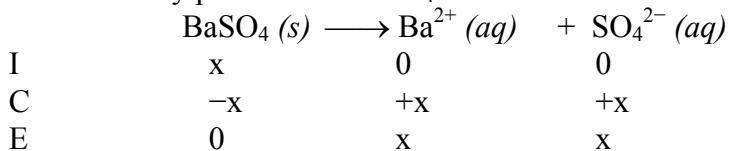


1. Which one of the following compounds has the lowest solubility in water, expressed in moles per liter? **Lowest solubility would mean the one with the smallest K_{sp}. Remember that K_{sp} refers to the dissociation of a solid into its ions.**

A. SrF₂, K_{sp} = 2.8 x 10⁻⁹ **This is the smallest K_{sp} in the list.**

2. The solubility product for BaSO₄ is 1.1 x 10⁻¹⁰. Calculate the solubility of BaSO₄ in pure water.



$$K_{sp} = [\text{Ba}^{2+}] [\text{SO}_4^{2-}] = x^2 = 1.1 \times 10^{-10}$$

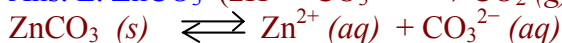
$$x = \sqrt{1.1 \times 10^{-10}} = 1.0 \times 10^{-5} \quad \text{Ans. B. } 1.0 \times 10^{-5} \text{ M}$$

3. Which of the following is least soluble in water according to solubility rules?

E. K₂CO₃ Remember to review your solubility rules (handout) for your Final Exam!

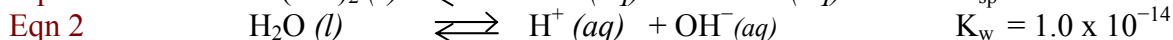
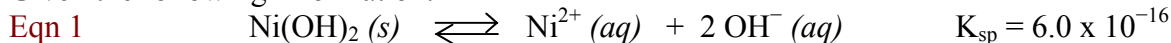
4. Which of the following solids would be more soluble in a strong acid solution than in pure water?

Ans. E. ZnCO₃ (2H⁺ + CO₃²⁻ → CO₂ (g) + H₂O (l)).



The acid reacting with the carbonate lowers the concentration of CO₃²⁻ and shifts the equilibrium to the right, thus making more of the solid ZnCO₃ dissolve.

5. Given the following information:

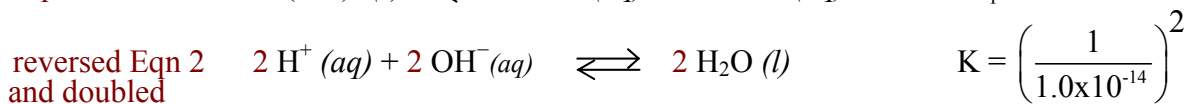
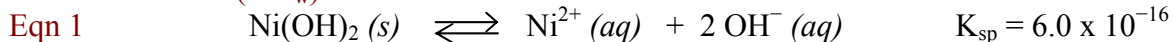


What is the equilibrium constant for the reaction shown below?



In order to get the desired Eqn 3, we keep Eqn 1 as is, but Eqn 2 must be reversed AND DOUBLED. Note that we need two OH⁻ in Eqn 2 to cancel with 2 OH⁻ in Eqn 1. You should ALWAYS check to make sure things cancel out properly.

K would become (1/K_w)².



C. K_{sp}/K_w²

6. The mathematical equation which expresses the first law of thermodynamics is

D. ΔE = q + w

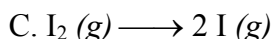
7. You are preparing to turn in homework and you staple your papers together (with the stapler that Lori lent you). Would this process give a positive or negative ΔS?

Stapler papers together involve more order. Remember that S is a measure of disorder.

S_{final} is lower than S_{initial} and ΔS = S_{final} - S_{initial}.

B. negative

8. Which reaction is accompanied by an increase in entropy? (Remember that means more disorder.)



9. For a certain endothermic chemical reaction where a gaseous reactant is converted into a solid product, we can conclude that...

ΔH would be + and ΔS would be -.

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

+ - -

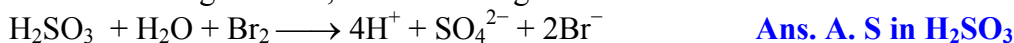
ΔG would be positive regardless of the temperature.

B. the reaction must not be spontaneous regardless of temperature.

10. What is the oxidation number of C in $\text{Na}_2\text{C}_2\text{O}_4$?

It is important to remember that oxidation number is defined as the charge per atom. C_2 is +6 so oxidation number of C is +3. **Ans. D. +3**

11. In the following reaction, which is being oxidized?



When there are other elements in the reaction, generally, H and O are not undergoing redox.

Check the other elements first! You should have seen that obviously $\text{Br}_2 \rightarrow \text{Br}^-$ is changing oxidation numbers. Also, when you see " SO_3 " changing to " SO_4 ", you know the S is undergoing oxidation (it gained an O).

In this rxn, S in H_2SO_3 is +4 and it became +6 in SO_4^{2-} . It is oxidized.

12. If you want to protect Zn, which of the following metal would you use as the "sacrificial" metal?

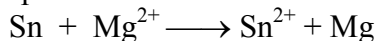
You would need a more active metal. That would be a metal below Zn in the reduction potential table.

Ans. D. Mg

$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	$E^\circ = +0.34 \text{ V}$
$\text{Sn}^{2+} + 2\text{e}^- \rightarrow \text{Sn}$	$E^\circ = -0.14 \text{ V}$
$\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe}$	$E^\circ = -0.44 \text{ V}$
$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$	$E^\circ = -0.76 \text{ V}$
$\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$	$E^\circ = -2.37 \text{ V}$

13. Examine the table of standard reduction potential above. Which metal is the most active metal? In a reduction potential table, the most active metal is the one with the lowest E° . (Remember that metals are oxidized, not reduced. The most active metal is the one that is most easily oxidized, and that would be the least readily reduced (lowest in the list). **Ans. E. Mg**

14. Examine the table of standard reduction potential above. Is the reaction expected to be spontaneous?



Break the eqn into two half-rxns:



$$E^\circ_{\text{cell}} = +0.14\text{V} - 2.37 \text{ V} = -2.23 \text{ V} \quad \text{Ans. B. It is not expected to be spontaneous.}$$

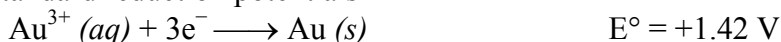
15. In the rusting process of the metal body of a car by the ocean, what is the reason that the presence of salt water would accelerate this process? **Ans. D. The salt water is acting as a salt bridge.**

15. Which statement is correct?

Ans. B. In an electrolytic cell, the electric current is flowing from the anode to the cathode.

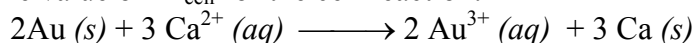
The electric current ALWAYS flow from the anode to the cathode, regardless of whether it is an electrolytic or galvanic cell. (Electric current refers to the electricity traveling through the wire.)

17. Using the standard reduction potentials





Calculate the value of E°_{cell} for the cell reaction:



Au to Au^{3+} has $E^{\circ} = -1.42\text{V}$.

Ca^{2+} to Ca has $E^{\circ} = -2.76\text{V}$

$E^{\circ}_{\text{cell}} = -1.42 \text{ V} - 2.76 \text{ V} = -4.18 \text{ V}$. Note that you don't double or triple E° even though you have to do that with the equations to balance the number of electrons. **Ans. C. -4.18 V**

18. Which statement is true of the galvanic cell with KNO_3 in the salt bridge?

Remember that the cation always move towards the cathode (anions to the anode). So K^{+} is expected to move towards the cathode, and the cathode is where reduction is taking place. (Red Cat). **Ans. B. K^{+} will move towards the electrode where reduction is taking place.**

19. For the following galvanic cell, which is the cathode?



The convention is this Line Cell Notation is the anode is always on the left and the cathode on the right.

Thus, Cu must be the cathode. **Ans. D. Cu**

20. In a galvanic cell shown in Question 19 above, concentration of which ion is building up at the anode for which the KNO_3 in the salt bridge must compensate?

One way to think about this is that the anion (NO_3^{-}) moves towards the anode. This means a positive charge is building up at that electrode. The anode is on the left side in the Line Cell Notation, so the positive charge must be due to the buildup of Cr^{3+} .

Another way is to think about what reaction is taking place at the anode. In the Line Cell Notation, it shows $\text{Cr} \rightarrow \text{Cr}^{3+}$. Chromium is converted to Cr^{3+} at the anode, so it must be Cr^{3+} .

Ans.A. Cr^{3+}

21. During the electrolysis of water, which gas is formed at the anode?

Recall that oxidation is at the anode. Which is at the anode?

$\text{H}_2\text{O} \rightarrow \text{O}_2$ or $\text{H}_2\text{O} \rightarrow \text{H}_2$? **Ans. A. O_2** (O in water is -2 and zero in O_2 . Oxid. # is increasing.)

21. (5 pts) Will a precipitate form when 20.0 mL of $1.8 \times 10^{-3} \text{ M}$ $\text{Pb}(\text{NO}_3)_2$ is added to 30.0 mL of $5.0 \times 10^{-4} \text{ M}$ Na_2SO_4 ? The K_{sp} of PbSO_4 is 6.3×10^{-7} . Show your work and give a brief explanation.

First you should recognize that in mixing the solutions, the concentrations of both reactants get diluted. Total volume = 20.0 mL + 30.0 mL = 0.0500 L

$$[\text{Pb}^{2+}] = \frac{0.0200 \text{ L} \times 1.8 \times 10^{-3} \text{ mol/L}}{0.0500 \text{ L}} = 7.2 \times 10^{-4} \text{ M Pb}^{2+}$$

$$[\text{SO}_4^{2-}] = \frac{0.0300 \text{ L} \times 5.0 \times 10^{-4} \text{ mol/L}}{0.0500 \text{ L}} = 3.0 \times 10^{-4} \text{ M SO}_4^{2-}$$

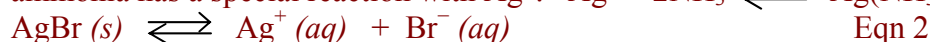
$$Q = (7.2 \times 10^{-4})(3.0 \times 10^{-4}) = 2.2 \times 10^{-7} < K_{\text{sp}} (6.3 \times 10^{-7})$$

This means the concentrations of ions are below the saturation point.

Ans. No, a precipitate will not form.

22. (3 pts) Explain why bubbling ammonia gas through a mixture of AgBr (s) in water would make it more soluble. Formation of what **complex ion** will affect its solubility? Include chemical equations in your explanation, but answer in full sentences.

Note that the question is asking about formation of a **complex ion**. You should recall that



Formation of the complex ion Ag(NH₃)₂⁺ in Eqn 1 decreases the concentration of Ag⁺ in Eqn 2. To relieve the stress, the equilibrium in Eqn 2 will shift to the right to compensate, thus making more of the solid AgBr dissolve.

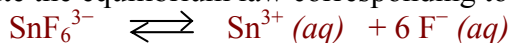
23. (6 pts) Write the chemical equation for the complex ion formation of SnF₆³⁻.



If the K_{form} of SnF₆³⁻ is 1 x 10²⁵, what is its K_{inst}?

K_{inst} is for the reverse of formation and so K_{inst} = 1/K_{form} Ans. 1/(1x10²⁵) = **1x10⁻²⁵**.

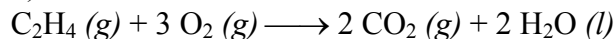
Write the equilibrium law corresponding to the K_{inst}.



The equilibrium law is...

$$K = \frac{[\text{Sn}^{3+}][\text{F}^-]^6}{[\text{SnF}_6^{3-}]}$$

24. (10 pts) Calculate ΔG° for the reaction below. **Show setups carefully.**



ΔH_f°(kJ/mol)

+51.9 0 -394 -286.0

ΔS°(J·mol⁻¹·K⁻¹)

219.8 205.0 213.6 69.96

$\Delta H^\circ = -2\text{mol}(394\text{kJ/mol}) - 2\text{mol}(286.0\text{kJ/mol}) - 1\text{mol}(51.9\text{kJ/mol}) = -1411.9 \text{ kJ}$

$\Delta S^\circ = 2\text{mol} (213.6 \text{ J}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}) + 2\text{mol}(269.96 \text{ J}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}) - 1\text{mol}(219.8 \text{ J}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}) - 3\text{mol}(205.0 \text{ J}\cdot\text{mol}^{-1}\cdot\text{K}^{-1})$

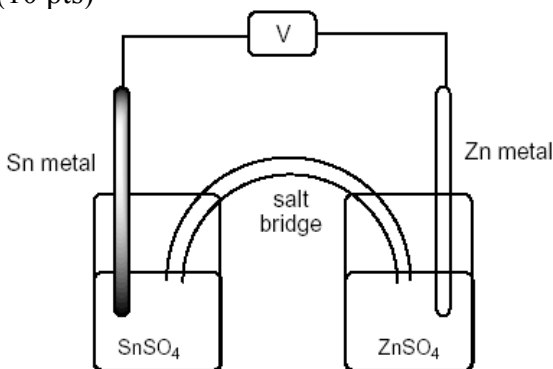
$= -267.6 \text{ J/K}$

$\Delta G^\circ = \Delta H - T\Delta S = -1411.9 \text{ kJ} - (298\text{K})(-267.6 \text{ J/K})(1 \text{ kJ}/10^3\text{J})$

$= -1411.9 \text{ kJ} + 79.7 \text{ kJ}$

$= -1332.1 \text{ kJ} \quad \text{Ans. } -1332 \text{ kJ}$

25. (10 pts)



$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	$E^\circ = +0.34 \text{ V}$
$\text{Sn}^{2+} + 2\text{e}^- \rightarrow \text{Sn}$	$E^\circ = -0.14 \text{ V}$
$\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe}$	$E^\circ = -0.44 \text{ V}$
$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$	$E^\circ = -0.76 \text{ V}$
$\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$	$E^\circ = -2.37 \text{ V}$

The decision is on whether it is $\text{Sn} + \text{Zn}^{2+} \rightarrow \text{Sn}^{2+} + \text{Zn}$
 or $\text{Zn} + \text{Sn}^{2+} \rightarrow \text{Zn}^{2+} + \text{Sn}$

The E°_{cell} for the first choice would be negative ($+0.14\text{V} - 0.76\text{V} = -0.62\text{V}$).

The E°_{cell} for the second choice would therefore be positive (reverse the sign): $+0.62\text{V}$.

This means the reaction is $\text{Zn} + \text{Sn}^{2+} \rightarrow \text{Zn}^{2+} + \text{Sn}$

and the half-rxns are $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ (oxidation, at the anode) at the right cell

and $\text{Sn}^{2+} + 2\text{e}^- \rightarrow \text{Sn}$ (reduction, at the cathode) at the left cell.

- (a) Using the standard reduction potentials provided, determine whether the anode is on the left or on the right in the diagram above.

The above analysis tells us that the anode is the Zn and in the diagram, Zn is on the right.

Circle one: The anode is on the right

- (b) What direction will the electric current be flowing? It always flows from anode to cathode.

We have already established the anode is on the right, cathode on the left, so the electric current is flowing from the right to the left.

- (c) Which electrode do you expect to get bigger as the reaction proceeds?

$\text{Zn} + \text{Sn}^{2+} \rightarrow \text{Zn}^{2+} + \text{Sn}$ We see that Sn is being made, so the answer is...
Sn electrode

- (d) If the salt in the salt bridge is KCl, which direction will the Cl^- be moving?

Anions move towards the anode. So Cl^- ions will move towards the right.

Ans. into the right compartment

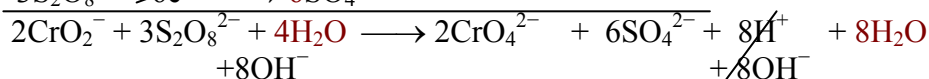
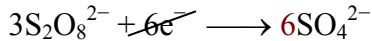
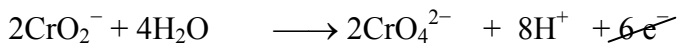
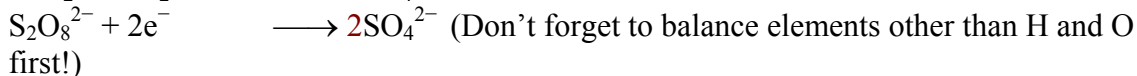
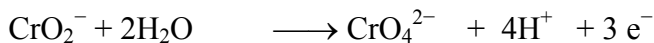
- (e) Write the standard cell notation (the Line Cell Notation) for this cell.

$\text{Zn} (s) | \text{Zn}^{2+} (aq) || \text{Sn}^{2+} (aq) | \text{Sn}$

Remember in the Line Cell Notation, the anode is ALWAYS on the left.

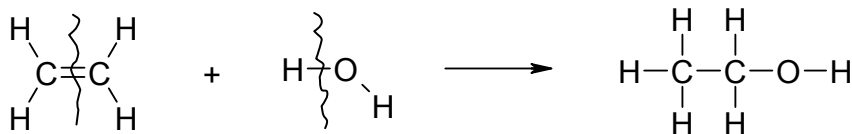
26. (6 pts) Balance the following reactions which is under BASIC conditions.

Show your work carefully.



Ans. $2\text{CrO}_2^- + 3\text{S}_2\text{O}_8^{2-} + 8\text{OH}^- \rightarrow 2\text{CrO}_4^{2-} + 6\text{SO}_4^{2-} + 4\text{H}_2\text{O}$

27. (6 pts) Using the bond energies provided, calculate ΔH° of reaction for



	Bond Energy (kJ mol ⁻¹)
C-C	348
C=C	612
C≡C	961
C-O	362
C=O	743
C-H	412
H-H	436
H-O	463
O=O	498

Show your work carefully. Do not break any bonds that you don't need to break! Watch your significant figures.

Break one C=C +612 kJ
 one H-O +463 kJ

Form one C-C -348 kJ
 one C-O -362 kJ
 one H-C -412 kJ

Total = (+612 + 463 -348 -362 -412)kJ = **-47 kJ**

28. (8 pts) For the system, $2 \text{NO}_2 (g) \rightleftharpoons \text{N}_2\text{O}_4 (g)$, $\Delta G^\circ = -5.40 \text{ kJ}$.

(a) Calculate the value of the equilibrium constant, K_p for this system. Show your work and watch your sig. fig.

$$\Delta G^\circ = -RT \ln K_p$$

Ans. $K_p = 8.8$

$$\ln K_p = \frac{\Delta G^\circ}{-RT} = \frac{-5.40 \text{ kJ}}{-(8.314 \text{ J mol}^{-1}\text{K}^{-1})(298\text{K})} \left(\frac{10^3 \text{ J}}{1 \text{ kJ}} \right) = 2.179$$

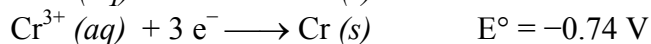
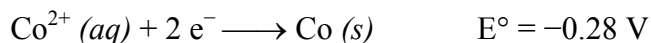
$$K_p = e^{2.179} = 8.8$$

(b) Give the equilibrium expression corresponding to this K_p .

Remember that K_p is the equilibrium constant written in terms of partial pressures of the gases.

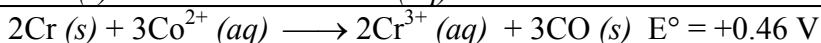
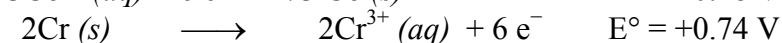
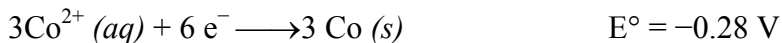
$$K_p = \frac{P_{\text{N}_2\text{O}_4}}{(P_{\text{NO}_2})^2}$$

29. (6 pts) A galvanic cell is composed of these two half-cells, with the standard reduction potentials shown:



The actual concentrations in the cell are: $[\text{Co}^{2+}] = 0.00100 \text{ M}$, $[\text{Cr}^{3+}] = 0.100 \text{ M}$. What is the potential of this galvanic cell?

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{nF} \ln Q, \text{ where } Q = \frac{[\text{Cr}^{3+}]^2}{[\text{Co}^{2+}]^3} = \frac{(0.100)^2}{(0.00100)^3} = 1.00 \times 10^7 \text{ and } \ln Q = 16.118$$



$$E_{\text{cell}} = +0.46 \text{ V} - \frac{(8.314 \text{ J mol}^{-1}\text{K}^{-1})(298\text{K})}{(6 \text{ mol})(9.65 \times 10^4 \text{ C mol}^{-1})} \left(\frac{1 \text{ V}}{1 \text{ J C}^{-1}} \right) (16.118)$$

$$E_{\text{cell}} = +0.46 \text{ V} - 0.06896 \text{ V} = \mathbf{+0.39 \text{ V}}$$

Bonus Points

(1pt) Make sure you have your name on every page, front and back.

(2 pts) Which are the two elements that are the most powerful oxidizing agents? F_2 and O_2 .

(1 pts) Are metals more likely to be reduced or oxidized? Ans. **oxidized**.

(2 pts) Briefly explain how water can be oxidized as well as reduced.

The H in water can undergo reduction as it converts to H_2 .

(Oxidation number changes from +1 to zero.)

The O in water can undergo oxidation as it converts to O_2 .

(Oxidation number changes from -2 to zero.)