# 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? _____

# B7 · Reduction potentials and the hydrogen reference

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

## 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.

# B7 · Reduction potentials and the hydrogen reference

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

## 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

Reduction potentials and the hydrogen reference | Questions 4-6 of 12

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.

# B8 · Electrolytic cells

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

Explain how an electrolytic cell uses electrical energy to drive a reaction that is not spontaneous, label its parts, write its reactions and compare it with a voltaic cell.

Name: _____ Date: _____

## 1. A power source drives the reaction

## A reaction that does not go by itself

Look at $Cu + Zn^{2+} \rightarrow Cu^{2+} + Zn$. $Zn^{2+}$ is reduced at the cathode (−0.76 V) and copper is oxidized at the anode (+0.34 V).

E°cell = −0.76 − (+0.34) = _____. A negative value means the reaction is _____ as written, so it will not go by itself.

But a power source can push it, like pushing a ball uphill. A cell that uses _____ energy to drive a reaction that is not spontaneous is an electrolytic cell.

The source must supply more than _____. We ignore the sign, because the size of the negative value is the minimum push needed.

On the diploma: give a minimum voltage as a positive number. A negative one is a common wrong choice.

Watch: A reaction that is not spontaneous (Tyler DeWitt, “Electrolysis”, 3:58)

Watch: Splitting water needs electrical energy (Alberta Chemistry Teacher, “Chemistry 30 - Predicting redox reaction from species in the reduction table”, 20:04)

Watch: Why the battery must push the electrons (Tyler DeWitt, “Introduction to Electrochemistry”, 13:46)

## Anode and cathode

In every cell, _____ happens at the cathode and _____ happens at the anode. Those names come from the chemical change, not from a plus or minus sign.

In an electrolytic cell, the power source sets the signs. A terminal is where a wire joins the source. The negative terminal pushes electrons out, and they arrive at the electrode joined to it.

Electrons arriving means that electrode gains electrons, which is reduction. So the cathode is joined to the _____ terminal and the anode to the _____ terminal.

Cations (positive ions) move toward the _____. Anions (negative ions) move toward the _____.

Electrons move only in the _____. In the liquid, the moving ions carry the charge. Both electrodes sit in one container, so no salt bridge is needed. In a drawing, a power source in the wire means electrolytic, and a voltmeter means voltaic.

On the diploma: a diagram may show only the power source. Label the terminals first, then the electrodes. The wrong option usually swaps the two electrodes.

Watch: Anode and cathode in an electrolytic cell (Tyler DeWitt, “Electrolysis”, 8:02)

Watch: A way to remember anode and cathode (Tyler DeWitt, “Introduction to Electrochemistry”, 10:44)

Cathode joins the _____ terminal; anode joins the _____ terminal.

Cathode (reduction): _____

Anode (oxidation): _____

E°cell = −2.71 − 1.36 = _____; source must exceed _____

# B8 · Electrolytic cells

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

## 2. Molten salts

### Molten sodium chloride

Molten NaCl is the ionic solid heated until it melts, so its ions can move. It has only $Na^+$ and $Cl^-$ ions and no water.

$Na^+$ is attracted to the negative _____ and gains one electron: $Na^+ + e^- \rightarrow$ _____.

$Cl^-$ goes to the positive anode. $Cl_2$ has two atoms, so two chloride ions each give up one electron: $2Cl^- \rightarrow$ _____ + $2e^-$.

The anode gives 2 electrons and the cathode takes only 1, so double the cathode half. The cell reaction is $2Na^+ + 2Cl^- \rightarrow$ _____.

$E°cell = -2.71 - (+1.36) =$ _____, so the source must supply more than _____. We do not double E° when we double a half-reaction.

On the diploma: list the ions first, then send the cation to the cathode. Doubling E° with the half-reaction is a common wrong choice.

Watch: Splitting sodium chloride (Tyler DeWitt, “Electrolysis”, 0:46)

Watch: A battery forces the reaction (Tyler DeWitt, “Electrolysis”, 4:46)

Watch: Why the salt must be molten (Tyler DeWitt, “Electrolysis”, 7:10)

Watch: The two half-reactions (Tyler DeWitt, “Electrolysis”, 11:56)

Watch: The final equation (Tyler DeWitt, “Electrolysis”, 14:19)

### Molten zinc bromide

$ZnBr_2(l)$ has only $Zn^{2+}$ and $Br^-$. $Zn^{2+}$ goes to the _____: $Zn^{2+} + 2e^- \rightarrow Zn$. $Br^-$ goes to the _____: $2Br^- \rightarrow Br_2 + 2e^-$.

Both halves move two electrons, so they cancel when we add them. The cell reaction is $Zn^{2+} + 2Br^- \rightarrow$ _____.

$E°cell = -0.76 - (+1.07) =$ _____. A negative value shows that a power _____ is needed, one that supplies more than 1.83 V.

On the diploma: reporting +1.83 V would suggest the cell runs by itself. Check the electrode order before you subtract, and give the sign.

# B8 · Electrolytic cells

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

## 3. Compare cells, plating and refining

### Side by side

The same two half-cells can run either way. In the Zn / Cu voltaic cell, zinc is oxidized and E°cell is +1.10 V. In the electrolytic cell, copper is oxidized and E°cell is −1.10 V for the forced reaction.

Voltaic cell: _____ energy to electrical energy, E°cell _____, anode _____. It needs a _____ or porous barrier.

Electrolytic cell: electrical energy to _____ energy, E°cell for the forced reaction _____, anode _____. It _____ need a salt bridge.

The voltaic anode is negative because that cell makes electrons there. The electrolytic anode is positive because it is joined to the positive terminal.

On the diploma: many wrong choices copy the voltaic sign onto an electrolytic cell. Check which cell type the question describes.

Watch: Electrolytic cell versus voltaic cell (Tyler DeWitt, “Introduction to Electrochemistry”, 11:30)

### Plating and refining

To plate a spoon with silver, make the spoon the _____. Reduction deposits metal there: $Ag^+ + e^- \rightarrow$ _____, so the spoon gains mass.

A silver anode _____ mass as $Ag \rightarrow Ag^+ + e^-$. It replaces the $Ag^+$ the cathode used up, so the ion concentration stays nearly constant. A carbon anode would not dissolve.

To refine copper, impure copper is the anode and a thin pure sheet is the cathode. Pure metal builds up on the _____. Impurities above copper on the table, like silver and gold, fall as solid. Those below it dissolve.

Two electrodes of the same metal in its own ion give E°cell = _____, because the same half-reaction is reversed at the anode. Only a small voltage is needed, plus a little extra to beat resistance.

On the diploma: the electrode that gains mass is where ions are reduced, the cathode. Same metal at both electrodes means E°cell = 0.00 V.

Watch: Plating a copper ring with silver (Tyler DeWitt, “Electroplating”, 1:30)

Watch: Which terminal the object connects to (Tyler DeWitt, “Electroplating”, 3:00)

Watch: Half-reaction at the ring (Tyler DeWitt, “Electroplating Part 2”, 1:27)

Watch: Half-reaction at the silver (Tyler DeWitt, “Electroplating Part 2”, 2:36)

Watch: Anode and cathode in electroplating (Tyler DeWitt, “Electroplating Part 2”, 6:23)

# B8 · Practice

Electrolytic cells | 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.

B8-P01 · Easy · Given $Ni^{2+} + 2e^- \rightarrow Ni$ (−0.26 V) and $Cu^{2+} + 2e^- \rightarrow Cu$ (+0.34 V). A power source is used to drive $Cu + Ni^{2+} \rightarrow Cu^{2+} + Ni$. Calculate E°cell and state the smallest voltage that the source must exceed.

B8-P02 · Easy · In an electrolytic cell, which electrode is joined to the negative terminal of the power source? Name the process there and give the direction in which cations move.

B8-P03 · Easy · Explain why electrons do not travel through the electrolyte, and why an electrolytic cell does not need a salt bridge.

# B8 · Practice

Electrolytic cells | Questions 4-6 of 12

B8-P04 · Easy · Write the half-reactions, the cell reaction and the product at each electrode for the electrolysis of molten NaCl.

B8-P05 · Medium · Given molten $ZnBr_2(l)$ with inert electrodes ($Zn^{2+}/Zn$ −0.76 V; $Br_2/Br^-$ +1.07 V). Write the cell reaction and calculate E°cell.

B8-P06 · Medium · Given molten $MgCl_2$ with inert electrodes ($Mg^{2+}/Mg$ −2.37 V; $Cl_2/Cl^-$ +1.36 V). Write the cell reaction and calculate the smallest voltage that the source must exceed.

# B8 · Practice

Electrolytic cells | Questions 7-9 of 12

B8-P07 · Medium · Compare a voltaic cell and an electrolytic cell for the Zn / Cu pair. Use the energy change, the sign of the cell value for the reaction that runs, and the sign of the anode.

B8-P08 · Medium · A student says the anode in an electrolytic cell is negative because it is negative in a voltaic cell. Explain the error.

B8-P09 · Challenging · A silver spoon is plated in $AgNO_3(aq)$ with a silver bar. Say which is the cathode, write the half-reaction at each electrode and predict the mass change of each.

# B8 · Practice

Electrolytic cells | Questions 10-12 of 12

B8-P10 · Challenging · Describe a cell for refining impure copper and name the electrode that gains mass.

B8-P11 · Challenging · Silver electrodes are placed in silver nitrate solution and connected to a power source. Calculate E°cell for the reaction in this cell and say how much voltage is needed in principle.

B8-P12 · Challenging · A student says that a negative E°cell means a reaction cannot happen at all. Use the idea of an electrolytic cell to correct this.

# B9 · Electrolysis of aqueous solutions

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

Decide what reacts at each electrode when water is present, explain the chloride exception and why E° must be used with caution, and predict products and observations.

Name: _____ Date: _____

## 1. Water joins the competition

### Choose by table value

In B8 the cells held a molten salt, so only its ions could react. In a water solution, water is in the beaker too, and it can react at either electrode.

Water can be reduced: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$, _____. Water can be oxidized: $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$, which is the $O_2$ row run backward, from the _____ row.

At the cathode, the strongest oxidizing agent is reduced, and it has the _____ value. At the anode, the strongest reducing agent is oxidized, and it has the _____ value.

$Na^+$ (−2.71 V) is lower than water, so _____ is reduced at the cathode. It makes $H_2$ gas and _____, so the solution turns basic and the pH _____. Red litmus turns blue there. Sodium metal never forms from a water solution.

On the diploma: list every ion, and water. At the cathode pick the highest value. At the anode pick the lowest. Picking the highest at the anode is a common wrong choice.

Watch: Electrolysis of water (Tyler DeWitt, “Electrolysis”, 19:02)

### Cathode examples

$Cu^{2+}$ (+0.34 V) is higher than water, so _____ is reduced to copper metal. The cathode _____ mass and the blue colour fades as $Cu^{2+}$ is used up.

The dividing line is −0.83 V. A cation above it, like $Ag^+$ or $Cu^{2+}$, is reduced before _____. A cation below it, like $Na^+$, $Mg^{2+}$ or $Al^{3+}$, loses to water.

$Zn^{2+}$ (−0.76 V) is just above the line, so it is predicted to win by a small margin. This is a prediction from the table, so compare numbers every time.

On the diploma: the wrong options usually reduce a metal that sits below water.

Cathode choices: $Cu^{2+}$ +0.34 V or water −0.83 V. Reduced: _____

Anode choices: sulfate (not oxidized) or water +1.23 V. Oxidized: _____

Products: cathode _____; anode _____

Near the anode the pH _____ because $H^+$ is made.

# B9 · Electrolysis of aqueous solutions

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

## 2. Anodes and the chloride exception

### Choose the anode reaction

Take NaI(aq) with inert electrodes, which means electrodes that do not react, like carbon or platinum. $I^-$ (+0.54 V) is lower than water (+1.23 V), so _____ is oxidized: $2I^- \rightarrow$ _____ + $2e^-$.

We observe a brown colour near the anode and infer that _____ forms. The $Na^+$ ions are spectator ions, which means they do not react at either electrode.

In $CuSO_4(aq)$, sulfate appears in the booklet only as a product, in the $S_2O_8^{2-}$ row at +2.01 V. That is higher than water, so sulfate is not oxidized and _____ is: $2H_2O \rightarrow O_2$ + _____ + $4e^-$.

Oxygen bubbles form, and $H^+$ is made, so the pH _____ near the anode. Blue litmus turns red there.

A reactive electrode can react itself. A copper anode (+0.34 V) is lower than water, so copper is the stronger reducing agent: $Cu \rightarrow Cu^{2+} + 2e^-$. The anode _____ and loses mass.

On the diploma: sulfate and nitrate are not oxidized, so water reacts. Always check what the electrode is made of. A wrong option ignores a reactive anode.

### The chloride anomaly

In NaCl(aq), $Cl^-$ is +1.36 V and water is +1.23 V. Water has the lower value, so the table predicts oxygen. But chlorine gas is _____: $2Cl^- \rightarrow Cl_2 + 2e^-$.

This is the chloride _____. The reason is overvoltage, the extra voltage some reactions need to run at a useful rate. Oxidizing water needs a _____ overvoltage than oxidizing chloride, so chloride reacts.

So think of a table value as a minimum, not a promise. E° values must be used with _____. The table minimum is _____ enough on its own, because an extra overvoltage must be applied.

On the diploma: write the table prediction first, then check for exceptions. In a chloride solution the observed product is chlorine gas. An option saying the table minimum is enough is a wrong choice.

# B9 · Electrolysis of aqueous solutions

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

## 3. Predict products and observations

### Sodium bromide

Take NaBr(aq) with inert electrodes. At the cathode, water (−0.83 V) beats $Na^+$ (−2.71 V), so _____ is reduced: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$.

At the anode, $Br^-$ (+1.07 V) is lower than water (+1.23 V), so _____ is oxidized: $2Br^- \rightarrow Br_2 + 2e^-$. This is not a chloride case, so there is no exception.

Both halves move two electrons, so no multiplier is needed. The cell reaction is $2Br^- + 2H_2O \rightarrow Br_2 + H_2 +$ _____. Atoms and charge check: Br 2, H 4, O 2, charge −2 on each side.

E°cell = −0.83 − 1.07 = _____. The pH rises near the _____, because $OH^-$ is made there.

On the diploma: E°cell is in volts, and the sign matters.

### Observations and inferences

What we can see is an _____: bubbles at the cathode and a brown-orange colour near the anode.

Hydrogen gas, bromine and a rising pH are _____ from the half-reactions. A pH _____, a dye that changes colour with pH, can test the pH.

On the diploma: keep observation and inference apart. You see colour and bubbles. From them you work out each product.

Watch: Twice as much hydrogen as oxygen (Tyler DeWitt, “Electrolysis”, 22:58)

### Copper sulfate and sodium sulfate

$CuSO_4(aq)$ with inert electrodes: copper forms at the _____, $Cu^{2+} + 2e^- \rightarrow Cu$. $O_2$ forms at the anode and the pH _____.

The water half moves four electrons and each copper half only two, so multiply the copper half by 2. Then add: $2Cu^{2+} + 2H_2O \rightarrow 2Cu + O_2 +$ _____.

E°cell = 0.34 − 1.23 = −0.89 V, so the source must supply more than 0.89 V. The blue colour fades and the solution turns more acidic.

$Na_2SO_4(aq)$: neither ion reacts, so water reacts at both electrodes. The ions only carry charge through the solution.

Double the cathode half to match four electrons: $4H_2O + 4e^- \rightarrow 2H_2 + 4OH^-$. Adding gives $6H_2O \rightarrow 2H_2 + O_2 + 4OH^- + 4H^+$.

The $4OH^-$ and $4H^+$ make $4H_2O$, so the overall reaction is $2H_2O \rightarrow$ _____ + $O_2$, a 2 : 1 ratio of $H_2$ to $O_2$.

On the diploma: expect the 2 to 1 ratio of $H_2$ to $O_2$ in a numerical-response item. $H^+$ on the product side means the pH goes down.

# B9 · Practice

Electrolysis of aqueous solutions | 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.

B9-P01 · Easy · Explain why there are more possible reactions in an aqueous solution than in a molten salt.

B9-P02 · Easy · Given $NaNO_3(aq)$ with inert electrodes ($Na^+$ −2.71 V; water −0.83 V). Decide what is reduced at the cathode, write the half-reaction and state the change in pH near the cathode.

B9-P03 · Easy · Given $CuSO_4(aq)$ with inert electrodes ($Cu^{2+}$ +0.34 V; water −0.83 V). Decide what is reduced at the cathode and predict the change in mass of the cathode and the colour of the solution near it.

# B9 · Practice

Electrolysis of aqueous solutions | Questions 4-6 of 12

B9-P04 · Easy · Given NaI(aq) with inert electrodes ($I_2/I^-$ +0.54 V; $O_2/H_2O$ +1.23 V). Decide what is oxidized at the anode, write the half-reaction and describe the observation and the inference.

B9-P05 · Medium · In $CuSO_4$(aq) with inert electrodes, water is oxidized at the anode. Write the half-reaction and predict the observation at the anode and the change in pH near it.

B9-P06 · Medium · Given $CuSO_4$(aq) with a copper anode and an inert cathode ($Cu^{2+}/Cu$ +0.34 V; $O_2/H_2O$ +1.23 V). Decide what is oxidized at the anode and predict the change in its mass.

# B9 · Practice

Electrolysis of aqueous solutions | Questions 7-9 of 12

B9-P07 · Medium · Aqueous NaCl is electrolyzed with inert electrodes. State the product at the anode that the table predicts, the product that is observed, and the reason for the difference.

B9-P08 · Medium · Given NaBr(aq) with inert electrodes (water −0.83 V; $Br_2/Br^-$ +1.07 V; $O_2/H_2O$ +1.23 V). Write the cell reaction and calculate E°cell.

B9-P09 · Challenging · Given $CuSO_4(aq)$ with inert electrodes. Match the electrons in the half-reactions, write the cell reaction and calculate E°cell ($Cu^{2+}/Cu$ +0.34 V; $O_2/H_2O$ +1.23 V).

# B9 · Practice

Electrolysis of aqueous solutions | Questions 10-12 of 12

B9-P10 · Challenging · Given $Na_2SO_4(aq)$ with inert electrodes. Write the half-reactions with matched electrons and the overall reaction, and say which gas forms in the larger amount.

B9-P11 · Challenging · A student says that applying the minimum voltage from the table guarantees the predicted reaction. Use aqueous NaCl to explain why this is not correct.

B9-P12 · Challenging · Given $AgNO_3(aq)$ with inert electrodes ($Ag^+$ +0.80 V; water −0.83 V; $O_2/H_2O$ +1.23 V). State what is reduced and what is oxidized and write the cell reaction.