Chem 111  2:30p section  Final Exam

This exam is composed of 50 questions, 14 of which require mathematics that require a calculator. Go initially through the exam and answer the questions you can answer quickly. Then go back and try the ones that are more challenging to you and/or that require calculations.

As discussed in the course syllabus, honesty and integrity are absolute essentials for this class. In fairness to others, dishonest behavior will be dealt with to the full extent of University regulations.

I hereby state that all answers on this exam are my own and that I have neither gained unfairly from others nor have I assisted others in obtaining an unfair advantage on this exam.

Signature


PERIODIC TABLE OF THE ELEMENTS

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- $PV = nRT$
- $E = hf = \frac{hc}{\lambda}$
- $\frac{E}{n} = E_a = 3RT$
- $K.E. = \frac{1}{2}mv^2$
- $R_H = 1.0974 \times 10^7 m^{-1}$
- $1 mL = 1 cm^3$
- $1 atm = 760 mm Hg$
- $\Delta H_{vap} (H_2O) = 40.65 \text{ kJ mol}^{-1}$
- $\Delta H_{fus} (H_2O) = 6.00 \text{ kJ mol}^{-1}$
- $d_{\text{water}} = 1.00 \text{ g mL}^{-1}$
- $\Delta E = q + w = \Delta H - P\Delta V$

- $h = 6.626 \times 10^{-34} \text{ J s}$
- $c = 2.998 \times 10^8 \text{ m s}^{-1}$
- $R = 0.0820 \text{ L atm K}^{-1} \text{ mol}^{-1}$
- $R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}$
- $J = kg m^2 s^{-2}$

__________________________

Name: _______________________________
Solubility Rules for some ionic compounds in water

Soluble Ionic Compounds
1. All sodium (Na⁺), potassium (K⁺), and ammonium (NH₄⁺) salts are SOLUBLE.
2. All nitrate (NO₃⁻), acetate (CH₃CO₂⁻), chlorate (ClO₃⁻), and perchlorate (ClO₄⁻) salts are SOLUBLE.
3. All chloride (Cl⁻), bromide (Br⁻), and iodide (I⁻) salts are SOLUBLE -- EXCEPT those also containing: lead, silver, or mercury (I) (Pb²⁺, Ag⁺, Hg₂⁺) which are NOT soluble.
4. All sulfate (SO₄²⁻) salts are SOLUBLE -- EXCEPT those also containing: calcium, silver, mercury (I), strontium, barium, or lead (Ca²⁺, Ag⁺, Hg₂⁺, Sr²⁺, Ba²⁺, Pb²⁺) which are NOT soluble.

Not Soluble Ionic Compounds
5. Hydroxide (OH⁻) and oxide (O²⁻) compounds are NOT SOLUBLE -- EXCEPT those also containing: sodium, potassium, or barium (Na⁺, K⁺, Ba²⁺) which are soluble.
6. Sulfide (S²⁻) salts are NOT SOLUBLE -- EXCEPT those also containing: sodium, potassium, ammonium, or barium (Na⁺, K⁺, NH₄⁺, Ba²⁺) which are soluble.
7. Carbonate (CO₃²⁻) and phosphate (PO₄³⁻) salts are NOT SOLUBLE -- EXCEPT those also containing: sodium, potassium, or ammonium (Na⁺, K⁺, NH₄⁺), which are soluble.

Some common ions:

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<th>PO₄³⁻</th>
<th>CN⁻</th>
<th>CH₃CO₂⁻</th>
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<td>SO₄²⁻</td>
<td>CrO₄²⁻</td>
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Bond Dissociation Energies (kJ mol⁻¹) (gas phase)

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<td>I-I</td>
<td>151</td>
<td>N-I</td>
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1. In an exothermic process:
   1) work is performed on the surroundings
   2) heat is transferred to the surroundings
   3) work is performed on the system
   4) heat is transferred to the system

2. A positive value of $\Delta E$ means that:
   1) heat is transferred to the surroundings
   2) heat is transferred to the system
   3) energy in the form of heat and/or work is transferred to the surroundings
   4) energy in the form of heat and/or work is transferred to the system

3. An automobile engine generates 2757 Joules of heat that must be carried away by the cooling system. The internal energy changes by -3852 Joules in this process. How much work to push the pistons is available in this process?
   1) 4918 J  2) 1095 J  3) 683 J  4) 6283 J  5) 1277 J

4. What is the minimum amount of ice at 0°C that must be added to the contents of a can of diet cola (340 mL) to cool it from 20°C to 5°C? Assume that diet cola has the properties of pure water ($\Delta H_{\text{fusion}}^{\text{water}} = 333 \text{ J g}^{-1}$, $d_{\text{water}} = 1.0 \text{ g mL}^{-1}$)
   1) 34.2 g  2) 64.1 g  3) 87.6 g  4) 10.2 g  5) 125 g
5 Given the following information:

\[ 2 \text{N}_2\text{O} (g) + 3 \text{O}_2 (g) \rightarrow 2 \text{N}_2\text{O}_4 (g) \quad \Delta H^\circ = -145.8 \text{ kJ} \]

\[ 2 \text{N}_2\text{O} (g) \rightarrow 2 \text{N}_2 (g) + \text{O}_2 (g) \quad \Delta H^\circ = -164.2 \text{ kJ} \]

what is the standard enthalpy change for the reaction:

\[ \text{N}_2 (g) + 2 \text{O}_2 (g) \rightarrow \text{N}_2\text{O}_4 (g) \quad \Delta H^\circ = ?? \text{ kJ} \]

1) 155 kJ mol\(^{-1}\)  
2) -146 kJ mol\(^{-1}\)  
3) 9.2 kJ mol\(^{-1}\)  
4) 146 kJ mol\(^{-1}\)  
5) not enough information to determine

6 The root mean square speed of molecules in a sample of F\(_2\) gas is 890 m/s. What is the temperature of the gas?

1) 513 K  
2) 890 K  
3) 127 K  
4) 1208 K  
5) 233 K

7 A 2.38 mol sample of He gas is confined in a 62.5 liter container at 62.5 °C. If 1.28 mol of F\(_2\) gas is added while maintaining constant temperature, the average kinetic energy per molecule will:

1) decrease  
2) remain the same  
3) increase  
4) not enough information  
5) I don’t have a clue
8. Which listing below correctly orders the molecules by increasing root mean square molecular speed (slowest → fastest)?

1) CO₂ < Ar < N₂ < H₂  
2) Ar < CO₂ < N₂ < H₂  
3) H₂ < N₂ < CO₂ < Ar  
4) H₂ < N₂ < Ar < CO₂

9. A sample of Cl₂ gas is confined in a 2.0 liter container at 50 °C. Then 2.5 mol of He is added, holding both the volume and temperature constant. The pressure will increase because:

1) With more molecules per unit volume, there are more molecules hitting the walls of the container.
2) As the number of molecule-wall collisions increases, the force per collision increases.
3) With more molecules in the container, the molecules have higher average speeds.
4) With higher average speeds, on average the molecules hit the walls of the container with more force.
5) None of the Above

10. What is the average kinetic energy of an N₂ molecule confined in 3.1 L at 1.0 atm and 25°C?

1) 5.71x10³ J  
2) 9.48x10³ J  
3) 6.17x10⁻²¹ J  
4) 5.71x10⁻²¹ J  
5) 3.21x10⁻²¹ J
Consider the molecular orbital energy diagram shown at right.

11. The energy level denoted “b” refers to:
   1) a bonding molecular orbital
   2) an antibonding molecular orbital
   3) a nonbonding molecular orbital
   4) an atomic orbital

12. The electrons in the orbital represented by energy level “b”:
   1) are distributed more toward X  
   2) are distributed more toward Y  
   3) are equally distributed between X and Y

13. The molecule XY is the diatomic (He-H)⁺. What is its bond order?
   1) 0.0  
   2) 0.5  
   3) 1.0  
   4) 1.5  
   5) 2.0

14. What is the energy of ultraviolet light with frequency 1.36x10¹⁵ Hz?
   1) 126 kJ mol⁻¹  
   2) 196 kJ mol⁻¹  
   3) 427 kJ mol⁻¹  
   4) 544 kJ mol⁻¹  
   5) 832 kJ mol⁻¹

15. Consider two cases for emission from the hydrogen atom:
   **Case 1:** 
   Electron goes from n=4 to n=3
   **Case 2:** 
   Electron goes from n=6 to n=2

   Compare the energies of the photons emitted:
   1) E_{case 1} > E_{case 2}  
   2) E_{case 1} < E_{case 2}  
   3) E_{case 1} = E_{case 2}
16. Consider the energy vs temperature diagram at right, describing the transitions of water from ice to steam:

The segment labeled (d) is described best with which parameter below:

1) \( \Delta H^\circ_{\text{fus}} \)  
2) \( \Delta H^\circ_{\text{vap}} \)  
3) \( C_{\text{ice}} \)  
4) \( C_{\text{liquid}} \)  
5) \( C_{\text{steam}} \)

17. The following information is given for bismuth, Bi, at 1 atm:

- boiling pt = 1627°C  
  \( H_{\text{vap}}^{1627^\circ \text{C},1\text{atm}} = 172 \text{ kJ mol}^{-1} \)  
  \( C_{\text{liquid Bi}} = 0.151 \text{ J g}^{-1} \text{ K}^{-1} \)

- melting pt = 271°C  
  \( H_{\text{fus}}^{271^\circ \text{C},1\text{atm}} = 11.0 \text{ kJ mol}^{-1} \)  
  \( C_{\text{solid Bi}} = 0.126 \text{ J g}^{-1} \text{ K}^{-1} \)

At a pressure of 1 atm, what amount of heat is needed to melt a 35.2 g sample of solid bismuth at its normal melting point of 271 °C?

1) 4.21 kJ  
2) 13.8 kJ  
3) 0.561 kJ  
4) 9.67 kJ  
5) 1.85 kJ

18. At a pressure of 1 atm, what amount of heat is needed to take a 35.2 g sample of bismuth from 200°C to 400°C?

1) 2.85 kJ  
2) 15.4 kJ  
3) 32.6 kJ  
4) 9.67 kJ  
5) 14.3 kJ
19. Which ion has the smallest radius?
   1) K⁺  2) Ca²⁺  3) P³⁻  4) S²⁻  5) all the same

20. Consider the following samples:
   a) 0.531 moles of CH₄ in a 6.18 L container at a temperature of 308K
   b) 0.281 moles of CH₄ in a 2.77 L container at a temperature of 388K
   c) 0.569 moles of CH₄ in a 1.42 L container at a temperature of 453K
   d) 0.212 moles of CH₄ in a 5.95 L container at a temperature of 298K

   Which has the lowest average molecular speed?
   1) a  2) b  3) c  4) d  5) all the same

21. HNO₃ is (data at the front of the exam provide a clue):
   1) a strong base  2) a weak acid  3) a weak base
   4) a strong acid  5) none of the above

22. Reactions in water that produce gases tend to be:
   1) favorable  2) ugly  3) unfavorable
   4) exothermic  5) endothermic

23. Which reaction below is a redox reaction?
   1) NaOH (aq) + HNO₃ (aq) → NaNO₃ (aq) + H₂O (l)
   2) Na₂CO₃ (aq) + 2 HClO₄ (aq) → CO₂ (g) + H₂O (l) + 2NaClO₄
   3) CdCl₂ (aq) + Na₂S (aq) → CdS (s) + 2 NaCl (aq)
   4) Zn(OH)₂ (s) + H₂SO₄ (aq) → ZnSO₄ (aq) + 2 H₂O (l)
   5) None of the above

24. The net ionic equation for the reaction of zinc sulfate and sodium hydroxide is:
   1) Zn²⁺ (aq) + 2 OH⁻ (aq) → Zn(OH)₂ (s) + Na₂SO₄ (aq)
   2) ZnSO₄ (aq) + 2 NaOH (aq) → Zn(OH)₂ (aq) + Na₂SO₄ (aq)
   3) Zn²⁺ (aq) + 2 OH⁻ (aq) → Zn(OH)₂ (s)
   4) Zn²⁺ (aq) + 2 OH⁻ (aq) → Zn(OH)₂ (aq)
   5) No net reaction occurs
25. Which element has the lowest ionization energy?
   1) In  2) Ga  3) Tl  4) B  5) all the same

26. Draw the Lewis structure for \( \text{CO}^{2-} \). What is the hybridization on oxygen?
   1) \( \text{sp}^3 \text{d} \)  2) \( \text{sp}^4 \)  3) \( \text{sp}^3 \)  4) \( \text{sp}^2 \)  5) \( \text{sp} \)

27. Draw the Lewis structure for \( \text{XeOF}_4 \) (Xe is the central atom). What is the hybridization on \( \text{Xe} \)?
   1) \( \text{sp}^3 \text{d}^2 \)  2) \( \text{sp}^3 \text{d}^3 \)  3) \( \text{sp}^3 \text{d} \)  4) \( \text{sp}^3 \)  5) \( \text{sp}^2 \)

28. The molecule \( \text{XeOF}_4 \) is:
   1) polar  2) nonpolar  3) can’t tell

29. The correct molecular formula for the molecule at right is:
   1) \( \text{C}_2\text{O}_2\text{H}_4 \)  2) \( \text{CO}_2\text{H}_4 \)  3) \( \text{C}_2\text{OH}_4 \)  4) \( \text{C}_2\text{O}_2\text{H}_3 \)

30. A specific isotope of an ion from a given element has 8 protons, 7 neutrons, and 10 electrons. The ion is:
   1) \( \text{O}^{2-} \)  2) \( \text{Ne}^{3-} \)  3) \( \text{P}^{3-} \)  4) \( \text{N}^{3-} \)  5) \( \text{Mn}^{3+} \)

31. What is the formula of the ionic compound formed in the reaction of elemental \( \text{Mg} \) and \( \text{O}_2 \)?
   1) \( \text{MgO} \)  2) \( \text{Mg}_2\text{O} \)  3) \( \text{Mg}_2\text{O}_3 \)  4) \( \text{Mg}_3\text{O}_2 \)  5) \( \text{MgO}_2 \)

32. What is the (mass) percent composition of \( \text{C} \) in \( \text{C}_4\text{H}_8 \)?
   1) 85.6%  2) 14.4%  3) 50.0%  4) 88.3%  5) 11.7%
33. What is the wavelength of ultraviolet light with frequency $1.18 \times 10^{15}$ Hz?
   1) 209 nm  2) 254 nm  3) 280 nm  4) 190 nm  5) 350 nm

34. What is the maximum number of orbitals that can be identified by the set of quantum numbers $n=+5 \ l=+3$?
   1) 2  2) 3  3) 5  4) 6  5) 7

35. Consider the molecule ClF$_4^-$ How many lone pairs are on the central atom?
   1) 1  2) 2  3) 3  4) 6  5) 0

36. Light is given off by a sodium or mercury containing street light when the atoms are excited. The light you see arises for which of the following reasons?
   1) Electrons are moving from a given energy level to one of lower $n$
   2) Electrons are moving from a given energy level to one of higher $n$
   3) Electrons are being removed from the atom, thereby creating a metal cation

37. Consider the molecule ClF$_4^-$ What is the electron pair geometry?
   1) Trigonal bipyramidal  2) Octahedral  3) linear
   4) Trigonal planer  5) Tetrahedral
38. Which of the following has the highest affinity for electrons?
   1) B  2) N  3) As  4) P  5) Ge

39. In ionizing elemental lithium to Li⁺, from which orbital is an electron removed?
   1) 1s  2) 2s  3) 3s  4) 2p  5) 3p

40. In the molecule formaldehyde CH₂O, what is the approximate HCO bond angle?
   1) 180°  2) 90°  3) 109°  4) 120°  5) 60°

41. If you completely react 0.678 g of iodine (I₂), what mass of NI₃ can be produced?
   1) 0.276 g  2) 0.678 g  3) 0.226 g  4) 0.876 g  5) 0.351 g
42 Nitrogen triiodide (NI$_3$) is unstable, reacting to form N$_2$ (g) and I$_2$ (g), and evolving heat.

\[ 2 \text{NI}_3 \ (s) \rightarrow \text{N}_2 \ (g) + 3 \text{I}_2 \ (g) \]

Spontaneous decomposition of 0.5 g of NI$_3$ (s) produces what volume of gas at 200°C and 1 atm pressure?

1) 28.7 L  2) 0.197 L  3) 0.098 L  4) 14.4 L  5) 0.731 L

43 Using the Table of Bond Dissociation Energies at the front of the exam, predict ΔH° for the spontaneous decomposition of nitrogen triiodide above.

1) -256 kJ mol$^{-1}$  2) -927 kJ mol$^{-1}$  3) -384 kJ mol$^{-1}$
4) -35 kJ mol$^{-1}$  5) +927 kJ mol$^{-1}$

44 What is the molecular geometry of nitrogen triiodide?

1) tetrahedral  2) square planar  3) trigonal planar
4) octahedral  5) trigonal pyramidal

45 What is the hybridization on N in nitrogen triiodide?

1) sp  2) sp$^2$  3) sp$^3$  4) sp$^4$  5) sp$^3$d
46 Which do you expect to have the longest bond length?
1) NI₃  2) NBr₃  3) NCl₃  4) NF₃  5) can’t tell

47 In class, we saw the following reaction (unbalanced).

\[ \text{Al (s) + Br}_2 (l) \rightarrow \text{AlBr}_3 (s) \]

In the correctly balanced reaction, what is the stoichiometry coefficient preceding \( \text{Br}_2 \) (all coefficients should be integral)?
1) 1  2) 2  3) 3  4) 4  5) 6

48 In the reaction above of aluminum and bromine, which is the oxidizing agent?
1) Al (s)  2) Br₂ (l)

49 What is the electron pair geometry in AlBr₃?
1) tetrahedral  2) square planar  3) trigonal planar
4) octahedral  5) trigonal pyramidal

50 What is the catalog number for this class?
1) 123  2) 111  3) 86  4) 345  5) 68.6 g