notes 2/3

2. What the table values mean

Energy per charge, measured from one starting point

The symbol E means **electric potential**, measured in volts. The small ° means **standard conditions**: 1.0 mol/L, 25.00 °C and 101.325 kPa.

A voltage is the **energy for each unit of charge**. One volt means **one joule per coulomb**, and a coulomb is a unit of electric charge. For 150 J moved by 100 C, 150 ÷ 100 = **1.50 V**. Double both and 300 ÷ 200 is still **1.50 V**.

Every table value is measured from the same starting point, the hydrogen half-reaction, $2H^+ + 2e^- \rightarrow H_2$, which is given **0.00 V**. Each half-reaction is printed as a **reduction**. A more positive value means a stronger pull on electrons under table conditions.

A table value is not a voltage you could read from one beaker. Voltage compares a **reducing agent–oxidizing agent pair**. In this lesson, putting a metal in a solution gives us a prediction only. It does not measure the voltage of a real cell.

On the diploma: remember the table lists reductions only. Changing a coefficient never changes E°.

Pick the roles first, then subtract

Use E°reaction = E° of the **oxidizing agent** minus E° of the **reducing agent**. We subtract because the reducing agent's half runs **backward as oxidation**. Do not just take larger minus smaller.

A positive answer means the written direction is spontaneous under standard conditions. A negative answer means **the reverse** is spontaneous. It tells you nothing about **speed**. Doubling the amounts does not change the energy per charge, so never multiply E° by the coefficients.

On the diploma: a common wrong option gives $Cu + Zn^{2+}$ a positive E° by subtracting in the wrong order. Check which one is the reducing agent first.

Watch: List everything that is present (Alberta Chemistry Teacher, "Chemistry 30 - Predicting redox reaction from species in the reduction table", 0:03)

Watch: Label the oxidizing and reducing agents (Alberta Chemistry Teacher, "Chemistry 30 - Predicting redox reaction from species in the reduction table", 1:33)

Watch: Find the best of each (Alberta Chemistry Teacher, "Chemistry 30 - Predicting redox reaction from species in the reduction table", 2:24)

A mixture changes the choice, not the method

Suppose a beaker holds Zn(s), Cu(s), Ag^+ and Cu^{2+}. Use only what is present. The strongest reducing agent is **Zn** and the strongest oxidizing agent is **Ag^+**. Choose them before you write any equation.

The reaction is $Zn + 2Ag^+ \rightarrow Zn^{2+} + 2Ag$. $E° = +0.80 - (-0.76) =$ **+1.56 V**. Cu and Cu^{2+} are still in the mixture, but they are not part of this **net equation**.

This predicts which reaction the strongest pair starts with. It does not tell you the final mixture, or the order other reactions would follow.

Watch: Zinc in hydrochloric acid: list everything present (Alberta Chemistry Teacher, “Chemistry 30 - Predicting redox reaction from species in the reduction table”, 6:02)

Watch: Watch for pairs such as chloride and water (Alberta Chemistry Teacher, “Chemistry 30 - Predicting redox reaction from species in the reduction table”, 9:37)

Watch: Build the list from a longer formula list (Alberta Chemistry Teacher, “Chemistry 30 - Example: predicting redox reaction and observations of reactants and products”, 0:36)

Watch: Tin(II) is both an oxidizing and a reducing agent (Alberta Chemistry Teacher, “Chemistry 30 - Example: predicting redox reaction and observations of reactants and products”, 2:22)

Watch: The common mistake: missing tin(II) as a reducing agent (Alberta Chemistry Teacher, “Chemistry 30 - Example: predicting redox reaction and observations of reactants and products”, 7:42)

B4 · Predict reactions and investigate

Completed notes | Notes 3/3

3. Predict, then test the claim

Table values

From the Alberta Chemistry 30 Data Booklet, printed as reductions: Ag^+/Ag **+0.80 V**; Cu^{2+}/Cu **+0.34 V**; Zn^{2+}/Zn **-0.76 V**. These hold at standard table conditions, which are 1.0 mol/L solutions at 25.00 °C. A more positive value means a stronger pull on electrons.

For the reaction as written, E°reaction = E° of the **oxidizing agent** minus E° of the **reducing agent**. We subtract because the reducing agent runs its table row backwards, as an oxidation. A positive result means the written direction is spontaneous. A negative result means the reverse is spontaneous.

For $Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$, Ag^+ is the oxidizing agent and Cu is the reducing agent: +0.80 - (+0.34) = **+0.46 V**. A volt is one joule of energy per coulomb of charge, so it is a ratio. Do not multiply E° when you multiply a half-reaction's coefficients.

On the diploma: check which value is the oxidizing agent's and which is the reducing agent's before you subtract.

Direction matters

For the proposed $Cu + Zn^{2+} \rightarrow Cu^{2+} + Zn$, Zn^{2+} is the oxidizing agent and Cu is the reducing agent: -0.76 - (+0.34) = **-1.10 V**. Always taking larger minus smaller skips the choice of reducing agent and oxidizing agent. It can answer the wrong direction.

The negative sign means the reverse, Zn with Cu^{2+}, is spontaneous at +1.10 V. The sign tells direction, not speed. No visible change in a short trial is **not proof that reaction is impossible**, because the reaction might be too slow to see.

On the diploma: a sign-flipped answer is a common wrong option, so match the sign to the reducing agent and the oxidizing agent.

Watch: The spontaneity rule: oxidizing agent above reducing agent (Alberta Chemistry Teacher, "Chemistry 30 - Predicting redox reaction from species in the reduction table", 4:59)

Watch: Below means it does not happen by itself (Alberta Chemistry Teacher, "Chemistry 30 - Predicting redox reaction from species in the reduction table", 5:18)

Watch: Check the spontaneity rule at the end (Alberta Chemistry Teacher, "Chemistry 30 - Example: predicting redox reaction and observations of reactants and products", 10:09)

A fair comparison and a supported conclusion

In a fair test, change one metal-ion pair at a time. Keep the concentration, volume, strip size, temperature and time **the same**. Add a **control**: a test where nothing should happen, such as Cu in Cu^{2+}.

A good conclusion gives a claim, then the evidence, then an **explanation of who gave electrons**. Name the metal, the ion that changed and who gave the electrons. The simulation follows fixed rules, so running it again adds no new **evidence**. It does not test hands-on lab skill.

On the diploma: tell an observation, such as seeing a deposit, from an inference, such as saying a metal is the stronger reducing agent. Cite the actual comparisons in your answer.

B4 · Worked practice

Predict reactions and investigate | 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.

B4-P01 · Easy · In the reaction $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$, name the reducing agent and the oxidizing agent. Which metal is the stronger reducing agent, Zn or Cu?

1. Zn gives electrons and Cu^{2+} takes them. Zinc reduces the copper ions to Cu metal.

2. So Zn is a stronger reducing agent than Cu. Do not put Cu^{2+} in the list of reducing agents. It is the oxidizing agent.

B4-P02 · Easy · Test results: Zn gives electrons to Cu^{2+}, and Cu gives electrons to Ag^{+}. Put Zn, Cu and Ag in order as reducing agents, strongest first. Say which two results support your order.

1. Zn gives electrons to Cu^{2+}, so Zn > Cu. Cu gives electrons to Ag^{+}, so Cu > Ag.

2. Copper is in both results, so it links them: Zn > Cu > Ag. This order comes from two tests, not three.

B4 · Worked practice

Predict reactions and investigate | Questions 3-4 of 12

B4-P03 · Easy · The metals are in the order Zn > Cu > Ag as reducing agents. Put their ions, Ag^+, Zn^{2+} and Cu^{2+}, in order as oxidizing agents, strongest first.

1. The order of the oxidizing agents is $Ag^+ > Cu^{2+} > Zn^{2+}$.

2. Each ion takes electrons and becomes its metal. A metal that gives electrons easily makes an ion that takes them back poorly. So the two orders run opposite ways.

B4-P04 · Easy · A beaker contains Cu(s), Ag^+(aq) and nitrate ions. The nitrate ions are spectator ions here. Which two substances could react in a redox reaction? Explain why Zn cannot be one of them.

1. The pair is Cu metal, as the reducing agent, and Ag^+, as the oxidizing agent. Nitrate is a spectator ion, so it does not change.

2. Zn is not in the beaker. A stronger reducing agent in the table cannot react if it is not there.

B4 · Worked practice

Predict reactions and investigate | Questions 5-6 of 12

B4-P05 · Medium · Use these rows from the Data Booklet table of standard electrode potentials: $Ag^+ + e^- \rightarrow Ag$, +0.80 V; $Cu^{2+} + 2e^- \rightarrow Cu$, +0.34 V; $Zn^{2+} + 2e^- \rightarrow Zn$, -0.76 V. The conditions are standard: 25 °C and 1.0 mol/L solutions. Copper metal is put in a solution of Ag^+. Calculate E° for the reaction and write the balanced net equation. Is the reaction spontaneous?

1. Ag^+ is the oxidizing agent and Cu is the reducing agent. In the table, Ag^+ is above Cu. E° = +0.80 - (+0.34) = +0.46 V.

2. $Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$. The value is positive, so the reaction is spontaneous under standard conditions.

B4-P06 · Medium · Use these rows from the Data Booklet table of standard electrode potentials: $Ag^+ + e^- \rightarrow Ag$, +0.80 V; $Cu^{2+} + 2e^- \rightarrow Cu$, +0.34 V; $Zn^{2+} + 2e^- \rightarrow Zn$, -0.76 V. The conditions are standard: 25 °C and 1.0 mol/L solutions. Calculate E° for the reaction $Cu + Zn^{2+} \rightarrow Cu^{2+} + Zn$. Is the reaction spontaneous as written?

1. In the reaction as written, Zn^{2+} is the oxidizing agent and Cu is the reducing agent. E° = -0.76 - (+0.34) = -1.10 V.

2. The value is negative, so the reaction is not spontaneous as written. The reverse reaction is spontaneous. Subtracting the smaller number from the larger would answer the reverse question.

B4 · Worked practice

Predict reactions and investigate | Questions 7-8 of 12

B4-P07 · Medium · Use these rows from the Data Booklet table of standard electrode potentials: $Ag^+ + e^- \rightarrow Ag$, +0.80 V; $Cu^{2+} + 2e^- \rightarrow Cu$, +0.34 V; $Zn^{2+} + 2e^- \rightarrow Zn$, -0.76 V. The conditions are standard: 25 °C and 1.0 mol/L solutions. Zinc metal is put in a solution of Ag^+. Write the balanced equation and calculate E°. A student doubles +0.80 V because two Ag^+ are needed. Explain the mistake.

1. $Zn + 2Ag^+ \rightarrow Zn^{2+} + 2Ag$. Zn loses two electrons, and two silver ions take one each.

2. E° = +0.80 - (-0.76) = +1.56 V. Do not multiply E° when you multiply a half-reaction. E° is the energy for each unit of charge. More silver does not change it.

B4-P08 · Medium · A simulation shows the same result every time a test is replayed. Can five replays tell you how much a real experiment changes from try to try? Say what the replays can check, and one thing they cannot measure.

1. Five replays can check that the simulation gives the same result for the same input.

2. They are not separate real measurements. So they cannot show the effect of a dirty metal surface, the timing, the temperature or the concentration. A simulation follows fixed rules.

B4 · Worked practice

Predict reactions and investigate | Questions 9-10 of 12

B4-P09 · Challenging · Test results, all under the same conditions: a strip of metal A in B^{2+} solution gets a coating of metal B. A strip of metal B in C^{+} solution gets a coating of metal C. Predict what happens when A is put in C^{+} solution. Write the balanced equation. A forms A^{2+}. Explain how the two results lead to your prediction.

1. The tests show A > B and B > C as reducing agents. So A > C, and A will reduce C^{+}.

2. $A(s) + 2C^{+}(aq) \rightarrow A^{2+}(aq) + 2C(s)$. Two C^{+} ions take one electron each, which matches the two electrons A loses. This prediction is for the same conditions as the tests.

B4-P10 · Challenging · Two tests compare Zn in Cu^{2+} solution with Cu in Ag^{+} solution. They use different concentrations, different surface areas and different times. Name three things to keep the same so the comparison is fair. Explain why a test with a metal in a solution of its own ion helps.

1. Keep the ion concentration, the surface area of the metal and the time the same. Also keep the temperature and the volume of solution the same.

2. Zn in Zn^{2+}, or Cu in Cu^{2+}, shows no change. This is a control. It shows that any change in the other tests comes from the metal and the ion being different. A control does not remove every possible error.

B4 · Worked practice

Predict reactions and investigate | Questions 11-12 of 12

B4-P11 · Challenging · Use these rows from the Data Booklet table of standard electrode potentials: $Ag^+ + e^- \rightarrow Ag$, +0.80 V; $Cu^{2+} + 2e^- \rightarrow Cu$, +0.34 V; $Zn^{2+} + 2e^- \rightarrow Zn$, -0.76 V. The conditions are standard: 25 °C and 1.0 mol/L solutions. A student reports $E° = +1.10$ V for $Cu + Zn^{2+} \rightarrow Cu^{2+} + Zn$, by subtracting -0.76 from +0.34. Explain the mistake. Give the correct value with its sign. Which reaction does +1.10 V really describe?

1. The student used Cu^{2+} as the oxidizing agent, but the beaker has Cu metal. In the reaction as written, the oxidizing agent is Zn^{2+}.

2. For the reaction as written, -0.76 - (+0.34) = -1.10 V. The value +1.10 V is for $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$, the reverse reaction.

B4-P12 · Challenging · A lab record says: a clean Cu strip in Ag^+ solution gets a silver-coloured coating. In a control, Cu in Cu^{2+} solution shows no coating. Write a conclusion about the electrons that the record supports. Give two limits of the evidence.

1. The coating fits Ag^+ taking electrons from Cu and forming Ag metal, while Cu is oxidized to Cu^{2+}. The control supports this, because the coating forms only when the metal and the ion are different.

2. Colour alone does not prove what the coating is. A chemical test would make the claim stronger. The record also gives no speed and no change in mass.

3. No visible change in the control is an observation. It does not prove that no reaction at all can happen.