Chem. 1B Final

March 20, 2008

Name: ____________________________

   Last Name            First Name

Perm # ____________________________

INSTRUCTIONS: No hats allowed. No sharing of calculators. Cell Phones, iPods, headsets, etc. must be turned off and put away.

SCANTRON FORM: Use a PENCIL
   1) Write your name
   2) Bubble in FORM A
   3) Bubble in your PERM number (7 digits only, no extra numbers)

INFORMATION PAGE: An information page is provided separately. No other notes or books are allowed.

There are 8 pages, 32 questions. Each question is worth 5 points. SHOW ALL YOUR WORK on the exam.

Turn in the Scantron form and your Exam.

1. For a reaction, A → B + C, ΔH = −30 kJ/mol and the activation energy, E_a = 45 kJ/mol. What is the activation energy for the reverse reaction, B + C → A?
   a) 75 kJ/mol
   b) 45 kJ/mol
   c) 15 kJ/mol
   d) 30 kJ/mol
   e) None of these

   ANSWERS are given on the last page

2. The decomposition of KClO_3 produces O_2 (g) according to the following reaction.

   \[ \text{KClO}_3 (s) \rightarrow \text{KCl} (s) + \frac{3}{2} \text{O}_2 (g) \]

   For the reaction, ΔH^o = −44.7 kJ and ΔS^o = +59.1 J/K. This reaction will be
   a) spontaneous at all temperatures
   b) spontaneous only when T > 756 K
   c) spontaneous only when T < 756 K
   d) nonspontaneous at all temperatures
   e) spontaneous only when T = 756 K
3. Given the following data at 25 °C:

\[
\begin{align*}
\text{CH}_3\text{OH} + 3/2 \text{O}_2 & \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O} & \Delta H = -726.6 \text{ kJ} \\
\text{C (s, gr)} + 1/2 \text{O}_2 & \rightarrow \text{CO} & \Delta H = -110.5 \text{ kJ} \\
\text{C (s, gr)} + \text{O}_2 & \rightarrow \text{CO}_2 & \Delta H = -393.5 \text{ kJ} \\
\text{H}_2(\text{g}) + 1/2 \text{O}_2 & \rightarrow \text{H}_2\text{O} & \Delta H = -285.8 \text{ kJ}
\end{align*}
\]

Calculate the $\Delta H$ for the following reaction.

\[
\text{CO} + 2 \text{H}_2 \rightarrow \text{CH}_3\text{OH}
\]

a) + 157.8 kJ  

b) - 349.0 kJ  

c) + 128.0 kJ  

d) - 157.8 kJ  

e) - 128.0 kJ

4. Consider the following reaction at 25°C and 1 atm.

\[
\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2 \text{NO}_2(\text{g})
\]

For this system at equilibrium, how will raising the temperature affect the amount of NO2 present?

a) The amount of NO2 increases  

b) The amount of NO2 decreases  

c) The amount of NO2 stays the same  

d) More information is needed to answer this question.

5. For the combustion of ethene gas, \(\text{C}_2\text{H}_4(\text{g})\), the standard enthalpy change, \(\Delta H^\circ = -1411.1 \text{ kJ/mol}\) at 298 K.

Given the following enthalpies of formation, calculate \(\Delta H^\circ_f\) for \(\text{C}_2\text{H}_4(\text{g})\).

\[
\begin{align*}
\text{CO}_2(\text{g}) & \quad -393.5 \\
\text{H}_2\text{O}(\text{l}) & \quad -285.9
\end{align*}
\]

a) + 52.3 kJ/mol  

b) - 52.3 kJ/mol  

c) + 731.7 kJ/mol  

d) - 731.7 kJ/mol  

e) - 2769.9 kJ/mol
6. The enthalpy of vaporization for 1.00 mole of ethanol is 38.7 kJ/mol at its normal boiling point of 78°C. Predict the sign or value of w, q, and ΔE when 1.00 mol of ethanol is vaporized reversibly at 78°C and 1 atm. **Hint:** Calculate ΔE to determine if it is less than or greater than zero.

   a) \( w = 0 \) \( q > 0 \) \( \Delta E > 0 \)
   b) \( w > 0 \) \( q > 0 \) \( \Delta E > 0 \)
   c) \( w < 0 \) \( q > 0 \) \( \Delta E < 0 \)
   d) \( w = 0 \) \( q > 0 \) \( \Delta E > 0 \)
   e) \( w < 0 \) \( q > 0 \) \( \Delta E > 0 \)

7. Same problem continued. The enthalpy of vaporization for 1.00 mole of ethanol is 38.7 kJ/mol at its normal boiling point of 78°C. Predict the sign or value of \( \Delta S \) and \( \Delta G \) when 1.00 mol of ethanol is vaporized reversibly at 78°C and 1 atm.

   a) \( \Delta S < 0 \) \( \Delta G > 0 \)
   b) \( \Delta S > 0 \) \( \Delta G < 0 \)
   c) \( \Delta S < 0 \) \( \Delta G = 0 \)
   d) \( \Delta S > 0 \) \( \Delta G = 0 \)
   e) \( \Delta S = 0 \) \( \Delta G = 0 \)

8. Consider the dissolution of \( \text{CaCl}_2 \).

\[
\text{CaCl}_2(s) \rightarrow \text{Ca}^{2+}(aq) + 2 \text{Cl}^-(aq) \quad \Delta H = -81.5 \text{ kJ}
\]

A 12.5 g sample of \( \text{CaCl}_2 \) is dissolved in water, with both substances at 25.0 °C. Calculate the final temperature of the solution assuming no heat is lost to the surroundings, the final mixture has a mass of 100.0 g and the solution has a specific heat capacity of \( 4.18 \text{ J}^{\circ} \text{C}^{-1} \text{g}^{-1} \).

   a) 47.0 °C
   b) 25.2 °C
   c) 36.7 °C
   d) 3.04 °C
   e) none of these
9. Consider the combustion of ethanol in air at constant temperature and pressure:

\[ 2 \text{CH}_3\text{OH}(l) + 3 \text{O}_2(g) \rightarrow 2 \text{CO}_2(g) + 4 \text{H}_2\text{O}(l) \]

How is the value of \( \Delta E \) for this reaction related to the value of \( \Delta H \)?

a) \( \Delta E \) is not related to \( \Delta H \)
b) \( \Delta E \) is less negative than \( \Delta H \)
c) \( \Delta E \) is more negative than \( \Delta H \)
d) \( \Delta E \) is equal to \( \Delta H \)

10. What is the oxidation state of Be in Be\(_2\)O\(_3\)\(^{2-}\)?

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

11. In the reaction below, which substance is the oxidizing agent?

\[ \text{Pb} \text{(s)} + \text{PbO}_2 \text{(s)} + 2 \text{H}^+(\text{aq}) + 2\text{HSO}_4^- \text{(g)} \rightarrow 2 \text{PbSO}_4 \text{(s)} + 2 \text{H}_2\text{O}(l) \]

a) Pb  
b) PbO\(_2\)  
c) H\(^+\)  
d) HSO\(_4^-\)  
e) PbSO\(_4\)

12. A galvanic cell is constructed with a copper electrode in a CuSO\(_4\) (aq) solution and the lead electrode in a Pb(NO\(_3\))\(_2\) (aq) solution at 25°C. The standard reduction potentials are:

\[ \text{Pb}^{2+} + 2 \text{e}^- \rightarrow \text{Pb} \text{(s)} \quad E^\circ = -0.13 \text{ V} \]
\[ \text{Cu}^{2+} + 2 \text{e}^- \rightarrow \text{Cu} \text{(s)} \quad E^\circ = +0.34 \text{ V} \]

When sulfuric acid is added to the Pb(NO\(_3\))\(_2\) (aq) solution, a PbSO\(_4\) (s) precipitate is formed. This will cause the cell potential to

a) increase  
b) decrease  
c) remain unchanged  
d) more information is needed to determine the effect

13. Consider the following reaction for galvanic cell at 25°C. Initially [Au\(^{3+}\)] = 5.00 M and [Ni\(^{2+}\)] = 4.00 M. Calculate the cell potential after the cell has operated long enough for [Au\(^{3+}\)] to have changed by 2.00 M.

\[ 2 \text{Au}^{2+} \text{(aq)} + 3 \text{Ni} \text{(s)} \rightarrow 3 \text{Ni}^{2+} \text{(aq)} + 2 \text{Au} \text{(s)} \quad E^\circ = 1.73 \text{ V} \]

a) 1.71 V  
b) 1.74 V  
c) 1.68 V  
d) 1.78 V  
e) 1.69 V
14. For the following cell reaction, which will increase the cell voltage the most?

$$2 \text{Fe}^{2+}(aq) + \text{Cu}^{2+}(aq) \rightarrow 2 \text{Fe}^{3+}(aq) + \text{Cu}(s)$$

a) halve [Cu$$^{2+}$$]

b) halve [Fe$$^{2+}$$]

c) double [Cu$$^{2+}$$]

d) double [Fe$$^{2+}$$]

e) cut the Cu electrode in half

15. Given the following two half reactions:

$$\text{Fe}^{3+} + 3 \text{e}^- \rightarrow \text{Fe}(s) \quad E^0 = -0.034 \text{ V}$$

$$\text{Zn}^{2+} + 2 \text{e}^- \rightarrow \text{Zn}(s) \quad E^0 = -0.76 \text{ V}$$

Calculate the $\Delta G^\circ$, for the cell reaction at 25 °C.

$$\text{Zn}(s) \mid \text{Zn}^{2+}(aq) \parallel \text{Fe}^{3+}(aq) \mid \text{Fe}(s)$$

a) −210 kJ

b) −140 kJ

c) −70.0 kJ

d) −463 kJ

e) −420 kJ

16. Which of the following atoms has the smallest radius?

a) Mg

b) Al

c) P

d) S

e) Cl

17. Which of the following elements is the most electron negative?

a) P

b) C

c) N

d) Se

e) B

18. Which of the following bonds is the most polar?

a) O − N

b) H − Br

c) O − H

d) C − H

e) N − H

19. Which of the following atoms or ions has 3 unpaired electrons?

a) O

b) Al

c) S$$^{2-}$$

d) Zn$$^{2+}$$

e) Co

20. What is the maximum number of electrons in an atom that can have the quantum numbers $n = 4$, $m_l = +1$.

a) 2

b) 4

c) 6

d) 8

e) 10
21. Which of the following electronic configurations correspond to an excited state?
   I. 1s²2s²3p¹
   II. 1s²2s²2p⁴
   III. 1s²2s²2p⁶3s²
   IV. [Ar]4s²3d⁶4p³
   a) I and II
   b) I, II, and III
   c) I, II, III, and IV
   d) I, III, and IV
   e) II, III, and IV

22. Identify the atom with the electron configuration [Ar]4s²3d⁶4p³.
   a) As  b) Fe  c) Cu  d) Zn  e) Co

23. The successive ionization energies for an unknown element are shown below.
   IE₁ = 580 kJ/mol
   IE₂ = 1815 kJ/mol
   IE₃ = 2740 kJ/mol
   IE₄ = 11,600 kJ/mol
   To which family in the periodic table does the unknown element most likely belong?
   a) Group 1A  b) Group 2A  c) Group 3A  d) Group 4A  e) Group 5A

24. Order the ions in size from smallest to largest.
   a) Ca²⁺ < K⁺ < Cl⁻ < S²⁻ < Al³⁺ < Te²⁻
   b) Al³⁺ < S²⁻ < Cl⁻ < K⁺ < Ca²⁺ < Te²⁻
   c) Te²⁻ < Ca²⁺ < K⁺ < S²⁻ < Cl⁻ < Al³⁺
   d) Cl⁻ < S²⁻ < Al³⁺ < K⁺ < Ca²⁺ < Te²⁻
   e) Al³⁺ < Ca²⁺ < K⁺ < Cl⁻ < S²⁻ < Te²⁻

25. In the Lewis structure for SO₂, how many bonds are formed and how many lone pairs of electrons are there on the sulfur atom.
   a) 2 bonds, O – S – O, no lone pairs of electrons on S
   b) 2 bonds, O – S – O, two lone pairs of electrons on S
   c) 3 bonds, O = S – O, one lone pair of electrons on S
   d) 3 bonds, O = S – O, no lone pairs of electrons on S
   e) 4 bonds, O = S = O, no lone pairs of electrons on S

26. In the Lewis structure for H₂CO, what is the total number of valence electrons available for bonding and how many bonds are formed?
   a) 12 valence electrons, 3 bonds
   b) 12 valence electrons, 4 bonds
   c) 16 valence electrons, 3 bonds
   d) 16 valence electrons, 4 bonds
   e) 14 valence electrons, 3 bonds
27. When you draw the Lewis structure for SO₃, and any resonance structures (with sulfur as the central atom) what is the total number of structures?

a) 1
b) 2
c) 3
d) 4

28. Consider the reaction: \( A \rightarrow B + C \)
The reaction is zero order with respect to \( A \), and the rate constant for the reaction, \( k = 0.07 \text{ M/s at 25 } ^\circ \text{C.} \)
The initial concentration \( [A]_0 = 1.0 \times 10^{-3} \text{ M.} \) Calculate the concentrations of \( B \) after \( 4.0 \times 10^{-3} \) sec has elapsed.

a) \([B] = 7.2 \times 10^{-4} \text{ M}\)
b) \([B] = 2.8 \times 10^{-4} \text{ M}\)
c) \([B] = 1.3 \times 10^{-3} \text{ M}\)
d) \([B] = 1.0 \times 10^{-3} \text{ M}\)
e) none of these

29. For the reaction: \( A + B \rightarrow C \) The following data were obtained.

<table>
<thead>
<tr>
<th>Experiment</th>
<th>[A] (M)</th>
<th>[B] (M)</th>
<th>Initial Rate (M/s)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.030</td>
<td>0.060</td>
<td>(2.5 \times 10^{-5})</td>
</tr>
<tr>
<td>2</td>
<td>0.030</td>
<td>0.020</td>
<td>(2.5 \times 10^{-5})</td>
</tr>
<tr>
<td>3</td>
<td>0.060</td>
<td>0.060</td>
<td>(10.0 \times 10^{-5})</td>
</tr>
</tbody>
</table>

a) \(\text{Rate} = k [A][B]\)
b) \(\text{Rate} = k [A]^2[B]\)
c) \(\text{Rate} = k [A]^2[B]^2\)
d) \(\text{Rate} = k [A][B]^2\)
e) \(\text{Rate} = k [A]^2\)

30. Consider the reaction: \( 2 \text{N}_2\text{O}_5 \rightarrow 2 \text{N}_2\text{O}_4 + \text{O}_2 \) For this reaction, the activation energy, \( E_a = 106 \text{ kJ/mol}, \) and the rate constant, \(k = 3.46 \times 10^{-5} \text{ s}^{-1}\) at \( 298 \text{ K}. \) What is the rate constant at \( 305 \text{ K}. \)

a) \(2.4 \times 10^{-5} \text{ s}^{-1}\)
b) \(4.8 \times 10^{-5} \text{ s}^{-1}\)
c) \(6.0 \times 10^{-5} \text{ s}^{-1}\)
d) \(1.2 \times 10^{-5} \text{ s}^{-1}\)
e) \(9.2 \times 10^{-5} \text{ s}^{-1}\)

(More problems on next page)
31. A mixture of hydrogen and chlorine gas remains unreacted until it is exposed to ultraviolet light from a burning magnesium strip. Then the following reaction occurs very rapidly.

\[
\text{H}_2 (g) + \text{Cl}_2 (g) \rightarrow 2\text{HCl} (g) \quad \Delta G = -45.54 \text{ kJ} \\
\Delta H = -44.12 \text{ kJ} \\
\Delta S = -4.76 \text{ J} / \text{K} 
\]

Select the statement below that best explains this behavior.

a) The reactants are thermodynamically more stable than the products.
b) The reaction has a small equilibrium constant.
c) The negative value for \(\Delta S\) slows down the reaction.
d) The reaction is spontaneous, but the kinetics are unfavorable.
e) The negative value for \(\Delta G\) increases the rate of the reaction.

32. The figure below represents part of the emission spectrum for a one-electron ion in the gas phase. All the lines result from electronic transitions from excited states to the \(n = 3\) state. In the spectrum, Line C corresponds to the electronic transition from the \(n = 4\) to the \(n = 3\) state.

![Spectrum Diagram]

If the wavelength of line B is 111.7 nm, calculate the wavelength of A. \(1 \text{ nm} = 1 \times 10^{-9} \text{ m}\)

a) 104.3 nm 
b) 121.6 nm 
c) 106.1 nm 
d) 102.6 nm 
e) none of these

1. (4 pts) Which of the following is a properly written formation reaction?

a) \( 3 \text{H}_2 (g) + \text{N}_2 (g) \rightarrow 2 \text{NH}_3 (g) \)

b) \( 2 \text{H} (g) + \text{O} (g) \rightarrow \text{H}_2 \text{O} (g) \)

c) \( \text{CaO} (s) + \text{CO}_2 (g) \rightarrow \text{CaCO}_3 (s) \)

d) \( \text{Ti} (s) + \text{O}_2 (g) \rightarrow \text{TiO}_2 (g) \)

e) more than one of these

2. (4 pts) When ammonium chloride is dissolved in water (both initially at the same temperature), the resulting solution decreases in temperature. What is the sign of \( \Delta H \) for the dissolution reaction?

a) positive  
b) negative  
c) \( \Delta H = 0 \)  
d) more information needed

3. (4 pts) For a given process, \( \Delta H = -3.7 \text{kJ/mol} \) and \( \Delta S = -6.7 \text{ J/K} \text{mol}^{-1} \). Under what conditions will this process be spontaneous?

a) High temperature  
b) Low temperature  
c) The reaction spontaneous at all temperatures  
d) The reaction is not spontaneous at all temperatures  
e) More information is needed
4. (4 pts) Which of the following atoms has exactly two unpaired electrons?
   a) Be
   b) N\(^{2-}\)
   c) Al\(^{+}\)
   d) S
   e) more than one of these has exactly two unpaired electrons

5. (4 pts) Determine the maximum number of electrons that can have the quantum numbers \( n = 3, m_s = +1/2 \)
   a) 3
   b) 6
   c) 9
   d) 18
   e) 1

Questions 6-7: Use the information below to answer the following questions:

\[
\begin{array}{l}
F_2 + 2e^- \longrightarrow 2F^- & E^0 (V) = 2.87 \\
Au^{3+} + 3e^- \longrightarrow Au & E^0 (V) = 1.50 \\
Cl_2 + 2e^- \longrightarrow 2Cl^- & E^0 (V) = 1.33 \\
Ag^+ + e^- \longrightarrow Ag & E^0 (V) = 0.80 \\
I_2 + 2e^- \longrightarrow 2I^- & E^0 (V) = 0.53 \\
Cu^{2+} + 2e^- \longrightarrow Cu & E^0 (V) = 0.34 \\
Pb^{2+} + 2e^- \longrightarrow Pb & E^0 (V) = -0.13 \\
\end{array}
\]

6. Which of the following can reduce Au\(^{3+}\) but not Cu\(^{2+}\)?
   A) Pb\(^{2+}\)
   B) I\(^-\)
   C) Ag\(^+\)
   D) Pb
   E) F

7. Which of the following can be oxidized by I\(_2\)?
   A) Pb\(^{2+}\)
   B) Cl\(_2\)
   C) Au
   D) Pb
   E) none of these
8. (5 pts) Consider the Lewis structure shown below:

\[
\begin{array}{c}
O=\cdot X \cdot - Cl : \\
\end{array}
\]

Element X must belong to which group of the periodic table?

a) Group 1A  
b) Group 5A  
c) Group 6A  
d) Group 7A  
e) Group 8A

Questions 9-10: Consider the following energy diagram for a reaction:

9. (4 pts) Which of the following represents the activation energy for the forward reaction?

a) A  
b) B  
c) C  
d) none of these  
e) more information is needed

10. (4 pts) This reaction is:

a) exothermic  
b) endothermic  
c) neither endothermic nor exothermic  
d) both endothermic and exothermic  
e) more information is needed
11. (5 pts) Which is the most polar bond?
   a) C-F 
   b) N-F 
   c) O-F 
   d) F-F 

12. (5 pts) When a lead acid battery in a car is discharged, the following reaction takes place:

   \[ \text{Pb (s) + PbO}_2 \text{ (s) + 2 H}_2\text{SO}_4 \text{(aq) \rightarrow 2 PbSO}_4 \text{(s) + 2 H}_2\text{O (l)}} \]

   How many electrons are transferred in this reaction?
   A) 0 
   B) 1 
   C) 2 
   D) 4 
   E) 6 

13. (5 pts) Consider the following rate law: \( \text{Rate} = k \ [A][B]^2[C]^2 \)

   What are the units of the rate constant \( k \)?
   A) \( \text{M}^5 \text{s}^{-1} \)
   B) \( \text{M}^{-3} \text{s}^{-1} \)
   C) \( \text{M} \text{s}^{-1} \)
   D) \( \text{M}^{-4} \text{s}^{-1} \)
   E) none of these 

14. (5 pts) Consider the ions \( \text{NO}_2^- \) and \( \text{NO}_3^- \). Which ion has the longest bond lengths between nitrogen and oxygen?
   a) \( \text{NO}_2^- \)
   b) \( \text{NO}_3^- \)
   c) \( \text{NO}_2^- \) and \( \text{NO}_3^- \) have the same bond lengths between nitrogen and oxygen
   d) more information is needed
15. (5 pts) Which of the following has the largest radius?
   a) Br⁻
   b) Kr
   c) Rb⁺

16. (6 pts) For the following overall reaction:

\[ \text{Cl}_2 + \text{CO} \rightarrow \text{COCl}_2 \]

The following reaction mechanism has been proposed:

\[
\begin{align*}
\text{Cl}_2 & \leftrightarrow 2 \text{Cl} \quad \text{fast} \quad \text{Forward rate constant} = k_1 \quad \text{Reverse rate constant} = k_1^{-1} \\
\text{Cl} + \text{CO} & \leftrightarrow \text{COCl} \quad \text{fast} \quad \text{Forward rate constant} = k_2 \quad \text{Reverse rate constant} = k_2^{-1} \\
\text{COCl} + \text{Cl}_2 & \rightarrow \text{COCl}_2 + \text{Cl} \quad \text{slow} \quad \text{Rate constant} = k_3 \\
2 \text{Cl} & \rightarrow \text{Cl}_2 \quad \text{fast} \quad \text{Rate constant} = k_4
\end{align*}
\]

Derive a rate law for the overall reaction that is consistent with the proposed mechanism.

A) rate = k[Cl₂][CO]
B) rate = k[COCl][Cl₂]
C) rate = k [Cl₂]^(1/2)[CO]
D) rate = k [Cl₂]^(1/2)[CO]
E) none of these

17. (4 pts) Identify the atom with the electron configuration \([\text{Ar}]4s^23d^5\)

a) Cr
b) Mn
c) Tc
d) Ru
e) Br

18. (4 pts) Which of the following orbital designations is not possible?

a) 4s
b) 2p
c) 4d
d) 2f
e) more than one of these orbital designations is not possible
19. (4 pts) Which of the following is isoelectronic with N\(^{2-}\)?
   a) B
   b) N\(^{2-}\)
   c) Ne\(^+\)
   d) F
   e) more than one of these is isoelectronic with N\(^{2-}\)

20. (5 pts) In the Lewis structure for the molecule NOF, how many lone pairs are on the nitrogen atom, and what is the formal charge on the nitrogen atom?
   a) 2 lone pairs, formal charge = -3
   b) 0 lone pairs, formal charge = +1
   c) 1 lone pair, formal charge = 0
   d) 2 lone pairs, formal charge = -1
   e) 1 lone pair, formal charge = -3

21. (5 pts) The standard enthalpy of formation for HCl (g) is -92.3 kJ/mol. What is the standard enthalpy of the following reaction:

\[
2 \text{HCl (g)} \rightarrow \text{H}_2 (g) + \text{Cl}_2 (g)
\]

   a) 92.3 kJ
   b) 92.3 kJ
   c) -184.6 kJ
   d) 184.6 kJ
   e) none of the above

22. (6 pts) Calculate the de Broglie wavelength of a sulfur atom traveling at 2.39 \(\times\) 10\(^5\) m/s.
   a) 8.65 \(\times\) 10\(^{-38}\) m
   b) 8.65 \(\times\) 10\(^{-41}\) m
   c) 1.44 \(\times\) 10\(^{-61}\) m
   d) 5.21 \(\times\) 10\(^{-14}\) m
   e) 5.21 \(\times\) 10\(^{-37}\) m
23. (6 pts) Consider the following reaction:

\[ \text{Fe}_3\text{O}_4 (s) + \text{CO} (g) \rightarrow 3 \text{FeO} (s) + \text{CO}_2 (g) \quad \Delta H = ??? \]

Calculate \(\Delta H\) for this reaction, given the following reactions and their corresponding \(\Delta H\) values:

\[ 2 \text{Fe} (s) + 3 \text{CO}_2 (g) \rightarrow \text{Fe}_2\text{O}_3 (s) + 3 \text{CO} (g) \quad \Delta H = 14 \text{ kJ} \]
\[ 2 \text{Fe}_3\text{O}_4 (s) + \text{CO}_2 (g) \rightarrow 3 \text{Fe}_2\text{O}_3 (s) + \text{CO} (g) \quad \Delta H = 23 \text{ kJ} \]
\[ \text{Fe} (s) + \text{CO}_2 (g) \rightarrow \text{FeO} (s) + \text{CO} (g) \quad \Delta H = -1.0 \text{ kJ} \]

a) -16.8 kJ  
b) 36 kJ  
c) 29.5 kJ  
d) -12.5 kJ  
e) none of these

24. (5 pts) For the following process involving compound X:

\[ X (s) \rightarrow X (l) \]

\(\Delta H^\circ = 8.8 \text{ kJ/mol}\) and \(\Delta S^\circ = 43.0 \text{ J mol}^{-1} \text{ K}^{-1}\). What is the normal melting point for compound X?

a) 205 °C  
b) 68°C  
c) -68°C  
d) -205°C  
e) none of these

25. (5 pts) In an isothermal reversible transformation of 4.5 moles of an ideal gas at 298K, 8.0 kJ of heat is released from the system. If the final volume of the gas was 6.3 L, what was the initial volume?

a) 8.0 L  
b) 6.4 L  
c) 13 L  
d) 3.0 L  
e) none of these
26. (5 pts) Consider the Galvanic cell at 25°C described below:

\[ X | X^{2+} | | Y^{3+} | Y \]

Where \( X \) and \( Y \) are unknown metals. Given that the equilibrium constant \( K \) for the reaction in this Galvanic cell is \( 3.7 \times 10^8 \) and the standard reduction potential for \( X^{2+} \)

\[ X^{2+} + 2 \text{e}^- \rightarrow X \quad E^0 = 0.34 \text{ V} \]

What is the standard reduction potential of \( Y^{3+} \)?

a) \(-0.26 \text{ V}\)
b) \(0.42 \text{ V}\)
c) \(0.85 \text{ V}\)
d) \(0.34 \text{ V}\)
e) none of these

27. (5 pts) Estimate \( \Delta H \) for the following reaction

\[ \text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O} \]

Given the following bond energy data:

<table>
<thead>
<tr>
<th>Bond</th>
<th>( D ) (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C-H</td>
<td>413</td>
</tr>
<tr>
<td>O=O</td>
<td>495</td>
</tr>
<tr>
<td>C=O</td>
<td>799</td>
</tr>
<tr>
<td>C-O</td>
<td>358</td>
</tr>
<tr>
<td>O-H</td>
<td>467</td>
</tr>
</tbody>
</table>

a) \(-883 \text{ kJ}\)
b) \(-824 \text{ kJ}\)
c) \(-383 \text{ kJ}\)
d) \(110 \text{ kJ}\)
e) none of these
28. (5 pts) A first-order reaction is 45% complete at the end of 19 min. What is the value of the rate constant?

a) 0.096 min⁻¹  
b) 0.20 min⁻¹  
c) 3.1 x 10⁻² min⁻¹  
d) 1.4 x 10⁻² min⁻¹  
e) 4.2 x 10⁻² min⁻¹

29. (6 pts) An unknown metal X is electrolyzed. It took 6.2 hours for a current of 16.8 A to plate out 90.32 g of the metal from a solution containing X(NO₃)₃. Identify the metal. Note: the nitrate ion is NO₃⁻.

a) Na  
b) Fe  
c) Ga  
d) Mo  
e) none of these

30. (6 pts) In the hydrogen atom spectrum, an emission line is observed at a wavelength of 1282 nm. If this line corresponds to an electronic transition to a final state of n = 3, what was the initial state? Note: 1 nm = 10⁻⁹ m.

a) 6  
b) 4  
c) 5  
d) 2  
e) none of these
31. (5 pts) A galvanic cell is constructed with copper electrodes and Cu²⁺ in each compartment. In one compartment, \([\text{Cu}^{2+}] = 4.0 \times 10^{-3}\) M, and in the other compartment, \([\text{Cu}^{2+}] = 1.3\) M. Calculate the potential for this cell at 25°C.

a) 0.41 V  

b) 0.074 V  

c) 0.75 V  

d) −0.074 V  

e) −0.41 V

32. (6 pts) A 20.0 gram sample of liquid propane \((\text{C}_3\text{H}_8)\) at −60°C is heated to its boiling point at −42°C and completely vaporized at −42°C and a constant pressure of 1 atm. Calculate \(\Delta S\) for this process. The \(C_p\) for liquid propane is 98.4 J K⁻¹ mol⁻¹. The enthalpy of vaporization for propane is 15.7 kJ/mol.

a) 75.9 J/K  

b) 34.5 J/K  

c) 3.70 J/K  

d) −186 J/K  

e) none of these
33. (6 pts) Consider the decomposition of ethanol (C₂H₅OH) on an aluminum oxide surface:

\[ \text{C}_2\text{H}_5\text{OH} (g) \rightarrow \text{C}_2\text{H}_4 (g) + \text{H}_2\text{O} (g) \]

This reaction is zero order with respect to [C₂H₅OH] with a rate constant of 0.0036 M⁻¹ s⁻¹.

If the reaction is run with an initial [C₂H₅OH] of 0.398 M, how long will it take for all of the C₂H₅OH to react?

a) 192.5 sec  
b) the reaction will take infinite time  
c) 0.0036 sec  
d) 110.6 sec  
e) none of these

34. (6 pts) Consider the following reaction:

\[ \text{A} + 2 \text{B} \rightleftharpoons 3 \text{C} \]

Initially, [A] = 1.42 M, [B] = 2.78 M, and [C] = 0 M. If the equilibrium concentration of B is 1.32 M at 298 K, then what is \( \Delta G^\circ \) for this reaction?

a) −2.17 kJ  
b) 3.85 kJ  
c) 0.55 kJ  
d) −5.37 kJ  
e) none of these
35. (6 pts) 3 moles of an ideal monatomic gas initially at 2.00 atm and 20. L is expanded isothermally from its initial state to a final state \( (V_f = 36 \text{ L}, P_f = 1.11 \text{ atm}) \) in two steps as shown below.

   **Step 1:** Expand to 32 L at a pressure of 1.25 atm  
   **Step 2:** Expand to 36 L at a pressure of 1.11 atm

   Calculate the ratio of the work done in this two-step process to the theoretical maximum amount of work that could be done in a transformation from the initial state to the final state.

   a) 0.41 (41%)  
   b) 0.83 (83%)  
   c) 0.87 (87%)  
   d) 0.0082 (0.82%)  
   e) none of these

36. (6 pts) An 83 gram sample of aluminum (heat capacity = 0.789 J g\(^{-1}\)°C\(^{-1}\)) initially at 98.3°C is added to 141 grams of water (heat capacity = 4.184 J g\(^{-1}\)°C\(^{-1}\)) initially at 25.7°C. What is the final temperature of the water and metal? Assume no heat is lost to the calorimeter or the surroundings.

   a) 16.6°C  
   b) 15.0°C  
   c) 72.6°C  
   d) 33.0°C  
   e) 36.4°C
37. (6 pts) Consider the Galvanic cell at 298 K described by the following line notation: \( \text{Sn} | \text{Sn}^{2+} | | \text{Ag}^+ | \text{Ag} \)

Initially, \([\text{Sn}^{2+}] = [\text{Ag}^+] = 0.97 \text{ M}\). Calculate the cell potential after the reaction has operated long enough for \([\text{Ag}^+]\) to change by 0.82 M.

\[
\begin{align*}
\text{Ag}^+ + e^- & \rightarrow \text{Ag} (s) \quad E^o = 0.80 \text{ V} \\
\text{Sn}^{2+} + 2 e^- & \rightarrow \text{Sn} (s) \quad E^o = -0.14 \text{ V}
\end{align*}
\]

a) 0.89 V 
b) 0.91 V 
c) 0.97 V 
d) 1.00 V 
e) none of these

38. (6 pts) For the vaporization of N\(_2\) at -197°C and 1 atm, \( \Delta H = 5.56 \text{ kJ/mol} \). If 2.3 moles of N\(_2\) is vaporized under these conditions, what is the change in internal energy? Assume ideal behavior for N\(_2\) gas.

a) 12.8 kJ 
b) -1.5 kJ 
c) -1440 kJ 
d) 11.3 kJ 
e) none of these
39. (6 pts) When a photon with a wavelength of 323 nm strikes the surface of a metal, an electron is ejected with a kinetic energy of 2.4x10^{-18} J. What is the binding energy of the metal, in units of kJ/mol?

a) 3.8x10^{-22} kJ/mol
b) 6.2x10^{-46} kJ/mol
c) 370 kJ/mol
d) 144 kJ/mol
e) 226 kJ/mol

40. (6 pts) Consider the reaction: \[ A \rightarrow \text{Products} \] The first, second, and third half lives for this reaction are 10 sec, 20 sec, and 40 sec, respectively. If the initial concentration of A is 2.1 M, what will be [A] after 16 seconds have elapsed?

a) 0.69 M
b) 0.14 M
c) 0.30 M
d) 0.42 M
e) 0.81 M
<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1) D – Heat of formation (Lecture 5)</td>
<td>21) D – Properties of ΔH and ΔH°, (Lecture 5)</td>
</tr>
<tr>
<td>2) A – Heat of reaction sign conventions (Lec. 2 &amp; 9)</td>
<td>22) D – de Broglie equation (Lectures 19-20)</td>
</tr>
<tr>
<td>3) B – Spontaneity temperature dependence (Lecture 8)</td>
<td>23) D – Hess’s Law (Lecture 5)</td>
</tr>
<tr>
<td>4) D – Electron configurations (Lecture 22)</td>
<td>24) C – ΔH and ΔS for phase transitions (Lecture 7)</td>
</tr>
<tr>
<td>5) C – Quantum numbers (Lecture 21, Book #12.75)</td>
<td>25) C – Isothermal reversible processes (Lecture 7)</td>
</tr>
<tr>
<td>6) B – Table of reduction potentials (Lecture 12)</td>
<td>26) B – Reduction potentials, Galvanic cells (Lec. 11)</td>
</tr>
<tr>
<td>7) D – Table of reduction potentials (Lecture 12)</td>
<td>27) B – ΔH from bond energy (Lecture 25 or 26)</td>
</tr>
<tr>
<td>9) A – Activation energy (Lecture 16)</td>
<td>29) C – Electrolysis (Lecture 11)</td>
</tr>
<tr>
<td>10) A – Energy diagrams for reactions (Lecture 1)</td>
<td>30) C – Bohr Model (Lecture 20)</td>
</tr>
<tr>
<td>11) A – Bond polarity (Lecture 23-24)</td>
<td>31) B – Concentration cell (Lecture 14)</td>
</tr>
<tr>
<td>12) C – Writing half reactions (Lecture 10)</td>
<td>32) B – Calculating ΔS for heating/cooling (Lecture 7)</td>
</tr>
<tr>
<td>13) D – Units of rate constant (Lecture 15)</td>
<td>33) D – Integrated rate laws (Lecture 15)</td>
</tr>
<tr>
<td>14) A – Bond order from Lewis Structure (Lecture 25 or 26)</td>
<td>34) D – Connection between ΔG &amp; K (ICE Table) (Lec. 10)</td>
</tr>
<tr>
<td>15) A – Trends in ionic radius (Lecture 24)</td>
<td>35) B – Calculate irreversible work (Lec. 2) Reversible work = maximum work! (Lec. 7)</td>
</tr>
<tr>
<td>16) C – Mechanisms (Lectures 17-18)</td>
<td>36) D – Calorimetry (Lecture 4-5)</td>
</tr>
<tr>
<td>17) B – Electron configurations (Lecture 22)</td>
<td>37) A – Nernst equation (Lecture 13)</td>
</tr>
<tr>
<td>18) D – Quantum #’s (Lecture 21, Book #12.61)</td>
<td>38) D – Relate ΔH and ΔE (Lecture 4)</td>
</tr>
<tr>
<td>19) C – Isoelectronic species (Lecture 24)</td>
<td>39) E – Photoelectric effect (Lecture 19)</td>
</tr>
<tr>
<td>20) C – Lewis structures &amp; Formal Charge (Lectures 24-25)</td>
<td>40) E – Relationship between t½ and integrated rate law (Lecture 15 &amp; book #15.46)</td>
</tr>
</tbody>
</table>
1. (8 pts) Use the following information to determine $\Delta H^\circ$ of CH$_3$OH (l).

\[
\begin{array}{ll}
\text{CO}_2 \text{ (g)} & -393.5 \\
\text{H}_2\text{O} \text{ (g)} & -241.8 \\
\end{array}
\]

\[
2 \text{ CH}_3\text{OH} \text{ (l)} + 3 \text{ O}_2 \text{ (g)} \rightarrow 2 \text{ CO}_2 \text{ (g)} + 4 \text{ H}_2\text{O} \text{ (g)} \quad \Delta H^\circ = -1277 \text{ kJ/mol}
\]

2. (9 pts) In class we observed the following reaction at room temperature and pressure:

\[
\text{Ba(OH)}_2 \cdot 8 \text{ H}_2\text{O} \text{ (s)} + 2 \text{ NH}_4\text{SCN} \text{ (s)} \rightarrow \text{Ba(SCN)}_2 \text{ (s)} + 2 \text{ NH}_3 \text{ (g)} + 10 \text{ H}_2\text{O} \text{ (l)}
\]

The solution produced was so cold that it froze the water underneath the beaker, adhering a block of wood to the beaker. Predict the sign of $\Delta G$, $\Delta H$ and $\Delta S$ for this reaction. Circle the answers.

$\Delta G < 0$ $\Delta G > 0$ $\Delta G = 0$

$\Delta H < 0$ $\Delta H > 0$ $\Delta H = 0$

$\Delta S < 0$ $\Delta S > 0$ $\Delta S = 0$

3. (8 pts) The major source of aluminum in the world is bauxite, a compound that consists primarily of aluminum oxide. The thermal decomposition of aluminum oxide to produce aluminum occurs via the following reaction.

\[
2 \text{ Al}_2\text{O}_3 \text{ (s)} \rightarrow 4 \text{ Al} \text{ (s)} + 3 \text{ O}_2 \text{ (g)}
\]

Given the standard enthalpy of formation of Al$_2$O$_3$ (s), $\Delta H^\circ_f = -1676 \text{ kJ/mol}$, determine how many grams of aluminum will be produced when 1000 kJ of heat is consumed.
4. (6 pts) Write the reactions that correspond to the following enthalpy changes:
   a) \( \Delta H_f^o \) for solid hydrogen cyanide, HCN (g).
   
   b) \( \Delta H_f^o \) for calcium hydroxide.

5. (6 pts) Write the electronic configuration and indicate the number of unpaired electrons for each of the following species.

<table>
<thead>
<tr>
<th>electronic configuration</th>
<th>number of unpaired electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ga</td>
<td></td>
</tr>
<tr>
<td>Co</td>
<td></td>
</tr>
<tr>
<td>O⁺</td>
<td></td>
</tr>
</tbody>
</table>

6. (4 pts) An atom is in an excited electronic state with the electron configuration [Ar]4s²3d⁶4p¹. Identify the atom.

7. a) (8 pts) Draw the Lewis structures, all three resonance forms, for N₂O. The central atom is a nitrogen. Assign formal charges on each of the atoms in all three resonance structures.

   b) (3 pts) Based on the assigned formal charges, which of the resonance structures is least favored.

   c) (3 pts) Based on the assigned formal charges, which of the resonance structures is most favored.
8. (12 pts) Circle the formula that best fits each of the following descriptions:

   a) smallest radius
      Mg, Al, Al³⁺, Cl⁻, P³⁻

   b) largest ionic radius
      Cs⁺, Ba²⁺, I⁻, Te²⁻

   c) most polar bond
      Li – Cl, Na – Cl, K – Cl

   d) longest bond length
      NO, CO⁻, N₂

   e) greatest electronegativity
      Si, Li, B, C

   f) largest second ionization energy
      Mg, Na, Al, Ca

9. a) (6 pts) Draw the Lewis structures for NO and NO₂.

   b) (4 pts) For which molecule is the NO bond length the longest, NO or NO₂? Circle the answer.
      NO, NO₂

10. (4 pts) At 1 atm pressure and 25°C, the following phase change occurs. N₂(l) → N₂(g)
    In the direction written, the phase change is
    a) spontaneous at all temperatures
    b) more spontaneous at low temperatures
    c) more spontaneous at high temperatures
    d) can not be determined from the information given

11. (4 pts) How many electrons are transferred in the following reaction?
    5 H₂O₂ + 2 IO₃⁻ + 2 H⁺ → I₂ + 5 O₂(g) + 6 H₂O
12. In class we observed the Briggs-Rauber Reaction, causing the solution to switch colors back and forth from blue to clear for several minutes. The reactions that cause the color change are as follows.

\[
\begin{align*}
5 \text{H}_2\text{O}_2 + 2 \text{IO}_3^- + 2 \text{H}^+ & \rightarrow \text{I}_2 + 5 \text{O}_2 (g) + 6 \text{H}_2\text{O} \\
5 \text{H}_2\text{O}_2 + \text{I}_2 & \rightarrow 2 \text{IO}_3^- + 2 \text{H}^+ + 4 \text{H}_2\text{O}
\end{align*}
\]

\[10 \text{H}_2\text{O}_2 (l) \rightarrow 5 \text{O}_2 (g) + 10 \text{H}_2\text{O} (l)\]

a) (4 pts) Identify the substance that is reduced in the first reaction.

b) (4 pts) Identify the substance that is oxidized in the second reaction.

13. Given the following reaction mechanism,

\[
\begin{align*}
\text{H}_2\text{O}_2 & \rightarrow \text{H}_2\text{O} + \text{O} \\
\text{O} + \text{CF}_2\text{Cl}_2 & \rightarrow \text{ClO} + \text{CF}_2\text{Cl} \\
\text{ClO} + \text{O}_3 & \rightarrow \text{Cl} + 2 \text{O}_2 \\
\text{Cl} + \text{CF}_2\text{Cl} & \rightarrow \text{CF}_2\text{Cl}_2
\end{align*}
\]

a) (3 pts) Write the overall equation for the reaction?

b) (3 pts) Identify the reaction intermediate(s).

c) (3 pts) Identify the catalyts(s).

14. (10 pts) One mole of an ideal gas is heated from 25°C to 35°C at constant pressure.

Circle the correct answer for each of the following quantities: \( w \), \( q \), \( \Delta H \), \( \Delta E \), and \( \Delta S \).

\[
\begin{align*}
w < 0 & \quad w > 0 & \quad w = 0 \\
q < 0 & \quad q > 0 & \quad q = 0 \\
\Delta H < 0 & \quad \Delta H > 0 & \quad \Delta H = 0 \\
\Delta E < 0 & \quad \Delta E > 0 & \quad \Delta E = 0 \\
\Delta S < 0 & \quad \Delta S > 0 & \quad \Delta S = 0
\end{align*}
\]
15. Answer the following questions using the data given below:

<table>
<thead>
<tr>
<th>Half Reaction</th>
<th>$E^\circ$ (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{Pb}^{2+}$ + 2 e$^-$ $\longrightarrow$ Pb (s)</td>
<td>$-0.125$</td>
</tr>
<tr>
<td>$\text{Sn}^{2+}$ + 2 e$^-$ $\longrightarrow$ Sn (s)</td>
<td>$-0.137$</td>
</tr>
</tbody>
</table>

Consider a galvanic cell where Sn(s) is oxidized to Sn$^{2+}$ and Pb$^{2+}$ is reduced to Pb (s) at 25°C.

Sn (s) | Sn$^{2+}$ | $\mid$ | Pb$^{2+}$ | Pb (s)  
Initial concentrations: $[\text{Sn}^{2+}] = 0.075 \text{ M}$ and $[\text{Pb}^{2+}] = 0.600 \text{ M}$

a) (6 pts) Calculate the initial cell potential, $E_{\text{cell}}$, when the concentrations are: $[\text{Sn}^{2+}] = 0.075 \text{ M}$ and $[\text{Pb}^{2+}] = 0.600 \text{ M}$

b) (4 pts) If the cell is allowed to operate spontaneously, will $E_{\text{cell}}$ increase, remain the same or decrease. Circle the answer.

increase  remain the same  decrease

c) (8 pts) If the cell is allowed to operate spontaneously, what will $E_{\text{cell}}$ be when $[\text{Pb}^{2+}]$ has decreased to 0.500 M?

d) (8 pts) If the cell is allowed to operate spontaneously, what will $[\text{Sn}^{2+}]$ be at the point at which $E_{\text{cell}} = 0.020 \text{ V}$?
16. The hydrolysis of table sugar (sucrose) to produce glucose and sucrose occurs by the following overall reaction.

\[ \text{C}_{12}\text{H}_{22}\text{O}_{11} (s) + \text{H}_2\text{O} (l) \rightarrow \text{C}_6\text{H}_12\text{O}_6 (s) + \text{C}_6\text{H}_{12}\text{O}_6 (s) \]

The following data and the rate law for this reaction were determined. \( \text{rate} = k [\text{C}_{12}\text{H}_{22}\text{O}_{11}] \)

<table>
<thead>
<tr>
<th>[C_{12}H_{22}O_{11}] (mol/L)</th>
<th>time (hr)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.451</td>
<td>0.50</td>
</tr>
<tr>
<td>0.353</td>
<td>1.50</td>
</tr>
</tbody>
</table>

a) (6 pts) Use this data to determine the rate constant for this reaction.

b) (6 pts) How long does it take to hydrolyze 75% of the sucrose?

17. Consider the following reaction.

\[ \text{NO} (g) + \text{Cl}_2 (g) \rightarrow \text{NOCl} (g) + \text{Cl} (g) \]

This reaction is endothermic, \( \Delta H^\circ = 83 \text{ kJ/mol} \) and the forward activation energy, \( E_a \), is 88 kJ/mol.

a) (6 pts) Draw a potential energy profile for this reaction. Your diagram should show the reactants, products, transition state \( [X]^\dagger \), \( E_a \) and \( \Delta H^\circ \).

\[ \begin{array}{c}
\text{Energy} \\
\mid \\
\longrightarrow \\
\text{Reaction Coordinate}
\end{array} \]

b) (4 pts) What is the activation energy for the reverse reaction: \( \text{NOCl} (g) + \text{Cl} (g) \rightarrow \text{NO} (g) + \text{Cl}_2 (g) \)
1. Using the following data:

\[ \begin{array}{c|c}
\text{Substance} & \Delta G^\circ \text{ (kJ/mol)} \\
\hline
\text{N}_2\text{H}_4 (l) & 149 \\
\text{NO}_2 (g) & 52 \\
\text{H}_2\text{O} (l) & -237 \\
\text{O}_2 (g) & 0 \\
\end{array} \]

a) (6 pts) Calculate \( \Delta G^\circ \) for the following reaction at 298 K: \( \text{N}_2\text{H}_4 (l) + 3 \text{ O}_2 (g) \rightarrow 2 \text{ NO}_2 (g) + 2 \text{ H}_2\text{O} (l) \).

b) (6 pts) Predict the sign of \( \Delta S \) and \( \Delta H \). Circle the answer.

\[ \begin{array}{ccc}
\Delta S & < 0 & > 0 & = 0 \\
\Delta H & < 0 & > 0 & = 0 \\
\end{array} \]

c) (3 pts) Will the reaction be more spontaneous at higher or lower temperature? Circle the answer.

Higher T  Lower T

2. (8 pts) The standard enthalpy of combustion of 1 mol of ethanol, \( \text{CH}_3\text{CH}_2\text{OH} (l) \), is \(-1368\) kJ at 298 K. Given the following enthalpies of formation, calculate \( \Delta H_f^\circ \) for \( \text{CH}_3\text{CH}_2\text{OH} (l) \).

\[ \begin{array}{c|c}
\text{Substance} & \Delta H_f^\circ \text{ (kJ/mol)} \\
\hline
\text{H}_2\text{O} (l) & -286 \\
\text{CO}_2 (g) & -394 \\
\end{array} \]

3. (6 pts) Write the reactions that correspond to the following enthalpy changes:

a) \( \Delta H_f^\circ \) for solid iron(III) oxide.

b) \( \Delta H_f^\circ \) for calcium carbonate, \( \text{CaCO}_3 \).
4. (4 pts) Write the electronic configuration and indicate the number of unpaired electrons for each of the following species.

<table>
<thead>
<tr>
<th>electronic configuration</th>
<th>number of unpaired electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>a) S</td>
<td></td>
</tr>
<tr>
<td>b) Se⁻</td>
<td></td>
</tr>
</tbody>
</table>

5. a) (6 pts) Draw the Lewis structures, all three resonance forms, for the cyanate ion, NCO⁻, assign formal charges to all atoms. Carbon is the central atom.

b) (4 pts) Based on the formal charges, which resonance structure is least favored?

6. (6 pts) Draw the Lewis structure for each of the following molecules:

   - SCl₃
   - N₂F₄

7. (16 pts) Circle the formula that best fits each of the following descriptions:
   a) smallest radius
      \[ \text{O}^+ \quad \text{O}^- \quad \text{O} \]
   b) largest ionic radius
      \[ \text{Cl}^- \quad \text{S}^{2-} \quad \text{K}^+ \quad \text{Ca}^{2+} \]
   c) most polar bond
      \[ \text{H} - \text{F} \quad \text{C} - \text{O} \quad \text{H} - \text{Cl} \]
   d) longest bond length
      \[ \text{F}_2 \quad \text{I}_2 \quad \text{Br}_2 \quad \text{Cl}_2 \]
   e) greatest electronegativity
      \[ \text{Si} \quad \text{N} \quad \text{B} \quad \text{H} \]
   f) smallest first ionization energy
      \[ \text{Mg} \quad \text{P} \quad \text{S} \quad \text{Cl} \]
   g) contains the longest oxygen-oxygen bond length
      \[ \text{O}_2 \quad \text{H}_2\text{O}_2 \quad \text{O}_3 \]
   h) an excited state electron configuration
      \[ \text{[Ar]}4s^23d^5 \quad 1s^22s^22p^4 \quad \text{[Ar]}4s^23d^44p^1 \]
8. (4 pts) At constant pressure, the following reaction is exothermic. 
\[ 2 \text{NO}_2(g) \rightarrow \text{N}_2\text{O}_4(g) \]
In the direction written, the reaction is 
\( a) \) always spontaneous. 
\( b) \) spontaneous at low temperatures 
\( c) \) spontaneous at high temperatures 
\( d) \) never spontaneous.

9. (4 pts) For which of the following processes would \( \Delta S^\circ \) be expected to be most positive? 
\( a) \) \( 4 \text{ClO}_2(g) \rightarrow 2 \text{Cl}_2\text{O}(g) + 3 \text{O}_2(g) \) 
\( b) \) \( \text{I}_2\text{O}(l) \rightarrow \text{I}_2\text{O}(s) \) 
\( c) \) \( \text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(g) \) 
\( d) \) \( \text{N}_2\text{O}_4(g) \rightarrow 2 \text{NO}_2(g) \) 
\( e) \) \( 2 \text{N}_2\text{O}_5(g) \rightarrow 4 \text{NO}_2(g) + \text{O}_2(g) \)

10. (4 pts) How many electrons are transferred in the following reaction? 
\[ 6 \text{Ag} + 3 \text{HS}^- + 2 \text{CrO}_4^{2-} + 5 \text{H}_2\text{O} \rightarrow 3 \text{Ag}_2\text{S} + 2 \text{Cr(OH)}_3 + 7 \text{OH}^- \]

11. (10 pts) Answer the following question using the data given below:

\[
\begin{array}{c|c|c}
\text{Half Reaction} & E^\circ (V) \\
\hline
\text{Co}^{2+} + 2 \text{e}^- \rightarrow \text{Co} (s) & -0.28 \\
\text{Ni}^{2+} + 2 \text{e}^- \rightarrow \text{Ni} (s) & -0.23 \\
\end{array}
\]
A piece of cobalt is placed in 0.15 M Ni(NO_3)_2. Calculate the concentrations of Co^{2+} and Ni^{2+} at equilibrium, at 25°C.
12. (6 pts) One mole of an ideal gas is heated from 25°C to 35°C under two different conditions. First, the gas is heated from 25°C to 35°C at constant volume then the gas is heated from 25°C to 35°C at constant pressure.
   a) What is the heat absorbed at constant volume, \( q_v \), relative to the heat absorbed at constant pressure, \( q_p \)? Circle the answer.
      \[ q_v = q_p \quad q_v > q_p \quad q_v < q_p \]
   b) What is the change in enthalpy relative to the change in energy for these processes? Circle the answer.
      \[ \Delta H = \Delta E \quad \Delta H > \Delta E \quad \Delta H < \Delta E \]

13. Consider the following reaction mechanism:
   \[ O_3 \rightleftharpoons O_2 + O \quad \text{(fast equilibrium)} \]
   \[ O + O_3 \rightarrow 2 O_2 \quad \text{(slow)} \]
   a) (2 pts) What is the overall reaction?

   b) (8 pts) Derive the rate law. Be sure to eliminate intermediates from your answer.

   c) (4 pts) If the concentration of \( O_2 \) is doubled, while the concentration of \( O_3 \) is kept constant, how will this affect the rate of reaction?

14. (8 pts) In a common car battery, six identical cells each carry out the following reaction.
   \[ \text{Pb (s)} + \text{PbO}_2 + \text{HSO}_4^- + 2 \text{H}^+ \rightarrow 2 \text{PbSO}_4 + 2 \text{H}_2\text{O} \quad E^o = 2.04 \text{V} \]
   Suppose that to start a car on a cold morning, 125 A is drawn for 15.0 seconds from this cell. How many grams of \( \text{Pb (s)} \) would be consumed?
15. The molecular formula for ozone is O₃. We can draw two possible Lewis structures that satisfy the octet rule.

a) (2 pts) Show all the valence electrons in the following Lewis structures for ozone?

\[ \text{O} \quad \text{O} \quad \text{O} \]

b) (3 pts) Which structure, if any, is favored (more stable, lower energy)? Circle the answer.

i. Both are equally favored.  
ii. The cyclic structure is favored.  
iii. The bent structure is favored.

16. The rate of the reaction, \( \text{NO}_2 (g) + \text{CO} (g) \rightarrow \text{NO} (g) + \text{CO}_2 (g) \)
depends only on the concentration of nitrogen dioxide at 200°C. The following data were collected at 200°C.

<table>
<thead>
<tr>
<th>Time (min)</th>
<th>([\text{NO}_2]) (mol/L)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>0.500</td>
</tr>
<tr>
<td>20</td>
<td>0.444</td>
</tr>
<tr>
<td>50</td>
<td>0.381</td>
</tr>
<tr>
<td>75</td>
<td>0.340</td>
</tr>
<tr>
<td>150</td>
<td>0.250</td>
</tr>
<tr>
<td>300</td>
<td>0.174</td>
</tr>
</tbody>
</table>

a) (10 pts) If the reaction is second order with respect to nitrogen dioxide, determine the value of the rate constant at this temperature. Be sure to include units.

b) (4 pts) Calculate the concentration of nitrogen dioxide, \([\text{NO}_2]\), after 450.0 minutes.
17. (10 pts) Use the bond energies given to predict $\Delta H$ for the combustion of ethanol.

$$
\begin{align*}
\text{H} & \quad \text{H} \\
\text{H} - \text{C} - \text{C} - \text{O} - \text{H} & + \quad 3 \text{ O}_2 \quad \rightarrow \quad 2 \text{ CO}_2 \quad + \quad 3 \text{ H}_2\text{O}
\end{align*}
$$

<table>
<thead>
<tr>
<th>Bond Energy (kJ/mol)</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>C=O</td>
<td>358</td>
</tr>
<tr>
<td>C=O</td>
<td>799</td>
</tr>
<tr>
<td>C-C</td>
<td>347</td>
</tr>
<tr>
<td>O=O</td>
<td>495</td>
</tr>
<tr>
<td>C-H</td>
<td>413</td>
</tr>
<tr>
<td>O-H</td>
<td>463</td>
</tr>
</tbody>
</table>

18. The figure below represents part of the emission spectrum for a one-electron ion in the gas phase. All the lines result from electronic transitions from excited states to the $n = 3$ state.

```
A
B

Wavelength
```

a) (6 pts) What electronic transitions correspond to lines A and B?

Line A:

Line B:

b) (10 pts) If the wavelength of line B is 121.6 nm, calculate the wavelength of A.