Moles: Review Questions for the Test

Alice said, " Dear Rabbit, I see holes everywhere and not a mole to be found".
Rabbit answered, "Dearest Alice, why on a Sunday, the moles in the holes attend molar-mass, come down and join in the fun and games".

1. Calculate the number of molecules in:
   (a) 125.9 g of nitrogen gas (2.706×10^{24})
   (b) 0.678 mol of $NaCl_{(s)}$ (4.08×10^{23})
   (c) 0.00789 mol of $NaOH_{(aq)}$ (4.75×10^{21})
   (d) 265.0 g of $Ca(NO_3)_2$ (9.722×10^{23})

2. Calculate the number of atoms in:
   (a) 2.65 g of $SO_{2(g)}$ (7.47 x 1022)
   (b) 235.9 g of $CaBr_{2(s)}$ (2.311×10^{24})
   (c) 0.679 mol $H_2SO_{4(aq)}$ (2.86×10^{24})
   (d) 0.500 mol of $MgSO_{4(s)}$ (1.81×10^{24})

3. Calculate the mass of the following:
   (a) 0.50 mol $Ca(ClO_3)_2$ (100 g)
   (b) 1.875 mol copper(II)acetate (340.6 g)
   (c) 0.0097 mol $KOH_{(aq)}$ (0.54 g)
   (d) 0.235 mol $CuSO_4$ x 8 $H_2O_{(s)}$ (71.4 g)

4. Calculate the number of moles for each of the following:
   (a) 159.70 g of $K_2CrO_{4(s)}$ (0.82238 mol)
   (b) 73.89 g of lithium sulfate (1.608 mol)

5. Magnesium metal burns in air to produce magnesium oxide.
   (a) Write a balanced chemical chemical equation.
   (b) How many moles of oxygen are required to react with 9.70 g of magnesium to produce magnesium oxide? (O_2: 0.202 mol)
   (c) What mass of oxygen is required? (O_2: 6.47 g)
   (d) What mass of magnesium oxide is formed? (MgO: 16.2 g)

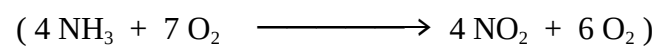
6. Consider the chemical reaction below:

$$MnO_{2(s)} + 4\,HCl_{(aq)} \longrightarrow MnCl_{2(aq)} + Cl_{2(g)} + 2\,H_2O_{(l)}$$

   (a) Name MnO_2 and $MnCl_2$.
   (b) How many moles of $Cl_{2(g)}$ are produced if 2.00 mol of HCl are consumed?
   (c) How many molecules of HCl are consumed if 3.01×10^{23} molecules of H_2O are formed?
   (d) What mass of $MnCl_2$ can be produced if 66.6 g of MnO_2 is reacted with excess HCl?

7. Potassium chlorate undergoes a decomposition reaction upon heating to produce potassium chloride and oxygen gas.
   (a) Write a balanced chemical equation for the decomposition of potassium chlorate.
   (b) How many grams of $KClO_3$ must be decomposed to yield 0.96 g of oxygen? (2.5 g)
   (b) How many moles of KCl will be produced during this same reaction? (0.020 mol)

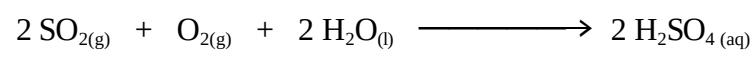
8. Ammonia undergoes combustion in oxygen gas to form nitrogen dioxide and water.
(a) Write a balanced chemical equation for the reaction of ammonia.
(b) How many mols of ammonia will be required to produce 10.5 mols of water?
(c) What mass of oxygen is required to produce 4.62 g of nitrogen dioxide?



($4\ NH_3 + 7\ O_2 \longrightarrow 4\ NO_2 + 6\ O_2$)

9. Iron (III) oxide is reduced by the use of carbon monoxide gas to produce iron metal and carbon dioxide gas.
(a) Write a balanced chemical equation for the reduction of iron(III) oxide by carbon monoxide to produce iron metal and carbon dioxide gas.
(b) When 479.1 g of iron (III) oxide reacts with excess carbon monoxide:
(i) how many moles of iron are produced?
(ii) how many grams of iron are produced?

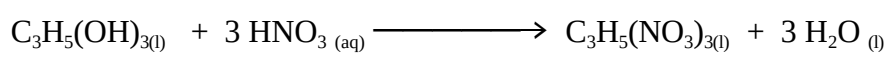
10. Sulfuric acid can be prepared by reacting sulfur dioxide, oxygen, and water.
The **unbalanced** chemical reaction is:



$2\ SO_{2(g)} + O_{2(g)} + 2\ H_2O_{(l)} \longrightarrow 2\ H_2SO_{4\ (aq)}$

If 15.0 g of oxygen and 50.0 g of sulfur dioxide are reacted with an unlimited quantity of water:
(i) What is the limiting reagent? (O_2)
(ii) How much sulfuric acid will be formed? (76.5 g)
(iii) What is the mass of the excess reactant remaining?

11. A chemist makes nitroglycerin, $C_3H_5(NO_3)_3$ from glycerol $C_3H_5(OH)_3$ and HNO_3. The balanced chemical reaction is listed below:

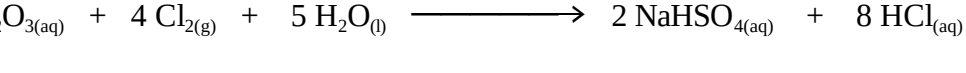


$C_3H_5(OH)_{3(l)} + 3\ HNO_{3\ (aq)} \longrightarrow C_3H_5(NO_3)_{3(l)} + 3\ H_2O_{\ (l)}$

If 4.15 g of glycerol and 13.55 g of HNO_3 are used to produce 8.88 g of nitroglycerin:

(a) What is the limiting reagent?
(b) What is the theoretical yield of nitroglycerin?
(c) What is the actual yield of nitroglycerin?
(d) What is the percentage yield of nitroglycerin?
(e) What mass of the excess reactant remains unconsumed?

12. Chlorine is used in the dry cleaning business to bleach fabrics. Any toxic chlorine gas is then destroyed by reacting it with sodium thiosulphate solution, $Na_2S_2O_{3(aq)}$, according to the following equation:



$Na_2S_2O_{3(aq)} + 4\ Cl_{2(g)} + 5\ H_2O_{(l)} \longrightarrow 2\ NaHSO_{4(aq)} + 8\ HCl_{(aq)}$

If 135.0 g of sodium thiosulphate solution, $Na_2S_2O_{3(aq)}$ reacts with 150.0 g chlorine gas and 238.0 g of water:
(a) Name $NaHSO_4$.
(b) Which reactant is the limiting reagent?
(c) What mass of sodium of $NaHSO_4$ is formed?

13. Consider the following compounds of nitrogen:

NH_3, N_2H_4, HNO_3, NO

a. Name the above compounds.
b. Which of the compound has the highest mass percentage of nitrogen?

14. A compound has the following percentage composition: 2.00 % H, 32.7 % S, and 65.3 % O. The molar mass of the compound is 98.0 g mol^{-1}. What is the empirical formula and the molecular formula of the compound ? (H_2SO_4)

15. A sample of a liquid used in dry-cleaning was found to consist of 10.06 % carbon, 89.10 % chlorine, with the remainder being hydrogen. The molar mass of the compound was determined to be 119.6 g mol^{-1}. Determine the empirical and the molecular formula of the compound.

16. A compound contains 12.8 % carbon, 2.10 % hydrogen, and 85.1 % bromine. Calculate the empirical formula and the molecular formula of the compound given that the molar mass is 188 g mol^{-1}.

17. The combustion of a 5.048 g of a compound containing carbon, hydrogen and oxygen gave: 7.406 g carbon dioxide, 3.027 g if water. Calculate the empirical formula. If the molar mass of the compound was determined to be 180.0 g mol^{-1}, determine the correct molecular formula.

18. 0.5438 g sample of a liquid consisting only of carbon, hydrogen, and oxygen was burned in pure oxygen and 1.039 g of carbon dioxide, and 0.6369 g of water was obtained. What is the empirical formula of the liquid? (C_2H_6O)

[Methodology for the type of problems # 17 & # 18:
a. g CO_2 —> mol CO_2 —> mol carbon —> g carbon —> % carbon
b. g H_2O —> mol H_2O —> 2 x mol hydrogen —> g hydrogen —> % hydrogen
c. % oxygen = 100.0 — (% carbon + % hydrogen) = % oxygen
Hence, calculate the empirical formula]

19. When 9.416 g of a hydrate of cobalt (II) thiocyanate, $Co(SCN)_2$ C x H_2O were heated, 7.197 g of the anhydrous solid remained. Calculate the number of water molecules in the formula of the hydrated salt.