Exam 3

# Part 1: Multiple Choice (2 points each)

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

1. Safety goggles must be worn when
2. you are using chemicals.
3. anyone is using flames.
4. anyone is using chemicals.
5. if my teacher is wearing them.
6. all of the above
7. A ligand is any \_\_\_\_\_\_\_ forming a coordinate bond to a metal cation.
	1. Lewis acid
	2. ion
	3. Lewis base
	4. organic compound
	5. species
8. Which ligand below is not correctly named?
	1. H2O, hydro
	2. NH3, ammine
	3. CN-, cyano
	4. CO, carbonyl
	5. OH-, hydroxo
9. 1,10-phenanthroline, which is shown below, is an example of a \_\_\_\_\_\_\_\_ ligand. 
	1. monodentate
	2. bidentate
	3. tridentate
	4. good
	5. poor



1. A pH meter uses an electrode arrangement that provides a voltage that depends on [H+], in accord with the Nernst equation. The following plot illustrates this dependence. What effect would changing temperature have on the measurement and this plot?
	1. A change of temperature would have no effect.
	2. An increase in temperature would increase the slope.
	3. An increase in temperature would decrease the slope.
	4. An increase in temperature would increase the *y* intercept.
	5. An increase in temperature would decrease the *y* intercept.
2. For an octahedral complex what metal d orbitals are directly towards the ligand?
	1. dxy, dxz
	2. dxy, dxz, dyz
	3. $d\_{z^{2}}, d\_{x^{2}-y^{2}}$
	4. $d\_{z^{2}}, d\_{xz}, d\_{yz}$
	5. none of the above
3. The unit of current, ampere (A), is defined as \_\_\_\_\_\_\_\_
	1. 1 C.
	2. 1 C/s.
	3. 1 mol of electrons.
	4. 10 C/s.
	5. 96,500 C/s.
4. The standard cell potential for the nickel–cadmium battery is 1.35 V, and the cell reaction can be written as

2 NiO(OH) (*s*) + 2 H2O (*l*) + Cd (s)  2 Ni(OH)2 (*s*) + Cd(OH)2 (*s*)

Which one of the following statements do you expect to be true based on the Nernst equation? Note *Q* = reaction quotient, and *K* = equilibrium constant for the cell reaction.

* 1. As the battery is used, the cell voltage approaches zero because *Q* approaches *K*, in value.
	2. When the battery no longer works, the cell voltage is zero because *Q* = *K*.
	3. As the battery is used, the cell voltage does not change because *Q* equals 1.
	4. When the battery is fully charged, *Q* > *K*.
	5. When the battery is fully charged, *Q* < *K*.



1. Which statement about this concentration cell is *incorrect*?
	1. The electrode on the left side is the anode.
	2. The electrode on the left side is the electron.
	3. Anions in the salt bridge migrate to the left side.
	4. Hydrogen ions are reduced on the right side.
	5. The pH of the solution on the right side is higher than the pH on the left side.
2. Which statement about corrosion is *not* correct?
	1. Corrosion is promoted by the presence of water.
	2. Corrosion is promoted by the presence of electrolytes.
	3. Corrosion is promoted by contact between dissimilar metals.
	4. Surface oxidation protects metals such as iron, stainless steel, and aluminum from corrosion.
	5. Cathodic protection prevents corrosion of an object by attaching it to a metal that serves as the cathode of the resulting electrochemical cell.

# 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.

1. What determines whether a transition metal ion is in a high-spin or a low-spin electron configuration (3 points)?

The magnitude of the crystal field splitting (A) compared to the electron-pairing energy in a *d* orbital (B). If A is larger than B, low spin results. If B is larger than A, then high spin results.

1. Briefly account for each of the following observations (4 points):
	1. Atomic radii decreases in the order Sc > Ti > V

The decrease in radii with increasing atomic number is expected because the added d electrons only partially shield the added nuclear charge. As a result, Zeff increases. With increasing Zeff, the electrons are more strongly attracted to the nucleus, and atomic size decreases.

* 1. Densities increase in the order Ti < V < Cr

The densities of the transition metals are inversely related to their atomic radii.

1. Although Cl- is a weak-field ligand and CN- is a strong-field ligand, [CrCl6]3­- and [Cr(CN)6 ]3- exhibit approximately the same amount of paramagnetism (7 points).
	1. Explain.

Cr3+ is a 3d3 ion. Regardless of the crystal field splitting energy, the three electrons singly occupy the three lower energy d orbitals.

* 1. Name each complex.

[CrCl6]3­- hexachlorochromate(III) ion

[Cr(CN)6 ]3- hexacyanochromate(III) ion

* 1. What is the geometry around each ion? \_\_\_\_\_\_\_\_\_\_\_\_\_\_octahedral\_\_\_\_\_\_\_\_\_\_
1. The Ni2+ (aq) cation is green, but Zn2+ (aq) is colorless. Explain (4 points).

Ni2+ (aq) is green because the Ni2+ ion can absorb light, which promotes electrons from the filled d orbitals to the higher energy half-filled d orbitals. Zn2+ (aq) is colorless because the d orbitals are completely filled and no electrons can be promoted, so no light is absorbed.

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



1. In water, the titanium(III) ion, in hexaaquatitanium(III), has a broad absorption band at 500 nm (10 points).
	1. What color light is absorbed by the ion? \_\_\_blue-green\_\_\_\_\_\_
	2. What color light is transmitted by the ion? \_\_red-violet\_\_\_\_\_\_
	3. What the formula of the ion? \_\_\_[Ti(H2O)6]3+\_\_\_\_\_\_
	4. Calculate the crystal field splitting energy, Δ, in kJ/mol.

$$∆=\frac{hc}{λ}=\frac{\left(6.626×10^{-34} Js\right)(3.00×10^{8}\frac{m}{s})}{500×10^{-9} m}=3.9756×10^{-19}\frac{J}{ion}$$

$$3.9756×10^{-19}\frac{J}{ion}×\frac{6.022×10^{23} ions}{1 mol}×\frac{1 kJ}{1000 J}=239\frac{kJ}{mol}$$

1. Would you expect the following compounds to have cis-trans isomers, mer-fac isomers, or no geometric isomers (5 points)?
	1. [CuCl2Br2]2- cis-trans isomers
	2. [Fe(NH3)5Cl]2+ no geometric isomers
	3. [Ti(H2O)4(OH)2]2+ cis-trans isomers
	4. [Co(NH3)3Cl3] mer-fac isomers
2. 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. The following cell reactions occur spontaneously (8 points):

B + A+ → B+ + A

C + A+ → C++ A

B + C+ → B+ + C

1. Arrange the following reaction half-reactions in order of decreasing tendency to occur: A++ e- → A, B++ e- → B, C+ + e- → C.

Looking at the equations above B > A, C > A, B > C, therefore, B > C > A, so

A++ e- → A

C+ + e- → C

B++ e- → B

1. Which of these substances (A, A+, B, B+, C, C+) is the strongest oxidizing agent? \_\_A+
2. Which is the strongest reducing agent? \_\_B
3. Which of the three cell reactions delivers the highest voltage? \_\_\_\_ B + A+ → B+ + A
4. A current of 15.0 A electroplated 50.0 g of hafnium metal from an aqueous solution in 2.00 hours. What was the oxidation number of hafnium in the solution (8 points)?

$$15.0 A×\frac{1 C}{1 A s}×\frac{3600 S}{1 hr}×\frac{2.00 hr}{1}×\frac{1 mol e^{-}}{96500 C}×\frac{178.5 g Hf}{1 mol Hf}×\frac{1}{50.0 g Hf}=\frac{4 mol e^{-}}{1 mol Hf}$$

Hf4+

1. An electrochemical cell is based on the following two half-reactions:

Use the following reduction potentials in the problems below:

MnO4- (aq) + 4 H+ (aq) + 3 e- → MnO2 (s) + 2 H2O (l) E° = 1.68 V

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

Red: MnO4- (aq) (1.50 M) + 4 H+ (aq) (2.0 M) + 3 e- → MnO2 (s) + 2 H2O (l)

Ox: Pb (s) → Pb2+ (aq) (0.10 M) + 2 e-

Calculate the cell potential at 25 °C (8 points).

Red: (MnO4- (aq) + 4 H+ (aq) + 3 e- → MnO2 (s) + 2 H2O (l)) × 2 E° = 1.68 V

Ox: + (Pb (s) → Pb2+ (aq) + 2 e-) × 3 E° = 0.13 V

2 MnO4- (aq) + 8 H+ (aq) + 6 e- + 3 Pb (s) → 2 MnO2 (s) + 4 H2O (l) + 3 Pb2+ (aq) + 6 e-

2 MnO4- (aq) + 8 H+ (aq) + 3 Pb (s) → 2 MnO2 (s) + 4 H2O (l) + 3 Pb2+ (aq) E° = 1.81 V

$$E\_{cell}=E\_{cell}^{°}+\frac{RT}{nF} lnQ$$

$$E\_{cell}=E\_{cell}^{°}+\frac{RT}{nF} ln\left(\frac{\left[Pb^{2+}\right]^{3}}{\left[MnO\_{4}^{-}\right]^{2}\left[H^{+}\right]^{8}}\right)$$

$$E\_{cell}=1.81 V+\frac{\left(8.314\frac{J}{mol K}\right)(298 K)}{\left(6 mol e^{-}\right)\left(96485\frac{C}{mol e^{-}}\right)}×\frac{1 C V}{1 J} ln\left(\frac{\left[0.10\right]^{3}}{\left[1.50\right]^{2}\left[2.0\right]^{8}}\right)$$

$$E\_{cell}=1.81 V+\frac{\left(8.314\frac{J}{mol K}\right)(298 K)}{\left(6 mol e^{-}\right)\left(96485\frac{C}{mol e^{-}}\right)}×\frac{1 C V}{1 J} ln\left(1.736×10^{-6}\right)$$

$$E\_{cell}=1.81 V+0.057 V≈1.87 V$$

1. The following half-reactions that when combined result in a spontaneous reaction conditions (12 points):

 NO3- (aq) + 4 H+ (aq) + 3 e- 🡪 NO (g) + 2 H2O (l) E° = +0.96 V

 Hg22+ (aq) + 2 e- 🡪 2 Hg­ (l) E° = +0.79 V

1. Determine the net ionic equation and the standard cell potential.

Oxidation: (2 Hg­ (l) 🡪 Hg22+ (aq) + 2 e-) x 3 E° = -0.79 V

Reduction: (NO3- (aq) + 4 H+ (aq) + 3 e- 🡪 NO (g) + 2 H2O (l) ) x 2 E° = +0.96 V

2 NO3­- (aq) + 8 H+ (aq) + 6 Hg (l) 🡪 3 Hg22+ (aq) + 2 NO (g) + 4 H2O (l) E° = 0.17 V

1. Write the cell notation for the reaction.

Pt (s) | Hg (l) | Hg22+ (aq) || NO3- (aq), H+ (aq) | NO (g) | Pt­ (s)

1. Calculate the ΔG°.

$$∆G^{°}=-nFE^{°}=-\left(6 mol e^{-}\right)\left(96485\frac{C}{mol e^{-}}\right)\left(0.17 V\right)\left(\frac{1 J}{1 C V}\right)\left(\frac{1 kJ}{1000 J}\right)=-98 kJ$$