B7 · Reduction potentials and the hydrogen reference

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

Use the hydrogen reference and the reduction table to build a cell from any two half-cells, write its notation, predict observations and say what E° means.

Name: _____ Date: _____

1. The hydrogen reference and the table

One reference for every half-cell

A voltmeter reads the push between _____ half-cells, so one half-cell alone can't have a measured value.

So chemists picked the standard hydrogen half-cell ($2H^+ + 2e^- \rightarrow H_2$, 1.0 mol/L, H_2 gas at 101.325 kPa) as the reference. They set it to exactly _____, like sea level for heights.

Join zinc to it and the meter reads 0.76 V, with electrons flowing from zinc to hydrogen. Electrons leave the anode, so zinc is the _____.

Use $E°cell = E°cathode - E°anode$. That gives $0.76 = 0.00 - E°(Zn^{2+}/Zn)$, so $E°(Zn^{2+}/Zn)$ = _____. Negative means zinc ions are poorer than hydrogen ions at gaining electrons.

Now join copper. The meter reads 0.34 V, but electrons flow toward copper. Electrons go toward the cathode, so copper is the _____.

So $0.34 = E°(Cu^{2+}/Cu) - 0.00$ and $E°(Cu^{2+}/Cu)$ = _____. Positive means copper ions are better than hydrogen ions at gaining electrons.

On the diploma: electron flow tells you which half is the anode, and that decides the sign. Name the anode before you write the equation.

Read the table

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

A higher value means the particle on the left gains electrons better, so it is a stronger _____. A lower value means the particle on the right gives electrons better, so it is a stronger reducing agent.

Ag^+ (+0.80 V) is a stronger oxidizing agent than Cu^{2+} (+0.34 V). Zn is the strongest reducing agent of the three, because Zn^{2+}/Zn is lowest at −0.76 V. You built this same ranking from experiments in B4.

If something is oxidized in a cell, we still use its printed _____ value and we never change its _____.

The subtraction in E°cell = E°cathode − E°anode already handles the sign. If you flip the sign first, you count it twice.

On the diploma: the strongest oxidizing agent has the highest value on the left. The strongest reducing agent has the lowest value on the right. Flipping the anode's sign is a common wrong option.

Reference: $2H^+ + 2e^- \rightarrow H_2$ is set at _____

Zn half-cell joined to the reference reads 0.76 V and Zn is the anode: E° = _____

Fe^{3+}/Fe^{2+} +0.77 V cathode; I_2/I^- +0.54 V anode: E°cell = _____

E°cell = E°cathode − E°anode. Do you multiply E° by the electron count? _____

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2. Cells from any two half-cells

Choose the cathode and the anode

Given $Fe^{3+} + e^- \rightarrow Fe^{2+}$ (+0.77 V) and $I_2 + 2e^- \rightarrow 2I^-$ (+0.54 V). It is the same rule as the Zn / Cu cell: the half with the _____ value is reduced at the cathode.

So _____ is reduced and _____ is oxidized. The iodine row runs backward at the anode.

Each I_2 made gives 2 electrons, and each Fe^{3+} takes only 1. So multiply the iron half by _____ to match the electrons.

Add the halves and cancel the electrons. The cell reaction is $2Fe^{3+} + 2I^- \rightarrow$ _____. Check the charge: +6 − 2 = +4 on the left and +4 on the right.

$E°_{cell}$ = _____ − (+0.54) = _____. We did not double 0.77 V. A potential is a push per electron, not a total, so never _____ a table value by a coefficient.

On the diploma: write the two values side by side before anything else, and count electrons before you balance. Multiplying E° by a coefficient is a typical wrong option.

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

Inert electrodes and notation

In B6 both halves had a solid metal to connect the wire. The Fe^{3+}/Fe^{2+} half-cell has only dissolved ions, so it needs an _____ electrode, which means one that does not react.

_____ is the usual choice, written C(s). It only gives electrons a surface to enter or leave. Iron metal would not work, because it could react itself.

Take MnO_4^-/Mn^{2+} (+1.51 V) and Fe^{3+}/Fe^{2+} (+0.77 V). The higher value is the cathode, so permanganate is reduced and the iron half is the anode.

C(s) | Fe^{2+}(aq), Fe^{3+}(aq) || MnO_4^-(aq), H^+(aq), Mn^{2+}(aq) | C(s). The _____ is on the far left, and a _____ separates ions in one solution. H^+ is there because the permanganate half needs it.

The reduction is $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$. Each Fe^{2+} gives 1 electron and each MnO_4^- takes 5, so multiply the iron half by 5.

$E°_{cell}$ = 1.51 − 0.77 = _____. This is the same reaction as the titration of iron(II) with permanganate.

On the diploma: put the anode on the left, the double line in the middle, and each dissolved ion in its own solution. A half-cell with no solid metal needs a carbon electrode.

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3. Predict and interpret

Predict observations

We predict what we would see the same way as in B6. Read each half-reaction and ask what is used up and what is made.

At the cathode, purple MnO_4^- is reduced to almost colourless _____, so the purple fades. The half-reaction has $8H^+$ on the left, so H^+ is used up and the pH _____.

At the anode, lime green Fe^{2+} is oxidized to Fe^{3+}, and the solution turns _____.

Carbon is inert, so it is in no half-reaction and its _____ does not change. A mass change only happens if a solid is made or used up.

On the diploma: pH goes up when H^+ is used up and down when H^+ is made. Say which one the half-reaction does.

What the sign of E° means

Try $2Ag + Cu^{2+} \rightarrow 2Ag^+ + Cu$. Cu^{2+} is reduced, so Cu^{2+}/Cu (+0.34 V) is the cathode. Silver is oxidized, so Ag^+/Ag (+0.80 V) is the anode.

$E°_{cell} = 0.34 - 0.80 =$ _____. The reaction is not _____ as written. The reverse, $Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$, has E° = +0.46 V, so it is spontaneous.

E° predicts the spontaneous _____ under standard conditions, meaning 1.0 mol/L and 25 °C. It does not predict _____, because a spontaneous reaction can still be very slow.

E° also says nothing about other concentrations. The Nernst equation is _____ part of this course.

On the diploma: if $E°_{cell}$ is negative, say the reaction is not spontaneous as written and give the reverse reaction. A claim that positive E° means fast is a wrong option.

B7 · Practice

Reduction potentials and the hydrogen reference | 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.

B7-P01 · Easy · A silver half-cell is joined to the standard hydrogen half-cell and the voltmeter reads 0.80 V. Electrons flow toward the silver electrode. Find E° for $Ag^+ + e^- \rightarrow Ag$ and name the anode.

B7-P02 · Easy · Explain why a table cannot list a measured voltage for a single half-cell, and what the standard hydrogen half-cell is used for.

B7-P03 · Easy · Given $Cl_2 + 2e^- \rightarrow 2Cl^-$ (+1.36 V), $Br_2 + 2e^- \rightarrow 2Br^-$ (+1.07 V) and $I_2 + 2e^- \rightarrow 2I^-$ (+0.54 V). Name the strongest oxidizing agent and the strongest reducing agent among Cl_2, Br_2, I_2, Cl^-, Br^- and I^-.

B7 · Practice

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B7-P04 · Easy · A student reverses the zinc row to $Zn \rightarrow Zn^{2+} + 2e^-$ and uses +0.76 V for the cell calculation. Which value should be used, and why?

B7-P05 · Medium · Given $Fe^{3+} + e^- \rightarrow Fe^{2+}$ (+0.77 V) and $I_2 + 2e^- \rightarrow 2I^-$ (+0.54 V). Write the cell reaction and calculate E°cell.

B7-P06 · Medium · Given $Cl_2 + 2e^- \rightarrow 2Cl^-$ (+1.36 V) and $I_2 + 2e^- \rightarrow 2I^-$ (+0.54 V). Write the cell reaction in which chlorine oxidizes iodide and calculate E°cell.

B7 · Practice

Reduction potentials and the hydrogen reference | Questions 7-9 of 12

B7-P07 · Medium · In the cell $2Fe^{3+} + 2I^- \rightarrow 2Fe^{2+} + I_2$, a student multiplies 0.77 V by 2 because the iron half was doubled. Explain the error.

B7-P08 · Medium · Why does the Fe^{3+}/Fe^{2+} half-cell need an inert electrode, and why is carbon a suitable choice?

B7-P09 · Challenging · Write the cell notation for the cell in which acidified MnO_4^- oxidizes Fe^{2+}, using carbon electrodes. Name the anode and give E°cell (MnO_4^-/Mn^{2+} +1.51 V; Fe^{3+}/Fe^{2+} +0.77 V).

B7 · Practice

Reduction potentials and the hydrogen reference | Questions 10-12 of 12

B7-P10 · Challenging · Given $Br_2 + 2e^- \rightarrow 2Br^-$ (+1.07 V) and $Fe^{3+} + e^- \rightarrow Fe^{2+}$ (+0.77 V). Will Br_2 oxidize Fe^{2+} under standard conditions? Write the equation and E°cell.

B7-P11 · Challenging · In the cell C(s) | $Fe^{2+}(aq)$, $Fe^{3+}(aq)$ || $MnO_4^-(aq)$, $H^+(aq)$, $Mn^{2+}(aq)$ | C(s), predict three things. Give the pH change near the cathode, the colour change at each electrode and the mass change of the electrodes.

B7-P12 · Challenging · A student calculates E°cell = −0.46 V for $2Ag + Cu^{2+} \rightarrow 2Ag^+ + Cu$ and says the reaction is impossible and very fast in reverse. Explain what the sign does and does not tell us.