Chem 115 POGIL Worksheet - Week 4
Moles & Stoichiometry

Why?

Chemists are concerned with mass relationships in chemical reactions, usually run on a macroscopic scale (grams, kilograms, etc.). To deal with the very large numbers of atoms and molecules in such samples, chemists developed the unit of the *mole* (abbreviated mol) and a unit of measure called the *molar mass*, which has units of g/mol. Next to the atomic theory, the mole concept is the most fundamental unifying idea in all chemistry.

Learning Objective

- Understand the relationship between the mole and Avogadro's number
- Understand the meaning of molar mass of a substance
- Understand how the mole concept is applied to determining empirical formulas from analytical data
- Understand how the mole concept allows prediction of the mass relationships between reactants and products in a chemical reaction
- Understand the concept of the limiting reagent

Success Criteria

- Convert between numbers of atoms, moles, and mass of sample by using Avogadro's number and the appropriate molar mass
- Calculate the empirical formula of a compound from percent composition data
- Calculate mass relationships between reactants and products, based on a balanced chemical equation
- Be able to determine the limiting reagent in a chemical reaction and calculate mass relationships for a chemical reaction based on it

Prerequisite

- Have read Chapter 3 in the text

Information

One of the most important ideas in chemistry is the mole concept. A mole of substance is that amount that contains as many elementary units (atoms, molecules, or formula units, depending on the nature of the substance) as there are atoms in exactly 12 grams of an isotopically pure sample of ^{12}C.

$$12 \text{ g } ^{12}\text{C (exactly)} = \text{mole } ^{12}\text{C atoms}$$

The number of atoms in such a sample defines **Avogadro's number** (symbol $\boldsymbol{N_A}$), which has been experimentally determined to be $N_A = 6.0221367 \times 10^{23}$. For most of our needs, a value of 6.022×10^{23} will be sufficiently precise. It follows that if we had a mole of atoms of some other element, that sample would weigh its atomic weight in grams. For example, for aluminum (at. wt. = 26.981538 u), a sample weighing 26.981538 g would contain a mole of aluminum atoms; i.e.,

$$26.981538 \text{ g Al} = \text{mole Al atoms}$$

The mass in grams of one mole of substance is called its **molar mass**. For an element or compound that is composed of molecules, the molar mass in grams is numerically equal to its molecular weight in atomic mass units. The molar mass of a molecular substance contains an Avogadro's number of molecules of the substance. For CO_2 (m.w. = 44.01 u),

$$\text{mole } CO_2 = 44.01 \text{ g } CO_2 = 6.022 \times 10^{23} \; CO_2 \text{ molecules}$$

Because each CO_2 molecule is composed of one carbon atom and two oxygen atoms, we could say that a mole of CO_2 contains one mole of carbon atoms and two moles of oxygen atoms. In general, it is useful to think of a mole as just an Avogadro's number of things. In the case of molecular compounds, that number of molecules has a mass in grams that is numerically equal to the substance's molecular weight.

For a compound described by an empirical formula (e.g., ionic compound, network solid, empirical formula unit of a molecular compound), the molar mass in grams is numerically equal to the formula weight in atomic mass units. The molar mass based on a formula weight contains an Avogadro's number of formula units of the substance. For NaCl (f.w. = 58.44 u),

$$\text{mol NaCl} = 58.44 \text{ g NaCl} = 6.022 \times 10^{23} \text{ NaCl formula units}$$

(Ionic compound - no molecules!)

Key Questions

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?

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?

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

Exercises

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?

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

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

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?

Information

The elemental composition of a compound can be determined experimentally by a variety of techniques. The results of chemical analysis are usually expressed in terms of weight percentages of each element in the compound, which can be converted into masses of each element for a given sample. The masses of each element can be used to calculate the numbers of moles of each element, from which the lowest whole number ratios between the moles of elements can be determined. These ratios are the same as the ratios between the numbers of individual atoms of each elements in the empirical formula. The strategy for converting analytical data into an empirical formula generally uses the following steps:

1. Convert weight percentages into grams of each element. Often it is helpful to assume a sample size of exactly 100 grams; then the given percentages are numerically equal to the number of grams of each element.

2. Convert grams of each element into moles of each element, using atomic weights.

3. Find the lowest whole number ratios among the moles of elements. To do this, start by dividing the smallest number of moles into each of the numbers of moles of elements (i.e., set the smallest number to 1). This may yield integers, or it may yield decimal results that correspond closely to rational fractions. For example, 1.25 : 2.75 = 1¼ : 2¾ = 5 : 11

4. Write the empirical formula, using the same whole-number ratios between atoms of each element as the ratios among moles of elements.

5. If the molecular weight is known, divide the formula weight for the empirical formula into the molecular weight to determine the number of formula units in the molecular formula. Using this integer factor, multiply all the subscripts (including any implied 1's) in the empirical formula to obtain the molecular formula.

Key Questions

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?

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

Exercises

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?

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

Information

One experimental method for determining the composition of organic compounds is **combustion analysis**, in which a weighed sample of the compound is burned in excess oxygen. In all cases all the carbon in the compound is converted to CO_2, and all the hydrogen is converted to H_2O, which can be separated from each other and weighed. The masses of carbon and hydrogen in the original sample can be calculated from the weights of CO_2 and H_2O. If the compound also contains oxygen, its amount can be obtained by subtracting the found masses of carbon and hydrogen from the total mass of the sample. These masses can then be converted into moles, from which the empirical formula can be obtained.

Exercise

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]

Information

For the burning (combustion) of propane gas the balanced equation is

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

When we first encountered reaction equations, we thought of this in terms of ratios among reactant and product species; e.g.,

> "For every molecule of $C_3H_8(g)$, five molecules of $O_2(g)$ are required to produce three molecules of $CO_2(g)$ and four molecules of $H_2O(l)$."

The relationships between individual reactant and product species and between moles of those species is multiplication by the constant Avogadro's number. Therefore, the ratios among moles of reactants and products are the same as between individual reactant and product species; e.g.,

> "For every mole of $C_3H_8(g)$, five moles of $O_2(g)$ are required to produce three moles of $CO_2(g)$ and four moles of $H_2O(l)$."

Thus, we can routinely interpret balanced chemical equations in terms of mole relationships. Once we know the numbers of moles, we can use the relationships between moles and molar masses of the various species to calculate masses of reactants and/or products, as needed. These mass relationships, made through moles, are called *stoichiometry* (Gk *stoicheon*, element + -*metry*, measure).

Using mole and mass relationships, we can calculate the mass of product that should be produced from a given amount of reactant when it is completely consumed in the reaction. This calculated amount of product is called the *theoretical yield.* In running real chemical reactions, it often occurs that less product is obtained than expected, for a variety of reasons. The amount obtained is the *actual yield*. A comparison of the actual yield to the theoretical yield, expressed as a percentage, is a statement of the *percent yield*; i.e.,

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Key Questions

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

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

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

Exercise

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)

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?

Information

Very often when we run a reaction between two or more substances, the amounts of reactants are not present in precisely the stoichiometric ratio indicated by the balanced chemical equation. In such cases, one reactant may be present in short supply, while other reactants may be present in abundance. Assuming complete reaction, the reactant in shortest supply will be completely consumed, but some amounts of the other reagents will be left over after the reaction is finished. In such cases, the amount of product obtained is limited by the reactant in shortest supply, which is called the **limiting reagent**. It is important to realize that the limiting reagent is present in shortest supply *on the basis of the stoichiometry of the balanced chemical equation in moles*; i.e., the mole ratios implied by the balanced equation. In some cases, the limiting reagent may be the substance present with larger absolute amount (either in grams or moles), but used in greater quantity in the balanced equation. In any case, the theoretical yield of product always will be limited by the stoichiometric relationship between the limiting reagent and products. Therefore, in any case where amounts of reactants are specified, determine the moles of each present, and then determine which reactant is the limiting reagent. All calculations of the theoretical yield for the reaction (or any other stoichiometric calculations) *must* be based on the amount of the limiting reagent, using the stoichiometric relationships in the balanced chemical equation.

How do we know which of two or more reactants is limiting? There are a number of ways to determine this. One of the most efficient is to see the amounts of each reagent in terms stoichiometric units, or what we might call “sets”(for want of a better term). For example, suppose we were building toy wagons and had 24 wheels and 15 wagon bodies, We would take the wheels in sets of four and the bodies in sets of one to build each wagon. Therefore we have 24/4 = 6 sets of wheels and 15/1 = 15 sets of bodies. As we assemble wagons the wheels will run out (the “limiting reagent”) before the bodies. Based on the wheels as the “limiting reagent” and their “stoichiometric” relationship to completed wagons (4 wheels/wagon), we could make only six wagons. In doing this, we would use six bodies, and we would have 15 - 6 = 9 bodies left over. Applying this approach to chemical reactions, if we take the number of moles of each reagent and divide that by its stoichiometric coefficient in the balanced equation, we will have a number for each that represents its number of reaction “sets”. The reagent that has the smallest number by this calculation is the limiting reagent; any other reagent is an excess reagent. We then use the number of moles of the limiting reagent (*not* its calculated number of “sets”) as the basis for all our further calculations, such as theoretical yield or amount of non-limiting reagent used. In short, *all calculations are based on the moles of the limiting reagent and the stoichiometric relationships implied by the balanced chemical equation*.

Key Questions

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

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?

Exercise

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]

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]