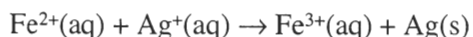


Exam3C152S07

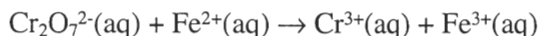
Be sure the person sitting on either side of you has a different colored exam. You may write on the exam, but be sure to put your name and exam color on your Green Scantron. Identify the letter of the choice that best completes the statement or answers the question.

1. In the following reaction,



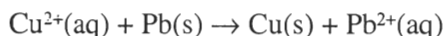
- a. Ag<sup>+</sup> is oxidized and Fe<sup>3+</sup> is reduced.
- b. Fe<sup>2+</sup> is oxidized and Fe<sup>3+</sup> is reduced.
- c. Fe<sup>2+</sup> is oxidized and Ag<sup>+</sup> is reduced.
- d. Ag<sup>+</sup> is oxidized and Fe<sup>2+</sup> is reduced.
- e. Ag<sup>+</sup> is oxidized and Ag(s) is reduced.

2. Write a balanced chemical equation for the following reaction in an acidic solution.



- a.  $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 6 \text{Fe}^{2+}(\text{aq}) + 7 \text{H}^+(\text{aq}) \rightarrow 2 \text{Cr}^{3+}(\text{aq}) + 6 \text{Fe}^{3+}(\text{aq}) + 7 \text{OH}^-(\text{aq})$
- b.  $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + \text{Fe}^{2+}(\text{aq}) + 14 \text{H}^+(\text{aq}) \rightarrow 2 \text{Cr}^{3+}(\text{aq}) + \text{Fe}^{3+}(\text{aq}) + 7 \text{H}_2\text{O}(\ell)$
- c.  $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + \text{Fe}^{2+}(\text{aq}) \rightarrow 2 \text{Cr}^{3+}(\text{aq}) + \text{Fe}^{3+}(\text{aq})$
- d.  $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + \text{Fe}^{2+}(\text{aq}) + 7 \text{H}^+(\text{aq}) \rightarrow 2 \text{Cr}^{3+}(\text{aq}) + \text{Fe}^{3+}(\text{aq}) + 7 \text{OH}^-(\text{aq})$
- e.  $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 6 \text{Fe}^{2+}(\text{aq}) + 14 \text{H}^+(\text{aq}) \rightarrow 6 \text{Fe}^{3+}(\text{aq}) + 2 \text{Cr}^{3+}(\text{aq}) + 7 \text{H}_2\text{O}(\ell)$

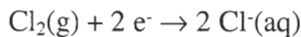
3. What is the correct cell notation for the reaction below?



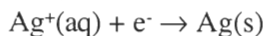
- a.  $\text{Cu} | \text{Pb}^{2+}(\text{aq}) || \text{Cu}^{2+}(\text{aq}) | \text{Pb}$
- b.  $\text{Pb} | \text{Cu}^{2+}(\text{aq}) || \text{Pb}^{2+}(\text{aq}) | \text{Cu}$
- c.  $\text{Pb} | \text{Cu}(\text{s}) || \text{Pb}^{2+}(\text{aq}) | \text{Cu}^{2+}$
- d.  $\text{Pb} | \text{Pb}^{2+}(\text{aq}) || \text{Cu}^{2+}(\text{aq}) | \text{Cu}$
- e.  $\text{Cu} | \text{Cu}^{2+}(\text{aq}) || \text{Pb}^{2+}(\text{aq}) | \text{Pb}$

OX → red or Anode → Cathode  
 $\text{Pb}(\text{s}) | \text{Pb}^{2+}(\text{aq}) || \text{Cu}^{2+}(\text{aq}) | \text{Cu}(\text{s})$

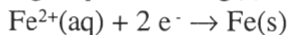
4. Use the standard reduction potentials below to determine which compound or ion is the best oxidizing agent?



$$E^\circ = +1.36 \text{ V} \leftarrow \text{most } (+)$$



$$E^\circ = +0.80 \text{ V}$$

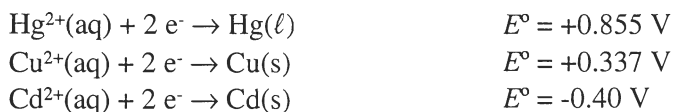


$$E^\circ = -0.44 \text{ V}$$

↑  
Gains e<sup>-</sup>

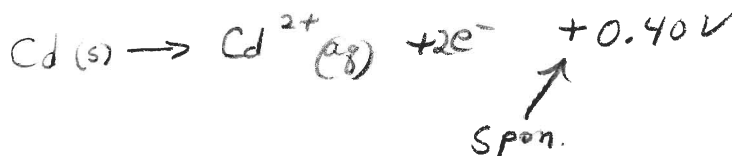
- a. Fe
- b. Fe<sup>2+</sup>
- c. Ag
- d. Cl<sub>2</sub>
- e. Cl<sup>-</sup>

5. Use the standard reduction potentials below to determine which compound or ion is the best reducing agent?

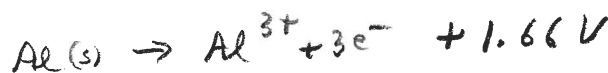
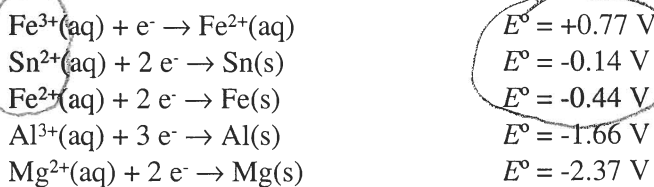


Loses  $e^{-}$ s

- a.  $\text{Cu}^{2+}$
- b.  $\text{Cd}^{2+}$
- c.  $\text{Hg}^{2+}$
- d. Cd**
- e.  $\text{Hg}(\ell)$



6. Consider the following half-reactions:

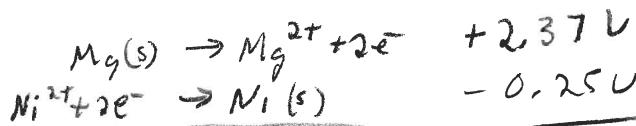
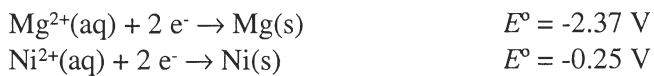


Which of the above metals or metal ions are able to oxidize Al(s)?

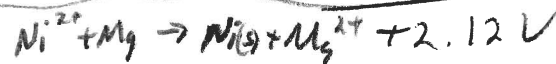
- a.  $\text{Fe}^{3+}$  and  $\text{Sn}^{2+}$
- b. Mg and  $\text{Mg}^{2+}$
- c.  $\text{Mg}^{2+}$  only
- d.  $\text{Fe}^{2+}$ , Sn, and Fe
- e.  $\text{Fe}^{3+}$ ,  $\text{Sn}^{2+}$ , and  $\text{Fe}^{2+}$**

↳ GAIN  $e^{-}$ s from Al

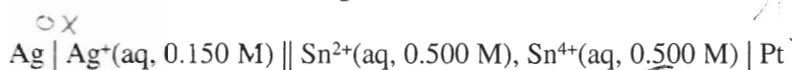
7. Given the following two half-reactions, determine which overall reaction is spontaneous and calculate the cell potential.



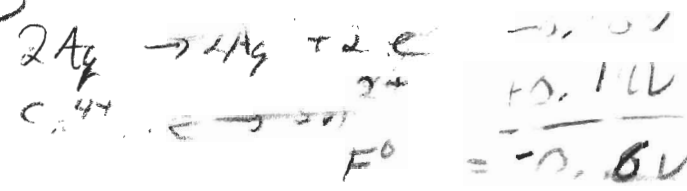
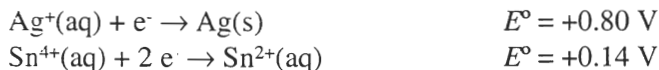
- a.  $\text{Mg}^{2+}(\text{aq}) + \text{Ni}(\text{s}) \rightarrow \text{Mg}(\text{s}) + \text{Ni}^{2+}(\text{aq}) \quad E^{\circ}_{\text{cell}} = +2.12 \text{ V}$
- b.  $\text{Mg}(\text{s}) + \text{Ni}^{2+}(\text{aq}) \rightarrow \text{Ni}(\text{s}) + \text{Mg}^{2+}(\text{aq}) \quad E^{\circ}_{\text{cell}} = +2.12 \text{ V}$**
- c.  $\text{Mg}^{2+}(\text{aq}) + \text{Ni}^{2+}(\text{aq}) \rightarrow \text{Ni}(\text{s}) + \text{Mg}(\text{s}) \quad E^{\circ}_{\text{cell}} = -2.12 \text{ V}$
- d.  $\text{Mg}^{2+}(\text{aq}) + \text{Ni}(\text{s}) \rightarrow \text{Mg}(\text{s}) + \text{Ni}^{2+}(\text{aq}) \quad E^{\circ}_{\text{cell}} = -2.62 \text{ V}$
- e.  $\text{Mg}^{2+}(\text{aq}) + \text{Ni}^{2+}(\text{aq}) \rightarrow \text{Mg}(\text{s}) + \text{Ni}(\text{s}) \quad E^{\circ}_{\text{cell}} = +2.62 \text{ V}$



8. Calculate  $E$  for the following electrochemical cell at 25°C



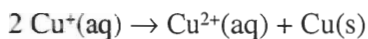
given the following standard reduction potentials.



- a. -0.915 V
- b. -0.61 V**
- c. +0.89 V
- d. +0.99 V
- e. +0.915 V

$E = -0.66 \text{ V} - \frac{0.0591}{2} \log \left( \frac{0.5}{(0.1)^2 (0.5)} \right)$   
 $= -0.61 \text{ V}$

9. Calculate  $\Delta G^\circ$  for the disproportionation of  $\text{Cu}^+$ ,

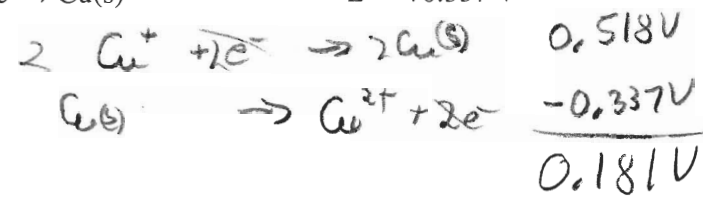


given the following standard reduction potentials.



$\Delta G = -nFE$   
 $\Delta G = -2(9.65 \times 10^4 \text{ C})(0.181 \text{ V})$   
 $\Delta G = -34933 \text{ J}$   
 $J = C \cdot V$

- a. -165 kJ
- b. -175 kJ
- c. -34.9 kJ**
- d. -56.8 kJ
- e. -1180 kJ



10.  $\text{Al}^{3+}$  is reduced to  $\text{Al}(\text{s})$  at an electrode. If a current of 1.00 ampere is passed for 24 hours, what mass of aluminum is deposited at the electrode? Assume 100 % current efficiency.

- a. 1.87 g
- b. 54.1 g
- c. 24.2 g
- d. 72.5 g
- e. 8.05 g**

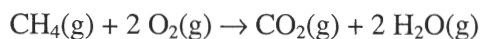
$\text{Al}^{3+} + 3e^- \rightarrow \text{Al}$

24 hrs	3600 sec	1.00 A	mol e <sup>-</sup>	1 mol Al	27.0 g
1 hr	sec	9.65 x 10 <sup>4</sup> C	3 mol e <sup>-</sup>	mol	

11. Thermodynamics can be used to determine all of the following EXCEPT

- a. the direction in which a reaction is spontaneous.
- b. the enthalpy change of a reaction.
- c. the rate of reaction.
- d. the extent to which a reaction occurs.
- e. the temperature at which a reaction is spontaneous.**

12. Calculate the standard molar entropy change for the combustion of methane.

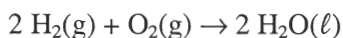


Species	$S^\circ$ (J/K·mol)
CH <sub>4</sub> (g)	186.3
O <sub>2</sub> (g)	205.1
CO <sub>2</sub> (g)	213.7
H <sub>2</sub> O(g)	188.8

$$591.3 - 596.5$$

- a. -1.0 J/K  
 b. +1.0 J/K  
 c. +11.1 J/K  
 d. +5.2 J/K  
 e. -5.2 J/K

13. Predict the signs of  $\Delta H$ ,  $\Delta S$ , and  $\Delta G$  for the combustion of hydrogen gas at 25°C.



- a.  $\Delta H > 0, \Delta S < 0, \Delta G < 0$   
 b.  $\Delta H < 0, \Delta S > 0, \Delta G < 0$   
 c.  $\Delta H > 0, \Delta S < 0, \Delta G > 0$   
 d.  $\Delta H < 0, \Delta S > 0, \Delta G < 0$   
 e.  $\Delta H < 0, \Delta S < 0, \Delta G < 0$

14. Predict the signs of  $\Delta H$ ,  $\Delta S$ , and  $\Delta G$  for the melting of ice at 50°C.

- a.  $\Delta H > 0, \Delta S > 0, \Delta G < 0$   
 b.  $\Delta H > 0, \Delta S < 0, \Delta G < 0$   
 c.  $\Delta H < 0, \Delta S < 0, \Delta G < 0$   
 d.  $\Delta H < 0, \Delta S > 0, \Delta G < 0$   
 e.  $\Delta H < 0, \Delta S > 0, \Delta G < 0$

$$\Delta H +, \Delta S +, \Delta G -$$

15. Above what temperature would you expect a reaction to become spontaneous if  $\Delta H = +322$  kJ and  $\Delta S = +531$  J/K?

- a. The reaction will NOT be spontaneous at any temperature.  
 b. 171 K  
 c. 606 K  
 d. 209 K  
 e. The reaction will be spontaneous at any temperature.

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

$$\Delta G = 0 \text{ @ Eq 50: } 0 = \Delta H - T\Delta S$$

$$T = \frac{\Delta H}{\Delta S} = \frac{+322 \times 10^3 \text{ J}}{531 \text{ J/K}}$$

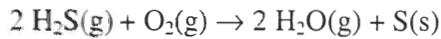
16. At what temperature would you expect a reaction to become spontaneous if  $\Delta H = +67.0$  kJ and  $\Delta S = -131$  J/K?

- a. The reaction will NOT be spontaneous at any temperature.  
 b.  $T > 511$  K  
 c.  $T < -511$  K  
 d. The reaction will be spontaneous at any temperature.  
 e.  $T > 238$  K

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

non spontan. non spontan

17. Calculate  $\Delta G_{\text{rxn}}^{\circ}$  for the reaction below at 25.0°C

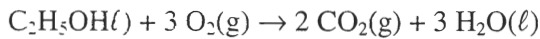


$$\Delta G = \Delta H - T\Delta S$$

given  $\Delta H_{\text{rxn}}^{\circ} = -442.4 \text{ kJ}$ , and  $\Delta S_{\text{rxn}}^{\circ} = -175.4 \text{ J/K}$ .

- a. -438.0 kJ
- b. +3943 kJ
- c. -390.1 kJ
- d. -321.9 kJ
- e. +5182 kJ

18. Calculate  $\Delta G^{\circ}$  for the reaction below at 25.0°C.



Species	$\Delta H_f^{\circ}(\text{kJ/mol})$	$S_f^{\circ}(\text{J/Kmol})$
$\text{C}_2\text{H}_5\text{OH}(\ell)$	-277.7	160.7
$\text{O}_2(\text{g})$	0	205.1
$\text{CO}_2(\text{g})$	-393.5	213.7
$\text{H}_2\text{O}(\ell)$	-285.8	69.1

$$\Delta S = 534.7 - 776. = -141.3 \text{ J/K}$$

$$\Delta H = -1644.4 - -277.7 = -1366.7 \text{ kJ}$$

- a. +2435 kJ
- b. 141.3 kJ
- c. +1038 kJ
- d. 1325 kJ
- e. 365.1 kJ

$$\Delta G = -1366.7 - 298(-0.1413)$$

$$\Delta G = -1325$$

19. We have a solution of benzoic acid,  $\text{C}_6\text{H}_5\text{CO}_2\text{H}$ . What is the effect of adding sodium hydroxide to this solution?

- 1. The pH increases.
- 2. The concentration of  $\text{C}_6\text{H}_5\text{CO}_2\text{H}$  decreases.
- 3. The concentration of  $\text{H}_3\text{O}^+$  increases.

- a. 1 and 2
- b. 1 only
- c. 2 only
- d. 2 and 3
- e. 3 only

20. What is the pH of a mixture containing 0.30 M  $\text{HNO}_2$  and 0.15 M  $\text{NaNO}_2$ ? ( $K_a$  for  $\text{HNO}_2 = 4.5 \times 10^{-4}$ )

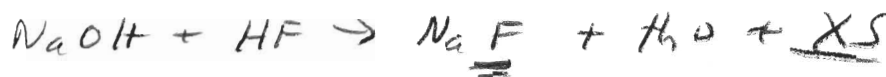
- a. 3.05
- b. 4.35
- c. 4.65
- d. 5.01
- e. 4.05

$$\text{pH} = -\log K_a + \log \frac{0.15}{0.30}$$

21. What is the pH of the solution that results from adding 15 mL of 0.50 M NaOH to 25 mL of 0.50 M HF? ( $K_a$  for HF =  $7.2 \times 10^{-4}$ )

- a. 3.32  
b. 10.86  
c. 3.49  
d. 4.61  
e. 7.53

$$pH = -\log K_a + \log \frac{0.0075}{0.0125 - 0.0075}$$



22. Which of the following combinations would be the best to buffer the pH to 9.0?

- a.  $CH_3CO_2H$  and  $CH_3COO^-$ ,  $K_a = 18 \times 10^{-5}$   
b.  $H_3PO_4$  and  $H_2PO_4^-$ ,  $K_a = 7.5 \times 10^{-3}$   
c.  $HNO_2$  and  $NO_2^-$ ,  $K_a = 4.5 \times 10^{-4}$   
d.  $H_2PO_4^-$  and  $HPO_4^{2-}$ ,  $K_a = 6.2 \times 10^{-8}$

- e.  $NH_4^+$  and  $NH_3$ ,  $K_a = 5.7 \times 10^{-10}$

$$pK_a = 9.2$$

23. Which of the following equations is the solubility product for  $Ca(IO_3)_2$ ?

a.  $K_{sp} = [Ca^{2+}][I^-]^2[3O^{2-}]^2$

b.  $K_{sp} = [Ca^{2+}]^2[IO_3^-]$

c.  $K_{sp} = [Ca^{2+}][IO_3^-]^2$

d.  $K_{sp} = [Ca^{2+}][I^-]^2[O^{2-}]^6$

e.  $K_{sp} = [Ca^{2+}][IO_3^-]$

24. The  $K_{sp}$  for  $BaF_2$  is  $1.7 \times 10^{-6}$ . What is the concentration of  $Ba^{2+}$  in a saturated solution of  $BaF_2$ ?

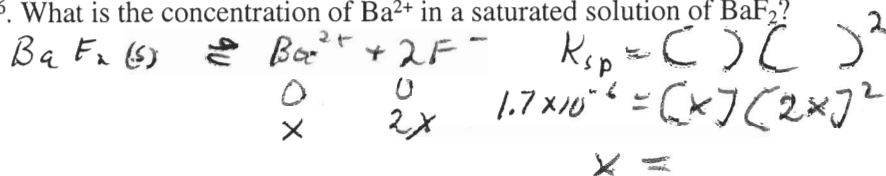
a.  $1.0 \times 10^{-2}$  M

b.  $5.7 \times 10^{-7}$  M

c.  $1.3 \times 10^{-3}$  M

d.  $7.5 \times 10^{-3}$  M

e.  $1.7 \times 10^{-6}$  M



25. At what pH will a solution 0.150 M  $Cu^{2+}$  begin to precipitate as  $Cu(OH)_2$ ? The  $K_{sp}$  for  $Cu(OH)_2$  is  $1.6 \times 10^{-19}$ .

a. 1.80

b. 7.23

c. 5.01

d. 8.99

e. 13.18



$$K_{sp} = [Cu^{2+}][OH^-]^2$$

$$1.6 \times 10^{-19} = [0.150][OH^-]^2$$

$$[OH^-] = 1.03 \times 10^{-9}$$

$$pOH = 8.99$$

$$pH = 5.01$$