

CHM 101 GENERAL CHEMISTRY

FALL QUARTER 2008

Section 2

Lecture Notes – 10/15/2008

(last revised: 10/15/08, 4:30 PM)

3.3 The Mole: The mole (abbreviated mol) is a unit of measure that greatly facilitates our ability to count atoms by weighing them.

- The **mole** is defined as the number of atoms of ${}^{12}_6\text{C}$ in exactly 12 g of pure ${}^{12}_6\text{C}$.
- The **numerical value** of the mole is known as **Avogadro's number**. Here it is to 6 significant figures:

$$n = 6.02214 \times 10^{23}$$

- Four significant figures will generally be sufficient for our calculations:

$$n = 6.022 \times 10^{23}$$

- **How big is a mole?** One mole of marbles could cover the surface of the earth to a depth of 50 miles!
- How do chemists use the mole? The simple answer is that we use the mole to count molecules and atoms by weighing them. For example, if we weigh out 22.99 g of sodium, we have also weighed out one mole of sodium atoms, since sodium has an atomic mass of 22.99. And if we compare 22.99 g of sodium to 12.01 g of carbon, we have the same numbers of atoms (6.022×10^{23}) in both samples.
- **To summarize:** *The mole is defined such that a sample of a natural element with a mass equal to the element's atomic mass expressed in grams contains one mole of atoms.*
- **Sample Exercises** (taken from the text) **to work out:**
 - What is the mass in grams of 6 atoms of americium (Am)? (3.2, p. 83)
 - How many moles are there in 10.0 g of aluminum? How many atoms? (3.3, pp. 84-5)

- How many silicon atoms are there in a silicon chip whose mass is 5.68 g? (3.4, p. 85)
- How many moles are there in a sample of cobalt containing 5.00×10^{20} atoms? What is the mass in grams? (3.5, p. 85)

3.4 Molar Mass: We can use the mole to help us count molecules by weighing them.

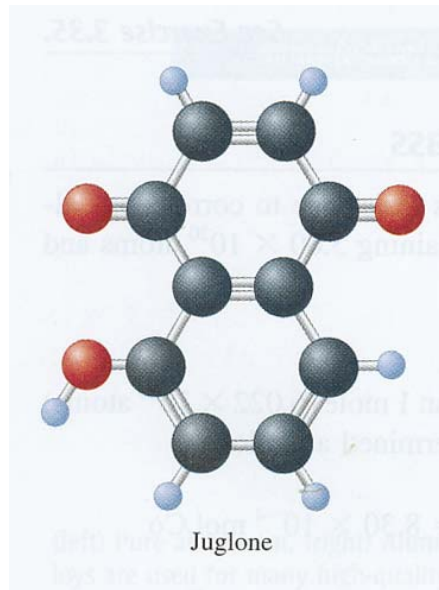
- Just as one mole of atoms contains 6.022×10^{23} atoms, one mole of molecules contains 6.022×10^{23} molecules. For example, one mole of methane (CH_4) contains 6.022×10^{23} molecules of CH_4 . To calculate the mass of a mole of CH_4 in grams, we add the atomic masses of its constituent atoms:

Mass of 1 mol C	=		=	12.01 g
Mass of 4 mol H	=	$4 \times 1.008 \text{ g}$	=	4.03 g
Mass of 1 mol CH_4	=		=	16.04 g

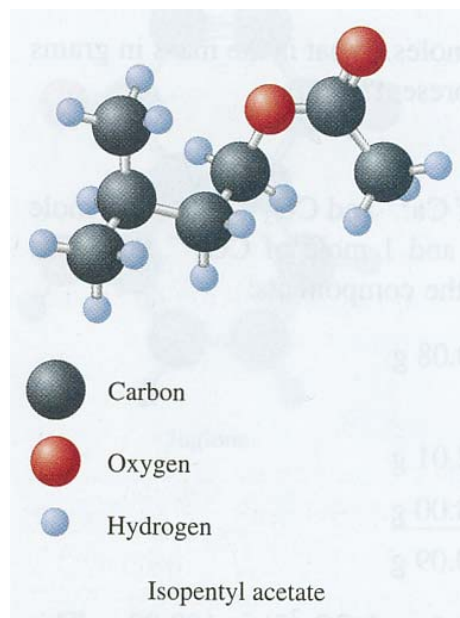
- **To summarize:** *The molar mass of a substance is the mass in grams of one mole of the substance. (An older name for molar mass is *gram-molecular-weight*, but we will avoid using this name.)*

- **Sample Exercises** (taken from the text) **to work out:**

- The molecular formula for juglone is $\text{C}_{10}\text{H}_6\text{O}_3$. (3.6, p. 86)
 - a) Calculate the molar mass of juglone.
 - b) How many moles of juglone are there in a 1.56×10^{-2} g sample of juglone?
- The formula for calcium carbonate is CaCO_3 . (3.7, p. 86-8)
 - a) Calculate the molar mass of calcium carbonate.
 - b) What is the mass (in grams) of 4.86 mols of calcium carbonate?



- c) What is the mass of the CO_3^{2-} ions in 4.86 mols of calcium carbonate?
- Isopentyl acetate ($\text{C}_7\text{H}_{14}\text{O}_2$) is the molecule responsible for the odor of bananas. Bees release about $1\ \mu\text{g}$ ($1 \times 10^{-6}\ \text{g}$) of isopentyl acetate when they sting, in order to attract other bees to the attack. (3.8, p. 88)
 - a) How many molecules of isopentyl acetate are released in a typical bee sting?
 - b) How many carbon atoms are there in this amount of isopentyl acetate?



3.5 Percent Composition

- We can specify the composition of a chemical compound in two ways:
 - By the numbers of each atom present in the molecular formula.
 - By the mass percent of each element in the compound. We can obtain these mass percentages from the molecular formula by the following procedure.
 - First we compute the molar mass (in grams/mol) of the compound. This also gives us the molar masses of each element present in the compound.
 - Then we compute the ratios of the masses of each element to the total molar mass.
 - When we express these ratios as percentages, we are done.

The example in the text is ethanol ($\text{C}_2\text{H}_5\text{OH}$). Let's work it out.

$$\text{Mass of 2 mol C} = 2 \times 12.01\ \text{g} = 24.02\ \text{g}$$

$$\text{Mass of 6 mol H} = 6 \times 1.008\ \text{g} = 6.05\ \text{g}$$

$$\begin{array}{r} \text{Mass of 1 mol O} = 1 \times 16.00 \text{ g} = 16.00 \text{ g} \\ \hline \text{Mass of 1 mol C}_2\text{H}_5\text{OH} = \phantom{1 \times 16.00 \text{ g}} = 46.07 \text{ g} \end{array}$$

$$\text{Mass \% C} = \frac{\text{mass of C (g)}}{\text{mass of C}_2\text{H}_5\text{OH (g)}} = \frac{24.02}{46.07} = 0.5214 = 52.14\%$$

$$\text{Mass \% H} = \frac{\text{mass of H (g)}}{\text{mass of C}_2\text{H}_5\text{OH (g)}} = \frac{6.048}{46.07} = 0.1313 = 13.13\%$$

$$\text{Mass \% O} = \frac{\text{mass of O (g)}}{\text{mass of C}_2\text{H}_5\text{OH (g)}} = \frac{16.00}{46.07} = 0.3473 = 34.73\%$$

We can check our arithmetic by adding the 3 percentages together. They should (and do) sum to 100.00%.

- **Sample Exercises** (taken from the text) **to work out:**
 - Calculate the mass percent of each element in carvone (C₁₀H₁₄O) (3.9, pp. 89-90)
 - Calculate the mass percent of each element in penicillin F (C₁₄H₂₀N₂SO₂) (3.10, pp. 90-1)

3.6 Determining the Formula of a Compound:

- If we know the mass percent composition of a compound, we can easily determine its empirical formula. Assume we have (exactly) 100 g of the compound and we know the mass % of each element in the compound.
 - First we calculate the masses (in grams) of each element in the 100 g sample.
 - Second we calculate the number of moles of each element in the sample by dividing the masses by the atomic masses.
 - Finally, we reduce the result to a set of small whole numbers. This gives us the empirical formula of the compound.

Let's work out the empirical formula for ethanol.

Element	C	H	O
% Composition	52.14 %	13.13%	34.73%

g/100 g of compound	52.14 g	13.13 g	34.73 g
molar mass element (g/mol)	12.01	1.008	16.00
mols/100 g compound	52.14/12.01	13.13/1.008	34.73/16.00
mols/100 g compound	4.3413	13.025	2.1706
<u>(above values)</u> 2.1706	2.000	6.000	1.000

The result is C_2H_6O . If we compare this result against the known formula for ethanol, (C_2H_5OH), we see we have the correct numbers of atoms, but no information about the structure. For all we can determine from the % composition, the true formula could be CH_3OCH_3 (dimethyl ether), or $C_4H_{12}O_2$, or any of an infinite number of possibilities where the C:H:O ratio is 2:6:1.

- We do not always start from the mass percent composition. The example in the book on p. 91 is a compound, newly prepared and purified in the laboratory. It is analyzed by burning it in an apparatus that allows the combustion products to be collected and weighed. Figure 3.5 shows a diagram of such a combustion device.

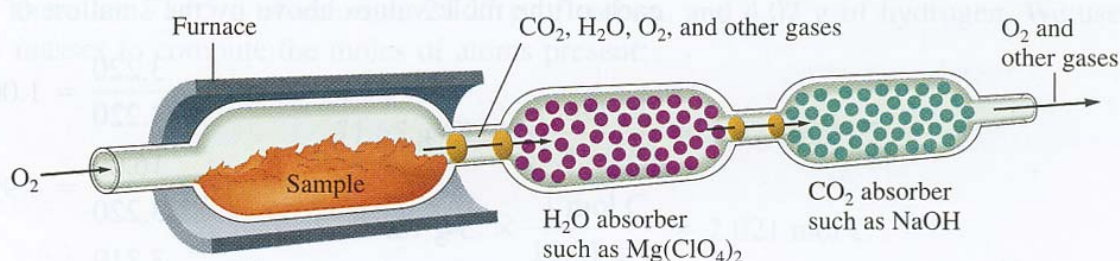


FIGURE 3.5

A schematic diagram of the combustion device used to analyze substances for carbon and hydrogen. The sample is burned in the presence of excess oxygen, which converts all its carbon to carbon dioxide and all its hydrogen to water. These products are collected by absorption using appropriate materials, and their amounts are determined by measuring the increase in masses of the absorbers.

In this example we are told that we start with 0.1156 g of a purified compound containing only carbon, nitrogen, and hydrogen. When we burn it in the presence of enough oxygen to make it burn completely, we find that the apparatus has

collected 0.1638 g of CO₂ and 0.1676 g of H₂O. From this information we are asked to determine the (empirical) formula of the compound. This is a complex problem, so we need to break it into a set of simple problems.

1. How many grams of carbon are there in 0.1638 g of CO₂?
What percent is this of the total mass of the sample?
2. How many grams of hydrogen are there in 0.1676 g of H₂O?
3. Since the only other element in the starting sample is nitrogen, this will be the 0.1156 g that we burned, less the masses of hydrogen and carbon that we determined in steps 1 & 2.
4. What is the mass percent composition of the starting compound. We will calculate this from the masses of C, H, and N that we determined in steps 1-3 and from the mass of compound we started with.
5. Now we can solve for the empirical formula in the same way we found the empirical formula for ethanol.

Let's work these simple problems in order:

1. 1 mol of CO₂ contains 1 mol (12.01 g) of C plus 2 mols (32.00) of O, giving us a molar mass of 44.01 g/mol for CO₂. Thus we can convert from g CO₂ to g C:

$$0.1638 \text{ g CO}_2 \times \frac{12.01 \text{ g C}}{44.01 \text{ g CO}_2} = 0.04470 \text{ g C}$$

And we can compute the % C in the starting compound:

$$\% \text{ C} = \frac{\text{mass of C recovered (g)}}{\text{mass of compound (g)}} = \frac{0.04470 \text{ g C}}{0.1156 \text{ g compound}} = 0.3867 \text{ or } 38.67\%$$

2. 1 mol of H₂O contains 2 mols (2.016 g) of H and 1 mol (16.00 g) of O, giving us a molar mass of 18.02 g/mol for H₂O. Thus we can convert from g H₂O to g H:

$$0.1676 \text{ g H}_2\text{O} \times \frac{2.016 \text{ g H}}{18.02 \text{ g H}_2\text{O}} = 0.01875 \text{ g H}$$

And we can compute the % H in the starting compound:

$$\% \text{ H} = \frac{\text{mass of H recovered (g)}}{\text{mass of compound (g)}} = \frac{0.01875 \text{ g H}}{0.1156 \text{ g compound}} = 0.1622 \text{ or } 16.22\%$$

3. Then the % N is:

$$\% \text{ N} = 100.00\% - 38.67\% - 16.22\% = 45.11\% \text{ N}$$

4. Now we can work out the mass percent composition of the compound.

Element	C	H	N
% Composition	38.67 %	16.22%	45.11%
g/100 g of compound	38.67 g	16.22 g	45.11 g
molar mass element (g/mol)	12.01	1.008	14.01
mols/100 g compound	38.67/12.01	16.22/1.008	45.11/14.01
mols/100 g compound	3.220	16.09	3.219
<u>(above values)</u> 3.219	1.000	4.998	1.000

5. Thus the empirical formula (formula determined by experiment) is CH_5N .

Suppose somehow that we know that the molar mass of the compound is 31.06 g/mol. How do we work out the true formula? Let's start by determining the molar mass of CH_5N .

Mass of 1 mol C	=	1 x 12.01 g	=	12.01 g
Mass of 5 mol H	=	5 x 1.008 g	=	5.04 g
Mass of 1 mol N	=	1 x 14.01 g	=	14.01 g
<hr/>				
Mass of 1 mol CH_5N	=		=	31.06 g

Since we know that the molar mass of the compound is the same as the molar mass of CH_5N , we can conclude that the true formula of the compound is (also) CH_5N .

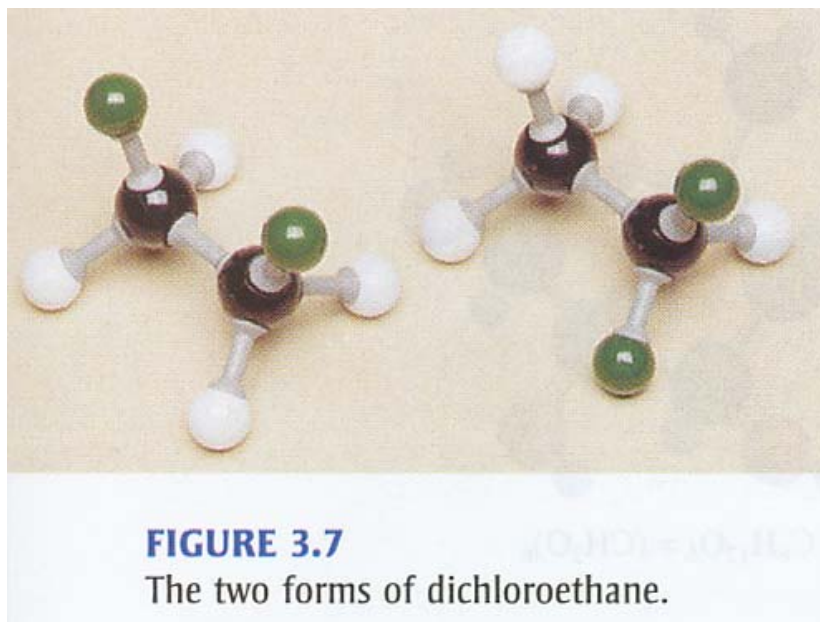
- Sample exercise 3.11 (p. 93): A compound with a molar mass of 98.96 g/mol has a mass percent analysis given by the top lines of the following table. The problem is to determine the empirical formula and the true formula of the compound:

Element	Cl	C	H
% Composition	71.65 %	24.27 %	4.07 %
g/100 g of compound	71.65 g	24.27 g	4.07 g
molar mass element (g/mol)	35.45	12.01	1.008
mols/100 g compound	71.65/35.45	24.27/12.01	4.07/1.008
mols/100 g compound	2.021	2.021	4.04
<u>(above values)</u> (smallest)	1	1	2

The empirical formula is ClCH_2 . Now we can determine the formula mass (i. e., the molar mass corresponding to the empirical formula):

Mass of 1 mol Cl	=	1 x 35.45 g	=	35.45 g
Mass of 1 mol C	=	1 x 12.01 g	=	12.01 g
Mass of 2 mol H	=	1 x 1.008 g	=	2.02 g
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Mass of 1 mol ClCH_2	=		=	49.48 g

If we divide the known molar mass (98.96 g/mol) by this formula mass (49.48 g/mol) we get 2. Thus the true formula of the compound is twice the empirical formula: $(\text{ClCH}_2)_2$ or $\text{Cl}_2\text{C}_2\text{H}_4$. FYI: Two possible structures for this molecule are shown in Figure 2.7:



- Sample exercise 3.12 (p. 94): A compound with a molar mass of 283.88 g/mol has a mass percent analysis given by the top lines of the following table. The problem is to determine the empirical formula and the true formula of the compound:

Element	P	O
% Composition	43.64 %	56.36 %
g/100 g of compound	43.64 g	56.36 g
molar mass element (g/mol)	30.97	16.00
mols/100 g compound	43.64/30.97	56.36/16.00
mols/100 g compound	1.409	3.523
<u>(above values)</u> (smallest)	1	2.5 = 5/2

In order that our empirical formula contain only whole numbers, we multiply by 2 to get 2 P atoms and 5 O atoms. Thus the empirical formula is P_2O_5 . Now we can determine the formula mass:

Mass of 2 mol P	=	2 x 30.97 g	=	61.94 g
Mass of 5 mol O	=	5 x 16.00 g	=	80.00 g
Mass of 1 mol P_2O_5	=		=	141.94 g

If we divide the known molar mass (283.88 g/mol) by this formula mass (141.94 g/mol) we get 2. Thus the true formula of the compound is twice the empirical formula: $(\text{P}_2\text{O}_5)_2$ or P_4O_{10} . FYI: The structure of this interesting molecule is shown in Figure 2.8:

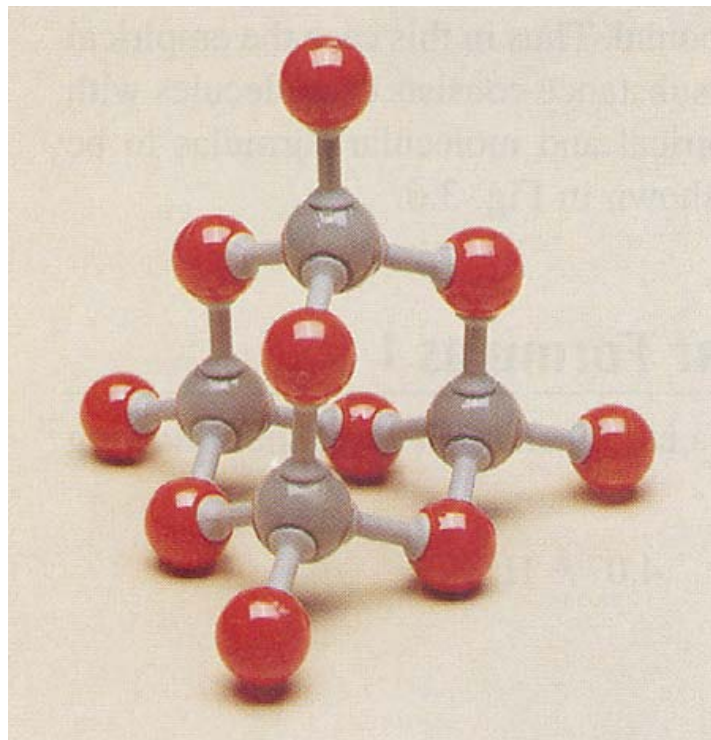


FIGURE 3.8

The structure of P_4O_{10} . Note that some of the oxygen atoms act as “bridges” between the phosphorus atoms. This compound has a great affinity for water and is often used as a desiccant, or drying agent.

- Let's take second look at sample exercise 3.12 (p. 94): Since we know that the compound has a molar mass of 283.88 g/mol, we can compute the true formula directly from the mass percent composition:

Element	P	O
% Composition	43.64 %	56.36 %
molar mass of compound (g/mol)	283.88	283.88

mass of element per mol of compound (g/mol)	0.4364 x 283.88	0.5636 x 283.88
mass of element per mol of compound (g/mol)	123.89	159.99
molar mass of element (g/mol)	30.97	16.00
mols element per mol compound	123.89/30.97	159.99/16.00
mols element per mol compound	4.000	10.00

Thus each mole of the compound contains 4 moles of P and 10 moles of O. This is just another way of saying that the true formula of the compound is P_4O_{10} .

- Sample exercise 3.13 (p. 95): Caffeine with a molar mass of 194.2 g/mol has a mass percent analysis given by the top lines of the following table. The problem is to determine the true formula of the compound:

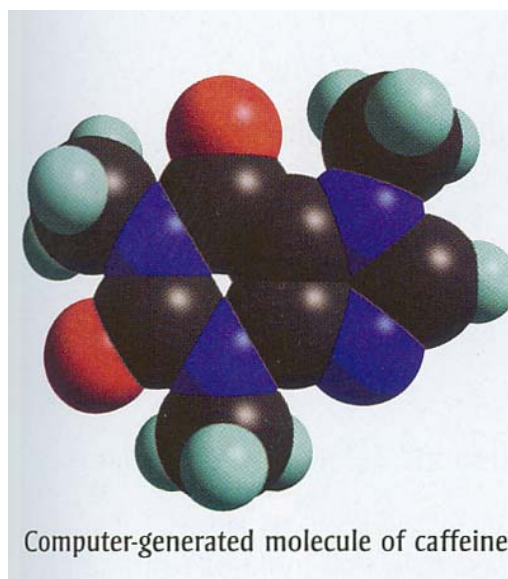
Element	C	H	N	O
% Composition	49.48 %	5.15 %	28.87 %	16.49%
molar mass of caffeine (g/mol)	194.2	194.2	194.2	194.2
mass of element per mol of caffeine (g/mol)	0.4948 x 194.2	0.0515 x 194.2	0.2887 x 194.2	0.1649 x 194.2
mass of element per mol of caffeine (g/mol)	96.09	10.00	56.07	32.02
molar mass of element (g/mol)	12.01	1.008	14.01	16.00

mols element				
per mol	96.09/12.01	10.00/1.008	56.07/14.01	32.02/16.00
caffeine				

mols element				
per mol	8.001	9.92	4.002	2.001
caffeine				

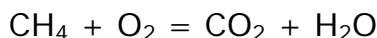
We round these results to the nearest integer and get the result. Caffeine has a true formula of $C_8H_{10}N_4O_2$. The figure illustrates its structure.

- Your text has a good summary of how to determine empirical and molecular (true) formulas. You will find it on p. 96.



3.7 Chemical Equations

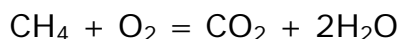
- Chemical Reactions:** A chemical reaction takes place whenever the atoms of one or more substances undergo reorganization. For example: when methane (CH_4) burns in air (It is actually reacting with the oxygen, O_2 , in the air). the methane and the oxygen that react together are replaced by carbon dioxide (CO_2) and water (H_2O).
- Chemical Equations:** We represent chemical reactions by writing chemical equations. In a chemical equation, the reactants are written on the left and the products of the reaction are written on the right. In the case of the reaction of methane with oxygen to produce carbon dioxide and water, the (unbalanced) equation is written:



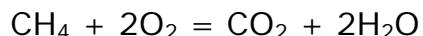
In this equation, the reactants are CH_4 and O_2 and the products are CO_2 and H_2O . Notice that the carbon-hydrogen bonds from the methane and the oxygen-oxygen bond in the oxygen have all disappeared and that new carbon-oxygen and hydrogen-oxygen bonds have been formed. (If you have sharp eyes, you might notice that the text uses a right pointing arrow in its version of this equation, where I use an equals sign. Both ways are correct, but I use the equals sign to remind you that this *is*

an equation whereas the text uses the arrow to indicate the direction of the reaction.)

- **Balancing Chemical Equations:** Recall that a chemical reaction does not create or destroy atoms; it simply rearranges them by breaking bonds and forming new bonds. Every atom that appears among the reactants must also appear among the products. Notice, however, that the equation we just wrote for the reaction of methane with oxygen does not obey this rule. We have 4 hydrogens on the left, but only 2 on the right, and we have 2 oxygen atoms on the left, but 3 on the right. We must fix the equation by balancing it. Carbons are already balanced, so we leave them alone (at least for now). We can start by fixing the imbalance of hydrogens. If we change the number of water molecules on the right from 1 to 2, we obtain:



Now there are 4 hydrogens on each side. And we now have 4 oxygens on the right, but we still have only 2 on the left. Let's change the number of oxygen molecules on the left from 1 to 2. This gives us:

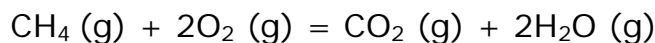


Now there are 4 oxygens on each side. And we left the numbers of carbons and hydrogens unchanged. Each side still has 1 carbon and 2 hydrogens. The equation is now balanced.

- **Information Content in a Balanced Chemical Equation:** A balanced chemical equation nominally tells us the numbers of molecules of each reactant that take part in the reaction, and it tells us the numbers of molecules of each product that are generated by the reaction. However, it also tells us the numbers of moles involved in the reaction, and since we already know how to convert between moles and grams, it also tells us the relative masses of the reactants and the products. For example, we can say that the balanced chemical equation of the previous bullet point tells us that 16 g of CH_4 will react with 64 g of O_2 to generate 44 g of CO_2 and 36 g of H_2O .
- **Physical States of Reactants and Products:** In addition to telling us the relative numbers of reactants and products involved in a reaction, a (balanced) chemical equation can also indicate the state of each reactant and product, i. e., whether a given reactant or product is a
 - gas (g),

- o liquid (l),
- o solid (s), or
- o dissolved in water (aq)

We do this by putting the appropriate abbreviation to the right of the chemical formula. Thus the methane/oxygen reaction can be written:



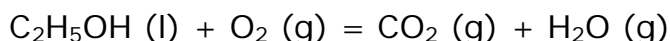
An example of a reaction involving all four of the states is the reaction of aqueous hydrochloric acid (HCl dissolved in water) with solid sodium hydrogen carbonate to generate carbon dioxide, water, and sodium chloride:



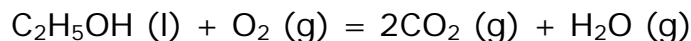
3.8 Rules for Balancing Chemical Equations: An unbalanced chemical equation is not very useful. You should get in the habit of checking any chemical equation you encounter to see if it is balanced. There are two fundamental rules to observe and a systematic trial-and-error procedure to follow when balancing an equation.

- **Fixed Identities:** The identities of the reactants and the products in a chemical reaction are fixed and must not be changed. For example, in a reaction that generates carbon dioxide, you cannot change its formula from CO_2 to CO_3 . The only way we are allowed to change numbers of atoms is to change the coefficients on the molecules.
- **Conservation of Atoms:** The same number of each type of atom must be found among the reactants as among the products. This is because no chemical reaction ever creates or destroys atoms. The reaction merely changes or rearranges the way they are bonded together.
- **A System for Balancing Equations:**
 - o Balancing a unbalanced chemical equation is a little like solving a Sudoku puzzle. You start by looking for an empty cell where only one of the nine possible digits can fit, i. e., an empty cell with a *uniquely determined* number. By filling the cell with this number, you eliminate the number as a choice for several other cells, and hopefully, one of these cells now has become uniquely determined. You keep repeating the process until the puzzle is done.

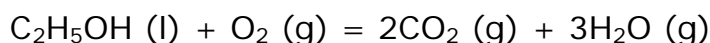
- o Now back to equation balancing. We start with the most complicated molecule among the reactants and products. (It helps if that molecule also contains all the atoms of a given type to be found on one of the sides of the equation. We can call these *unique occurrence* atoms.) We will use the example from p. 99 of your text: ethanol (C₂H₅OH) reacting with oxygen (O₂). We write out the unbalanced reaction. Then we choose ethanol as being the most complicated molecule, and we observe that it contains all the carbon atoms and all the hydrogen atoms to be found among the reactants:



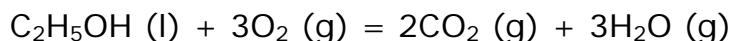
- o From the chosen molecule (ethanol in this case), we select a unique occurrence atom. In this example, we have our choice between C and H. We'll do them one at a time, starting with carbon. We have 2 carbon atoms on the left, both of them in the ethanol molecule, but only 1 on the right (in CO₂). We fix this by changing the coefficient of CO₂ from 1 to 2:



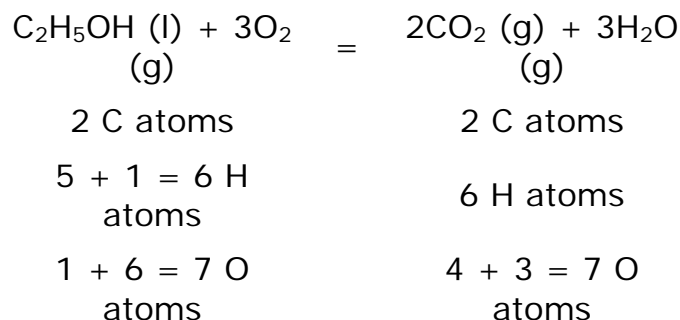
- o This balances carbon, and it increases oxygen on the right hand side, but it does not affect hydrogen. Now we can balance hydrogen. We have 6 hydrogen atoms on the left, all in the ethanol molecule, but only 2 on the right (in H₂O). We fix this by changing the coefficient of H₂O from 1 to 3:



- o Now we have balanced both of the unique occurrence type atoms from ethanol, the most complicated molecule in the reaction. What types of atom does this leave us to balance? In this case, there is only oxygen. The equation as it is now written has 7 oxygen atoms on the right, all of them in molecules that are balanced with respect to their other atoms. We have only 3 oxygens on the left, but if we try to increase the number to 7 by changing the coefficient of ethanol from 1 to 4, we lose our carbon and hydrogen balance. The only coefficient we can adjust is the 1 on the O₂. However, if we change it to 3, we gain the 4 oxygen atoms we need to finish balancing the equation:



- o Now we can check our final balance by tallying the atoms on each side:



- o Here is a summary of the process:
 - o Determine the reaction that occurs, i. e., identify the reactants, the products, and their physical states.
 - o Write the unbalanced equation.
 - o Balance the equation by inspection. Select the most complex of the molecules in the reaction and leave its coefficient set at 1. (However, you may need to adjust it later.) Starting with any unique occurrence type atoms in the selected molecule, balance each atom, type by type, reviewing each molecular coefficient for possible adjustment.

- **Sample Exercises from the Text:**

- o Exercise 3.14 (p. 100): Solid ammonium dichromate, $(NH_4)_2Cr_2O_7$, reacts spectacularly when it is ignited. Assuming the reaction products are solid chromium (iii) oxide, nitrogen gas (N_2) and water vapor, balance the reaction.
- o Exercise 3.15 (p. 101): At 1000 °C, ammonia gas (NH_3) reacts with oxygen gas (O_2) to form gaseous nitric oxide (nitrogen monoxide, NO) and water vapor. Balance the reaction.

3.9

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3.10