Exam 3

# Part 1: Multiple Choice (2 points each)

## Directions: Please circle the *best* answer for each of the following questions.

1. A complex ion contains
	1. a central nonmetal ion bound to one or more ligands.
	2. an anion bound to water.
	3. water bound to one or more ligands.
	4. a central metal ion bound to one or more ligands.
	5. ammonia bound to water.
2. Which has the smallest atomic radius?
3. Os
4. La
5. W
6. Hg
7. not enough information
8. For a dead battery E is \_\_\_\_\_\_ and ΔG is \_\_\_\_\_ for the cell reaction.
	1. negative, negative
	2. negative, positive
	3. positive, negative
	4. positive, positive
	5. zero
9. Which of the following pairs of coordination compounds or complex ions are examples of linkage isomers?
10. [Fe(CO)5NO2]2+ and [Fe(CO)5ONO]2+
11. [Fe(NH3)2(H2O)4]Cl2 and [Fe(NH3)2(H2O)4]Br2
12. [Fe(NH3)2(H2O)4]Cl2 and [Fe(NH3)4(H2O)2]Cl2
13. [Cu(CO)5Br]Cl and [Cu(CO)5Cl]Br
14. [MnCl3Br]2- and [MnClBr3]2-
15. Which of the following can function as a bidentate ligand?
	1. CN-
	2. SH-
	3. H2O
	4. H2NCH2CH2NH2
	5. CO
16. ΔG = ΔGº for a reaction
	1. if Q = K
	2. if Q = 1
	3. at STP
	4. at the start of a reaction
	5. at the end of a reaction
17. Doubling all the coefficients in the equation for the cell reaction
	1. Doubles both E° and ΔG°
	2. Double E°, but does not change ΔG°
	3. Double ΔG°, but does not change E°
	4. Does not change E° or ΔG°
	5. Not enough information
18. Determine the cell notation for the redox reaction given below:

3 Cl2 (g) + 2 Fe (s) → 6 Cl- (aq) + 2 Fe3+ (aq)

1. Cl2 (g) | Cl- (aq) | Pt (s) || Fe (s) | Fe3+ (aq)
2. Cl- (aq) | Cl2 (g) | Pt (s) || Fe3+ (aq) | Fe (s)
3. Fe3+ (aq) | Fe (s) || Cl- (aq) | Cl2 (g) | Pt (s)
4. Fe (s) | Cl2 (g) || Fe3+ (aq) | Cl- (aq) | Pt (s)
5. Fe (s) | Fe3+ (aq) || Cl2 (g) | Cl- (aq) | Pt (s)
6. Apart from environmental concerns, why is it unwise to dispose of chemical down the sink?
	1. Solids that are insoluble in water will remain the sink trap and perhaps react with subsequent chemical poured down the sink.
	2. Some chemicals can produce toxic gases when mixed with other chemicals in the sink trap.
	3. Some chemical may react unexpectedly with other chemicals in the sink trap.
	4. b and c
	5. All of the above
7. When using the balances
	1. you should never scoop reagents into or out of the balance pan.
	2. you should close the doors and windows to get a stable reading.
	3. you should always tare your weigh boat before adding reagents.
	4. a and b only
	5. all of the above

# Part 2: Short Answer

## Directions: Answer each of the following questions. Be sure to use complete sentences where appropriate. For full credit be sure to show all of your work.

Use the following reduction potentials in the problems below:

Cl2 (g) + 2 e- → 2 Cl- (aq) E° = 1.36 V

Ag+ (aq) + e- → Ag (s) E° = 0.7994 V

Cu2+ (aq) + 2 e- → Cu (s) E° = 0.337 V

Pb2+ (aq) + 2 e- → Pb (s) E° = -0.13 V

Ni2+ (aq) + 2 e- → Ni (s) E° = -0.257 V

Al3+ (aq) + 3 e- → Al (s) E° = -1.662 V

1. A Voltaic cell is made for the reduction of silver ions with copper metal at 25 °C (6 points).
	1. Write the balanced reaction:

Oxidation: Cu (s) → Cu2+ (aq) + 2 e-

Reduction: + **(**Ag+ (aq) + e- → 2 Ag (s)**) × 2**

2 Ag+ (aq) + 2 e- + Cu (s) → Cu2+ (aq) + 2 e- + 2 Ag (s)

2 Ag+ (aq) + Cu (s) → Cu2+ (aq) + 2 Ag (s)

* 1. Calculate the standard cell potential, E°cell, at 25 °C.

$$E\_{cell}^{°}=E\_{cathode}^{°}-E\_{anode}^{°}$$

$$E\_{cell}^{°}=0.7994 V-0.337 V=0.4624 V≈0.462 V$$

* 1. Calculate ∆rG° in kJ for this reaction.

$$∆\_{r}G^{°}=-nFE^{°}$$

$$∆\_{r}G^{°}=-\left(2 mol e^{-}\right)\left(96485 \frac{C}{mol e^{-}}\right)\left(0.462 V\right)=-89152.14 C V=-89.2 kJ$$

1. Why can’t a wire perform the same function as a porous separator in an electrochemical cell (4 points)?

A wire cannot allow the movement of ions from one side of the voltaic cell to the other to maintain electrical neutrality in the half-cells.

1. How can the corrosion of iron be prevented (4 points)?

The most obvious way to prevent rust formation is to keep iron dry. Without water, the redox reaction cannot occur.

A second way of preventing rust formation is to coat the iron with a substance that is impervious to water. Cars, for example, are painted and sealed to prevent rust. A scratch in the paint, however, can lead to rusting of the underlying iron.

Rust can also be prevented by placing a sacrificial electrode in electrical contact with the iron. The sacrificial electrode must be composed of a metal that oxidizes more easily than iron. The sacrificial electrode then oxidizes in place of the iron (just as the more easily oxidizable species in a mixture will be the one to oxidize), protecting the iron from oxidation.

Another way to protect iron from rusting is to coat it with a metal that oxidizes more easily than iron. Galvanized nails, for example, are coated with a thin layer of zinc. Since zinc has a more positive half-cell potential for oxidation, it will oxidize in place of the underlying iron (just as a sacrificial electrode does). The oxide of zinc is not crumbly and remains on the nail as a protective coating.

1. A Voltaic cell is set up at 25 °C with the half-cells Al3+(0.0010 M) |Al and Ni2+(0.50 M) |Ni (8 points).
2. Write an equation for the reaction that occurs when the cell generates an electric current.

Oxidation: **(**Al (s) → Al3+ (aq) + 3 e-**) × 2**

Reduction: + **(**Ni2+ (aq) + 2 e- → Ni (s)**) × 3**

2 Al (s) 3 Ni2+ (aq) + 6 e- → 2 Al3+ (aq) + 6 e- + 3 Ni (s)

2 Al (s) 3 Ni2+ (aq) → 2 Al3+ (aq) + 3 Ni (s)

1. What is the standard cell potential, E°cell?

$$E\_{cell}^{°}=E\_{cathode}^{°}-E\_{anode}^{°}$$

$$E\_{cell}^{°}=-0.257 V-\left(-1.662 V\right)=1.405 V$$

1. Determine the cell potential of the concentration cell.

$$E\_{cell}=E^{°}-\frac{RT}{nF}lnQ=E^{°}-\frac{RT}{nF}ln\left(\frac{[Al^{3+}]^{2}}{[Ni^{2+}]^{3}}\right)$$

$$E\_{cell}=1.405 V-\frac{\left(8.3145\frac{J}{mol K}\right)\left(298 K\right)}{\left(6 mol e^{-}\right)\left(96485\frac{C}{mol e^{-}}\right)}ln\left(\frac{(0.0010 M)^{2}}{(0.50 M)^{3}}\right)$$

$$E\_{cell}=1.405 V-(-0.050 V)=1.455 V$$

1. Consider the following voltaic cell (12 points):



* 1. Determine the direction electron flow and label the anode and cathode.
	2. Label each electrode as positive or negative.
	3. Indicate the direction of anion and cation flow in the salt bridge.



* 1. Calculate the standard cell potential, E°cell.

$$E\_{cell}^{°}=E\_{cathode}^{°}-E\_{anode}^{°}$$

$$E\_{cell}^{°}=1.36 V-\left(-0.13 V\right)=1.49 V$$

* 1. Write a balanced equation for the overall reaction.

 Oxidation: Pb (s) → Pb2+ (aq) + 2 e-

Reduction: + Cl2 (g) + 2 e- → 2 Cl- (aq)

Pb (s) + Cl2 (g) → Pb2+ (aq) + 2 Cl- (aq)

* 1. Write the line notation for the reaction. Pb (s) | Pb2+ (aq) || Cl- (aq) | Cl2 (g) | Pt (s)
1. A Downs cell operating at 77.0 A produces 31.0 kg of Na (10 points):

2 NaCl (s) → 2 Na (s) + Cl2 (g)

* 1. What volume of Cl2 (g) is produced at 1.0 atm and 540 °C?

$$31.0 kg Na×\frac{1000 g}{1 kg}×\frac{1 mol Na}{22.990 mol }×\frac{1 mol Cl\_{2}}{2 mol Na}=674.2061766 mol Cl\_{2}$$

$$PV=nRT⟹V=\frac{nRT}{P}=\frac{\left(674.2061766 mol \right)\left(0.08206 \frac{L atm}{mol K}\right)\left(540+273\right)}{1.0 atm}=4.5×10^{4} L$$

* 1. How many coulombs were passed through the cell?

$$674.2061766 mol Cl\_{2}×\frac{2 mol e^{-}}{1 mol Cl\_{2}}×\frac{96485 C}{mol e^{-}}=1.30×10^{8} C$$

Or

$$31.0 kg Na×\frac{1000 g}{1 kg}×\frac{1 mol Na}{22.990 mol }×\frac{1 mol e^{-}}{1 mol Na}×\frac{96485 C}{mol e^{-}}=1.30×10^{8} C$$

* 1. How long did it operate?

$$1.30×10^{8} C×\frac{s A}{1 C}×\frac{1}{77.0 A}=1.69×10^{6} s or 19.6 days $$

1. The anode in an electrochemical cell is defined as the electrode where oxidation takes place. Why is the anode in an electrolytic cell connected to the positive (+) terminal of an external supply, whereas the anode in a voltaic cell battery is connected to the negative (-) terminal (5 points).

The signs of the electrodes in a cell indicate the direction of electron flow. In a voltaic cell, the electrons are produced at the anode so a negative (-) charge builds up there; in an electrolytic cell, electrons are being forced onto the cathode so that it builds up a negative (-) charge with the anode having a positive (+) charge. The flow of electrons in the outside circuit is reversed in an electrolytic cell compared to the flow in a voltaic cell.

1. For each species below complete the table (10 points):

|  |  |  |  |
| --- | --- | --- | --- |
|  | Name | Coordination Number of Central Atom | Geometry of Complex Ion |
| [Rh(NH3)5I]I2(Rh is rhodium) | Pentaammineiodiorhodium(III) iodide | 6 | octahedral |
| [Au(CN)2]- | dicyanoaurate(I) ion | 2 | linear |
| [Ni(H2O)2Cl2] (only one isomer) | Diaquadichloronickel(II) | 4 | tetrahedral |

1. Draw all possible diastereoisomers of [Cr(C2O4)2(H2O)2]-. Which can exist as a pair of enantiomers (6 points)?

10. Deposits of elemental sulfur are often seen near active volcanoes. Their presence there may be due to the following reaction of SO2 with H2S (15 points):

SO2 (g) + 2 H2S (g) → 3 S (s) + 2 H2O (l)

|  |  |  |
| --- | --- | --- |
| Substance | ∆H° | ∆S° |
| SO2 (g) | -296.8 kJ/mol | 248.2 J/mol K |
| H2S (g) | -20.17 kJ/mol | 205.6 J/mol K |
| S (s) | 0 kJ/mol | 32.1 J/mol K |
| H2O (l) | -241.8 kJ/mol | * 1. J/mol K
 |

* + - 1. Calculate ∆H° for this reaction.

∆Hrxn° = [(3 mol S)(∆H° S (s)) + (2 mol H2O)(∆H° H2O (l))] - [(1 mol SO2 (g))(∆H° SO2 (g)) + (1 mol H2S)(∆H° H2S (g))]

∆Hrxn° = [(3 mol S)(0 kJ/mol) + (2 mol H2O)(-241.8 kJ/mol)] - [(1 mol SO2 (g))(-296.8 kJ/mol) + (2 mol H2S)(-20.17 kJ/mol)]

∆Hrxn° = -146.46 kJ ≈ -146.5 kJ

* + - 1. Is this reaction exothermic or endothermic? \_\_\_\_\_exothermic\_\_\_\_\_\_\_
			2. Calculate ∆S° for this reaction.

∆Srxn° = [(3 mol S)(∆S° S (s)) + (2 mol H2O)(∆S° H2O (l))] - [(1 mol SO2 (g))(∆S° SO2 (g)) + (1 mol H2S)(∆S° H2S (g))]

∆Srxn° = [(3 mol S)(32.1 J/mol K) + (2 mol H2O)(188.8 J/mol K)] - [(1 mol SO2 (g))(248.2 J/mol K) + (2 mol H2S)(205.6 J/mol K)]

∆Srxn° = -185.5 J/K

* + - 1. Does this reaction result in an increase or decrease in entropy? Decrease
			2. Predict the temperature range over which the reaction is spontaneous.

$$∆G\_{rxn}^{°}= ∆H\_{rxn}^{°}-T∆S\_{rxn}^{°}=0 ⇒T= \frac{∆H\_{rxn}^{°}}{∆S\_{rxn}^{°}}$$

$$T=\frac{-146.5 kJ}{-185.5\frac{J}{K}}×\frac{1000 J}{1 kJ}=789.5 K or 516.4 ℃$$

The reaction is spontaneous for temperatures less than 516.4 °C.