CHAPTER 2: STOICHIOMETRY

INTRODUCTION

The first chapter looked at compounds in isolation: a compound's chemical formula and relative atomic and molecular masses. This chapter looks at compounds in reaction with each other or as the products of reactions. We learn how to represent a reaction as a balanced chemical equation. We learn how a chemical equation tells us, given the amount of some reactant, how much of some other reactant must be added to carry the reaction to completion. We learn how to predict what amount of product is formed by a reaction, even when reactants are not present in exactly balanced amounts. We learn to deal quantitatively with reactions involving solutions and gases.

- Note on studying Chapter 2: This chapter emphasizes calculations. Much of the text is included in the Examples. Be sure you work through these with a pencil and paper; don’t try to just read through them. Pay attention to the notation used, and be sure you understand how to use the unit-factors that are key to solving the problems.

REVIEW

This chapter requires a firm grasp of the ideas behind the law of definite proportions (also called the law of constant composition) and Dalton's atomic theory, covered in Section 1-3. You should also be sure you understand the law of multiple proportions covered in Section 1-4. These ideas are fundamental to understanding empirical and molecular formulas and balanced chemical equations, introduced in Sections 2-1 to 2-3 and used throughout the rest of the book.

In preparation for reading Section 2-2 on you should review the law of combining volumes and Avogadro's hypothesis (Section 1-4) and the measurement of relative atomic and formula masses and of chemical amounts (Sections 1-6 and 1-7).

2-1 Writing Balanced Chemical Equations

A balanced chemical equation is our main tool for understanding a chemical reaction and making quantitative statements about the substances consumed or produced by a reaction. The main idea is that each substance's coefficient in the balanced chemical reaction tells us the relative number of moles of that substance that react or are produced by the reaction. Notice that the coefficients give us the relative numbers of molecules, if molecular substances are involved.
Balancing Equations by Inspection

A chemical equation must be in balanced form to satisfy the law of conservation of mass, as expressed in Dalton's atomic theory (section 1-3). Since atoms are indestructible and retain their identity in reactions, the atoms of any element must be in equal numbers on both sides of the equation. Since each substance's chemical formula has fixed subscripts, we balance the elements on both sides by choosing the right coefficients. The symbols (s), (l), (g), (aq) and so on inform us what phase each reactant or product is in, but are not necessary in the balancing process.

Most of the equations you will deal with up to Chapter 12 can be balanced by inspection (trial and error). Some rules to make this easier are given on p. 54. Follow the Example 2-1 and its Exercise. Then practice on a number of the Problems at the end of the chapter. If groups of atoms appear in parentheses in a formula, be sure to count all the atoms within parentheses the number of times required by the subscript on the parentheses. The formula Cu(NO₃)₂ contains one Cu, two N, and 6 O atoms.

Balancing Equations Algebraically

Balancing equations algebraically is a sophisticated method that’s useful for very complicated equations. It requires you to set up a system of linear equations to keep track of all the kinds of atoms and find the coefficients. It will give you a chance to use your scientific calculator, if you have one, or to write an algorithm for your computer. Follow through the examples using the rules on p. 56 to learn how it works. Note, however, this method is not used anywhere else in the text.

Chemical Equations That Cannot Be Uniquely Balanced

Note this section on pp. 58-59. If it’s possible to balance an equation using several different sets of coefficients, and the coefficients don’t differ by a simple factor such as 2 or 3, then you have two or more independent reactions added in together. Such “equations” cannot be used for meaningful chemical calculations.

2-2 Using Balanced Chemical Equations

We now learn to use a balanced chemical equation to do useful calculations. The equation tells us the relative number of moles of each substance. We can convert that number into relative amounts in grams or liters of volume, etc.

If the substances involved are composed of molecules, and are represented in the equation by molecular formulas, then the coefficients directly tell us the relative number of molecules involved in the reaction. If the substances involved are gases, and if the temperature and pressure do not change, then according to the law of combining volumes (Section 1-4) the coefficients also tell us the relative volumes of gases involved in the reaction. Finally, if we
multiply each substance's coefficient by that substance's molar mass, we then have the relative mass in grams of substances involved in the reaction:

- coefficient of X = relative number of moles of X involved in the reaction,
- molar mass of X = (grams of X) / (mol of X),
- \[(\text{coefficient of X}) \times \text{(molar mass of X)} = \text{relative number of grams of X}\].

- Notice the notation used for mass and chemical amount of a substance:
  - mass in grams of water is given by \(m_{H_2O}\)
  - chemical amount in moles is given by \(n_{H_2O}\)

Throughout your study of chemistry you will commonly encounter problems in which you are given a chemical equation, an amount (in moles, grams, or liters) of some substance consumed or produced in that reaction, and you must calculate the amount (in moles, grams, or liters) of some other substance consumed or produced in the reaction. These are called stoichiometry problems. The balanced chemical equation is the central switching station that connects your starting information with your answer. You need to become adept at using the unit-factor method to convert between different measures of a given substance and between chemically equivalent amounts of different substances (see Appendix B for additional help). Note that unit-factor calculations can be done a step at a time or in one grand set-up with many unit factors multiplied together. Use whichever method works best for you.

Figure 2-3 outlines the general procedure to use in all mass/mass stoichiometric calculations. Again, be sure you work through the examples and enough Problems at the end of the chapter to be confident you can handle them.

### 2-3 Limiting Reactant and Percentage Yield

Chemical reaction equations are usually introduced as if the reactants were present in the exact ratios that guarantee they are all completely consumed. But in practical cases the reaction usually stops when one of the reactants, the limiting reactant, is completely consumed. Other reactants, those in excess, are not completely used up. Given a chemical equation and the available amounts (in moles, grams, or liters) of reactants, you need to be able to identify the limiting reactant and the amount of product formed.

A chemical equation shows the maximum possible amount of product that can be formed from the amount of limiting reactant present. This amount is called the theoretical yield. But often reactions stop short of completion, or reactants are partly drawn off into competing reactions, for reasons discussed in depth in Chapter 7 (Chemical equilibrium). As a result, less than the theoretical yield is obtained. We use the actual yield and the percentage yield to describe the amount of product formed in these less than ideal conditions. Follow through the
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Examples and work the Exercises and some chapter-end Problems to master this kind of calculation.

Following the Progress of Chemical Reactions

The table of changes is introduced in the section on pp. 70-71. Such a table helps to identify a limiting reactant and to find the theoretical yield of a reaction. It comes in handy here, and it will again in Chapters 7, 8, and 9, where we deal with equilibrium reactions that aren’t expected to go to completion. As usual, use pencil and paper to follow through the examples.

2-4 The Stoichiometry of Reactions in Solution

So far we have worked with reactions between pure substances, solids, liquids, and gases. In this section the reactants are solutes, substances dissolved in water to form aqueous solutions. You need to learn how to deal with solution calculations, because many practical chemical procedures involve solutions.

In the laboratory and in automated testing equipment we often measure out accurate volumes of solution in order to obtain a desired chemical amount of solute, \( n_{\text{solute}} \). The most common unit of concentration used by chemists is moles of solute per liter of solution, which has the name molarity, and the symbol \( M \). Often we write square brackets around the formula of the solute as a shorthand way of indicating the molarity. We denote the concentration of solute as \( c_{\text{solute}} \). Then we can combine the equations in the text to summarize them:

\[
c_{\text{solute}} = \frac{n_{\text{solute}}}{V} = \frac{\text{moles solute}}{\text{liters solution}} = [\text{solute}] = \text{mol L}^{-1} = \text{molarity M}
\]

Rearrangement gives the amount of solute in a volume \( V \) (liters) as \( n_{\text{solute}} = c_{\text{solute}} \times V \).

Example 2-9 illustrates the basic calculation of concentration.

Other measures of concentration will be explored in Chapter 7, but you will be using molarity throughout the course. The big majority of solutions in this course are aqueous. Solutions in solvents other than water are met occasionally, and the calculations are essentially the same.

The Preparation of Solutions of Accurately Known Molarity

Be sure you understand why you cannot prepare a 0.50 molar solution of solute A by simply adding 0.50 moles of solute A to 1.00 liter of solvent. Read the text to get an idea of the care that must be taken when you dissolve a known amount of solute in water in a volumetric flask and then add the final few drops “up to the mark”.

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Dilutions

The key to dilution calculations is that the amount of solute remains constant when solvent is added to a solution of initial concentration \( c_i \). Then letting subscript \( i \) indicate the initial and subscript \( f \) indicate the final solution, we can write \( n_{\text{solute}} = c_i \times V_i = c_f \times V_f \). Rearranging this equation permits calculation of the required final volume or concentration from the data given in a problem. Follow through Example 2-10 and its Exercise for practice.

Using the Molarity in Calculations

In a stoichiometry problem you are always given a measurable quantity \( X_A \) of some reactant or product \( A \), and are asked to calculate the measurable quantity \( X_B \) of some other reactant or product \( B \). \( X_A \) and \( X_B \) were masses earlier in Chapter 2. Here they are volumes of solution. Solving the problem always requires converting the given measurable quantity \( X_A \) into the chemical amount \( n_A \) in moles. Then we use the coefficients in the balanced chemical equation to find the chemical amount \( n_B \) in moles. Finally we convert \( n_B \) into the measurable quantity \( X_B \). Just as molar mass is a conversion factor between the chemical amount in moles and the measurable quantity of mass in grams, molarity is a conversion factor between the chemical amount in moles and the measurable amount of solution volume in liters.

In many stoichiometry problems involving solutions, molarity itself is the unknown to be solved for. Both the example in the text of the \( \text{NaBr(aq)} + \text{Cl}_2(aq) \) reaction and Example 2-11 require this kind of calculation. Work through these with pencil and paper, then practice on a few problems from the chapter’s end.

2-5 The Scale of Chemical Processes

College laboratory experiments are usually carried out at the scale of grams, which are easy to handle. But you should be aware that reactions can take place at the tiny scale of nanograms (billionths of a gram), common in biochemical processes, up to the vast scale of millions of kilograms, common in industrial processes. Changing the scale of a reaction can introduce new problems, especially the dissipation of heat in large industrial processes. You should make sure you are familiar with the various prefixes denoting powers of ten and be comfortable with interconverting among them.

The Chemistry in Your Life article on Green Chemistry emphasizes the way the modern chemical industry attempts to utilize products of side-reactions and what used to be considered waste materials, both to economize and, perhaps more importantly, to preserve a safe and clean environment.
LEARNING GOALS . . . CHAPTER 2

KEY TERMS - Define or explain the following terms in your own words.

- stoichiometry
- word equation
- chemical equation, balanced chemical equation
- reactants, products, coefficients
- material balance
- system of linear equations
- limiting reactant, excess reactant
- theoretical yield, actual yield
- percentage yield
- completion of a reaction
- table of changes
- solute, solvent
- concentration
- molarity
- volumetric flask
- green chemistry
- microscale, industrial scale processes

STUDY QUESTIONS AND PROBLEM TYPES

Section 2-1

1. Which of Dalton’s principles of atomic theory (Section 1-3) is expressed by a balanced chemical equation? In balancing a chemical equation, do we modify the coefficients, the formula subscripts, or both? Explain.

2. Practice balancing many chemical equations. Choose an equation from the text. Write it out omitting all coefficients (in effect, setting all coefficients to one). Apply the procedure outlined in the text to balance the equation. Practice balancing many equations until you are quick and confident.

3. Pick a simple equation and then a complicated one to balance algebraically. Ask your instructor if you will be required to use this method on an exam or homework assignment.
Section 2-2

4. Review the various quantitative relationships in a chemical reaction described in Section 2-4. How are these relationships derived from a balanced chemical equation? What is meant by the term stoichiometry?

5. Given a chemical equation and a chemical amount (in moles) of a reactant, compute how many moles of another reactant are required, and how many moles of product are produced.

6. In a gas phase reaction (all reactants and products are gases), what is the relationship between coefficients in the chemical equation and the gas volumes? What additional conditions must be held constant? Given the volume of any one reactant or any one product, calculate the volumes of all other chemical species in the reaction.

7. Given a chemical equation and the mass of any one reactant or product, calculate the mass of all other chemical species involved.

Section 2-3

8. What is a limiting reactant? What can you say about the amounts of other reactants? What happens if the supply of a limiting reactant is increased slightly?

9. Given a chemical equation and the masses of reactants, decide which is the limiting reactant. Calculate the theoretical yield of products.

10. What can cause a reaction’s actual yield to be less than its theoretical yield? Calculate the theoretical and percentage yields in several of the end-of-chapter problems 29-44. Given the percentage yield of the consecutive reaction steps in a synthetic process, calculate the overall yield of the process.

11. You have an industrial process in which the cost per mole of the reactants differ dramatically. Which would you choose to be the limiting reactant and why? What if you only know the cost per gram?

Section 2-4

12. What is meant by the concentration of a solution? Define molarity. Given the molarity, calculate the number of moles of solute in a given volume of solution. Calculate the volume required to deliver a required chemical amount of solute, given the concentration of the solution. Recalling the units of molarity (mol L\(^{-1}\)) should keep its definition at your fingertips:

\[ c_{\text{solute}} = \frac{n_{\text{solute}}}{V} \]
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13. Explain how to make a given volume of a solution of a given molarity. Why can you not make a 0.500 M NaCl\((aq)\) solution by simply adding 0.500 moles of NaCl to 1.000 liters of water?

14. Be able to solve stoichiometry problems involving solution concentrations. Typical problems might give a chemical reaction and ask for the volume of one solution required to react with the volume of another, or ask for the number of moles of product produced, given the volumes of two reactants.

Section 2-5

14. State one problem you might encounter when a reaction is carried out at a large industrial scale, that you do not usually encounter at laboratory scale.

CHAPTER 2: PRACTICE EXAM

Section 2-1

1. Balance the following reactions, for which all reactants and products are given.
   (a) Decomposition of potassium chlorate to give oxygen gas:
       \[ \text{KClO}_3(s) \rightarrow \text{KCl}(s) + \text{O}_2(g) \]
   (b) Action of TUMS in the tummy:
       \[ \text{HCl}(aq) + \text{CaCO}_3(s) \rightarrow \text{CaCl}_2(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(aq) \]
   (c) Formation of phosphoric acid from diphosphoruspentaoxide
       \[ \text{P}_2\text{O}_5(s) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{PO}_4(aq) \]
   (d) Combustion of butane gas
       \[ \text{C}_4\text{H}_10(g) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g) \]

Section 2-2

2. How many moles of sulfuric acid \((\text{H}_2\text{SO}_4)\) would be needed to produce 4.80 moles of iodine, according to the following balanced equation?
   \[ 10 \text{HI (aq)} + 2 \text{KMnO}_4(aq) + 3 \text{H}_2\text{SO}_4 \rightarrow 5 \text{I}_2(s) + 2 \text{MnSO}_4(aq) + \text{K}_2\text{SO}_4(aq) + 8 \text{H}_2\text{O (l)} \]
   (a) 3.00 moles \(\text{H}_2\text{SO}_4\)  (c) 2.88 moles \(\text{H}_2\text{SO}_4\)
   (b) 1.60 moles \(\text{H}_2\text{SO}_4\)  (d) 8.00 moles \(\text{H}_2\text{SO}_4\)

3. All but one of the following statements about the burning of sulfur to form sulfur dioxide ("fire and brimstone") are directly indicated by the balanced equation:
   \[ \text{S(s)} + \text{O}_2(g) \rightarrow \text{SO}_2(g) \]
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Pick the incorrect statement.
(a) One mole of sulfur combines with one mole of oxygen molecules.
(b) One atom of sulfur reacts with two atoms of oxygen.
(c) One gram of sulfur reacts with one gram of oxygen.
(d) One mole of molecular oxygen reacts to form one mole of SO₂.
(e) One liter of O₂(g) will produce one liter of SO₂(g) (at the same T, P).

4. Gaseous ammonia (NH₃) burns with oxygen gas (O₂) to form gaseous NO plus water vapor. The volume of oxygen required per volume of ammonia (at the same T and P) is:
(a) 5  (b) 4  (c) 1.25  (d) 1  (e) 0.8

5. How many grams of sodium chloride can be made from 2.30 g of sodium metal in the reaction with chlorine? Molar masses: Na 23.0 g mol⁻¹, Cl 35.5 g mol⁻¹

\[ 2 \text{Na} (s) + \text{Cl}_2 (g) \rightarrow 2 \text{NaCl} (s) \]

(a) 2.93 g  (c) 11.70 g  (b) 5.85 g  (d) none of the above

Section 2-3

6. Silicon was reacted with nitrogen to produce the “space age” ceramic silicon nitride, Si₃N₄. Atomic masses are Si 28.09 and N 14.01 g/mol. The reaction is

\[ 3 \text{Si} (s) + 2 \text{N}_2 (g) \rightarrow \text{Si}_3\text{N}_4 (s) \]

If 79.0 g Si were used, and 125 g Si₃N₄ were produced, the percentage yield was closest to:
(a) 98 %  (d) 85 %
(b) 95 %  (e) 80 %
(c) 90 %

7. For the reaction (not balanced) producing sulfur trioxide,

\[
\underline{\text{____S}} (s) + \underline{\text{____O}_2} (g) \rightarrow \underline{\text{____SO}_3} (g)
\]

if one has 32 grams of S and 32 grams of O₂, which reactant is in excess?
(a) S
(b) O₂
(c) neither; these are equivalent amounts (within 2 significant figures)
(d) can't answer without more data

8. How many grams of ammonia (NH₃) could be formed if 10.0 g H₂ and 40.0 g of N₂ were allowed to react by the following reaction?

\[ \text{N}_2 (g) + 3 \text{H}_2 (g) \rightarrow 2 \text{NH}_3 (g) \]

9. Iron is reduced from its oxide by coke in the blast furnace, by the reaction
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\[ \text{Fe}_2\text{O}_3(s) + 3 \text{C}(s) \rightarrow 2 \text{Fe}(l) + 3 \text{CO}(g) \]

Molar masses are: \( \text{Fe}_2\text{O}_3 \) 159.7 g/mol \( \text{C} \) 12.01 g/mol \( \text{Fe} \) 55.85 g/mol

If \( 2.86 \times 10^4 \) kg of \( \text{Fe}_2\text{O}_3 \) is reacted with \( 9.82 \times 10^3 \) kg of \( \text{C} \), calculate the theoretical yield of \( \text{Fe} \) in kg.

Section 2-4

10. What is the molarity of a solution that contains 0.2922 grams of \( \text{NaCl} \) (molar mass = 58.44 g/mole) in 500.0 mL of solution?
   (a) 0.500 M  (d) 0.0500 M
   (b) 0.200 M  (e) None of the above is correct.
   (c) 0.100 M

11. A solution is prepared by dissolving 5.000 g of \( \text{KCl}(s) \) in enough water to make exactly 750 mL of solution. What is the molarity of the solution?
   (a) 0.0894 M
   (b) 0.894 M
   (c) 0.0671 M
   (d) 0.671 M
   (e) None of the above is correct to two significant figures.

12. A 15.0% by weight solution of ethanol (molar mass = 46.07 g/mol) in water (molar mass = 18.01 g/mol) has a density of 0.977 g/mL. What is the molar concentration of ethanol?
   (a) 7.07 M  (d) 3.18 M
   (b) 3.83 M  (e) 2.27 M
   (c) 3.26 M

13. How many grams of lithium hydroxide are required to prepare 600 ml of 1.5 molar solution?
   (a) 39.8  (d) 17.3
   (b) 27.1  (e) 14.4
   (c) 21.5

Section 2-5

14. 3,420 kg of sucrose (molar mass 342 g mol\(^{-1}\)) is dissolved in enough water to give 200 m\(^3\) of solution. What is the molarity of this solution?

Chapter 2: Practice Exam Answer Key

1. (a) \( 2 \text{KClO}_3(s) \rightarrow 2 \text{KCl}(s) + 3 \text{O}_2(g) \)
   (b) \( 2 \text{HCl}(aq) + \text{CaCO}_3(s) \rightarrow \text{CaCl}_2(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(aq) \)
(c) \( \text{P}_2\text{O}_5(s) + 3 \text{H}_2\text{O}(l) \rightarrow 2 \text{H}_3\text{PO}_4(aq) \)

(d) \( 2 \text{C}_4\text{H}_{10}(g) + 13 \text{O}_2(g) \rightarrow 8 \text{CO}_2(g) + 10 \text{H}_2\text{O}(g) \)

2. (c) \( 4.80 \text{ mol I}_2 \left( \frac{3 \text{ mol H}_2\text{SO}_4}{5 \text{ mol I}_2} \right) = 2.88 \text{ mol H}_2\text{SO}_4 \)

3. (c)

4. (c) \( 4 \text{NH}_3 + 5 \text{O}_2 \rightarrow 4 \text{NO} + 6 \text{H}_2\text{O} \) Ratio of volumes is same as ratio of moles, \( \frac{\text{O}}{\text{NH}_3} = \frac{5}{4} = 1.25 \).

5. (b)

Molar mass of NaCl = \( (23.0 \text{ g Na / mol Na}) + (35.5 \text{ g Cl / mol Cl}) = 58.5 \text{ g / mol} \)

\[ 2.30 \text{ g Na (1 mol Na / 23.0 g Na) (2 mol NaCl / 2 mol Na) (58.5 g NaCl / mol NaCl)} = 5.85 \text{ g NaCl} \]

6. (b) Th. yield = \( 79.0 \times \frac{1 \text{ mol Si}}{28.09 \text{ g Si}} \times \frac{1 \text{ mol Si}_3\text{N}_4}{3 \text{ mol Si}} \times \frac{140.3 \text{ g Si}_3\text{N}_4}{1 \text{ mol Si}_3\text{N}_4} = 131.5 \text{ g Si}_3\text{N}_4 \)

Pct. yield = \( \frac{\text{actual yield}}{\text{theor. yield}} \times 100\% = \frac{125 g}{131.5 g} \times 100\% = 95.1\% \)

7. (a)

Balanced equation. \( 2 \text{S} + 3 \text{O}_2 \rightarrow 2 \text{SO}_3 \) The given masses are \( 1.0 \text{ mole S, 1.0 mole O}_2 \).

The given mass ratio is \( 1.0 \text{ mol S} / 1.0 \text{ mol O}_2 = 3/3 \), whereas the equation requires a \( 2/3 \) ratio. S is in excess and \( \text{O}_2 \) is the limiting reactant.

8.

\[ 10.0 \text{ g H}_2 \left( \frac{1 \text{ mol H}_2}{2.02 \text{ g H}_2} \right) = 4.95 \text{ mol H}_2 \text{ provided.} \]

\[ 40.0 \text{ g N}_2 \left( \frac{1 \text{ mol N}_2}{28.0 \text{ g N}_2} \right) = 1.43 \text{ mol N}_2 \text{ provided} \]

1.43 mol N\(_2\) reacts with \( (1.43 \text{ mol N}_2)( \frac{3 \text{ mol H}_2}{1 \text{ mol N}_2} ) = 4.29 \text{ mol H}_2 \), so H\(_2\) is in excess and N\(_2\) is the limiting reactant.

9.

\[ (2.86 \times 10^4 \text{ kg Fe}_2\text{O}_3) \left( \frac{10^3 \text{ g}}{\text{kg}} \right) / (159.7 \text{ g / mol}) = 1.79 \times 10^5 \text{ mol Fe}_2\text{O}_3 \text{ provided.} \]

\[ (9.82 \times 10^3 \text{ kg C}) \left( \frac{10^3 \text{ g}}{\text{kg}} \right) / (12.01 \text{ g / mol}) = 8.18 \times 10^5 \text{ mol C} \text{ provided.} \]

C is in excess and Fe\(_2\)O\(_3\) is the limiting reactant. For each mole of Fe\(_2\)O\(_3\) consumed, 2 moles of Fe are produced as theoretical yield:

\[ (2 \times 1.79 \times 10^5 \text{ mol Fe}) \times (55.85 \text{ g/mol}) = 2.00 \times 10^7 \text{ g Fe} = 2.00 \times 10^4 \text{ kg Fe} \text{ produced as theoretical yield.} \]
10. (e) \( \frac{0.2922 \text{ g}}{58.44 \text{ g/mol}} = 5.000 \times 10^{-3} \text{ mol NaCl} \).  
\( (5.000 \times 10^{-3} \text{ mol NaCl}) / .5000 \text{ L} = 1.000 \times 10^{-2} \text{ M NaCl} \).

11. (a) molarity = (moles of solute)/(volume of solution)  
Molar mass KCl = 39.10 + 35.45 = 74.55 g/mol.  
\( \frac{5.000 \text{ g KCl}}{74.55 \text{ g KCl/mol KCl}} = 0.06707 \text{ mol KCl} \).  
Molarity = \( \frac{0.06707 \text{ mol KCl}}{0.750 \text{ L solution}} = 0.0894 \text{ M} \).

12. (d) First find the mass, then the amount in moles of ethanol in 1 L (977 g) of solution.  
In 1.000 L there are 977 \times 0.15 = 14.6 g ethanol  
mol ethanol per liter = \( \frac{146.6 \text{ g}}{46.07 \text{ g/mol}} = 3.18 \text{ mol/L} \).

13. (c) First find the amount of LiOH in moles, then the molarity.  
Molar mass of LiOH = 6.94 + 16.00 + 1.01 = 23.95 g / mol.  
\( (1.5 \text{ mol LiOH/L})(0.600 \text{ L}) = 0.900 \text{ mol LiOH required} \).  
\( (0.900 \text{ mol}) \times (23.95 \text{ g/mol}) = 21.5 \text{ g LiOH required} \).

14. First convert 3,420 kg to moles:  
\( 3.42 \times 10^3 \text{ kg} \times \frac{10^3 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol}}{3.43 \times 10^3 \text{ g}} = 1.00 \times 10^4 \text{ mol} \).  

Then convert 200 m³ to L:  
\( 200 \text{ m}^3 \times \frac{100^3 \text{ cm}^3}{\text{m}^3} \times \frac{1 \text{ L}}{10^3 \text{ cm}^3} = 2.00 \times 10^5 \text{ L} \).  

Finally,  
\( c_{\text{sucrose}} = \frac{\text{no. of mol}}{\text{no. of L}} = \frac{1.00 \times 10^4 \text{ mol}}{2.00 \times 10^5 \text{ L}} = 5.00 \times 10^{-2} \frac{\text{mol}}{\text{L}} = 0.0500 \text{ M} \).