CHEM 1314 3;30 pm Theory
Exam I
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September 11, 2002

Name _____

TA's Name _____

Lab Section _____

INSTRUCTIONS:

1. This examination consists of a total of 6 different pages. The last page include a periodic table and some useful equations. All work should be done in this booklet.

2. PRINT your name, TA's name and your lab section number now in the space at the top of this sheet. DO NOT SEPARATE THESE PAGES.

3. Answer all questions that you can and whenever called for show your work clearly. Your method of solving problems should pattern the approach used in lecture.

4. No credit will be awarded if your work is not shown in problems 2, 4 and 7a.

5. Point values are shown next to the problem number.

6. Budget your time for each of the questions. Some problems may have a low point value yet be very challenging. If you do not recognize the solution to a question quickly, skip it, and return to the question after completing the easier problems.

7. Look through the exam before beginning; plan your work; then begin.

8. Relax and do well.

	Page 2	Page 2	Page 3	Page 4	Page 5	TOTAL
SCORES	_____	_____	_____	_____	_____	_____
	(1)	(30)	(24)	(29)	(16)	(100)

(12) 1. Write the chemical formula(s) of the product(s) and balance all of the following reactions. Identify all products phases as either (g)as, (l)iquid, (s)olid or (aq)ueous

a) $3Mg_{(s)} + N_{2(g)} \rightarrow \mathbf{Mg_3N_{2(s)}}$

b) $S_{8(g)} + 8O_{2(g)} \rightarrow \mathbf{8SO_{2(g)}}$

c) $\mathbf{2}C_6H_{14(l)} + \mathbf{19}O_{2(g)} \rightarrow \mathbf{12CO_{2(g)} + 14H_2O_{(g)}}$

d) $\mathbf{2}K_{(s)} + Cl_{2(g)} \rightarrow \mathbf{2KCl_{(s)}}$

Grading: 3 points for the correct products, balance and phases. If the products are <u>all</u> correct, –1 for balance and –1 for phases

(18) 2. Perform the following conversions. Be sure your answer has the correct number of significant figures.

a) 400. cm^3 to liters

$$\mathbf{400\ cm^3\left(\frac{1\ mL}{1\ cm^3}\right)\left(\frac{1\ L}{1000\ mL}\right) = 0.400\ L}$$

b) the volume of a ball is 2.00 ft^3. Convert to km^3. (use at least 3 conversions factors.)

$$\mathbf{2.00\ ft^3\left(\frac{12\ in}{1\ foot}\right)^3\left(\frac{2.54\ cm}{1\ inch}\right)^3\left(\frac{1\ m}{100\ cm}\right)^3\left(\frac{1\ km}{1000\ m}\right)^3 = 5.66\ x\ 10^{-11}\ km}$$

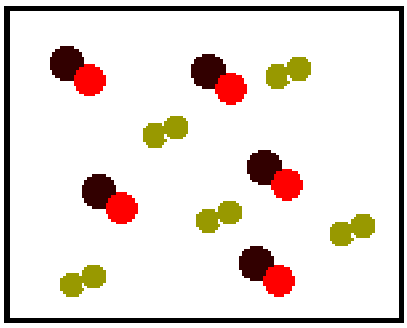
c) What is 10.0 Kelvin on the Fahrenheit scale?

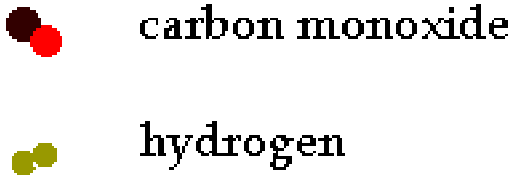
$$\mathbf{^\circ C = \frac{5}{9}(^\circ F - 32) = \frac{5}{9}(98 - 32) = \frac{5}{9}(66) = 36.7}$$

$$\mathbf{K = ^\circ C + 273.15 = 36.7 + 273.15 = 309.85 = 310.\ K}$$

Grading: each part is worth 6 points. Deduct 1 point for sig fig, 1 point for units, 1 point for math error. Deduct 3 points for a 'simple' conversion error. If more than 1 conversion error deduct all points.

(6) 3. Diagram the following system as viewed at the atomic level in the space provided. Be sure to clearly label each of the substances in your diagram.





A homogeneous mixture of carbon monoxide and hydrogen at room temperature.

Grading: 2 points each for the correct diatomic for hydrogen and carbon monoxide. 1 point for gas phase and 1 point for homogeneous mixture

(18) 4. A particular homogeneous mixture of ethanol (C_2H_6O) and water has a density of 0.945 g · ml^{-1}. This ethanol solution is also 36.0% (by mass) ethanol. The density of pure ethanol is 0.789 g · ml^{-1}.

a) A sample of this ethanol solution has a mass of 250. g, calculate the volume of the solution.

$$\textbf{250. g ethanol solution}\left(\frac{\textbf{1 mL ethanol solution}}{\textbf{0.945 g ethanol solution}}\right)\textbf{ = 265 mL of ethanol solution}$$

b) How many grams of water and ethanol are in the sample in a)?

$$\textbf{250. g ethanol solution}\left(\frac{\textbf{36.0 g pure ethanol}}{\textbf{100 g ethanol solution}}\right)\textbf{ = 90.0 g of pure ethanol}$$

250. grams – 90.0 grams = 160. grams of water

c) Calculate the volumes of pure ethanol and water that must be mixed to prepare the solution in part a)?

$$\textbf{90.0 g pure ethanol}\left(\frac{\textbf{1 mL pure ethanol}}{\textbf{0.789 g pure ethanol}}\right)\textbf{ = 114 mL of pure ethanol}$$

$$\textbf{160.0 g pure water}\left(\frac{\textbf{1 mL pure water}}{\textbf{1.00 g pure water}}\right)\textbf{ = 160 mL of pure water}$$

d) What is interesting about the answer in part a) and the answer(s) in part c)?

To prepare the solution of ethanol and water we would mix 114 mLs ethanol with 160 mLs of water. What is interesting is the volume of the <u>mixture</u> is smaller than the volume of the pure liquids added together. So upon mixing there is a contraction of the total volume. Why this happens is discussed in CHEM 1515.

Grading: Parts a, b and c are worth 5 points each. Pretty much R/W. 1 point for the correct units in part a). Part d) is worth 3 points. It was relatively easy to get 1 point (assuming parts a and c were correct), a little harder to get two points and 3 points were pretty difficult.

(6) 5. Predict a reasonable formula for the compound formed from each of the following combinations of elements or polyatomic ions.

a) sodium and sulfate **Na_2SO_4**

b) iron and bromide **$FeBr_2$ or $FeBr_3$**

c) carbon and fluorine **CF_4**

Grading: R/W 2 points each

(11) 6. Complete the following table;

Name of the compound	Formula of the compound	Ionic or Covalent Compound
sodium sulfide	**Na_2S**	**ionic**
Phosphorus pentachloride	**PCl_5**	**covalent**
Nickel (II) phosphate	$Ni_3(PO_4)_2$	**ionic**
Potassium peroxide		**ionic**
Nitric acid	$H_2SO_{4(aq)}$	
octane	C_8H_{18}	**covalent**

Grading: R/W 1 point each. Spelling and symbols had to be perfect for the point.

(12) 7. The relative weighted average atomic mass of gallium is 69.723 u. Gallium has two naturally occurring isotopes. One of those isotopes has a mass of 70.9247 u and a percent abundance of 39.89%.

a) Calculate the mass of the other isotope of gallium.

The relative average atomic mass of gallium from the periodic table is 69.723 u. The fractional abundance of the unknown Ga isotope must be 1 - 0.3989 = 0.6011, because only two isotopes of Ga exist.

Average atomic mass = Σ(mass$_{isotope}$ · fractional abundance$_{isotope}$)

69.723 u = (70.9247 u)(0.3989) + *x*(0.6011)

69.723 u = 28.292 + *x*(0.6011)

41.431 u = *x*(0.6011)

***x* = 68.93 u**

b) How many neutrons, protons and electrons are in each of the isotopes of gallium?

Isotope	Protons	Neutrons	Electrons
$^{71}_{31}$**Ga**	**31**	**40**	**31**
$^{69}_{31}$**Ga**	**31**	**38**	**31**

(6) 8. Complete each calculation and report the answer to the correct number of significant figures.

a) $\left(\frac{3.05 \times 10^{20}}{8.340 \times 10^{24}}\right) = \mathbf{3.66 \times 10^{-5}}$

b) $\left(\frac{1}{65}\right) + \left(\frac{1}{1234}\right) = \mathbf{0.1538 + 0.00081037 = 0.016}$

c) $(8.34 \cdot 0.450) - \frac{2.001}{0.83} = \mathbf{3.753 - 2.411 = 1.3}$

grading: R/W 2 points each

(10) 9. Complete the following table

Formula	*M*, Molar Mass $\left(\frac{g}{mol}\right)$	*m*, mass of sample (gms)	*n*, moles of sample (mol)	*N*, number of atoms, molecules, or formula units
P_2O_5	**142**	284	**2.00**	**1.20×10^{24} f.u.**
X	**95.94**	8.478	**8.84×10^{-2}**	5.32×10^{22} atoms
$NaBrO_3$	**151**	**0.484**	3.21×10^{-3}	**1.93×10^{21} f.u.**

P_2O_5: $\mathbf{284\ g \left(\frac{1\ mol}{142\ g}\right) = 2.00\ mol}$ $\mathbf{2.00\ mol \left(\frac{6.02 \times 10^{23}\ f.u.}{1\ mol}\right) = 1.20 \times 10^{24}\ f.u.}$

X: $\mathbf{5.32 \times 10^{22}\ atoms \left(\frac{1\ mol}{6.02 \times 10^{23}\ atoms}\right) = 8.84 \times 10^{-2}\ mol}$

$\mathbf{\left(\frac{8.478\ g}{8.84 \times 10^{-2}\ mol}\right) = 95.94 \frac{g}{mol}}$ **X is Mo (molybdenum)**

$NaBrO_3$: $\mathbf{3.21 \times 10^{-3}\ mol \left(\frac{151\ g}{1\ mol}\right) = 0.484\ g}$

$\mathbf{3.21 \times 10^{-3}\ mol \left(\frac{6.02 \times 10^{23}\ f.u.}{1\ mol}\right) = 1.93 \times 10^{21}\ f.u.}$

What is the symbol for the unknown element, X? **Mo**

Periodic Table of the Elements

	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII	VIII	VIII	IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
1	1 H 1.008																	2 He 4.00
2	3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.30											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
5	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 (98)	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
6	55 Cs 132.9	56 Ba 137.3	57 La 138.9	72 Hf 178.5	73 Ta 180.9	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)
7	87 Fr (223)	88 Ra 226.0	89 Ac 227.0	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hs (269)	109 Mt (268)	110 (271)	111 (272)	112 (277)		114 (285)		116 (289)		

Lanthanides	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (145)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.2	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
Actinides	90 Th 232.0	91 Pa 231.0	92 U 238.0	93 Np 237.0	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

Useful Information

1 pound (lb) = 453.59237 gram (gm)

1 liter (L) = 1.056718 quart (qt)

1 inch (in) = 2.54 centimeters (cm)

$^{\circ}C = \frac{5}{9}(^{\circ}F - 32)$

$K = {}^{\circ}C + 273.15$

relative weighted average atomic mass = Σ(isotopic mass · fractional abundance)

Avogadro's number = 6.02×10^{23}

density of water = $1.00 \frac{g}{mL}$