B7 · Reduction potentials and the hydrogen reference

Student notes - follow the slides, fill the blanks, work the questions

Goals: Explain why every potential is measured against the standard hydrogen reference. Use the table to build a cell from any two half-cells, with inert electrodes when needed. Calculate E°cell, predict observations and say what the sign does and does not tell us.

Name: _____ Date: _____

Slide 1 of 8 · One reference point for every half-cell

A voltmeter reads the push between _____ half-cells, so chemists chose one half-cell as the reference and set its value to exactly _____ V.

The standard hydrogen half-cell is $2H^+ + 2e^- \rightarrow H_2$, with H^+ at _____ mol/L and H_2 gas at _____ kPa.

Electrons flow away from the _____ and toward the cathode. Use E°cell = E°cathode − E°anode and solve for the unknown half-cell value.

When the reference is the cathode, the unknown value is _____. When the reference is the anode, the unknown value is _____.

Example. A zinc half-cell is joined to the standard hydrogen half-cell. The voltmeter reads 0.76 V and electrons flow from the zinc half-cell toward the hydrogen half-cell. Find E°(Zn^{2+}/Zn).

Work:

Answer: _____

Try it. A copper half-cell is joined to the standard hydrogen half-cell. The voltmeter reads 0.34 V and electrons flow toward the copper half-cell. Find E°(Cu^{2+}/Cu).

Work:

Answer: _____

Sketch (draw and label the diagram):

Slide 2 of 8 · Reading the table: all values are reductions

Every row in the table is written as a _____, with the electrons on the left side.

The strongest oxidizing agent is on the _____ side and has the _____ value.

The strongest reducing agent is on the _____ side and has the _____ value.

If something is oxidized in a cell, still use its reduction row value as _____. Flipping the sign and then subtracting counts the sign change _____.

Example. Given $Ag^+ + e^- \rightarrow Ag$ (+0.80 V), $Cu^{2+} + 2e^- \rightarrow Cu$ (+0.34 V) and $Zn^{2+} + 2e^- \rightarrow Zn$ (−0.76 V). Which of these is the strongest oxidizing agent and which is the strongest reducing agent? Rank each group.

Work:

Answer: _____

Try it. Given $Br_2 + 2e^- \rightarrow 2Br^-$ (+1.07 V), $Fe^{3+} + e^- \rightarrow Fe^{2+}$ (+0.77 V) and $I_2 + 2e^- \rightarrow 2I^-$ (+0.54 V). Name the strongest oxidizing agent and the strongest reducing agent among all six.

Work:

Answer: _____

Slide 3 of 8 · A cell from any two half-cells

Compare the two table values. The half with the _____ value is reduced at the cathode. The other half runs _____ as the oxidation at the anode.

The anode half $2I^- \rightarrow I_2 + 2e^-$ gives _____ electrons, and each Fe^{3+} takes _____, so multiply the iron half by 2.

The cell reaction is _____. Check the charge: both sides total +4.

$E°_{cell} = E°_{cathode} - E°_{anode}$ = _____ V. E° is energy for each unit of charge, so it is _____ multiplied when a half-reaction is.

Example. Given $Fe^{3+} + e^- \rightarrow Fe^{2+}$ (+0.77 V) and $I_2 + 2e^- \rightarrow 2I^-$ (+0.54 V). Build the cell reaction and find E°cell.

Work:

Answer: _____

Try it. Given $Fe^{3+} + e^- \rightarrow Fe^{2+}$ (+0.77 V) and $Cu^{2+} + 2e^- \rightarrow Cu$ (+0.34 V). Build the cell reaction in which Fe^{3+} is reduced and copper is oxidized, and find E°cell.

Work:

Answer: _____

Slide 4 of 8 · Method: how to find E°cell from any two half-cells

Step 1. Write the two table _____. Each one is E° for its pair.

Step 2. The _____ E° is the anode, so that half is oxidized.

Step 3. The _____ E° is the cathode, so that half is reduced.

Step 4. E°cell = E°_____ − E°anode. Do not multiply by _____.

Step 5. A _____ E°cell means the cell runs by itself.

Example. Use the method on $Fe^{3+} + e^- \rightarrow Fe^{2+}$ (+0.77 V) and $I_2 + 2e^- \rightarrow 2I^-$ (+0.54 V). Find E°cell.

Work:

Answer: _____

Try it. Use the method on $Br_2 + 2e^- \rightarrow 2Br^-$ (+1.07 V) and $Fe^{3+} + e^- \rightarrow Fe^{2+}$ (+0.77 V). Find E°cell.

Work:

Answer: _____

Slide 5 of 8 · Inert electrodes and notation

An _____ electrode does not react. It only gives electrons a surface to enter or leave. Use one when a half-cell has no solid _____.

Carbon is written _____ in cell notation. Iron metal would not work, because it could _____.

In cell notation, a _____ separates ions in the same solution. The anode goes on the far left.

Each Fe^{2+} gives _____ electron and each MnO_4^- takes _____, so multiply the iron half by 5. $E°cell = 1.51 - 0.77 =$ _____ V.

Example. Given $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ (+1.51 V) and $Fe^{3+} + e^- \rightarrow Fe^{2+}$ (+0.77 V). Neither half-cell has a solid metal. Name the electrode, write the cell notation and the overall reaction, and find E°cell.

Work:

Answer: _____

Try it. Given $Fe^{3+} + e^- \rightarrow Fe^{2+}$ (+0.77 V) and $Cu^{2+} + 2e^- \rightarrow Cu$ (+0.34 V). Write the cell notation for the cell in which Fe^{3+} is reduced. Say which electrode is carbon, and find E°cell.

Work:

Answer: _____

Sketch (draw and label the diagram):

Slide 6 of 8 · Your turn: will bromine oxidize iron(II)?

To oxidize something, an oxidizing agent must itself be _____, because an oxidizing agent takes electrons.

Br_2 (+1.07 V) has a _____ value than Fe^{3+}/Fe^{2+} (+0.77 V), so Br_2 is reduced at the cathode.

The iron half gives 1 electron and bromine takes 2, so multiply the iron half by _____. Overall: $Br_2 + 2Fe^{2+} \rightarrow$ _____.

$E°cell = 1.07 - 0.77 =$ _____ V. A positive sign says the reaction is _____. It predicts direction only, not _____.

Example. Given $Br_2 + 2e^- \rightarrow 2Br^-$ (+1.07 V) and $Fe^{3+} + e^- \rightarrow Fe^{2+}$ (+0.77 V). Does Br_2 oxidize Fe^{2+} to Fe^{3+} under standard conditions? Write the cell reaction and find E°cell.

Work:

Answer: _____

Try it. Given $I_2 + 2e^- \rightarrow 2I^-$ (+0.54 V) and $Fe^{3+} + e^- \rightarrow Fe^{2+}$ (+0.77 V). Does I_2 oxidize Fe^{2+} to Fe^{3+} under standard conditions? Write the reaction and find E°cell for it as written.

Work:

Answer: _____

Slide 7 of 8 · Predict what you would see in a cell

At the cathode, purple MnO_4^- is reduced to Mn^{2+}, which is _____ at 0.010 mol/L (pale pink at 1.0 mol/L), so the purple fades.

$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ uses up H^+, so the pH _____. Less H^+ means a higher pH.

At the anode, $Fe^{2+} \rightarrow Fe^{3+} + e^-$. The Data Booklet (p. 11) lists iron(III) as _____, so the solution turns that colour.

Carbon electrodes are _____, so their mass does not change. A mass change needs a _____ that is made or used up.

Example. The cell C(s) | Fe^{2+}(aq), Fe^{3+}(aq) || MnO_4^-(aq), H^+(aq), Mn^{2+}(aq) | C(s) is running. Predict the colour change at each electrode, the pH change near the cathode and the mass change of the electrodes.

Work:

Answer: _____

Try it. The cell Cu(s) | Cu^{2+}(aq) || Fe^{3+}(aq), Fe^{2+}(aq) | C(s) is running. Predict the mass change of each electrode and the change in the blue colour of the copper solution.

Work:

Answer: _____

Slide 8 of 8 · What E° tells us, and what it does not

In $2Ag + Cu^{2+} \rightarrow 2Ag^{+} + Cu$, Cu^{2+} is reduced, so Cu^{2+}/Cu (+0.34 V) is the _____ and Ag^{+}/Ag (+0.80 V) is the anode.

$E°cell = 0.34 - 0.80 =$ _____ V. A negative value means the reaction is _____ as written under standard conditions.

The reverse reaction, $Cu + 2Ag^{+} \rightarrow Cu^{2+} + 2Ag$, has E°cell = _____ V, so it is spontaneous.

E° gives the spontaneous direction under _____ conditions. It does not give _____, and it does not give voltages at other concentrations.

Example. Given $E°(Ag^{+}/Ag) = +0.80$ V and $E°(Cu^{2+}/Cu) = +0.34$ V. A student studies $2Ag + Cu^{2+} \rightarrow 2Ag^{+} + Cu$. Calculate E°cell for the reaction as written. What does the sign tell you, and which reaction is spontaneous?

Work:

Answer: _____

Try it. Given $Fe^{3+} + e^{-} \rightarrow Fe^{2+}$ (+0.77 V) and $Zn^{2+} + 2e^{-} \rightarrow Zn$ (−0.76 V). A student studies $Zn^{2+} + 2Fe^{2+} \rightarrow Zn + 2Fe^{3+}$. Calculate E°cell as written. Is the reaction spontaneous, and does the sign tell us its speed?

Work:

Answer: _____

B7 · Practice

Reduction potentials and the hydrogen reference | Questions 1-2 of 7

1. A cell is built from the standard hydrogen half-cell and a nickel half-cell. The voltmeter reads 0.26 V and electrons flow from the nickel electrode toward the hydrogen half-cell. Find E° for $Ni^{2+} + 2e^- \rightarrow Ni$.

Work:

Final answer (with unit): _____

2. Given these Data Booklet table rows: $Ag^+ + e^- \rightarrow Ag$ (+0.80 V), $Cu^{2+} + 2e^- \rightarrow Cu$ (+0.34 V) and $Zn^{2+} + 2e^- \rightarrow Zn$ (−0.76 V). Which of these is the strongest reducing agent, and which is the strongest oxidizing agent?

Work:

Final answer (with unit): _____

B7 · Practice

Reduction potentials and the hydrogen reference | Questions 3-4 of 7

3. Given these Data Booklet table rows: $I_2 + 2e^- \rightarrow 2I^-$ (+0.54 V) and $Cl_2 + 2e^- \rightarrow 2Cl^-$ (+1.36 V). Write the cell reaction for the cell in which chlorine oxidizes iodide and calculate E°cell.

Work:

Final answer (with unit): _____

4. Write the cell notation for the cell in which MnO_4^- (acidic) oxidizes Fe^{2+}, using carbon electrodes. Name the anode and give E°cell (Data Booklet table values: MnO_4^-/Mn^{2+} +1.51 V; Fe^{3+}/Fe^{2+} +0.77 V).

Work:

Final answer (with unit): _____

B7 · Practice

Reduction potentials and the hydrogen reference | Questions 5-6 of 7

5. Given these Data Booklet table rows: $Cl_2 + 2e^- \rightarrow 2Cl^-$ (+1.36 V) and $Br_2 + 2e^- \rightarrow 2Br^-$ (+1.07 V). Will Cl_2 oxidize Br^- to Br_2 under standard conditions? Write the equation and $E°_{cell}$.

Work:

Final answer (with unit): _____

6. In the cell C(s) | $Fe^{2+}(aq)$, $Fe^{3+}(aq)$ || $MnO_4^-(aq)$, $H^+(aq)$, $Mn^{2+}(aq)$ | C(s), predict the change in pH and in colour near the cathode. Explain with a half-reaction.

Work:

Final answer (with unit): _____

B7 · Practice

Reduction potentials and the hydrogen reference | Questions 7-7 of 7

7. Challenge: Given the Data Booklet table values $E°(Zn^{2+}/Zn) = -0.76$ V and $E°(Cu^{2+}/Cu) = +0.34$ V. Suppose chemists had chosen the zinc half-cell as the zero reference instead of hydrogen. What would $E°(Cu^{2+}/Cu)$ be on that new scale, in volts? Would $E°_{cell}$ for the zinc and copper cell change? Explain.

Work:

Final answer (with unit): _____