Chemistry – A Molecular Sciences Appendices

Appendix A
Stoichiometry of Elements and Compounds

A.1 INTRODUCTION

In this appendix, we will look at some of the quantitative relationships associated with the mass of elements and compounds. This subset of chemistry is called stoichiometry, a word derived from the Greek word "stoikheion", meaning element. Before we begin, we need to say a few words about the approach that we will take to the calculations in this appendix.

A.2 THE CONVERSION FACTOR APPROACH TO CALCULATIONS

If someone told you that she was "six", you might have a little trouble deciding what was meant. That person could be six years old, but if she were a college student, that would probably not be correct. She could weigh six tons or be six inches tall, but probably not. She is more likely six feet tall. The point is that there are two parts to a measurement or a piece of quantitative information; the "number" and the "unit". Six inches, six feet and six meters all have the same number, but are clearly different lengths. In scientific measurements or calculations, we must pay attention to both the number and the unit.

In order to convert the height of six feet to inches, most of us would say "multiply by 12" to give an answer of 72 inches. Although the result is correct, we did not multiply by 12; we actually multiplied by 1! Here's how:

We know the following equality: 12 in = 1 ft

Divide both sides by 1 ft: $\boxed{\frac{\textbf{12 in}}{\textbf{1 ft}}} = \frac{\textbf{1 ft}}{\textbf{1 ft}} = \textbf{1}$

The fraction in the box is called a "conversion factor" and it is equal to 1. In converting six feet to inches, we actually performed the following operation:

$$\textbf{6 ft} \times \frac{\textbf{12 in}}{\textbf{1 ft}} = \textbf{72 in}$$

In the above, the distance in feet is multiplied by a conversion factor to produce a distance in inches.

Notice that we did the operation on the numbers (six times twelve divided by one equals seventy-two) and on the units as well (feet times inches divided by feet equals inches; feet "cancel out"). Both the number and the unit changed, but the height did not (going from six feet to 72 inches, the person did not grow or shrink). This is what we would expect upon multiplication by 1. Although this is a trivial example, we will use this same **conversion factor** or **factor label** approach for nearly all of the stoichiometric calculations in this book.

A.3 MOLAR MASSES AND ATOMIC WEIGHTS OF THE ELEMENTS

The number under the symbol of an element on the periodic table is the element's atomic weight. It represents the "average atomic weight" or "average atomic mass" of the element because it is determined from the masses and abundance of the different isotopes of the element. Although there is a technical difference between weight and mass (weight depends on the gravitational force where you do the measurement), the two terms are often used interchangeably. We will use the term mass here, although your instructor may refer to "atomic weight". The mass of a single atom is the element's atomic weight expressed in units of amu, atomic mass units. The mass of a mole of atoms is the atomic weight expressed in units of grams/mol. The latter is referred to as the **molar mass** of the element. In this book, we use the symbol M_m to represent molar mass.

A.4 RELATING GRAMS, MOLES AND MOLAR MASS

Molar mass can also be used as a conversion factor. Using carbon as an example, the molar mass can be expressed as a fraction:

$$M_m = \frac{12.01\ \text{g}}{1\ \text{mol}}$$

The above is a conversion factor and is also equal to 1 because 12.01 grams of carbon and one mole of carbon are the same amount of carbon. The molar mass of an element can be used to convert between grams and moles.

Example 1

How many grams of sulfur are present in 0.250 moles of sulfur?

Solution:

The molar mass of sulfur from the periodic table is 32.07 g/mol. We start with the given information and apply the molar mass as a conversion factor.

$$0.250\ \text{moles} \times \frac{32.07\ \text{g}}{1\ \text{mol}} = 8.02\ \text{g}$$

Example 2

How many moles of copper are contained in 525 g of copper?

Solution:

The molar mass of copper from the periodic table is 63.55 g/mol. We again start with the given information and apply the molar mass as a conversion factor, but this time we use it in its reciprocal form (turn it "upside down") with moles in the numerator and grams in the denominator such that grams cancel out.

$$525\ \text{g} \times \frac{1\ \text{mol}}{63.55\ \text{g}} = 8.26\ \text{mol}$$

Comment:

In examples 1 and 2, a conversion factor was applied, that changed the number and the unit, but not the amount of substance (which is what you would expect upon "multiplying by 1"). 525 g of copper and 8.26 moles of copper are the same amount of copper, expressed in two different units.

A.5 COUNTING INDIVIDUAL ATOMS

Avogadro's number, which is $6.02\text{x}10^{23}\ \text{mol}^{-1}$, is the number of items present in a mole. Whether you are counting individual atoms, molecules, or trees, Avogadro's number can be used to convert between the number of items and the number of moles of items; *i.e.*, it is just another conversion factor.

$$\frac{6.02 \times 10^{23}\ \text{atoms}}{1\ \text{mole of atoms}} \qquad \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \text{mole of molecules}} \qquad \frac{6.02 \times 10^{23}\ \text{trees}}{1\ \text{mole of trees}}$$

Example 3

How many calcium atoms are in 2.25 moles of calcium?

Solution:

Start with the known information and apply Avogadro's number as a conversion factor:

$$2.25\ \text{mol Ca} \times \frac{6.02 \times 10^{23}\ \text{Ca atoms}}{1\ \text{mol Ca}} = 1.35 \times 10^{24}\ \text{Ca atoms}$$

Example 4

How many calcium atoms are in 2.25 grams of calcium?

Solution:

Avogadro's number tells us how many calcium atoms are in a mole of calcium, but the given information in this example is grams of calcium. Therefore, we must use the molar mass of calcium from the periodic table (40.08 g/mol) to convert grams to moles of calcium. We then apply Avogadro's number to obtain the number of atoms.

$$\textbf{2.25 g Ca} \times \frac{\textbf{1 mol Ca}}{\textbf{40.08 g Ca}} \times \frac{\textbf{6.02} \times \textbf{10}^{\textbf{23}} \textbf{ Ca atoms}}{\textbf{1 mol Ca}} = \textbf{3.38} \times \textbf{10}^{\textbf{22}} \textbf{ Ca atoms}$$

Comment:

Here we have combined two separate calculations (grams to moles and moles to number of atoms) by stringing together two conversion factors. We could, of course, have done the two separate calculations on two separate lines.

Example 5

What is the mass of $1.00\text{x}10^{22}$ bromine atoms?

Solution:

We apply Avogadro's number to determine the number of moles of bromine in the given number of atoms. Next we use, the molar mass of bromine (79.90 g/mol from the periodic table), to convert moles into grams.

$$\textbf{1.00} \times \textbf{10}^{\textbf{22}} \textbf{ Br atoms} \times \frac{\textbf{1 mol Br}}{\textbf{6.02} \times \textbf{10}^{\textbf{23}} \textbf{ Br atoms}} \times \frac{\textbf{79.90 g Br}}{\textbf{1 mol Br}} = \textbf{1.33 g Br}$$

Comment:

Again, two separate calculations were combined in one step.

A.6 CHEMICAL FORMULAS OF COMPOUNDS

A compound is a pure substance that is made up of more than one element. Compounds can be *ionic* (CAMS Chapter 4) or *covalent* as described in (CAMS Chapter 5). Covalent compounds are said to be *molecular* because they exist as discrete molecules, but ionic compounds exist as extended three-dimensional arrays of ions and not as discrete molecules.

The molecular formula of a compound tells us how many atoms of each element are in one molecule. A carbon dioxide molecule, which has the formula CO_2, contains one carbon atom and two oxygen atoms. This information is contained in the subscripts after each element. A molecule of sucrose ($C_{12}H_{22}O_{11}$) has 12 carbon atoms, 22 hydrogen atoms and 11 oxygen atoms. The subscripts also indicate the ratios of the elements. A dozen CO_2 molecules have one dozen carbon atoms and two dozen oxygen atoms. A million CO_2 molecules have one million carbon atoms and two million oxygen atoms. A mole of CO_2 molecules (we usually just say "a mole of CO_2") has one mole of carbon atoms and two moles of oxygen atoms. The atom ratio and the mole ratio of the elements are identical!

The chemical formula of an ionic compound does not tell us the number of atoms in a molecule because ionic substances are not molecular. However, it still gives the mole ratio of the elements. One mole of NaCl contains one mole of Na^{1+} ions and one mole of Cl^{1-} ions. Ionic compounds with polyatomic ions are somewhat more complicated. One mole of sodium sulfate, Na_2SO_4, contains two moles of Na^{1+} ions and one mole of SO_4^{2-} ions or two moles of sodium, one mole of sulfur and four moles of oxygen. An additional complexity comes from the way we write formulas of compounds containing polyatomic ions. Iron (III) nitrate has the formula $Fe(NO_3)_3$. This tells us that for every iron(III) ion (Fe^{3+}), there are three nitrate ions (NO_3^{1-}). Each nitrate ion contains one nitrogen atom and three oxygen atoms. Therefore, one mole of iron(III) nitrate contains one mole of iron, three moles of nitrogen and nine moles of oxygen.

A.7 MOLAR MASSES, MOLECULAR WEIGHTS AND FORMULA WEIGHTS OF COMPOUNDS

The molar mass of a compound can be determined from its chemical formula and the periodic table. The number obtained is sometimes referred to as the *molecular weight* or the *formula weight*. Molecular weight refers to the weight or mass of one molecule, in units of amu. Formula weight refers to the same quantity, but can be applied to substances that are not molecular, e.g., NaCl. Thus, even though iron(III) nitrate is not molecular, we can still talk about the mass of one formula unit, that is, one iron, three nitrogens and nine oxygens. All three terms are often used interchangeably, although molar mass is technically the only one that is in units of grams/mole.

To calculate a molar mass, we simply sum up the contributions of each element or atom. For carbon dioxide, CO_2, one carbon atom contributes 12.01 g/mol, the two oxygens together contribute (2)(16.00) = 32.00 g/mol. The molar mass is then 12.01 + 32.00 = 44.01 g/mol.

The same procedure is followed, whether the compound is covalent or ionic, as shown in the following examples.

Example 6

Calculate the molar mass of sucrose, $C_{12}H_{22}O_{11}$.

Solution:

Find the molar mass of each element on the periodic table and sum up the contributions.

Contribution from C:	(12 mol)(12.01 g/mol)	=	144.12 g
Contribution from H:	(22 mol)(1.008 g/mol)	=	22.18 g
Contribution from O:	(11 mol)(16.00 g/mol)	=	<u>176.00 g</u>
	Total	=	342.30 g/mol

Comment:

Notice that the molar mass of each element was multiplied by the number of times that element appeared in the chemical formula (as indicated by the subscripts in the formula).

Example 7

Calculate the molar mass of cobalt(II) phosphate, $Co_3(PO_4)_2$.

Solution:

Determine the contribution from each element and sum.

cobalt(II) phosphate :

Co:	(3 mol)(58.93 g/mol)	=	176.79 g
P:	(2 mol)(30.97 g/mol)	=	61.94 g
O:	(8 mol)(16.00 g/mol)	=	<u>128.00 g</u>
	Total	=	366.73 g/mol

Comment:

Even though the compound contains cobalt(II) ions, we use the molar mass of cobalt atoms. The difference in mass between a cobalt atom and a cobalt(II) ion is negligible, because the mass of an electron is so small compared to the mass of an atom.

A.8 RELATING GRAMS, MOLES AND MOLAR MASS OF COMPOUNDS

The molar mass of a compound can be used as a conversion factor in the same way as the molar mass of an element.

Example 8

How many grams of sucrose (M_m = 342.30 g/mol) are present in 0.125 moles of sucrose?

Solution:

We start with the given information and apply the molar mass as a conversion factor.

$$\mathbf{0.125\ mol \times \frac{342.30\ g}{1\ mol} = 42.8\ g}$$

Example 9

What is the mass of 2.50 moles of $Mg(ClO_4)_2$?

Solution:

In order to convert moles to grams, we need a molar mass from the chemical formula and the periodic table. Then we apply the molar mass as a conversion factor.

Magnesium perchlorate : $Mg(ClO_4)_2$

Molar mass :

Mg:	(1 mol)(24.31 g/mol)	=	24.31 g
Cl:	(2 mol)(35.45 g/mol)	=	70.90 g
O:	(8 mol)(16.00 g/mol)	=	<u>128.00 g</u>
	Total	=	223.21 g/mol

$$\mathbf{2.50\ mol \times \frac{223.21\ g}{1\ mol} = 558\ g}$$

Comment:

This example involves putting together several individual skills that you have learned to solve a problem. Rarely does an experiment or problem in science require only one skill. Often new discoveries and new applications result from putting known information together in new ways!

A.9 PERCENT COMPOSITION OF COMPOUNDS

The method we have used to calculate molar masses gives us a simple way to figure out the percentage of each element in a compound. Let's use ammonium nitrate as an example. This compound has the formula NH_4NO_3 and its molar mass is:

N:	(2 mol)(14.01 g/mol)	= 28.02 g
H:	(4 mol)(1.008 g/mol)	= 4.03 g
O:	(3 mol)(16.00 g/mol)	= <u>48.00 g</u>
		Total = 80.05 g/mol

This calculation not only tells us the mass of a mole of the compound , but it also tells us how many grams of each element are in a mole of the compound. For instance, in every 80.05 grams of the compound, there are 28.02 grams of nitrogen. Taking the ratio of grams of nitrogen to total grams of compound, we find:

$$\mathbf{\frac{28.02\ g}{80.05\ g} = 0.3500}$$

This number, 0.3500, is the mass fraction of nitrogen in the compound. This can easily be converted into a percent by multiplying by 100:

(0.3500)(100) = 35.00 %

Ammonium nitrate is 35.00% nitrogen by mass.

The ratio of grams of nitrogen to total grams of compound can be used as a conversion factor, too. If we want to know how many grams of nitrogen there are in 275 grams of ammonium nitrate, we would do the following :

$$\mathbf{275\ g\ NH_4NO_3 \times \frac{28.02\ g\ N}{80.05\ g\ NH_4NO_3} = 96.3\ g\ N}$$

Notice that grams of NH_4NO_3 cancel out in this calculation, leaving grams of N in the product. Also notice that the ratio 28.02/80.05 is identical to the ratio 96.3/275 (which is identical to 35/100 from the percent composition calculation). This technique is summed up in the following example.

Example 10

a) What is the percent nitrogen in potassium nitrate (KNO_3)?

Solution:

Percent nitrogen can be found from the data in a molar mass calculation**.** First, we need the chemical formula of potassium nitrate.

Potassium nitrate : KNO3

Molar mass:

K:	(1 mol)(39.10 g/mol)	=	39.10 g
N:	(1 mol)(14.01 g/mol)	=	14.01 g
O:	(3 mol)(16.00 g/mol)	=	48.00 g
	Total	=	101.11 g/mol

Percent nitrogen:

$$\mathbf{\frac{14.01\ g\ N}{101.11\ g\ total} \times 100\% = 13.86\ \%\ N}$$

b) How many grams of potassium are present in 137 g of KNO_3?

Solution:

The molar mass calculation tells us that, for every 101.11 g of compound, there are 39.10 g of K. Start with the given information and apply the ratio of grams of K to grams of compound as a conversion factor.

$$\mathbf{137\ g\ KNO_3 \times \frac{39.10\ g\ K}{101.11\ g\ KNO_3} = 53.0\ g\ K}$$

c) What mass of KNO_3 contains 125 g of potassium?

Solution:

The given quantity is 125 g of K. Clearly, we need more than 125 g of compound to give us 125 g of K, since the compound is only about 39% K. We can apply the same ratio as in part b, but turning it “upside down” to give us a result in units of grams of KNO_3.

$$\mathbf{125\ g\ K \times \frac{101.11\ g\ KNO_3}{39.10\ g\ K} = 323\ g\ KNO_3}$$

Comment:

We said that multiplication by a conversion factor is multiplication by 1. In this case, 101.11 g of KNO_3 is the same as “the amount of KNO_3 that contains 39.10 g of K”. In other words, the numerator and denominator represent the same amount of compound. This gives the effect of multiplying by 1.

A.10 COUNTING ATOMS, IONS AND MOLECULES IN COMPOUNDS

We use Avogadro’s number, 6.02×10^{23} mol^{-1}, to “count” the number of individual particles in a sample. The key in our calculations is to first find the number of moles of whatever item it is we wish to count, and then apply Avogadro’s number to convert from moles to individual items. The following examples all involve multiple steps, each of which we have discussed separately.

Example 11

a) How many water molecules are contained in 10.0 g of H_2O?

Solution:

In order to “count” water molecules, we first find how many moles of water are present in 10.0 g (M_m = 18.02 g/mol), and then apply Avogadro’s number.

$$\mathbf{10.0\ g\ H_2O \times \frac{1\ mol}{18.02\ g} \times \frac{6.02 \times 10^{23}\ molecules}{1\ mol} = 3.34 \times 10^{23}\ molecules}$$

b) How many hydrogen atoms are contained in 10.0 g of H_2O?

Solution to part b:

In order to "count" hydrogen atoms, we first find how many moles of hydrogen atoms are present in 10.0 g of water, and then apply Avogadro's number.

$$\mathbf{10.0\ g\ H_2O \times \frac{1\ mol\ H_2O}{18.02\ g\ H_2O} \times \frac{2\ mol\ H}{1\ mol\ H_2O} \times \frac{6.02 \times 10^{23}\ H\ atom}{1\ mol\ H} = 6.68 \times 10^{23}\ H\ atoms}$$

Comment:

Note that the two calculations are almost identical, the only difference being the term that converts moles of H_2O to moles of H in part b. This extra step was necessary because we are "counting" hydrogen atoms, not water molecules. The conversion factor comes directly from the chemical formula, where the subscript 2 (after the H) indicates that there are two hydrogen atoms per water molecule and two moles of hydrogen atoms per mole of water molecules. Note also that each calculation simply strings together a series of conversion factors. Each conversion factor is applied such that the term in the denominator "cancels out" the unit from the previous step. For example, the first conversion factor in each part is the molar mass of water simply turned "upside down" to put grams of water in the denominator.

Example 12

How many iron(III) ions are contained in 68.4 g of $Fe_2(SO_4)_3$?

Solution:

Many chemistry students find that the hardest part of a problem like this is figuring out where to start. Let's use some stepwise logic to figure out what we need for each step, working backwards to see where we should start. In order to count iron(III) ions, we need moles of iron(III). We can get moles of iron(III) if we know how many moles of iron(III) sulfate we have. We can get the moles of iron(III) sulfate from the 68.4 g and the molar mass. The molar mass is obtained from the chemical formula.

Molar mass :

Fe:	(2 mol)(55.85 g/mol)	=	111.70 g
S:	(3 mol)(32.07 g/mol)	=	96.21 g
O:	(12 mol)(16.00 g/mol)	=	192.00 g
	Total	=	399.91 g/mol

$$\mathbf{68.4\ g\ Fe_2(SO_4)_3 \times \frac{1\ mol\ Fe_2(SO_4)_3}{399.91\ g\ Fe_2(SO_4)_3} \times \frac{2\ mol\ Fe^{3+}}{1\ mol\ Fe_2(SO_4)_3} \times \frac{6.02 \times 10^{23}\ Fe^{3+}\ ions}{mol\ Fe^{3+}\ ions} = 2.06 \times 10^{23}\ Fe^{3+}\ ions}$$

Comment:

As is often the case, there are several steps required to solve the problem. Each individual step is not that hard; it is putting them together in the proper order that is the stumbling block for some students. Our approach was to use some logic, starting at the end and working back to the beginning, keeping track of everything needed along the way to solve the problem.

Example 13

How many grams of sucrose ($C_{12}H_{22}O_{11}$) contain 4.75x10^{25} carbon atoms?

Solution:

We are given the number of C atoms and are asked to find the mass of sucrose that contains them. (Notice that this problem works in the reverse direction to Examples 11 and 12, where grams were given and individual atoms were sought.) In order to calculate grams of sucrose, we need moles of sucrose (342.30 g/mol from Example 6). Moles of sucrose can be found from the moles of carbon atoms and the subscripts in the chemical formula. Moles of carbon atoms can be found from the number of individual atoms and Avogadro's number.

$$\mathbf{4.75 \times 10^{25}\ C\ atoms \times \frac{1\ mol\ C}{6.02\ x\ 10^{23}\ C\ atoms} \times \frac{1\ mol\ C_{12}H_{22}O_{11}}{12\ mol\ C} \times \frac{342.30\ g\ C_{12}H_{22}O_{11}}{1\ mol\ C_{12}H_{22}O_{11}} = 2.25 \times 10^{3}\ g = 2.25\ kg\ C_{12}H_{22}O_{11}}$$

Comment: As always, each conversion factor cancels the units from the previous step.

A.11 EXERCISES

Basic skills / elements

1. Arrange each of the following sets of elements in order of increasing atomic mass.
   a) phosphorus, sodium, iron, carbon
   b) manganese, potassium, fluorine, copper
   c) selenium, beryllium, arsenic, iron
   d) chlorine, zinc, scandium, helium
2. How many moles of titanium are contained in 15.5 g of titanium?
3. How many grams of sodium are contained in 1.25 moles of sodium?
4. How many moles of vanadium does $6.02x10^{22}$ vanadium atoms represent?
5. How many grams of magnesium are contained in 0.52 moles of magnesium?
6. How many moles of cobalt are contained in 66 kg of cobalt?
7. How many nickel atoms are contained in 0.50 moles of nickel?
8. How many moles of carbon are contained in 2.85 g of carbon?
9. How many moles of nitrogen does $7.5x10^{21}$ nitrogen atoms represent?
10. How many moles of germanium are contained in 25 mg of germanium?
11. How many oxygen atoms are contained in 6.25 moles of oxygen atoms?
12. How many grams of chromium are contained in $2.5x10^{-4}$ moles of chromium?
13. How many krypton atoms are contained in 1.22 moles of krypton?
14. How many kilograms of phosphorus atoms are contained in 38 moles of phosphorus atoms?
15. How many moles of silicon does $8.8x10^{24}$ silicon atoms represent?

Combined skills / elements

16. How many lithium atoms are contained in 1.0 g of lithium?
17. What is the mass of $2.5x10^{21}$ argon atoms?
18. How many aluminum atoms are contained in 1.5 kg of aluminum.
19. What is the mass of $3.5x10^{25}$ iron atoms?
20. What is the mass, in grams, of $8.25x10^{23}$ silver atoms?
21. How many barium atoms are contained in 0.050 g of barium?

Basic skills / compounds

22. Calculate the molar mass of the following compounds:
    a) hydrazine, N_2H_4
    b) acetic acid, $HC_2H_3O_2$
    c) pyridine, C_5H_5N
    d) succinic acid, $C_4H_6O_4$
    e) TNT, $C_7H_5N_3O_6$
    f) calcium nitrate, $Ca(NO_3)_2$
    g) potassium chromate, K_2CrO_4
    h) cobalt(II)citrate, $Co_3(C_6H_8O_7)_2$
23. Use the following molar masses to answer these questions:
    cobalt(II) iodide, CoI_2 : 312.74 g/mol
    morphine, $C_{17}H_{19}NO_3$: 285.35 g/mol
    a) How many moles of CoI_2 are contained in 10.0 g of CoI_2?
    b) How many grams of morphine are contained in $2.0x10^{-5}$ moles of morphine?
    c) How many moles of morphine are contained in 35 mg of morphine?
    d) How many kilograms of CoI_2 are contained in 12.0 moles of CoI_2?
24. What is the % carbon in each of the following compounds:
    a) CH_4 b) $C_6H_{12}O_6$ c) C_7H_8

Combined skills / compounds

25. How many grams of potassium are contained in 8.00 g of KCl?

26. How many chloride ions are contained in 8.00 g of KCl?

27. How many grams of sulfur are contained in 1.00 kg of thiophene, C_4H_4S (M_m = 84.14 g/mol)?

28. How many sulfur atoms are contained in 3.55 g of C_4H_4S?

29. How many carbon atoms are contained in 3.55 g of C_4H_4S?

30. If you wanted to obtain $1.00x10^5$ g of nitrogen, what mass of NH_3 would you need?

31. If you wanted to obtain $1.00x10^5$ g of nitrogen, what mass of NH_4NO_3 would you need?

32. How many oxygen atoms are contained in 5.25 g of $Fe(NO_3)_3$?

33. How many carbon atoms are contained in 65 g of quinine, $C_{20}H_{24}N_2O_2$ (M_m = 324.41 g/mol)?

34. How many grams of carbon are contained in 65 g of quinine?

Answers:

1. **a)** C(12) < Na(23) < P(31) < Fe(56)
b) F(19) < K(39) < Mn(55) < Cu(64)
c) Be(9) < Fe(56) < As(75) < Se(79)
d) He(4) < Cl(35) < Sc(45) < Zn(65)

2. 0.324 mol Ti

3. 28.8 g Na

4. 0.100 mol V

5. 13 g Mg

6. $1.1x10^3$ mol Co

7. $3.0x10^{23}$ atoms of Ni

8. 0.237 mol C

9. 0.012 mol N (or 0.0062 mol N_2

10. $3.4x10^{-4}$ mol Ge = 0.34 mmol Ge

11. $3.76x10^{24}$ atoms of O

12. 0.013 g Cr

13. $7.35x10^{23}$ atoms of Kr

14. 1.2 kg of P atoms

15. 15 mol Si

16. $8.7x10^{22}$ atoms of Li

17. 0.17 g Ar

18. $3.3x10^{25}$ atoms of Al

19. 3.2 kg Fe

20. 148 g Ag

21. $2.2x10^{20}$ atoms of Ba

22. **a)** 32.05 **b)** 60.34 **c)** 79.1 **d)** 118.1
e) 227.1 **f)** 164.1 **g)** 194.2 **h)** 560.9

23. **a)** 0.0320 mol CoI_2 **b)** 0.0057 mol morphine
c) $1.2x10^{-4}$ g morphine **d)** 3.75 kg CoI_2

24. **a)** 75.0% **b)** 40.0 % **c)** 91.3 %

25. 4.19 g K

26. $6.46x10^{22}$ Cl^{1-} ions

27. 381 g S

28. $2.54x10^{22}$ atoms of S

29. $1.02x10^{23}$ atoms of C

30. 122 kg NH_3

31. 286 kg NH_4NO_3

32. $1.18x10^{23}$ atoms of O

33. $2.4x10^{24}$ atoms of C

34. 48

Appendix B
Gases

B.1 GAS MOLECULES ACTING COLLECTIVELY

According to the Kinetic Molecular Theory (CAMS Section 7.2), gases are in constant random motion and the average kinetic energy is proportional to the absolute temperature. The kinetic energy of a molecule is $1/2\ mv^2$, so the average speed of the molecules also depends on the absolute temperature. The average O_2 molecule moves at about 1,000 mph on a nice day.

However, it is the collective action of large numbers of molecules that we sense or measure as gases not the individual molecules. When you fan your face, you feel some wind, which is the effect of molecules in the air hitting your face. You cannot sense the individual molecules hitting your skin, for they are much too small, but you can feel their collective action.

The fact that a balloon expands when it is filled with a gas also shows how gas molecules act collectively. The molecules in the balloon are moving around with an average kinetic energy dictated by the temperature. When a molecule strikes the inside wall of the balloon, it exerts a force on the balloon and pushes it outward. The collective forces of all of the molecules inside the balloon pushing outward cause it to stay inflated. At the same time, the gas molecules in the outside air are striking the outer surface of the balloon exerting a force pushing inward. The size of the balloon adjusts until the force from the “strikes” on the outside balances the force from the “strikes” on the inside.

The collective force of all of the molecules pushing on the inside wall of the balloon results in pressure. The collective force per unit area or pressure of the gas depends on the number of collisions with the walls per second and the force of each collision. The common units are pounds per square inch, atmosphere (atm), the millimeter of mercury (mm Hg or torr) and the pascal, the SI unit of pressure (N/m^2).

B.2 RELATIONSHIP OF PRESSURE TO OTHER GAS PROPERTIES

Let’s analyze what happens to the pressure of a gas as the temperature, the number of molecules and volume of the gas change. Imagine a cylinder with a movable piston. The gas molecules in the piston have kinetic energy (are moving) and are hitting the walls of the cylinder, the piston and each other.

If the temperature of the gas is increased, the molecules will move faster and will strike the piston more frequently and with more force. Consequently the pressure increases. If more molecules are added to the cylinder (moles of gas increase), the frequency of collisions and therefore the pressure increases. Finally, if we push the piston down and compress the gas to a smaller volume, the gas molecules have less distance to travel before they hit the piston, and they collide with the piston more frequently. Thus, a decrease in volume will result in an increase in pressure.

The relationships among the pressure, volume, number of moles, and temperature of a gas are summed up quantitatively in the *ideal gas law*:

$$PV = nRT$$

P is the pressure in atmospheres (atm), V the volume in liters (L), n the number of moles, T the absolute temperature in kelving, and R is a constant called the ideal gas constant, which is $0.0821\ L{\cdot}atm{\cdot}K^{-1}{\cdot}mol^{-1}$. However, when using SI units, P is expressed in pascals, V in m^3, and $R = 8.314\ J{\cdot}mol^{-1}{\cdot}K^{-1}$.

B.3 USING THE IDEAL GAS LAW

The ideal gas law contains four experimental quantities: pressure, volume, temperature, and number of moles. If we know three of the quantities, we can solve for the fourth. The first step is always to ensure that the units on the known quantities are consistent with our value of R. The following examples show how this can be done.

Example 1

What is the volume of 1.00 mole of gas at 1.00 atm and 0 ºC?

Solution:

n is in mol and P in atm, but T is in ºC, not K. To convert from ºC to K, we add 273: 0 ºC + 273 = 273 K = T. Next, rearrange the ideal gas law to solve for the unknown, which in this case is the volume.

$$V = \frac{nRT}{P} = \frac{(1.00\ \text{mol})\ (0.0821\ \text{L}\cdot\text{atm}\cdot\text{K}^{-1}\cdot\text{mol}^{-1})\ (273\ \text{K})}{1.00\ \text{atm}} = 22.4\ \text{L}$$

Comment:

The conditions 0 ºC and 1.00 atm are often referred to as the Standard Temperature and Pressure (STP) for a gas. The volume at STP is 22.4 L, which is an experimental, not theoretical, number that students often remember from high school chemistry.

Example 2

An experiment yields 5.67 mL of CO_2(g) at 26 ºC and 782 mm Hg. How many grams of CO_2 is this?

Solution:

We are given V, T, and P, but none have the correct units for our value of R, so we convert each into the proper units.

$$V = 5.67\ \text{mL} \times \frac{1\ \text{L}}{1000\ \text{mL}} = 0.00567\ \text{L} \qquad T = 26\ ^{o}\text{C} + 273 = 299\ \text{K}$$

$$P = 782\ \text{mm Hg} \times \frac{1.00\ \text{atm}}{760\ \text{mm Hg}} = 1.03\ \text{atm}$$

Next, rearrange the ideal gas law to solve for n.

$$n = \frac{PV}{RT} = \frac{(1.03\ \text{atm})\ (0.00567\ \text{L})}{(0.0821\ \text{L}\cdot\text{atm}\cdot\text{K}^{-1}\cdot\text{mol}^{-1})\ (299\ \text{K})} = 2.38 \times 10^{-4}\ \text{mol}\ CO_2$$

To find grams, apply the molar mass (M_m = 44.01 g/mol) as a conversion factor as done in Appendix A,

$$2.38 \times 10^{-4}\ \text{mol}\ CO_2 \times \frac{44.01\ \text{g}\ CO_2}{\text{mol}\ CO_2} = 0.0105\ \text{g}\ CO_2$$

Comment:

Notice that five and two-thirds mL of a gas seems like a very small amount in terms of mass! That is because we are used to weighing out solids and liquids, which are much denser than gases.

B.4 EXERCISES

1. What volume does 0.50 moles of CO_2 occupy at 725 mm Hg and 25 ^{o}C?
2. How many moles of He occupy a 2.50-L flask whose pressure is 945 mm Hg at 75 ^{o}C?
3. What is the pressure exerted by 28.8 g of N_2 contained in a 4.25 L-flask at 0 ^{o}C?
4. What volume does 5.8 moles of O_2 occupy at 285 mm Hg and -78 ^{o}C?
5. What is the temperature of 5.0 moles of N_2 contained in a 20.0 L-tank at a pressure of 7.5 atm?
6. What volume does 6.32 g of NH_3 occupy at 745 mm Hg and 25 ^{o}C?
7. How many moles of CH_4 occupy a 10.0-L tank whose pressure is 3.5 atm at 30 ^{o}C?
8. What volume does 0.45 g of Ar occupy at 1.25 atm and 27 ^{o}C?
9. What is the pressure exerted by 3.5 moles of H_2 contained in a 2.0-L tank at 27 ^{o}C?
10. What volume does 0.75 moles of N_2 occupy at 760 mm Hg and 0 ^{o}C?
11. What is the temperature of 7.65 g of He contained in a 6.25 L flask at a pressure of 1.75 atm?
12. How many moles of HCl gas occupy a 4.5 L tank whose pressure is 1875 mm Hg at 27 ^{o}C?
13. For this question, note that M_m = g/mol and density, d = mass/volume.
    a) What is the density of helium in g/L at 1.00 atm and 27 ^{o}C?
    b) What is the density of nitrogen in g/L at 1.00 atm and 27 ^{o}C?

ANSWERS:

1. 13 L
2. 0.109 mol
3. 5.42 atm
4. 2.5×10^2 L
5. 365 K = 92 ^oC
6. 9.26 L
7. 1.4 mol
8. 0.22 L
9. 43 atm
10. 17 L
11. 69.7 K = -203 ^oC
12. 0.45 mol
13. a) 0.163 g/L
13. b) 1.14 g/L

Appendix C
Molarity

C.1 MOLARITY AND THE MOLE

The molar mass is the mass of a mole of a pure substance while the *molarity*, M, is the number of moles of a pure substance contained in a liter of a *solution*.

$$\text{molarity} = \frac{\text{moles}}{\text{liter}} = \frac{n}{V}$$

One liter of a solution that contains 0.1 moles of sugar ($C_{12}H_{22}O_{11}$) is 0.1 M, or the solution is 0.1 molar in sugar. It can also be represented as: $[C_{12}H_{22}O_{11}] = 0.1$ M, which is read as "the molar concentration of sugar is 0.1 molar."

C.2 MOLARITY AS A CONVERSION FACTOR

Molarity is used to convert between moles of substance and liters of solution.

Example 1

How many moles of NaCl are in 325 mL of 0.25 M NaCl solution?

Solution:

We first convert mL to L, and then apply molarity as a conversion factor.

$$325\text{ mL} \times \frac{1\text{ L}}{10^3\text{ mL}} \times \frac{0.25\text{ moles NaCl}}{1\text{ L}} = 0.081\text{ mol NaCl}$$

Example 2

How many mL of 5.0 M HCl contains 0.15 moles of HCl?

Solution:

Our known quantities are moles of HCl and molarity. We start with moles and apply molarity as a conversion factor. The final step is to convert liters to milliliters.

$$0.15\text{ mol HCl} \times \frac{1\text{ L}}{5.0\text{ mol HCl}} \times \frac{10^3\text{ mL}}{\text{L}} = 30\text{ mL}$$

Comment:

Note that our definition of molarity is turned upside down, and we were careful to write the units L and moles HCl in the numerator and denominator. Do not use M as the units of the conversion factor.

C.3 CONCENTRATIONS OF IONS

When ionic compounds dissolve, individual solvated ions are formed. (Recall that there are no molecules in ionic compounds.) When we refer to a 0.1 M NaCl solution, we mean that the solution has 0.1 moles of NaCl units in every liter. We can also determine the concentrations of the individual ions from the chemical formula. As discussed in Appendix A, the chemical formula relates moles of compound to moles of each element in the compound. In one mole of Na_2SO_4 there are two moles of Na^{1+} ions and one mole of SO_4^{2-} ions.

Example 3

a) What is the concentrations of Na^{1+} ion and b) the total concentration of all ions in 0.25 M NaCl?

Solution:

Part a), knowing the molarity of the compound and the formula, we can easily see that for every mole of NaCl, there is one mole of Na^{1+}.

$$\frac{\textbf{0.25 mol NaCl}}{\textbf{1 L}} \times \frac{\textbf{1 mol Na}^{1+}}{\textbf{1 mol NaCl}} = \frac{\textbf{0.25 mol Na}^{1+}}{\textbf{1 L}} = \textbf{0.25 M Na}^{1+}$$

Part b), for every mole of NaCl, there is one mole of Na^{1+} ions and one mole of Cl^{1-} ions, which adds up to two moles total of ions. Again, start with the concentration of the compound and find the concentration of ions.

$$\frac{\textbf{0.25 mol NaCl}}{\textbf{1 L}} \times \frac{\textbf{2 mol ions}}{\textbf{1 mol NaCl}} = \frac{\textbf{0.50 mol ions}}{\textbf{1 L}} = \textbf{0.50 M ions}$$

Comment:

In each step, we have used the moles of ions per mole of compound as a conversion factor much like we did in Examples 11-13 in Appendix A. Notice that this conversion gets us directly to the molarity of the ions (moles of ions per liter).

Example 4

a) What is the concentration of chloride ions in a 0.1 M $CaCl_2$?

b) What is the total concentration of ions in a 0.1 M solution of $CaCl_2$?

Solution:

As in Example 3, we start with the solution concentration and apply a conversion factor that converts moles of compound to moles of individual ions. In $CaCl_2$, the chemical formula tells us that there are two moles of Cl^{1-} in every mole of compound, and three moles of total ions (1 mole Ca^{2+}, 2 moles Cl^{1-}) in every mole of compound.

$$\frac{\textbf{0.1 mol CaCl}_2}{\textbf{1 L}} \times \frac{\textbf{2 mol Cl}^{1-}}{\textbf{1 mol CaCl}_2} = \frac{\textbf{0.2 mol Cl}^{1-}}{\textbf{1 L}} = \textbf{0.2 M Cl}^{1-}$$

$$\frac{\textbf{0.1 mol CaCl}_2}{\textbf{1 L}} \times \frac{\textbf{3 mol ions}}{\textbf{1 mol CaCl}_2} = \frac{\textbf{0.3 mol ions}}{\textbf{1 L}} = \textbf{0.3 M ions}$$

Comment:

We have taken some care to write down the units in detail for each conversion, but once you understand chemical formulas and the fact that ionic compounds dissolve to form individual ions, you will be able to do these calculations in your head!

C.4 MAKING SOLUTIONS

One of the most common tasks in the chemistry laboratory is making solutions of desired concentrations. In this section, we will explore how to make solutions starting with a solid solute.

Example 5

How many grams of Na_2CO_3 (M_m = 105.99 g/mol) are required to make 0.500 L of a 0.10 M Na_2CO_3 solution?

Solution:

We have a target volume and molarity for our solution, and so we can calculate the necessary moles of Na_2CO_3. We can then use the molar mass to calculate the necessary grams of Na_2CO_3.

$$\textbf{0.500 L} \times \frac{\textbf{0.10 mol}}{\textbf{1 L}} \times \frac{\textbf{106 g}}{\textbf{1 mol}} = \textbf{5.3 g}$$

Example 6

What is the molar concentration of a 2.5 L of solution that contains 254 g of Na_2CO_3?

Solution:

In this problem, we have a known mass of solid and a molar mass, enough information to calculate moles of Na_2CO_3. We can then use the relationship between moles and volume to calculate molarity.

$$254\ g \times \frac{1\ mol}{106\ g} = 2.40\ mol$$

$$concentration = \frac{2.40\ mol}{2.5\ L} = 0.96\ M$$

Comment:

Note that Examples 5 and 6 start from opposite ends of the same type of calculation. In each case, we have enough information to calculate moles of Na_2CO_3. In Example 5, we had a target volume and molarity; in Example 6, we had a mass and a molar mass. Determining the amount of solute that is needed to make a desired solution or the concentration of a particular solution by knowing how it was made are two types of calculations that are performed routinely in the chemistry laboratory.

C.5 DILUTION OF SOLUTIONS

In the previous section, solution concentrations were related to the mass of the solute. That type of calculation is appropriate when the solutes come from a chemical supply house in solid form. Some compounds are supplied as concentrated solutions. HCl is a good example. Most HCl in the laboratory is purchased as 'concentrated hydrochloric acid', which is often called a "stock solution". The dilution of stock solutions to give new solutions of desired concentrations is another very common laboratory procedure. The quantitative aspects will be detailed here.

When calculating molarity and volume of diluted solutions, we can take a shortcut if we are simply diluting with pure solvent. (Be careful, this shortcut does not work for experiments where you dilute with another solution, see Example 9, or for reaction stoichiometry, see Appendix D.) The shortcut is based on the idea that in diluting a concentrated solution with pure solvent, you are not changing the number of moles of solute. The molarity changes of course, because the volume changes. Since the number of moles of solute in the concentrated stock solution (n_c) equals the number of moles of solute in the diluted solution (n_d), we can write that $n_c = n_d$.

Rearranging the relationship M = n/V, we find that n = MV, so:

$$M_cV_c = M_dV_d$$

The only restriction on the units of the volumes is that they must be the same.

Example 7

What is the concentration of the solution prepared by diluting 25 mL of 12 M HCl to 1.0 L with pure water?

Solution:

The volume of the concentrated solution (before dilution) is 25 mL, and its concentration is 12.0 M solution. The volume of the diluted solution is 1.0 L, but its concentration is unknown. Remember that the volumes must have the same units. Using our shortcut, we write:

$$(12\ M)\ (25\ mL) = M_d\ (1000\ mL)$$

$$M_d = 0.30\ M$$

Comment:

Notice that, upon rearranging the equation to solve for M_d, the mL units cancel out. Whether we use 25 and 1000 mL or 0.025 and 1.0 L for the two volumes, the results are the same. Also, we write the final result as 0.30 M. We could have written it out the long way, 0.30 moles/liter.

Example 8

An experiment requires 250. mL of 0.25 M KCl. How many mL of a 1.5 M stock solution of KCl must be used to prepare this solution?

Solution:

We are given the volume and molarity of a dilute solution and are asked for the volume of a stock (concentrated) solution of known molarity.

$$(1.5\text{ M})\ V_c = (0.25\text{ M})\ (250.\text{ mL}) \quad \text{or} \quad V_c = 42\text{ mL}$$

Comment:

We need to dilute 42 mL of the concentrated solution to 250. mL to give the desired solution.

Example 9

15 mL of a 12 M solution of HCl was diluted with 100. mL of a 0.50 M solution of HCl. What is the concentration of the resulting solution? Assume that the volumes are additive.

Solution:

Note here that we are not diluting a stock solution with pure solvent as was done in Examples 7 and 8. We will not be able to use our shortcut because both the 12.0 M and the 0.50 M solutions contribute some moles of HCl to the final solution. Instead, we add up the total number of moles of HCl and divide by the total volume.

The first solution contributes:

$$15\text{ mL} \times \frac{1\text{ L}}{10^3\text{ mL}} \times \frac{12.0\text{ mol HCl}}{1\text{ L}} = 0.18\text{ mol HCl}$$

The second solution contributes:

$$100\text{ mL} \times \frac{1\text{ L}}{10^3\text{ mL}} \times \frac{0.50\text{ mol HCl}}{1\text{ L}} = 0.050\text{ mol HCl}$$

The total number of moles of HCl in the final solution is:

$$0.18 + 0.050 = 0.23\text{ moles}$$

The total volume is: **15 mL + 100. mL = 115 mL = 0.115 L**

The concentration of the final solution is:

$$\frac{0.23\text{ moles}}{0.115\text{ L}} = 2.0\text{ M}$$

Comment:

We found the total number of moles contributed by both initial solutions by first finding the number of moles of HCl in each solution separately and adding them together. Because we are not using the shortcut, volumes had to be in liters, and the appropriate conversions were applied. This problem also points out that a small volume of a more concentrated solution (the first one) often contains more of the solute than a larger volume of a less concentrated solution.

C.6 EXERCISES

1. What is the concentration of NaCl when 25.0 g of NaCl is dissolved in water to make 450. mL of solution?
2. How many mL of a 5.0 M solution of HCl contains 0.10 moles of HCl?
3. How many moles of K_2SO_4 are contained 100. mL of a 1.35 M solution?
4. What is the concentration of K^{1+} ions in 500. mL of a 0.125 M solution of K_2SO_4?
5. How many mL of a 0.10 M solution of NaCl contains 6.2×10^{-3} moles of NaCl?
6. How many grams of $CaCl_2$ are required to make 10.0 mL of 1.00 M $CaCl_2$ solution?
7. How many moles of Cl^{1-} ions are contained in 250. mL of a 0.552 M solution of $MgCl_2$?
8. How many mL of a 0.80 M solution of Na_2CO_3 contains 0.20 moles of Na^{1+} ions?
9. How many moles of Li_2CO_3 are contained in 25.0 mL of a 1.15 M solution?
10. An experiment calls for 1.00 L of a 0.150 M KCl solution. How many mL of a 4.00 M stock solution of KCl must be used to prepare this solution?
11. How many moles of Cl^{1-} ions are contained in 18.5 mL of a 1.28 M solution of NaCl?
12. What is the concentration of K^{1+} ions in 25.0 mL of a 1.00 M solution of KCl?
13. What is the concentration of NaCl when 5.75 g of NaCl is dissolved in water to make 1.86 L of solution?
14. How many grams of LiCl are required to make 125 mL of 0.100 M LiCl solution?
15. How many mL of a 1.25 M solution of KCl contains 2.35 g of KCl?
16. How many grams of LiCl are required to make 625 mL of 2.87 M LiCl solution?
17. 10.0 mL of a 3.25 M solution of HCl is diluted with 200 mL of a 0.100 M solution of HCl. What is the concentration of the resulting solution? Assume that the volumes are additive.
18. What is the concentration of the solution prepared by diluting 25 mL of a 0.50 M solution of HCl to 125 mL with pure water?
19. What is the concentration of the solution prepared by diluting 5.0 mL of a 6.25 M solution of HCl to 65 mL with pure water?
20. How many moles of Li^{1+} ions are contained in 0.500 L of a 2.25 M solution of Li_2CO_3?
21. An experiment calls for 125 mL of a 0.625 M HCl solution. How many mL of a 12.0 M stock solution of HCl must be used to prepare this solution?
22. 12.5 mL of a 12.0 M stock solution of HCl is diluted with 85.0 mL of a 0.200 M solution of HCl. What is the concentration of the resulting solution? Assume that the volumes are additive.

ANSWERS:

1. 0.950 M	**6.** 1.11 g	**11.** 0.0237 mol	**16.** 76.1 g	**21.** 6.51 mL
2. 20 mL	**7.** 0.276 mol	**12.** 1.00 M	**17.** 0.250 M	**22.** 1.71 M
3. 0.135 mol	**8.** 125 mL	**13.** 0.0528 M	**18.** 0.10 M	
4. 0.250 M	**9.** 0.0288 mol	**14.** 0.530 g	**19.** 0.48 M	
5. 62 mL	**10.** 37.5 mL	**15.** 25.2 mL	**20.** 2.25 mol	

Appendix D
Reaction Stoichiometry

D.1 INTRODUCTION

In Appendix A, the stoichiometry of elements and compounds was presented. There, the relationships among grams, moles and number of atoms and molecules were reviewed. A similar relationship exists for chemical reactions, and we will now extend this concept of stoichiometry to reactions. In reaction stoichiometry, we are interested in the quantitative relationships between the amounts of reactants and products in a reaction. We will find, as we did in Appendix A, that the mole is the central character in these calculations.

D.2 QUANTITATIVE RELATIONSHIPS IN REACTIONS

Chemical reactions surround us. Chemists use shorthand notation to describe them in a sentence called a chemical equation. The chemical equation that describes the combustion of benzene is

$$2C_6H_6(l) + 15O_2(g) \rightarrow 12CO_2(g) + 6H_2O(g)$$

This equation implies that for every two benzene molecules that react, 15 dioxygen molecules must also react and 12 carbon dioxide molecules and six water molecules will be produced. In other words, it tells us about the “stoichiometry”, or the amounts of reactants and products involved. The coefficients, or numbers in front of each chemical formula, tell us the relative number of molecules involved in the reaction. They also tell us the relative number of *moles* involved in the reaction. Two moles of benzene will react with 15 moles of dioxygen to form twelve moles of carbon dioxide and six moles of water. It is important to note that the equation does not give us any direct information about the number of grams of each reactant or product, only the moles. If we want to know about a measurable quantity like grams, we will have to do some conversions.

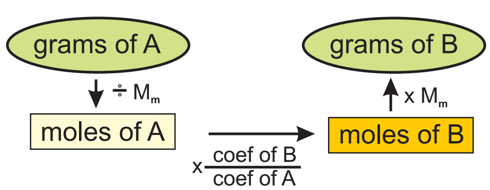
The first requirement for any stoichiometric calculation is a balanced equation. Once we have a balanced equation, the calculations will be performed by following three steps.

1. Convert the given quantitative information to moles. Experiments are always set up (thus chemistry problems are always written) such that number of moles of at least one reactant or product can be determined.
2. Use the balanced equation to convert from the moles of given substance to the moles of desired substance. Remember, we’re given quantitative information on one reactant or product, we desire quantitative information on another. This is the heart of all stoichiometry problems.
3. Convert from moles of the desired compound to the appropriate quantity.

These three steps are purposely vague. In the three previous appendices, we have discussed the conversion of moles to grams for a solid, moles to pressure, volume and temperature for a gas, and moles to volume and molarity for a solution. As you might guess, there are several variations on this three-step theme. The following examples will show some of the variety.

D.3 REACTION STOICHIOMETRY INVOLVING GRAMS

For calculations involving grams of reactants and products, our three-step scheme looks like this:



In the preceding diagram, A and B are products and/or reactants. The gram to mole conversion is achieved through application of the molar mass, the mole to mole conversion comes from the coefficients in the balanced equation.

Example 1

Consider the combustion of benzene:

$$2C_6H_6(l) + 15O_2(g) \rightarrow 12CO_2(g) + 6H_2O(g)$$

What is the maximum mass of CO_2 (M_m = 44.0 g/mol) that can be produced from the combustion of 10.0 g of C_6H_6 (M_m = 78.1 g/mol)?

Solution:

The given information is grams of benzene. The desired information is grams of CO_2. The road map above tells us we must convert grams of benzene to moles, use the mole ratio from the equation to give moles of carbon dioxide, and then convert back to grams of carbon dioxide.

$$10.0 \text{ g } C_6H_6 \times \frac{1 \text{ mol } C_6H_6}{78.1 \text{ g } C_6H_6} = 0.128 \text{ mol } C_6H_6$$

$$0.128 \text{ mol } C_6H_6 \times \frac{12 \text{ mol } CO_2}{2 \text{ mol } C_6H_6} = 0.768 \text{ mol } CO_2$$

$$0.768 \text{ mol } CO_2 \times \frac{44.0 \text{ g } CO_2}{\text{mol } CO_2} = 33.8 \text{ g } CO_2$$

Comment:

As we have seen several times in the previous appendices, each individual step is not difficult. Putting the steps together in a logical manner is the challenge. In this example, we have done each of the three steps separately. However, we could have strung together the conversion factors to save ourselves some writing. We will do that in the next example.

Example 2

How many grams of O_2 (M_m = 32.0 g/mol) are required to completely react with 10.0 g of C_6H_6?

Solution:

The given information is the mass of benzene, the desired information is the mass of dioxygen. Start with the given information and apply conversion factors following the road map.

$$10.0 \text{ g } C_6H_6 \times \frac{1 \text{ mol } C_6H_6}{78.1 \text{ g } C_6H_6} \times \frac{15 \text{ mol } O_2}{2 \text{ mol } C_6H_6} \times \frac{32.0 \text{ g } O_2}{1 \text{ mol } O_2} = 30.7 \text{ g } O_2$$

Comment:

Stoichiometry problems do not always relate reactants to products. Here is a situation where both the given and desired information deals with reactants. In this example, we have done our three step calculation by stringing together the three conversion factors appropriate to the three steps. Note that the order of operation can be determined by using the units because the units of the denominator of each conversion factor must be the same as the units of the previous numerator. Using the units to help is called the factor-label method.

D.4 LIMITING REACTANTS

We know from experience that the amount of product that is formed depends on the amount of reactant that is consumed. You can drive a car only as long as it has gasoline. The gasoline is the limiting reactant because it dictates how much product (miles) can be achieved. The amount of gasoline determines not only how far you can go, but it also determines how much CO_2 and H_2O (the reaction products of the combustion of gasoline) can be made.

In any chemical reaction, the amount of products that are made is limited by the amount of reactants. When any one reactant runs out, the reaction stops. The reactant that runs out is called the *limiting reactant* or *limiting reagent*. Any reactants that do not run out are said to be in excess. In most chemical reactions, one or more of the reactants is in excess. In the gasoline combustion reaction, there is certainly more oxygen available then there is gasoline in the gas tank, and so the oxygen is in excess.

In calculating the amount of product formed in a reaction, we always have to identify the limiting reactant. In some cases it is obvious. In Example 1 above, we read that 10.0 g of C_6H_6 reacts with excess O_2. Clearly, C_6H_6 is the limiting reactant, and O_2 is the excess reactant. But consider the following example.

Calcium hydride reacts with water to form calcium hydroxide and hydrogen gas, *via* the following reaction

$CaH_2(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + 2H_2(g)$

If 10.0 g of CaH_2 reacts with 9.00 g of H_2O, what mass of $Ca(OH)_2$ can be formed? We start by determining the limiting reactant, but we cannot tell which reactant will limit the amount of product just by comparing the grams of each reactant, we must determine how much $Ca(OH)_2$ we can make from each reactant. Based on the amount of CaH_2 we start with, we can make:

$$10.0\text{ g CaH}_2 \times \frac{1\text{ mol CaH}_2}{42.10\text{ g CaH}_2} \times \frac{1\text{ mol Ca(OH)}_2}{1\text{ mol CaH}_2} \times \frac{74.10\text{ g Ca(OH)}_2}{1\text{ mol Ca(OH)}_2} = 17.6\text{ g Ca(OH)}_2$$

Based on the amount of H_2O available, we can make:

$$9.0\text{ g H}_2\text{O} \times \frac{1\text{ mol H}_2\text{O}}{18.0\text{ g H}_2\text{O}} \times \frac{1\text{ mol Ca(OH)}_2}{2\text{ mol H}_2\text{O}} \times \frac{74.10\text{ g Ca(OH)}_2}{1\text{ mol Ca(OH)}_2} = 18.5\text{ g Ca(OH)}_2$$

Even though we have enough water to make 18.5 g of $Ca(OH)_2$, there is only enough calcium hydride to make 17.6 g. In this case, CaH_2 is the limiting reactant, H_2O is in excess, and 17.6 g of $Ca(OH)_2$ would be produced.

Let's determine how much of the excess reactant remains. CaH_2 is the limiting reactant and all amounts are calculated from it, and so we must now determine how much water reacts with the CaH_2.

$$10.0\text{ g CaH}_2 \times \frac{1\text{ mol CaH}_2}{42.10\text{ g CaH}_2} \times \frac{2\text{ mol H}_2\text{O}}{1\text{ mol CaH}_2} \times \frac{18.0\text{ g H}_2\text{O}}{1\text{ mol H}_2\text{O}} = 8.56\text{ g H}_2\text{O}$$

The above is how much water reacts, we now determine how much remains by subtracting the amount that reacts from the initial amount.

9.00 - 8.56 = 0.44 g H_2O remains

Let's summarize limiting reactants with another example.

Example 3

25.0 g of Na_2SO_4 is added to 7.00 g of carbon and allowed to react according to the following equation:

$Na_2SO_4(s) + 4C(s) \rightarrow Na_2S(s) + 4CO(g)$

a) What is the limiting reactant?

b) How many grams of Na_2S can be formed?

c) How many grams of the excess reactant will be leftover?

Solution:

Perform the three-step calculation twice, starting from the information given for each reactant.

$$25.0\text{ g Na}_2\text{SO}_4 \times \frac{1\text{ mol Na}_2\text{SO}_4}{142.05\text{ g Na}_2\text{SO}_4} \times \frac{1\text{ mol Na}_2\text{S}}{1\text{ mol Na}_2\text{SO}_4} \times \frac{78.05\text{ g Na}_2\text{S}}{1\text{ mol Na}_2\text{S}} = 13.7\text{ g Na}_2\text{S}$$

$$7.00\text{ g C} \times \frac{1\text{ mol C}}{12.01\text{ g C}} \times \frac{1\text{ mol Na}_2\text{S}}{4\text{ mol C}} \times \frac{78.05\text{ g Na}_2\text{S}}{1\text{ mol Na}_2\text{S}} = 11.4\text{ g Na}_2\text{S}$$

Comparing the two calculations leads us to conclude that C is the limiting reactant and Na_2SO_4 is in excess. 11.4 g of Na_2S can be produced. Next, find the amount of excess reactant that reacts.

$$7.00\text{ g C} \times \frac{1\text{ mol C}}{12.01\text{ g C}} \times \frac{1\text{ mol Na}_2\text{SO}_4}{4\text{ mol C}} \times \frac{142.05\text{ g Na}_2\text{SO}_4}{1\text{ mol Na}_2\text{SO}_4} = 20.8\text{ g Na}_2\text{SO}_4$$

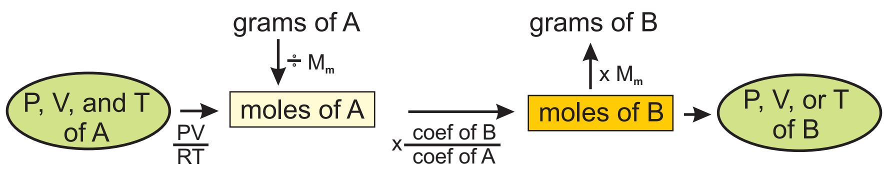
The amount remaining is given by the difference

25.0 g initially - 20.8 g consumed = 4.2 g of Na_2SO_4 remain

D.5 REACTIONS INVOLVING GASES

Examples 1 and 2 dealt with the combustion of benzene. In each, a known amount of benzene was burned, and we calculated the mass of CO_2 or O_2 produced. However, the measurable quantities of gases are pressure, volume and temperature, not mass. In this section, we use the treatment presented in Appendix B to introduce these quantities into our stoichiometric calculations.

If there is known quantitative information about the pressure, volume and temperature of a gas, we can calculate the moles of that gas, using PV = nRT. We can use this relationship in the first or last step of our calculation. We summarize this through the following road map or flowchart:



Notice that this flowchart is identical to the one that appeared earlier in this appendix, except that the use of P,V, and T information for a gas has been added as an entry into the scheme at the left and as a result from the scheme at the right. The following examples show how the ideal gas law can be used in reaction stoichiometry calculations.

Example 4

Octane combusts via the following chemical reaction:

$$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$$

How many liters of CO_2, collected at a pressure of 1.00 atm and a temperature of 25 ^oC, can be produced by the combustion of 35.0 kg of octane (the amount held by a typical car gasoline tank)?

Solution:

We first must recognize this as a reaction stoichiometry problem. We know the mass and therefore the number of moles of octane. We desire the volume of the carbon dioxide, so PV = nRT will be our third step. It is somewhat complicated to use the ideal gas law as a conversion factor, so we will string together the first two steps, and then do the third step separately, solving for volume.

$$35 \text{ kg } C_8H_{18} \times \frac{10^3 \text{ g}}{\text{kg}} \times \frac{1 \text{ mol } C_8H_{18}}{114 \text{ g } C_8H_{18}} \times \frac{16 \text{ mol } CO_2}{2 \text{ mol } C_8H_{18}} = 2.45\text{x}10^3 \text{ mol } CO_2$$

Next apply the ideal gas law, PV = nRT.

$$V = \frac{nRT}{P} = \frac{(2.45 \times 10^3 \text{ mol})(0.0821 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1} \cdot \text{mol}^{-1})(298 \text{ K})}{1.00 \text{ atm}} = 5.99\text{x}10^4 \text{ L}$$

Comment:

Burning a tank of gasoline generates enough carbon dioxide at 1.00 atm and 25 ^oC to fill a 16 ft $\times$ 16 ft room with an 8 ft ceiling.

Example 5

The following reaction is used to quickly inflate some car airbags:

$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(s)$$

How many grams of NaN_3 must be used if you wish to fill a 20.0 L airbag with dinitrogen to a pressure of 1.25 atm at 25 ^oC?

Solution:

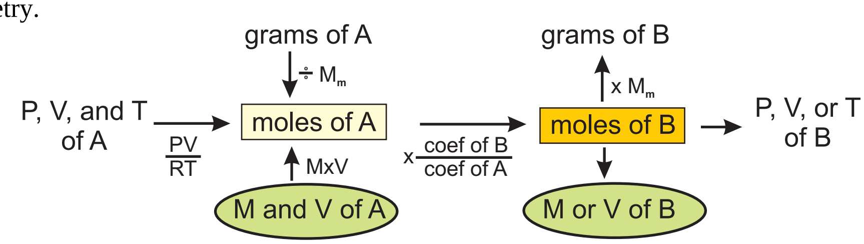
In this problem the known information involves a gaseous product (a pressure, volume and temperature are all given) and information on the reactant is desired. This time, the first of the three steps involves the ideal gas law, with the second and third steps being simple application of conversion factors.

$$n = \frac{PV}{RT} = \frac{(1.25\ \text{atm})(20.0\ \text{L})}{(0.0821\ \text{L}\cdot\text{atm}\cdot\text{K}^{-1}\cdot\text{mol}^{-1})(298\ \text{K})} = 1.02\ \text{mol}\ N_2$$

$$1.02\ \text{mol}\ N_2 \times \frac{2\ \text{mol}\ NaN_3}{3\ \text{mol}\ N_2} \times \frac{65.0\ \text{g}\ NaN_3}{1\ \text{mol}\ NaN_3} = 44.2\ \text{g}\ NaN_3$$

D.6 REACTIONS INVOLVING SOLUTIONS

In the previous sections, we have discussed the quantitative relationships between reactants and products in a reaction. In all cases, the heart of the problem was the mole ratio, and the only difference in the problems lies in how we get to the mole ratio and what we do after we have applied it. We will now consider our last conversion to and from moles. If a reactant or product is dissolved in solution, the moles of that compound are related to the molarity and the solution volume, M = n/V. (See Appendix C if you do not remember this relationship.) We can add this route to our road map of reaction stoichiometry.



The following examples show how solution data can be manipulated along with masses and data on gases to give information on reaction stoichiometry.

Example 6

The Pb^{2+} ions from water soluble $Pb(NO_3)_2$ can be precipitated by the addition of KI, forming insoluble PbI_2:

$$Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$$

If 25.0 mL of 0.375 M $Pb(NO_3)_2$ is reacted with excess KI, how many grams of PbI_2 will be produced?

Solution:

The moles of $Pb(NO_3)_2$ can be found since the volume and molarity of the solution are known. We convert the volume to liters then apply molarity to determine moles. The final two steps of the calculation are the mole to mole conversion from the balanced equation, and the conversion from moles of PbI_2 to grams, using the molar mass of 461.0 g/mol.

$$25.0\ \text{mL} \times \frac{0.375\ \text{mol}\ Pb(NO_3)_2}{1000\ \text{mL}} \times \frac{1\ \text{mol}\ PbI_2}{1\ \text{mol}\ Pb(NO_3)_2} \times \frac{461.0\ \text{g}\ PbI_2}{1\ \text{mol}\ PbI_2} = 4.32\ \text{g}\ PbI_2$$

Example 7

The silver ions in $AgNO_3$ solution can be precipitated by the addition of aqueous NaCl:

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

When 35.0 mL of a $AgNO_3$ solution of unknown concentration is reacted with excess NaCl solution, 8.53 g of AgCl is formed. What is the concentration of the $AgNO_3$ solution?

Solution:

In this problem, there is some information about the $AgNO_3$ solution and some information about the solid AgCl. After rereading the problem, it should become clear that the desired quantity is molarity of the $AgNO_3$ solution. Thus, we need to start at the other end, with the mass of AgCl. You should also recognize that whenever mass data is presented along with either a molar mass or a chemical formula (from which we can get a molar mass), we have an entry into our road map. Here, we will do the first two steps of the calculation in the usual manner, and then use the M = n/V relationship as our third step.

$$8.53\text{ g AgCl} \times \frac{1\text{ mol AgCl}}{143.35\text{ g AgCl}} \times \frac{1\text{ mol AgNO}_3}{1\text{ mol AgCl}} = 0.0595\text{ mol AgNO}_3$$

$$M = \frac{n}{V} = \frac{0.0595\text{ mol}}{0.0350\text{ L}} = 1.70\text{ mol/L} = 1.70\text{ M}$$

Comment:

The first two steps are as we have done many times now. The third step, using the molarity relationship, may at first seem a little unusual. Molarity is always the ratio of moles of a substance to the volume in liters. The first two steps tell us that there are 0.0595 moles of $AgNO_3$ contained in the 35.0 mL of solution. We simply take the ratio of these two numbers, first converting 35.0 mL to 0.0350 L, since molarity is moles per liter. Note that the experiment did not have to be done on a 1 L scale in order to calculate the molarity!

Example 8

Potassium permanganate solutions can react with acidic hydrogen peroxide solutions via the following balanced equation in water:

$$2KMnO_4 + 5H_2O_2 + 6HCl \rightarrow 2MnCl_2 + 5O_2(g) + 2KCl + 8H_2O(l)$$

When 15.0 mL of 0.0200 M $KMnO_4$ reacts with excess H_2O_2 and HCl, how many liters of O_2, collected at a total pressure of 1.00 atm and a temperature of 27 °C, will be formed?

Solution:

The volume and molarity data on $KMnO_4$ allow us to enter into the road map. The desired information is volume of O_2. We start by converting mL to L, and then proceed through the first two of the three steps in the calculation to find moles of O_2. We will do the third step, manipulation of PV = nRT, separately.

$$15.0\text{ mL} \times \frac{0.0200\text{ mol KMnO}_4}{1000\text{ m L}} \times \frac{5\text{ mol O}_2}{2\text{ mol KMnO}_4} = 7.50 \times 10^{-4}\text{ mol O}_2$$

$$V = \frac{nRT}{P} = \frac{(7.50 \times 10^{-4}\text{ mol})\ (0.0821\text{ L} \cdot \text{atm} \cdot \text{K}^{-1} \cdot \text{mol}^{-1})\ (300\text{ K})}{1.00\text{ atm}} = 0.0185\text{ L}$$

Comment:

As you can see, there are many possible routes along our road map for solving stoichiometry problems.

Example 9

How many mL of a 0.250 M NaOH solution are required to completely react with 40.0 mL of a 0.150 M H_2SO_4 solution. The overall reaction is:

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow 2H_2O(l) + Na_2SO_4(aq)$$

Solution:

We see that there is enough information to calculate moles of H_2SO_4. We desire information on NaOH. Start with the volume of 0.0400 L of H_2SO_4 solution and proceed as usual.

$$0.0400\text{ L} \times \frac{0.150\text{ mol H}_2\text{SO}_4}{1\text{ L}} \times \frac{2\text{ mol NaOH}}{1\text{ mol H}_2\text{SO}_4} \times \frac{1\text{ L}}{0.250\text{ mol NaOH}} \times \frac{10^3\text{ L}}{\text{mL}} = 48.0\text{ mL}$$

Comment:

The mL to L conversion was done at the beginning and the end of the problem since the volume information is in mL, but the concentration (molarity) is, of course, in moles per liter. Both molarities were used as conversion factors; 0.150 M converted volume of solution to moles of H_2SO_4, 0.250 M was “turned upside down” to convert moles of NaOH to liters of solution.

Example 10

Aqueous HCl and NaOH react in the following manner:

$$HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$$

When 30.0 mL of 0.100 M HCl are mixed with 20.0 mL of 0.125 M NaOH, what is concentration of the excess reagent?

Solution:

This is a limiting reactant problem. At first glance, it would appear quite different than the limiting reactant problems we saw earlier. However, upon careful reading of the experiment, we see that we have quantitative information on both reactants, enough to calculate moles of both. The desired quantity the concentration of the excess reactant. In essence, we are reacting an acid and a base, and need to determine which reactant is limiting, and find how much of the excess reactant is leftover, as was done in Example 3 above. In order to determine the limiting reactant, we calculate how much product can be made from each reactant. It doesn't matter which product we choose. Let's pick water.

$$\textbf{30 mL} \times \frac{\textbf{1 L}}{\textbf{10}^3\textbf{ mL}} \times \frac{\textbf{0.100 mol HCl}}{\textbf{1 L}} \times \frac{\textbf{1 mol H}_2\textbf{O}}{\textbf{1 mol HCl}} = \textbf{0.0030 mol H}_2\textbf{O}$$

$$\textbf{20 mL} \times \frac{\textbf{1 L}}{\textbf{10}^3\textbf{ mL}} \times \frac{\textbf{0.125 mol NaOH}}{\textbf{1 L}} \times \frac{\textbf{1 mol H}_2\textbf{O}}{\textbf{1 mol NaOH}} = \textbf{0.0025 mol H}_2\textbf{O}$$

Fewer moles of water can be made from the NaOH, so NaOH is the limiting reactant, HCl is the excess reactant. Notice that it was not necessary to go all the way through and calculate the grams of water. Clearly, if we multiply each result by 18.02 g/mol (the molar mass of H_2O), the conclusion is the same, NaOH is limiting. In order to calculate molarity of HCl, we need the number of <u>moles</u> of HCl which were leftover, and the <u>*total*</u> solution volume.

<u>Moles of HCl at the start</u>:

$$\textbf{30 mL} \times \frac{\textbf{1 L}}{\textbf{10}^3\textbf{ mL}} \times \frac{\textbf{0.100 mol HCl}}{\textbf{1 L}} = \textbf{0.0030 mol HCl}$$

The number of moles of HCl consumed is based on the amount of limiting reactant consumed:

$$\textbf{20 mL} \times \frac{\textbf{1 L}}{\textbf{10}^3\textbf{ mL}} \times \frac{\textbf{0.125 mol NaOH}}{\textbf{1 L}} \times \frac{\textbf{1 mol HCl}}{\textbf{1 mol NaOH}} = \textbf{0.0025 mol HCl react}$$

<u>Moles of HCl remaining</u>:

0.0030 - 0.0025 = 0.0005 moles of HCl remain

<u>Concentration of HCl at the end</u>

$$\frac{\textbf{moles of HCl}}{\textbf{total volume}} = \frac{\textbf{0.0005 mol HCl}}{\textbf{0.050 L}} = \textbf{0.010 M}$$

Comment:

This problem is actually very similar to the limiting reactant problems we did before. The difference is that, instead of finding the grams of the leftover reactant, we had to find the concentration, which involved a calculation of the number of moles of the leftover reactant.

D.7 EXERCISES

Use the following molar masses to do the following problems:

C_4H_8 : 56.10 g/mol	C_4H_9OH : 74.12 g/mol
Fe_2O_3 : 159.70 g/mol	Al_2O_3 : 101.96 g/mol
V_2O_5 : 181.88 g/mol	NH_4VO_3 : 116.98 g/mol
NH_3 : 17.03 g/mol	V_2O_3 : 149.88 g/mol
Cu_2S : 159.17 g/mol	CuO : 79.55 g/mol
Cu_2O : 95.55 g/mol	AgCl : 143.4 g/mol

1. In the presence of acids, water can react with alkenes to form alcohols:

$$C_4H_8 + H_2O \rightarrow C_4H_9OH$$

If 250 g of C_4H_8 reacts with excess H_2O, how many grams of C_4H_9OH can be produced?

2. Aluminum reacts with iron(III) oxide in the "thermite reaction":

$$2Al(s) + Fe_2O_3(s) \rightarrow 2Fe(s) + Al_2O_3(s)$$

a) If 10.0 g of Al reacts with excess Fe_2O_3, how many grams of Al_2O_3 can be produced?

b) If 25.0 g of Al reacts with 10.0 g of Fe_2O_3, how many grams of Al_2O_3 can be produced?

c) In the experiment in part b, what is the mass of the excess reactant remaining after complete reaction?

3. Vanadium(V) oxide reacts with ammonia and water as follows:

$$V_2O_5 + 2NH_3 + H_2O \rightarrow 2NH_4VO_3$$

a) If 50.0 g of V_2O_5 is reacted with excess ammonia and water, how many grams of NH_4VO_3 can be produced?

b) How many grams of NH_3 are required to completely react with 50.0 g of V_2O_5?

4. Vanadium(III) oxide can be made by reduction of vanadium(V) oxide with hydrogen:

$$V_2O_5(s) + 2H_2(g) \rightarrow V_2O_3(s) + 2H_2O(l)$$

a) How many liters of H_2, measured at 1.00 atm and 30 ^oC, are required to completely react with 75.0 g of V_2O_5?

b) If 10.0 g of V_2O_5 reacts with 1.65 L of H_2, measured at 1.00 atm and 30 ^oC, how many grams of V_2O_3 can be produced?

5. Copper(I) sulfide is prepared by heating copper and sulfur in the absence of air:

$$2Cu(s) + S(s) \rightarrow Cu_2S(s)$$

a) How many grams of Cu_2S can be produced from the reaction of 25.0 g of Cu with excess S?

b) How many grams of sulfur are required to form 75.0 g of Cu_2S?

c) If a mixture of 135 g of Cu and 45 g of S is allowed to react, how many grams of Cu_2S could be produced?

d) How many grams of the excess reactant remain in the experiment in part c?

6. Copper(I) oxide can be prepared by thermal decomposition of copper(II) oxide:

$$4CuO(s) \rightarrow 2Cu_2O(s) + O_2(g)$$

a) How many grams of Cu_2O can be produced upon the decomposition of 450 g of CuO?

b) How many liters of O_2, collected at 1.00 atm and 27 ^oC, can be produced by the decomposition of 450 g of CuO?

7. The silver ions in aqueous silver sulfate can be precipitated by addition of excess chloride:

$$Ag_2SO_4(aq) + 2NaCl(aq) \rightarrow 2AgCl(s) + Na_2SO_4(aq)$$

a) How many grams of silver chloride can be formed when 35.0 mL of a 0.100 M Ag_2SO_4 solution is reacted with excess sodium chloride solution?

b) If 22.7 mL of a silver sulfate solution of unknown concentration yields 0.985 g of AgCl upon reaction with excess sodium chloride solution, what is the concentration of the silver sulfate solution?

8. Zn metal reacts with hydrochloric acid to produce hydrogen gas and zinc(II) chloride:

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

a) If 15.0 g of Zn are added to excess HCl(aq), how many liters of $H_2(g)$, collected at 27 ^oC and 725 mm Hg, are produced?

b) If excess Zn is added to 25.0 mL of 0.025 M HCl(aq), how many liters $H_2(g)$, collected at 27 ^oC and 725 mm Hg, can be produced?

9. Potassium permanganate and iron(II) chloride undergo an electron transfer reaction in acid solution:

$$KMnO_4(aq) + 5FeCl_2(aq) + 8HCl \rightarrow MnCl_2(aq) + 5FeCl_3(aq) + KCl(aq) + 4\ H_2O(l)$$

How many mL of 0.150 M $FeCl_2(aq)$ are needed to completely react with 13.7 mL of 0.110 M $KMnO_4$?

10. Citric acid reacts with sodium hydroxide in a proton transfer reaction:

$H_3C_6H_5O_7(aq) + 3NaOH(aq) \rightarrow 3H_2O(l) + Na_3C_6H_5O_7(aq)$

a) How many mL of 0.125 M NaOH(aq) are required to completely react with 25.0 mL of 0.0695 M citric acid?

b) If 37.5 mL of 1.25 M NaOH(aq) is needed to completely react with 22.5 mL of a citric acid solution, what is the concentration of the citric acid solution?

ANSWERS:

1.	330 g			
2.	**a)** 18.9 g	**b)** 6.38 g	**c)** 21.6 g	
3.	**a)** 64.3 g	**b)** 9.34 g		
4.	**a)** 20.5 L	**b)** 4.97 g		
5.	**a)** 31.3 g	**b)** 15.1 g	**c)** 169 g	**d)** 11 g
6.	**a)** 405 g	**b)** 34.8 L		
7.	**a)** 1.00 g	**b)** 0.151 M		
8.	**a)** 5.92 L	**b)** 8.07 mL		
9.	50.2 mL			
10.	**a)** 41.7 mL	**b)** 0.694 M		