

## 2011 Test 2 Version 1 – Solutions

1. The density of a **noble gas** is  $2.71 \text{ g L}^{-1}$  at  $3.00 \text{ atm}$  and  $0^\circ\text{C}$ . What is the gas?

- A) Xe  
**B) Ne**  
 C) He  
 D) Kr  
 E) Ar

The number of moles in  $1.00 \text{ L}$  of this gas is

$$n = \frac{PV}{RT} = \frac{3.00 \text{ atm} \times 1.00 \text{ L}}{0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 273.15 \text{ K}} = 0.133_8 \text{ mol}$$

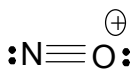
$1.00 \text{ L}$  of gas has mass,  $m = 2.71 \text{ g L}^{-1} \times 1.00 \text{ L} = 2.71 \text{ g}$ .

Therefore, the molar mass is  $M = 2.71 \text{ g} / 0.1338 \text{ mol} = 20.2 \text{ g mol}^{-1}$ .

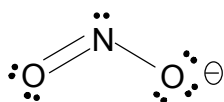
The gas is clearly neon, Ne.

2. For the species  $\text{NO}^+$ ,  $\text{NO}_2^-$ ,  $\text{NO}_3^-$ , what is the correct order of decreasing N-O **bond length**?

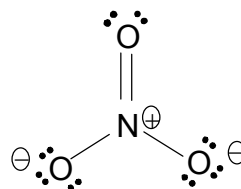
- A)  $\text{NO}_2^- > \text{NO}_3^- > \text{NO}^+$   
 B)  $\text{NO}_2^- > \text{NO}^+ > \text{NO}_3^-$   
 C)  $\text{NO}^+ > \text{NO}_3^- > \text{NO}_2^-$   
**D)  $\text{NO}_3^- > \text{NO}_2^- > \text{NO}^+$**   
 E)  $\text{NO}_3^- > \text{NO}^+ > \text{NO}_2^-$



ave bond order = 3



ave bond order =  $3/2$



ave bond order =  $4/3$

Bonds of higher order are shorter. Thus,  $\text{NO}^+$  has the shortest bond (its order is 3), while  $\text{NO}_2^-$  has the next shortest bonds (bond order = 1.5).  $\text{NO}_3^-$  has the longest bonds (bond order = 1.33).

3. Of the series: Al, Ba, Ca, Mg, Si, which atom has the highest **second** ionization energy:

- A) Ca
- B) Mg
- C) Ba
- D) Si
- E) Al

Second ionization energy is the energy required to ionize the +1 cation. The electron configuration of a +1 cation of the element with atomic number Z is the same as that of a neutral atom with atomic number Z-1. The ordering of the second ionization energies of Al, Ba, Ca, Mg and Si is thus the same as the ordering of the first ionization energies of Mg, Cs, K, Na and Al, respectively. According to first ionization, these atoms are ordered as follows:  $\text{Cs} < \text{K} < \text{Na} < \text{Al} < \text{Mg}$ . The first three elements are ascending group 1. Mg and Al are to the right of Na in period 3. Mg has the higher first ionization because the filled subshell effect. The electron ejected from Mg is from the 3s subshell rather than the 3p subshell as it is for Al. Since Mg has the same electron configuration as  $\text{Al}^+$ , it is Al that has the highest second ionization energy.

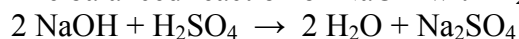
4. When 25.0 mL of a 0.610 M NaOH solution reacts with 20.0 mL of a 0.245 M  $\text{H}_2\text{SO}_4$  solution,

- A) the limiting reagent is  $\text{H}_2\text{SO}_4$  and the resulting solution is acidic.
- B) the resulting solution is pH neutral.
- C) the limiting reagent is NaOH and the resulting solution is basic.
- D) the limiting reagent is NaOH and the resulting solution is acidic.
- E) the limiting reagent is  $\text{H}_2\text{SO}_4$  and the resulting solution is basic

$$\text{Moles of NaOH} = 0.0250 \text{ L} \times 0.610 \text{ mol L}^{-1} = 0.01525 \text{ mol}$$

$$\text{Moles of H}_2\text{SO}_4 = 0.0200 \text{ mL} \times 0.245 \text{ mol L}^{-1} = 0.00490 \text{ mol}$$

The balanced reaction of NaOH with  $\text{H}_2\text{SO}_4$  is



Thus, 0.00490 mol  $\text{H}_2\text{SO}_4$  consumes 0.00980 mol NaOH. NaOH is thus in excess.  $\text{H}_2\text{SO}_4$  is the limiting reactant and the resulting solution is basic because of the excess NaOH.

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5. A hydrogen atom in the  $n = 5$  state emits a photon with a frequency of  $3.1558 \times 10^{15}$  Hz. What would be the **final state** of this atom?

- A) 7  
B) 3  
C) 1  
D) 2  
E) 6

The emitted photon carries the energy that is lost by the H atom.

$$h\nu = R_H \left( \frac{1}{n_f^2} - \frac{1}{5^2} \right)$$

So

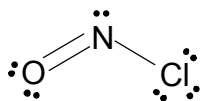
$$\begin{aligned} \frac{1}{n_f^2} &= \frac{1}{25} + \frac{h\nu}{R_H} \\ &= 0.04 + \frac{6.6256 \times 10^{-34} \text{ J s} \times 3.1558 \times 10^{15} \text{ s}^{-1}}{2.178 \times 10^{-18} \text{ J}} \\ &= 1.000 \end{aligned}$$

So,

$$n_f = \sqrt{\frac{1}{1.000}} = 1$$

The final state is  $n_f = 1$ .

6. Choose the **TRUE** statement concerning the Lewis structure of nitrosyl chloride, ONCl (N is the central atom):
- A) The molecule contains 5 nonbonding electron pairs.  
B) The molecule contains 4 bonding electron pairs.  
C) The molecule is linear.  
D) All formal charges can be minimized to zero.  
E) None of the above statements are true.



ONCl is bent with zero formal charges on all atoms. There are 6 nonbonding pairs of electrons and 3 bonding pairs.

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7. Select the one **FALSE** statement for the following equilibrium in a constant volume container:

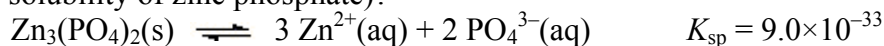


The concentration of  $\text{SO}_2(\text{g})$  will increase if:

- A) An inert gas is added to increase the total pressure.
- B) The reaction mixture is transferred to a larger volume container.
- C)  $\text{O}_2$  is removed.
- D)  $\text{SO}_3$  is added.
- E) The temperature is increased.

Transferring the reaction mixture to a larger volume decreases the partial pressures of all the gases by the same factor. Since there are 3 moles of reactant gas versus 2 moles of product gas, the denominator of the reaction quotient decreases more than the numerator. Therefore,  $Q$  increases away from the equilibrium constant. The reverse reaction is favored and increases the  $\text{SO}_2(\text{g})$  partial pressure (and concentration). Removing  $\text{O}_2$ , a reactant, also favors the reverse reaction. Adding  $\text{SO}_3$ , a product, again favors the reverse reaction. Increasing temperature favors the endothermic reaction – again, the reverse reaction. Adding an inert gas to the mixture does not affect the partial pressures of the reactants and products, and therefore has no effect on the equilibrium.

8. What is the **concentration** (in  $\text{mol L}^{-1}$ ) of a saturated solution of  $\text{Zn}_3(\text{PO}_4)_2$  (i.e. what is the molar solubility of zinc phosphate)?



- A)  $3.5 \times 10^{-6}$
- B)  $1.9 \times 10^{-8}$
- C)  $1.5 \times 10^{-7}$
- D)  $2.7 \times 10^{-9}$
- E)  $2.1 \times 10^{-5}$

	$\text{Zn}_3(\text{PO}_4)_2(\text{s}) \rightleftharpoons$	$3 \text{Zn}^{2+}(\text{aq}) +$	$2 \text{PO}_4^{3-}(\text{aq})$
Initial		0	0
Change		$3x$	$2x$
Equilibrium		$3x$	$2x$

At equilibrium,  $K_{\text{sp}} = [\text{Zn}^{2+}]^3 [\text{PO}_4^{3-}]^2 = (3x)^3 (2x)^2 = 108 x^5 = 9.0 \times 10^{-33}$

Therefore, the extent of reaction is

$$x = (9.0 \times 10^{-33} / 108)^{1/5} = 1.5 \times 10^{-7}$$

This is the molar solubility of  $\text{Zn}_3(\text{PO}_4)_2$  in  $\text{mol L}^{-1}$ .

Name: \_\_\_\_\_

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9. A vessel is filled with  $\text{N}_2\text{O}_4(\text{g})$  to an initial pressure of 3.01 atm. Some of this gas decomposes into  $\text{NO}_2(\text{g})$ . At equilibrium, the partial pressure of  $\text{N}_2\text{O}_4(\text{g})$  is found to be 2.71 atm. What is the value of the **equilibrium constant** for the decomposition of  $\text{N}_2\text{O}_4(\text{g})$  into  $\text{NO}_2(\text{g})$ , at the temperature of this experiment?

- A) 1.07  
 B) 0.277  
 C) 1.49  
**D) 0.133**  
 E) 3.41

	$\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons$	$2 \text{NO}_2(\text{g})$
Initial	3.01	0
Change	$-x$	$2x$
Equilibrium	$3.01 - x$	$2x$

At equilibrium,  $P_{\text{N}_2\text{O}_4} = 2.71 \text{ atm} = 3.01 - x \text{ atm}$ .

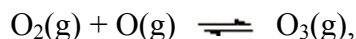
Therefore,  $x = 3.01 - 2.71 = 0.30$

and  $P_{\text{NO}_2} = 2x \text{ atm} = 0.60 \text{ atm}$ .

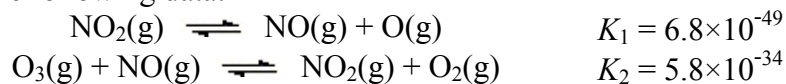
The equilibrium constant is

$$K = \frac{P_{\text{NO}_2}^2}{P_{\text{N}_2\text{O}_4}} = \frac{0.60^2}{2.71} = 0.133$$

10. Determine the **equilibrium constant** for the formation of ozone,



from the following data:



- A)  $3.7 \times 10^{-82}$   
 B)  $5.6 \times 10^{83}$   
 C)  $8.1 \times 10^{-81}$   
**D)  $2.5 \times 10^{81}$**   
 E)  $1.9 \times 10^{82}$

The target reaction can be written as  $-(\text{Rxn 1}) - (\text{Rxn 2})$ . Therefore,

$$K = K_1^{-1} K_2^{-1} = (6.8 \times 10^{-49})^{-1} (5.8 \times 10^{-34})^{-1} = 2.5 \times 10^{81}$$

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11. At 125°C,  $K = 0.25$  for the decomposition of sodium hydrogen carbonate,



If 10.0 g of  $\text{NaHCO}_3$  is added to a 1.00 L flask which is then heated to 125°C, what is the **partial pressure** of  $\text{CO}_2$  in the flask when the reaction comes to equilibrium?

- A) 0.75 atm  
 B) 1.25 atm  
 C) 0.25 atm  
 D) 1.0 atm  
 E) 0.50 atm

Initially, the partial pressure of the gases is zero.

	$2 \text{NaHCO}_3(\text{s}) \rightleftharpoons$	$\text{Na}_2\text{CO}_3(\text{s}) +$	$\text{CO}_2(\text{g}) +$	$\text{H}_2\text{O}(\text{g})$
Initial			0	0
Change			$x$	$x$
Equilibrium			$x$	$x$

At equilibrium,  $K = x^2 = 0.25$ .

Therefore,  $x = 0.50$  and  $P_{\text{CO}_2} = 0.50 \text{ atm}$ .

12. Which one of the following equilibria is **unaffected** when the volume of the reaction vessel is doubled?

- A)  $\text{CO}_2(\text{s}) \rightleftharpoons \text{CO}_2(\text{g})$   
 B)  $\text{Br}_2(\text{g}) \rightleftharpoons 2 \text{Br}(\text{g})$   
 C)  $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{NO}(\text{g})$   
 D)  $2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{SO}_3(\text{g})$   
 E)  $2 \text{NO}_2(\text{g}) + 7 \text{H}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g}) + 4 \text{H}_2\text{O}(\text{l})$

The only equilibrium that is unaffected is the one with equal numbers of moles of reactant and product gases (ignore solids and liquids) – i.e.  $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{NO}(\text{g})$ . Doubling the volume decreases all partial pressures (they are  $\frac{1}{2}$  as big). If there are more product gases,  $Q$  decreases and the forward reaction is favored. If there are more reactant gases,  $Q$  increases and the reverse reaction is favored.

13. Aqueous solutions of aluminum nitrate and sodium hydroxide are mixed to give a white precipitate of aluminum hydroxide and a solution of sodium nitrate. Which of the following equations represents a **balanced chemical equation** for this reaction?

- A)  $\text{Al}(\text{NO}_3)_2(\text{aq}) + 2 \text{NaOH}(\text{aq}) \rightarrow 2 \text{NaNO}_3(\text{aq}) + \text{Al}(\text{OH})_2(\text{s})$   
 B)  $3 \text{Al}(\text{NO}_3)_3(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{NaNO}_3(\text{aq}) + 3 \text{Al}(\text{OH})_3(\text{s})$   
 C)  $\text{Al}(\text{NO}_3)_3(\text{aq}) + 3 \text{NaOH}(\text{aq}) \rightarrow 3 \text{NaNO}_3(\text{aq}) + \text{Al}(\text{OH})_3(\text{s})$   
 D)  $\text{Al}(\text{NO}_2)_3(\text{aq}) + 2 \text{NaOH}(\text{aq}) \rightarrow 2 \text{NaNO}_2(\text{aq}) + \text{Al}(\text{OH})_3(\text{s})$   
 E)  $\text{Al}(\text{NO}_3)_2(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow 2 \text{NaNO}_3(\text{aq}) + \text{Al}(\text{OH})_2(\text{s})$

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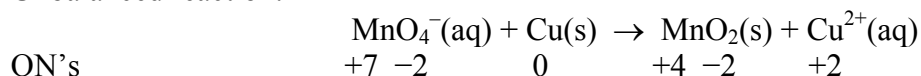
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Since Al has a valence of three, aluminum nitrate and aluminum hydroxide must be  $\text{Al}(\text{NO}_3)_3$  and  $\text{Al}(\text{OH})_3$ , respectively. Therefore, three moles of  $\text{NaOH}(\text{aq})$  are required to react with one mole of  $\text{Al}(\text{NO}_3)_3(\text{aq})$  to give one mole of  $\text{Al}(\text{OH})_3(\text{s})$  and three moles of  $\text{NaNO}_3(\text{aq})$  (the spectator ions).

14. Permanganate ions,  $\text{MnO}_4^-(\text{aq})$ , react with copper metal in basic solution to produce  $\text{MnO}_2(\text{s})$  and  $\text{Cu}^{2+}(\text{aq})$  ions. When the reaction is balanced, such that the smallest possible integers appear as stoichiometric coefficients, the coefficient for  $\text{Cu}^{2+}$  is:

- A) 4  
B) 3  
C) 1  
D) 6  
E) 2

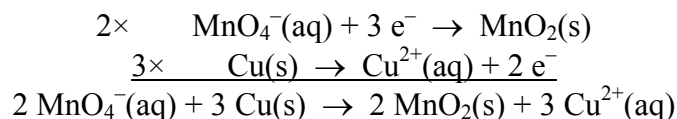
Unbalanced reaction:



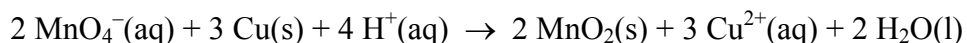
Unbalanced half reactions:



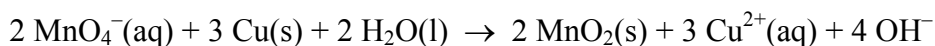
Balance the electrons:



Balance O's (add 2  $\text{H}_2\text{O}$  to the product side) and H's (add 4  $\text{H}^+$  to the reactant side) as though in acid:



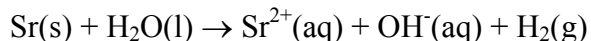
Neutralize the  $\text{H}^+$  (add 4  $\text{OH}^-$  to both sides – two waters cancel from both sides):



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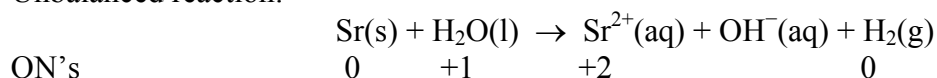
15. Which of the following statement(s) is(are) **FALSE** concerning the following reaction when it is balanced?



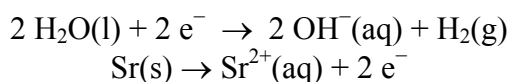
- (i) The indicator phenolphthalein turns pink as the reaction proceeds.
- (ii) Water acts as an oxidizing agent.
- (iii) Strontium gets oxidized.
- (iv) Two moles of  $\text{H}_2(\text{g})$  are produced for each mole of Sr consumed. **F**
- (v) Electrons are transferred to strontium. **F**
- (vi) Two moles of  $\text{H}_2\text{O(l)}$  react for each mole of Sr consumed.

- A) ii, iii
- B) iv, v**
- C) i, ii
- D) iii, vi
- E) iv

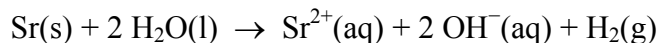
Unbalanced reaction:



Half reactions:



Balanced reaction:



Phenolphthalein turns pink due to the  $\text{OH}^-$  produced. Water is reduced - i.e. it is the oxidizing agent. Strontium gets oxidized. One mole of  $\text{H}_2(\text{g})$  is produced for each mole of Sr consumed. Electrons are transferred away from strontium. Two moles of  $\text{H}_2\text{O(l)}$  react for each mole of Sr consumed.



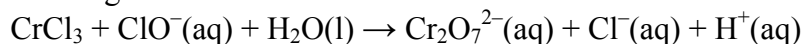
16. Which of the following statements are **TRUE**?

- (i) Alkali metal hydrides produce  $\text{H}_2(\text{g})$  upon reacting with water. **T**
- (ii) The formation of acidic solutions is a characteristic of the reaction of most metal oxides with water.
- (iii) Redox reactions between alkali metals (Group 1) and halogens (Group 17) produce ionic compounds which are generally soluble in water. **T**
- (iv) A transfer of  $\text{OH}^-$  is the definition of a Bronsted-Lowry acid-base reaction.
- (v) Mixing saturated aqueous solutions of phosphoric acid and calcium hydroxide produces no visible reaction.

- A) iv, v
- B) i, ii, iii
- C) i, iii, v
- D) i, iii**
- E) ii, iv

Alkali metal hydrides contain  $\text{H}^-$  ions.  $\text{H}^-$  is a strong base that removes  $\text{H}^+$  from water to form  $\text{H}_2$  gas. Most metal oxides form basic solutions when added to water. Alkali metal halides are generally soluble in water. A Bronsted-Lowry acid-base reaction is a  $\text{H}^+$  transfer reaction. Mixing saturated aqueous solutions of phosphoric acid and calcium hydroxide produces calcium phosphate precipitate.

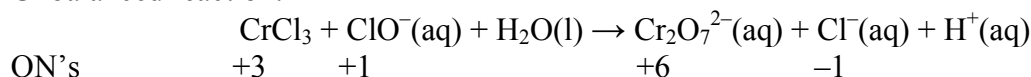
17. Balance the following redox reaction in **acidic** solution.



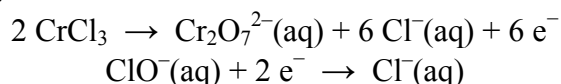
When this has been done correctly, the stoichiometric **coefficient** for  $\text{H}^+$  is:

- A) 8**
- B) 6
- C) 7
- D) 9
- E) 5

Unbalanced reaction:



Unbalanced half reactions:



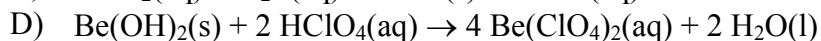
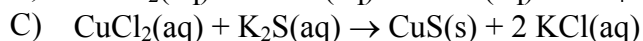
note there must be 2  $\text{CrCl}_3$  to balance the two Cr atoms in  $\text{Cr}_2\text{O}_7^{2-}$   
- this means 6 electrons are liberated

also, there are six spectator chloride ions in the oxidation half reaction

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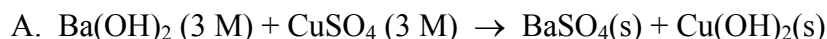
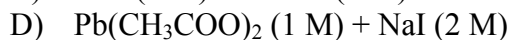
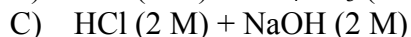
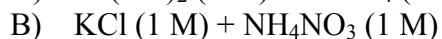
$$2 \text{CrCl}_3 + 3 \text{ClO}^-(\text{aq}) \rightarrow \text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 9 \text{Cl}^-(\text{aq})$$
$$2 \text{CrCl}_3 + 3 \text{ClO}^-(\text{aq}) + 4 \text{H}_2\text{O} \rightarrow \text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 9 \text{Cl}^-(\text{aq}) + 8 \text{H}^+(\text{aq})$$

A)  $\text{NH}_4\text{NO}_3(\text{s}) + \text{heat} \rightarrow \text{N}_2\text{O}(\text{g}) + 2 \text{H}_2\text{O}(\text{g})$

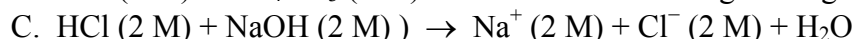

$$\text{NH}_4\text{NO}_3(\text{s}) + \text{heat} \rightarrow \text{N}_2\text{O}(\text{g}) + 2 \text{H}_2\text{O}(\text{g})$$
$$-3 \quad +5 \quad +1$$

Reactions B and D are acid-base reactions. Reaction C is a precipitation. Reaction E is a hydration reaction.

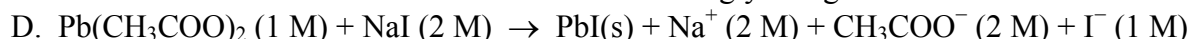
A)  $\text{Ba}(\text{OH})_2$  (3 M) +  $\text{CuSO}_4$  (3 M)



no ions here



4 M of singly charged ions



5 M of singly charged ions



20. Determine the **FALSE** statement regarding a concentration cell based on  $H^+$  (aq). Assume that the partial pressure of hydrogen is 1 atm in both half cells.

- A)  $Q = [\text{concentrated}] / [\text{dilute}]$   
 B) When  $[\text{dilute}] = [\text{concentrated}]$  the cell potential will become 0 V  
 C)  $E_{\text{cell}}$  is always greater than  $E^\circ_{\text{cell}}$  while the cell is operating.  
 D) Electrons flow from the dilute to the concentrated solution.  
 E) The concentrated solution is at the cathode

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{zF} \ln Q = -\frac{RT}{zF} \ln Q$$

in the case of a concentration cell (for which  $E^\circ_{\text{cell}} = 0$ ). A positive cell potential ( $E_{\text{cell}} > 0$ ) results when  $Q < 1$ . Thus,  $Q = [\text{dilute}] / [\text{concentrated}]$  and the cell reaction is

concentrated  $\rightarrow$  dilute

i.e. the spontaneous movement of  $H^+$ (aq) is against the concentration gradient (having the effect of reducing the concentration gradient).

$Q = 1$  and  $E_{\text{cell}} = 0$  when  $[\text{dilute}] = [\text{concentrated}]$ .

$H^+$  ions are produced at the anode (dilute solution) and consumed at the cathode (concentrated).

The electrons liberated at the anode flow through the wire to the cathode (i.e from dilute to concentrated).

21. Determine the **FALSE** statement regarding the following cell diagram:



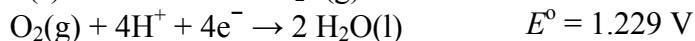
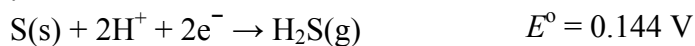
- A) Cu (s) is being oxidized.  
 B) || represents the salt bridge  
 C) At 25°C this is a standard electrochemical cell.  
 D) Zinc is the anode  
 E)  $NO_3^-$  is a spectator ion

Anode is on the left – i.e. it is zinc. Cathode is on the right – i.e. it is copper. Copper is reduced at the cathode – not oxidized. Since the ion concentrations are both 1.0 M, the cell is the standard cell at 25°C. Nitrate is a spectator ion – needed only for charge neutrality.

Name: \_\_\_\_\_

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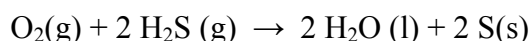
22. A student creates an electrochemical cell based on  $5.00 \times 10^{-3}$  mol of both  $O_2(g)$  and  $H_2S(g)$  in separate 2.0 L vessels outfitted with Pt (s) electrodes in acidic solution at 298 K. What is the **voltage** produced?



- A) 1.145 V  
 B) 0.984 V  
 C) 1.088 V  
**D) 1.031 V**  
 E) 1.002 V

$$E^\circ_{\text{cell}} = E^\circ[O_2(g)/H_2O(l)] - E^\circ[S(s)/H_2S(g)] = 1.229 - 0.144 \text{ V} = 1.085 \text{ V}$$

Cell reaction:



$$Q = \frac{1}{P_{O_2} P_{H_2S}^2} \quad \text{and} \quad z = 4$$

$$P_{O_2} = P_{H_2S} = \frac{nRT}{V} = \frac{5.00 \times 10^{-3} \text{ mol} \times 0.08206 \text{ L atm K mol}^{-1} \times 298 \text{ K}}{2.0 \text{ L}}$$

$$= 0.061_1 \text{ atm}$$

So,  $Q = 4.3_8 \times 10^3$  and

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0592 \text{ V}}{z} \log_{10} Q$$

$$= 1.085 \text{ V} - \frac{0.0592 \text{ V}}{4} \log_{10} 4.3_8 \times 10^3$$

$$= 1.031 \text{ V}$$

23. In the event of an uncontrollable fire in the lab, the **safest** course of action for a student is:

- A) Exit the laboratory and building, pulling the fire alarm while exiting the building**  
 B) Attempt to extinguish the fire  
 C) Collect all personal items and exit the building  
 D) Exit the lab and wait in the hallway until safe to return  
 E) Find the Laboratory Coordinator

Name: \_\_\_\_\_

Student number: \_\_\_\_\_

24. What **concentration** of 5.0 L of  $\text{Fe}(\text{NO}_3)_3$  would need to be added to 5.0 L of 1.0 M KSCN to produce an equilibrium concentration of  $\text{FeSCN}^{2+} = 0.40$  M. The equilibrium constant for the production of  $\text{FeSCN}^{2+} = 40$ .

- A) 5.0 M  
**B) 1.0 M**  
 C) 4.0 M  
 D) 2.0 M  
 E) 3.0 M

Note that when the 5.0 L of  $y$  M  $\text{Fe}(\text{NO}_3)_3$  solution is added to 5.0 L of 1.0 M KSCN solution both solutions are diluted by a factor of 2 (the total volume is 10.0 L).

	$\text{Fe}^{3+}(\text{aq}) +$	$\text{SCN}^-(\text{aq}) \rightleftharpoons$	$\text{FeSCN}^{2+}(\text{aq})$
Initial	$y/2$	0.50	0
Change	$-x$	$-x$	$x$
Equilibrium	$y/2 - x$	$0.50 - x$	$x = 0.40$

Since  $x = 0.40$ ,  $[\text{SCN}^-] = 0.10$  M and  $[\text{Fe}^{3+}] = y/2 - 0.40$  M.  
 At equilibrium,

$$K = \frac{[\text{FeSCN}^{2+}]}{[\text{Fe}^{3+}][\text{SCN}^-]} = \frac{0.40}{(y/2 - 0.40) \times 0.10} = 40$$

Therefore,  $y/2 - 0.40 = 0.10$  and  $y = 1.0$ .

The concentration of the  $\text{Fe}(\text{NO}_3)_3$  solution must have been 1.0 M.

25. Choose the **CORRECT** statement based on the following reduction potentials.

$\text{Ag}^+/\text{Ag}$	0.80 V	$\text{Fe}^{2+}/\text{Fe}$	-0.44 V
$\text{Cu}^{2+}/\text{Cu}$	0.34 V	$\text{Zn}^{2+}/\text{Zn}$	-0.76 V
$\text{Sn}^{2+}/\text{Sn}$	-0.14 V	$\text{Mg}^{2+}/\text{Mg}$	-2.37 V

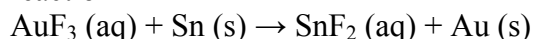
- A) Fe will displace  $\text{H}^+$  from solution**  
 B) Mg will not displace  $\text{Zn}^{2+}$  from solution  
 C) Cu will displace  $\text{Sn}^{2+}$  from solution  
 D) Sn will not displace  $\text{Ag}^+$  from solution  
 E) Fe will displace  $\text{Zn}^{2+}$  from solution

Fe will displace  $\text{H}^+$  from solution. This means  $\text{H}^+$  can oxidize Fe. This is true because the reduction potential of  $\text{H}^+$  (0 V) is greater than that of  $\text{Fe}^{2+}$  (-0.44 V). Mg will displace  $\text{Zn}^{2+}$  ( $\text{Mg}^{2+}$  has a lower reduction potential). Cu will not displace  $\text{Sn}^{2+}$ , Sn will displace  $\text{Ag}^+$  and Fe will not displace  $\text{Zn}^{2+}$ .

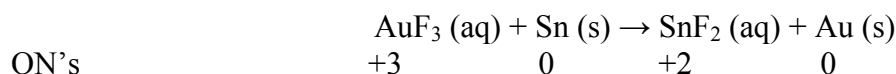
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26. Determine the **FALSE** statement regarding the following spontaneous electrochemical reaction

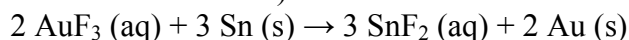


- A)  $\text{Au}^{3+}$  is acting as the oxidizing agent  
 B)  $\text{Au} (\text{s})$  is a weaker reducing agent than  $\text{Sn} (\text{s})$   
 C)  $Q$  for the reaction is  $[\text{SnF}_2]^3 / [\text{AuF}_3]^2$   
**D) In the balanced equation with smallest whole number coefficients, there is a transfer of 3 electrons**  
 E) The reduction potential for  $\text{Sn}^{2+}$  is lower than for  $\text{Au}^{3+}$



$\text{Au}^{3+}$  is reduced to  $\text{Au}$  – i.e.  $\text{Au}^{3+}$  is the oxidizing agent. The reaction proceeds away from the stronger reducing (or oxidizing agent).  $\text{Sn}$  is the reducing agent on the left, while  $\text{Au}$  is the reducing agent on the right. The reaction is towards  $\text{Au}$ , the weaker reducing agent.

Balanced reaction (6 electrons are transferred):

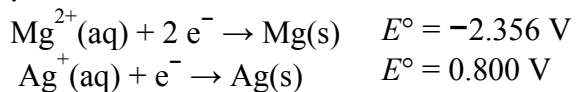


Clearly,

$$Q = \frac{[\text{SnF}_2]^3}{[\text{AuF}_3]^2}$$

Reduction of  $\text{Au}^{3+}$  is more favorable than  $\text{Sn}^{2+}$  – i.e.  $\text{Sn}^{2+}$  has a lower reduction potential.

27. Determine  $E^\circ_{\text{cell}}$  for the reaction:  $2 \text{Ag}^+ + \text{Mg} \rightarrow 2 \text{Ag} + \text{Mg}^{2+}$ . The half reactions are:



- A) 1.556 V  
 B) -2.356 V  
 C) -1.556 V  
**D) 3.156 V**  
 E) 3.956 V

$$E^\circ_{\text{cell}} = 0.800 - (-2.356) \text{ V} = 3.156 \text{ V}.$$

28. Choose the one **FALSE** statement from among the following:

- A) The process,  $\text{CO}_2(\text{g}) \rightarrow \text{CO}_2(\text{s})$ , releases energy.
- B) For all processes of a closed system,  $q = 0$ .**
- C) If two samples absorb the same quantity of heat, the sample with the lower heat capacity has the larger increase in temperature.
- D) The process,  $\text{O}_2(\text{g}) \rightarrow 2 \text{O}(\text{g})$ , requires an input of energy.
- E) All diatomic molecules have molar heat capacities that are larger than their specific heat capacities.

$\text{CO}_2(\text{g}) \rightarrow \text{CO}_2(\text{s})$  is a “deposition” process. It involves the making of the intermolecular bonds in  $\text{CO}_2(\text{s})$ . Bonds are only made - no bonds are broken. The process is exothermic – i.e. energy is released.  $q = 0$  only for isolated systems.

Closed systems have no matter lost or gained, but  $q$  can have any value.

Since  $q = C \Delta T$ ,  $\Delta T = q / C$  which is greater if  $C$  is smaller ( $q$  is fixed).

29. In an ice calorimeter, 0.3233 g of Mn are oxidized to  $\text{Mn}^{2+}$  in an exothermic reaction. As a result, 3.87 g of ice melts, producing water. How much **heat** (in kJ) was released by the reaction, **per mole of manganese (Mn)**?  
Molar heat of fusion (i.e. melting) for  $\text{H}_2\text{O}(\text{l}) = 6.01 \text{ kJ/mol}$ .

- A) 219**
- B) 108
- C) 1.29
- D) 70.8
- E) 3.99

The heat released in an ice calorimeter is determined by the amount of ice melted (or water frozen, in the case of an endothermic reaction). Note that the heat of fusion is given in on a per mole basis. Thus,

$$q = n \Delta H_{\text{fus}} = (3.87 \text{ g} / 18.015 \text{ g mol}^{-1}) 6.01 \text{ kJ mol}^{-1} \\ = 1.29_1 \text{ kJ}$$

$$\text{Moles of Mn oxidized} = 0.3233 \text{ g} / 54.938 \text{ g mol}^{-1} = 0.005885 \text{ mol}$$

$$\text{Heat released per mole of Mn} = 1.29_1 \text{ kJ} / 0.005885 \text{ mol} = 219 \text{ kJ mol}^{-1}$$

Name: \_\_\_\_\_

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30. A coffee cup calorimeter, including the water it contains, has a heat capacity of 425 J/K, and is initially at a temperature of 23.07°C. A 16.9 gram piece of nickel metal, initially at a temperature of 4.0°C is placed in the calorimeter. The final temperature of the calorimeter and the metal is 22.74°C. What is the **specific heat of nickel metal (in J g<sup>-1</sup> K<sup>-1</sup>)** ?

- A) 0.17  
B) 145  
**C) 0.44**  
D) 1.4  
E) 25

The heat lost by the calorimeter = the heat gained by the nickel

$$C_{\text{calorimeter}} \Delta T_{\text{calorimeter}} = -C_{\text{nickel}} \Delta T_{\text{nickel}}$$
$$425 \text{ J/K} \times (22.74 - 23.07 \text{ K}) = -C_{\text{nickel}} (22.74 - 4.0 \text{ K})$$

So,

$$C_{\text{nickel}} = 140 \text{ J} / 18.74 \text{ K} = 7.48 \text{ J/K}$$

The specific heat of nickel is its heat capacity per g – i.e.

$$c_{\text{nickel}} = 7.48 \text{ J/K} / 16.9 \text{ g} = 0.44 \text{ J g}^{-1} \text{ K}^{-1}$$