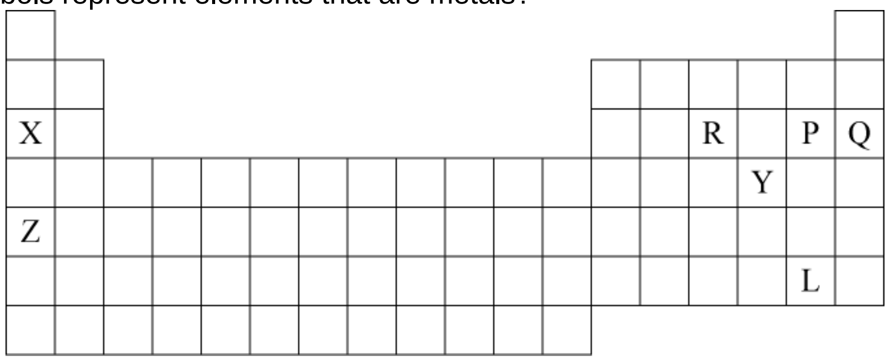
CHEM 100 **Name**________________________________
Exam 1 **Summer 2011**

Part I. Multiple choice. Circle the correct answer for each problem. 3 points each

1. Which symbols represent elements that are metals?



a. X and Z b. X and Q
c. P and L d. X, R, P, and Q

2. In the hydrogen chloride molecule, HCl, the chlorine end of the molecule is more negative than the hydrogen end because

a. hydrogen transfers an electron to chlorine.
b. chlorine is more electronegative than hydrogen.
c. hydrogen and chlorine have the same electronegativity.
d. hydrogen is more electronegative than chlorine.

3. Which is an IMPOSSIBLE electron configuration?

a. $1s^2 2s^2 2p^4$ b. $1s^2 2s^2$
c. $1s^2 2s^2 2p^6 3s^1$ d. $1s^2 2s^2 2d^{10} 2p^4$

4. In a Lewis formula the dots represent:

a. whatever number of electrons are needed to satisfy the octet rule.
b. all the electrons in the atoms.
c. the valence electrons in all the atoms.
d. only the electrons that are being transferred or shared.

5. The atomic number is the

a. same as the mass number of an atom. b. number of protons in a nucleus.
c. number of protons and neutrons in the nucleus. d. number of neutrons in a nucleus.

6. Which is an isotope of cobalt? (X may represent any elemental symbol)

a. $^{59}_{29}X$ $\boxed{\text{b. } ^{60}_{27}X}$ c. $^{63}_{29}X$ d. all of these

7. If a central atom has a total of three groups and no lone pairs attached to it, the geometry about the central atom is

a. linear.
b. tetrahedral.
c. triangular (*aka* trigonal planar).
d. pyramidal (*aka* trigonal pyramidal).

8. The chemical properties of the elements are chiefly due to

a. the number of protons.
b. the number and distribution of the outer electrons.
c. the number and distribution of the inner electrons.
d. the number and distribution of the neutrons.

9. What type of forces exist between oxygen molecules in liquid oxygen (O_2)?

a. hydrogen bonding
b. dispersion forces
c. dipole forces
d. a mixture of all of these forces.

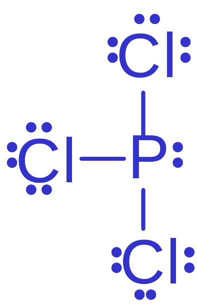
10. For which of the following is hydrogen bonding **NOT** a factor?

a. CH_3OH
b. CH_4
c. NH_3
d. H_2O

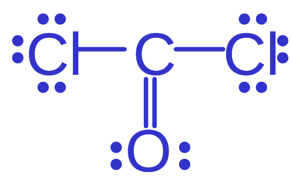
Part II Short Answer. Complete each of the following. Point values are noted by each question.

11. Draw dot (Lewis) structures for each of the following: (4 points each)

a. PCl_3



b. $COCl_2$ *(C is in the middle)*



12. Complete the following table: (1 point per blank, 10 total)

Name	Symbol	# Protons	# Neutrons	# Electrons	Atomic #	Mass #
bromide ion	$^{79}Br^-$	**35**	**44**	36	35	79
magnesium	^{24}Mg	**12**	**12**	**12**	**12**	**24**

13. Outline the main points of Dalton's atomic theory. Identify one area of his theory that we now understand to be incorrect. (8 points)

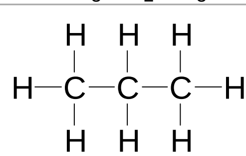
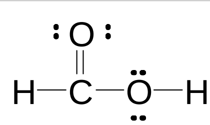
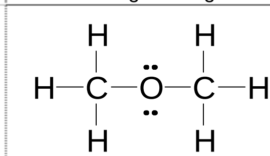
Components of Dalton's theory are below. Items in parentheses are "errors" or inaccuracies that are included in our modern atomic models.

1. **matter is made up of atoms, which are indivisible (we now know that atoms are made of smaller particles**
2. **atoms of the same element are identical (we now know about isotopes)**
3. **compounds are formed when atoms of different elements combined is certain proportions**
4. **during chemical reactions, atoms are rearranged, not destroyed (we now know about nuclear reactions)**

14. Convert the following formulas to names or names to formulas: (4 points each)

Name	**Formula**
Potassium carbonate	**K_2CO_3**
copper (I) oxide	Cu_2O
dinitrogen tetrasulfide	N_2S_4

15. Rank the compounds below in order of boiling point, from lowest to, middle, to highest. Clearly explain your reasoning. Ignore any effects of molecular mass. (8 points)

Compound	Propane	Formic Acid	Methyl Ether
Formula	$CH_3CH_2CH_3$	CHCOOH	CH_3OCH_3
Lewis Structure			
Rank (low, middle, high)	Lowest	Highest	Middle
IM Forces Present	Dispersion	Dispersion Dipole-Dipole Hydrogen Bonding	Dispersion Dipole-Dipole

The key here is that an increase in the strength of the intermolecular forces present in a material will lead to an increase in boiling point. Since propane only experiences dispersion forces, it has the lowest boiling point. Methyl ether can undergo stronger dipole-dipole interactions and has a higher boiling point. Formic acid undergoes hydrogen bonding as well, and therefore has the highest boiling point.

IMPORTANT: COVALENT BONDS ARE NOT BROKEN WHEN SOMETHING BOILS

Part III. Complete 3 of the following 4 problems. Clearly mark the problem you do not want graded. Each problem is worth ten (10) points. You must show you work on calculations to receive partial credit. Report numerical results to the correct number of significant figures and with the appropriate units.

16. Describe the similarities and differences in the electronic structures of fluorine and bromine. Include and electron configuration for each of the atoms. Why to both atoms tend to form anions with a charge of negative one (1-)?

    Consider the electron configurations for both atoms. For fluorine, the configuration is $1s^2 2s^2 2p^5$ and for bromine is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$. So, both of the atoms have seven valence electrons in a configuration $ns^2 np^5$. Bromine, however has electrons in a d orbital as well, but fluorine doesn't. Since each atom is one electron short of a filled valence shell, it is not surprising that they undergo reduction reactions, gaining a single electron to achieve a filled valence shell. In order to attain noble gas configuration by an oxidation process, the atoms would need to lose seven electrons, which is very unfavorable.

17. In class we noted that one of the requirements of a polar molecule is the presence of polar bonds. Describe how to characterize a bond as polar or not and explain how a molecule that contains polar bonds can actually be *nonpolar*.

    Your discussion should address the following

    1. **Polar bonds result from differences in electronegativity of atoms at each end of the bond.**
    2. **These bond dipoles have a size and an orientation (direction)**
    3. **Polar molecules arise when individual bond dipoles do not cancel each other out.**
    4. **If the dipoles are symmetrical (same size and opposite direction), the can cancel so a molecule may have polar bonds, but its overall structure may allow them to cancel, producing a nonpolar molecule.**

18. An oxide of tin with the formula SnO consists of 0.742 g tin and 0.100 g oxygen.

a. If 1.327 g tin and 0.325 g oxygen are allowed to react, how much SnO could be produced?

$$1.327 \text{ g tin} \times \frac{0.842 \text{ g SnO}}{0.742 \text{ g tin}} = 1.506 \text{ g SnO possible if all tin runs out first.}$$

$$0.325 \text{ g oxygen} \times \frac{0.842 \text{ g SnO}}{0.100 \text{ g oxygen}} = 2.737 \text{ g SnO possible if all oxygen runs out first.}$$

So, the oxygen will run out first, producing 1.506 g SnO

b. A sample of a second oxide of tin consists of 0.555 g tin and 0.150 g oxygen. What is the formula for this second oxide?

The second oxide has 0.555 g tin/0.150 g oxygen = 3.7 g tin/ 1 g oxygen. The first oxide has 0.742 g tin/0.100 g oxygen = 7.42 g tin/1 g oxygen. The second oxide has twice as much oxygen per tin as the first, therefore a reasonable formula is $\underline{SnO_2}$.

19. Perform the following conversions:

a. Express 7.5 x 10^{-6} m, the diameter of a red blood cell, in nonscientific (decimal) notation. (3 points)

0.0000075 m

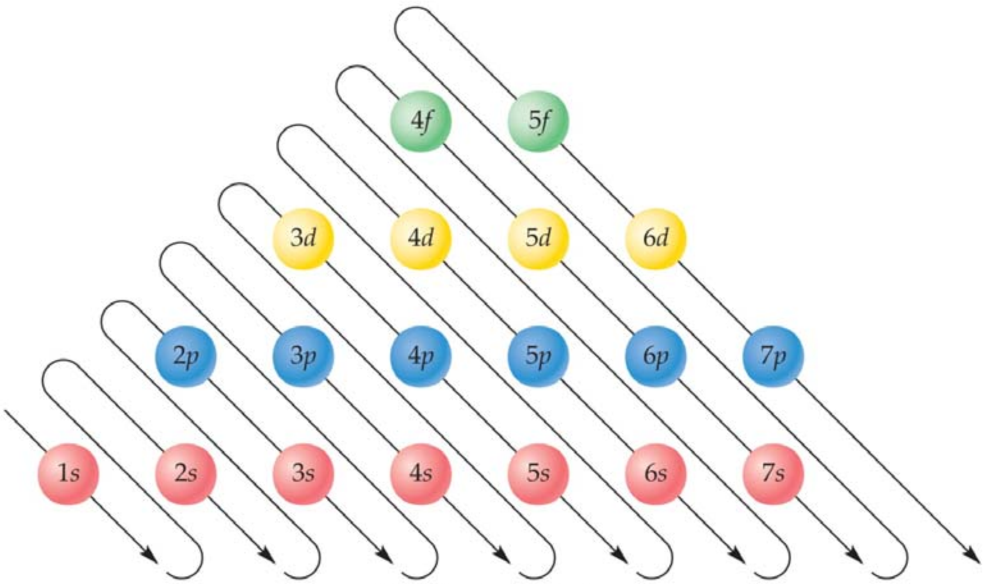
b. Convert 32.5 inches to mm. (1 inch = 0.0254 m) (4 points)

$$32.5 \cancel{\text{in}} \times \frac{0.0254 \cancel{\text{m}}}{1 \cancel{\text{in}}} \times \frac{1000 \text{ mm}}{1 \cancel{\text{m}}} = 825.5 = \mathbf{826 \text{ mm}}$$

c. Convert 3.2 x 10^5 nm to mm. (3 points)

$$3.2 \times 10^5 \cancel{\text{nm}} \times \frac{10^{-9} \cancel{\text{m}}}{1 \cancel{\text{nm}}} \times \frac{1 \text{ mm}}{10^{-3} \cancel{\text{m}}} = \mathbf{0.32 \text{ mm}}$$

Possibly Useful Information



PERIODIC CHART OF THE ELEMENTS

IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII	VIII	VIII	IB	IIB	IIIA	IVA	VA	VIA	VIIA	INERT GASES
1 H 1.00797																1 H 1.00797	2 He 4.0026
3 Li 6.939	4 Be 9.0122											5 B 10.811	6 C 12.0112	7 N 14.0067	8 O 15.9994	9 F 18.9984	10 Ne 20.183
11 Na 22.9898	12 Mg 24.312											13 Al 26.9815	14 Si 28.086	15 P 30.9738	16 S 32.064	17 Cl 35.453	18 Ar 39.948
19 K 39.102	20 Ca 40.08	21 Sc 44.956	22 Ti 47.90	23 V 50.942	24 Cr 51.996	25 Mn 54.9380	26 Fe 55.847	27 Co 58.9332	28 Ni 58.71	29 Cu 63.54	30 Zn 65.37	31 Ga 69.72	32 Ge 72.59	33 As 74.9216	34 Se 78.96	35 Br 79.909	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.905	40 Zr 91.22	41 Nb 92.906	42 Mo 95.94	43 Tc (99)	44 Ru 101.07	45 Rh 102.905	46 Pd 106.4	47 Ag 107.870	48 Cd 112.40	49 In 114.82	50 Sn 118.69	51 Sb 121.75	52 Te 127.60	53 I 126.904	54 Xe 131.30
55 Cs 132.905	56 Ba 137.34	*57 La 138.91	72 Hf 178.49	73 Ta 180.948	74 W 183.85	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.09	79 Au 196.967	80 Hg 200.59	81 Tl 204.37	82 Pb 207.19	83 Bi 208.980	84 Po (210)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	†89 Ac (227)	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 ? (271)	111 ? (272)	112 ? (277)						

Numbers in parenthesis are mass numbers of most stable or most common isotope.

Atomic weights corrected to conform to the 1963 values of the Commission on Atomic Weights.

The group designations used here are the former Chemical Abstract Service numbers.

* Lanthanide Series

58 Ce 140.12	59 Pr 140.907	60 Nd 144.24	61 Pm (147)	62 Sm 150.35	63 Eu 151.96	64 Gd 157.25	65 Tb 158.924	66 Dy 162.50	67 Ho 164.930	68 Er 167.26	69 Tm 168.934	70 Yb 173.04	71 Lu 174.97

† Actinide Series

90 Th 232.038	91 Pa (231)	92 U 238.03	93 Np (237)	94 Pu (242)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (249)	99 Es (254)	100 Fm (253)	101 Md (256)	102 No (256)	103 Lr (257)