

IB electrochemistry

topics

redox

oxidation states

nomenclature

half-equations

agents

activity series

redox titrations

winkler method

batteries (electrochemical cells)

voltaic

cell diagrams

standard electrode potentials

calculating cell potential

electrolytic cells

electrolysis

electroplating

electroplating lab

In this lab you will take a brass key and electroplate copper onto its surface. The electroplating process is a 14 billion dollar industry as of 2015. To understand the chemistry involved please complete the prelab before you begin.

Faraday's Laws of Electrolysis

First Law :

The mass of a substance produced by electrolysis is proportional to the quantity of electricity used. In order to produce more substance electrolytically we must use more electricity.

S.I. unit of electric charge:

coulomb

= amount of charge in 1 ampere of current for one second ($C = 1 A s$ or $1A = 1 C/s$)

= 6.242×10^{18} electrons

Second Law:

one mole of electrons contains 96,500 coulombs of electricity

$F = \text{faraday constant} = 96,484 C/mol e^-$

Faraday's Laws of Electrolysis Calculations

$$Q = n_e F$$

$Q =$ quantity of electricity measured in coulombs (C)

$n_e =$ moles of electrons used

$F =$ the Faraday (Faraday constant) = 96,500 C/mol e^-

Examples

note that all of these problems may be solved by unit conversion using the Faraday constant.

- Calculate the quantity of electricity in coulombs obtained from 2 moles of electrons

$$96,500 C/mol \times 2 mol = 193,000C$$

- Calculate the moles of electrons obtained from 250. C of electricity

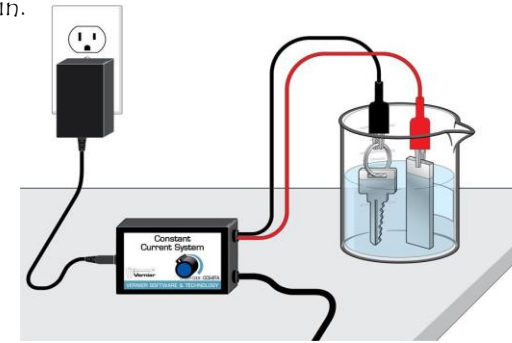
$$mol/96,500C \times 250C = 0.00259 mol$$

- Calculate the moles of copper metal that can be produced by the electrolysis of molten copper sulfate using 500. C of electricity.

$$mol/96,500C \times 500C = 0.00518 mol$$

however since $Cu^{2+} + 2e^- \rightarrow Cu(s)$ it takes 2 mol of electrons to produce 1 mol $Cu(s)$

$$0.00518 mol e^- \times mol Cu(s)/2 mol e^- = 0.00259 mol Cu$$



- Calculate the mass of silver that can be produced by the electrolysis of 1M $AgCN(aq)$ using 800 C of electricity.

note that one mol of electrons can plate out one mole of silver ($Ag^+ + e^- \rightarrow Ag(s)$)

$$\frac{mol e^-}{96,500C} \times 800C \times \frac{mol Ag}{mol e^-} \times \frac{107.9 g Ag}{mol Ag} = 0.895g Ag$$

- An electroplating device containing copper (II) sulfate is set to run at a current of 0.50 amps. How much copper will be deposited in ten seconds?

note that 1 amp = 1 coulomb / second and it takes 2 mol of electrons to deposit one mol of copper ($Cu^{2+} + 2e^- \rightarrow Cu$)

$$0.5 \text{ amps} \times \frac{\text{coulomb/second}}{\text{amp}} \times 10 \text{ seconds} = 5 C$$

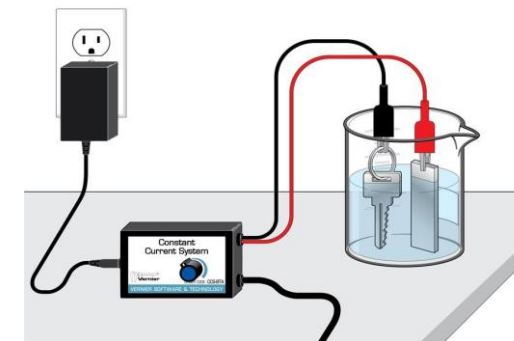
$$5 C \times \frac{mol e^-}{96,500 C} \times \frac{mol Cu}{2 mol e^-} \times \frac{63.55 g Cu}{mol Cu} = 0.00165 g Cu$$

- How long will it take to deposit 56 g of silver from a silver nitrate solution using a current of 4.5 A?

$$4.5 \text{ amps} \times \frac{\text{coulomb/sec}}{\text{amp}} \times \frac{107.9 g Ag}{96,500 C} = 0.005032 g Ag/second$$

$$\frac{\text{second}}{0.005032 g Ag} \times 56 \text{ seconds} = 11,130 \text{ seconds} \\ = 3.1 \text{ hours}$$

materials	
LabQuest	electrolyte solution (CuSO_4 in H_2SO_4)
LabQuest App	vinegar
Vernier Constant Current System	1 cm \times 10 cm copper strip
steel wool	brass key
two 250 mL beakers	solid sodium chloride, NaCl
distilled water	analytical balance
	bare copper wire, 20–22 gauge



procedure

Both Constant Current System and Generic Power Supply Users

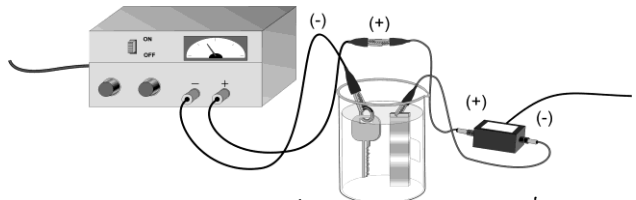
1. Obtain and wear goggles.
2. Use steel wool to clean a brass key and a strip of copper, which will be the electrodes of the electrochemical cell.
3. Mix 3 g of NaCl with 15 mL of vinegar in a 250 mL beaker. Wash the key and the copper strip in this salt/vinegar solution. Rinse the key and copper strip with distilled water and dry each metal piece.
4. Use an analytical balance to determine the mass of the key and the mass of the copper strip. Record these two masses in your data table.
5. Fill a 250 mL beaker about 3/4 full with the electrolyte solution. *CAUTION: The electrolyte solution in this experiment is prepared in H_2SO_4 and should be handled with care.*
6. Attach a 7 cm length of bare copper wire to the brass key to act as a handle, so that it is easier to completely immerse the key in the electrolyte solution.

Constant Current System Users Only (Others proceed to the Generic Power Supply Section)

7. Gently turn the dial of the Constant Current System counterclockwise to confirm that it is in the minimum current position. Clip the wire handle on the key and the copper strip to the proper leads (See Figure 1). Important: You will not place the electrodes in the cell until Step 10.
8. Plug the Constant Current System into a powered electrical outlet. Connect the sensor cable to Channel 1 of the LabQuest and choose New from the File menu.
9. Change the data-collection rate to 0.05 samples/second and the data collection duration to 1800 seconds.

Generic Power Supply Users

7. Obtain a DC power supply and a Vernier Current Probe. Use connecting wires, with alligator clips, to connect the DC power supply, Current Probe, and the electrodes (See Figure 2). Important: You will not place the electrodes in the cell until Step 10.
8. Connect the Current Probe to LabQuest and choose New from the File menu.
9. Change the data-collection rate to 0.05 samples/second and the data collection duration to 1800 seconds.



Both Constant Current System and Generic Power Supply Users

10. Place the key and the copper strip into the electrolyte solution in the beaker. Make sure that the key is completely immersed in the solution, and keep the two electrodes as far apart as possible.
11. Start data collection. Allow two data points to be collected to establish a baseline. Adjust the current to the 0.2–0.3 A range. Observe the electrolysis. Note the slow deposition of copper on the surface of the key. Data collection will run for 30 minutes.
12. After data collection has run for ~28 minutes, turn off the power supply. Allow the data collection to run to completion. After data collection stops, carefully remove the copper strip and the key from the electrolyte solution. Rinse the two metals with distilled water. Gently dry the copper strip and key so as not to remove copper.
13. Measure and record the mass of the dry copper strip and key.
14. Analyze the graphed data to determine the average current applied during the experiment. Record the average current in your data table.
15. Discard the electrolyte solution and take care of the electrochemical cell as directed.

Initial mass of copper electrode (g)	
Final mass of copper electrode (g)	
Initial mass of key (g)	
Final mass of key (g)	
Average current (A)	
Time of current application (s)	

1. Calculate the number of coulombs of charge that passed through the electrolytic cell. (**Hint:** make sure you are using the proper unit of time for your calculation.).
2. Use your current vs. time data to calculate the theoretical number of moles of copper that should have plated out onto the brass key.
3. Use your data from measuring the mass of the electrode and the key to calculate the actual number of moles of copper that plated out.
4. Calculate the percent yield of copper.
5. Suggest the sources of error in your experiment.
6. Write the oxidation and reduction half-reactions for this process.

electrochemistry formulas

to calculate:

E (cell voltage)

Use this formula

$$E_{\text{cell}} = E_{\text{anode}} - E_{\text{cathode}}$$

(add them up using diagonal rule to make it positive, unless it is electrolysis)

Helpful notes

Spontaneous = positive E

Diagonal rule:

"flip the one that is lower down on the standard reduction table, not related to the periodic table (don't use the formula)

ΔG

$$\Delta G = -RT \ln K = -nFE$$

Spontaneous = positive G

R = 8.314 J/K mol

n = number of electrons transferred (usually 1, 2, or 3)

F = 96,485 c

Determine k

$$k = e^{-\Delta G/RT} = e^{nE/.0257 V}$$

Spontaneous = $k > 1$

How long will it take?

1 coulomb = 1 amp second

1 volt = 1 joule/coulomb

1. Get mol of electrons involved
2. X 96485 = C
3. /A = seconds.

Current required

1 mol electrons = 96,500 coulombs

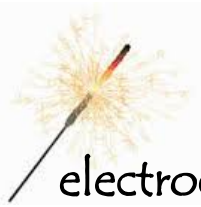
Get coulombs as above, multiply by seconds

Grams required

balancing redox reactions

1. write the net ionic equation
2. separate the equation into two half reactions
3. balance each half-reaction for number and type of atoms and charges. Add H₂O to balance the O atoms and H⁺ to balance the H atoms
4. add the two half reactions together, multiplying each reaction as needed to balance the electrons.
5. if in basic medium for each H⁺ ion add one OH⁻ ion to each side of the equation.

1. Get coulombs (= amps x seconds)
2. Get moles, then g

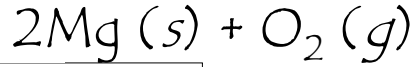


electrochemistry

electrochemical processes are:

oxidation and reduction

redox ex
charges:



losing electrons =

gaining electrons =

oxidizing and reducing agent

What was oxidized?

What was reduced?

What is the oxidizing agent?

What is the reducing agent?

half reactions

write what happened for each reactant. oxidation:

write the reduction half reaction, don't forget to balance electrons

note that combining them yields the complete reaction.

activity series

Mg
Al
Zn
Fe
Pb
Cu
Ag

strongest
reducing agent

most readily
oxidized

weakest
reducing agent

least readily
becomes
oxidized

in general the strongest oxidizing agent is the element with the highest desire for electrons:

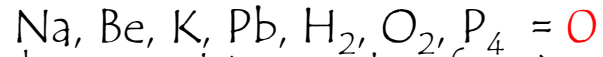
since not all the substance involved are always ions, it is helpful to approximate the charge of all atoms or ions using

oxidation number

definition:

rule for free elements

free elements and single element molecules have an oxidation number of zero.



which substances below have an oxidation number of zero?

rule for ions

for monatomic ions, the oxidation number is equal to the charge on the ion.

rule for oxygen the oxidation number of oxygen is **usually** -2. In H_2O_2 and O_2^{2-} it is -1.

rule for hydrogen The oxidation number of hydrogen is +1 *except* when it is bonded to metals in binary compounds, such as LiH. In these cases, its oxidation number is -1.

rule for group 1 and 2
ions, and fluorine

Group IA metals are +1, IIA metals are +2 and fluorine is always -1.

rule for total oxidation number The sum of the oxidation numbers of all the atoms in a molecule or ion is equal to the charge on the molecule or ion.

Find the oxidation state of carbon in the acetate ion:

The oxidation states are usually in whole numbers, but in some cases, they are in fractional numbers. Find the oxidation state of iron in

(In most cases, atoms of the same element in a given compound have the same oxidation states, but each atom can have a different state than the other. For Fe_3O_4 , two Fe atoms have an oxidation state of +3 and one of +2,)

Provide the oxidation state of each atom or ion.

Ca CaF_2 CaO CaH_2 SO_2 SO_2^- SO_4^{2-} $\text{Cr}_2\text{O}_7^{2-}$ H_2SO_4 S_8

how to name polyvalent compounds? **include the charge using roman numerals**

titanium (IV) chloride

NaCl

NaCN

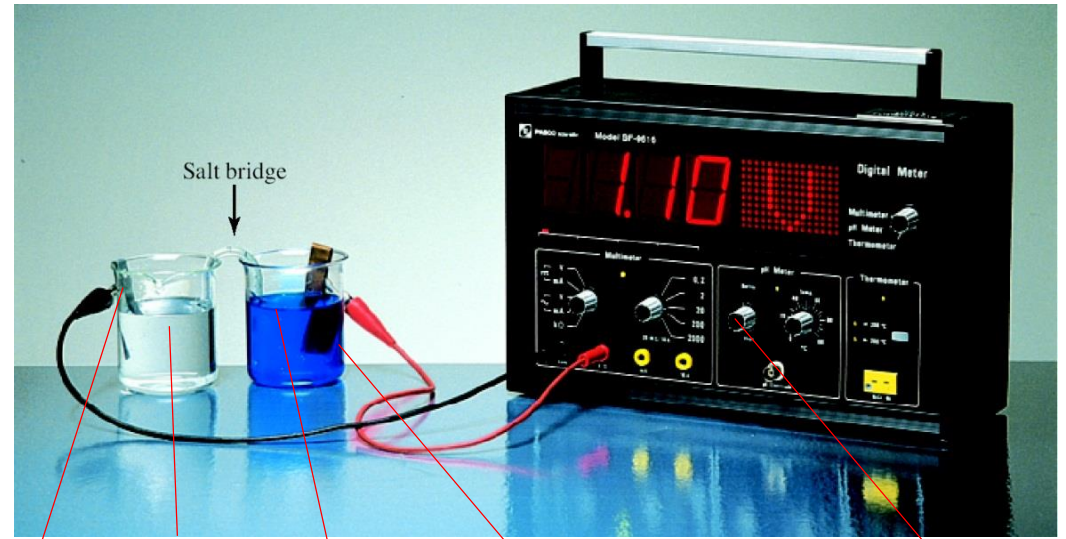
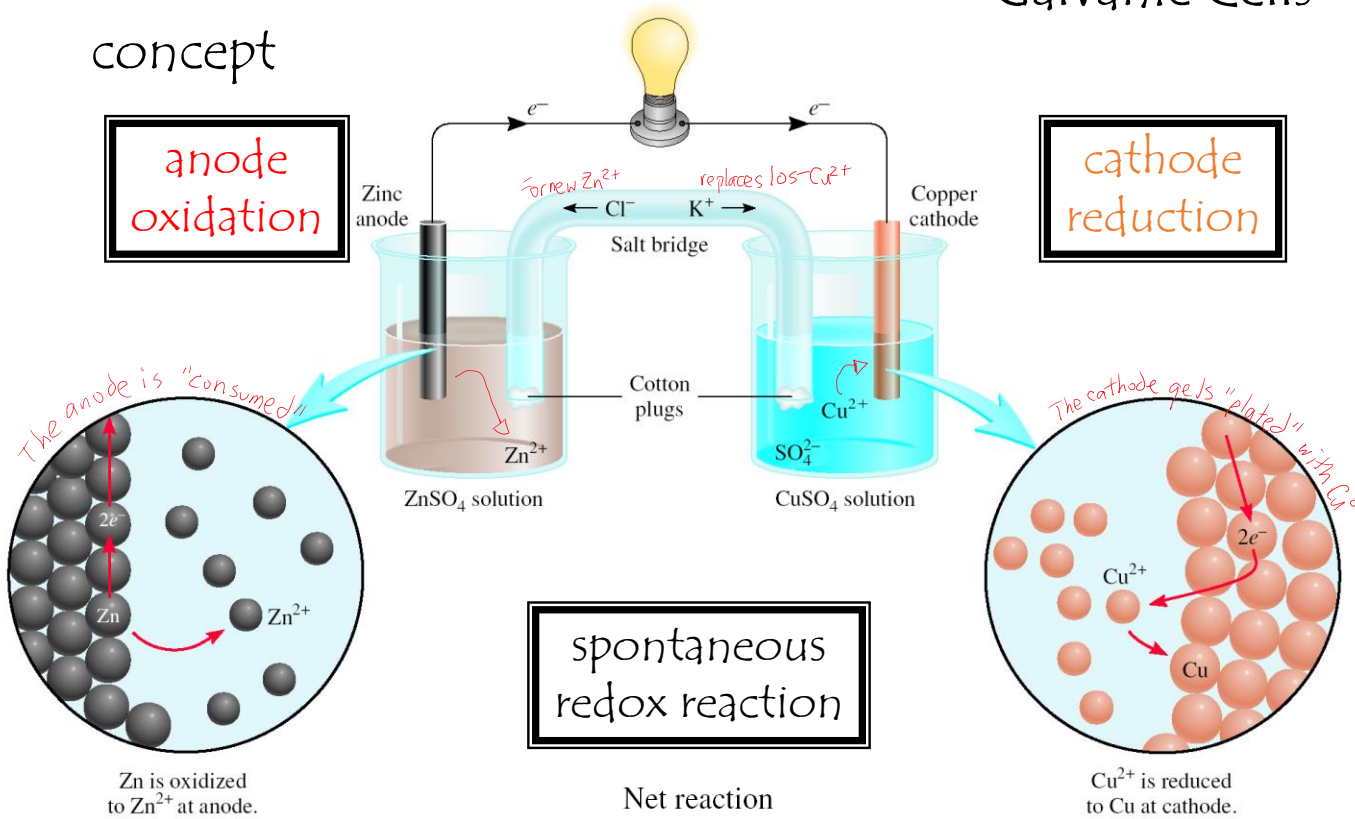
TiCl_2

iron (II) cyanide

Galvanic Cells

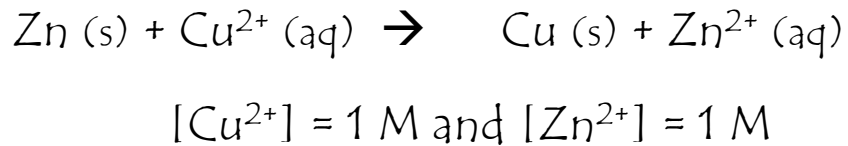
practice

concept

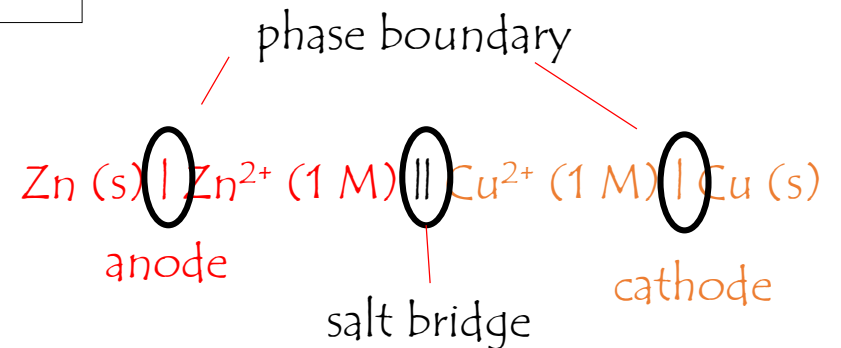


cell voltage = electromotive force (emf) = cell potential

notation



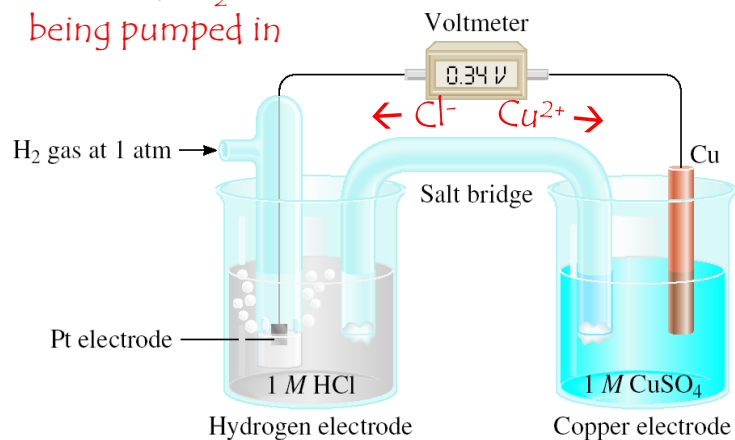
cell Diagram:



Standard Reduction Potentials

$$E_{\text{cell}}^{\circ} = 0.34 \text{ V}$$

Note that H_2 is being pumped in



$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$$

$$E_{\text{cell}}^{\circ} = E_{\text{Cu}^{2+}/\text{Cu}}^{\circ} - E_{\text{H}^+/\text{H}_2}^{\circ}$$

$$0.34 = E_{\text{Cu}^{2+}/\text{Cu}}^{\circ} - 0$$

$$E_{\text{Cu}^{2+}/\text{Cu}}^{\circ} = 0.34 \text{ V}$$

Table 19.1 Standard Reduction Potentials at 25°C*

Half-Reaction	$E^{\circ}(\text{V})$
$\text{F}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{F}^-(\text{aq})$	+2.87
$\text{O}_3(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{O}_2(\text{g}) + \text{H}_2\text{O}$	+2.07
$\text{Co}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Co}^{2+}(\text{aq})$	+1.82
$\text{H}_2\text{O}_2(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow 2\text{H}_2\text{O}$	+1.77
$\text{PbO}_2(\text{s}) + 4\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + 2\text{e}^- \rightarrow \text{PbSO}_4(\text{s}) + 2\text{H}_2\text{O}$	+1.70
$\text{Ce}^{4+}(\text{aq}) + \text{e}^- \rightarrow \text{Ce}^{3+}(\text{aq})$	+1.61
$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}$	+1.51
$\text{Au}^{3+}(\text{aq}) + 3\text{e}^- \rightarrow \text{Au}(\text{s})$	+1.50
$\text{Cl}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{Cl}^-(\text{aq})$	+1.36
$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}$	+1.33
$\text{MnO}_2(\text{s}) + 4\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 2\text{H}_2\text{O}$	+1.23
$\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}$	+1.23
$\text{Br}_2(\text{l}) + 2\text{e}^- \rightarrow 2\text{Br}^-(\text{aq})$	+1.07
$\text{NO}_3^-(\text{aq}) + 4\text{H}^+(\text{aq}) + 3\text{e}^- \rightarrow \text{NO}(\text{g}) + 2\text{H}_2\text{O}$	+0.96
$2\text{Hg}_2^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Hg}_2^{2+}(\text{aq})$	+0.92
$\text{Hg}_2^{2+}(\text{aq}) + 2\text{e}^- \rightarrow 2\text{Hg}(\text{l})$	+0.85
$\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$	+0.80
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq})$	+0.77
$\text{O}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2\text{O}_2(\text{aq})$	+0.68
$\text{MnO}_4^-(\text{aq}) + 2\text{H}_2\text{O} + 3\text{e}^- \rightarrow \text{MnO}_2(\text{s}) + 4\text{OH}^-(\text{aq})$	+0.59
$\text{I}_2(\text{s}) + 2\text{e}^- \rightarrow 2\text{I}^-(\text{aq})$	+0.53
$\text{O}_2(\text{g}) + 2\text{H}_2\text{O} + 4\text{e}^- \rightarrow 4\text{OH}^-(\text{aq})$	+0.40
$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$	+0.34
$\text{AgCl}(\text{s}) + \text{e}^- \rightarrow \text{Ag}(\text{s}) + \text{Cl}^-(\text{aq})$	+0.22
$\text{SO}_4^{2-}(\text{aq}) + 4\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{SO}_2(\text{g}) + 2\text{H}_2\text{O}$	+0.20
$\text{Cu}^{2+}(\text{aq}) + \text{e}^- \rightarrow \text{Cu}^+(\text{aq})$	+0.15
$\text{Sn}^{4+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Sn}^{2+}(\text{aq})$	+0.13
$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$	0.00
$\text{Pb}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Pb}(\text{s})$	-0.13
$\text{Sn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Sn}(\text{s})$	-0.14
$\text{Ni}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Ni}(\text{s})$	-0.25
$\text{Co}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Co}(\text{s})$	-0.28
$\text{PbSO}_4(\text{s}) + 2\text{e}^- \rightarrow \text{Pb}(\text{s}) + \text{SO}_4^{2-}(\text{aq})$	-0.31
$\text{Cd}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cd}(\text{s})$	-0.40
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Fe}(\text{s})$	-0.44
$\text{Cr}^{3+}(\text{aq}) + 3\text{e}^- \rightarrow \text{Cr}(\text{s})$	-0.74
$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn}(\text{s})$	-0.76
$2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) + 2\text{OH}^-(\text{aq})$	-0.83
$\text{Mn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Mn}(\text{s})$	-1.18
$\text{Al}^{3+}(\text{aq}) + 3\text{e}^- \rightarrow \text{Al}(\text{s})$	-1.66
$\text{Be}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Be}(\text{s})$	-1.85
$\text{Mg}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Mg}(\text{s})$	-2.37
$\text{Na}^+(\text{aq}) + \text{e}^- \rightarrow \text{Na}(\text{s})$	-2.71
$\text{Ca}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Ca}(\text{s})$	-2.87
$\text{Sr}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Sr}(\text{s})$	-2.89
$\text{Ba}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Ba}(\text{s})$	-2.90
$\text{K}^+(\text{aq}) + \text{e}^- \rightarrow \text{K}(\text{s})$	-2.93
$\text{Li}^+(\text{aq}) + \text{e}^- \rightarrow \text{Li}(\text{s})$	-3.05

E° is for the reaction as written

The more positive E° the greater the tendency for the substance to be reduced

The half-cell reactions are reversible

The sign of E° changes when the reaction is reversed

Changing the stoichiometric coefficients of a half-cell reaction does not change the value of E°

Using a standard reduction potential table, predict what will happen if molecular bromine (Br_2) is added to a solution containing NaCl and NaI at 25°C . Assume all species are in their standard states.

Possible reactions:	$2\text{NaCl} + \text{Br}_2 \rightarrow 2\text{NaBr} + \text{Cl}_2$	$2\text{NaI} + \text{Br}_2 \rightarrow 2\text{NaBr} + \text{I}_2$
What is happening:	Cl^- is being oxidized, Br_2 reduced	I^- is being oxidized, Br_2 reduced
Half reactions	$2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^- \quad -1.36\text{V}$	$2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^- \quad -0.54\text{V}$
With Standard potentials (sign changed for ox)	$\text{Br}_2 + 2\text{e}^- \rightarrow 2\text{Br}^- \quad 1.09\text{V}$	$\text{Br}_2 + 2\text{e}^- \rightarrow 2\text{Br}^- \quad 1.09\text{V}$
	Negative: wont happen... Will go in the other direction	Positive: the above reaction will happen

note: these reactions don't have to be balanced since it does not affect the voltages.

can also solve using the "diagonal rule"



bromine can oxidize iodide, but not chloride.
or base it on electronegativity position (Cl is the easiest to reduce)

A galvanic cell consists of a Mg electrode in a 1.0 M $\text{Mg}(\text{NO}_3)_2$ solution and a Ag electrode in a 1.0 M AgNO_3 solution. Using a standard reduction potential table, calculate the standard emf of this cell at 25°C.

Spontaneity of Redox Reactions

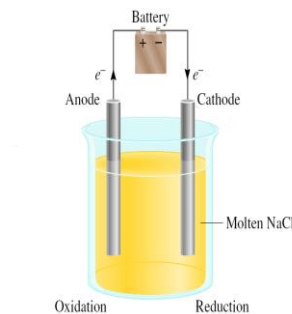
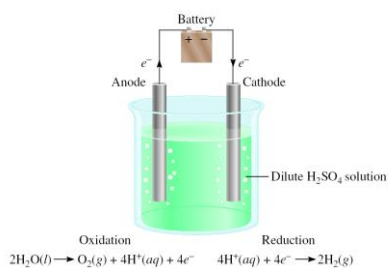
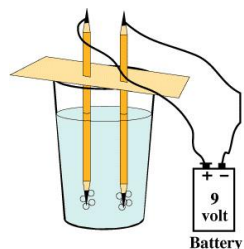
$$\Delta G = -nFE_{\text{cell}} \quad n = \text{number of moles of electrons in reaction}$$

$$\Delta G^{\circ} = -nFE_{\text{cell}}^{\circ} \quad F = 96,500 \frac{\text{J}}{\text{V} \cdot \text{mol}} = 96,500 \text{ C/mol}$$

Table 19.2 Relationships Among ΔG° , K , and E_{cell}°

ΔG°	K	E_{cell}°	Reaction Under Standard-State Conditions
Negative	>1	Positive	Favors formation of products.
0	$=1$	0	Reactants and products are equally favored.
Positive	<1	Negative	Favors formation of reactants.

a non-spontaneous electrochemical process can be performed by providing a power source, this is called electrolysis



Calculate the standard free-energy change for the following reaction at 25°C:

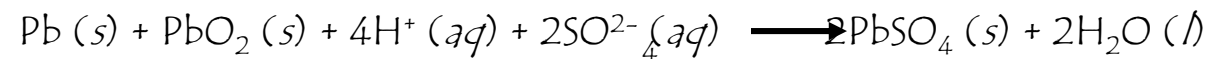
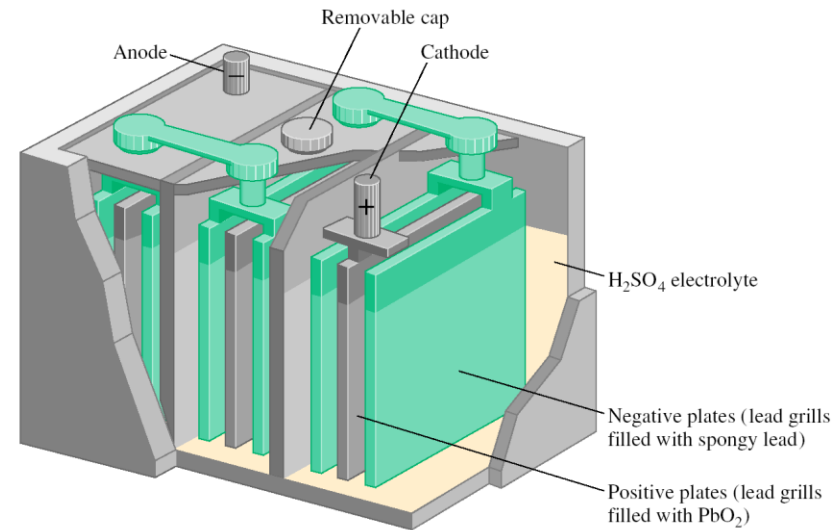


look up the reduction potential and reverse it.
Ignore coefficients (intensive property)

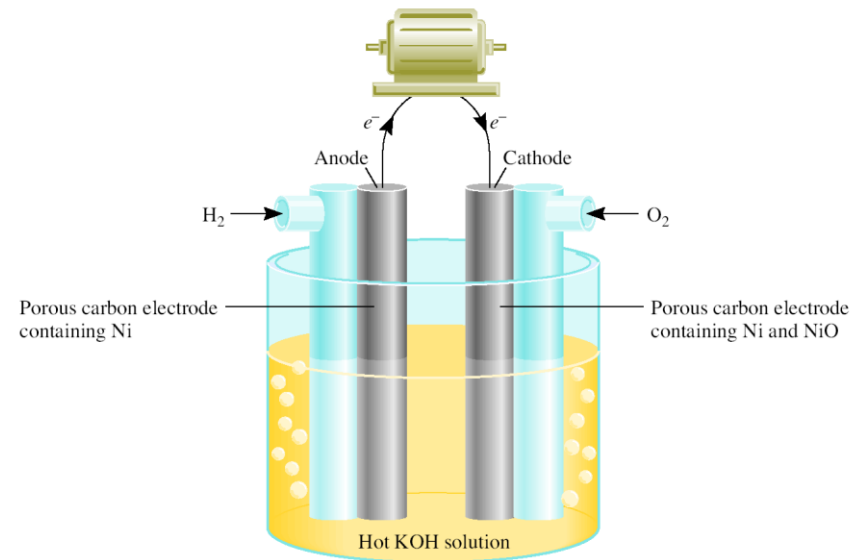
do not triple the voltage, it is an intensive property

Batteries

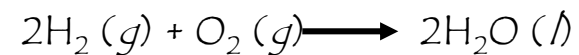
Lead storage
battery



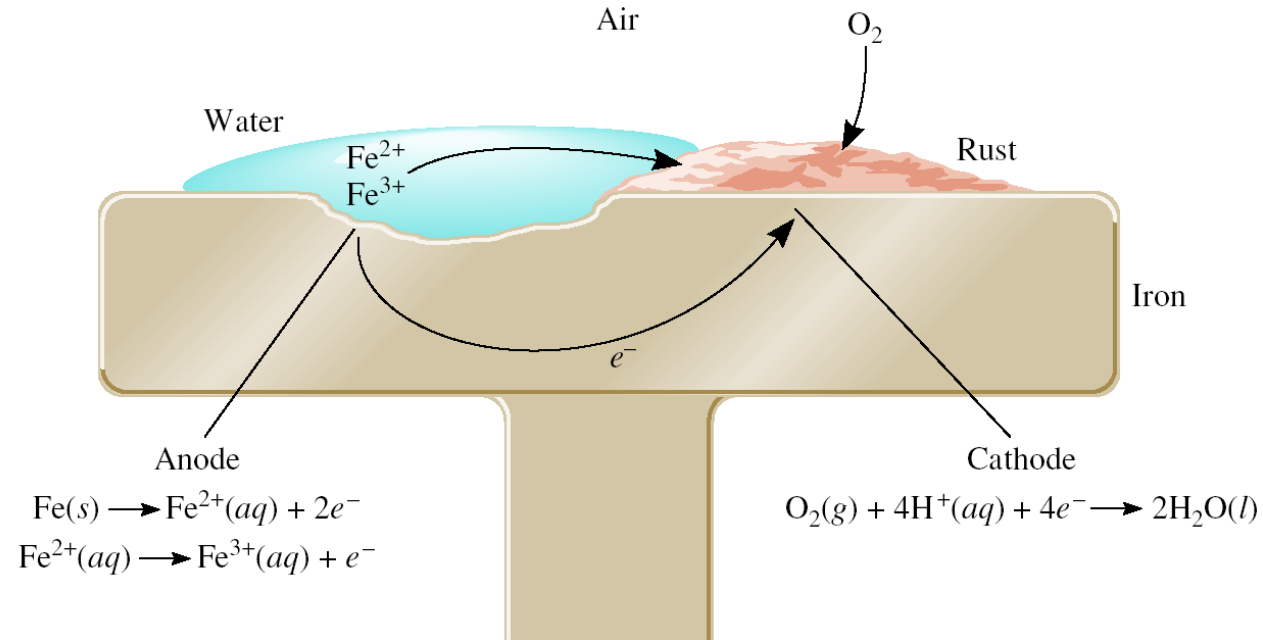
Batteries



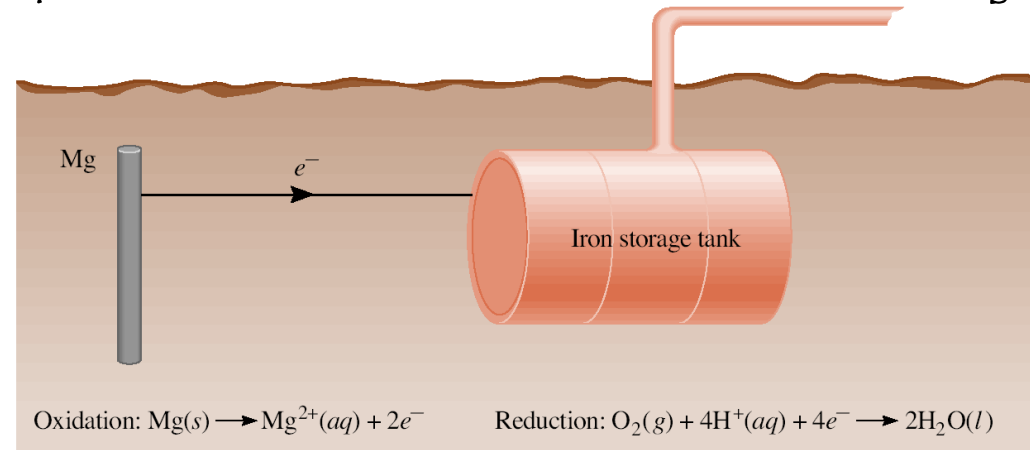
A *fuel cell* is a galvanic cell that requires a continuous supply of reactants to keep functioning



Corrosion:
the deterioration of metals by an electrochemical process.



Example: Cathodic Protection of an Iron Storage Tank



electrochemical measurements
and calculations

volts: voltage
difference in electrical potential
analogy: pressure applied

$$1 \text{ amp} = 1 \text{ coulomb/second}$$

$$1 \text{ joule} = \text{volt} \times \text{coulomb}$$

joules energy

$$1 \text{ mole of electrons} = 96,485 \text{ coulombs}$$

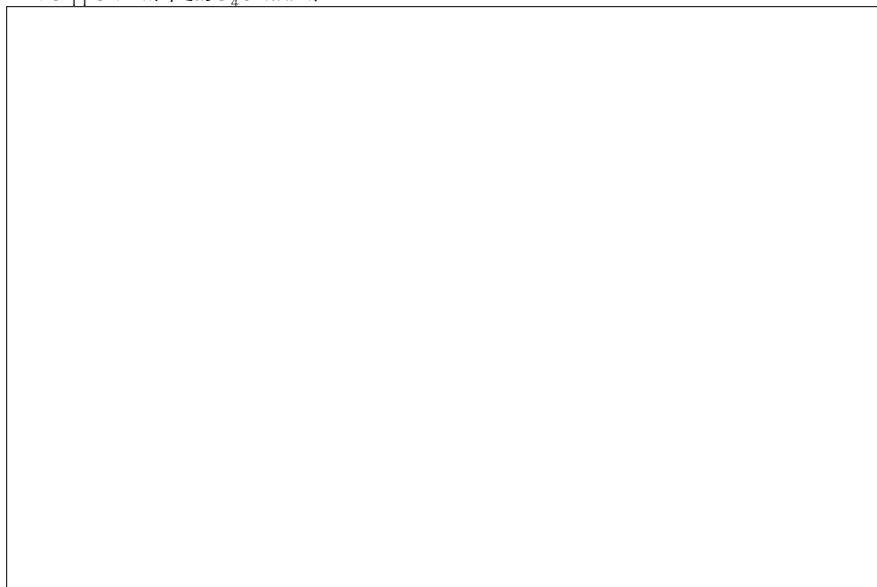
amps current

coulombs charge per electron

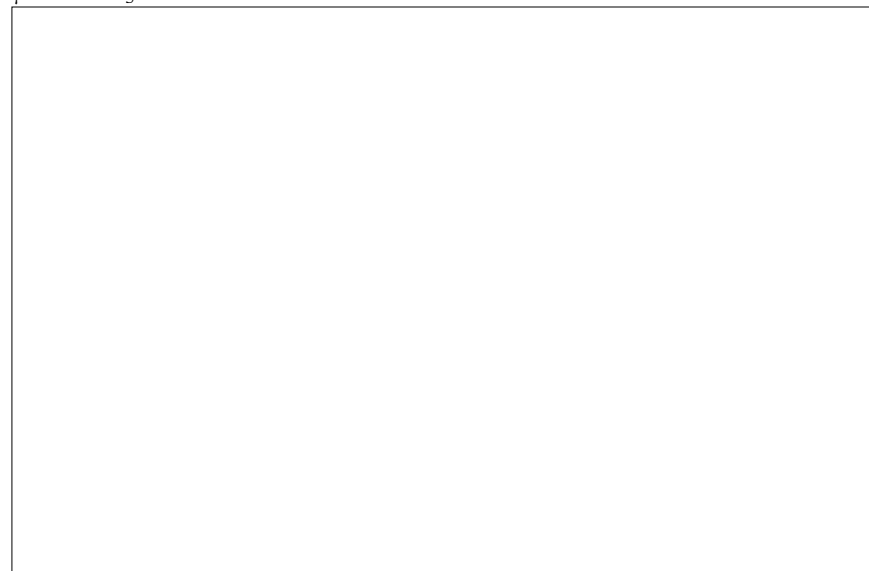
A neat analogy to help understand these terms is a system of plumbing pipes. The voltage is equivalent to the water pressure, the current is equivalent to the flow rate, and the resistance is like the pipe size

Electrochemistry solved problems by Chemteam Present one problem (see following pages)

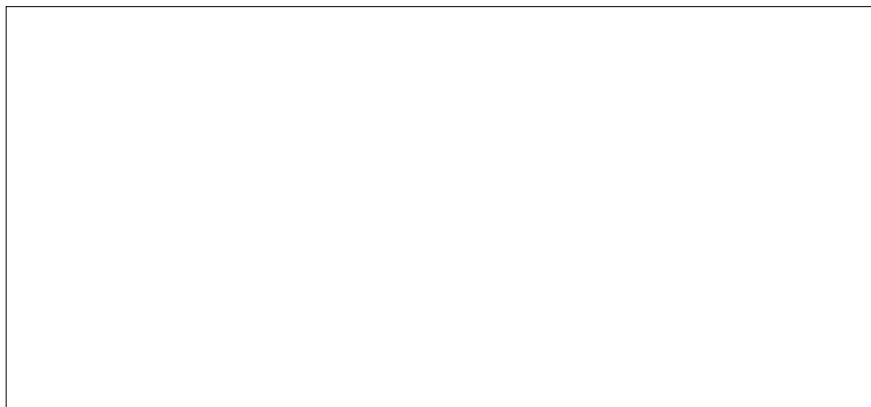
Problem #1: Calculate the quantity of electricity (Coulombs) necessary to deposit 100.00 g of copper from a CuSO_4 solution.



Problem #3: What is the equivalent weight of a metal if a current of 0.2500 amp causes 0.5240 g of metal to plate out a solution undergoing electrolysis in 1 hour? (Comment: One mole of electrons will plate out one equivalent weight of metal.)



Problem #2: How many minutes will take to plate out 40.00 g of Ni from a solution of NiSO_4 using a current of 3.450 amp?



Problem #4: How many hours will it take to plate out copper in 200.0 mL of a 0.0 M Cu^{2+} solution using a current of 0.200 amp?



Problem #1: Calculate the quantity of electricity (Coulombs) necessary to deposit 100.00 g of copper from a CuSO_4 solution.

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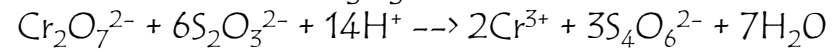
Problem #3: What is the equivalent weight of a metal if a current of 0.2500 amp causes 0.5240 g of metal to plate out a solution undergoing electrolysis in 1 hour? (Comment: One mole of electrons will plate out one equivalent weight of metal.)

Problem #4: How many hours will it take to plate out copper in 200.0 mL of a 0.0 M Cu^{2+} solution using a current of 0.200 amp?

Problem #5: A constant electric current deposits 0.3650 g of silver metal in 12960 seconds from a solution of silver nitrate. What is the current? What is the half reaction for the deposition of silver?

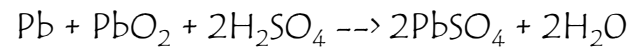


1. What is the reducing agent in the reaction below?



- a) $\text{Cr}_2\text{O}_7^{2-}$
- b) $\text{S}_2\text{O}_3^{2-}$
- c) H^+
- d) Cr^{3+}
- e) $\text{S}_4\text{O}_6^{2-}$

2. Which substance is the oxidizing agent in the reaction below?



- a) Pb
- b) H_2SO_4
- c) PbO_2
- d) PbSO_4

3. What is the oxidation number of each sulfur atom in osmium(V) sulfate?

- a) -2
- b) +3
- c) +6
- d) -1
- e) 0

4. The electrode at which oxidation occurs is called the

- a) oxidizing agent
- b) cathode
- c) reducing agent
- d) anode

5. Which transformation could take place at the anode of an electrochemical cell?

- a) Cr^{3+} to $\text{Cr}_2\text{O}_7^{2-}$
- b) F_2 to F^-
- c) O_2 to H_2O
- d) HAsO_2 to As

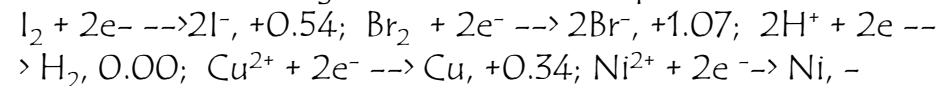
11. Which of the following types of elements is most likely to be good oxidizing agents?

- a) alkali metals
- b) lanthanides
- c) alkaline earth elements
- d) transition metals
- e) halogens

12. Which one of the following is not a good reducing agent?

- a) H_2
- b) Na
- c) O_2
- d) Li
- e) Ca

13. Use the following standard reduction potentials in V:



Which one of the following species could be used to oxidize I^- to I_2 ?

- a) Br_2
- b) H^+
- c) Cu^{2+}
- d) Ni^{2+}

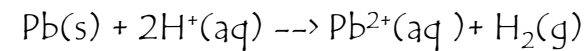
14. Calculate the ΔG° for the reaction of elemental bromine with chloride ion. The E° for the reduction of Br_2 to bromide ion is 1.09V and the E° for the reduction of Cl_2 to chloride ion is 1.36V.

- a) $2.10 \times 10^5 \text{ J}$
- b) $-2.62 \times 10^5 \text{ J}$
- c) 0.27 J
- d) $5.21 \times 10^4 \text{ J}$

15. A spontaneous electrochemical reaction has

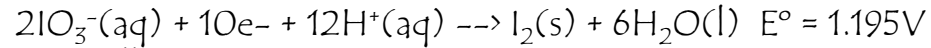
- a) $\Delta G^\circ = 0$, $E^\circ = 0$ and $K \gg 1$
- b) $\Delta G^\circ < 0$, $E^\circ > 0$ and $K > 1$
- c) $\Delta G^\circ > 0$, $E^\circ < 0$ and $K < 1$
- d) $\Delta G^\circ > 0$, $E^\circ < 0$ and $K > 1$
- e) $\Delta G^\circ < 0$, $E^\circ = 0$ and $K \gg 1$

16. E° for the following reaction is 0.13 V. What is the value of ΔG° (in kJ) for the reaction?



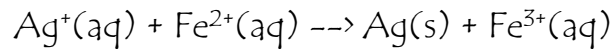
- a) -25
- b) 25
- c) -12
- d) 12

17. Consider the half reaction below. What will happen to the voltage of the half cell if the iodate concentration (IO_3^-) is increased?

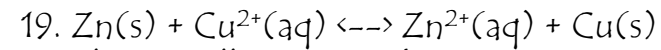


- a) it will increase
- b) it will decrease
- c) it will not change

18. The standard reduction potentials in volts, for Ag^+ to Ag and Fe^{3+} to Fe^{2+} are +0.80 and +0.77 respectively. Calculate K , the equilibrium constant, for the following reaction at 25°C . ($R = 8.3145 \text{ J/K}\cdot\text{mole}$)



- a) 10.0
- b) 2.0
- c) 3.2
- d) 1.0

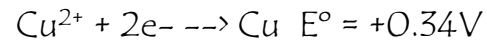
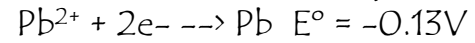


A galvanic cell based on the reaction above was constructed from zinc and copper half-cells. The observed voltage was found to be 1.00 volt instead of the standard cell potential, E° , of 1.10 volts. Which of the following could correctly account for this observation?

- a) The copper electrode was larger than the zinc electrode.
- b) The Zn^{2+} electrolyte was $\text{Zn}(\text{NO}_3)_2$, while the Cu^{2+} electrolyte was CuSO_4 .
- c) The Zn^{2+} solution was more concentrated than the Cu^{2+} solution.
- d) The solutions in the half-cells had different volumes.
- e) The Cu^{2+} solution was more concentrated than the Zn^{2+} solution.

20. A cell is set up with copper and lead electrodes in contact with $\text{CuSO}_4(\text{aq})$ and $\text{Pb}(\text{NO}_3)_2(\text{aq})$, respectively, at 25°C .

The standard reduction potentials are



If sulfuric acid is added to the $\text{Pb}(\text{NO}_3)_2$ solution, forming a precipitate of PbSO_4 , the cell potential

- a) increases
- b) decreases
- c) is unchanged
- d) can't tell what will happen

21. How long will it take to plate out 2.19 g of chromium metal from a solution of Cr^{3+} using a current of 35.2 amps?

- a) 5.77 minutes
- b) 346 minutes
- c) 115 minutes
- d) 1.92 minutes

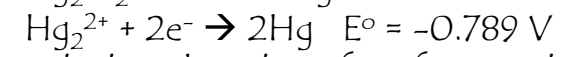
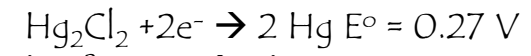
22. What current is required to plate out 1.22 g of nickel from a solution of Ni^{2+} in 1.0 hour?

- a) 65.4 amps
- b) 4010 amps
- c) 1.17 amps
- d) 12.9 amps

23. How many grams of Ca metal could be produced by the electrolysis of molten CaBr_2 using a current of 30.0 amp for 10.0 hours?

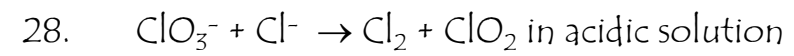
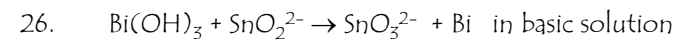
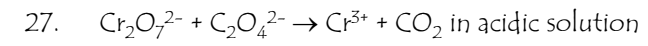
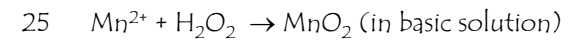
- a) 22.4 g
- b) 448 g
- c) 0.0622 g
- d) 224 g

24. Given



Calculate the value of K_{sp} for Hg_2Cl_2 . Please show your work.

Balance the following redox reaction by the half-reaction method:



29. What is a cell diagram? Write the cell diagram for a galvanic cell consisting of an Al electrode placed in a $1\text{ M Al(NO}_3)_3$ solution and a Ag electrode placed in a 1 M AgNO_3 solution.

32. Calculate the standard emf of a cell that uses Ag/Ag⁺ and Al/Al³⁺ half-cell reactions. Write the cell reaction that occurs under standard-state conditions.

30. Describe the basic features of a galvanic cell. Why are the two components of the cell separated from each other?

31. Discuss the spontaneity of an electrochemical reaction in terms of its standard emf (E°_{cell}).

33. Which of the following reagents can oxidize H₂O to O₂(g) under standard-state conditions? H⁺(aq), Cl⁻(aq), Cl₂(g), Cu²⁺(aq), Pb²⁺(aq), (aq) (in acid).

34. Predict whether the following reactions would occur spontaneously in aqueous solution at 25°C. Assume that the initial concentrations of dissolved species are all 1.0 M.

36. Calculate ΔG° and K_c for the following reactions at 25°C:



35. Which species in each pair is a better reducing agent under standard-state conditions? (a) Na or Li, (b) H₂ or I₂, (c) Fe²⁺ or Ag, (d) Br⁻ or Co²⁺.

37. FRQ 2000A2

2. Answer the following questions that relate to electrochemical reactions.

(a) Under standard conditions at 25°C, $\text{Zn}(s)$ reacts with $\text{Co}^{2+}(aq)$ to produce $\text{Co}(s)$.

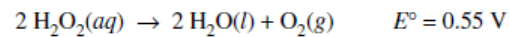
(i) Write the balanced equation for the oxidation half reaction.

(ii) Write the balanced net-ionic equation for the overall reaction.

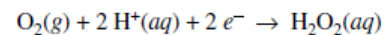
(iii) Calculate the standard potential, E° , for the overall reaction at 25°C.

38. frq

(b) At 25°C, H_2O_2 decomposes according to the following equation.



- (i) Determine the value of the standard free energy change, ΔG° , for the reaction at 25°C.
- (ii) Determine the value of the equilibrium constant, K_{eq} , for the reaction at 25°C.
- (iii) The standard reduction potential, E° , for the half reaction $\text{O}_2(g) + 4 \text{H}^+(aq) + 4 e^- \rightarrow 2 \text{H}_2\text{O}(l)$ has a value of 1.23 V. Using this information in addition to the information given above, determine the value of the standard reduction potential, E° , for the half reaction below.



- (c) In an electrolytic cell, $\text{Cu}(s)$ is produced by the electrolysis of $\text{CuSO}_4(aq)$. Calculate the maximum mass of $\text{Cu}(s)$ that can be deposited by a direct current of 100. amperes passed through 5.00 L of 2.00 M $\text{CuSO}_4(aq)$ for a period of 1.00 hour.

electrochemistry summary problem

you have some copper and silver nitrate

1. Write what will happen when they are mixed
2. type of reaction: _____
3. what got oxidized?
4. what got reduced?
5. What is the oxidizing agent?
6. What is the reducing agent?
7. Draw a design for a battery (galvanic cell) using these reactants
label as many parts as possible
8. write the reaction using the line, double line notation
9. Calculate the standard voltage of your cell
10. Calculate the standard free energy change for your cell
11. Calculate the rate constant for this cell

from Cambridge IB chem

12. A series of experiments was carried out in order to work out an activity series for some metals.

Different metals were added to salt solutions:

1. Zinc and CuSO_4 reacted
2. Zinc and $\text{Pb}(\text{NO}_3)_2$ reacted
3. copper and $\text{Pb}(\text{NO}_3)_2$ did not react
4. zinc and magnesium did not react

solution:

1. Zinc and CuSO_4 reacted (zinc is more reactive than copper)
2. Zinc and $\text{Pb}(\text{NO}_3)_2$ reacted (zinc is more reactive than lead)
3. copper and $\text{Pb}(\text{NO}_3)_2$ did not react (lead is more reactive than copper)
4. zinc and magnesium did not react (magnesium is more reactive than zinc)

magnesium is higher than zinc, zinc is higher than lead, lead is higher than copper (no surprise that zinc is also higher than copper)