

last name

first name

signature

1 1 H 1.008																	18 2 He 4.003
3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.64	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57 La 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.20	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	89 Ac (227)	104 Rf (267)	105 Db (268)	106 Sg (269)	107 Bh (270)	108 Hs (270)	109 Mt (278)	110 Ds (281)	111 Rg (282)	112 Cn (285)	113 Nh (286)	114 Fl (289)	115 Mc (290)	116 Lv (293)	117 Ts (294)	118 Og (294)

58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (266)

constants

$R = 0.08206 \text{ L atm/mol K}$

$R = 8.314 \text{ J/mol K}$

$F = 96485 \text{ C/mol e}^-$

$N_A = 6.022 \times 10^{23} / \text{mol}$

$h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$

$c = 3.00 \times 10^8 \text{ m/s}$

$g = 9.81 \text{ m/s}^2$

conversions

$1 \text{ atm} = 760 \text{ torr}$

$1 \text{ atm} = 101325 \text{ Pa}$

$1 \text{ atm} = 1.01325 \text{ bar}$

$1 \text{ bar} = 10^5 \text{ Pa}$

$^\circ\text{F} = ^\circ\text{C}(1.8) + 32$

$\text{K} = ^\circ\text{C} + 273.15$

conversions

$1 \text{ in} = 2.54 \text{ cm}$

$1 \text{ ft} = 12 \text{ in}$

$1 \text{ yd} = 3 \text{ ft}$

$1 \text{ mi} = 5280 \text{ ft}$

$1 \text{ lb} = 453.6 \text{ g}$

$1 \text{ ton} = 2000 \text{ lbs}$

$1 \text{ tonne} = 1000 \text{ kg}$

$1 \text{ gal} = 3.785 \text{ L}$

$1 \text{ gal} = 231 \text{ in}^3$

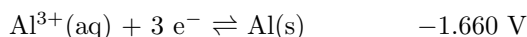
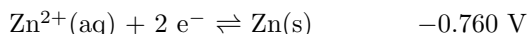
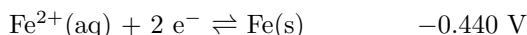
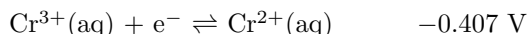
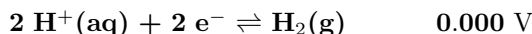
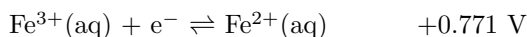
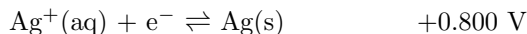
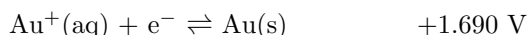
$1 \text{ gal} = 128 \text{ fl oz}$

$1 \text{ fl oz} = 29.57 \text{ mL}$

$1 \text{ J} = 1 \text{ V C}$

$1 \text{ W} = 1 \text{ J/s}$

$1 \text{ A} = 1 \text{ C/s}$

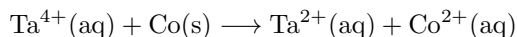
standard potentials at 25 °C  $E^\circ$  (V)

This exam should have exactly 25 questions. Each question is equally weighted at 4 points each. You will enter your answer choices on the virtual bubbleseet after you have finished. Your score is based on what you submit on the virtual bubblesheet and not what is circled on the exam.

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1. Consider the following redox reaction:



What is the reducing agent?

- a. Co
- b.  $\text{Co}^{2+}$
- c.  $\text{Ta}^{2+}$
- d.  $\text{Ta}^{4+}$

**Explanation:** The reducing agent is the species being oxidized, which is  $\text{Co}(\text{s})$  being oxidized to  $\text{Co}^{2+}$ .

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2. Which of the following metals is most easily oxidized?

- a. Cu
- b. Al
- c. Au
- d. Ni
- e. Ag

**Explanation:** The metal that is most easily oxidized is the one with the greatest oxidation potential. The table provided shows reduction potentials for ions associated with these metals. The oxidation potential of the metal is the opposite of the standard reduction potential of the associated ions.

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3. Consider the cell:



What is serving as the anode in this electrochemical cell?

- a.  $\text{H}_2(\text{g})$
- b.  $\text{Cl}_2(\text{g})$
- c.  $\text{C}(\text{s})$
- d.  $\text{Cl}^-(\text{aq})$
- e.  $\text{Pt}(\text{s})$
- f.  $\text{H}^+(\text{aq})$

**Explanation:** When using standard cell notation, the solid to the furthest left side of the notation is solid substance serving as the anode.

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4. You hook up a voltmeter to a standard electrochemical cell and it reads +0.313 V. With this information, which of the following is true regarding this cell?

- a.  $\Delta G$  is positive and  $K$  is greater than 1.
- b.  $\Delta G$  is negative and  $K$  is less than 1.
- c.  $\Delta G$  is zero and  $K$  is equal to 1.
- d.  $\Delta G$  is negative and  $K$  is greater than 1.
- e.  $\Delta G$  is positive and  $K$  is less than 1.

**Explanation:** Recall the triformula,  $-nF\mathcal{E}^\circ = \Delta G^\circ = -RT \ln K$ . If we measure a positive potential, this must correspond to a negative  $\Delta G^\circ$  (a spontaneous process). This must also correspond to an equilibrium constant which is greater than 1.

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5. Identify the oxidation number for each element in  $\text{MnSO}_4$ . List your answer in the order Mn, S, and O.

- a. +4, +3, -1
- b. +2, +8, -2
- c. +2, +5, -2
- d. +2, +6, -2
- e. +2, +4, -6

**Explanation:** Oxygen has an oxidation number of -2. The polyatomic sulfate anion has an overall charge of -2, so sulfur's oxidation number be -6 in order to compensate for the four oxygen ions. The manganese is given as manganese(II) ion, so its oxidation number in considered to be +2.

6. What is the cell potential for the following nonstandard cell?



- a. 0.771 V
- b. 1.160 V
- c. 0.901 V
- d. 0.622 V
- e. 0.332 V
- f. 0.460 V

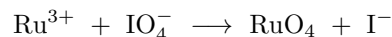
**Explanation:**  $E^\circ = +0.800 - -0.230 = +1.03 \text{ V}$ .  
 $Q = \frac{1}{0.0065^2}$  and  $n = 2$   
 $E = 1.03 - 0.05916/2 \log\left(\frac{1}{0.0065^2}\right)$   
 $E = 0.901 \text{ V}$

7. I put a piece of copper wire into a solution of 1 M  $\text{Ag}^+$ . What happens?

- a. The copper ions in solution are reduced and plate onto the copper wire
- b. The two reactants will fuse together to make Zn metal.
- c. Nothing will happen.
- d. The copper wire reduces the  $\text{Ag}^+$  ions to silver metal.

**Explanation:** The standard potential for  $\text{Cu(s)}$  and  $\text{Ag}^+(\text{aq})$  is 0.46 V. This is spontaneous, so the copper wire will be oxidized and the silver ions will be reduced.

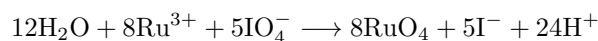
8. Balance the following redox reaction in acidic solution:



What is the coefficient of  $\text{Ru}^{3+}$  in the simplest, overall reaction?

- a. 8
- b. 5
- c. 4
- d. 1
- e. 6
- f. 2

**Explanation:**



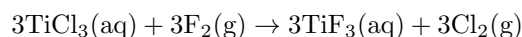
Coefficient on  $\text{Ru}^{3+}$  is 8.

9. A proposed solution for grid-scale energy storage is the use of so-called 'flow batteries'. They are essentially just warehouse sized electrochemical cells! They can be charged by running the cell electrolytically, and can later be discharged by running the cell voltaically. A common cell design uses the  $\text{Cr}^{2+}/\text{Cr}^{3+}$  and  $\text{Fe}^{2+}/\text{Fe}^{3+}$  redox couples. What is the standard cell potential for a voltaic cell constructed using these redox couples?

- a. 0.033 V
- b. +1.18 V
- c. -0.36 V
- d. 1.68 V
- e. -1.18 V
- f. +0.36 V

**Explanation:** The standard reduction potentials are -0.407 V for  $\text{Cr}^{2+}/\text{Cr}^{3+}$  and 0.771 for  $\text{Fe}^{2+}/\text{Fe}^{3+}$ . We are running a voltaic cell, so we assign the more negative potential to be the anode ( $\text{Cr}^{2+}/\text{Cr}^{3+}$ ) and the more positive potential to be the cathode ( $\text{Fe}^{2+}/\text{Fe}^{3+}$ ), giving us a potential of +1.18 V.

10. For the following reaction of titanium(III) chloride with fluorine, what is the oxidizing agent?



- a.  $\text{F}_2(\text{g})$
- b.  $\text{Cl}_2(\text{g})$
- c.  $\text{Ti}(\text{aq})$
- d.  $\text{TiCl}_3(\text{aq})$
- e.  $\text{TiF}_3(\text{aq})$

**Explanation:** Fluorine is being reduced from an oxidation state of 0 to a state of -1, so it is causing the chloride ion to be oxidized.

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11. Which of the following is true of a standard hydrogen electrode?

- I. It cannot be part of a voltaic cell since its standard reduction potential is zero.
  - II. It contains a solution with a concentration of  $[\text{H}^+] = 1 \text{ M}$
  - III. It uses  $\text{H}_2$  gas at a pressure of 1 atm.
- a. only I
  - b. I, II and III
  - c. only III
  - d. only II
  - e. II and III
  - f. I and III
  - g. I and II

**Explanation:** SHE is used as a reference for all other half-reactions, and is zero by definition. That does not mean that it cannot be used in a cell. By definition, the solution must have 1 M  $\text{H}^+$  and the  $\text{H}_2$  pressure must be 1 atm. Thus, the correct answers are only II and III.

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12. Salts of chromate ( $\text{CrO}_4^{2-}$ ) are incredibly toxic and carcinogenic because they mimic phosphate ( $\text{PO}_4^{3-}$ ), which forms backbone of DNA. What is the oxidation number of Cr in chromate and P in phosphate, respectively?

- a. +6 and +6
- b. +5 and +6
- c. +6 and +5
- d. +8 and +8
- e. +5 and +5

**Explanation:** For  $\text{CrO}_4^{2-}$ , each oxygen is 2-, but the overall complex has a 2- charge, making the Cr 6+. For  $\text{PO}_4^{3-}$ , each oxygen is again -2, but this time the overall molecule has a 3- charge, making the P 5+.

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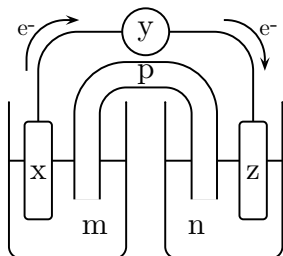
13. In an electrolytic cell, the positive terminal is the \_\_\_\_\_ and is the site of the \_\_\_\_\_ half-reaction.

- a. cathode; reduction
- b. anode; oxidation
- c. anode; reduction
- d. cathode; oxidation

**Explanation:** In an electrolytic cell (analogous to charging a battery), the cathode is always attributed a negative sign and the anode a positive sign. By definition, the cathode is the site of reduction, and the anode is the site of oxidation.

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14. (part 1 of 2): Consider the electrochemical cell diagram shown below which is running with the electron flow direction as shown. What must be true for this electrochemical cell to run a complete closed-loop (circular) circuit without creating an imbalance of charge?



- anions must flow from  $y$  toward  $z$ , cations flow from  $y$  toward  $x$
- anions must flow from  $p$  toward  $x$ , cations flow from  $y$  toward  $z$
- anions must flow from  $y$  toward  $x$ , cations flow from  $y$  toward  $z$
- anions must flow from  $p$  toward  $x$ , cations flow from  $p$  toward  $z$
- anions must flow from  $p$  toward  $z$ , cations flow from  $p$  toward  $x$

**Explanation:**  $p$  = salt bridge; cations will flow towards the cathode which has to be the  $z$  component due to the fact electrons are flowing towards it in the external circuit. Anions will flow to  $x$ , which is the anode compartment.

15. (part 2 of 2): Suppose the previously shown cell diagram represents an electrolytic cell. What is the role and assigned charge of the component labeled  $z$ ?

- cathode, positive
- cathode, negative
- anode, positive
- anode, negative
- cathode, no charge

**Explanation:** Electrons always travel from anode to cathode. For the diagram given, the electrons travel from  $x$  (anode) to  $z$  (cathode). For an electrolytic cell, the anode is positive ( $x$ ) and the cathode is negative ( $z$ ).

16. How long will it take to deposit 0.0231 moles of solid promethium (Pm, atomic number 61) by the electrolysis of  $\text{Pm}(\text{NO}_3)_3(\text{aq})$  using a current of 1.31 A?

- 85 minutes
- 75 minutes
- 60 minutes
- 65 minutes
- 70 minutes
- 80 minutes

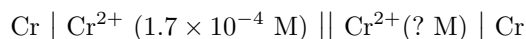
**Explanation:** First note that the oxidation number of promethium is +3 in the aqueous solution given, so when it gets reduced to solid promethium, 3 moles of electrons will be transferred. Now use  $t = \frac{\text{mol} \cdot n \cdot F}{I} = \frac{0.0231(3)(96485)}{1.31} = 5100 \text{ s} = 85 \text{ minutes}$

17. Which of the following species is the strongest reducing agent?

- $\text{Cu}(\text{s})$
- $\text{Ag}(\text{s})$
- $\text{F}_2(\text{g})$
- $\text{Zn}(\text{s})$
- $\text{Mg}^{2+}(\text{aq})$

**Explanation:** A good reducing agent needs to push electrons away which means they will be on the right side of a standard potential table and will have the most negative  $E^\circ$ . That means that Zn is the best of those listed.

18. The following concentration cell has a measured potential of +84.8 mV. What is the unknown chromium(II) ion concentration in this cell?



- a. 0.075 M
- b.  $3.2 \times 10^{-8}$  M
- c. 0.0023 M
- d. 0.14 M
- e.  $2.0 \times 10^{-6}$  M
- f. 0.125 M

**Explanation:** Use Nernst equation:

$$E = E^\circ - \frac{0.05916}{2} \log Q$$

$$0.0848 = 0 - (0.02958) \log\left(\frac{1.7 \times 10^{-4}}{x}\right)$$

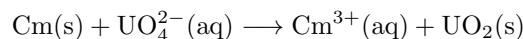
solve for  $x = 0.125 \text{ M} = [\text{Cr}^{2+}]$  in the cathode compartment

19. What is the reducing agent in an alkaline cell?

- a.  $\text{H}_2\text{O} (\ell)$
- b.  $\text{ZnO} (\text{s})$
- c.  $\text{MnO}_2 (\text{s})$
- d.  $\text{Zn} (\text{s})$
- e.  $\text{Mn}_2\text{O}_3 (\text{s})$

**Explanation:** The overall reaction for an alkaline cell is  $2\text{MnO}_2 + \text{Zn} \rightarrow \text{Mn}_2\text{O}_3 + \text{ZnO}$ . Zn changes its oxidation state from 0 to +2 and is thus oxidized and the reducing agent is the reactant that itself undergoes oxidation. It's the reactant that "gives away" electrons.

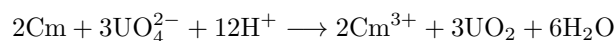
20. Consider the following redox reaction involving curium (An element whose namesake honors Marie Curie, a pioneering chemist and physicist who studied radioactivity) and the unfortunately named Uranate ion,  $\text{UO}_4^{2-}$ .



Using the smallest possible integer coefficients, balance the reaction in acidic aqueous solution. What is the coefficient for  $\text{H}^+$ , and which side of the reaction is it on?

- a. 10, right
- b. 2, left
- c. 4, left
- d. 12, left
- e. 10, left
- f. 12, right

**Explanation:**



Coefficient on  $\text{H}^+$  is 12, and is on the left.

21. You go to the hardware store to buy some root killer (copper(II) sulfate, 159.6 g/mol) and to your dismay it is all sold out. Using your electrochemistry knowledge, you decide to make some yourself. You instead buy copper piping and sulfuric acid to construct an electrolytic cell. Copper(II) sulfate is deposited on the copper electrode from a solution of sulfuric acid. What mass of copper(II) sulfate is made when 27.0 amps of current is run continuously for 17.0 hours?

- a. 672 g
- b. 308 g
- c. 22.8 g
- d. 1.37 kg
- e. 1.76 kg
- f. 1.13 kg

**Explanation:** The oxidation number of copper sulfate is +2, so you need to lose 2 mols of electrons per mol of copper to make Cu(II). Now use  $\text{mol} = \frac{I \cdot t}{n \cdot F} = \frac{(27) \cdot (17 \cdot 60 \cdot 60)}{2 \cdot 96458} = 8.56$  mols of  $\text{CuSO}_4$ . Use MW to convert into g and then into kg to find 1.37 kg.

22. What is the standard free energy change for the following standard cell?



- a. +45.3 kJ
- b. -45.3 kJ
- c. -90.7 kJ
- d. +40.5 kJ
- e. -40.5 kJ
- f. +90.7 kJ

**Explanation:** We can relate the cell potential to the standard free energy change using one of the triforme relationships:  $-nF\mathcal{E}^\circ = \Delta G^\circ$ .

Recognize that Pb is the anode and Cu is the cathode:

$$\mathcal{E}^\circ = \mathcal{E}_{\text{cathode}}^\circ - \mathcal{E}_{\text{anode}}^\circ = 0.340 \text{ V} - (-0.130 \text{ V}) = 0.470 \text{ V}$$

Rearrange the above relationship and solve for  $\Delta G^\circ = -90.7 \text{ kJ}$

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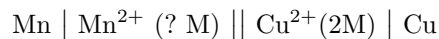
23. Car batteries from the 1970's were able to produce about 300 cold cranking amps. Modern car batteries are able to produce well over 600 cold cranking amps. This improvement in performance is mostly due to:

- a. The design of improved electrodes with higher surface area.
- b. Batteries from the 1970's used lead (II) sulfate, modern batteries now use lead (II) nitrate.
- c. Lead from the 1970's was just lazy.
- d. The use of higher purity lead in batteries.
- e. Batteries from the 1970's used  $\text{PbO}_2$  as the oxidant, modern batteries now use  $\text{TiO}_2$ .

**Explanation:** Car batteries from the 1970's and modern car batteries operate off of the exact same electrochemical reaction. The improvement in current output is largely due to the design improved electrodes with increased surface area.

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24. We construct the following cell below and measure the potential to be 1.610 V. What is the concentration of  $\text{Mn}^{2+}$  ions?



- a. 0.110 M
- b. 0.0073 M
- c. 0.0018 M
- d. 0.256 M
- e. 0.767 M

**Explanation:** Use Nernst equation:

$$E = E^\circ - \frac{0.05916}{2} \log Q$$

$$E^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ = 0.340\text{V} - (-1.180\text{V}) = 1.520\text{V}$$

$$1.610 = 1.520 - \frac{0.05916}{2} \log\left(\frac{x}{2}\right)$$

Solve for  $x = 0.0020 \text{ M}$

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25. Which type of battery chemistry listed below results in a battery that is not rechargeable?

- a. Pb-acid
- b. NiCd
- c. NiMH

explin: NiCd, Pb-acid, Li-ion, and NiMH are all examples of secondary cells (rechargeable) batteries. The only one that is a primary (non-rechargeable) battery is an alkaline cell.

- d. Li-ion
- e. alkaline

**Explanation:**

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After you are finished and have all your answers circled, go to the front of the room and then use the QR code show below to pull up the virtual answer page for your exam. Enter the appropriate info plus all your answers - click the SUBMIT button. Double check your choices on the next page. Once your are sure, click the submit button on that page to enter your answers. Make sure you get the confirmation screen (different background color!) and show it to the TA or proctor. After that, turn in your exam and scratch paper. You're free to leave after that.



<https://mccord.cm.utexas.edu/iron>

\* \* \* Remember your Version Number \* \* \*