CHEMISTRY 123-01
Practice exam #2 – answer key
September 14, 2007

PART I: MULTIPLE CHOICE

1. An atom of bromine has a mass about four times greater than that of an atom of neon. How many grams of neon will contain the same number of atoms as 1,000 g of bromine?
   a. 4 g Ne
   b. 250 g Ne
   c. 400 g Ne
   d. 1,000 g Ne
   e. 4,000 g Ne

2. The mass of 1.21×10^{20} atoms of sulfur is:
   a. 3.88×10^{21} g.
   b. 2.00 mg.
   c. 32.06 g.
   d. 6.44 mg.
   e. 2.00×10^{-4} g.

3. One mole of iron:
   a. is heavier than one mole of lead (Pb).
   b. is 77.0 g of iron.
   c. is 26.0 g of iron.
   d. weighs the same as one mole of lead.
   e. is none of these.

4. How many moles of Cl atoms are there in 65.2 g $CHCl_3$?
   a. 0.548 mol
   b. 1.09 mol
   c. 3.3×10^{23} mol
   d. 1.64 mol
   e. 3.0 mol

5. An average atom of uranium (U) is approximately how many times heavier than an atom of potassium?
   a. 6.1 times
   b. 4.8 times
   c. 2.4 times
   d. 12.5 times
   e. 7.7 times

6. The element oxygen consists of three naturally occurring isotopes: ^{16}O, ^{17}O, and ^{18}O. The atomic mass of oxygen is 16.0 amu. What can be implied about the relative abundances of these isotopes?
   a. More than 50% of all O atoms are ^{17}O.
   b. Almost all O atoms are ^{18}O.
   c. Almost all O atoms are ^{17}O.
   d. The isotopes all have the same abundance, i.e. 33.3%.
   e. The abundances of ^{17}O and ^{18}O are very small.

7. Which of the following cannot be determined from a balanced chemical equation?
   a. The number of moles of reactants and products.
   b. The number of molecules of reactants and products.
   c. The relative mass of each reactant and product.
   d. Whether the reaction will proceed as written.
   e. The number of atoms of each element reacting.

8. In the reaction given below, how many grams of water are consumed if 2.0 g of hydrogen gas and 16.0 g of oxygen gas are produced?

$$2\ H_2O\ (l) \longrightarrow .\ 2\ H_2(g) + O_2(g)$$

a. 2.0 g
b. 4.0 g
c. 18.0 g
d. 20.0 g
e. 36.0 g

9. In the reaction given below, for every unit of aluminum oxide consumed, how many atoms of aluminum are produced?

$$2\ Al_2O_3 \longrightarrow \_\ 4\ Al + 3\ O_2$$

a. 1
b. 2
c. 4
d. 7
e. 8

10. The Roman numerals in the reaction given represent the coefficients in the balanced chemical equation. What are the values of the coefficients?

$$I\ C_3H_8 + II\ O_2 \longrightarrow \_\ III\ CO_2 + IV\ H_2O$$

I II III IV
a. 2 10 6 4
b. 1 10 6 4
c. 1 10 3 4
d. 1 5 3 4
e. 3 5 3 4

11. The Roman numerals in the reaction given represent the coefficients in the balanced chemical equation. What are the values of the coefficients?

$$I\ H_2O + II\ CO_2 \longrightarrow \_\ III\ C_6H_{12}O_6 + IV\ O_2$$

I II III IV
a. 6 6 1 6
b. 6 6 2 6
c. 1 6 1 6
d. 3 3 1 3
e. 1 3 1 3

12. The Roman numerals in the reaction given represent the coefficients in the balanced chemical equation. What are the values of the coefficients?

$$I\ KO_2 + II\ H_2O + III\ CO_2 \longrightarrow \_\ IV\ KHCO_3 + V\ O_2$$

I II III IV V
a. 4 4 4 4 3
b. 4 2 4 4 3
c. 4 4 4 2 3
d. 2 2 2 2 3
e. 4 2 4 2 3

13. How many grams of CuCl are formed from the complete reaction of 0.750 mol of Cu with excess Cl_2 (Pay attention to stoichiometry!!)?

$$Cu + Cl_2 \longrightarrow CuCl$$

a. 32.2 g
b. 43.6 g
c. 74.5 g
d. 101 g
e. 148.5 g

14. Which statement about the reaction of 10.0 g of carbon with 10.0 g of Cl_2 is true?

$$C + Cl_2 \longrightarrow CCl_4$$

a. The equation, as written, is balanced.
b. Carbon is the limiting reactant.
c. Chlorine is the limiting reactant
d. There will be 126 g of CCl_4 produced.
e. There will be 39.5 g of CCl_4 produced.

15. What is the maximum possible quantity of product obtained from a chemical reaction called?
a. limiting reactant
b. theoretical yield
c. percent yield
d. molecular weight of the product
e. stoichiometric coefficients

PART II: CALCULATION PROBLEMS (Show your work in its entirety. Do not provide just a single number! Pay attention to significant figures!).

16. The complete reaction of 16.12 g of titanium with 23.88 g of chlorine (Cl_2) produces a compound with the formula Ti_xCl_y. What is the empirical formula of the compound?

Answer: We need to establish the ratio of moles of Ti and Cl, which will automatically give us the ratio of atoms (i.e. the empirical formula). Thus we start we conversion of grams into moles:

mol Ti = 16.12 g Ti x (1 mol/47.88 g Ti) = 0.3367 mol Ti
mol Cl = 23.88 g Cl_2 x (1 mol/70.90 g Cl_2) x (2 mol Cl/1 mol Cl_2) = 0.6736 mol Cl (notice that we had to convert mol of Cl_2 into mol Cl !!)

The next step is to find the ratio of moles:

0.6736 mol Cl/0.3367 mol Ti = 2/1

The empirical formula is $TiCl_2$

17. How many grams of MgI_2 are produced by the reaction of 75.0 g of Mg with 75.0 g of I_2?

$$Mg + I_2 \longrightarrow MgI_2$$

Answer: We first need to establish which is the limiting reactant. Let us start with assuming that we react the entire amount of Mg. It will require:

75.0 g Mg x (1 mol/24.31 g Mg) x (1 mol I_2/1 mol Mg) x (253.80 g I_2/1 mol) = 783g I_2

In other words, it will require a lot more I_2 than present, in order to consume the entire amount of Mg. At the point when all 75.0 g of I_2 is consumed, there will still be a considerable amount of Mg left. Therefore I_2 is the limiting reactant and the subsequent calculation on the theoretical yield of Mg I_2 must be based on the quantity of I_2.

75.0 g I_2 x (1 mol/253.80 g I_2) x (1 mol MgI_2/1 mol I_2) x (278.11 g MgI_2/1 mol) = 82.2 g MgI_2

The reaction will lead to the generation of 82.2 g MgI_2.

18. Allicin, a sulfur-containing compound found in garlic, is a potent antibacterial agent. A sample of 5.00 mg of allicin was burnt to produce 8.13 mg of CO_2, 3.95 mg of SO_2 and 2.76 mg of H_2O. The molar mass of allicin is 160 g/mol. What is its molecular formula?

Answer: First, let us convert the mg into gram quantities, using the conversion factor 1g/1000 mg.

Sample: 0.00500 g CO_2: 0.00813 g SO_2: 0.00395 g H_2O: 0.00276 g

The product quantities can be used to find the amounts of C, S and H:

0.00813 g CO_2 x (1 mol/44.01 g CO_2) x (1 mol C/1 mol CO_2) x (12.01 g C/1 mol) = 0.00222 g C
0.00395 g SO_2 x (1 mol/64.06 g SO_2) x (1 mol S/1 mol SO_2) x (32.06 g S/1 mol) = 0.00198 g S
0.00276 g H_2O x (1 mol/18.02 g H_2O) x (2 mol H/1 mol H_2O) x (1.01 g H/1 mol) = 0.000309 g H

If any oxygen was present in the compound, its quantity can be found from the mass of the sample:

g O = 0.00500 g – (0.00222 g + 0.00198 g + 0.000309 g) = 0.000491 g O

Next we find the number of moles of each element:

Mol C = 0.00222 g C x (1 mol C/12.01 g) = 0.000185 mol C
Mol S = 0.00198 g S x (1 mol S/32.06 g) = 0.0000617 mol S
Mol H = 0.000309 g H x (1 mol H/1.01 g) = 0.000306 mol H
Mol O = 0.000491 g O x (1 mol O/16.00 g) = 0.0000306 mol O

Construct ratios with common denominator:

0.000185 mol C/0.0000306 mol O = 6/1

0.000306 mol H/0.0000306 mol O = 10/1

0.0000617 mol S/0.0000306 mol O = 2/1

This leads to an empirical formula: $C_6H_{10}OS_2$

The weight of this unit is: 6 x 12.01 + 10 x 1.01 + 16.00 + 2 x 32.06 = 162.28, which means that the empirical formula is also the molecular formula for this compound.

The molecular formula is: $C_6H_{10}OS_2$

19. A hydrocarbon contains 82.7% carbon and the remainder hydrogen. Determine the empirical formula of the compound.

Solution: Knowing the % of C and H, we know the gram quantities in 100 g of sample:

82.7 g of C per 100 g of sample
17.3 g of H per 100 g of sample

Next we have to convert the gram quantities into moles. We have to use molar masses as conversion factors:

moles of C = 82.7 g x (1 mol/12.01 g) = 6.89 mol
moles of H = 17.3 g x (1 mol/1.01 g) = 17.1 mol

Carbon and hydrogen are then in the mol ratio:

mol C/mol H = 6.89/17.1 = 1/2.5 = 2/5

The mol ratio is 2 mol C : 5 mol H, and the ratio of atoms is then 2 atoms C : 5 atoms H. Thus the empirical formula is: C_2H_5

Answer: C_2H_5