



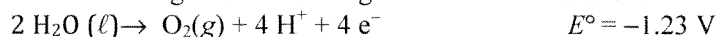
Electrochemistry Cheat Sheet	
E°_{cell} ; reduction and oxidizing agent; cell potential; reduction or oxidation; anode; cathode; salt bridge; electron flow; voltage; electromotive force; galvanic/voltaic; electrode; battery; current; amps; time; grams (mass); plate/deposit; electroplating; identity of metal; coulombs of charge	
Electrolytic Cell Relationships	
The Mnemonics apply to Electrolytic cells too: ANOX; REDCAT; FATCAT...	Electrolytic Positive Anode
“the more positive reduction potential gets to be reduced”	96,500 Coulombs = 1 mole of electrons
“the more positive oxidation potential gets to be oxidized”	# electrons in balanced equations = # moles of electrons transferred
Reduction of H_2O $2 \text{H}_2\text{O}(\ell) + 2 \text{e}^- \rightarrow \text{H}_2(\text{g}) + 2 \text{OH}^- \quad E^\circ = -0.83 \text{ V}$	All time measurements must be in sec for electroplating/electrolysis problems
Oxidation of H_2O $2 \text{H}_2\text{O} \rightarrow \text{O}_2 + 4 \text{H}^+ + 4 \text{e}^- \quad E^\circ = -1.23 \text{ V}$	$\frac{It}{n\mathfrak{S}}$ (Molar Mass) = grams plated
No alkali or alkaline earth metal can be reduced in an aqueous solution - water is more easily reduced.	Polyatomic ions are typically NOT oxidized in an aqueous solution - water is more easily oxidized.
	$\text{amp}(I) = \frac{\text{Coulomb}(q)}{\text{sec}(t)}$
$E^\circ_{\text{cell}} = -$; not thermodynamically favored; $\Delta G = (+)$; $K < 1$	$\text{Volt} = 1 \frac{\text{Joule}(J)}{\text{Coulomb}(c)}$
$\Delta G^\circ = -n\mathfrak{S} E^\circ$	$\text{Amp} = \frac{\text{Coulomb}(c)}{\text{sec}(t)}$
$\Delta G^\circ = -RT \ln K$	$\text{Faraday}(\mathfrak{S}) = 96,500 \frac{\text{Coulomb}(c)}{\text{mol of e}^-}$
Inert electrodes are typically used in gas and ion to ion galvanic cells and in electrolytic cells	
Connections	
Thermo and Equilibrium	Stoichiometry
Potential Pitfalls	
Watch signs on voltages!!	BE SURE units cancel out in your calculations.
Balancing overall reactions – make sure # of electrons is the same in both half reactions.	Units on E° are volts

NMSI SUPER PROBLEM

An electric current is applied to two separate solutions for 30 minutes, under the same conditions using inert electrodes. Observations are noted in the table below.

Solution A – 1.0 M K ₂ SO ₄	Solution B – 1.0 M CuSO ₄
Anode: gas bubbles	Anode: gas bubbles
Cathode: gas bubbles	Cathode: dark flakes formed on the electrode

In both reactions, water is oxidized according to the following oxidation half-reaction.



(a) Write the balanced equation for the half-reaction that occurs at the *cathode* in

(i) Solution A

$2 \text{H}_2\text{O} + 2 \text{e}^- \rightarrow \text{H}_2 + 2 \text{OH}^- \quad E^\circ = -0.83 \text{ V}$ Either K ⁺ or H ₂ O will be reduced. $\text{K}^+ + \text{e}^- \rightarrow \text{K} \quad E^\circ = -2.92 \text{ V}$ $2 \text{H}_2\text{O}(\ell) + 2 \text{e}^- \rightarrow \text{H}_2(\text{g}) + 2 \text{OH}^- \quad E^\circ = -0.83 \text{ V}$ Since H ₂ O has a more positive reduction potential it is more easily reduced than H ₂ O. H ₂ O is always reduced over an alkali metal	1 point for a correctly indicating the reduction of H ₂ O over K ⁺ 1 point for the correct reduction half-reaction.
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(ii) Solution B

$\text{Cu}^{2+} + 2 \text{e}^- \rightarrow \text{Cu} \quad E^\circ = +0.34 \text{ V}$ Either Cu ²⁺ or H ₂ O will be reduced. $\text{Cu}^{2+} + 2 \text{e}^- \rightarrow \text{Cu} \quad E^\circ = +0.34 \text{ V}$ $2 \text{H}_2\text{O}(\ell) + 2 \text{e}^- \rightarrow \text{H}_2(\text{g}) + 2 \text{OH}^- \quad E^\circ = -0.83 \text{ V}$ Since Cu ²⁺ has a more positive reduction potential it is more easily reduced than H ₂ O.	1 point for a correctly balanced reaction showing the reduction of Cu ²⁺ over H ₂ O 1 point for the correct reduction half-reaction.
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(b) For Solution A, is the reaction thermodynamically favorable or not thermodynamically favorable? Justify your answer.

Non-thermodynamically favorable, since the reaction requires electric current the E° is negative (–) for this reaction, therefore ΔG is positive (+).	1 point for a correctly identifying Cu metal
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- (c) In the electrolysis of the K_2SO_4 solution, identify the gas produced and describe a test that can be used to identify the gas at the

(i) anode

At the anode, the oxidation of H_2O releases O_2 gas	1 point for a correctly identifying O_2 gas
An identifying test for O_2 gas is trapping the gas in a test tube and inserting a glowing wooden splint. In the presence of O_2 gas a glowing wooden splint will re-ignite.	1 point for a correctly identifying the O_2 test.

(ii) cathode

At the cathode, the reduction of H_2O releases H_2 gas	1 point for a correctly identifying H_2 gas
An identifying test for H_2 gas is trapping the gas in a test tube and inserting a burning wooden splint. In the presence of H_2 gas a popping sound will be observed.	1 point for a correctly identifying the H_2 test.

- (d) Describe in the box below, what observations, if any, would be noted if a couple of drops of phenolphthalein indicator were added around the cathode of *both* solutions. Phenolphthalein indicator is colorless in acidic solutions and turns pink in basic solutions.

Justify your observations.

Solution A – 1.0 M K_2SO_4	Solution B – 1.0 M $CuSO_4$

In solution A, since H_2O is reduced the OH^- ion is produced. This will create an alkaline solution that will turn pink in the presence of phenolphthalein.	1 point for a correctly indicating the solution will turn pink with justification.
In solution B, Cu^{2+} is reduced to Cu. The phenolphthalein will have no observed effect on this solution since it is not becoming significantly more basic.	1 point for a indicating the solution will not change with justification.

- (e) The dark flakes formed on the electrode in the electrolysis of Solution B were collected and dried. The mass of these flakes was determined to be 1.019 grams.

(i) Identify the flakes.

The flakes are Cu metal, as Cu^{2+} is reduced to Cu.	1 point for a correctly identifying Cu metal
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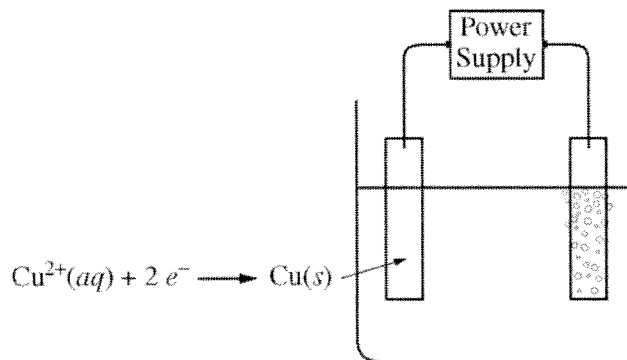


(ii) Calculate the amount of current that was passed through Solution B.

$\frac{It}{n\mathfrak{S}}(\text{Molar Mass}) = g$ $I = \frac{gn\mathfrak{S}}{(\text{Molar Mass})t} = \frac{(1.019)(2)(96500)}{(63.55)(30)(60)} = 1.72 \text{ amps}$	<p>1 point for the correct number of Coulombs</p> <p>1 point for the correct number of amps of current</p> <p>(both points are earned using the equation shown)</p>
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Question 3



An external direct-current power supply is connected to two platinum electrodes immersed in a beaker containing 1.0 M $\text{CuSO}_4(\text{aq})$ at 25°C, as shown in the diagram above. As the cell operates, copper metal is deposited onto one electrode and $\text{O}_2(\text{g})$ is produced at the other electrode. The two reduction half-reactions for the overall reaction that occurs in the cell are shown in the table below.

Half-Reaction	$E^\circ(\text{V})$
$\text{O}_2(\text{g}) + 4 \text{H}^+(\text{aq}) + 4 e^{-} \rightarrow 2 \text{H}_2\text{O}(\text{l})$	+1.23
$\text{Cu}^{2+}(\text{aq}) + 2 e^{-} \rightarrow \text{Cu}(\text{s})$	+0.34

(a) On the diagram, indicate the direction of electron flow in the wire.

The electron flow in the wire is from the right toward the left (counterclockwise).	One point is earned for the correct direction.
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(b) Write a balanced net ionic equation for the electrolysis reaction that occurs in the cell.

$2 \text{H}_2\text{O}(\text{l}) + 2 \text{Cu}^{2+}(\text{aq}) \rightarrow 4 \text{H}^+(\text{aq}) + 2 \text{Cu}(\text{s}) + \text{O}_2(\text{g})$	<p>One point is earned for all three products.</p> <p>One point is earned for balancing the equation.</p>
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(c) Predict the algebraic sign of ΔG° for the reaction. Justify your prediction.

The sign of ΔG° would be positive because the reaction is NOT spontaneous.	One point is earned for indicating that ΔG° is greater than zero <u>and</u> supplying a correct explanation.
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Question 3 (continued)

(d) Calculate the value of ΔG° for the reaction.

$E^\circ = -1.23 \text{ V} + 0.34 \text{ V} = -0.89 \text{ V} = -0.89 \text{ J C}^{-1}$ $\Delta G^\circ = -n \mathcal{F} E^\circ = -4 (96,500 \text{ C mol}^{-1})(-0.89 \text{ J C}^{-1})$ $= +340,000 \text{ J mol}^{-1} = +340 \text{ kJ mol}^{-1}$	<p>One point is earned for calculating E°.</p> <p>One point is earned for calculating ΔG° (consistent with the calculated E°).</p>
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An electric current of 1.50 amps passes through the cell for 40.0 minutes.

(e) Calculate the mass, in grams, of the Cu(s) that is deposited on the electrode.

$q = (1.50 \text{ C s}^{-1})(40.0 \text{ min}) \times \frac{60 \text{ s}}{1 \text{ minute}} = 3,600 \text{ C}$ $\text{mass Cu} = (3,600 \text{ C}) \times \frac{1 \text{ mol } e^-}{96,500 \text{ C}} \times \frac{1 \text{ mol Cu}}{2 \text{ mol } e^-} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}}$ $= 1.19 \text{ g Cu}$	<p>One point is earned for calculating q.</p> <p>One point is earned for calculating the mass of copper deposited.</p> <p style="text-align: center;">OR</p> <p>Two points are earned for calculating the mass of copper in one step.</p>
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(f) Calculate the dry volume, in liters measured at 25°C and 1.16 atm, of the O₂(g) that is produced.

$n_{\text{O}_2} = (1.19 \text{ g Cu}) \times \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} \times \frac{1 \text{ mol O}_2}{2 \text{ mol Cu}} = 0.00936 \text{ mol O}_2$ $V = \frac{nRT}{P} = \frac{(0.00936 \text{ mol})(0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1})(298 \text{ K})}{1.16 \text{ atm}}$ $= 0.197 \text{ L}$	<p>One point is earned for calculating the number of moles of O₂.</p> <p>One point is earned for calculating V (consistent with previous calculations).</p>
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1991 AP[®] CHEMISTRY FREE-RESPONSE QUESTIONS

Question 7 – Modified into a Short Free Response

Explain each of the following.

- (a) When an aqueous solution of NaCl is electrolyzed, $\text{Cl}_2(\text{g})$ is produced at the anode, but no $\text{Na}(\text{s})$ is produced at the cathode.

Cl^- is more easily oxidized than H_2O	1 point is earned for the correct justification for oxidation.
H_2O is more easily reduced than Na^+	1 point is earned for the correct justification for reduction.

- (b) The mass of $\text{Fe}(\text{s})$ produced when 1 faraday is used to reduce a solution of FeSO_4 is 1.5 times the mass of $\text{Fe}(\text{s})$ produced when 1 faraday is used to reduce a solution of FeCl_3 .

Fe^{2+} requires 2 Faraday per mole of $\text{Fe}(\text{s})$ produced	1 point is earned for the identifying the inverse relationship between the number of Faraday and mole of Fe produced. 1 point is earned for connecting the flow of electrons to moles produced.
Fe^{3+} requires 3 Faraday per mole of $\text{Fe}(\text{s})$ produced	
Therefore the 2:3 ratio indicates the production of 1.5 times	

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Question 8 (continued)

Half-reaction	E° (V)
$\text{O}_2(\text{g}) + 4 \text{H}^+ + 4 \text{e}^- \rightarrow 2 \text{H}_2\text{O}(\text{l})$	1.23
$\text{I}_2(\text{s}) + 2 \text{e}^- \rightarrow 2 \text{I}^-$	0.53
$2 \text{H}_2\text{O}(\text{l}) + 2 \text{e}^- \rightarrow \text{H}_2(\text{g}) + 2 \text{OH}^-$	-0.83
$\text{Na}^+ + \text{e}^- \rightarrow \text{Na}(\text{s})$	-2.71

(d) An electric current is applied to a 1.0 M NaI solution.

(i) Write the balanced oxidation half-reaction for the reaction that takes place.

$2 \text{I}^- \rightarrow \text{I}_2(\text{s}) + 2 \text{e}^-$	One point is earned for the correct half-reaction.
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(ii) Write the balanced reduction half-reaction for the reaction that takes place.

$2 \text{H}_2\text{O}(\text{l}) + 2 \text{e}^- \rightarrow \text{H}_2(\text{g}) + \text{OH}^-$	One point is earned for the correct half-reaction.
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(iii) Which reaction takes place at the anode, the oxidation reaction or the reduction reaction?

The oxidation half-reaction occurs at the anode.	One point is earned for the correct choice.
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(iv) All electrolysis reactions have the same sign for ΔG° . Is the sign positive or negative? Justify your answer.

The sign of ΔG for all electrolysis reactions is positive. Because electrolysis reactions are non-spontaneous, energy in the form of applied electrical current (electrical work) must be applied to make the reaction occur.	One point is earned for the correct sign of ΔG and a correct explanation.
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Question 3
(9 points)

(a) $2 \text{Cl}^- \rightarrow \text{Cl}_2 + 2 e^-$ (equation need not be balanced) 1 point

(b) $\frac{0.250 \text{ coul/sec} \times 7,200 \text{ sec}}{96,500 \text{ coul/mol } e^-} = \frac{1,800 \text{ coul}}{96,500 \text{ coul/mol } e^-} = 0.01865 \text{ mol } e^-$ 1 point

$\text{mol Fe} = \frac{0.521 \text{ g Fe}}{55.85 \text{ g/mol Fe}} = 0.00933 \text{ mol Fe}$ 1 point

$\frac{\text{mol } e^-}{\text{mol Fe}} = \frac{1.865 \times 10^{-2} \text{ mol } e^-}{9.33 \times 10^{-3} \text{ mol Fe}} \approx 2 e^- \text{ per Fe atom} \Rightarrow \text{FeCl}_2$ 1 point



Notes: "FeCl₂(aq)" accepted for reactants.

Any balanced equation corresponding to answer in part (b) earns 1 point.

(d) $\text{moles Fe}^{2+} = \text{moles Cl}_2 = 9.33 \times 10^{-3} \text{ mol Cl}_2$ 1 point

$$V = \frac{nRT}{P} = \frac{0.00933 \text{ mol Cl}_2 \times 0.0821 \text{ L} \cdot \text{atm} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \times 298 \text{ K}}{\left(\frac{750}{760}\right) \text{ atm}}$$

= 0.231 L (or 231 mL) 1 point

(e) $\frac{3.00 \text{ g Cl}_2 / 71 \text{ g} \cdot \text{mol}^{-1}}{3,600 \text{ sec}} = \frac{0.0423 \text{ mol Cl}_2}{3,600 \text{ sec}} = 1.17 \times 10^{-5} \frac{\text{mol Cl}_2}{\text{sec}}$ 1 point

$$\text{current (in amperes)} = \frac{2 \text{ mol } e^-}{\text{mol Cl}_2} \times 1.17 \times 10^{-5} \frac{\text{mol Cl}_2}{\text{sec}} \times \frac{96,500 \text{ coul}}{1 \text{ mol } e^-}$$

= 2.27 amp (coul/sec) 1 point

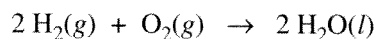
Alternate solution: $0.00933 \text{ mole Cl}_2 / 2 \text{ hrs} = 0.662 \text{ g Cl}_2 / 2 \text{ hrs} = 0.331 \text{ g Cl}_2 / \text{hr}$

$$\frac{0.250 \text{ amp}}{0.331 \text{ g Cl}_2} = \frac{x}{3.00 \text{ g Cl}_2}$$

$$x = \frac{3.00 \text{ g} \times 0.250 \text{ amp}}{0.331 \text{ g}} = 2.27 \text{ amp}$$

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Question 3



In a hydrogen-oxygen fuel cell, energy is produced by the overall reaction represented above.

- (a) When the fuel cell operates at 25°C and 1.00 atm for 78.0 minutes, 0.0746 mol of $\text{O}_2(\text{g})$ is consumed. Calculate the volume of $\text{H}_2(\text{g})$ consumed during the same time period. Express your answer in liters measured at 25°C and 1.00 atm.

$(0.0746 \text{ mol O}_2) \times \frac{2 \text{ mol H}_2}{1 \text{ mol O}_2} = 0.149 \text{ mol H}_2$ $V = \frac{n_{\text{H}_2} RT}{P} = \frac{(0.149 \text{ mol H}_2)(0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1})(298 \text{ K})}{1.00 \text{ atm}}$ $= 3.65 \text{ L H}_2$	<p>One point is earned for calculation of moles of H_2.</p> <p>One point is earned for substitution into $PV = nRT$.</p> <p>One point is earned for the answer.</p>
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- (b) Given that the fuel cell reaction takes place in an acidic medium,
- (i) write the two half reactions that occur as the cell operates,

$\text{O}_2 + 4 \text{H}^+ + 4 e^- \rightarrow 2 \text{H}_2\text{O}$ $\text{H}_2 \rightarrow 2 \text{H}^+ + 2 e^-$	<p>One point is earned for each of the two half reactions.</p>
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- (ii) identify the half reaction that takes place at the cathode, and

$\text{O}_2 + 4 \text{H}^+ + 4 e^- \rightarrow 2 \text{H}_2\text{O}$	<p>One point is earned for either the equation of the correct half reaction, or for indicating “the reduction half reaction” if the correct equation is given in (b)(i).</p>
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- (iii) determine the value of the standard potential, E° , of the cell.

$E^\circ = 1.23 \text{ V} + 0.00 \text{ V} = 1.23 \text{ V}$	<p>One point is earned for the standard potential.</p>
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Question 3 (continued)

- (c) Calculate the charge, in coulombs, that passes through the cell during the 78.0 minutes of operation as described in part (a).

$(0.0746 \text{ mol O}_2) \times \frac{4 \text{ mol } e^-}{1 \text{ mol O}_2} \times \frac{96,500 \text{ C}}{1 \text{ mol } e^-} = 2.88 \times 10^4 \text{ C}$	<p>One point is earned for the stoichiometry.</p> <p>One point is earned for the answer.</p>
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