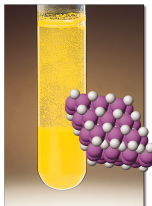
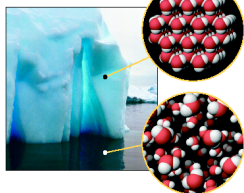
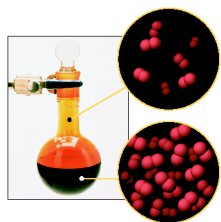


## Chemistry 223 Exam II Review Chapters 15, 16 and 17



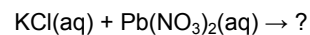
Chemistry 223  
Professor Michael Russell

MAR

Last update:  
6/14/23

MAR

Decide if a precipitate will form when mixing the indicated reagents (all concentrations are 1.0 M).



- A. Yes
- B. No
- C. Who knows!

If the solubility of  $\text{BaF}_2$  is  $3.6 \times 10^{-3}$ , a reasonable value for  $K_{\text{sp}}$  for  $\text{BaF}_2$  is

- A.  $3.6 \times 10^{-3}$
- B.  $7.2 \times 10^{-3}$
- C.  $1.1 \times 10^{-2}$
- D.  $1.9 \times 10^{-7}$
- E.  $4.7 \times 10^{-8}$

MAR

Which lead salt has the greatest molar solubility in water at 25 °C?

- A.  $\text{PbCO}_3$        $K_{\text{sp}} = 1.5 \times 10^{-13}$
- B.  $\text{PbS}$           $K_{\text{sp}} = 8.4 \times 10^{-28}$
- C.  $\text{PbSO}_4$        $K_{\text{sp}} = 1.8 \times 10^{-4}$

MAR

A solution contains 0.10 M  $\text{K}_2\text{SO}_3$  and 0.30 M  $\text{Na}_2\text{SO}_4$ . Solid  $\text{Ca(NO}_3)_2$  is added slowly. Which precipitates first,  $\text{CaSO}_3$  or  $\text{CaSO}_4$ ?

$$K_{\text{sp}} \text{ for } \text{CaSO}_3 = 1.3 \times 10^{-8}$$

$$K_{\text{sp}} \text{ for } \text{CaSO}_4 = 2.4 \times 10^{-5}$$

- A.  $\text{CaSO}_3$
- B.  $\text{CaSO}_4$
- C. 42

MAR

$\text{CaSO}_3$  precipitates first as  $\text{Ca}^{2+}$  ions are added to a solution containing 0.10 M  $\text{K}_2\text{SO}_3$  and 0.30 M  $\text{Na}_2\text{SO}_4$ . What is  $[\text{SO}_3^{2-}]$  as the  $\text{CaSO}_4$  begins to precipitate?

$$K_{\text{sp}}(\text{CaSO}_3) = 1.3 \times 10^{-8} \quad K_{\text{sp}}(\text{CaSO}_4) = 2.4 \times 10^{-5}$$

- A. 0.10 M
- B. 0.30 M
- C.  $1.6 \times 10^{-4}$  M
- D.  $5.4 \times 10^{-4}$  M
- E. 42

MAR

What is the pH of a saturated solution of  $\text{Mg}(\text{OH})_2$ ? ( $K_{\text{sp}}(\text{Mg}(\text{OH})_2) = 5.6 \times 10^{-12}$ )

- A. 3.65
- B. 8.37
- C. 10.35
- D. 0.15
- E. 11.25

MAR

A solution has  $[\text{Pb}^{2+}] = 0.0012 \text{ M}$  and  $[\text{Cl}^-] = 0.010 \text{ M}$ . Will  $\text{PbCl}_2$  precipitate?  
 $K_{\text{sp}}(\text{PbCl}_2) = 1.7 \times 10^{-5}$

- A. Yes,  $\text{PbCl}_2$  precipitates
- B. No,  $\text{PbCl}_2$  does NOT precipitate

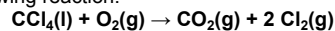
MAR

Which of the following shows the correct formation constant ( $K_f$ ) equation for  $\text{Cr}(\text{CN})_6^{3-}$ ?

- A.  $\text{Cr}(\text{CN})_3(\text{s}) + 3 \text{CN}^{-1}(\text{aq}) \rightleftharpoons \text{Cr}(\text{CN})_6^{3-}(\text{aq})$
- B.  $\text{Cr}(\text{NO}_3)_3(\text{s}) + 6 \text{NaCN}(\text{aq}) \rightleftharpoons \text{Cr}(\text{CN})_6^{3-}(\text{aq}) + 3 \text{NaNO}_3(\text{aq}) + 3 \text{Na}^+(\text{aq})$
- C.  $\text{Cr}(\text{CN})_6^{3-}(\text{s}) \rightleftharpoons \text{Cr}^{3+}(\text{aq}) + 6 \text{CN}^{-1}(\text{aq})$
- D.  $\text{Cr}(\text{CN})_6^{3-}(\text{aq}) \rightleftharpoons \text{Cr}^{3+}(\text{aq}) + 6 \text{CN}^{-1}(\text{aq})$
- E.  $\text{Cr}^{3+}(\text{aq}) + 6 \text{CN}^{-1}(\text{aq}) \rightleftharpoons \text{Cr}(\text{CN})_6^{3-}(\text{aq})$

MAR

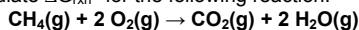
Calculate the standard entropy change for the following reaction:



- \*  $S^\circ[\text{CCl}_4(\text{l})] = 214.39 \text{ J/K} \cdot \text{mol}$
  - \*  $S^\circ[\text{CO}_2(\text{g})] = 213.74 \text{ J/K} \cdot \text{mol}$
  - \*  $S^\circ[\text{O}_2(\text{g})] = 205.07 \text{ J/K} \cdot \text{mol}$
  - \*  $S^\circ[\text{Cl}_2(\text{g})] = 223.08 \text{ J/K} \cdot \text{mol}$
- A. -17.36 J/K
  - B. +17.36 J/K
  - C. +240.44 J/K
  - D. -25.78 J/K
  - E. 42

MAR

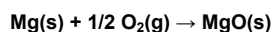
Calculate  $\Delta G_{\text{rxn}}^\circ$  for the following reaction:



- \*  $\Delta G_f^\circ[\text{CO}_2(\text{g})] = -394.4 \text{ kJ/mol}$
  - \*  $\Delta G_f^\circ[\text{CH}_4(\text{g})] = -50.8 \text{ kJ/mol}$
  - \*  $\Delta G_f^\circ[\text{H}_2\text{O}(\text{g})] = -228.6 \text{ kJ/mol}$
- A. 572.2 kJ/mol·rxn
  - B. -673.7 kJ/mol·rxn
  - C. -572.2 kJ/mol·rxn
  - D. -436.4 kJ/mol·rxn
  - E. -800.8 kJ/mol·rxn

MAR

Given the following information, calculate  $\Delta G^\circ$  for the reaction below at 25 °C:



- $\Delta H^\circ = -601.24 \text{ kJ/mol} \cdot \text{rxn}$
  - $\Delta S^\circ = -108.36 \text{ J/K} \cdot \text{rxn}$
- A. 664.5 kJ/mol·rxn
  - B. -568.9 kJ/mol·rxn
  - C. 31700 kJ/mol·rxn
  - D. -528.3 kJ/mol·rxn
  - E. 42

MAR

A reaction has a  $\Delta H^\circ$  which is **positive** and a  $\Delta S^\circ$  which is **positive**. What can be said about the reaction spontaneity at different temperatures?

- A. product favored at all temperatures
- B. product favored only at high temperature
- C. product favored only at low temperature
- D. not product favored at any temperature

MAR

Calculate  $\Delta G^\circ$  at 25 °C for:  $2 \text{H}_2\text{O}_2(\text{l}) \rightarrow 2 \text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$

	$\Delta H^\circ(\text{kJ/mol})$	$S^\circ(\text{J/K} \cdot \text{mol})$
$\text{H}_2\text{O}_2(\text{l})$	-187.8	109.6
$\text{H}_2\text{O}(\text{l})$	-285.8	69.9
$\text{O}_2(\text{g})$	-----	205.1

- A. -157.9 kJ/mol•rxn
- B. -192.3 kJ/mol•rxn
- C. -37700 kJ/mol•rxn
- D. -233.5 kJ/mol•rxn
- E. 42

MAR

A reaction has  $\Delta H = -96.0 \text{ kJ/mol}$  and  $\Delta S = -12.6 \text{ J/K} \cdot \text{mol}$ . If the temperature is increased slowly, at what temperature will this reaction become nonspontaneous?

- A. It will never be spontaneous
- B. 7162 °C
- C. 762 K
- D. 7620 K
- E. -57.6 K

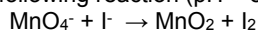
MAR

Barium sulfite is poorly soluble in water with a  $K_{\text{sp}}$  value of  $8.0 \times 10^{-7}$ . What is  $\Delta G^\circ$  at 25 °C?

- A. 15.1 kJ/mol•rxn
- B. 34.8 kJ/mol•rxn
- C. -34.8 kJ/mol•rxn
- D. 343 kJ/mol•rxn
- E. 42

MAR

Balance the following reaction (pH = 8.37):



- A.  $\text{MnO}_4^- + 2 \text{I}^- \rightarrow \text{MnO}_2 + \text{I}_2 + \text{O}_2^{2-}$
- B.  $8 \text{H}^+ + 2 \text{MnO}_4^- + 6 \text{I}^- \rightarrow 2 \text{MnO}_2 + 3 \text{I}_2 + 4 \text{H}_2\text{O}$
- C.  $8 \text{H}_2\text{O} + 4 \text{MnO}_4^- + 12 \text{I}^- \rightarrow 4 \text{MnO}_2 + 6 \text{I}_2 + 16 \text{OH}^-$
- D.  $4 \text{H}_2\text{O} + 2 \text{MnO}_4^- + 6 \text{I}^- \rightarrow 2 \text{MnO}_2 + 3 \text{I}_2 + 8 \text{OH}^-$

MAR

What is the strongest reducing agent in the list?

Half-Reaction	$E^\circ(\text{V})$
$\text{Ce}^{4+}(\text{aq}) + \text{e}^- \rightarrow \text{Ce}^{3+}(\text{aq})$	+1.61
$\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$	+0.80
$\text{Hg}_2^{2+}(\text{aq}) + 2 \text{e}^- \rightarrow 2 \text{Hg}(\text{l})$	+0.79
$\text{Sn}^{2+}(\text{aq}) + 2 \text{e}^- \rightarrow \text{Sn}(\text{s})$	-0.14
$\text{Ni}^{2+}(\text{aq}) + 2 \text{e}^- \rightarrow \text{Ni}(\text{s})$	-0.25
$\text{Al}^{3+}(\text{aq}) + 3 \text{e}^- \rightarrow \text{Al}(\text{s})$	-1.66

- A.  $\text{Ce}^{4+}$
- B.  $\text{Al}^{3+}$
- C. Sn
- D. Al
- E. Jq

MAR

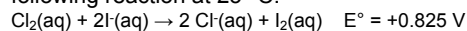
Will Sn(s) reduce Ag<sup>+</sup>(aq) to Ag(s)?

Half-Reaction	E°(V)
Ag <sup>+</sup> (aq) + e <sup>-</sup> → Ag(s)	+0.80
Sn <sup>2+</sup> (aq) + 2 e <sup>-</sup> → Sn(s)	-0.14

- A. Yes  
 B. No  
 C. Only if it feels like it

MAR

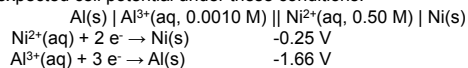
Determine the equilibrium constant for the following reaction at 25 °C:



- A.  $1.31 \times 10^{-28}$   
 B.  $8.74 \times 10^{13}$   
 C.  $8.03 \times 10^{27}$   
 D. 0.217  
 E.  $-1.16 \times 10^5$

MAR

A voltaic cell is created using the information below to be used in Alaska where the average temperature is 5.00 °C. Calculate the expected cell potential under these conditions.



- A. 1.46 V  
 B. 1.31 V  
 C. 1.17 V  
 D. 0.51 V  
 E. -1.91 V

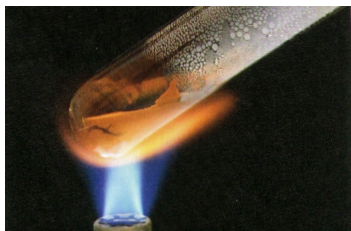
MAR

How long must a 2.00 amp current flow through a gold solution to convert 0.0100 mol of Au<sup>3+</sup>(aq) into Au(s)?

- A. 483 s  
 B.  $4.83 \times 10^4$  s  
 C. 965 s  
 D. 1450 s  
 E. 1 zillion s

MAR

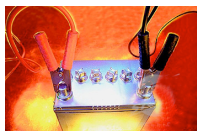
**End of  
 Review -  
 good luck  
 with your  
 studying!**



Need more practice?

- Practice Problem Sets (online)
- Concept Guides (Companion and online)
- Chapter Guides (online)
- End of Chapter Problems in Textbook (every other question has answer at end)

Good luck with your studying!



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