 Name: ________________________________

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

$PV = nRT \quad N_o = 6.022x10^{23}\,mol^{-1}$	$1\ mL = 1\ cm^3$	$h = 6.626x10^{-34}\ J\ s$
$E = h\nu = \frac{hc}{\lambda} \quad \overline{u^2} = \frac{3RT}{M} \quad \overline{K.E.} = \frac{1}{2}m\overline{u^2}$	$1\ atm = 760\ mm\ Hg$	$c = 2.998x10^8\ m\ s^{-1}$
$E_n^{H-atom} = -\frac{R_H hc}{n^2} \quad R_H hc = 1312\ kJ\ mol^{-1}$	$\Delta H_{vap}(H_2O) = 40.65\ kJ\ mol^{-1}$	$R = 0.0820\ L\ atm\ K^{-1}\ mol^{-1}$
$R_H = 1.0974x10^7\ m^{-1}$	$\Delta H_{fus}(H_2O) = 6.00\ kJ\ mol^{-1}$	$R = 8.314\ J\ K^{-1}\ mol^{-1}$
	$d_{water} = 1.00\ g\ mL^{-1}$	$J = kg\ m^2\ s^{-2}$
	$\Delta E = q + w = \Delta H - P\Delta V$	

PERIODIC TABLE OF THE ELEMENTS

1A	2A	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	3A	4A	5A	6A	7A	8A
1 **H** 1.008																	2 **He** 4.003
3 **Li** 6.939	4 **Be** 9.012											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.90	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.71	29 **Cu** 63.55	30 **Zn** 65.39	31 **Ga** 69.72	32 **Ge** 72.61	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (99)	44 **Ru** 101.1	45 **Rh** 102.9	46 **Pd** 106.4	47 **Ag** 107.9	48 **Cd** 112.4	49 **In** 114.8	50 **Sn** 118.7	51 **Sb** 121.8	52 **Te** 127.6	53 **I** 126.9	54 **Xe** 131.3
55 **Cs** 132.9	56 **Ba** 137.3	57 **La** 138.9	72 **Hf** 178.5	73 **Ta** 181.0	74 **W** 183.8	75 **Re** 186.2	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.1	79 **Au** 197.0	80 **Hg** 200.6	81 **Tl** 204.4	82 **Pb** 207.2	83 **Bi** 209.0	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** 226.0	89 **Ac** 227.0	104 **Unq** (261)	105 **Unp** (262)	106 **Unh** (263)	107 **Uns** (262)	108 **Uno** (265)	109 **Une** (266)									

Solubility Rules for some ionic compounds in water

Soluble Ionic Compounds

1. All sodium (Na^+), potassium (K^+), and ammonium (NH_4^+) salts are SOLUBLE.
2. All nitrate (NO_3^-), acetate ($CH_3CO_2^-$), chlorate (ClO_3^-), and perchlorate (ClO_4^-) salts are SOLUBLE.
3. All chloride (Cl^-), bromide (Br^-), and iodide (I^-) salts are SOLUBLE -- EXCEPT those also containing: lead, silver, or mercury (I) (Pb^{2+},Ag^+, Hg^{2+}) which are NOT soluble.
4. All sulfate (SO_4^{2-}) salts are SOLUBLE - - EXCEPT those also containing: calcium, silver, mercury (I), strontium, barium, or lead (Ca^{2+}, Ag^+, Hg_2^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}) which are NOT soluble.

Not Soluble Ionic Compounds

5. Hydroxide (OH^-) and oxide (O^{2-}) compounds are NOT SOLUBLE -- EXCEPT those also containing: sodium, potassium, or barium (Na^+, K^+, Ba^{2+}) which are soluble.
6. Sulfide (S^{2-}) salts are NOT SOLUBLE -- EXCEPT those also containing: sodium, potassium, ammonium, or barium (Na^+, K^+, NH_4^+, Ba^{2+}) which are soluble.
7. Carbonate (CO_3^{2-}) and phosphate (PO_4^{3-}) salts are NOT SOLUBLE -- EXCEPT those also containing: sodium, potassium, or ammonium (Na^+, K^+, NH_4^+), which are soluble.

Some common ions:

PO_4^{3-} CN^- $CH_3CO_2^-$ NO_2^- NO_3^-

CO_3^{2-} SO_3^{2-} SO_4^{2-} CrO_4^{2-} MnO_4^-

Bond Dissociation Energies (kJ mol^{-1}) (gas phase)

Bond	D	Bond	D	Bond	D
H-H	436	C-C	346	N-N	163
C-H	413	C=C	610	N=N	418
C≡O	1046	C≡N	887	N≡N	945
N-H	391	O-O	146	C-O	358
O-H	463	O=O	498	C=O	745
C-F	485	F-F	155	N-F	283
C-Cl	339	Cl-Cl	242	N-Cl	192
C-I	213	I-I	151	N-I	169

1. In an exothermic process:
   1) work is performed on the system
   2) heat is transferred to the system
   3) work is performed on the surroundings
   4) heat is transferred to the surroundings

   (4)

Ch 6.1 – Energy: basic principles

2. A negative value of ΔE means that:
   1) heat is tranferred to the surroundings
   2) heat is transfered 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)

Ch 6.4 – Energy: 1st Law of Thermo

3. An automobile engine generates **2575** Joules of heat that must be carried away by the cooling system. The internal energy changes by **–3258** Joules in this process.

   How much work to push the pistons is available in this process?

   1) 4918 J 2) 5833 J 3) 683 J 4) 6283 J 5) 1277 J

   $$\Delta E = q + w \qquad w = \Delta E - q = (-3258\ J) - (-2575\ J) = -683\ J$$

   (3) w is negative. The system does work on the surroundings.

Ch 6.4 – Energy: 1st Law of Thermo

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 ($\Delta H^{water}_{fusion} = 333\ J\,g^{-1}$ $d_{water} = 1.0\ g\,mL^{-1}$)

   1) 34.2 g 2) 64.1 g 3) 87.6 g 4) 10.2 g 5) 125 g

   $$m_{cola} = (340 mL)(1.00\ g\,mL^{-1}) = 340g \qquad m_{ice} = ?$$

   $$\Delta T = [(20+273)-(5+273)]K = -15K$$

   (2) $$\Delta H^{water}_{fusion} = 333\ J\,g^{-1}$$

   $$q_{cola} + q_{ice} = 0 = [C_{cola} m_{cola} \Delta T] + [m_{ice} \Delta H^{water}_{fusion}]$$

   $$m_{ice} \Delta H^{water}_{fusion} = -C_{cola} m_{cola} \Delta T$$

   $$m_{ice} = \frac{-C_{cola} m_{cola} \Delta T}{\Delta H^{water}_{fusion}} = \frac{-(4.184\ J\,g^{-1})(340g)(-15K)}{333\ J\,g^{-1}} = 64.1g$$

   (5)

Ch 6.3 – Specific heat capacity and phase change, see Example 6.4 on page 249

5. Given the following information:

   $2\ N_2O\ (g) + 3\ O_2\ (g) \rightarrow 2\ N_2O_4\ (g)$ ΔH° = -145.8

   $2\ N_2O\ (g) \rightarrow 2\ N_2\ (g) + O_2\ (g)$ ΔH° = -164.2 kJ

   what is the standard enthalpy change for the reaction:

   $N_2\ (g) + 2\ O_2\ (g) \rightarrow N_2O_4\ (g)$ ΔH° = ?? kJ

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

   $N_2\ (g) + 1/2\ O_2\ (g) \rightarrow N_2O\ (g)$ ΔH° = -1/2 * -164.2 kJ

   $N_2O\ (g) + 3/2\ O_2\ (g) \rightarrow N_2O_4\ (g)$ ΔH° = 1/2 * -145.8 kJ

   (1) ΔH° = +9.2 kJ

Ch 6.7 – Hess's Law (Owl, Unit 6-6c)

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

$$\overline{u^2} = \frac{3RT}{M}$$

(4)

$$T = \frac{M\sqrt{\overline{u^2}}^{\,2}}{3\ R} = \frac{38.0\ g\ mol^{-1}\left(890\ m\ s^{-1}\right)^2}{3\left(8.314\ J\ K^{-1}\ mol^{-1}\right)}\frac{J}{kg\ m^2\ s^{-2}}\frac{kg}{10^3\ g} = 1208K$$

Ch 12.6 – Kinetic theory of gases

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

   (2) Temperature determines average kinetic energy (Chapter 12)

Ch 12.6 – Kinetic theory of gases

8d. Which listing below correctly orders the molecules by increasing root mean square molecular speed (slowest → fastest)?

1) $CO_2 < Xe < N_2 < H_2$ 2) $Xe < CO_2 < N_2 < H_2$

3) $H_2 < N_2 < CO_2 < Xe$ 4) $H_2 < N_2 < Xe < CO_2$

(2) $\sqrt{\overline{u^2}} = \sqrt{\frac{3RT}{M}}$ **Molar masses: 131 > 48 > 28 > 2 (OWL 12-x)**

Ch 12.6 – kinetic theory, rms speed and molar mass.

9. A sample of Cl_2 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) As the number of molecule-wall collisions increases, the force per collision increases.
2) With more molecules per unit volume, there are more molecules hitting the walls of the container.
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

(2) (Chapter 12)

Ch 12.6 – Kinetic theory of gases

10. What is the average kinetic energy of an N_2 molecule confined in 3.1 L at 1.0 atm and 25°C?

1) $5.71\text{x}10^{3}$ J 2) $9.48\text{x}10^{3}$ J 3) $5.71\text{x}10^{-21}$ J 4) $6.17\text{x}10^{-21}$ J 5) $3.21\text{x}10^{-21}$ J

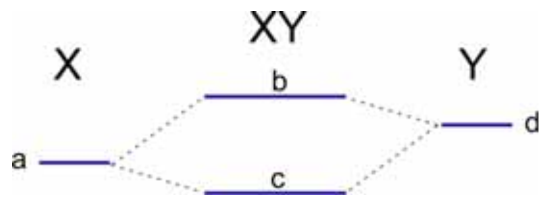
(4)

$$\overline{K.E.} = \frac{1}{2} m\overline{u^2} \text{ (per molecule)} \qquad \overline{u^2} = \frac{3RT}{M}$$

$$\overline{K.E.} = \frac{1}{2} m \frac{3RT}{M}\left(\frac{M\ N_o^{-1}}{m}\right) = \frac{3}{2}\frac{RT}{N_o} = \frac{3}{2}\frac{\left(8.314\ J\,K^{-1}\ mol^{-1}\right)\left(25+273\right)K}{\left(6.022\text{x}10^{23}\,mol^{-1}\right)} = 6.17x10^{-21}\,J$$

Ch 12.6 – Kinetic theory of gases

Consider the molecular orbital energy diagram shown at right.



11. The energy level denoted "**c**" refers to:

1) a bonding molecular orbital

2) an antibonding molecular orbital

3) a nonbonding molecular orbital

4) an atomic orbital

(1) **(OWL question)**

Ch 10.3 – basic concepts of molecular orbitals

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

(2) **(OWL question)**

Ch 10, but also 8, 9 – concepts of electronegativities and energy. Covered in class.

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

(2)

Ch 10, but also 8, 9 – concepts of electronegativities and energy. Covered in class.

14d. What is the energy of near IR light with frequency 3.16×10^{14} 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}

(1) $E = h\nu = (6.626x10^{-34}\ J\ s)(3.16x10^{14}\ Hz)\left(\frac{s^{-1}}{Hz}\right)(6.022x10^{23}\ photons\ mol^{-1}) = 126000\ J\ mol^{-1}$ **(OWL question)**

Ch 7.2 - light and energy.

15d. 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) $E_{case\ 1} = E_{case\ 2}$

$$E_n^{H-atom} = -\frac{R_H hc}{n^2} \qquad E_n^{H-atom} \propto -\frac{1}{n^2}$$

$$\therefore \quad \Delta E = E_f^{H-atom} - E_i^{H-atom} \propto -\frac{1}{n_f^2} - \left(-\frac{1}{n_i^2}\right) = \frac{1}{n_i^2} - \left(\frac{1}{n_f^2}\right)$$

$$\Delta E_{case1} \propto \frac{1}{3^2} - \left(\frac{1}{4^2}\right) = 20.6 \qquad \Delta E_{case2} \propto \frac{1}{2^2} - \left(\frac{1}{6^2}\right) = 4.5$$

(1)

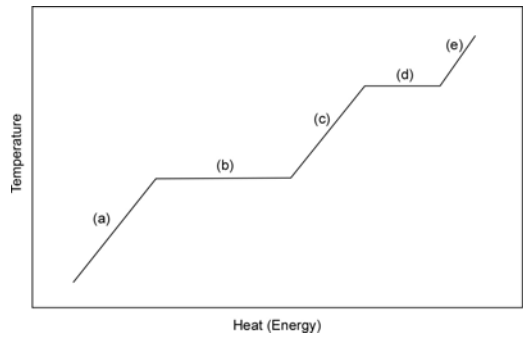
Ch 7.3 – hydrogen atom and Rydberg.

16d. Consider the energy vs temperature diagram at right, describing the transitions of water from ice to steam:

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

1) $\Delta H°_{fus}$ 2) $\Delta H°_{vap}$ 3) C_{ice}

4) C_{liquid} 5) C_{steam}



(4) heat capacity of liquid water

Ch 6.x – phase changes and heat capacities.

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

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

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

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

(5) $q = nH_{fus}^{271°C,1atm} = \frac{m}{M} H_{fus}^{271°C,1atm} = \left(35.2 \frac{mol}{209.0g}\right)\left(11.0\ kJ\ mol^{-1}\right) = 1.85kJ$

Ch 6.x – phase changes and heat capacities.

18d. 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

(1)

$$q = \frac{m}{M} C_{solid\ Bi} + \frac{m}{M} H_{fus}^{271°C,1atm} + \frac{m}{M} C_{liquid\ Bi}$$

$$= \left[0.126\ J\ g^{-1}\ K^{-1}(35.2g)(271-200)K\right]\left(\frac{kJ}{10^3\ J}\right) + 1.85kJ + \left[0.151\ J\ g^{-1}\ K^{-1}(35.2g)(400-271)K\right]\left(\frac{kJ}{10^3\ J}\right)$$

$$= (0.315 + 1.85 + 0.686)kJ = 2.85kJ$$

Ch 6.x – phase changes and heat capacities.

19d. Which ion has the smallest radius?

1) K^+ 2) Ca^{2+} 3) P^{3-} 4) S^{2-} 5) all the same

(2) – all are isoelectronic. Ca^{2+} has the largest nuclear charge, therefore attracts its electrons the most (OWL 8-12c)

Ch 8.x – ionic radii trends

20d. Consider the following samples:

a) 0.212 moles of CH_4 in a 5.95 L container at a temperature of 298K
b) 0.531 moles of CH_4 in a 6.18 L container at a temperature of 308K
c) 0.281 moles of CH_4 in a 2.77 L container at a temperature of 388K
d) 0.569 moles of CH_4 in a 1.42 L container at a temperature of 453K

Which has the lowest average molecular speed?

1) a 2) b 3) c 4) d 5) all the same

(1) $\sqrt{\overline{u^2}} = \sqrt{\frac{3RT}{M}}$ **M all the same; lowest T, lowest rms speed**

Ch 12.6 – kinetic theory, rms speed and molar mass.

21d. HNO_3 is (data at the front of the exam provide a clue):

1) a strong base 2) a weak base 3) a weak acid
4) a strong acid 5) none of the above

(4) **Chapter 5**

Ch 5.3 – Acids, but also solubility

22d. Reactions in water that produce gases tend to be:

1) unfavorable 2) ugly 3) favorable
4) endothermic 5) exothermic

(3) **Chapter 5**

Ch 5.5 – Gas forming rxns, but also Ch 6 concepts

23. Which reaction below is a redox reaction?

1) $NaOH$ (aq) + HNO_3 (aq) → $NaNO_3$ (aq) + H_2O (l)

2) Na_2CO_3 (aq) + 2 $HClO_4$ (aq) → CO_2 (g) + H_2O (l) + $2NaClO_4$

3) $CdCl_2$ (aq) + Na_2S (aq) → CdS (s) + 2 $NaCl$ (aq)

4) $Zn(OH)_2$ (s) + H_2SO_4 (aq) → $ZnSO_4$ (aq) + 2 H_2O (l)

5) None of the above

(5) Look at redox changes – there are none. **Chapt 5 inspired by book**

Ch 5.7 – Redox

24. The net ionic equation for the reaction of zinc sulfate and sodium hydroxide is:

1) Zn^{2+} (aq) + 2 OH^- (aq) → $Zn(OH)_2$ (s) + Na_2SO_4 (aq)

2) $ZnSO_4$ (aq) + 2 $NaOH$ (aq) → $Zn(OH)_2$ (aq) + Na_2SO_4 (aq)

3) Zn^{2+} (aq) + 2 OH^- (aq) → $Zn(OH)_2$ (s)

4) Zn^{2+} (aq) + 2 OH^- (aq) → $Zn(OH)_2$ (aq)

5) No *net* reaction occurs

(3) hydroxide salts are generally insoluble **(OWL 5-2c)**

Ch 5.2 – Precipitation rxns

25d. Which element has the lowest ionization energy?

1) In 2) Ga 3) Tl 4) B 5) all the same

(3) – IE trends (OWL 8-9b)

Ch 8.x – ionization energy trends

26d. Draw the Lewis structure for $\mathbf{CO^{2-}}$. What is the hybridization on oxygen?

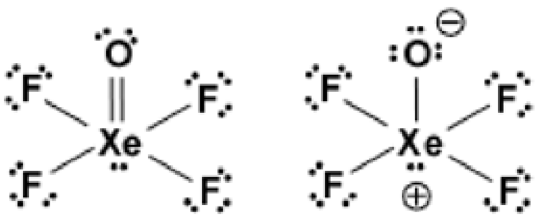
1) sp^4 2) sp^3 3) sp^2 4) sp 5) sp^3d

(3) sp^2  **(12 valence electrons) OWL 9-xx & 10**

Ch 10.x and 9 – Lewis structures and hybridization

27d. Draw the Lewis structure for $\mathbf{XeOF_4}$ (Xe is the central atom). What is the hybridization on **Xe**?

1) sp^2 2) sp^3 3) sp^3d 4) sp^3d^2 5) sp^3d^3

(4)  **OWL 9-xx**

Ch 10.2 & 9.7 – Hybridization

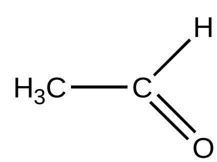
28d. The molecule $\mathbf{XeOF_4}$ is:

1) polar 2) nonpolar 3) can't tell

(1) polar – the individual dipoles do not cancel out. OWL 9-10b

Ch 9.9 – Polarity

29d. The correct molecular formula for the molecule at right is:



1) CO_2H_4 2) $C_2O_2H_4$ 3) $C_2O_2H_3$ 4) C_2OH_4

(4)

Ch 3 – molecular formulas

30d. A specific isotope of an ion from a given element has 7 protons, 8 neutrons, and 10 electrons. The ion is:

1) Mn^{3+} 2) N^{3-} 3) P^{3-} 4) Ne^{3-} 5) O^{2-}

(2) **(from an OWL question 3-3c)**

Ch 2.3 – atomic composition

31d. What is the formula of the ionic compound formed in the reaction of elemental **Cs** and $\mathbf{F_2}$?

1) CsF_2 2) Cs_2F 3) Cs_2F_3 4) CsF 5) Cs_3F_2

(4) **CsF - $Cs^+ + F^-$** **(OWL question)**

Ch 3.3 – ionic compounds

32d. What is the (mass) percent composition of **C** in $\mathbf{C_4H_8}$?

1) 88.3% 2) 85.6% 3) 50.0% 4) 14.4% 5) 11.7%

Mass of C in 1 mol of the compound: $(4\,mol)(12.01\ g\ mol^{-1}) = 48.04g$

Mass of 1 mol of the compound:

$(1mol)[4(12.011\ g\ mol^{-1}) + 8(1.0079\ g\ mol^{-1})] = 56.11\text{g}$

(2) **Percent composition:** $\frac{48.04g\ \text{C}}{56.11g\ \text{C}_4\text{H}_8}100\% = 85.6\%$ **(OWL question)**

Ch 3.6 – percent composition

33d. What is the wavelength of ultraviolet light with frequency $1.58\text{x}10^{15}$ Hz?

1) 209 nm 2) 254 nm 3) 280 nm 4) 190 nm 5) 350 nm

(4) $\lambda = \left(\frac{2.9998x10^8 m}{s}\right)\left(\frac{1}{1.58x10^{15}\,Hz}\right)\left(\frac{Hz}{s^{-1}}\right)\left(\frac{10^9\,nm}{m}\right) = 190nm$ **(OWL question)**

Ch 7.1 – wavelength & frequency

34d. What is the maximum number of orbitals that can be identified by the set of quantum numbers n=+6 *l*=+3 ?

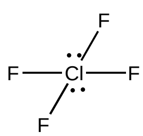
1) 7 2) 6 3) 5 4) 3 5) 2

(1) for *l* = 3, one can have m_l = -3, -2, -1, 0, +1, +2, +3 (7 orbitals)

Ch 7.5 – quantum numbers

35d. Consider the molecule ClF_4^- How many lone **pairs** are on the central atom?

1) 1 2) 2 3) 3 4) 6 5) 0



(2)

Ch 9.6 – octet rule beyond 2nd row

36d. 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 being removed from the atom, thereby creating a metal cation

3) Electrons are moving from a given energy level to one of higher n

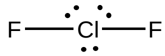
(1) **(end of chapter question)**

Ch 7.3 – atomic energy levels

37d. Consider the molecule ClF_2^- What is the electron pair geometry?

1) Trigonal bipyramidal 2) Octahedral 3) linear

4) Trigonal planer 5) Tetrahedral

(1) 

Ch 9.7 – electron pair geometry

38d. Which of the following has the highest affinity for electrons?

1) Ge 2) P 3) N 4) As 5) B

(3) (OWL 8-11)

Ch 8.6 – electron affinity

39d. In ionizing elemental lithium to Li^+, from which orbital is an electron removed?

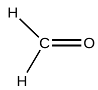
1) 1s 2) 2s 3) 3s 4) 2p 5) 3p

(2) (OWL 8-11)

Ch 8.x – electron configuration and ionization

40d. In the molecule **formaldehyde** CH_2O, what is the approximate HCO bond angle?

1) 120° 2) 109° 3) 90° 4) 180° 5) 60°

(1) 

- trigonal planar at the C

Ch 9.7 – molecular geometry

As we demonstrated in class, reaction of iodine (I_2) and aqueous ammonia (NH_3) produces nitrogen triiodide (NI_3) according to the following reaction:

$$3\ I_2\ (s) + 4\ NH_4OH\ (aq) \rightarrow NI_3\ (s) + 3\ NH_4I\ (aq) + 4\ H_2O$$

41. If you completely react 0.678 g of iodine (I_2), what mass of NI_3 can be produced?

1) 0.276 g 2) 0.351 g 3) 0.226 g 4) 0.876 g 5) 0.678 g

$$M_{I_2} = 2\left(126.9\ g\ mol^{-1}\right) = 253.8\ g\ mol^{-1} \quad n_{I_2} = \frac{0.678g}{253.8\ g\ mol^{-1}} = 2.67x10^{-3}\ mol$$

(2)

$$n_{NI_3} = \frac{1}{3} n_{I_2} = 8.90x10^{-4}\ mol \quad M_{NI_3} = 14.01 + 3\left(126.9\ g\ mol^{-1}\right) = 394.7\ g\ mol^{-1}$$

$$m_{NI_3} = n_{NI_3}\left(M_{NI_3}\right) = \left(8.90x10^{-4}\ mol\right)\left(394.7\ g\ mol^{-1}\right) = 0.351\ g$$

Ch 3 - Stoichiometry

42. Nitrogen triiodide (NI_3) is unstable, reacting to form N_2 (g) and I_2 (g), and evolving heat.

$$2\ NI_3\ (s) \rightarrow N_2\ (g) + 3\ I_2\ (g)$$

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) 0.731 L 3) 14.4 L 4) 0.098 L 5) 0.197 L

$$M = \left(14.0 + 3(126.9)\right) g\ mol^{-1} = 394.7\ g\ mol^{-1}$$

(4)

$$n_{NI_3} = \frac{0.5g}{394.7\ g\ mol^{-1}} = 1.27x10^{-3} mol$$

$$n_{gas} = n_{N_2} + n_{I_2} = \frac{1}{2} n_{NI_3} + \frac{3}{2} n_{NI_3} = 2n_{NI_3} = 2.53x10^{-3} mol$$

$$V = \frac{nRT}{P} = \frac{\left(2.53x10^{-3} mol\right)\left(0.082057\ L\ atm\ K^{-1}\ mol^{-1}\right)\left((200+273)K\right)}{1\ atm} = 0.0983\ L$$

Ch 3 & 12 – Stoichiometry and gases

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

(2)

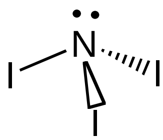
$$\Delta H° = \sum D_{(Bonds\ broken)} - \sum D_{(Bonds\ formed)}$$

$$\Delta H° = \left\{\left[6(169)\right] - \left[945 + 3(151)\right]\right\} kJ = (1014 - 1398) kJ = -384\ kJ$$

Ch 9.10 – Bond properties

44. What is the molecular geometry of nitrogen triiodide?

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



(5)

Ch 9.7 – Molecular Shapes

45. What is the hybridization on N in nitrogen trioiodide?

1) sp^3d
2) sp^4
3) sp^3
4) sp^2
5) sp

(3)

Ch 10.2 – Orbital hybridization

46. Which do you expect to have the shortest bond length?

1) NF_3
2) NCl_3
3) NBr_3
4) NI_3
5) can’t tell

(1)

Ch 8.6 – Atomic properties/trends

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

$$Al\ (s) + Br_2\ (l) \rightarrow AlBr_3\ (s)$$

In the correctly balanced reaction, what is the stoichiometry coefficient preceding Br_2 (all coefficients should be integral)?

1) 1 2) 2 3) 3 4) 4 5) 6

(3) 2 Al (s) + 3 Br_2 (l) → 2 $AlBr_3$ (s)

Ch 4 (but everywhere!) – Balancing reactions

48. In the reaction above of aluminum and bromine, which is the reducing agent?

1) Al (s) 2) Br_2 (l)

(1) Al is oxidized, therefore it is the reducing agent

Ch 5.7 - Redox reactions

49. What is the electron pair geometry in $AlBr_3$?

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

(2)

Ch 9.7 – Molecular Shapes

50. What is the catalog number for this class?

1) 123 2) 345 3) 86 4) 111 5) 68.6 g

(4)