City College, Chemistry Department
Chemistry 10301, sections T and T2, Prof. T. Lazaridis
Final exam, Dec 21, 2006
Name (last name first): _____
I.D. Number: _____
Workshop leader: _____

Note: There are 18 questions in this exam. Fill in your answer in the blank space provided immediately following each question. 1/2 point will be subtracted every time you report a numerical result with an incorrect number of significant figures.

A Data Sheet with useful information is at the end.

1. (4) Write the names of the elements below next to their atomic symbols:

Ga **Gallium**

Co **Cobalt**

Sn **Tin**

Na **Sodium**

2. (4) Write the molecular formula next to the names of the following compounds:

Ammonium Bromide **NH_4Br**

Potassium Chlorate **$KClO_3$**

Magnesium Sulfite **$MgSO_3$**

Sodium Phosphate **$(Na)_3PO_4$**

3. (8) Balance the following chemical equations:

$CaCO_3 + H_2O + CO_2 \rightarrow Ca(HCO_3)_2$ **(already balanced)**

2 $AgNO_3 + K_2CO_3 \rightarrow Ag_2CO_3 +$ **2** KNO_3

$XeF_6 +$ **3** $H_2O \rightarrow XeO_3 +$ **6** HF

2 $H_2S +$ **3** $O_2 \rightarrow$ **2** $SO_2 +$ **2** H_2O

4. (8) How many grams of oxygen are required to burn 75.0 g of ammonia? The reaction is

$$4\ NH_3 + 5\ O_2 \rightarrow 4\ NO + 6\ H_2O$$

Molar mass of ammonia: 14.00+3X1.008= 17.02 g/mol

Molar mass of O_2 : 16.00X2 = 32.00 g/mol

moles of NH_3 : 75.0 g/ 17.02 (g/mol) = 4.41 mol

moles of O_2 required: 4.41 mol NH_3 X 5 mol O_2/ 4 mol NH3 = 5.51 mol

convert to grams: 5.51 mol X 32.00 g/mol = 176 g O_2

5. (5) A 2.122 g sample of an unknown cobalt compound is dissolved in 50.0 mL of water. The concentration of Co^{2+} (aq) is found to be 0.071 M. Calculate the percentage by mass of cobalt in the unknown compound.

Atomic weight of Co: 58.93 g/mol

moles of Co in the solution: 0.071 mol/L X 50.0 X 10^{-3} L = 0.0036 mol

convert to gram: 0.0036 mol X 58.93 g/mol = 0.21 g

0.21/ 2.122 X 100% = 9.9 %

6. (6) How many mL of oxygen, measured at 0 °C and 1.00 atm, are required for the complete combustion of 5.00 g of liquid methanol (CH_3OH)?

$2\ CH_3OH + 3\ O_2 \rightarrow 2\ CO_2 + 4\ H_2O$

Molar mass of methanol: 12.01 + 4X1.008 + 16.00= 32.04 g /mol

Convert to moles: 5.00 g / 32.04 (g/mol) = 0.156 mol methanol

Moles O_2 required: 0.156 mol CH_3OH X 3 mol O_2 / 2 mol CH_3OH = 0.234 mol O_2

Use Ideal Gas Law to get the volume

PV= nRT => V = nRT/P = 0.234 mol X 0.082 L atm/ (mol K) X 273.15 K / 1 atm =

5.24 L

that is 5240 mL

7. (7) Use Hess's law to calculate the standard enthalpy for the reaction

$C_2H_2\ (g) + 2\ H_2\ (g) \rightarrow C_2H_6\ (g)$

using the following combustion data:

$C_2H_2\ (g) + 5/2\ O_2\ (g) \rightarrow 2CO_2\ (g) + H_2O\ (l)$	$\Delta H^o = -1300$ kJ
$H_2\ (g) + 1/2\ O_2\ (g) \rightarrow H_2O\ (l)$	$\Delta H^o = -286$ kJ
$C_2H_6\ (g) + 7/2\ O_2\ (g) \rightarrow 2CO_2\ (g) + 3H_2O\ (l)$	$\Delta H^o = -1560$ kJ

Flip the third equation, multiply the second by 2, and then add all three to get the desired reaction.

ΔH^o = -1300 + 2 X (-286) + 1560 = -312 kJ

8. (6) Use enthalpy of formation data to compute standard enthalpy changes for the following reactions at 25 ^{o}C. State whether the reactions are endothermic or exothermic.

a) NH_3 (g) + HCl (g) → NH_4Cl (s)

-314.4 – (-46.11) – (-92.31) = -176.0 kJ exothermic

b) C_2H_4 (g) + H_2 (g) → C_2H_6 (g)

-84.68 – 52.26 - 0 = -136.9 kJ exothermic

9. (4) Which of the following sets of quantum numbers are allowed, and which are not? State what is wrong with each set that is not allowed.

a) n=0	l=0	m_l=0	**NOT (n must be greater or equal to 1)**
b) n=1	l=0	m_l=0	**OK**
a) n=6	l=4	m_l=-5	**NOT (m_l goes from –l to +l)**
a) n=6	l=5	m_l=-4	**OK**

10. (4) A hygrogen electron drops from the n=4 level to the n=1 level. What is the wavelength of the emitted photon?

$E_n = -2.179 \times 10^{-18}$ J/n^2

$\Delta E = -2.179 \times 10^{-18}$ J X ($1/1^2 - 1/4^2$) = - 2.043×10^{-18} J

2.043×10^{-18} J = hv = h c/λ =>

λ = 6.626×10^{-34} Js X 2.9979×10^{8} m/s / 2.043×10^{-18} J

= 9.723×10^{-8} m = 97.23 nm

11. (4) Write an abbreviated (noble gas core) electron configuration for the following:

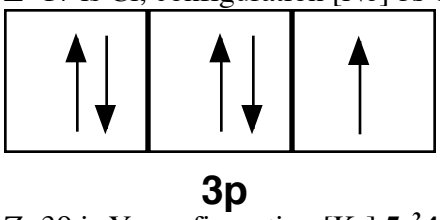
C	**[He] $2s^2 2p^2$**
Cl	**[Ne] $3s^2 3p^5$**
Te	**[Kr] $5s^2 4d^{10} 5p^4$**
Ge^{2+}	**[Ar] $4s^2 3d^{10}$**

12. (2) Place the atoms K, N, As in order of increasing atomic radius.

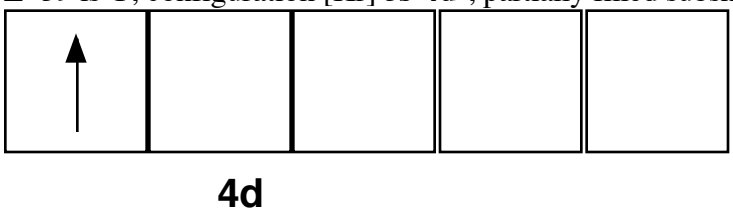
N < As < K

13. (4) Draw orbital diagrams for the partially filled subshells of the atoms with Z=17 and Z=39.

Z=17 is Cl, configuration [Ne] **$3s^2 3p^5$**, partially filled subshell: 3p

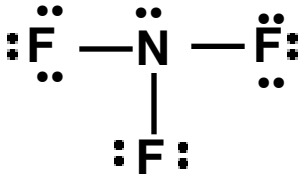


Z=39 is Y, configuration [Kr] **$5s^2 4d^1$**, partially filled subshell: 4d



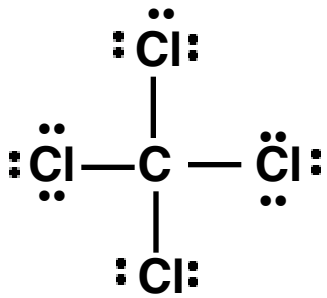
14. (9) Draw Lewis structures for the following molecules and give the molecular geometry and type of hybridization around each central atom.

a) NF_3



4 electron groups, molecular geometry: trigonal pyramidal, hybridization sp^3

(b) CCl_4



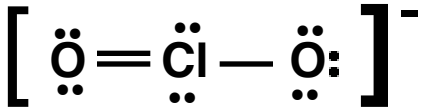
4 electron groups, molecular geometry: tetrahedral, hybridization sp^3

(c) HCN

2 electron groups, molecular geometry: linear, hybridization sp

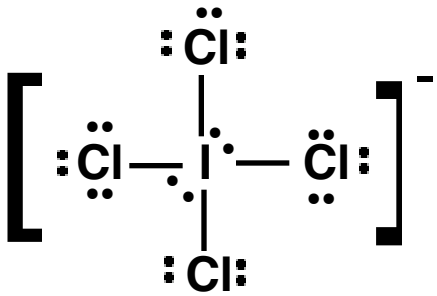
15. (10) Use Lewis structures and the VSEPR method to predict the molecular geometry and the bond angles of the following ions:

a) ClO_2^-



4 electron groups, molecular geometry bent (angular), bond angle close to 109.5 º

b) ICl_4^-



6 electron groups, molecular geometry square planar, bond angles close to 90 º

16. (5) Use MO theory to determine the bond order for N_2^+. Is N_2^+ more or less stable than N_2 ?

Look at figure 10.27. If we remove one electron from N_2, it is removed from a bonding orbital, so the molecule becomes LESS stable.

Bond order of N_2^+ = 1/2 (bonding electrons – antibonding electrons) =

= 1/2 (7 – 2) = 2.5

17. (5) Give the oxidation number of each element in the following compounds or ions:

a) H_3BO_3 **+1 for H, -2 for O, +3 for B**

b) BrF_3 **-1 for F, +3 for Br**

c) NH_3OH^+ **+1 for H, -2 for O, -1 for N**

d) AlO_2^- **-2 for O, +3 for Al**

e) $(CH_3)_2S$ **+1 for H, -2 for S, -2 for C**

18. (5) The specific heat of copper is 0.385 J/(g °C) and that of gold 0.129 J/(g °C). A 1.00-g sample of either metal adsorbs 10.0 J of heat. If both metals were originally at 25.0 °C, what will be their final temperature?

q = sp.heat X mass X ΔT => ΔT= q / (sp.heatXmass) =

10.0 J/ [0.385 J/(g °C) X 1g] = 26.0 °C for copper, final T = 25.0+26.0 = 51.0 °C

10.0 J/ [0.129 J/(g °C) X 1g] = 77.5 °C for gold, final T = 25.0+77.5 = 102.5 °C

DATA SHEET

Speed of light : 2.9979 X 10^8 m/s

Planck's constant : 6.626 X 10^{-34} Js

E_n = -2.179 X 10^{-18} J/n^2

Mass of a proton: 1.67262 X 10^{-24} g

Mass of an electron: 9.10939 X 10^{-28} g

R= 0.082058 L atm / (mol K) = 62.364 L torr / (mol K) = 8.3145 J / (mol K)

Substance	ΔH_f^o, kJ/mol
NH_3 (g)	-46.11
NH_4Cl (s)	-314.4
HCl (g)	-92.31
C_2H_4 (g)	52.26
C_2H_6 (g)	-84.68