B1 · Electron transfer and redox

Student notes - complete the underlined spaces | Notes 1/3

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

Name: _____ Date: _____

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 _____ it and B _____ it. A reaction that moves electrons like this is called an oxidation–reduction reaction, or redox for short.

_____ means loss of electrons, so A is oxidized. _____ 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)

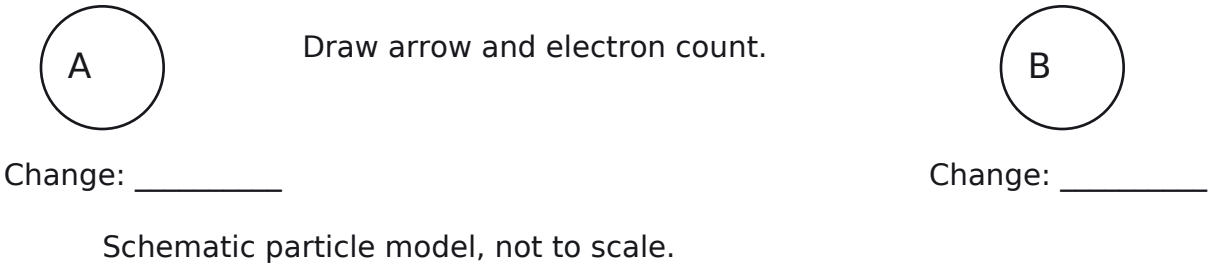
Why the charge changes

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

For example, 0 + (-1) = _____, and +3 + (-1) = _____. 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 _____.

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 = _____.

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



Schematic particle model, not to scale.

B1 · Electron transfer and redox

Student notes - complete the underlined spaces | Notes 2/3

2. Read the particles and the reaction

Particle words and notation

First, the words. An _____ is a nucleus with electrons around it. An _____ is a particle with a net charge, and it can be one atom or a bonded group. A _____ 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 _____. A subscript is the small lower number, and it counts _____. A superscript is the small raised number, and it gives the _____. So $2Cl^-$ means _____, but Cl_2 is one neutral molecule with _____.

The letters in brackets tell you the state. (s) means _____, (l) means _____, (g) means _____, and (aq) means _____.

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 _____. We are given Na as 2,8,1 and Cl as 2,8,7. When Na gives one electron to Cl, Na^+ becomes _____ and Cl^- becomes _____. The proton counts stay _____, 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 _____.

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 _____. The Cl_2 molecule has two Cl atoms, and each takes one electron, so together they take in _____.

The Na^+ and Cl^- ions that form attract each other and build an ionic _____, a repeating pattern of ions. Two +1 and two -1 balance, so the solid is neutral overall. The formula NaCl shows the _____. 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)

B1 · Electron transfer and redox

Student notes - complete the underlined spaces | Notes 3/3

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 + _____ → $2Cl^-$. The electrons go on the left because chlorine _____. Check it: two Cl on each side, and the total charge is _____ on each side.

Oxidation: $2Na \rightarrow 2Na^+$ + _____. The electrons go on the right because sodium _____. The right side is +2 - 2 = _____, which matches the neutral left side.

Now add the two halves. The same $2e^-$ sits on _____, so it cancels. The electrons are not destroyed. They move from sodium to chlorine. What is left is $2Na + Cl_2 \rightarrow 2Na^+ + 2Cl^-$, 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 _____. That makes Na the _____, even though Na itself is oxidized.

Cl_2 accepts the electrons, so sodium can be _____. That makes Cl_2 the _____, 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 _____.

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 _____ to _____. So Mg is the _____ and Cu^{2+} is the _____. 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 · Practice

Electron transfer and redox | Questions 1-3 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?

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?

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

B1 · Practice

Electron transfer and redox | Questions 4-6 of 12

B1-P04 · Easy · Complete Ag^{+} + ___ → Ag with electrons. Name the oxidizing agent and check the charge on each side.

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

Optional start: First draw an arrow from the reactant that gives electrons to the reactant that takes them.

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

B1 · Practice

Electron transfer and redox | Questions 7-9 of 12

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.

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.

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.

B1 · Practice

Electron transfer and redox | Questions 10-12 of 12

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.

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.

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?

B2 · Oxidation numbers and redox changes

Student notes - complete the underlined spaces | 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.

Name: _____ Date: _____

1. Rules and formula counting

An assigned value, not always a real ion charge

An oxidation number is an _____. 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 _____, including O_2 and F_2. A one-atom ion, called a monatomic ion, has oxidation number equal to _____. 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 _____, Group 2 is _____, and F is _____. Other halogens are usually _____, but Cl, Br and I can be positive when combined with O or F.

H is usually _____, but it is _____ in metal hydrides such as NaH. O is usually _____, but it is _____ 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 _____. A polyatomic ion, a charged group of bonded atoms, has sum equal to _____. 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 _____. 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) = _____. The total must be 0, so nitrogen is _____. 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: _____ Whole-ion charge: _____

Each O: _____ All O total: _____

Sum and sulfur: _____

B2 · Oxidation numbers and redox changes

Student notes - complete the underlined spaces | 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 _____ S and _____ O. The whole ion has a charge of _____. In $2SO_4^{2-}$ you have two copies, but each ion still has a charge of _____.

Brackets keep a group together. In $Ca(ClO_4)_2$ the 2 outside multiplies _____. So the full count is _____. No charge is written, so the atoms must add up to _____.

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 _____ and SO_4^{2-} is _____. An oxidation number is written with the sign first, then the number: calcium is _____ and sulfur in sulfate is _____.

The charge belongs to the _____. The oxidation number is a value we assign to _____. Try NH_4^+. Hydrogen is +1, so x + 4(+1) = +1. That makes N equal to _____. Check: -3 + 4 = _____, 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) = _____. The ion must total -2, so the two phosphorus atoms together must supply _____.

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

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 = _____, so each Cl is _____. Check: +2 + 14 - 16 = _____.

Look at one ClO_4^- group on its own: +7 + 4(-2) = _____. 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, _____ and _____. Then 2(+3) + 3x + 12(-2) = 0, so each S is _____. Check: +6 + 18 - 24 = _____.

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

Student notes - complete the underlined spaces | 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 _____. Check: +6 - 8 = _____. 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 = _____. NaH: +1 + x = 0 gives H = _____. OF_2: F is -1, so x + 2(-1) = 0 gives O = _____. 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 _____. 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 _____. A decrease means _____. 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 _____ and oxygen changes _____.

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 _____. The oxidizing agent is _____. 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 _____. The same reactant is both oxidized and reduced. This is _____. 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 · Practice

Oxidation numbers and redox changes | Questions 1-3 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.

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

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

B2 · Practice

Oxidation numbers and redox changes | Questions 4-6 of 12

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.

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.

Optional start: Circle the charge on the whole ion before you write the sum.

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?

B2 · Practice

Oxidation numbers and redox changes | Questions 7-9 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.

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.

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?

B2 · Practice

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

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.

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.

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.

NI · Net ionic equations

Student notes - complete the underlined spaces | 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.

Name: _____ Date: _____

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 _____.

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

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

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

Read the state before you split

Split only an ionic compound or a strong acid marked _____.

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

A solid ionic compound stays whole too. AgCl(s) is _____ split.

A strong acid in water is written as ions. HCl(aq) is _____ 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

Student notes - complete the underlined spaces | 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 _____.

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

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

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

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 _____: 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 _____. Ag^+ goes +1 → 0, so it is _____.

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 _____ 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 · Practice

Net ionic equations | Questions 1-3 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.

Optional start: The subscript tells you how many of that ion there are.

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

Optional start: The 2 outside the brackets counts whole NH_4^+ ions.

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.

Optional start: Read the state after each formula.

NI · Practice

Net ionic equations | Questions 4-6 of 12

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.

Optional start: Look for an ion that is written the same on both sides.

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.

Optional start: Split only the two compounds marked (aq).

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.

Optional start: Two $AgNO_3$ give two Ag^+ and two NO_3^-.

NI · Practice

Net ionic equations | Questions 7-9 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.

Optional start: A strong acid marked (aq) is written as $H^+(aq)$ and its negative ion.

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.

Optional start: $Br_2(l)$ and $I_2(s)$ are not dissolved ionic compounds. Keep them whole.

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.

Optional start: Multiply the coefficient by the subscript to count each ion.

NI · Practice

Net ionic equations | Questions 10-12 of 12

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.

Optional start: A solid product stays whole. Then compare the oxidation numbers before and after.

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.

Optional start: Add the charge 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?

Optional start: Count the nitrate ions from each compound separately, then add them.

B3 · Balance redox equations in acid

Student notes - complete the underlined spaces | Notes 1/3

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

Name: _____ Date: _____

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 _____ and the _____. Electrons carry charge but add _____. 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 _____. So multiply the whole Al half by _____ and the whole Cu half by _____.

Now add the halves: $2Al + 3Cu^{2+} + 6e^- \rightarrow 2Al^{3+} + 6e^- + 3Cu$. Only identical terms on _____ cancel, so the six electrons cancel. The net charge is _____ 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 _____ in permanganate to +2, which is reduction. Fe goes from +2 to +3, which is _____.

Write the real species in each skeleton: _____ and _____. 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 _____, then H with _____, then total charge with _____. 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

Student notes - complete the underlined spaces | 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 _____. Add _____ to the right. Four waters also bring in _____ on the right.

Add _____ 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 = _____, and the right is _____.

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

As a second check, one Mn goes from +7 to +2. That is _____. 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 _____. 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 _____.

The net equation is $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$. The charge is _____ on each side. H^+ and water stay in the equation because they have _____.

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: $MnO_4^- + ___ H^+ + ___ e^- \rightarrow Mn^{2+} + ___ H_2O$

Before e^-: left _____; right _____

After e^-: left _____; right _____

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

B3 · Balance redox equations in acid

Student notes - complete the underlined spaces | 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 _____. That puts two H on the left, so add _____. The charge is now -1 on the left and +1 on the right.

Add _____: $NO_2^- + H_2O \rightarrow NO_3^- + 2H^+ + 2e^-$. N goes from +3 to +5, which is _____. 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 _____. So multiply the first half by _____ and the second by _____.

Add the halves and cancel $10e^-$. The left has $16H^+$ and the right has $10H^+$, so _____ is left over on the left. The left has $5H_2O$ and the right has $8H_2O$, so _____ 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 _____ 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 _____ so the Cr atoms match. The left has seven O, so add _____ on the right. That brings 14 H, so add _____ on the left. The charges are +12 on the left and +6 on the right, so add _____.

Final check: Cr is 2, O is 7, H is 14 and the charge is _____. Each of the two Cr atoms drops by three. That accounts for _____.

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 _____ 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 _____, so just cancel them. You get $OCl^- + H_2O + 2I^- \rightarrow Cl^- + 2OH^- + I_2$.

The charge is _____ on each side. Use the basic form you were given. Do not turn it into an acidic one. Also, balancing does not tell you _____. 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 · Practice

Balance redox equations in acid | Questions 1-3 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.

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

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?

B3 · Practice

Balance redox equations in acid | Questions 4-6 of 12

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?

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.

Optional start: Make the electrons equal before you add the two halves.

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.

B3 · Practice

Balance redox equations in acid | Questions 7-9 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.

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.

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.

B3 · Practice

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

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?

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?

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.

B4 · Predict reactions and investigate

Student notes - complete the underlined spaces | Notes 1/3

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

Name: _____ Date: _____

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 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 _____ 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 _____. The matching ions run the other way: oxidizing agents rank _____. 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 _____. 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 _____.

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 _____ and _____ 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: _____

Ion acceptors, strongest first: _____

Example results. Your class did not measure them.

B4 · Predict reactions and investigate

Student notes - complete the underlined spaces | Notes 2/3

2. What the table values mean

Energy per charge, measured from one starting point

The symbol E means _____, measured in volts. The small ° means _____: 1.0 mol/L, 25.00 °C and 101.325 kPa.

A voltage is the _____. One volt means _____, and a coulomb is a unit of electric charge. For 150 J moved by 100 C, 150 ÷ 100 = _____. Double both and 300 ÷ 200 is still _____.

Every table value is measured from the same starting point, the hydrogen half-reaction, $2H^+ + 2e^- \rightarrow H_2$, which is given _____. Each half-reaction is printed as a _____. 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 _____. 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 _____ minus E° of the _____. We subtract because the reducing agent's half runs _____. Do not just take larger minus smaller.

A positive answer means the written direction is spontaneous under standard conditions. A negative answer means _____ is spontaneous. It tells you nothing about _____. 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 _____ and the strongest oxidizing agent is _____. Choose them before you write any equation.

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

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

Student notes - complete the underlined spaces | Notes 3/3

3. Predict, then test the claim

Table values

From the Alberta Chemistry 30 Data Booklet, printed as reductions: Ag^+/Ag _____; Cu^{2+}/Cu _____; Zn^{2+}/Zn _____. 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 _____ minus E° of the _____. 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) = _____. 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) = _____. 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 _____, 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 _____. Add a _____: a test where nothing should happen, such as Cu in Cu^{2+}.

A good conclusion gives a claim, then the evidence, then an _____. Name the metal, the ion that changed and who gave the electrons. The simulation follows fixed rules, so running it again adds no new _____. 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 · Practice

Predict reactions and investigate | Questions 1-3 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?

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.

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.

B4 · Practice

Predict reactions and investigate | Questions 4-6 of 12

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.

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?

Optional start: The ion in the solution is the oxidizing agent. The metal strip is the reducing agent.

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?

B4 · Practice

Predict reactions and investigate | Questions 7-9 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.

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.

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.

B4 · Practice

Predict reactions and investigate | Questions 10-12 of 12

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.

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?

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.