Chem 115 POGIL Worksheet - Week 4
Moles & Stoichiometry
Answers

Key Questions & Exercises

1. The atomic weight of carbon is 12.0107 u, so a mole of carbon has a mass of 12.0107 g. Why doesn't a mole of carbon weigh 12 g?

   The atomic weight refers to the weighted average of masses of the isotopes comprising a naturally occurring sample of carbon. A 12.0107-g sample of natural carbon would contain an Avogadro's number of carbon atoms with all the naturally occurring mass numbers, present in their natural fractional abundance.

2. The atomic weight of oxygen is 16.00 u. What is the mass of a mole of $O_2(g)$? How many O_2 molecules does a mole of $O_2(g)$ contain? How many moles of oxygen *atoms* does a mole of $O_2(g)$ contain?

   The molecular weight of O_2 is 32.00 u, so a mole of O_2 would have a mass of 32.00 g and would contain 6.022×10^{23} O_2 molecules. Each O_2 molecule is composed of two oxygen atoms, so one mole of O_2 contains two moles of oxygen atoms.

3. The mole is sometimes described as "the chemist's dozen". How is a mole like a dozen?

   A dozen is twelve things, and a mole is an Avogadro's number (6.022×10^{23}) of things. The difference is that a dozen things (12) is an exact number, but a mole of things (6.022×10^{23}) is an inexact (measured) number.

4. Consider a 15.00-g sample of CO_2 (m.w. = 44.01u). How many moles of CO_2 are there in this sample?

   $(15.00 \text{ g } CO_2)(1 \text{ mol } CO_2 / 44.01 \text{ g } CO_2) = 0.3408 \text{ mol } CO_2$

5. How many CO_2 molecules are there in a 15.00-g sample of carbon dioxide?

   $(0.3408 \text{ mol } CO_2)(6.022 \times 10^{23} \text{ } CO_2 \text{ molecules/mol } CO_2) = 2.052_5 \times 10^{23} \text{ } CO_2 \text{ molecules}$

6. How many oxygen atoms are there in a 15.00-g sample of carbon dioxide.

   $(2.0525 \times 10^{23} \text{ } CO_2 \text{ molecules})(2 \text{ O atoms/}CO_2 \text{ molecule}) = 4.105 \times 10^{23} \text{ O atoms}$

7. Fluorine consists of a single isotope, ^{19}F, with a mass of 19.00 u. What is the mass in grams of a single fluorine atom?

$$1 \text{ mol F} = 19.00 \text{ g} = 6.022 \times 10^{23} \text{ atoms}$$

$$\left(\frac{19.00\,\text{g}}{6.022\times10^{23}\ \text{F atoms}}\right) = 3.155\times10^{-23}\,\text{g/F atom}$$

8. If you have data for the percent composition of a compound, element by element, do you need to know the size of the sample in order to figure out the empirical formula? Why or why not?

   No, because the mass ratios of all the elements are the same, regardless of the sample size, as required by the Law of Constant Composition.

9. How is the molecular formula of a molecular compound related to its empirical formula?

   The molecular formula is always a whole-number multiple of the empirical formula of the compounds. Sometimes the whole-number multiple is 1, when the molecular and empirical formulas are the same. When there is a difference, the molecular formula is usually a small whole-number multiple of the empirical formula. For example, the molecular formula of benzene is C_6H_6, and its empirical formula is CH. The whole number multiplier in this case is the integer 6. Note that in going from the empirical formula to the molecular formula, only the subscripts (including implied 1's) are multiplied. It is *not* correct to write the molecular formula of benzene as 6 CH.

10. A compound is found to contain 54.52% C, 9.17% H, and 36.31% O. What is the empirical formula of the compound? If the compound is found to have molecular weight of 88.12 u, what is the molecular formula?

    Assume exactly 100 g of compound. Then the percentages are numerically equal to the numbers of grams of each element.

    C: (54.52 g C)(1 mol C/12.01 g C) = 4.540 mol C

    H: (9.17 g H)(1 mol H/1.01 g H) = 9.08 mol H

    O: (36.31 g O)(1 mol O/16.00 g O) = 2.269 mol O

    To find the simplest whole number ratio between these numbers of moles, take the smallest number and divide it into all the numbers.

C: 4.540 mol/2.269 mol = 2.001 = 2

H: 9.08 mol/2.269 mol = 4.00_2 = 4

O: 2.269 mol/2.269 mol = 1

From this we obtain the empirical formula C_2H_4O. The formula weight for this is

$$\text{f.w.} = 2(12.01\text{ u}) + 4(1.01\text{ u}) + 16.00\text{ u} = 44.06\text{ u}$$

Dividing into the given molecular weight

$$\text{m.w./f.w.} = 88.12/44.06 = 2$$

This shows that the molecular formula is twice the empirical formula; i.e., $C_4H_8O_2$.

11. What is the empirical formula of an oxide of nitrogen whose composition is 25.94% nitrogen?

    The compound contains only nitrogen and oxygen. By subtraction, %O = 100% – 25.94% = 74.06%

    Assume 100 g of compound.

    N: (25.94 g N)(1 mol N/14.01 g N) = 1.852 mol N $\Rightarrow$ 1.852 mol/1.852 mol = 1 $\Rightarrow$ 2

    O: (74.06 g O)(1 mol O/16.00 g O) = 4.629 mol O $\Rightarrow$ 4.629 mol/1.852 mol = 2.5 $\Rightarrow$ 5

    Therefore, the empirical formula is N_2O_5. [Note that the originally found ratio N:O = 1:2.5 needed to be multiplied by 2 to make the integer ratio N:O = 2:5. Empirical formulas must have integer subscripts.]

12. A 2.554-g sample of a certain hydrocarbon is burned in excess oxygen, producing 8.635 g $CO_2(g)$ and 1.768 g $H_2O(l)$. If the molecular weight of the hydrocarbon is found to be 78.11 u, what is its molecular formula? [m.w. CO_2 = 44.01 u; m.w. H_2O = 18.02 u]

    A hydrocarbon only contains carbon and hydrogen. When burned, all of the carbon ends up in the $CO_2(g)$ produced, and all of the hydrogen ends up in the $H_2O(l)$ produced. First find the moles of $CO_2(g)$ and $H_2O(l)$, and then the moles of C and H in each of them. Then determine the empirical formula from the lowest whole-number ratio between the moles of C and moles of H. Finally, by dividing the given molecular weight by the formula weight, determine the factor by which the empirical formula must be multiplied to give the molecular formula.

$$\text{mol C} = (8.635\ \text{g CO}_2)\left(\frac{\text{mol CO}_2}{44.01\ \text{g CO}_2}\right)\left(\frac{\text{mol C}}{\text{mol CO}_2}\right) = 0.1962\ \text{mol C}$$

$$\text{mol H} = (1.768\ \text{g H}_2\text{O})\left(\frac{\text{mol H}_2\text{O}}{18.02\ \text{g H}_2\text{O}}\right)\left(\frac{2\ \text{mol H}}{\text{mol H}_2\text{O}}\right) = 0.1962\ \text{mol H}$$

The empirical formula is CH, for which f.w. = 12.01 + 1.01 = 13.02.

m.w./f.w. = 78.11/13.02 = 6

Therefore, the molecular formula is C_6H_6.

13. In the complete combustion of propane, how many moles of $CO_2(g)$ are produced per mole of $O_2(g)$?

    From the balanced equation,

    $$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(l)$$

    we see that 3 moles of 3 $CO_2(g)$ are produced for every 5 moles of $O_2(g)$ consumed. Thus, for every mole of $O_2(g)$, 3/5 mole of $CO_2(g)$ is produced.

14. In the complete combustion of propane, how many moles of $H_2O(l)$ are produced per mole of $O_2(g)$?

    From the balanced equation,

    $$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(l)$$

    we see that 4 moles of 4 $H_2O(l)$ are produced for every 5 moles of $O_2(g)$ consumed. Thus, for every mole of $O_2(g)$, 4/5 mole of $H_2O(l)$ is produced.

15. In the complete combustion of propane, how many moles of $H_2O(l)$ are produced per mole of $CO_2(g)$?

    From the balanced equation,

    $$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(l)$$

    we see that 4 moles of 4 $H_2O(l)$ are produced for every 3 moles of $CO_2(g)$ that are produced. Thus, for every mole of $CO_2(g)$ produced, 1⅓ mole of $H_2O(l)$ is produced.

16. A 1.638-g sample of propane is burned in excess oxygen. What are the theoretical yields (in grams) of $CO_2(g)$ and $H_2O(l)$ expected from the reaction? [m.w. C_3H_8 = 44.09 u, m.w. CO_2 = 44.01 u, m.w. H_2O = 18.02 u)

The balanced equation is

$$C_3H_8(g) + 2\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(l)$$

Using the stoichiometric relationships,

$$\text{mol } C_3H_8 = (1.638\text{ g } C_3H_8)\left(\frac{\text{mol } C_3H_8}{44.09\text{ g } C_3H_8}\right) = 0.03715 \text{ mol } C_3H_8$$

$$\text{g } CO_2 = (0.03715 \text{ mol } C_3H_8)\left(\frac{3 \text{ mol } CO_2}{\text{mol } C_3H_8}\right)\left(\frac{44.01\text{ g } CO_2}{\text{mol } CO_2}\right) = 4.905\text{ g } CO_2$$

$$\text{g } H_2O = (0.03715 \text{ mol } C_3H_8)\left(\frac{4 \text{ mol } H_2O}{\text{mol } C_3H_8}\right)\left(\frac{18.02\text{ g } H_2O}{\text{mol } H_2O}\right) = 2.678 \text{ g} H_2O$$

17. If 4.750 g of $CO_2(g)$ was obtained from the combustion of 1.638 g of propane, what was the percent yield?

$$\%\text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{4.750\text{ g}}{4.905\text{ g}} \times 100\% = 96.84\%$$

18. Define what is meant by the terms limiting reagent and excess reagent.

The limiting agent is the reagent that is present in shortest supply, on the basis of the balanced chemical equation, and which will be completely consumed in a complete reaction. The excess reagent is present in more than a sufficient amount to react with the limiting reagent, and some of it will remain after a complete chemical reaction. The amount of excess reagent left can be determined by calculating the amount consumed on the basis of its stoichiometric relationship with the limiting reagent, and then subtracting that amount from the amount that was initially present.

19. In the reaction 2 A + 3 B → products, if you have 0.500 mol A and 0.500 mol B, which is the limiting reagent? How much of the excess reagent will be left over, if compete reaction takes place?

Dividing the amount of each reagent by its stoichiometric coefficient in the balanced equation, we see we have

(0.500 mol A)/(2 mol/”set”) = 0.250 “set” of A and (0.500 mol B)/(3 mol/”set”) = 0.167 “set” of B

Thus, B is the limiting reagent and will be completely consumed. Based on the balanced equation, 2 moles of A are consumed for every 3 mole of B, so the amount of A that is consumed will be

mol A used = (0.500 mol B)(2 mol A/3 mol B) = 0.333 mol A

Subtracting from the original0.500 mol A that was present,

mol A left = (0.500 – 0.333) mol A = 0.167 mol A

20. What is the theoretical yield of $Ca_3(PO_4)_2(s)$ by the reaction

$$3\ Ca(OH)_2(s) + 2\ H_3PO_4(l) \rightarrow Ca_3(PO_4)_2(s) + 3\ H_2O(l)$$

when 10.00 g $Ca(OH)_2$ and 10.00 g H_3PO_4 are mixed? [f.w. $Ca(OH)_2$ = 74.10 u; m.w. H_3PO_4 = 97.99 u; f.w. $Ca_3(PO_4)_2$ = 310.18 u]

mol $Ca(OH)_2$ = (10.00 g $Ca(OH)_2$)(1 mol $Ca(OH)_2$/74.10 g $Ca(OH)_2$)
= 0.1349_{53} mol $Ca(OH)_2$

mol H_3PO_4 = (10.00 g H_3PO_4)(1 mol H_3PO_4/97.99 g H_3PO_4)
= 0.1020_{51} mol H_3PO_4

To identify the limiting reagent divide each number of moles by its stoichiometric coefficient in the balanced equation, thereby determining the number of "sets" of each.

"sets" $Ca(OH)_2$ = (0.1349_{53} mol $Ca(OH)_2$)(1 "set" $Ca(OH)_2$/3 mol $Ca(OH)_2$)
= 0.04498 "set" $Ca(OH)_2$

"sets" H_3PO_4 = (0.1020_{51} mol H_3PO_4)(1 "set" H_3PO_4/2 mol H_3PO_4)
= 0.05103 "set" H_3PO_4

$\Rightarrow$ $Ca(OH)_2$ limits, because it has the fewer "sets".

Now, use the moles of $Ca(OH)_2$ (*not the number of "sets"!*) in all subsequent calculations.

g $Ca_3(PO_4)_2$ = (0.1349_{53} mol $Ca(OH)_2$)(1 mol $Ca_3(PO_4)_2$/3 mol $Ca(OH)_2$)
x (310.18 g $Ca_3(PO_4)_2$/mol $Ca_3(PO_4)_2$)
= 13.95 g $Ca_3(PO_4)_2$

Periodic Table of the Elements

1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8	8B 9	10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1 H 1.008																	2 He 4.003
3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.61	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc [98]	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
55 Cs 132.9	56 Ba 137.3	57 La 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po [209]	85 At [210]	86 Rn [222]
87 Fr [223]	88 Ra [226]	89 Ac [227]	104 Rf [261]	105 Db [262]	106 Sg [263]	107 Bh [264]	108 Hs [265]	109 Mt [268]	110 Uun [269]	111 Uuu [272]	112 Uub [277]						

58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm [145]	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
90 Th 232.0	91 Pa 231.0	92 U 238.0	93 Np [237]	94 Pu [244]	95 Am [243]	96 Cm [247]	97 Bk [247]	98 Cf [251]	99 Es [252]	100 Fm [257]	101 Md [258]	102 No [259]	103 Lr [262]