Practice MCQs Thermochemistry and Electrochemistry

Questions 1-2 refer to the following information.

The equation and standard cell potential for the decomposition of $H_2O_2(aq)$ in acidic solution at 25°C is given.
The reduction half-reactions for the process are also listed.

$2\,H_2O_2(aq) \rightarrow 2\,H_2O(l) + O_2(g) \qquad E° = 0.55\ \text{V}$

cat $O_2(g) + 4\,H^+(aq) + 4\,e^- \rightarrow 2\,H_2O(l) \quad E° = 1.23\ \text{V}$ (red)

~~$O_2(g) + 2\,H^+(aq) + 2\,e^- \rightarrow H_2O_2(aq)$~~ $E° = ?$

an $2H_2O_2 \rightarrow 2O_2 + 4H^+ + 4e^-$ (ox)

$E° = E_{cat} - E_{an}$

$0.55 = 1.23 - x$

$x = 0.68\ V$

1) Determine the missing standard reduction potential. A) −1.78 V B) −0.68 V C) +0.68 V D) +1.78 V

Note: The voltage value for the decomposition of hydrogen peroxide is incorrectly given as 0.55V – it should be 1.10V. Also the 2 half reactions exchange 4 electrons where the overall reaction only exchanges 2 electrons. This reaction should not have been used. The "correct" answer for the question can be found by treating it as if the exchange was equal.

2) Which of the following is true for the decomposition of $H_2O_2(aq)$?
A) $\Delta G° > 0$ and $K_{eq} > 1$ B) $\Delta G° > 0$ and $K_{eq} < 1$ C) $\Delta G° < 0$ and $K_{eq} > 1$ D) $\Delta G° < 0$ and $K_{eq} < 1$

$E° = +$ so $\Delta G = -$ and $K > 1$

3) A student conducted an experiment to determine $\Delta H°_{rxn}$ for the reaction between $HCl(aq)$ and $NaOH(aq)$. The student ran two trials using the volumes of $HCl(aq)$ and $NaOH(aq)$ indicated in the table, and determined the amount of heat released. Which of the following best explains the relationship between X and Y?

Trial	Volume of 0.10 M HCl	Volume of 0.10 M NaOH	Amount of Heat Released
1	50. mL	50. mL	X
2	100. mL	50. mL	Y

A) Y = 2X, because the volume of $HCl(aq)$ used in trial 2 is twice the volume used in trial 1.
B) Y = X, because the number of moles of acid and base reacting with each other is the same in both trials.
C) Y = 2X/3 because the heat is distributed over more particles in trial 2 than in trial 1.
D) The relationship between X and Y cannot be predicted.

1:1 ratio
100 mL would be excess

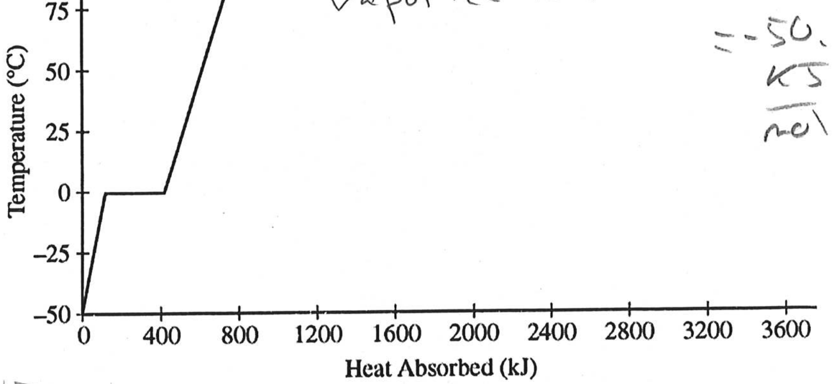
$0.0100\,\text{mL} \left|\frac{1\ \text{mol}}{1\ \text{L}}\right| = 0.0100\ \text{mol}$

4) A student mixes a 10.0 mL sample of 1.0 M $NaOH(aq)$ with a 10.0 mL sample of 1.0 M $HCl(aq)$ in a polystyrene container. The temperature of the solutions before mixing was 20.0°C. If the final temperature of the mixture is 26.0°C, what is the experimental value of $\Delta H°_{rxn}$?
(Assume that the solution mixture has a specific heat of 4.2 J/(g•K) and a density of 1.0 g/mL.)
A) −50. kJ/mol$_{rxn}$ B) −25 kJ/mol$_{rxn}$ C) -5.0×10^4 kJ/mol$_{rxn}$ D) -5.0×10^2 kJ/mol$_{rxn}$

$q = mc\Delta T = (20.0\ g)(4.2)(6°C) = 500\ J$ (water)
total

so rxn is −0.50 kJ / 0.010 mol = −50. kJ/mol

5) At 1.0 atm a sample of ice is heated to liquid water and then to water vapor. The heating curve is shown in the graph. Which of the following lists the signs for the changes in enthalpy and entropy for the process corresponding to segment X, going from left to right?



$\Delta H°$	$\Delta S°$
A) Positive	Negative
B) Positive	Positive more disorder in gas
C) Negative	Negative
D) Negative	Positive

heat absorbed to vaporize

Questions 6-9 refer to the following. $CH_3OH(g) \rightarrow CO(g) + 2\,H_2(g)$ $\Delta H° = +91$ kJ/mol$_{rxn}$

Endo

The reaction represented above goes essentially to completion. The reaction takes place in a rigid, insulated vessel that is initially at 600 K.

$-\Delta G$

6) What happens to the temperature of the contents of the vessel as the reaction occurs?
A) The temperature must increase, because according to Le Châtelier's principle, an increase in temperature causes more products to form.
B) The temperature must decrease, because the reaction takes place at a temperature above room temperature.
C) The temperature must decrease, because the reaction is endothermic.
D) The temperature does not change, because the vessel is insulated.

7) A sample of $CH_3OH(g)$ is placed in the previously evacuated vessel with a pressure of P_1 at 600 K. What is the final pressure in the vessel after the reaction is complete and the contents of the vessel are returned to 600 K?
A) $P_1/9$ B) $P_1/3$ C) P_1 D) $3P_1$

1 mol reactants → 3 mol products

8) What can be inferred about $\Delta S°$ for the reaction at 600 K?
A) It must be positive, since the reaction is thermodynamically unfavorable at 600 K. it isn't
B) It must be negative, since there are more moles of products than reactants. that would be positive
C) It must be positive, since $\Delta G°$ is negative and $\Delta H°$ is positive.
D) It must be negative, since $\Delta G°$ is positive and $\Delta H°$ is positive.

its not

9) Which of the following statements about the bonds in the reactants and products is most accurate?
A) The sum of the bond enthalpies of the bonds in the reactant is greater than the sum of the bond enthalpies of the bonds in the products.
B) The sum of the bond enthalpies of the bonds in the reactant is less than the sum of the bond enthalpies of the bonds in the products.
C) The length of the bond between carbon and oxygen in CH_3OH is shorter than the length of the bond between carbon and oxygen in CO.
D) All of the bonds in the reactant and products are polar.

ΔH = Broken − Formed
$+$ = Greater − smaller

more −

10) Based on the information in the table, which of the following shows the cell potential and the reaction that occurs in a standard Gibbs free energy change for the overall galvanic cell?

	Half-Reaction	$E°$ (V)
ox	$Mg^{2+}(aq) + 2\,e^- \rightarrow Mg(s)$	−2.37
red	$Cr^{3+}(aq) + 3\,e^- \rightarrow Cr(s)$	−0.74

6e⁻

	$E°cell$(V)	ΔG^0 (kJ/mol*rxn*)
A)	+1.63	−157
B)	+1.63	−944
C)	+5.63	−543
D)	+5.63	−3262

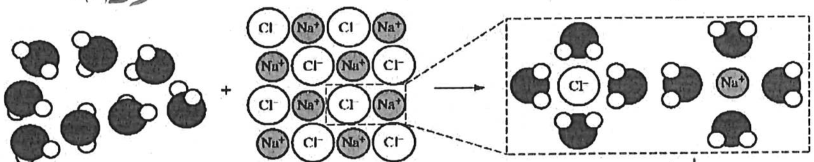
$E_{cell} = E_{red} - E_{ox}$
$= -0.74 - (-2.37)$
$= 1.63$

$\Delta G = -nFE$
$\Delta G = -(6)(96485)(1.63)$
$= -944000$ J
$= -944$ kJ

more random more random Less random

$+\Delta S$ $+\Delta S$ $-\Delta S$

11) The process of dissolution of NaCl(*s*) in $H_2O(l)$ is represented in the diagram above. Which of the following summarizes the signs of ΔH° and ΔS° for each part of the dissolution process?

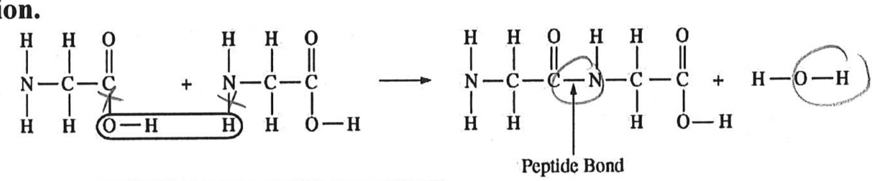


(E need to break IMFs/bonds) $+\Delta H$ $+\Delta H$ $-\Delta H$

	Breaking solvent-solvent interactions		Breaking solute-solute interactions		Forming solute-solvent interactions	
	ΔH°	ΔS°	ΔH°	ΔS°	ΔH°	ΔS°
A)	+	+	+	+	−	−
B)	+	+	+	+	−	+
C)	−	−	−	−	+	+
D)	−	+	−	+	+	−

Questions 12-14 refer to the following information.

Two molecules of the amino acid glycine join through the formation of a peptide bond, as shown. The thermodynamic data for the reaction are listed in the following table.



ΔG^o_{298}	ΔH^o_{298}	ΔS^o_{298}
+15 kJ/mol$_{rxn}$	+12 kJ/mol$_{rxn}$	−10 J/(K·mol$_{rxn}$)

$\Delta G = \Delta H - T\Delta S$

$+ \quad - \; (-)$

12) Under which of the following temperature conditions is the reaction thermodynamically favored?
A) It is only favored at high temperatures.
B) It is only favored at low temperatures.
C) It is favored at all temperatures.
D) It is not favored at any temperature.

13) Based on the bond energies listed in the table, which of the following is closest to the bond energy of the C–N bond?
A) 200 kJ/mol B) 300 kJ/mol C) 400 kJ/mol D) 500 kJ/mol

Bond	Bond Energy (kJ/mol)
C–O	360
N–H	390
O–H	460

$\Delta H = \text{Broken} - \text{Formed}$

$12 = (390 + 360) - (x + 460)$

(N-H) (C-O) (C-N) (O-H)

$x = 278$ kJ

14) Based on the thermodynamic data, which of the following is true at 298 K?
A) $K_{eq} = 0$ B) $0 < K_{eq} < 1$ C) $K_{eq} = 1$ D) $K_{eq} > 1$

$+\Delta G = K < 1$ (reaction favors left)

15) ΔS is positive for which of the following reactions?
A) $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$ 3 → 2 ×
B) $2Hg(l) + O_2(g) \rightarrow 2HgO(s)$ ×
C) $CO_2(g) \rightarrow CO_2(s)$ ×
D) $CaCl_2(s) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq)$ ✓

16) For which of the following processes would ΔS have a negative value?
A) I only B) I and II C) I and III D) II and III

I. $2Fe_2O_3(s) \rightarrow 4Fe(s) + 3O_2(g)$ ✓
II. $Mg^{2+} + 2OH^- \rightarrow Mg(OH)_2(s)$ ×
III. $H_2(g) + C_2H_4(g) \rightarrow C_2H_6(g)$ ×
2 → 1

$\Delta G = \Delta H - T\Delta S$

$-\quad -\quad -(-)$

17) $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

The reaction indicated above is thermodynamically spontaneous at 298 K, but becomes nonspontaneous at higher temperatures. Which of the following is true at 298 K?

~~A)~~ ΔG, ΔH, and ΔS are all positive.

B) ΔG, ΔH, and ΔS are all negative.

C) ΔG and ΔH are negative, but ΔS is positive.

D) ΔG and ΔS are negative, but ΔH is positive

only spontaneous @ low temps

18) Which of the following is always positive when a spontaneous process occurs?

A) ΔS(system)

B) ΔH(surroundings)

C) ΔS(universe)

D) ΔH(system)

Always increasing

Hydrogen gas burns in air according to the equation: $2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(l)$

a) Calculate the standard enthalpy change, ΔH°_{298} , for the reaction represented by the equation above. (The molar enthalpy of formation, ΔH°_f, for $H_2O(l)$ is -285.8 kJ mol^{-1} at 298 K.) *(1pt)*

$\Delta H = \Delta H_f(\text{products}) - \Delta H_f(\text{reactants})$

$= 2(-285.8) - 0$ (elements have a $\Delta H_f = 0$)

$\boxed{\Delta H = -571.6 \text{ kJ/mol}}$

b) Calculate the amount of heat, in kJ, that is released when 10.0 g of $H_2(g)$ is burned in air. *(2pts)*

$\frac{10.0 \text{ g } H_2}{} \cdot \frac{1 \text{ mol } H_2}{2.02 \text{ g } H_2} \cdot \frac{-571.6 \text{ kJ}}{2 \text{ mol } H_2} = \boxed{-1410 \text{ kJ}}$

c) Given that the molar enthalpy of vaporization, ΔH°_{vap}, for $H_2O(l)$ is 44.0 kJ mol^{-1} at 298 K, what is the standard enthalpy change, ΔH°_{298} , for the reaction $2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(g)$? *(1pt)* (Hess' Law)

$(H_2O(l) \rightarrow H_2O(g) \quad \Delta H = 44.0 \text{ kJ}) \times 2 = 88 \text{ kJ}$

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l) \quad \Delta H = -571.6 \text{ kJ}$

$2H_2(g) + O_2(g) \rightarrow H_2O(g) \quad \boxed{\Delta H = -483.6 \text{ kJ/mol}}$

A fuel cell is an electrochemical cell that converts the chemical energy stored in a fuel into electrical energy. A cell that uses H_2 as the fuel can be constructed based on the following half-reactions.

Half-reaction	E° (298 K)	
$2\,H_2O(l) + O_2(g) + 4\,e^- \rightarrow 4\,OH^-(aq)$	0.40 V	red
$2\,H_2O(l) + 2\,e^- \rightarrow H_2(g) + 2\,OH^-(aq)$	$-$ 0.83 V	ox (more neg)

d) Write the equation for the overall cell reaction. *(1pt)* *Show how the 2 "half-rxns add to produce the overall rxn and potential.*

$\cancel{2H_2O(l)} + O_2(g) + \cancel{4e^-} \rightarrow \cancel{4OH^-(aq)}$

$+\ 2H_2(g) + \cancel{4OH^-} \rightarrow \cancel{4}_2 H_2O(l) + \cancel{4e^-} \quad (\times -2)$

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

e) Calculate the potential E (the conditions are not STANDARD) for the cell at 298 K. *(1pt)*

$E_{cell} = E_{red} - E_{ox} = 0.40 - (-0.83) = \boxed{1.23 \text{ V}}$

f) Assume that 0.93 mol of $H_2(g)$ is consumed as the cell operates for 600. seconds.

i) Calculate the number of moles of electrons that pass through the cell. *(1pt)*

$0.93 \text{ mol } H_2 \cdot \frac{4 \text{ mol } e^-}{2 \text{ mol } H_2} = \boxed{1.9 \text{ mol } e^-}$

ii) Calculate the average current, in amperes, that passes through the cell. *(2pts) Hint: I =q/t (from reference sheet)*

$1.9 \text{ mol } e^- \cdot \frac{96485 \text{ C}}{1 \text{ mol } e^-} = 180000 \text{ C}$

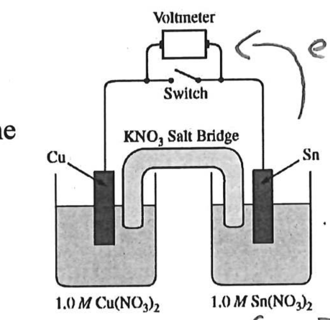
$I = \frac{q}{t} = \frac{180000 \text{ C}}{600 \text{ s}} = \boxed{3.0 \times 10^2 \text{ A}}$

g) Some fuel cells use butane gas, C_4H_{10}, rather than hydrogen gas. The overall reaction that occurs in a butane fuel cell is $2\,C_4H_{10}(g) + 13\,O_2(g) \rightarrow 8\,CO_2(g) + 10\,H_2O(l)$. What is one environmental advantage of using fuel cells that are based on hydrogen rather than on hydrocarbons such as butane? *(1pt)*

H Fuel cells only produce H_2O

Hydrocarbons produce CO_2 as well. This contributes to global warming

A student is given a standard galvanic cell, shown here, that has a Cu electrode and a Sn electrode. As current flows through the cell, the student determines that the Cu electrode increases in mass and the Sn electrode decreases in mass.



$Cu^{+2} + 2e^- \rightarrow Cu$ (red)

$Sn \rightarrow Sn^{2+} + 2e^-$ (ox)

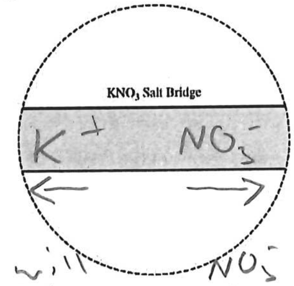
a) Identify the electrode at which oxidation is occurring. Explain your reasoning based on the student's observations. *(1pt)*

The Sn electrode. If Cu is gaining mass, Cu^{+2} is being reduced to Cu(s).

b) As the mass of the Sn electrode decreases, where does the mass go? *(1pt)*

Into the $Sn(NO_3)_2$ solution as it gets oxidized to Sn^{2+}.

c) In the expanded view of the center portion of the salt bridge, draw and label a particle view of what occurs in the salt bridge as the cell begins to operate. Omit solvent molecules and use arrows to show the movement of particles. *(2pts)*



K^+ will offset loss of Cu^{2+}

NO_3^- will offset excess Sn^{2+}

d) A nonstandard cell is made by replacing the 1.0 *M* solutions of $Cu(NO_3)_2$ and $Sn(NO_3)_2$ in the standard cell with 0.50 *M* solutions of $Cu(NO_3)_2$ and $Sn(NO_3)_2$. The volumes of solutions in the nonstandard cell are identical to those in the standard cell.

i) Is the cell potential of the nonstandard cell greater than, less than, or equal to the cell potential of the standard cell? Justify your answer. *(1pt)*

E_{cell} is the same as the standard cell.

Q = 1 in both cases so E_{cell} will be the same.

ii) Both the standard and nonstandard cells can be used to power an electronic device. Would the nonstandard cell power the device for the same time, a longer time, or a shorter time as compared with the standard cell? Justify your answer. *(1pt)*

The nonstandard cell would power the device for a shorter time. The Cu^{2+} would be consumed faster. (There is half as much Cu^{2+} ions.)

e) In another experiment, the student places a new Sn electrode into a fresh solution of 1.0 *M* $Cu(NO_3)_2$.

i) Using information from the table, write a net-ionic equation for the reaction between the Sn electrode and the $Cu(NO_3)_2$ solution that would be thermodynamically favorable. Justify that the reaction is thermodynamically favorable. *(2pts)* *(calculate E°)*

Half-Reaction	$E°$ (V)
$Cu^+ + e^- \rightarrow Cu(s)$	0.52
$Cu^{2+} + 2e^- \rightarrow Cu(s)$	0.34
$Sn^{4+} + 2e^- \rightarrow Sn^{2+}$	0.15
$Sn^{2+} + 2e^- \rightarrow Sn(s)$	−0.14

$Cu^{2+} + 2e^- \rightarrow Cu$ 0.34 V

$+ \quad Sn \rightarrow Sn^{2+} + 2e^-$ +0.14 V

$Cu^{2+} + Sn \rightarrow Cu + Sn^{2+}$

$\boxed{E_{cell} = 0.48\ V}$

(+ E = favorable)

ii) Calculate the value of ΔG°*(include units)* and *Keq* for the reaction. *(2-4pts)* *Explain what E°, ΔG, Keq tell about the rxn.*

$\Delta G = -nFE$

$\Delta G = -(2\ mol)(96485\ \frac{C}{mol})(0.48\ \frac{J}{C})$

$= -93000\ J$

$\boxed{\Delta G = -93\ kJ}$

$-\Delta G$ = spontaneous

$\Delta G = -RT\ln(K)$

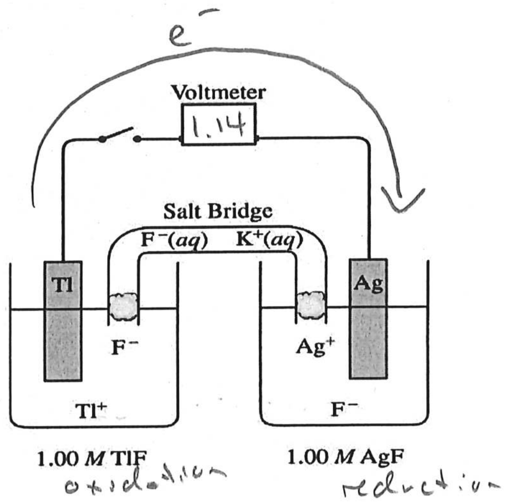
$-93000\ J = -(8.314\ \frac{J}{mol \cdot K})(298\ K)\ln(K)$

$e^{37.} = e^{\ln K}$

$\boxed{K = 1.7 \times 10^{16}}$

K > 1 = spontaneous

Answer the following questions about electrochemical cells. 2009B #6



It is observed that when silver metal is placed in aqueous thallium(I) fluoride, TlF, no reaction occurs. When the switch is closed in the cell represented above, the voltage reading is +1.14 V. Ag + TlF is nonspontaneous so silver must be reduced

a) Write the reduction half-reaction that occurs in the cell. *(1pt)*

$Ag^+ + e^- \rightarrow Ag$

b) Write the equation for the overall reaction that occurs in the cell. *(1pt)*

$Tl + Ag^+ \rightarrow Tl^+ + Ag$

c) Identify the anode in the cell. Justify your answer. *(1pt)*

Tl - it is the site of oxidation

$Tl \rightarrow Tl^+ + e^-$

d) On the diagram above, use an arrow to clearly indicate the direction of electron flow as the cell operates. *(1pt)*

e) Calculate the value of the standard reduction potential for the Tl^+/Tl half-reaction. *(see table in "f")* *(2pts)*

$E_{cell} = E_{red} + E_{ox}$ $E_{ox} = 0.34$ V

$1.14 = 0.80 + E_{ox}$ $\boxed{E_{red} = -0.34 \text{ V}}$

f) Assume that electrodes of pure Pt, Ag, and Ni are available as well as 1.00 M solutions of their salts. Three different electrochemical cells can be constructed using these materials. Identify the two metals that when used to make an electrochemical cell would produce the cell with the largest voltage. Explain how you arrived at your answer. *(1pt)*

Reduction	E^o
$Pt^{2+} + 2e^- \rightarrow Pt$	1.20 V
$Ni^{2+} + 2e^- \rightarrow Ni$	-0.25 V
$Ag^{1+} + e^- \rightarrow Ag$	0.80 V

Pt an Ni

If Pt^{2+} is reduced and Ni is oxidized, it results in the greatest cell potential

$E_{cell} = E_{red} + E_{ox}$

$= 1.20 + 0.25 = 1.45$ V

g) Predict whether Pt metal will react when it is placed in 1.00 M $AgNO_3(aq)$. Justify your answer. *(2pts)*

$Pt + 2Ag^+ \rightarrow Pt^{2+} + 2Ag$

(ox) red

No reaction - Pt^{2+} has a higher reduction potential than Ag^{+1}. It will not be oxidized (react.)

$E_{cell} = E_{red} + E_{ox}$

$= 0.80 + (-1.20) = -0.40$ V = Nonspontaneous

In a laboratory experiment, Pb and an unknown metal Q were immersed in solutions containing aqueous ions of unknown metals Q and X. The following reactions summarize the observations. 2012 #6

Observation 1: $Pb(s) + X^{2+}(aq) \rightarrow Pb^{2+}(aq) + X(s)$

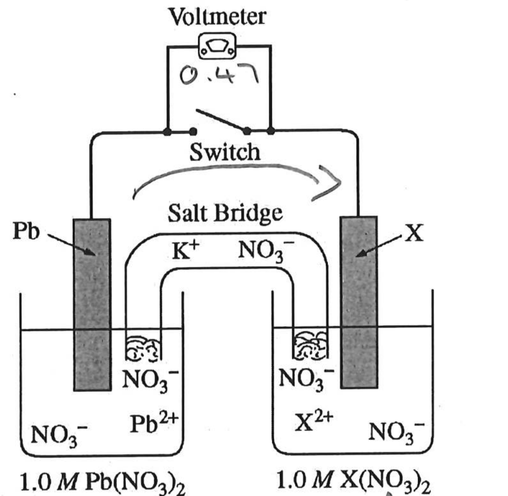
Observation 2: $Q(s) + X^{2+}(aq) \rightarrow$ no reaction

Observation 3: $Pb(s) + Q^{2+}(aq) \rightarrow Pb^{2+}(aq) + Q(s)$

oxidized with X & Q

a) On the basis of the reactions indicated above, arrange the three metals, Pb, Q, and X, in order from least reactive to most reactive on the lines provided. *(1pt)* Q (least reactive metal), X, Pb (most reactive metal)

The diagram below shows an electrochemical cell that is constructed with a Pb electrode immersed in 100. mL of 1.0 *M* $Pb(NO_3)_2(aq)$ and an electrode made of metal X immersed in 100. mL of 1.0 *M* $X(NO_3)_2(aq)$. A salt bridge containing saturated aqueous KNO_3 connects the anode compartment to the cathode compartment. The electrodes are connected to an external circuit containing a switch, which is open. When a voltmeter is connected to the circuit as shown, the reading on the voltmeter is 0.47 V. When the switch is closed, electrons flow through the switch from the Pb electrode toward the X electrode.



ox

red

$Cu^{2+} + 2e^- \rightarrow Cu$

b) Write the equation for the half-reaction that occurs at the anode. *(1pt)*

$2Pb \rightarrow Pb^{2+} + 2e^-$

c) The value of the standard potential for the cell, $E°$, is 0.47 V.

i) Determine the standard reduction potential for the half-reaction that occurs at the cathode. *(1pt)*

$0.47 = E_{red} + 0.13$ $\boxed{E_{red} = 0.34\ V}$

ii) Determine the identity of metal X. Cu *(1pt)* *(use the Table of Standard Reduction Potentials)*

d) Describe what happens to the mass of each electrode as the cell operates. *(1pt)*

The mass of Pb decreases (solid Pb oxidizes to Pb^{2+} in solution)

The mass of Cu(X) increases (Cu^{2+} is reduced to solid Cu)

e) During a laboratory session, students set up the electrochemical cell shown above. For each of the following three scenarios, choose the correct value of the cell voltage and justify your choice.

i) A student bumps the cell setup, resulting in the salt bridge losing contact with the solution in the cathode compartment. Is V equal to 0.47 or is V equal to 0 ? Justify your choice. *(1pt)*

0 - The K^+ no longer offsets the Cu^{2+} being reduced

The charge of that half-cell will become negative and the e^- will stop moving

ii) A student spills a small amount of 0.5 *M* $Na_2SO_4(aq)$ into the compartment with the Pb electrode, resulting in the formation of a precipitate. Is V less than 0.47 or is V greater than 0.47 ? Justify your choice. *(1pt)*

The formation of a precipitate will lower the concentration of Pb^{2+}. This will lower the value of Q and increase the value of Ecell.

$\uparrow E_{cell} = E° - \frac{RT}{nF}\ln(Q)\downarrow$

iii) After the laboratory session is over, a student leaves the switch closed. The next day, the student opens the switch and reads the voltmeter. Is V less than 0.47 or is V equal to 0.47 ? Justify your choice. *(1pt)*

Less than 0.47. The redox reaction will proceed forward causing $[Cu^{+2}]$ to decrease and $[Pb^{+2}]$ to increase. This will increase the value of Q and lower Ecell. $\downarrow E_{cell} = E° - \frac{RT}{nF}\ln(Q)\uparrow$