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Exam 4

Fall 2018

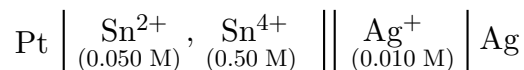Dec 5, 2018

Wednesday 7:30 - 9:00 PM

BUR 106

Remember to refer to the Periodic Table handout that is separate from this exam copy.
A table of standard potentials is on the back of the periodic table handout.

NOTE: Please keep this exam copy intact (all pages still stapled - including this cover page). You must turn in ALL the materials that were distributed. This means that you turn in your exam copy (name and signature included), bubble sheet, periodic table handout, and all scratch paper. Please also have your UT ID card ready to show as well.

This print-out should have 20 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

001　5.0 points
Consider the following voltaic cell:

$$\text{Pt} \left| \begin{array}{cc} \text{Sn}^{2+} & \text{Sn}^{4+} \\ (0.050 \text{ M}) & (0.50 \text{ M}) \end{array} \right. , \left\| \begin{array}{c} \text{Ag}^+ \\ (0.010 \text{ M}) \end{array} \right| \text{Ag}$$

The experimental cell potential for the cell is closest to

1. 0.502 V **correct**

2. 0.561 V

3. 0.739 V

4. 0.354 V

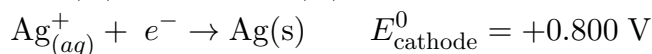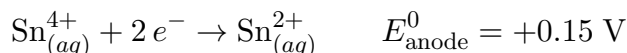**5.** 0.650 V

6. 0.798 V

7. 0.946 V

Explanation:
The overall cell reaction is:
$$2\text{Ag}^+ + \text{Sn}^{2+} \to 2\text{Ag} + \text{Sn}^{4+}$$
Calculate E°_{cell} first:

$$\text{Sn}^{4+}_{(aq)} + 2\,e^- \to \text{Sn}^{2+}_{(aq)} \qquad E^0_{\text{anode}} = +0.15 \text{ V}$$
$$\text{Ag}^+_{(aq)} + e^- \to \text{Ag(s)} \qquad E^0_{\text{cathode}} = +0.800 \text{ V}$$
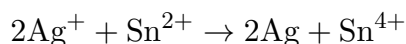
$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$E^\circ_{\text{cell}} = 0.800 - (+0.15) = 0.65 \text{ V}$$

The stated concentrations are not standard. This means that you must use the Nernst equation to calculate E_{cell} at these concentrations:

$$E = E^\circ - \frac{0.05916 \text{ V}}{n} \log\left(\frac{[\text{Sn}^{4+}]}{[\text{Sn}^{2+}][\text{Ag}^+]^2}\right)$$
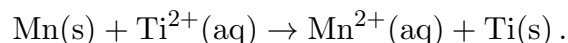
$$E = 0.65 \text{ V} - \frac{0.05916 \text{ V}}{2} \log\left(\frac{.50}{(0.050)(.01)^2}\right)$$

$$E = 0.65 \text{ V} - 0.1479 \text{ V}$$
$$E = 0.502 \text{ V}$$

002　5.0 points
Consider the half-reactions and the balanced equation for the cell reaction represented by the skeletal equation

$$\text{Mn(s)} + \text{Ti}^{2+}(\text{aq}) \to \text{Mn}^{2+}(\text{aq}) + \text{Ti(s)}.$$

What is its standard potential?

1. 1.405 V
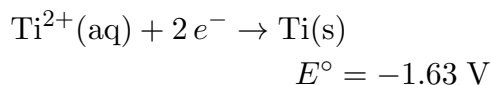
2. 2.81 V

3. −1.405 V

4. −0.45 V **correct**

5. 0.45 V

6. −2.81 V

Explanation:
The two half-reactions, written as reductions, are

$$\text{Mn}^{2+}(\text{aq}) + 2\,e^- \to \text{Mn(s)}$$
$$E^\circ = -1.18 \text{ V}$$
$$\text{Ti}^{2+}(\text{aq}) + 2\,e^- \to \text{Ti(s)}$$
$$E^\circ = -1.63 \text{ V}$$

Equate e^-:

$$\text{Ti}^{2+}(\text{aq}) + \cancel{2\,e^-} \to \text{Ti(s)} \quad E^\circ = -1.63 \text{ V}$$
$$\text{Mn(s)} \to \text{Mn}^{2+}(\text{aq}) + \cancel{2\,e^-}$$
$$E^\circ = +1.18 \text{ V}$$
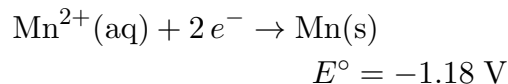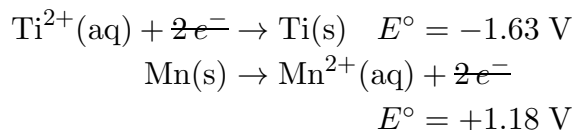
Add the balanced half reactions:

$$\text{Mn(s)} + \text{Ti}^{2+}(\text{aq}) \to \text{Mn}^{2+}(\text{aq}) + \text{Ti(s)}$$
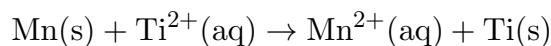
$$E^\circ_{\text{cell}} = -1.63 \text{ V} - (-1.18 \text{ V}) = -0.45 \text{ V}$$

003　5.0 points
A reaction has $k = 8.39 \text{ M}^{-1} \cdot \text{s}^{-1}$. How long

does it take for the reactant concentration to drop from 0.0840 M to 0.0220 M?

1. 4.00 s **correct**

2. 8.39 s

3. 5.42 s

4. 2.00 s

5. 1.42 s

Explanation:

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004   5.0 points

Suppose for some reaction the rate constant doubles in going from $T = 50°C$ to $70°C$. What is E_a for this reaction?

1. 24.0 kJ/mol

2. 8.12 kJ/mol

3. 39.2 kJ/mol

4. 16.1 kJ/mol

5. 1.01 kJ/mol

6. 32.0 kJ/mol **correct**

Explanation:
$T_1 = 50°C + 273.15 = 323.15$ K $k_2 = 2\,k_1$
$T_2 = 70°C + 273.15 = 343.15$ K

$$\ln \frac{k_2}{k_1} = \frac{E_a}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$$E_a = \frac{R \, \ln \dfrac{k_2}{k_1}}{\dfrac{1}{T_1} - \dfrac{1}{T_2}}$$

$$= \frac{(8.314 \text{ J/mol} \cdot \text{K}) \ln 2}{\dfrac{1}{323.15 \text{ K}} - \dfrac{1}{343.15 \text{ K}}}$$

$$= 31.95 \text{ kJ/mol}$$

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005   5.0 points

Indium metal, In(s), is electroplated from a concentrated solution of indium (III) chloride. 12.5 amps of current is passed for a total of 96 hours. What mass of indium is plated?

1. 1.30 kg

2. 1.71 kg **correct**

3. 4.24 kg

4. 5.14 kg

5. 3.92 kg

6. 2.57 kg

Explanation:
$\frac{I \cdot t}{n \cdot F}$ = moles of metal

$\frac{12.5(96 \cdot 3600)}{3 \cdot 96485}$ = 14.92 mol In

In = 114.8 g/mol

14.92(114.8) = 1713 g = 1.71 kg

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006   5.0 points

For the cell diagram

$$\text{Pt} \,|\, \text{H}_2(\text{g}) \,|\, \text{H}^+(\text{aq}) \,||\, \text{Co}^{3+}(\text{aq}), \text{Co}^{2+}(\text{aq}) \,|\, \text{Pt}$$

what reaction occurs at the cathode?

1. $\text{Co}^{2+}(\text{aq}) \to \text{Co}^{3+}(\text{aq}) + e^-$

2. $\text{Co}^{3+}(\text{aq}) + e^- \to \text{Co}^{2+}(\text{aq})$ **correct**

3. $\text{Co}^{3+}(\text{aq}) + 2\,e^- \to \text{Co}^+(\text{aq})$

4. $\frac{1}{2}\,\text{H}_2(\text{g}) \to \text{H}^+(\text{aq}) + e^-$

5. $2\,\text{H}_2\text{O}(\ell) + 2\,e^- \to \text{H}_2(\text{g}) + 2\,\text{OH}^-(\text{aq})$

Explanation:

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007   5.0 points

A catalyst works by

1. providing an alternate reaction pathway with a higher activation energy.

2. increasing the average kinetic energy of the reactant molecules.

3. making the reaction more exothermic and therefore more favorable.

4. providing an alternate reaction pathway with a lower activation energy. **correct**

Explanation:

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008   5.0 points

Consider the data collected for a chemical reaction between compounds A and B that is first order in A and first order in B:

| [A] | [B] | rate |
M	M	M/s
1 0.2	0.05	0.1
2 ?	0.05	0.4
3 0.4	?	0.8

From the information above for 3 experiments, determine the missing concentrations of A and B. Answers should be in the order of [A] then [B].

1. 0.40 M; 0.20 M

2. 0.80 M; 0.20 M **correct**

3. 0.20 M; 0.80 M

4. 0.40 M; 0.10 M

5. 1.60 M; 0.40 M

6. 0.80 M; 0.10 M

Explanation:

As we know the reaction is first order in A and first order in B, the rate law has the form

$$\text{Rate} = k[\text{A}][\text{B}]$$

To determine the rate constant k, consider trial 1:

$$\text{Rate} = k[\text{A}][\text{B}]$$
$$0.1 \text{ M/s} = k(0.2 \text{ M})(0.05 \text{ M})$$
$$k = 10 \text{ M}^{-1} \cdot \text{s}^{-1}$$

So, in the second trial,

$$\text{Rate} = k[\text{A}][\text{B}]$$
$$0.4 \text{ M/s} = 10 \text{ M}^{-1} \cdot \text{s}^{-1}[\text{A}](0.05 \text{ M})$$
$$[A] = 0.80 \text{ M}$$

And in the third trial,

$$\text{Rate} = k[\text{A}][\text{B}]$$
$$0.8 \text{ M/s} = 10 \text{ M}^{-1} \cdot \text{s}^{-1}(0.4 \text{ M})[\text{B}]$$
$$[B] = 0.20 \text{ M}$$

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009   5.0 points

Which type of widely used battery is NOT rechargeable?

1. alkaline **correct**

2. nickel-cadmium (NiCad)

3. lead-acid (storage batteries)

4. lithium-ion

Explanation:

Alkaline batteries were not designed to be rechargeable and do not do so efficiently, although there are some websites that disagree.

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010   5.0 points

Three separate experiments were performed on the rate of the reaction:

$$3 \text{ A}_2 + 2 \text{ B} \rightarrow 2 \text{ A}_3\text{B}$$

The measured initial concentrations of A_2 and B (in moles per liter) are shown below along with the measured initial rates of formation of A_3B (moles per liter per second).

Trial	Initial $[A_2]_0$ M	Initial $[B]_0$ M	Initial rate M/s
1	1.5	3.0	7.0×10^{-8}
2	1.5	1.5	1.75×10^{-8}
3	4.5	3.0	2.1×10^{-7}

The reaction is...

1. first order in $[A_2]$ and second order in $[B]$ **correct**

2. second order in $[A_2]$ and first order in $[B]$

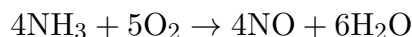**3.** first order in $[A_2]$ and first order in $[B]$

4. third order in $[A_2]$ and second order in $[B]$

5. None of these is correct.

Explanation:

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011   5.0 points
Consider the following reaction.

$$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$$

If the rate of change for NO is 0.016 M/s, what is the rate of change for O_2 ?

1. +0.018 M/s

2. +0.024 M/s

3. −0.024 M/s

4. −0.020 M/s **correct**

5. +0.020 M/s

6. −0.018 M/s

7. +0.013 M/s

8. −0.013 M/s

Explanation:
   The rate of change for O_2 will be 5/4 that of the NO rate based on stoichiometry and opposite in sign due to the fact that O_2 is a reactant and NO is a product. Therefore
   rate change $O_2 = -5/4(0.16) = -0.20$ M/s

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012   5.0 points
Which species will reduce Ag^+ but not Fe^{2+}?

1. Cr

2. Pt
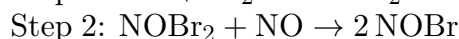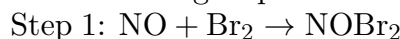
3. H_2 **correct**

4. Au

5. V

Explanation:

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013   5.0 points
What would you propose as the rate law for the reaction of bromine with nitric oxide if the second step of a proposed mechanism is the rate determining step?
   Step 1: $NO + Br_2 \rightarrow NOBr_2$
   Step 2: $NOBr_2 + NO \rightarrow 2\,NOBr$

1. $k\,[NO]^2$

2. $k\,[NO]\,[Br_2]\,[NOBr_2]$

3. $k\,[NO]^2\,[Br_2]\,[NOBr_2]^{-1}$

4. $k\,[NO]\,[Br_2]\,[NOBr_2]^{-1}$

5. $k\,[NO]^2\,[Br_2]$ **correct**

Explanation:

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014   5.0 points
For a first-order reaction, after 230 s, 33% of the reactants remain. Calculate the rate constant for the reaction.

1. 0.00209 s^{-1}

2. 0.00482 s^{-1} **correct**

3. 0.000756 s^{-1}

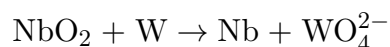**4.** 207 s^{-1}

5. 0.00174 s^{-1}

Explanation:

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015   5.0 points
Balance the following redox reaction in acidic solution. You will have to provide the H_2O

and the H^+ for the reaction. Make sure all the coefficients are whole numbers.

$$NbO_2 + W \rightarrow Nb + WO_4^{2-}$$

What is the coefficient for WO_4^{2-} in the balanced equation?

1. 2 **correct**

2. 4

3. 1

4. 6

5. 3

6. 5

Explanation:

$$3NbO_2 + 2W \quad \rightarrow \quad 3Nb + 2WO_4^{2-}$$
$$+ 2H_2O \qquad\qquad + 4H^+$$
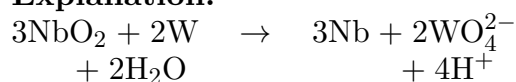
016   5.0 points

The following reaction

$$A + B \rightarrow C$$

is found to follow the rate law

rate $=k[A][B]$

when will a plot of ln[A] vs time yield a straight line?

1. when the $[B] >> [A]$ **correct**

2. when [B] = [A]

3. never

4. when the $[B] << [A]$

5. always

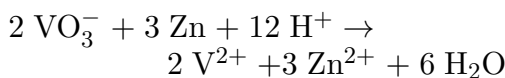**Explanation:**
The reaction is first order in A and B. For the plot of ln[A] vs time to yield a straight line plot the reaction must be only first order in A. This can be effectively achieved by making the concentration of [B] so high that during the course of the reaction it is not changing.

Then the reaction is pseudo-first order in A.

017   5.0 points

In the balanced reaction
$$2 VO_3^- + 3 Zn + 12 H^+ \rightarrow$$
$$2 V^{2+} + 3 Zn^{2+} + 6 H_2O$$
which species is the oxidizing agent?

1. H^+

2. H_2O

3. Zn^{2+}

4. VO_3^- **correct**

5. Zn

6. V^{2+}

Explanation:
In a redox reaction the reducing agent is the species that starts out in a reduced state and is oxidized as the reaction progresses from left to right. The complementary species to the reducing agent is the oxidizing agent, which is starts out in an oxidized state and is reduced as the reaction progresses from left to right. In this case, that species would be VO_3^- in which V undergoes a reduction from a +5 oxidation state to a +2 oxidation state in the form of V^{2+}.

018   5.0 points

A certain reaction (X $\rightarrow$ Y) is run and the half-life is found to be 7.0 minutes. The same reaction is run again using half the original amount of reactant X and the half-life changed to 3.5 minutes. What is the order of this reaction?

1. You must run one more experiment before you can tell.

2. first order

3. second order

4. Half lives are not indicators of order.

5. zero order **correct**

Explanation:
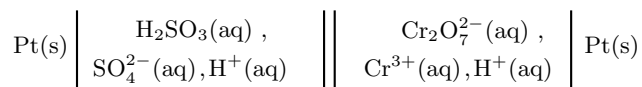For the first reaction
$t_{1/2_1} = 7.0$ min
For the second reaction
$[A]_0' = \frac{1}{2} [A]_0$ $t_{1/2}' = 3.5$ min
In zero order kinetics, half-lifes are DIRECTLY proportional to half-life.

$$t_{1/2} = \frac{[A]_0}{2\,a\,k}$$

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019   5.0 points
Calculate the emf of the following electrochemical cell:

$$Pt(s) \left| \begin{array}{c} H_2SO_3(aq)\,, \\ SO_4^{2-}(aq), H^+(aq) \end{array} \right| \left| \begin{array}{c} Cr_2O_7^{2-}(aq)\,, \\ Cr^{3+}(aq), H^+(aq) \end{array} \right| Pt(s)$$

1. -1.50 V

2. -1.16 V

3. -0.24 V

4. $+1.50$ V

5. $+0.24$ V

6. $+1.16$ V **correct**

Explanation:
Reduction potentials:
$SO_4^{2-} + 4H^+ + 2e^- \rightarrow H_2SO_3 + H_2O$, $E^\circ = +0.17$ V;
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$, $E^\circ = +1.33$ V
Summing the oxidation potential (-0.17) with the reduction potential $(+1.33)$ gives the cell potential $+1.16$ V
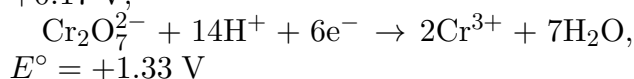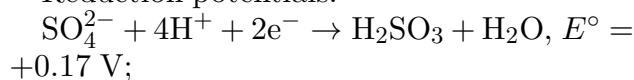
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020   5.0 points
For a given electrochemical cell, E^0 was calculated to be -0.7V for a two-electron process. What is K_{eq} for the reaction at $T = 298$ K?

1. 1.74×10^{30}

2. 1.45×10^{-12}

3. 2.94

4. 1.05×10^{-27}

5. 2.1×10^{-24} **correct**

Explanation:
$E^0 = -0.7$ V $T = 298$ K

$$-n\,E\,F = -R\,T\,\ln K$$
$$\ln K = \frac{n\,F\,E}{R\,T}$$
$$= \frac{2\text{ mol}\left(\dfrac{96485\text{ J}}{\text{V}\cdot\text{mol}}\right)(-0.7\text{ V})}{(8.314\text{ J/mol}\cdot\text{K})(298\text{ K})}$$
$$= -54.5207$$
$$K = 2.1 \times 10^{-24}$$