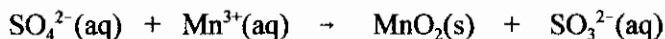


Consider the following *unbalanced* oxidation reduction reaction for the next three questions.



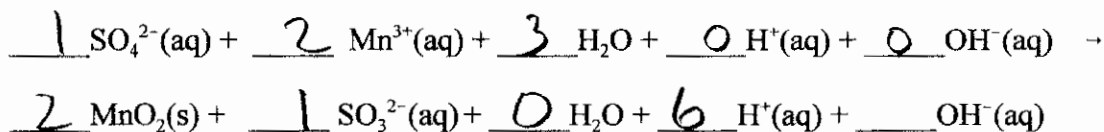
1. The oxidation number of S in  $\text{SO}_3^{2-}$  is \_\_\_\_.

- (a) +6      (b) +4      (c) -2      (d) +8      (e) +3

2. What is being oxidized in this reaction?

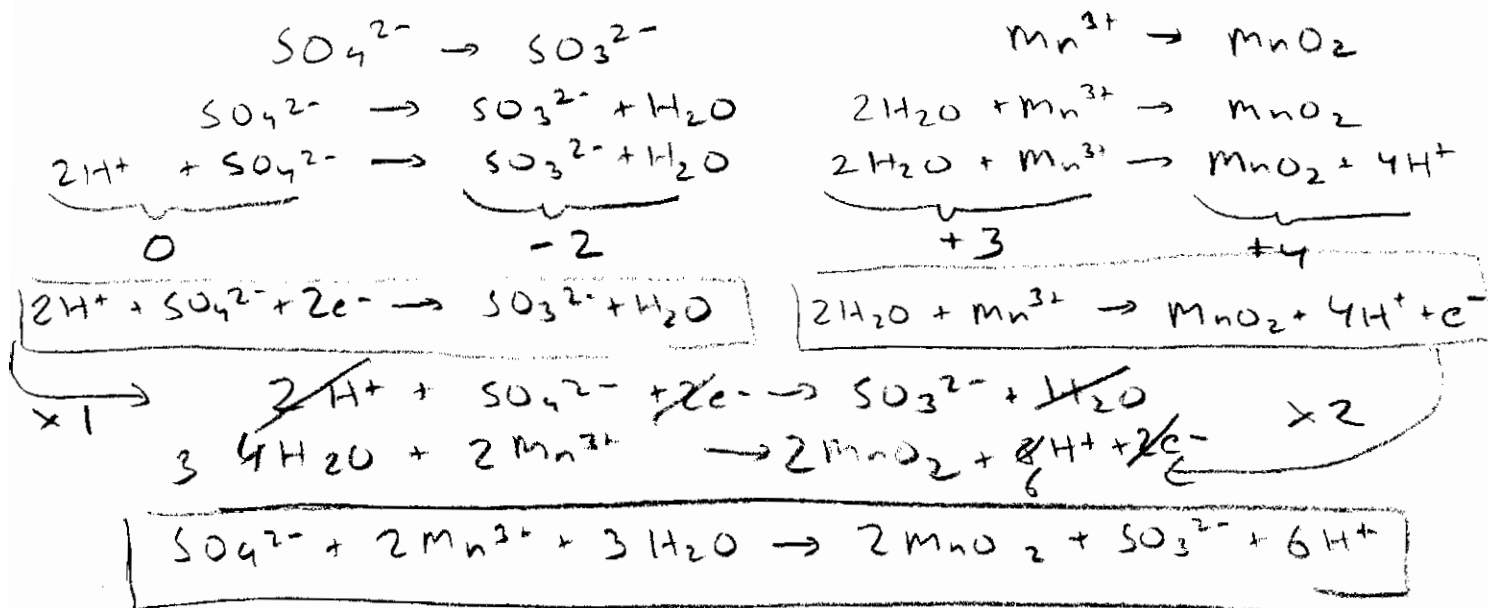
- (a)  $\text{SO}_4^{2-}$       (b)  $\text{MnO}_2$       (c)  $\text{Mn}^{3+}$       (d)  $\text{SO}_3^{2-}$
- $\begin{matrix} +6 \\ \text{SO}_4^{2-} \end{matrix} \rightarrow \begin{matrix} +4 \\ \text{SO}_3^{2-} \end{matrix}$       ox # ↓ reduction  
 $\begin{matrix} +3 \\ \text{Mn}^{3+} \end{matrix} \rightarrow \begin{matrix} +4 \\ \text{MnO}_2 \end{matrix}$       ox # ↑ oxidation

3. The complete balanced equation in **ACIDIC** media for the reaction has the following stoichiometry coefficients. (Note that some of the values may be 0 indicating that the compound does not appear!)

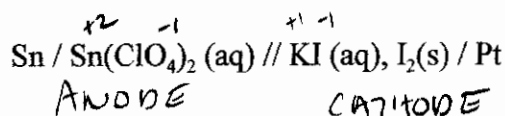


The stoichiometry coefficients, in order are

- (a) 1,1,1,0,0 → 1,1,0,2,0      (b) 1,1,0,2,0 → 1,1,1,0,0  
 (c) 1,2,3,0,0 → 2,1,0,6,0      (d) 2,1,0,6,0 → 1,2,3,0,0  
 (e) 1,2,0,0,6 → 2,1,3,0,0



Consider the following electrochemical cell for the next 3 questions:

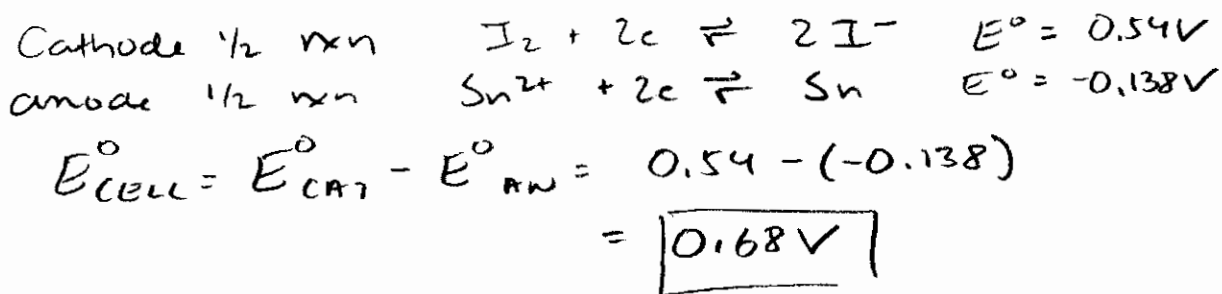


4. What is the name of the aqueous ionic compound in the anode compartment?

- (a) tin chloride                      (b) potassium iodide                      (c) sodium iodine  
(d) sodium chlorate                      (e) tin(II) perchlorate

5. What is  $E^\circ$  cell?

- (a) -0.40 V    (b) -0.68 V    (c) -0.34 V    (d) 0.68 V    (e) 0.34 V    (ab) +0.40 V



6. What is overall net ionic reaction occurring in the cell when current flows?

- (a)  $\text{Sn}(\text{s}) + 2\text{I}^-(\text{aq}) \rightarrow \text{Sn}^{2+}(\text{aq}) + \text{I}_2(\text{s})$   
 (b)  $\text{Sn}^{2+}(\text{aq}) + 2\text{I}^-(\text{aq}) \rightarrow \text{Sn}(\text{s}) + \text{I}_2(\text{s})$   
 (c)  $\text{Sn}(\text{s}) + 2\text{ClO}_4^-(\text{aq}) \rightarrow \text{Sn}(\text{ClO}_4)_2(\text{aq})$   
 (d)  $\text{I}_2(\text{s}) + 2\text{K}^+(\text{aq}) \rightarrow 2\text{KI}(\text{aq})$   
 (e)  $\text{Sn}(\text{s}) + \text{I}_2(\text{s}) \rightarrow \text{Sn}^{2+}(\text{aq}) + 2\text{I}^-(\text{aq})$
- $$\begin{array}{l} \text{I}_2 + 2\text{e}^- \rightleftharpoons 2\text{I}^- \\ \text{Sn} \rightleftharpoons \text{Sn}^{2+} + 2\text{e}^- \\ \hline \text{Sn} + \text{I}_2 \rightleftharpoons \text{Sn}^{2+} + 2\text{I}^- \end{array}$$

oxidation happens @ anode  
reduction @ cathode

7. What does it mean when the oxidation states do NOT change in a reaction?

- (a) The reaction is not an oxidation-reduction reaction. !  
 (b) The reaction is a very favorable oxidation-reduction reaction.  
 (c) The reaction is a very unfavorable oxidation-reduction reaction.  
 (d) The reaction is just an oxidation reaction.  
 (e) The reaction is just a reduction reaction.

8. Before the development of the pH electrode (in the 1950's), the following concentration cell was routinely used to measure pH:



If  $E_{\text{cell}} = -0.363 \text{ V}$  what is the pH of the solution in the compartment on the right?  
*E stand. conditions so  $E = E^\circ$*

- (a) 6.13      (b) 12.26      (c) 3.87      (d) 7.74      (e) 8.45



$$\begin{aligned} E_{\text{cell}} &= E_{\text{cath}} - E_{\text{an}} \\ &= \underbrace{E^\circ_{\text{H}^+/\text{H}_2}}_{E_{\text{cath}}} - \frac{0.0592}{2} \log \frac{P_{\text{H}_2}}{[\text{H}^+]^2} - \underbrace{E^\circ_{\text{H}^+/\text{H}_2}}_{E_{\text{an}}} \end{aligned}$$

$$\begin{aligned} -0.363 &= -\frac{0.0592}{2} \log \frac{1}{[\text{H}^+]^2} \\ &= -\frac{0.0592}{2} \log [\text{H}^+]^{-2} = -(-2) \frac{0.0592}{2} \log [\text{H}^+] \end{aligned}$$

$$\begin{aligned} -0.363 &= +0.0592 \log [\text{H}^+] \\ -\frac{0.363}{0.0592} &= \boxed{-6.13 = \log [\text{H}^+]} \end{aligned}$$

*since pH =  $-\log [\text{H}^+]$   
pH = 6.13*

9. Which one of the following is NOT necessarily TRUE for a chemical reaction at equilibrium?

- (a)  $\Delta G = 0$   
(b)  $\Delta G^\circ = 0$   
(c) The concentrations of reactants and products do not change with time.  
(d) The reaction quotient,  $Q = K$   
(e) The rate of the forward reaction = the rate of the reverse reaction.

Consider the following reaction for the next 6 questions. Hypochlorite,  $\text{ClO}^-$ , the active ingredient in bleach, decomposes to chloride and chlorate by the following reaction:



10. What is  $\Delta G^\circ$  of this reaction?  $\Delta G_f^\circ$  of  $\text{ClO}^-(\text{aq}) = -36.8 \text{ kJ}$ ;  $\Delta G_f^\circ$  of  $\text{Cl}^-(\text{aq}) = -131.3 \text{ kJ}$ ;  $\Delta G_f^\circ$  of  $\text{ClO}_3^-(\text{aq}) = -1.92 \text{ kJ}$ . ( $\Delta G_f^\circ$  values are at  $25^\circ\text{C}$ )

- (a) 96.4 kJ (b) 262.6 kJ (c) 154.1 kJ (d) -154.1 kJ (e) -262.6 kJ (ab) -96.4 kJ

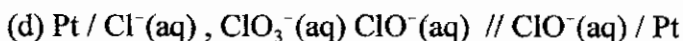
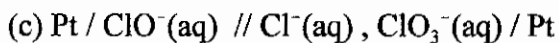
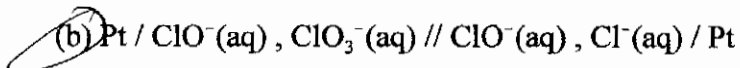
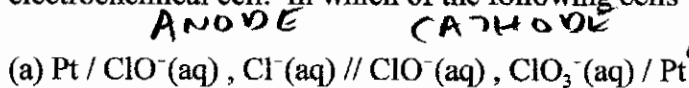
$$\begin{aligned} \Delta G^\circ &= 2\Delta G_f^\circ \text{Cl}^- + \Delta G_f^\circ \text{ClO}_3^- - 3\Delta G_f^\circ \text{ClO}^- \\ &= 2(-131.3 \text{ kJ}) + (-1.92 \text{ kJ}) - 3(-36.8 \text{ kJ}) \\ &= \boxed{-154.1 \text{ kJ}} \end{aligned}$$

11. What is K for this reaction at  $25^\circ\text{C}$ ?

- (a)  $1.8 \times 10^{-11}$  (b)  $1.0 \times 10^{-27}$  (c)  $5.6 \times 10^{10}$  (d)  $1.0 \times 10^{27}$   
(e)  $6.5 \times 10^{-7}$

$$\begin{aligned} K &= e^{-\Delta G^\circ / RT} = e^{\frac{-154.1 \text{ kJ}}{(8.314 \times 10^{-3} \text{ kJ/K} \cdot 298.15 \text{ K})}} \\ &= \boxed{1.0 \times 10^{27}} \end{aligned}$$

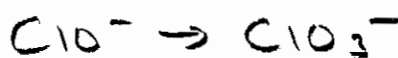
12. The above reaction is an oxidation/reduction reaction so it could take place in an electrochemical cell. In which of the following cells would this reaction occur?



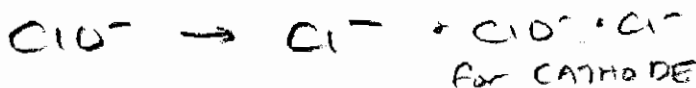
Use a Pt electrode because none of the components are metallic

Note that all components of the half-rxn need to be present to set the potential

oxidation half rxn (in anode)



reduction half rxn (in cathode)



so need  $\text{ClO}^-$  &  $\text{ClO}_3^-$  for ANODE  
for CATHODE

Continuation of the questions on the following reaction:



13. What does the  $\Delta G^\circ$  and K values say about the rate of this reaction?
- (a) This will be a very fast reaction.  
(b) This will be a very slow reaction.  
(c) The  $\Delta G^\circ$  and K values tell you nothing definite about the rate of the reaction.
14. Given the stoichiometry of the reaction, if, at a particular temperature, the rate of change of  $\text{ClO}^-$  was  $-9.0 \times 10^{-4} \text{ M/hr}$ , what would be the rate of change of  $\text{Cl}^-$  at the same time?

- (a)  $-3.0 \times 10^{-4} \text{ M/hr}$       (b)  $+9.0 \times 10^{-4} \text{ M/hr}$       (c)  $+3.0 \times 10^{-4} \text{ M/hr}$   
(d)  $+6.0 \times 10^{-4} \text{ M/hr}$       (e)  $+1.2 \times 10^{-3} \text{ M/hr}$

takes 3  $\text{ClO}^-$  to make 2  $\text{ClO}_3^-$   
so  $\text{ClO}_3^-$  must be increasing  $2/3$  as fast  
as  $\text{ClO}^-$  is decreasing

$$-2/3 (-9.0 \times 10^{-4} \text{ M/hr}) = 6.0 \times 10^{-4} \text{ M/hr}$$

15. It is found experimentally that the rate of the reaction increases 4 times when the concentration of  $\text{ClO}^-$  is doubled. Assuming the products are not in the rate law, what is the rate law of this reaction?

- (a) rate =  $k[\text{ClO}^-]^3$       (b) rate =  $k[\text{ClO}^-]^4$       (c) rate =  $k[\text{ClO}^-]$   
(d) rate =  $k[\text{ClO}^-]^{-1/2}$       (e) rate =  $k[\text{ClO}^-]^2$

$2^x = 4 \Rightarrow x = 2$

16. Consider the following reaction:



The rate law for this reaction has been determined by experiment to be

$$\text{rate} = k[\text{Cr}^{3+}]^2[\text{Ce}^{4+}]^1[\text{CrO}_4^{2-}]^{-1}$$

What is the overall kinetic order of the reaction? *Add up exponents*

- (a) 1<sup>st</sup>      (b) 2<sup>nd</sup>      (c) 4<sup>th</sup>      (d) 8<sup>th</sup>      (e) 3<sup>rd</sup>       $2 + 1 + (-1) = 2$

17. Given the rate law for the reaction, which one of the following statements best describes how the rate of the reaction will change with time?

- (a) The rate of the reaction will decrease with time in the normal fashion as the reactant concentration decreases.  
 (b) The rate of the reaction will increase with time in the normal fashion as the reactant concentration decreases.  
 (c) The rate of the reaction will decrease with time faster than normal because build up of the product  $\text{CrO}_4^{2-}$  slows the reaction further.  
 (d) The rate of the reaction will decrease with time slower than normal because the build up of the product  $\text{CrO}_4^{2-}$  increases the rate of the reaction.

*rate  $\propto \frac{1}{[\text{CrO}_4^{2-}]}$  so as  $\text{CrO}_4^{2-}$  is made rate  $\downarrow$*

18. The following rates were measured for the reaction  $\text{A} + \text{B} \rightarrow \text{C}$ :

Experiment	Initial [A], M	Initial [B], M	Initial rate, M/s
1	0.10	0.10	$1.50 \times 10^{-4}$
2	0.10	0.30	$4.05 \times 10^{-3}$
3	0.20	0.10	$3.00 \times 10^{-4}$

Based on the above information, what is the rate law for this reaction?

- (a) rate =  $k[\text{A}][\text{B}]^3$       (b) rate =  $k[\text{A}]^2[\text{B}]^2$       (c) rate =  $k[\text{B}]^2$   
 (d) rate =  $k[\text{A}]^2$       (e) rate =  $k[\text{A}][\text{B}]$       (ab) rate =  $k[\text{A}]^3$

*rate =  $k[\text{A}]^x[\text{B}]^y$*   
 to determine  $x$ , compare runs 1 & 3  
 $\frac{[\text{B}]_3}{[\text{B}]_1} = \left(\frac{0.30}{0.10}\right)^y = 3^y = \frac{\text{rate}_3}{\text{rate}_1} = \frac{3.00 \times 10^{-4}}{1.50 \times 10^{-4}} = 2$        $2^y = 2$  so  $y = 1$   
 to determine  $y$ , compare runs 2 & 1  
 $\frac{[\text{A}]_2}{[\text{A}]_1} = \left(\frac{0.20}{0.10}\right)^x = 2^x = \frac{\text{rate}_2}{\text{rate}_1} = \frac{4.05 \times 10^{-3}}{1.50 \times 10^{-4}} = 27$        $2^x = 27$  so  $x = 3$

19. The decay of radioactive isotopes follow first order kinetics. If the rate constant for the radioactive decay of  $^{131}\text{I} = 0.0862 \text{ day}^{-1}$ , what percentage of  $^{131}\text{I}$  in a sample will remain after 1 week?

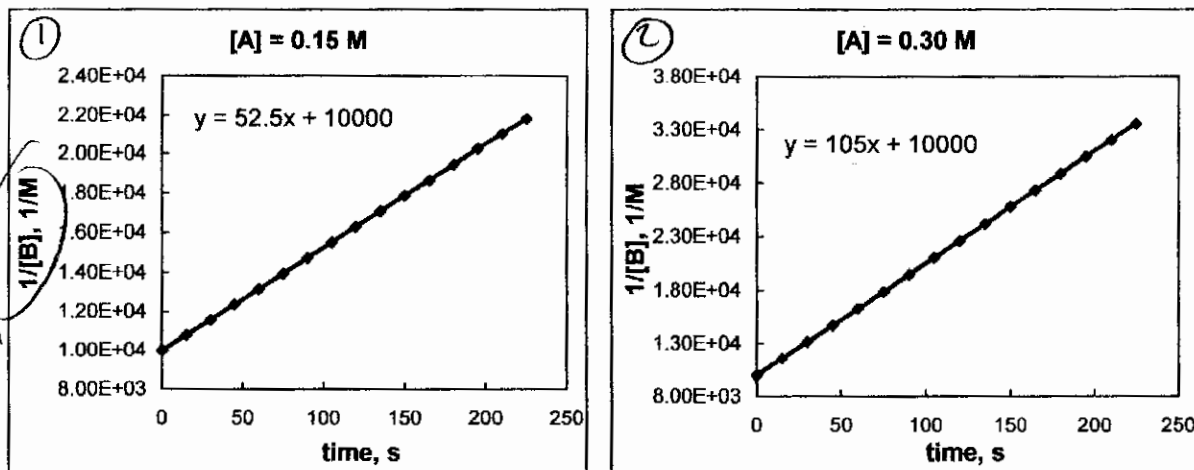
- (a) 84%    (b) 44%    (c) 16%    (d) 55%    (e) 92%

1<sup>st</sup> order eqn  $\ln\left(\frac{[A]_t}{[A]_0}\right) = -kt \Rightarrow \frac{[A]_t}{[A]_0} = e^{-kt}$  fraction that remains

$$\frac{[^{131}\text{I}]_t}{[^{131}\text{I}]_0} = e^{-(0.0862 \text{ day}^{-1})(1 \text{ week})(7 \text{ day/wk})}$$

$$= 0.547 \times 100\% \quad \boxed{54.7\%}$$

20. The rate law and rate constant for the reaction  $\text{A} + \text{B} \rightarrow \text{C}$  was determined using the method of pseudo-order kinetics (as in Exp. 8). To do this the reaction was run with a large excess of A compared to B. The concentration of B was measured as a function of time, for two different starting concentrations of A. The following two graphs were obtained from the data.



Based on the above plots, what is the rate law and what is the value of the true rate constant  $k$  for this reaction?

- (a) rate =  $k[\text{A}][\text{B}]^2$ ;  $k = 52.5 \text{ M}^{-2}\text{s}^{-1}$     (b) rate =  $k[\text{A}][\text{B}]$ ;  $k = 105 \text{ M}^{-1}\text{s}^{-1}$   
 (c) rate =  $k[\text{A}][\text{B}]^2$ ;  $k = 350 \text{ M}^{-2}\text{s}^{-1}$     (d) rate =  $k[\text{A}][\text{B}]$ ;  $k = 350 \text{ M}^{-1}\text{s}^{-1}$   
 (e) rate =  $k[\text{A}]^2[\text{B}]^2$ ;  $k = 105 \text{ M}^{-3}\text{s}^{-1}$

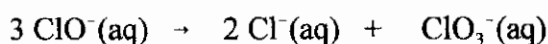
2<sup>nd</sup> order plot  
slope =  $k_{\text{obs}}$

rate =  $k[\text{A}]^x [\text{B}]^2$   
=  $k_{\text{obs}}$

$$\frac{\text{run 2}}{\text{run 1}} = \frac{k_{\text{obs}2}}{k_{\text{obs}1}} = \frac{k[0.30]^x}{k[0.15]^x} = 2^x = \frac{105}{52.5} = 2 \quad \text{so } x=1$$

$$k_{\text{obs}1} = k[0.15] = 52.5 \quad \text{so } k = \frac{52.5 \text{ M}^{-1}\text{s}^{-1}}{0.15 \text{ M}} = 350 \text{ M}^{-2}\text{s}^{-1}$$

Consider the following reaction for the next 6 questions. Hypochlorite,  $\text{ClO}^-$ , the active ingredient in bleach, decomposes to chloride and chlorate by the following reaction:



1. What is  $\Delta G^\circ$  of this reaction?  $\Delta G_f^\circ$  of  $\text{ClO}^-(\text{aq}) = -36.8 \text{ kJ}$ ;  $\Delta G_f^\circ$  of  $\text{Cl}^-(\text{aq}) = -131.3 \text{ kJ}$ ;  $\Delta G_f^\circ$  of  $\text{ClO}_3^-(\text{aq}) = -1.92 \text{ kJ}$ . ( $\Delta G_f^\circ$  values are at  $25^\circ\text{C}$ )

(a)  $-96.4 \text{ kJ}$  (b)  $262.6 \text{ kJ}$  (c)  $-154.1 \text{ kJ}$  (d)  $96.4 \text{ kJ}$  (e)  $-262.6 \text{ kJ}$  (ab)  $154.1 \text{ kJ}$

See Form A, # 10

2. What is  $K$  for this reaction at  $25^\circ\text{C}$ ?

(a)  $1.5 \times 10^6$  (b)  $1.8 \times 10^{-11}$  (c)  $5.6 \times 10^{10}$  (d)  $1.0 \times 10^{-27}$   
(e)  $1.0 \times 10^{27}$

See Form A, # 11

3. The above reaction is an oxidation/reduction reaction so it could take place in an electrochemical cell. In which of the following cells would this reaction occur?

(a)  $\text{Pt} / \text{ClO}^-(\text{aq}), \text{ClO}_3^-(\text{aq}) // \text{ClO}^-(\text{aq}), \text{Cl}^-(\text{aq}) / \text{Pt}$

(b)  $\text{Pt} / \text{Cl}^-(\text{aq}), \text{ClO}_3^-(\text{aq}) // \text{ClO}^-(\text{aq}) / \text{Pt}$

(c)  $\text{Pt} / \text{ClO}^-(\text{aq}) // \text{Cl}^-(\text{aq}), \text{ClO}_3^-(\text{aq}) / \text{Pt}$

(d)  $\text{Pt} / \text{ClO}^-(\text{aq}), \text{Cl}^-(\text{aq}) // \text{ClO}^-(\text{aq}), \text{ClO}_3^-(\text{aq}) / \text{Pt}$

See Form A, # 12

Continuation of the questions on the following reaction:



4. What does the  $\Delta G^\circ$  and K values say about the rate of this reaction?

- (a) The  $\Delta G^\circ$  and K values tell you nothing definite about the rate of the reaction.  
 (b) This will be a very slow reaction.  
 (c) This will be a very fast reaction.

5. Given the stoichiometry of the reaction, if, at a particular temperature, the rate of change of  $\text{ClO}^-$  was  $-9.0 \times 10^{-4} \text{ M/hr}$ , what would be the rate of change of  $\text{ClO}_3^-$  at the same time?

- (a)  $+6.0 \times 10^{-4} \text{ M/hr}$       (b)  $+9.0 \times 10^{-4} \text{ M/hr}$       (c)  $+3.0 \times 10^{-4} \text{ M/hr}$   
 (d)  $-6.0 \times 10^{-4} \text{ M/hr}$       (e)  $+2.7 \times 10^{-3} \text{ M/hr}$

takes 3  $\text{ClO}^-$  to make 1  $\text{ClO}_3^-$   
 so  $\text{ClO}_3^-$  must be increasing  $1/3$  as fast  
 as  $\text{ClO}^-$  is decreasing

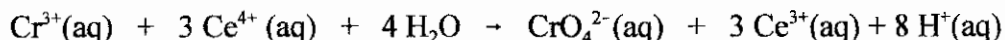
$$-1/3 (-9.0 \times 10^{-4} \text{ M/hr}) = \boxed{3.0 \times 10^{-4} \text{ M/hr}}$$

6. It is found experimentally that the rate of the reaction increases 4 times when the concentration of  $\text{ClO}^-$  is doubled. Assuming the products are not in the rate law, what is the rate law of this reaction?

- (a) rate =  $k[\text{ClO}^-]^{-1/2}$       (b) rate =  $k[\text{ClO}^-]^2$       (c) rate =  $k[\text{ClO}^-]$   
 (d) rate =  $k[\text{ClO}^-]^3$       (e) rate =  $k[\text{ClO}^-]^4$

See Form A, # 15

7. Consider the following reaction:



The rate law for this reaction has been determined by experiment to be

$$\text{rate} = k[\text{Cr}^{3+}]^2[\text{Ce}^{4+}]^1[\text{CrO}_4^{2-}]^{-1}$$

*add exponents*  
 $2 + 1 - 1 = 0$

What is the overall kinetic order of the reaction?

- (a) 4<sup>th</sup>      (b) 1<sup>st</sup>      (c) 3<sup>rd</sup>      (d) 2<sup>nd</sup>      (e) 8<sup>th</sup>

8. Given the rate law for the reaction, which one of the following statements best describes how the rate of the reaction will change with time?

- (a) The rate of the reaction will increase with time in the normal fashion as the reactant concentration decreases.  
 (b) The rate of the reaction will decrease with time in the normal fashion as the reactant concentration decreases.  
 (c) The rate of the reaction will decrease with time slower than normal because the build up of the product  $\text{CrO}_4^{2-}$  increases the rate of the reaction.  
(d) The rate of the reaction will decrease with time faster than normal because build up of the product  $\text{CrO}_4^{2-}$  slows the reaction further.

*rate  $\propto \frac{1}{[\text{CrO}_4^{2-}]}$  so as*

9. The following rates were measured for the reaction  $\text{A} + \text{B} \rightarrow \text{C}$ :

*$[\text{CrO}_4^{2-}] \uparrow$  rate  $\downarrow$*

Experiment	Initial [A], M	Initial [B], M	Initial rate, M/s
1	0.10	0.10	$1.50 \times 10^{-4}$
2	0.10	0.30	$4.50 \times 10^{-4}$
3	0.20	0.10	$3.00 \times 10^{-4}$

Based on the above information, what is the rate law for this reaction?

- (a) rate =  $k[\text{A}]^2$       (b) rate =  $k[\text{A}]^2[\text{B}]^2$       (c) rate =  $k[\text{A}]^3$   
 (d) rate =  $k[\text{A}][\text{B}]^3$       (e) rate =  $k[\text{A}][\text{B}]$       (ab) rate =  $k[\text{B}]^2$

$$\text{rate} = k[\text{A}]^x[\text{B}]^y$$

*to determine x, compare runs 1 & 3*

$$\frac{\text{run 3}}{\text{run 1}} = \frac{[\text{A}]_3}{[\text{A}]_1} = \frac{0.20^x}{0.10^x} = 2 \quad \frac{\text{rate 3}}{\text{rate 1}} = \frac{3.00 \times 10^{-4}}{1.50 \times 10^{-4}} = 2$$

$$2^x = 2 \Rightarrow x = 1$$

*to determine y, compare runs 1 & 2*

$$\frac{\text{run 2}}{\text{run 1}} = \frac{[\text{B}]_2}{[\text{B}]_1} = \frac{0.30^y}{0.10^y} = 3^y \quad \frac{\text{rate 2}}{\text{rate 1}} = \frac{4.50 \times 10^{-4}}{1.50 \times 10^{-4}} = 3$$

*so  $3^y = 3$   
 $y = 1$*

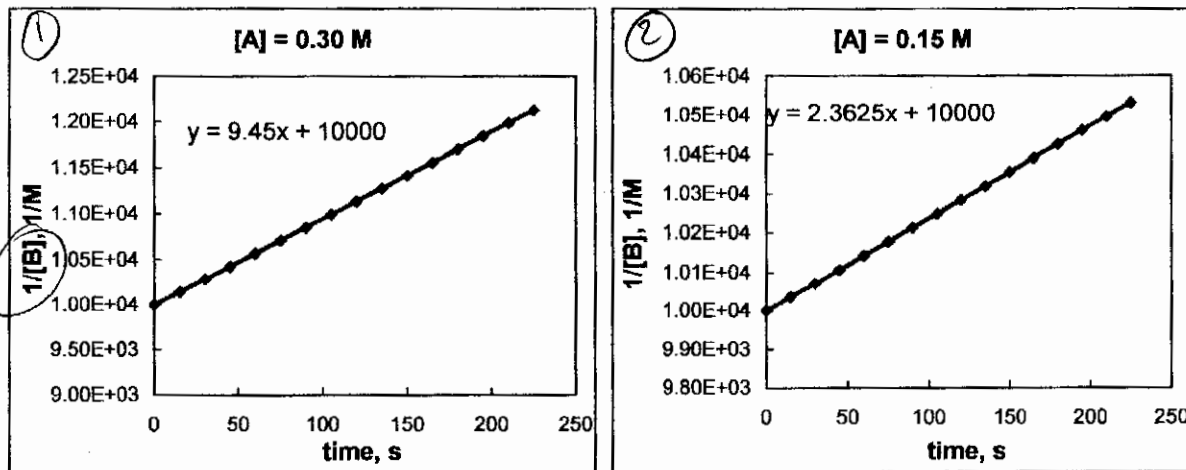
10. The decay of radioactive isotopes follow first order kinetics. If the rate constant for the radioactive decay of  $^{131}\text{I} = 0.0862 \text{ day}^{-1}$ , what percentage of  $^{131}\text{I}$  in a sample will remain after 3 weeks?

- (a) 55%    (b) 84%    (c) 16%    (d) 44%    (e) 77%

1<sup>st</sup> order eqn  $\ln\left(\frac{[A]_t}{[A]_0}\right) = -kt \Rightarrow \frac{[A]_t}{[A]_0} = e^{-kt}$

$$\frac{[^{131}\text{I}]_t}{[^{131}\text{I}]_0} = e^{-(0.0862 \text{ day}^{-1})(3 \text{ wk})(7 \text{ day/wk})} = 0.16 \times 100\% = \boxed{16\%}$$

11. The rate law and rate constant for the reaction  $\text{A} + \text{B} \rightarrow \text{C}$  was determined using the method of pseudo-order kinetics (as in Exp. 8). To do this the reaction was run with a large excess of A compared to B. The concentration of B was measured as a function of time, for two different starting concentrations of A. The following two graphs were obtained from the data.



Based on the above plots, what is the rate law and what is the value of the true rate constant  $k$  (not  $k_{\text{obs}}$ ) for this reaction?

- (a) rate =  $k[A][B]^2$ ;  $k = 350 \text{ M}^{-2}\text{s}^{-1}$   
 (b) rate =  $k[A][B]$ ;  $k = 2.36 \text{ M}^{-1}\text{s}^{-1}$   
 (c) rate =  $k[A][B]^2$ ;  $k = 9.45 \text{ M}^{-2}\text{s}^{-1}$   
 (d) rate =  $k[A][B]$ ;  $k = 9.45 \text{ M}^{-1}\text{s}^{-1}$   
 (e) rate =  $k[A]^2[B]^2$ ;  $k = 105 \text{ M}^{-3}\text{s}^{-1}$

2<sup>nd</sup> order plot so slope =  $k_{\text{obs}}$

rate =  $k_{\text{obs}} [A]^x [B]^2$

$$\frac{\text{run 1}}{\text{run 2}} = \frac{k_{\text{obs 1}}}{k_{\text{obs 2}}} = \frac{9.45}{2.3625} = 4 \Rightarrow x = 2$$

$$\text{run 1 } k_{\text{obs}} = k [0.30]^2 = 9.45 \Rightarrow k = \frac{9.45}{(0.30)^2} = \boxed{105 \text{ M}^{-3}\text{s}^{-1}}$$

Consider the following *unbalanced* oxidation reduction reaction for the next three questions.



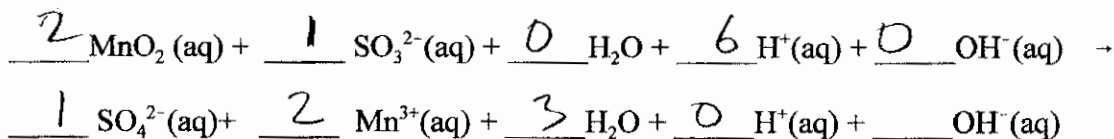
12. The oxidation number of S in  $\text{SO}_4^{2-}$  is +6.

- (a) +8      (b) -2      (c) +3      (d) +6      (e) +4
- $+6 - 8 = -2$

13. What is being oxidized in this reaction?

- (a)  $\text{SO}_4^{2-}$       (b)  $\text{SO}_3^{2-}$       (c)  $\text{Mn}^{3+}$       (d)  $\text{MnO}_2$
- $\text{MnO}_2 \rightarrow \text{Mn}^{3+}$       Mn being reduced, ox # ↓
- $\text{SO}_3^{2-} \rightarrow \text{SO}_4^{2-}$       S " oxidized, ox # ↑

14. The complete balanced equation in **ACIDIC** media for the reaction has the following stoichiometry coefficients. (Note that some of the values may be 0 indicating that the compound does not appear!)

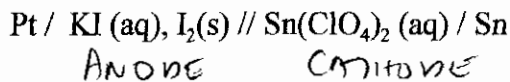


The stoichiometry coefficients, in order are

- (a) 2,1,0,6,0 → 1,2,3,0,0      (b) 1,1,0,2,0 → 1,1,1,0,0
- (c) 1,1,1,0,0 → 1,1,0,2,0      (d) 1,2,0,0,6 → 2,1,3,0,0
- (e) 1,2,3,0,0 → 2,1,0,6,0

See Form A # 3  
(rxn is just the opposite one)

Consider the following electrochemical cell for the next 3 questions:

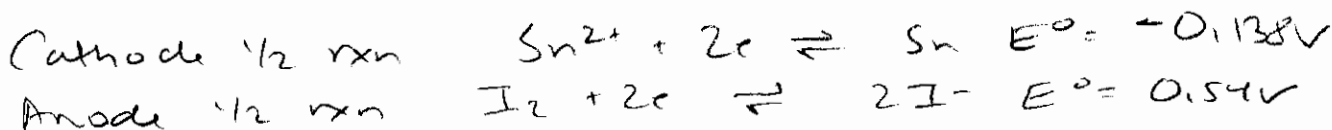


15. What is the name of the aqueous ionic compound in the anode compartment?

- (a) tin(II) perchlorate      (b) potassium iodide      (c) sodium iodate  
(d) sodium iodine          (e) tin chloride

16. What is  $E^\circ$  cell?

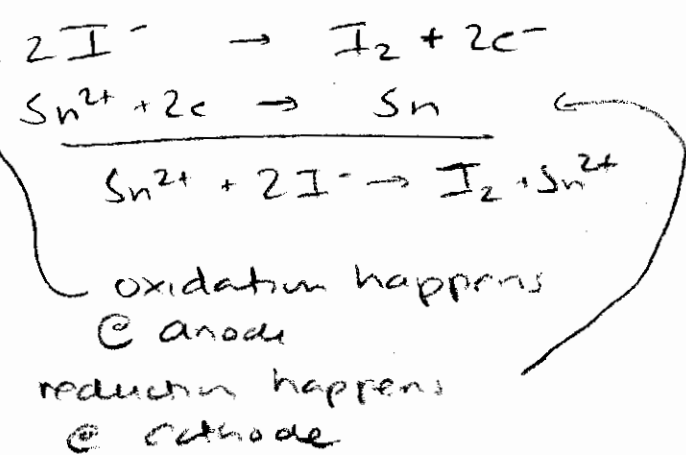
- (a) 0.68 V    (b) -0.34 V    (c) -0.40 V    (d) +0.40 V    (e) 0.34 V    (ab) -0.68 V



$$E^\circ_{\text{cell}} = E^\circ_{\text{cat}} - E^\circ_{\text{an}} = -0.138\text{V} - 0.54\text{V} = \boxed{-0.68\text{V}}$$

17. What is overall net ionic reaction occurring in the cell when current flows?

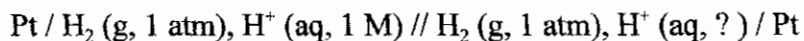
- (a)  $\text{Sn}(\text{s}) + \text{I}_2(\text{s}) \rightarrow \text{Sn}^{2+}(\text{aq}) + 2\text{I}^-(\text{aq})$   
 (b)  $\text{I}_2(\text{s}) + 2\text{K}^+(\text{aq}) \rightarrow 2\text{KI}(\text{aq})$   
 (c)  $\text{Sn}(\text{s}) + 2\text{ClO}_4^-(\text{aq}) \rightarrow \text{Sn}(\text{ClO}_4)_2(\text{aq})$   
 (d)  $\text{Sn}^{2+}(\text{aq}) + 2\text{I}^-(\text{aq}) \rightarrow \text{Sn}(\text{s}) + \text{I}_2(\text{s})$   
 (e)  $\text{Sn}(\text{s}) + 2\text{I}^-(\text{aq}) \rightarrow \text{Sn}^{2+}(\text{aq}) + \text{I}_2(\text{s})$



18. What does it mean when the oxidation states do NOT change in a reaction?

- (a) The reaction is a very unfavorable oxidation-reduction reaction.  
 (b) The reaction is a very favorable oxidation-reduction reaction.  
 (c) The reaction is not an oxidation-reduction reaction. !  
 (d) The reaction is just a reduction reaction.  
 (e) The reaction is just an oxidation reaction.

19. Before the development of the pH electrode (in the 1950's), the following concentration cell was routinely used to measure pH:



If  $E_{\text{cell}} = -0.229$  what is the pH of the solution in the compartment on the right?

- (a) 12.26    (b) 7.74    (c) 3.87    (d) 8.45    (e) 6.13

See Form A, # 8  
for setup

$$\frac{0.229 \text{ V}}{0.0592 \text{ V}} = -\log [\text{H}^+] = \text{pH}$$
$$= \boxed{3.87}$$

20. Which one of the following is NOT necessarily TRUE for a chemical reaction at equilibrium?

- (a) The concentrations of reactants and products do not change with time.  
(b) The rate of the forward reaction = the rate of the reverse reaction.  
(c)  $\Delta G = 0$   
(d) The reaction quotient,  $Q = K$   
(e)  $\Delta G^\circ = 0$

} ALL TRUE  
C equilibrium

## EQUATIONS and ADDITIONAL INFO

$$E_{\text{cell}} = E_{\text{cat}} - E_{\text{an}} \quad \text{oO} + n\text{e}^- \rightleftharpoons \text{rR}$$

$$\Delta G = -nFE_{\text{cell}} \quad E = E^\circ - \frac{0.0592V}{n} \log \frac{[R]^r}{[O]^o} \quad \text{at } 25^\circ\text{C}$$

$$\Delta G^\circ = -RT \ln K \quad \text{or} \quad \Delta G^\circ = -2.303 RT \log K$$

$$K = e^{-\frac{\Delta G^\circ}{RT}} \quad \text{or} \quad K = 10^{-\frac{\Delta G^\circ}{2.303RT}}$$

$$\ln \left( \frac{[A]_t}{[A]_0} \right) = -kt \quad \text{or} \quad \log \left( \frac{[A]_t}{[A]_0} \right) = -\frac{kt}{2.303}$$

$$\ln[A]_t = -kt + \ln[A]_0 \quad \text{or} \quad \log[A]_t = -kt/2.303 + \log[A]_0$$

$$1/[A]_t = kt + 1/[A]_0$$

$$F = 96485 \text{ coul/mol} \quad R = 8.314 \text{ J/(mol}\cdot\text{K)} \quad \text{Volt} = \text{Joule/coul} \quad 0^\circ\text{C} = 273.15 \text{ K}$$

### Standard Electrode Potentials at 25°C

<u>Half Reaction</u>	<u>E° (V)</u>
Ag <sup>+</sup> (aq) + e <sup>-</sup> ⇌ Ag(s)	+0.80 V
AgCl(s) + e <sup>-</sup> ⇌ Ag(s) + Cl <sup>-</sup> (aq)	+0.2223 V
AgBr(s) + e <sup>-</sup> ⇌ Ag(s) + Br <sup>-</sup> (aq)	+0.07 V
Br <sub>2</sub> (l) + 2 e <sup>-</sup> ⇌ 2 Br <sup>-</sup> (aq)	+1.09 V
Cl <sub>2</sub> (g) + 2 e <sup>-</sup> ⇌ 2 Cl <sup>-</sup> (aq)	+1.36
Cu <sup>+</sup> (aq) + e <sup>-</sup> ⇌ Cu(s)	+0.52 V
Cu <sup>2+</sup> (aq) + 2 e <sup>-</sup> ⇌ Cu(s)	+0.34 V
Cu <sup>2+</sup> (aq) + e <sup>-</sup> ⇌ Cu <sup>+</sup> (aq)	+0.159 V
Cu <sup>2+</sup> (aq) + 2 CN <sup>-</sup> (aq) + e <sup>-</sup> ⇌ Cu(CN) <sub>2</sub> <sup>-</sup> (aq)	+1.12 V
Fe <sup>2+</sup> (aq) + 2 e <sup>-</sup> ⇌ Fe(s)	-0.45 V
Fe <sup>3+</sup> (aq) + e <sup>-</sup> ⇌ Fe <sup>2+</sup> (aq)	+0.77 V
2 H <sup>+</sup> (aq) + 2 e <sup>-</sup> ⇌ H <sub>2</sub> (g)	0.000V
2 H <sub>2</sub> O(aq) + 2 e <sup>-</sup> ⇌ H <sub>2</sub> (g) + 2 OH <sup>-</sup> (aq)	-0.828 V
2 Hg <sup>2+</sup> (aq) + 2 e <sup>-</sup> ⇌ 2 Hg <sub>2</sub> <sup>2+</sup>	0.905 V
Hg <sub>2</sub> <sup>2+</sup> (aq) + 2 e <sup>-</sup> ⇌ 2 Hg(l)	0.7961 V
Hg <sub>2</sub> SO <sub>4</sub> (s) + 2 e <sup>-</sup> ⇌ 2 Hg(l) + SO <sub>4</sub> <sup>2-</sup>	0.6158 V
I <sub>2</sub> (s) + 2 e <sup>-</sup> ⇌ 2 I <sup>-</sup> (aq)	+0.54 V
Mg <sup>2+</sup> (aq) + 2 e <sup>-</sup> ⇌ Mg(s)	-2.37 V
Ni <sup>2+</sup> (aq) + 2 e <sup>-</sup> ⇌ Ni(s)	-0.257 V
Ni(OH) <sub>2</sub> (s) + 2 e <sup>-</sup> ⇌ Ni(s) + 2 OH <sup>-</sup> (aq)	-0.72 V
Pb <sup>2+</sup> (aq) + 2 e <sup>-</sup> ⇌ Pb(s)	-0.126 V
PbO <sub>2</sub> (s) + 3 H <sup>+</sup> (aq) + HSO <sub>4</sub> <sup>-</sup> (aq) + 2 e <sup>-</sup> ⇌ PbSO <sub>4</sub> (s) + 2 H <sub>2</sub> O(l)	+1.628 V
PbSO <sub>4</sub> (s) + H <sup>+</sup> + 2 e <sup>-</sup> ⇌ Pb(s) + HSO <sub>4</sub> <sup>-</sup>	-0.296 V
S <sub>2</sub> O <sub>8</sub> <sup>2-</sup> (aq) + 2 e <sup>-</sup> ⇌ 2 SO <sub>4</sub> <sup>2-</sup> (aq)	+2.01 V
Sn <sup>4+</sup> (aq) + 2 e <sup>-</sup> ⇌ Sn <sup>2+</sup>	+0.15 V
Sn <sup>2+</sup> (aq) + 2 e <sup>-</sup> ⇌ Sn(s)	-0.138 V
Zn <sup>2+</sup> (aq) + 2 e <sup>-</sup> ⇌ Zn(s)	-0.76 V