B4 · Predict reactions and investigate

Completed notes

Goals: Rank reducing agents and oxidizing agents from test results. Predict which way a reaction goes under standard conditions, using the Data Booklet table. Plan a virtual investigation and say what its results can and cannot show.

Slide 1 of 11 · Balanced does not mean spontaneous

$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$ and its reverse both balance for atoms and **charge**. Balancing cannot tell which direction happens.

A direction is **spontaneous** if it happens by itself, with no outside push of energy.

Spontaneous does not mean **fast**. A spontaneous reaction can still be slow.

To choose a direction we need evidence about which substance gives electrons more **easily**.

Example. Show that $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$ and its reverse both balance. Can the balanced equation alone tell you which direction is spontaneous under standard conditions?

1. Forward: Zn 1 = 1; Cu 1 = 1; charge +2 = +2. It balances.

2. Reverse: Zn 1 = 1; Cu 1 = 1; charge +2 = +2. It also balances.

3. Both directions balance for atoms and charge, so balancing cannot choose a direction.

4. We need evidence about reducing agents and oxidizing agents, which comes next.

Try it. A reaction is spontaneous under standard conditions, but a student sees no change after one minute. Is that a contradiction? Explain.

1. No. Spontaneous means it happens by itself, with no outside push of energy.

2. It does not mean fast. A spontaneous reaction can still be slow.

Slide 2 of 11 · From observations to reducing-agent strength

Example test results: Zn in Cu^{2+} gives a copper deposit. Cu in Ag^{+} gives a silver deposit. Cu in Zn^{2+} shows **no change**.

A copper deposit means Cu^{2+} gained electrons. The electrons came from **zinc**, so zinc is the stronger reducing agent.

Zn beats Cu and Cu beats Ag, so the order of reducing agents is **Zn > Cu > Ag**. Read > as "is a stronger reducing agent than".

To keep the test fair, control the solution, temperature, strip area and **time**.

Example. Given these example test results: Zn in Cu^{2+} gives a copper deposit; Cu in Ag^{+} gives a silver deposit; Cu in Zn^{2+} gives no change. Rank Zn, Cu and Ag as reducing agents.

1. Zn gives electrons to Cu^{2+}, so Zn is a stronger reducing agent than Cu.

2. Cu gives electrons to Ag^{+}, so Cu is a stronger reducing agent than Ag.

3. Copper links the two comparisons: Zn > Cu > Ag.

4. Cu in Zn^{2+} shows no change, which agrees with this order.

Try it. Given these example test results: Fe in Cu^{2+} gives a copper deposit; Zn in Fe^{2+} gives an iron deposit. Rank Zn, Fe and Cu as reducing agents.

1. Fe gives electrons to Cu^{2+}, so Fe is a stronger reducing agent than Cu.

2. Zn gives electrons to Fe^{2+}, so Zn is a stronger reducing agent than Fe.

3. Order of reducing agents: Zn > Fe > Cu.

Slide 3 of 11 · Keep reducing agents and oxidizing agents in separate columns

The metals are reducing agents: Zn > Cu > Ag. Their ions are oxidizing agents, and they rank the **opposite** way: $Ag^+ > Cu^{2+} > Zn^{2+}$.

Write the exact formula with its **charge**, such as Ag^+, not just silver.

Written as reductions, the strongest oxidizing agent goes at the **top**. The strongest reducing agent is then at the bottom right.

List what is in the beaker: every ion, any metal, and **water**. An acidic or “acidified” solution also has $\mathbf{H^+}$. A basic solution has OH^-.

Example. For the same pairs, the reducing agents rank Zn > Cu > Ag. Rank Ag^+, Cu^{2+} and Zn^{2+} as oxidizing agents and explain why the two lists run oppositely.

1. Oxidizing agents: $Ag^+ > Cu^{2+} > Zn^{2+}$.

2. Zn gives up electrons most easily, so Zn^{2+} is least willing to take them back.

3. Ag gives up electrons least easily, so Ag^+ is most willing to take them.

Try it. Given: Fe gives electrons to Cu^{2+}, and Zn gives electrons to Fe^{2+}. Rank Zn^{2+}, Fe^{2+} and Cu^{2+} as oxidizing agents.

1. Reducing agents: Zn > Fe > Cu.

2. The ions run the opposite way: $Cu^{2+} > Fe^{2+} > Zn^{2+}$.

Slide 4 of 11 · Read the Alberta reduction table

Every table row is a **reduction**: something gains electrons. The number is E°, in **volts**.

The oxidizing agents are on the **left**. The strongest is at the top. The reducing agents are on the **right**. The strongest is at the bottom.

A higher E° means a stronger **oxidizing agent**. So Ag^+ is stronger than Cu^{2+}. E° predicts direction, not **speed**.

Standard conditions are 1.0 mol/L solutions, **25.00 °C** and 101.325 kPa for gases. The table is on Data Booklet page **7**.

Example. Given from the Alberta Data Booklet: $E°(Ag^+/Ag) = +0.80$ V, $E°(Cu^{2+}/Cu) = +0.34$ V and $E°(Zn^{2+}/Zn) = -0.76$ V. Which ion pulls electrons most strongly, and which metal is the strongest reducing agent?

1. The highest E° is +0.80 V, so Ag^+ pulls electrons most strongly.

2. The lowest E° is −0.76 V, so Zn^{2+} pulls electrons least. Its metal, Zn, is the strongest reducing agent.

3. These values are for standard conditions: 1.0 mol/L, 25.00 °C and 101.325 kPa for gases.

Try it. Given: $E°(Mg^{2+}/Mg) = -2.37$ V, $E°(Fe^{2+}/Fe) = -0.45$ V and $E°(Cu^{2+}/Cu) = +0.34$ V. Which metal is the strongest reducing agent, and which ion is the strongest oxidizing agent?

1. The lowest E° is −2.37 V (Mg^{2+}/Mg), so Mg is the strongest reducing agent.

2. The highest E° is +0.34 V (Cu^{2+}/Cu), so Cu^{2+} is the strongest oxidizing agent.

Slide 5 of 11 · What the potential comparison means

One volt is one **joule** per coulomb. A coulomb (C) is a unit of electric **charge**.

Voltage is energy per charge: 150 J ÷ 100 C = **1.50** V. Doubling both energy and charge gives the **same** voltage.

Every table value is compared with the standard hydrogen electrode, which is defined as **0.00** V.

E°reaction = E°(oxidizing agent) − E°(reducing agent). A positive value means the written direction is **spontaneous**. A negative value means the **reverse** is spontaneous.

Example. The table lists E°(Ag^+/Ag) = +0.80 V and E°(Cu^{2+}/Cu) = +0.34 V. What baseline do these numbers compare to? If Ag^+ is the oxidizing agent and Cu is the reducing agent, what is E°reaction?

1. Both values are compared with the standard hydrogen electrode, defined as 0.00 V.

2. E°reaction = E°(oxidizing agent) − E°(reducing agent) = +0.80 − (+0.34) = +0.46 V.

3. The value is positive, so the written direction is spontaneous under standard conditions.

Try it. A reaction transfers 240 J of energy for every 160 C of charge. What is the potential difference in volts? What is it if the energy and the charge both triple?

1. 240 J ÷ 160 C = 1.50 J/C, which is 1.50 V.

2. Tripled: 720 J ÷ 480 C = 1.50 J/C, which is still 1.50 V.

3. Voltage is a ratio, so scaling both does not change it.

Slide 6 of 11 · Complete prediction: copper with silver ions

Ag^+ is the oxidizing agent. Cu is the reducing agent, so read its row **backwards**: Cu → Cu^{2+} + $2e^-$. In the table, Ag^+ on the left is **above** Cu on the right, so the reaction is spontaneous.

Multiply the silver half by **2**, add and cancel the $2e^-$: Cu + $2Ag^+$ → Cu^{2+} + 2Ag.

E°reaction = E°(oxidizing agent) − E°(reducing agent) = +0.80 − (+0.34) = **+0.46** V. Positive means the written direction is **spontaneous**.

Do not multiply E° by a coefficient. Predict a silver **coating**, a shrinking copper strip and a **blue** solution from Cu^{2+} (Data Booklet colours, printed p. 11).

Example. Given Cu(s) and Ag^+(aq) under standard conditions, with E°(Ag^+/Ag) = +0.80 V and E°(Cu^{2+}/Cu) = +0.34 V. Predict whether the reaction is spontaneous, write the net ionic equation and predict what you would see.

1. Ag^+ is the oxidizing agent: $Ag^+ + e^- \rightarrow Ag$. Cu is the reducing agent: Cu → Cu^{2+} + $2e^-$.

2. Multiply the silver half by 2 and add: Cu + $2Ag^+$ → Cu^{2+} + 2Ag.

3. E°reaction = +0.80 − (+0.34) = +0.46 V. It is positive, so it is spontaneous.

4. Check: charge +2 = +2; atoms Cu 1 and Ag 2 on each side.

5. Predictions: silver coating forms, the copper strip shrinks and Cu^{2+} makes the solution blue.

Try it. Given Mg(s) and Cu^{2+}(aq) under standard conditions, with E°(Cu^{2+}/Cu) = +0.34 V and E°(Mg^{2+}/Mg) = −2.37 V. Predict whether the reaction is spontaneous and write the net ionic equation.

1. Cu^{2+} is the oxidizing agent: $Cu^{2+} + 2e^- \rightarrow Cu$. Mg is the reducing agent: Mg → Mg^{2+} + $2e^-$.

2. Both halves use $2e^-$, so add: Mg + Cu^{2+} → Mg^{2+} + Cu.

3. E°reaction = +0.34 − (−2.37) = +2.71 V. It is positive, so it is spontaneous.

4. Check: charge +2 = +2; atoms Mg 1 and Cu 1 on each side.

Slide 7 of 11 · Method: how to predict a redox reaction with the table

Step 1. **List** what is in the beaker: every ion, any metal, and water. Add H^+ if it is acidic and OH^- if it is basic.

Step 2. Find the strongest **oxidizing agent** (highest on the left) and the strongest **reducing agent** (lowest on the right).

Step 3. Write the two halves. **Reverse** the reducing agent's row so it is an oxidation.

Step 4. Calculate $E°_{reaction} = E°(\text{oxidizing agent}) - E°(\text{reducing agent})$.

Step 5. A **positive** $E°_{reaction}$ is spontaneous. A negative $E°_{reaction}$ is not spontaneous.

Check: never **multiply** $E°$ by a coefficient.

Example. Use the method on Cu(s) and $Ag^+(aq)$ with $E°(Ag^+/Ag) = +0.80$ V and $E°(Cu^{2+}/Cu) = +0.34$ V. Is the reaction spontaneous?

1. Step 1: the beaker has Cu(s) and $Ag^+(aq)$.

2. Step 2: Ag^+ is the oxidizing agent. Cu is the reducing agent. Ag^+ is above Cu.

3. Step 3: $Ag^+ + e^- \rightarrow Ag$ and $Cu \rightarrow Cu^{2+} + 2e^-$.

4. Steps 4 and 5: $E°_{reaction} = +0.80 - (+0.34) = +0.46$ V. It is positive, so $Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$ is spontaneous.

Try it. Use the method on Zn(s) and $Cu^{2+}(aq)$ with $E°(Cu^{2+}/Cu) = +0.34$ V and $E°(Zn^{2+}/Zn) = -0.76$ V. Is the reaction spontaneous?

1. Cu^{2+} is the oxidizing agent. Zn is the reducing agent. Cu^{2+} is above Zn in the table.

2. $E°_{reaction} = +0.34 - (-0.76) = +1.10$ V.

3. It is positive, so $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$ is spontaneous.

Slide 8 of 11 · Select partners from a mixture before calculating

First list what is in the **beaker**. Then pick the strongest oxidizing agent and the strongest reducing agent only from that list.

Ag^+ has the higher E° (+0.80 V against +0.34 V), so it is the stronger **oxidizing agent**. Zn^{2+} has the lowest E°, so Zn is the stronger **reducing agent**.

$Zn + 2Ag^+ \rightarrow Zn^{2+} + 2Ag$. E°reaction = +0.80 − (−0.76) = **+1.56** V. Subtracting a negative is the same as **adding**.

Cu and Cu^{2+} are in the mixture but not in the net equation. We do not **force** every substance into the equation.

Example. Given a standard mixture of Zn(s), Cu(s), Ag^+(aq) and Cu^{2+}(aq), with E°(Ag^+/Ag) = +0.80 V, E°(Cu^{2+}/Cu) = +0.34 V and E°(Zn^{2+}/Zn) = −0.76 V. Select the strongest reducing agent and oxidizing agent, write the net ionic equation and find E°reaction.

1. Oxidizing agents present: Ag^+ (+0.80 V) and Cu^{2+} (+0.34 V). The higher E° is Ag^+, the strongest oxidizing agent.

2. Reducing agents present: Zn and Cu. Zn^{2+}/Zn has the lower E° (−0.76 V), so Zn is the strongest reducing agent.

3. $Zn \rightarrow Zn^{2+} + 2e^-$ and $2Ag^+ + 2e^- \rightarrow 2Ag$ give $Zn + 2Ag^+ \rightarrow Zn^{2+} + 2Ag$.

4. E°reaction = +0.80 − (−0.76) = +1.56 V, so it is spontaneous.

5. Check: charge +2 = +2; Zn 1 and Ag 2 on each side.

Try it. Given a standard mixture of Fe(s), Cu(s), Ag^+(aq) and Cu^{2+}(aq), with E°(Ag^+/Ag) = +0.80 V, E°(Cu^{2+}/Cu) = +0.34 V and E°(Fe^{2+}/Fe) = −0.45 V. Select the strongest reducing agent and oxidizing agent, write the net ionic equation and find E°reaction.

1. Oxidizing agents present: Ag^+ (+0.80 V) and Cu^{2+} (+0.34 V). Ag^+ has the higher E°, so it is the strongest oxidizing agent.

2. Reducing agents present: Fe and Cu. Fe^{2+}/Fe has the lower E° (−0.45 V), so Fe is the strongest reducing agent.

3. $Fe \rightarrow Fe^{2+} + 2e^-$ and $2Ag^+ + 2e^- \rightarrow 2Ag$ give $Fe + 2Ag^+ \rightarrow Fe^{2+} + 2Ag$.

4. E°reaction = +0.80 − (−0.45) = +1.25 V, so it is spontaneous.

5. Check: charge +2 = +2; Fe 1 and Ag 2 on each side.

Slide 9 of 11 · Your turn: copper with zinc ions

Write the roles for the direction as written. Cu is the **reducing agent** and Zn^{2+} is the **oxidizing agent**.

$E°_{reaction} = E°(\text{oxidizing agent}) - E°(\text{reducing agent}) = -0.76 - (+0.34) =$ **−1.10** V.

A negative value means the written direction is **not spontaneous**. The reverse, Zn with Cu^{2+}, gives +1.10 V and is spontaneous.

In the table, Zn^{2+} is below Cu, so nothing happens. A copper **container** can safely hold a Zn^{2+} solution.

Example. Given Cu(s) and $Zn^{2+}(aq)$, with $E°(Cu^{2+}/Cu) = +0.34$ V and $E°(Zn^{2+}/Zn) = -0.76$ V. Is $Cu + Zn^{2+} \rightarrow Cu^{2+} + Zn$ spontaneous under standard conditions? Justify with the numbers.

1. As written, Cu is the reducing agent and Zn^{2+} is the oxidizing agent.

2. $E°_{reaction} = E°(\text{oxidizing agent}) - E°(\text{reducing agent}) = -0.76 - (+0.34) = -1.10$ V.

3. The value is negative, so the reaction is not spontaneous as written.

4. The reverse, Zn with Cu^{2+}, gives +1.10 V and is spontaneous.

Try it. Given Cu(s) and $Fe^{2+}(aq)$, with $E°(Cu^{2+}/Cu) = +0.34$ V and $E°(Fe^{2+}/Fe) = -0.45$ V. Is $Cu + Fe^{2+} \rightarrow Cu^{2+} + Fe$ spontaneous under standard conditions? Justify with the numbers.

1. As written, Cu is the reducing agent and Fe^{2+} is the oxidizing agent.

2. $E°_{reaction} = -0.45 - (+0.34) = -0.79$ V.

3. The value is negative, so the reaction is not spontaneous as written.

Slide 10 of 11 · Plan a virtual metal–ion investigation

A fair test changes **one factor** at a time. Compare different metals in the same ion solution.

Keep equal: clean strips of equal **area**, concentration, time and temperature. Add a control, such as Cu in Cu^{2+}.

Record the deposit, metal loss and solution appearance. Also record **no change**, because it is a result.

Before each run make a **prediction** with a reason. The simulation is a fixed model, not real **data**. Running it again adds no new **evidence**.

Example. Question: which of Zn, Cu and Ag is the strongest reducing agent? Plan a fair investigation. State the factor you change, three factors you keep the same and what you record.

1. Change: the identity of the metal (Zn, Cu, Ag).

2. Keep the same: the ion solution and concentration, clean strips of equal area, time and temperature.

3. Control: the same metal and ion, such as Cu in Cu^{2+}.

4. Record: deposit, metal loss, solution appearance and any no-change results.

Try it. A student puts Zn in Cu^{2+} for 5 s and Ag in Zn^{2+} for 60 s. Why is this not a fair comparison, and how can it be fixed?

1. The metal, the solution and the time all differ, so any difference cannot be blamed on one factor.

2. Fix: put each metal in the same ion solution for the same time.

Slide 11 of 11 · Make an evidence-based claim

A good conclusion has a **claim**, then evidence, then an electron explanation.

As reducing agents, Zn > Cu > Ag. The ions rank **Ag^{+} > Cu^{2+} > Zn^{2+}** as oxidizing agents.

Zn reacts with Cu^{2+}, so Zn gives electrons to Cu^{2+}. Cu reacts with Ag^{+}, so Cu is a stronger reducing agent than **Ag**.

No visible change does not **prove** a reaction is impossible. A limitation can be an oxide coating or a reaction too **slow** to see.

Example. Given these example results: Zn reacts with Cu^{2+} and Ag^{+}; Cu reacts with Ag^{+}; Ag reacts with neither Zn^{2+} nor Cu^{2+}. State a ranking, cite two comparisons and name one limitation.

1. Claim: as reducing agents, Zn > Cu > Ag.

2. Evidence 1: Zn reacts with Cu^{2+}, so Zn gave electrons to Cu^{2+} and is the stronger reducing agent.

3. Evidence 2: Cu reacts with Ag^{+}, so Cu is a stronger reducing agent than Ag. This extends the order.

4. Limitation: an oxide coating or a slow reaction could hide a change, so no visible change does not prove impossibility.

Try it. Given these example results: Zn reacts with Fe^{2+}; Fe reacts with Cu^{2+}; Cu reacts with neither Fe^{2+} nor Zn^{2+}. State a ranking, cite two comparisons and name one limitation.

1. Claim: as reducing agents, Zn > Fe > Cu.

2. Evidence 1: Zn reacts with Fe^{2+}, so Zn gave electrons to Fe^{2+} and is a stronger reducing agent than Fe.

3. Evidence 2: Fe reacts with Cu^{2+}, so Fe is a stronger reducing agent than Cu.

4. Limitation: no visible change for Cu does not prove it is impossible, because a slow reaction could be missed.

B4 · Worked practice

Predict reactions and investigate | Questions 1-2 of 8

1. Given Zn(s), $Ag^+(aq)$, $E°(Ag^+/Ag) = +0.80$ V and $E°(Zn^{2+}/Zn) = -0.76$ V from the Alberta table. Predict which way the reaction goes under standard conditions. Write the balanced equation.

Use Method: how to predict a redox reaction with the table. Ag^+ is the oxidizing agent and Zn is the reducing agent. Ag^+ is above Zn in the table. $+0.80 - (-0.76) = +1.56$ V, so $Zn + 2Ag^+ \rightarrow Zn^{2+} + 2Ag$ is spontaneous. The charge is +2 on each side and the atoms match.

2. Given Ag(s), $Cu^{2+}(aq)$, $E°(Ag^+/Ag) = +0.80$ V and $E°(Cu^{2+}/Cu) = +0.34$ V from the Alberta table, decide whether $2Ag + Cu^{2+} \rightarrow 2Ag^+ + Cu$ is spontaneous under standard conditions. Show the E° values.

Use Method: how to predict a redox reaction with the table. Cu^{2+} would be the oxidizing agent and Ag the reducing agent. Cu^{2+} is below Ag in the table. $+0.34 - (+0.80) = -0.46$ V. The reaction as written is not spontaneous under standard conditions.

B4 · Worked practice

Predict reactions and investigate | Questions 3-4 of 8

3. Given test results for three unknown metals: metal P reacts with Q^{2+} and gets a coating of Q. Metal Q reacts with R^{2+} and gets a coating of R. Each metal forms one ion: P^{2+}, Q^{2+} or R^{2+}. The conditions are the same in each test. Rank the reducing agents and the oxidizing agents.

Reducing agents: P > Q > R. Oxidizing agents: $R^{2+} > Q^{2+} > P^{2+}$. P gives electrons more easily than Q, and Q more easily than R. The matching ions rank the other way.

4. A trial shows no visible deposit after ten seconds. Does this alone prove that the reaction cannot happen under standard conditions? Explain.

No. A slow reaction, a coating on the metal or a change too faint to see can hide it. Check the conditions, the metal surface and the time, and compare with the table prediction. No change seen is weak evidence.

B4 · Worked practice

Predict reactions and investigate | Questions 5-6 of 8

5. Plan a fair test to compare Zn(s) and Ag(s) as reducing agents with Cu^{2+}. Name the one thing you change, two things you keep the same and one thing you record.

Change the metal. Keep the Cu^{2+} concentration, the volume of solution, the temperature, the size of the clean metal strip and the time the same. Record any coating and any loss of metal. The blue colour of Cu^{2+} may fade. The table predicts that Zn reacts and Ag does not.

6. Given the table pair $Cu^{2+} + 2e^- \rightarrow Cu$ with $E° = +0.34$ V, a student triples it and reports +1.02 V. Explain the error.

Triple the coefficients to match the electrons, but leave E° at +0.34 V. A volt is energy for each unit of charge. Tripling both the energy and the charge does not change it.

B4 · Worked practice

Predict reactions and investigate | Questions 7-8 of 8

7. Given the Data Booklet table values $E°(Ag^+/Ag) = +0.80$ V, $E°(Cu^{2+}/Cu) = +0.34$ V and $E°(Zn^{2+}/Zn) = -0.76$ V. A solution that contains $Cu^{2+}(aq)$ must be stored. Can it be kept in a zinc container? Can it be kept in a silver container? Explain with the table.

Use Method: how to predict a redox reaction with the table. Cu^{2+} is the oxidizing agent. The metal of the container is the reducing agent. Zinc: Cu^{2+} is above Zn in the table. $E°_{reaction} = +0.34 - (-0.76) = +1.10$ V. It is spontaneous, so the zinc container reacts and is eaten away. Do not use zinc. Silver: Cu^{2+} is below Ag in the table. $E°_{reaction} = +0.34 - (+0.80) = -0.46$ V. It is not spontaneous, so the silver container is safe.

8. Challenge: Given a solution that contains both $Ag^+(aq)$ and $Cu^{2+}(aq)$. A strip of nickel is placed in it. Data Booklet table values: $E°(Ag^+/Ag) = +0.80$ V, $E°(Cu^{2+}/Cu) = +0.34$ V and $E°(Ni^{2+}/Ni) = -0.26$ V. Which ion reacts first? Write the balanced equation and find $E°_{reaction}$ in volts.

List what is in the beaker: Ag^+, Cu^{2+} and Ni(s). The strongest oxidizing agent is the ion with the highest table value, Ag^+. The only reducing agent is Ni. So Ag^+ reacts first: $2Ag^+(aq) + Ni(s) \rightarrow 2Ag(s) + Ni^{2+}(aq)$. $E°_{reaction} = E°_{(oxidizing\ agent)} - E°_{(reducing\ agent)} = +0.80 - (-0.26) = +1.06$ V. It is positive, so the reaction is spontaneous. The value is not multiplied by 2.