CHEM 1035 Homework #3 – Solutions

2.83 a. $CH_3O$
b. $HSO_4$

2.85 a. RbBr – Rubidium bromide
b. BaS – Barium sulfide
c. $CaF_2$ Calcium fluoride

2.91 a. Tin(II) sulfite
b. $K_2C_2O_7$
c. Iron(II) carbonate
d. $Cu(NO_3)_2$

2.107 a. $Fe(CH_3COO)_2 – 4H_2O$ FW=246.00 Da
b. $SCl_4$ FW=173.87 Da
c. $KMnO_4$ FW = 158.04 Da

2.120 This is a mixture. We know it is a mixture because the proportions of the "Drug" and the Vitamin C are not fixed – rather they are variable dependent upon the source of the street drug.

2.122 The strongest ionic bonds occur between ions with the highest "charge density" (e.g. the most charge per unit volume). For the cations, $Mg^{+2}$ has a +2 charge, and the smallest volume (radius of 72pm) – this cation has the highest charge density of the cations. Likewise, of the anions, $O^{-2}$ has the most charge (-2) and the smallest volume (radius of 140pm). The strongest ionic bond will be between $Mg^{+2}$ and $O^{-2}$ to form MgO.

The weakest ionic bond will be between the ions with the smallest charge density (e.g. the least amount of charge per unit volume). For the cations, this is found to be $Rb^+$, and for the anions it is $I^-$. The weakest ionic compound, therefore, is RbI.

3.13 a. $$3.8x10^{20}\,molecules NO_2 \times \frac{46.01 Da}{molecule} \times \frac{1.66054x10^{-27}\,kg}{Da} = 2.9 \times 10^{-5}\,kg NO_2$$

b. $$0.0425 g C_2H_4Cl_2 \times \frac{1 mole C_2H_4Cl_2}{98.95 g C_2H_4Cl_2} \times \frac{2 mole Cl}{1 mole C_2H_4Cl_2} = 8.59 \times 10^{-4}\,mole Cl$$

c.

$$4.92 g SrH_2 \times \frac{1 mole SrH_2}{89.64 g SrH_2} \times \frac{2 mole H^-}{1 mole SrH_2} \times \frac{6.022 \times 10^{23}\,H^- ions}{1 mole H^-} = 6.61 \times 10^{22}\,H^- ions$$

3.19 a. $Sr(IO_4)_2$ FW = 469.42 Da
I FW = 126.9 Da

But there are 2 Iodine atoms in each $Sr(IO_4)_2$ molecule – so I contributes 253.8 Da to the total mass of $Sr(IO_4)_2$

Mass % of I is: $\frac{253.8 Da I}{469.42 Da\ Sr(IO_4)_2} \times 100\% = 54.07\%\ I$

b. $KMnO_4$ FW = 158.04 Da
Mn FW = 54.94 Da

Mass % of Mn is: $\frac{54.94 Da Mn}{158.04 Da\ KMnO_4} \times 100\% = 34.76\%\ Mn$

3.24 FW of allyl sulfide = $C_6H_{10}S$, Molar mass =114.21gm/mole

a

$$1.63 mole C_6H_{10}S \times \frac{114.21 g C_6H_{10}S}{mole C_6H_{10}S} = 186 g\ C_6H_{10}S$$

b.

$$4.77 g\ C_6H_{10}S \times \frac{1 mole C_6H_{10}S}{114.21 g\ C_6H_{10}S} \times \frac{6 mole C}{1 mole C_6H_{10}S} \times \frac{6.022 \times 10^{23} C atoms}{1 mole C} = 1.51 \times 10^{23} C atoms$$

3.36 a Empirical formula: CH Molar mass = 78.11g/mole
The mass of the empirical formula is 13.008gm/mole. To determine the total number of atoms in the molecule, determine the ratio of the Empirical formula mass to that of the molar mass:

$\frac{78.11}{13.008} = 6$ Therefore, the Molecular formula is $C_6H_6$

b. Empirical Formula: $C_3H_6O_2$ Molar mass = 74.08gm/mole

Mass of the empirical formula is 74.08gm/mole

$\frac{74.08}{74.08} = 1$ Therefore, the Molecular formula is $C_3H_6O_2$

c. Empirical Formula: HgCl Molar mass = 472.1gm/mole
The mass of the empirical formula is 236.05gm/mole

$\frac{472.1}{236.05} = 2$ Therefore, the Molecular formula is $Hg_2Cl_2$

d. Empirical Formula: $C_7H_4O_2$ Molar mass = 240.20gm/mole

The mass of the empirical formula is 120.10gm/mole

$$\frac{240.20}{120.10} = 2$$ Therefore, the molecular formula is $C_{14}H_8O_4$

3.38 a. $Fe_{0.039}O_{0.052}$

Find the smallest whole number ratio of atoms:

$$\frac{0.052}{0.039} = 1.333$$

$Fe_1O_{1.333}$ - This, however, is not a whole number ratio. To get a whole number ratio, we need to multiply each of the subscripts by 3.

$Fe_3O_4$

b.

$$0.903 gP \times \frac{1 mole P}{30.97 gP} = 0.0292 mole P$$

$$6.99 gBr \times \frac{1 mole Br}{79.90 gBr} = 0.0875 mole Br$$

$P_{0.0292}Br_{0.0875}$ Again, this is not a whole number ratio:

$$\frac{0.0875}{0.0292} = 2.997 \approx 3$$ and the empirical formula is: $PBr_3$

c. A hydrocarbon (a compound that contains only C and H) has 79.9% C and 20.1% H. Assume that we start with 100 g of the compound. This says that we will have 79.9g of C and 20.1g of H

$$79.9 gC \times \frac{1 mole C}{12.01 gC} = 6.653 mole C$$

$$20.1 gH \times \frac{1 mole H}{1.008 gH} = 19.94 mole H$$

$C_{6.653}H_{19.94}$ Again, this is not a whole number ratio:

$$\frac{19.94}{6.653} = 2.997 \approx 3$$ and the empirical formula is $CH_3$

3.40 If a compound contains Si and Cl, and the mass percentage of Cl is 79.1%, then the mass percentage of Si is 20.9%. Assume 100gms of the compound

a. $$20.9gSi \times \frac{1moleSi}{28.09gSi} = 0.744moleSi$$

$$79.1gmCl \times \frac{1moleCl}{35.45gCl} = 2.23moleCl$$

Empirical Formula = $Si_{0.744}Cl_{2.23}$ This is not in whole number ratios

$$\frac{2.23}{0.744} = 2.997 \approx 3$$ And the empirical formula is $SiCl_3$

b. The Molar mass of the compound is 269g/mole.

The mass of the empirical formula is 134.45gm/mole

The molar mass is 2x that of the empirical formula mass. That tells us that the molecular formula is: $Si_2Cl_6$