Chem 1515
Review Problem Set
Fall 2001

Name______________________________

TA Name _________________________

Lab Section #_______

ALL work must be shown to receive full credit. **Due in lecture, at 2:30 p.m. on Friday, August 31, 2001.**

RPS.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) **2**Na*(s)* + Cl_2*(g)* $\xrightarrow{H_2O}$ **$2NaCl_{(s)}$**

b) **2**C_4H_{10}*(l)* + **13**oxygen*(g)* → **$8CO_{2(g)}$ + $10H_2O_{(l)}$ or (g)**

c) silver bromide*(s)* + sodium thiosulfate*(aq)* → → **$Ag(S_2O_3)_2^{3-}{}_{(aq)}$ + $4Na^+_{(aq)}$ + $Br^-_{(aq)}$**

d) **2**NaCl*(s)* + H_2SO_4*(l)* $\xrightarrow{\Delta}$ **$Na_2SO_{4(s)}$ + $2HCl_{(g)}$**

e) SiO_2*(l)* + C*(s)* $\xrightarrow{\Delta}$ **Si*(s)* + CO*(g)* or CO_2*(g)***

f) $AgNO_3$*(aq)* + NaCl*(aq)* → **$AgCl_{(s)}$ + $Na^+_{(aq)}$ + $NO_3^-{}_{(aq)}$**

g) **3**iron metal*(s)* + **8**warm dilute nitric acid*(aq)* → **$2NO_{(g)}$ + $3Fe^{2+}_{(aq)}$ + $6NO_3^-{}_{(aq)}$+ $4H_2O_{(l)}$**

h) HgS*(s)* + O_2*(g)* $\xrightarrow{\Delta}$ **Hg*(l)* + SO_2*(g)***

RPS.2. Write the ionic and net ionic chemical equations for 1a), 1c), 1d), 1f) and 1g).

1a) **2Na*(s)* + Cl_2*(g)* $\xrightarrow{H_2O}$ $2NaCl_{(s)}$**
ionic and net ionic are the same

1c) **AgBr*(s)* + $4Na^+_{(aq)}$ + $2S_2O_3^{2-}{}_{(aq)}$ → $Ag(S_2O_3)_2^{3-}{}_{(aq)}$ + $4Na^+_{(aq)}$ + $Br^-_{(aq)}$ ionic**
AgBr*(s)* + $2S_2O_3^{2-}{}_{(aq)}$ → $Ag(S_2O_3)_2^{3-}{}_{(aq)}$ + $Br^-_{(aq)}$ net ionic

1d) **2NaCl*(s)* + H_2SO_4*(l)* $\xrightarrow{\Delta}$ $Na_2SO_{4(s)}$ + $2HCl_{(g)}$**
ionic and net ionic are the same

1f) **$Ag^+_{(aq)}$ + $NO_3^-{}_{(aq)}$ + $Na^+_{(aq)}$ + $Cl^-_{(aq)}$ → $AgCl_{(s)}$ + $Na^+_{(aq)}$ + $NO_3^-{}_{(aq)}$ ionic**
$Ag^+_{(aq)}$ + $Cl^-_{(aq)}$ → $AgCl_{(s)}$ net ionic

1g) **3Fe*(s)* + $8H^+$*(aq)* + $8NO_3^-$*(aq)* → $2NO_{(g)}$ + $3Fe^{2+}_{(aq)}$ + $6NO_3^-{}_{(aq}$+ $4H_2O_{(l)}$ ionic**
3Fe*(s)* + $8H^+$*(aq)* + $2NO_3^-$*(aq)* → $2NO_{(g)}$ + $3Fe^{2+}_{(aq)}$ + $4H_2O_{(l)}$ net ionic

RPS.3. Figure I shows a glass cylinder containing four liquids each of different density. Two of the liquids have been identified. A table containing a list of substances and their density (at 25 ˚C) has been provided. From the list select a substance for Liquid #1 and Liquid #3. Briefly explain the reason(s) for your selections and for the remaining substances the reason they were not selected.

Substance	Density ($\frac{g}{mL}$)	Liquid #
Mercury	13.5	Liquid #4
Water	1.0	Liquid #2
Hexane	0.660	**Liquid #1**
Ethyl alcohol	0.789	
Dichloromethane	1.33	**Liquid #3**
Aluminum	2.699	
Bromine	2.928	
Gold	19.3	

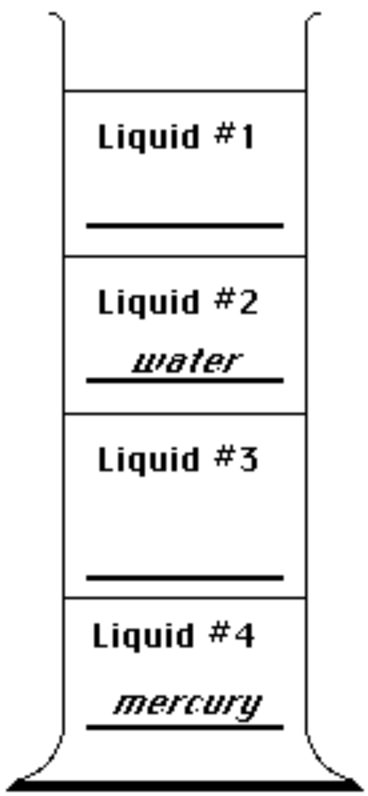


Figure I.

Liquid #1 must have a density less than the density of water. So it can be either hexane or ethyl alcohol. Ethyl alcohol is very soluble in water and would form a homogeneous solution, rather than a heterogeneous mixture. Hexane is insoluble in water and will form a heterogeneous mixture. Liquid #3 must have a density between water and mercury. So it can be dichloromethane, aluminum, or bromine. Aluminum is a solid at room temperature and bromine is a very reactive liquid and extremely poisonous. Therefore Liquid #3 must be dichloromethane. Gold is also a solid at room temperature. It has a density greater than that of mercury. Mercury is soluble in gold so a homogeneous mixture would be formed.

RPS.4. The amount of copper in a sample can be determined by dissolving the sample in water and reacting with zinc metal, according to the following reaction;

$$CuSO_4(aq) + Zn(s) \rightarrow Cu(s) + ZnSO_4(aq)$$

The metallic copper can be weighed after separating it from the solution. To insure complete conversion of the copper an excess of zinc is typically added to the solution. Any zinc which remains unreacted can be converted to a soluble form by adding an acid like H_2SO_4.

a) Write the reaction which you would expect to occur between zinc metal and sulfuric acid.

$$\mathbf{H_2SO_4(aq) + Zn(s) \rightarrow ZnSO_4(aq) + H_2(g)}$$

b) Calculate the value of *x* in the formula $CuSO_4 \cdot xH_2O$, if a 1.20 g sample of the hydrate reacts with excess zinc metal followed by addition of sulfuric acid yields 0.306 g of copper metal.

The hydrate of copper(II) sulfate contains copper(II) sulfate which reacts with the Zn and some water. Knowing how much copper metal is formed allows to back calculate the amount of copper(II) sulfate in the original sample of the hydrate.

$$\mathbf{0.306\ g\ Cu \left(\frac{1\ mol\ Cu}{63.6\ g\ Cu}\right)\left(\frac{1\ mol\ Cu}{1\ mol\ CuSO_4}\right)\left(\frac{159.6\ g\ CuSO_4}{1\ CuSO_4}\right) = }$$
$$\mathbf{0.768\ g\ CuSO_4}$$

$$\mathbf{1.20\ g\ CuSO_4 \cdot xH_2O\ -\ 0.768\ g\ CuSO_4 = 0.432\ g\ xH_2O}$$

Now that we know the mass of water in the original sample. Now the problem is identical to an empirical formula problem. We need to determine the mols of $CuSO_4$ and of H_2O.

$$\mathbf{0.768\ g\ CuSO_4 \left(\frac{1\ mol\ CuSO_4}{159.6\ g}\right) = 0.00481\ mol\ CuSO_4}$$
$$\mathbf{0.432\ g\ H_2O \left(\frac{1\ mol\ H_2O}{18.0\ g}\right) = 0.0240\ mol\ H_2O}$$

$$\mathbf{0.00481\ mol\ CuSO_4 : 0.0240\ mol\ H_2O}$$
$$\mathbf{\left(\frac{0.00481\ mol\ CuSO_4}{0.00481\ mol\ CuSO_4}\right) : \left(\frac{0.0240\ mol\ H_2O}{0.00481\ mol\ CuSO_4}\right)}$$
$$\mathbf{1\ mol\ CuSO_4 : 5\ mol\ H_2O}$$

$$\mathbf{CuSO_4 \cdot 5H_2O}$$

RPS.5. A gaseous mixture in a 5.00 L reaction vessel containing 114.0 g of $CS_{2(g)}$, 3.500 g of $H_{2(g)}$ and 88.00 g of $CH_{4(g)}$ at a particular temperature is allowed to react. After the reaction occurs, analysis shows 94.34 g of CH_4 are present. The equation which describes the reaction is;

$$CS_{2(g)} + 4H_{2(g)} \rightarrow CH_{4(g)} + 2H_2S_{(g)}$$

Calculate the mass of CS_2 and H_2 reacting and the mass of CS_2 and H_2 remaining.

This problem looks like a limiting reagent problem, but it is really easier than that. In limiting reagent problems amounts of both reactants are given and the amount of a particular product must be calculated. In this problem amounts of both reactants are given (just like a limiting reagent problem) plus the amount of one of the product (something extra), but then I tell you the reaction occurs AND some amount of a product is produced. In this case we are told after the reaction occur 94.34 g of CH_4 are present. Before the reaction the amount of CH_4 was 88.00 g. The difference tells us how much CH_4 is formed in the reaction, then we can calculate how much of each reactant, reacted.

Another way to word this question (Gelder wanted to word it this way, but Geldar wanted it as above):
a) Calculate the mass of CS_2 and H_2 that must react to form 6.34 g of CH_4 in the reaction above.
THEN,
b) Calculate the mass of CS_2 and H_2 remaining if we started with containing 114.0 g of $CS_{2(g)}$, 3.500 g of $H_{2(g)}$ and 88.00 g of $CH_{4(g)}$ at a particular temperature.
Oh, by the way the volume is not needed for any of the calculations in this problem.

$$6.34 \text{ g } CH_4 \left(\frac{1 \text{ mol } CH_4}{16.0 \text{ g } CH_4}\right)\left(\frac{1 \text{ mol } CS_2}{1 \text{ mol } CH_4}\right)\left(\frac{76.0 \text{ g } CS_2}{1 \text{ } CS_2}\right) = 30.1 \text{ g } CS_2$$

$$6.34 \text{ g } CH_4 \left(\frac{1 \text{ mol } CH_4}{16.0 \text{ g } CH_4}\right)\left(\frac{4 \text{ mol } H_2}{1 \text{ mol } CH_4}\right)\left(\frac{2.02 \text{ g } H_2}{1 \text{ } H_2}\right) = 3.20 \text{ g } H_2$$

114.0 g of $CS_{2(g)}$ – 30.1 g CS_2 = 83.9 g CS_2
3.500 g of H_2 – 3.20 g H_2 = 0.30 g H_2

RPS.6. Using a table of Standard Enthalpies of Formation (Appendix C, p 1012 in Brown, LeMay and Bursten), calculate the enthalpy of reaction for each of the following;

a) $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$

$$\Delta H^\circ_{rxn} = \Sigma\Delta H^\circ_f \text{ (P)} - \Sigma\Delta H^\circ_f \text{ (R)}$$

$$\Delta H^\circ_{rxn} = (-986 \text{ kJ}) - [(-285.8 \text{ kJ}) + (-635 \text{ kJ})]$$

$$= -65.2.8 \text{ kJ}$$

b) $C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$ (Note: $\Delta H^\circ_f(C_6H_{12}O_6(s)) = -1254 \frac{kJ}{mol}$)

$$\Delta H^\circ_{rxn} = \Sigma\Delta H^\circ_f \text{ (P)} - \Sigma\Delta H^\circ_f \text{ (R)}$$

$$\Delta H^\circ_{rxn} = 6(-393.5 \text{ kJ}) + 6(-285.8 \text{ kJ}) - [-1254 \text{ kJ}]$$

$$= -2821.8 \text{ kJ}$$

RPS.7. In a particular version of a solar heating system the radiation from the sun is used in the following chemical conversion,

$$Na_2SO_4 \cdot 10H_2O(s) \rightarrow Na^+(aq) + SO_4^{2-}(aq) + 10H_2O(l)$$

The enthalpy, ΔH°, for this reaction is +78.7 kJ. If clouds form, or during the evening the outside temperature drops and the reverse reaction occurs. The heat produced when the reverse reaction occurs heats water in a storage tank. If 1.00 kg of $Na_2SO_4 \cdot 10H_2O(s)$ is formed, calculate the final temperature of a 10.0 gallon storage tank containing water at 25.0 ˚C.

$1{,}000 \text{ kg} \left(\frac{1 \text{ mol}}{322 \text{ g}}\right) = 3.11 \text{ mol} \left(\frac{78700 \text{ J}}{1 \text{ mol}}\right) = 2.44 \times 10^5 \text{ J}$ heat released

$10 \text{ gal} \left(\frac{4 \text{ qt}}{1 \text{ gal}}\right)\left(\frac{1 \text{ L}}{1.0567 \text{ qt}}\right)\left(\frac{1000 \text{ mL}}{1 \text{ Ll}}\right)\left(\frac{1 \text{ g}}{1 \text{ mL}}\right) = 3.79 \times 10^4 \text{ mL}$

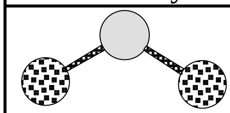
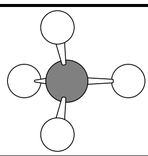
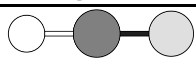
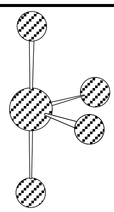
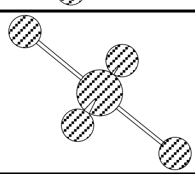
heat (g) = mass * S.H. * ΔT

$$2.44 \times 10^5 \text{ J} = 37850 \text{ gm} \left(\frac{4.184 \text{ J}}{g \cdot {}^\circ C}\right) \cdot \Delta T = 3{,}785 \text{ mL} \left(\frac{1 \text{ gm}}{1 \text{ ml}}\right)$$

$$1.540 \ {}^\circ C = \Delta T$$

$T_o = 25 \ {}^\circ C$ so $T_f = 26.5 \ {}^\circ C$

RPS.8a. Complete the following table

Geometry	Compound	Number of bonding groups on central atom	Number of non-bonding pairs on central atom	Name of the molecular geometry	Bond Angle(s)
	NO_2^-	**2**	**1**	**bent**	**~120˚**
	CH_4	**4**	**0**	**tetrahedral**	**109.5˚**
	HCN	**2**	**0**	**linear**	**180˚**
	SF_4	**4**	**1**	**seesaw**	**90˚ , ~119˚ , ~178˚**
	XeF_4	**4**	**2**	**square planar**	**90˚**

b. Indicate which of the molecular substances in part 8a) is polar and which are nonpolar. Support your conclusions with a brief explanation.

CH_4 and XeF_4 are nonpolar. For CH_4 the central carbon atom has no lone pairs of electrons and all of the terminal atoms are identical. For XeF_4 the central atom has two lone pairs of electrons, but the position of the electrons are above and below the plane of the atoms, so the geometry of the molecule cancels the electron pairs out. This is one exception where a central atom with two lone pairs is nonpolar. Another case where a molecule can be nonpolar is when a linear molecule with three lone pairs on the central atom.

HCN and SF_4 are polar. The central carbon atom in HCN has no lone pairs, but the terminal atoms are different. In the case of SF_4 the central atom has a lone pair.

NO_2^- is an ion and the question is really not appropriate for an ion.

RPS.9a. Write the electron configuration for S, Ba, O, Fe, Cl and Bi.

S: ***$1s^2 2s^2 2p^6 3s^2 3p^4$ or [Ne]$3s^2 3p^4$***

Ba: ***$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$ or [Xe]$6s^2$***

O: ***$1s^2 2s^2 2p^4$ or [He]$2s^2 2p^4$***

Fe: ***$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$ or [Ar]$4s^2 3d^6$***

Cl: ***$1s^2 2s^2 2p^6 3s^2 3p^5$ or [Ne]$3s^2 3p^5$***

Bi: ***$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^3$ or [Xe]$6s^2 4f^{14} 5d^{10} 6p^3$***

b) Which elements in part a) are metals and which are nonmetals?

O, S and Cl are nonmetals

Ba, Fe, and Bi are metals

c) As it relates to electron gain or loss, explain the difference between metals and nonmetals. Use the electron configuration of a neutral atom and its ion to support your explanation.

Metals lose electrons to attain an octet. For example, Ba has two valence electrons. The loss of the two electrons in the *6s* subshell leaves the atom with eight electrons in its outer-most level.	**Ba:** ***$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$*** **$Ba^{2+}$:** ***$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6$***
Nonmetals may gain or lose electrons. When combined with metals, nonmetals gain electrons. When combined with more electronegative nonmetals another nonmetal will lose electrons.	**S:** ***$1s^2 2s^2 2p^6 3s^2 3p^4$*** **In FeS** **S^{2-}:** ***$1s^2 2s^2 2p^6 3s^2 3p^6$*** **In SO_2** **S^{4+}:** ***$1s^2 2s^2 2p^6 3s^2$***

d) By combining a metal and a nonmetal, or a nonmetal and a nonmetal, from the elements listed in part a), write the formula and name of at least eight compounds. The compounds should include 5 ionic and 3 covalent examples.

Ionic Compounds:

$BaCl_2$	**barium chloride**	$FeCl_2$	**iron (II) chloride**
BaS	**barium sulfide**	FeS	**iron (II) sulfide**
BaO	**barium oxide**	FeO	**iron (II) oxide**
$BiCl_3$	**bismuth(III) chloride**	$FeCl_3$	**iron (III) chloride**
Bi_2S_3	**bismuth(III) sulfide**	Fe_2S_3	**iron (III) sulfide**
Bi_2O_3	**bismuth(III) oxide**	Fe_2O_3	**iron(III) oxide**

Covalent Compounds:

SO_2	**sulfur dioxide**	Cl_2O	**dichlorine oxide**
SO_3	**sulfur trioxide**	Cl_2O_7	**dichlorine heptoxide**

e) Use an extra sheet of paper to describe each of the compounds in d). Provide me with some of its physical and chemical properties and brief discussion of what makes each compound interesting/useful.

There can be all kinds of answers here.

RPS.10. Solve

a) $\log 6.57 \times 10^{-4}$ = **-3.18**

b) $\log 3.51 \times 10^{4}$ = **4.54**

c) $-\log 8.67 \times 10^{-7}$ = **6.06**

d) antilog (-10.004) = **9.91 x 10^{-11}**

e) antilog (.789) = **6.15**

f) ln 500 = **6.21**

g) ln 0.0159 = **-4.14**

h) $e^{-4.14}$ = **.0159**

i) $e^{3.90}$ = **49.40**

j) $\ln\left(\frac{452}{235}\right)$ = **.654**

k) $\ln\left(\frac{348}{x}\right) = 0.941$ Solve for x

$$e^{\ln\left(\frac{348}{x}\right)} = e^{0.941}$$

$$\frac{348}{x} = 2.56$$

x = 136

l) $\frac{1}{0.150} - \frac{1}{x} = 5.02$ Solve for x

$$\frac{1}{.150} - 5.02 = \frac{1}{x}$$ **x = 0.607**

$$1.647 = \frac{1}{x}$$

m) $0.954 = 1.57 - \frac{0.0591}{2} \log\left(\frac{1}{1 \cdot x^8}\right)$ Solve for x

$$\mathbf{(0.954 - 1.57) \bullet \left(\frac{-2}{.0591}\right) = 10g\ \frac{1}{x^8}}$$

$$\mathbf{20.846 = 10g\ \frac{1}{x^8}}$$

$$\mathbf{-20.846 = 10g\ x^8}$$

$$\mathbf{\frac{-20.846}{8} = 10g\ x}$$

$$\mathbf{-2.606 = 10g\ x}$$

$$\mathbf{2.48\ x\ 10^{-3} = x}$$

n) $x^2 + 5x - 20 = 0$ Solve for x

$$\mathbf{\frac{-b \pm \sqrt{b^2 - 4ac}}{2a}}$$

$$\mathbf{\frac{-5 \pm \sqrt{(5)^2 - 4(1)(-20)}}{2}}$$

x = +2.62

= -7.62

o) $x^3 - 0.1x^2 - 1.06 \text{ x } 10^{-2}x - 9.37 \text{ x } 10^{-4} = 0$ Solve for x

Since this is a third order polynomial there are three roots to the equation. Using a graphing calculator (I used a TI-85) the three roots are: 0.184; -0.0424; and –0.0424. When we use polynomials in this course we will only be interested in the positive root.