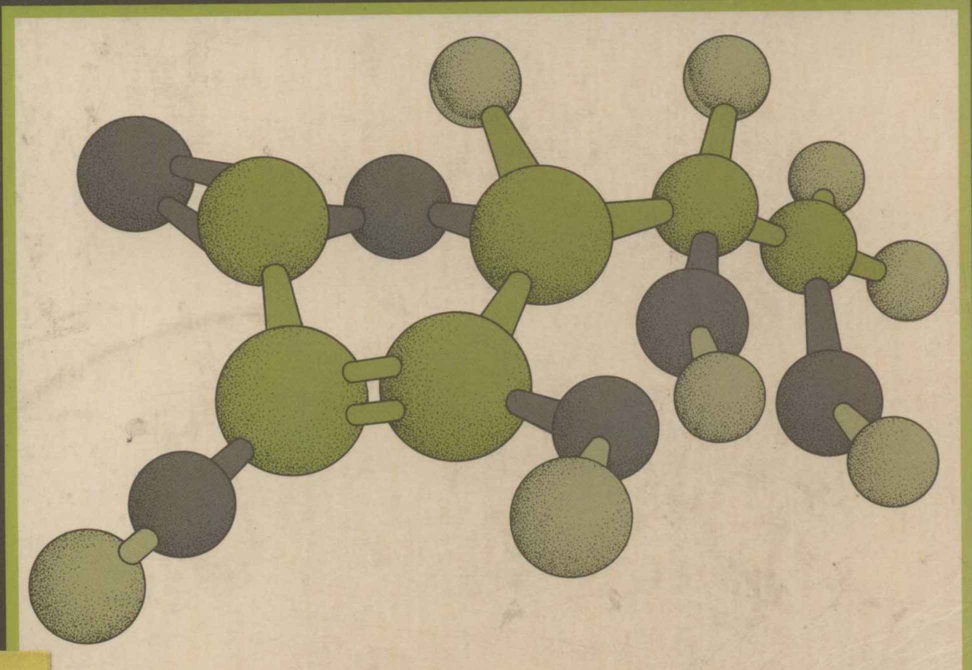
STUDY GUIDE

FRED H. REDMORE

INTRODUCTORY CHEMISTRY

KARL F. KUMLI



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PREFACE

This study guide has been written for the student in an introductory chemistry course, to accompany "Introductory Chemistry" by Karl F. Kumli.

What has been done here is first of all to point out what the objectives are for each chapter, since the first step in learning anything is to understand just what it is you are trying to learn. This is followed by a somewhat brief summary of the main points; that is, a short discussion of each topic pointing out the major aspects, clarifying points that often cause trouble and summarizing in various types of tables and charts when applicable. The material has been written in outline form, since I feel that this is the most logical way to present a summary of anything. This is what you, the student, should learn to do as you study chemistry or any other subject; that is, try to outline the material covered in your textbook and in the lecture and summarize it in your own words. The material in the first eleven chapters is more complete than in the remainder of the book, since these are the basic principles of chemistry which the rest of the material is based on.

And finally, each chapter concludes with a list of questions and problems, and a sample quiz. The answers to all of these are found in the back of the book so that you can check your work as you go along. The first question in each set consists of a list of terms for which you should write out definitions (the answers for these are not given). This should be done in your own words, not just copied from a book. If you can write out a definition in your own words it will mean much more to you, even if it doesn't sound as good as the way it is stated in a textbook.

This book has been written for you, the student, to help you in your study of beginning chemistry. It should not be used as a substitute for the textbook and lecture material, but as a guide to help you in your study of material covered in the text and lecture.

In using this study guide, you may find it helpful to read through the material covered here before it is discussed in class, or even before you read the text, to get a general idea of what topics are to be covered and what the general objectives are (there may well be other objectives that your instructor may wish you to master, in addition to these or possibly instead of some of these). After the material has been covered in class, the study guide information, along with the lecture notes that you take in class and any outside reading that you do, should be used to help you organize the topics to be learned and apply them by doing the exercises included. Answers to all of these exercises are included in the appendix. If you cannot understand the material or cannot work the exercises, get some help as soon as possible; it is essential that you understand each topic before going on to the next one.

Also in the appendix, along with the answers to all exercises, there is a section on nomenclature of inorganic compounds (to be used whenever this topic is covered in the course--probably near the beginning), a section on balancing oxidation-reduction equations, and a set of sample quizzes (one for each chapter) with all answers given for these also.

To summarize, I would suggest that you attend class regularly, read your text, and ask questions about things that you don't understand. Then use the material in this book to review and check yourself as you go along. With some effort on your part I feel that you will find the study of chemistry interesting.

Any comments you may have on this study guide or suggestions for ways that it could be made more useful would be greatly appreciated.

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Chapter 1

THE NATURE AND SCOPE OF CHEMISTRY

OBJECTIVES

I. To become familiar with the following terms:
   A. Matter
   B. Energy
   C. Chemistry
   D. Scientific Method

II. To understand the classification of matter into the following areas (including becoming familiar with all terms used)
   A. Pure substances
      1. Elements (composed of atoms)
      2. Compounds (composed of molecules or ions)
   B. Mixtures
      1. Homogeneous (solutions)
      2. Heterogeneous

III. To learn the units used in the measurement of matter, and how to convert from one unit to another
   A. Metric system
      1. Prefixes used
      2. Conversion between the metric and British systems
   B. Conversion factors
      1. Converting between any units within either of the systems
      2. Converting units between systems
      3. Other types of conversions
   C. Density
      1. Definition
      2. How it is used in calculations
   D. Exponents and scientific notation
      1. Writing numbers in exponential form
      2. Carrying out operations of multiplication, division, addition and subtraction and taking square and cube roots, with exponential numbers
      3. Scientific notation
   E. Significant figures
      1. Which digits are significant
      2. Use of correct number of significant digits in calculations

I. Introduction. The first step in studying chemistry is to become familiar with some basic terms, including the following:

A. Matter. All of the material that the universe is made up of; that is, it is anything that has weight and occupies space.

B. Energy. The other quantity in the universe. All changes that take place in nature are accompanied by changes in energy. It is now known that matter and energy are interchangeable; that is, small amounts of matter can be converted to energy. (This subject will be discussed in more detail later). Also, it is now known that the total amount of energy (or matter - energy) in the universe is constant and therefore the law of conservation of energy (or matter - energy) says that energy can neither be created nor destroyed, it can only be changed from one form to another.

C. Chemistry. This is the study of matter and is concerned with the changes that it undergoes and the energy transformations that accompany these changes.

D. The scientific method. A method of gaining knowledge by observing, attempting to explain the observations, making predictions based on these observations, and finding some method of verifying the prediction. This process, of course, includes many other things such as intuition, previous knowledge, and so forth.

II. Classification of matter. The classification of matter can be summarized by the chart shown in Fig. I-1. But to understand this classification you must first make sure that you are familiar with all of the terms used, including the following:

A. Pure substance. A material which has a definite and constant composition. This will either be an <u>element</u> or <u>compound</u>.

B. Atom. The smallest particle of matter that can undergo change in any ordinary chemical reaction. We shall see later that atoms are made up of smaller particles, called electrons, protons, and neutrons. But the study of chemistry is really only concerned with the atoms that matter is made up of and not the decomposition of atoms (in what are normally referred to as <u>nuclear</u> reactions).

C. Element. A type of material which contains only one kind of atom, and does not undergo decomposition to a simpler substance under ordinary conditions. (Again, the only changes this could undergo would be in a nuclear reaction.) There are presently only 105 different elements known; or 105 different kinds of atoms, and only 88 of these are known to occur in nature, the other 17 being man-made.

D. Compound. A substance made up of two or more elements in a definite and constant composition, which can be broken down to simpler substances by chemical means. For example, water, which has the formula H_2O, is always made up of two atoms of hydrogen (H) and one atom of oxygen (O). A formula gives the simplest whole number ratio of atoms that make up a compound. A symbol is used to represent each type of atom--that is, there is a symbol for each of the 105 elements. Therefore, another way of identifying a compound is by the fact that there is a formula for it. (The distinction between molecular and ionic compounds will be discussed in the next chapter.) We shall see later that since a particular element has a definite, relative weight, a compound will always have the same % composition by weight. For example, water is always 11.1% hydrogen and 88.9% oxygen, by weight.

E. Mixture. A combination of two or more substances in variable proportions. These can be homogeneous, in which case any sample taken will be the same, or heterogeneous, in which case samples will differ. For example, if salt is dissolved in water, to make a homogeneous mixture (called a solution) any portion of the solution will be the same, but there are an infinite number of different solutions of salt in water; that is, there is no formula for salt water. If sand is mixed with water a heterogeneous mixture results.

F. Chemical reaction. The process whereby two or more elements or compounds combine or rearrange to form one or more new compounds or elements. (Chemical reaction will be discussed in more detail in Chapter 4.)

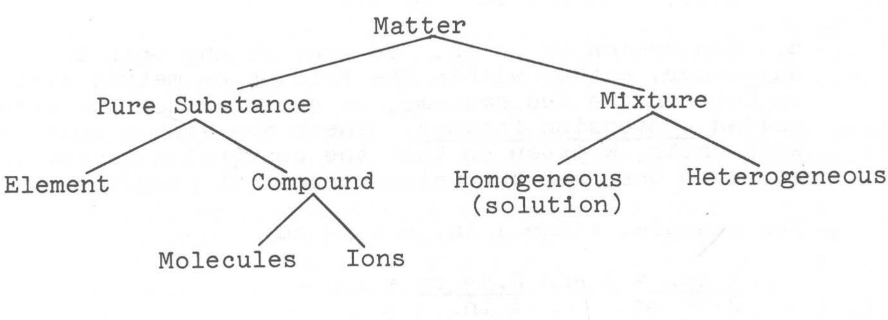

Fig. I-1

III. Measurement of matter. There are three types of measurements that we shall be concerned with. These are: mass (or weight), length, and volume.

A. Metric system. In the metric system, which is used in all scientific work, and which will eventually be used exclusively in the United States as it now is in most of the rest of the world, the prefixes in table I-1 are used. The basic units used are the gram (g), meter (m), and liter (l). The prefixes listed here are used with all types of units and must therefore be learned (there are also others, which can be found in any chemistry book, but these are the ones most commonly used).

kilo (k) = 1000 units	(1 kg = 1000 g)
deci (d) = 1/10 unit	(1 m = 10 dm)
centi (c) = 1/100 unit	(1 m = 100 cm)
milli (m) = 1/1000 unit	(1 l = 1000 ml)
micro (μ) = 1/1,000,000 unit	(1 μm = 10^{-6} m)

Table I-1

The following three basic conversion units must also be learned (or, any other conversion unit between the two systems, for each type of measurement).

1 lb. = 454 g

1 in. = 2.54 cm

1 l = 1.06 qt

Again, there are many others also, but if these are learned along with the following method of carrying out conversions this is all that is necessary.

Also, 1 ml = 1 cm^3 (or cc)

B. Conversion of units. To convert any unit to any other, either within the British or metric system, or between the two systems, we can use what are often called <u>conversion factors</u>. These are simply numbers, with units, written so that the conversion factor is equal to one, and the units cancel out properly.

For example, since 1 in. = 2.54 cm,

$$\frac{1\text{ in.}}{2.54\text{ cm}} = 1 \text{ and } \frac{2.54\text{ cm}}{1\text{ in.}} = 1$$

Therefore, to convert 5.8 cm to inches you simply multiply the 5.8 cm by the proper conversion factor to give the unit in., which is:

$$5.8\ \cancel{cm} \times \frac{1\ in.}{2.54\ \cancel{cm}} = 2.3\ in.$$

In this same way any quantity involving units can be made to give conversion factors, and any calculations involving relationships between quantities can be solved in this way, even though several steps may be involved.

Consider the following examples:

(1) Convert a speed of 50 mi/hr to cm/sec.

$$50\ \frac{\cancel{mi}}{\cancel{hr}} \times \frac{1\ \cancel{hr}}{60\ \cancel{min}} \times \frac{1\ \cancel{min}}{60\ sec} \times \frac{5280\ \cancel{ft}}{1\ \cancel{mi}} \times \frac{12\ \cancel{in.}}{1\ \cancel{ft}} \times \frac{2.54\ cm}{1\ \cancel{in.}}$$

$= 2{,}235$ cm/sec.

(2) Convert 1.20 g/cm^3 to lb/ft^3

$$\frac{1.20\ \cancel{g}}{\cancel{cm^3}} \times \frac{1\ lb}{4.54\ \cancel{g}} \times \frac{(2.54)^3\ \cancel{cm^3}}{13\ \cancel{in.^3}} \times \frac{(12)^3\ \cancel{in.^3}}{1^3\ ft^3}$$

$$= 74.9\ \frac{lb}{ft^3}$$

Remember that in conversions of this type, both the number and the unit must be cubed--for example 1 cubic foot (ft^3) does not equal 12 $in.^3$, but $(12)^3\ in^3$. Of course 1^3 is still one, but it is put in here to emphasize that both the number and the unit must be cubed.

We will see later that this same method can be used for many types of calculations, but not all calculations--only those actually involving equivalent quantities.

C. Density. Another commonly used measurement of matter is density. This gives the relationship between the mass and a given volume, or the "mass per unit volume", expressed as mass/volume. In the metric system this is almost always expressed as g/ml or g/cm^3 for liquids and solids, and g/l for gases. For example, the density of water is 1 g/ml and the density of sulfuric acid is about 1.8 g/ml. This means that one ml of water weighs one gram but one ml of sulfuric acid weighs 1.8 g. To use density in conversion problems, we can simply say, for water 1 g = 1 ml, and for sulfuric acid 1.8 g = 1 ml, and then these can be used as conversion factors, either as 1 g/1 ml or 1 ml/1 g for water and 1.8 g/1 ml or 1 ml/1.8 g for sulfuric acid.

D. Exponents and scientific notation. In exponential notation numbers are written in two parts, a whole number or decimal and a power of 10. For example, 2×10^2, 2.5×10^8, 3.2×10^{-3}, and so forth. Recall that a positive exponent (that is, power of 10) means that the decimal point must be moved to the right and a negative exponent means that it is moved to the left. That is, a negative exponent indicates a small number (and the larger the negative exponent the smaller the number is) and a positive exponent indicates a large number. Consider the following examples:

$10 = 10^1$ | $200 = 2 \times 10^2$

$100 = 10^2$ | $.0032 = 3.2 \times 10^{-3}$

$\frac{1}{10} = .1 = 10^{-1}$ | $.25 \times 10^2 = 25$

$\frac{1}{100} = .01 = 10^{-2}$ | $25 \times 10^{-3} = .025$

Operations with numbers involving exponents.

1. Multiplication. Multiply the numbers and add the exponents. For example:

$$(2 \times 10^2)\ (2 \times 10^3) = 4 \times 10^5$$

$$(3 \times 10^3)\ (2 \times 10^{-7}) = 6 \times 10^{-4}$$

2. Division. Divide the numbers and subtract the exponents algebraically. For example:

$$\frac{4 \times 10^5}{2 \times 10^2} = 2 \times 10^3$$

$$\frac{6 \times 10^{-8}}{3 \times 10^{-5}} = 2 \times 10^{-3}$$

3. Addition and subtraction. Here the exponents must be to the same power; for example, to add 2.50×10^2 and 1.20×10^3 one of these must be changed to: $2.50 \times 10^2 + 12.0 \times 10^2 = 14.5 \times 10^2 = 1.45 \times 10^3$ or, $.250 \times 10^3 + 1.20 \times 10^3 = 1.45 \times 10^3$

4. Square root, cube root, and so forth. To take a square root the exponent must be divisible by 2; to take a cube root it must be divisible by 3, and so forth. For example:

$$\sqrt{4 \times 10^{-8}} = 2 \times 10^{-4}$$

$$\sqrt{1.6 \times 10^3} = \sqrt{16 \times 10^4} = 4 \times 10^2$$

$$\sqrt[3]{8 \times 10^{-6}} = 2 \times 10^{-2}$$

$$\sqrt[3]{.27 \times 10^{-7}} = \sqrt[3]{27 \times 10^{-9}} = 3 \times 10^{-3}$$

In scientific notation the numbers are always written as a number greater than one and less than ten, with an exponent. Also, all significant digits, and only significant digits, are included. For example:

$2500 = 2.500 \times 10^3$ (not 2.5×10^3)

$0.00250 = 2.50 \times 10^{-3}$ (not 2.5×10^{-3})

$0.25 \times 10^{-4} = 2.5 \times 10^{-5}$ (not 2.50×10^{-5})

E. Significant figures. Numbers written in such a way that only the last digit is in doubt. For example, a value of 10.22 g means that this was measured to the closest one hundredth of a gram--likewise, a value of 10.00 g means the same thing, and therefore should not be written as 10 g or 10.0 g or 10.0000 g. A zero is only significant if it is NOT needed to show where the decimal point is (see the above example.) In doing calculations, the answer must have the same number of significant digits as the least number in any of the values used. For example:

(1.202) (3.01) (2.0) = 7.2, since the first number has four significant digits, the second has three, and the last only two, the answer must have two, but no more than two.

PROBLEM SET 1

1. Define the following terms (in your own words):

matter	element	symbol
energy	compound	formula
chemistry	mixture	equation
atom	density	conversion factor
mass	meter	significant digits
weight	gram	scientific method

2. Classify each of the following as: element, compound, or mixture:

A. Water	D. Sodium
B. Aluminum	E. Gasoline
C. Milk	F. Chlorine

3. Fill in the blanks:

A. 1 l = _____ ml

B. 1 lb. = _____ g

C. 1 g = _____ mg

D. 1 m = _____ cm

E. 1 l = _____ qt

F. 1 kg = _____ g

G. 1 in. = _____ cm

4. Carry out the following conversions:

A. 5.2 in. to cm

B. 2.5 l to qt

C. 1.2 kg to g

D. 10 mg to ounces

E. 10 m^2 to cm^2

F. 2 gal to ml

G. 5 ton to mg

H. 5 ft/min to cm/sec

5. Solve the following and express the answer in scientific notation:

A. $(2.5 \times 10^{-8})\ (2 \times 10^{5})$

B. $\dfrac{7 \times 10^{-5}}{2 \times 10^{-3}}$

C. $\dfrac{8 \times 10^{7}}{0.002}$

D. $(0.02)\ (80{,}000)$

E. $\dfrac{(8 \times 10^{6})\ (5 \times 10^{3})}{2 \times 10^{-4}}$

F. $(0.003)\ (0.004)\ (0.0005)$

G. $\dfrac{2 \times 10^{-3}}{5 \times 10^{-8}}$

6. Tell how many significant digits there are in each of the following:

A. 262.5 B. 250.2 C. 0.042
D. 201 E. 0.0020 F. 20.00

7. If 6 g of a substance occupy a volume of 1.5 ml, what is the density?

Chapter 2

ATOMIC STRUCTURE

OBJECTIVES

I. To become familiar with the three basic particles that the atom is composed of (electrons, protons, and neutrons) and learn what the relative mass and charge of each of these is

II. To learn the general arrangement of these three particles in the atom

III. To understand what the atomic number tells about an atom

IV. To learn what is meant by atomic mass, what isotopes are, and what the notation ${}^{A}_{Z}X$ represents

V. To understand what the atomic weight scale is and how mass defect affects it

VI. To learn how the electrons are arranged in the atom in terms of the following things
   A. Energy levels
   B. Sublevels and orbitals
   C. Electron spin
   D. Electron configuration
   E. Sequence of occupancy of orbitals
   F. Unpaired electrons
   G. Valence electrons

VII. To learn how the electron arrangement of atoms is related to the arrangement of the periodic table

VIII. To become familiar with some periodic variations in properties, such as
   A. Atomic size
   B. Ionization energy (ionization potential)
   C. Electron affinity

I. Subatomic particles. In discussing atoms it was stated that an atom is made up of electrons, protons, and neutrons. We now want to become familiar with these three types of particles and then learn how they are arranged in the atom. Table II-1 summarizes the relative properties. It is essential that you learn these, since a large part of the study of chemistry is concerned with the atom and the changes that atoms undergo.

In discussing mass (or weight) of atoms and particles that make up the atom, relative weights are used, since a single atom is too small to weigh. For example, it is known that oxygen is about 16 times as heavy as hydrogen and therefore on a relative weight scale we can say that hydrogen is about 1 and oxygen is about 16. These relative units are sometimes referred to as atomic mass units (amu). We will look at atomic weights and related topics later in the chapter, but for now you should understand what is meant by these relative units called amu.

particle	symbol	approximate mass (amu)	relative charge
electron	e^-	0	-1
proton	p	1	+1
neutron	n	1	0

Table II-1

II. Arrangement of particles in the atom. At the center of the atom is the nucleus, which contains the protons and neutrons, and hence has a + charge equal to the number of protons that it contains. Surrounding the nucleus is a cloud of electrons. The nucleus contains most of the mass of the atom, but only a very small fraction of the size; that is, the electron cloud makes up most of the volume of the atom.

III. Atomic number. Since an atom is a neutral species, there must always be the same number of electrons as there are protons, and the number of protons (or electrons) is called the atomic number. For example, all atoms of hydrogen, whose atomic number is 1, have 1 proton and 1 electron; all atoms of He (atomic number 2) have 2 protons and 2 electrons; and so forth.

IV. Atomic mass and isotopes. As was stated earlier this is based on a relative scale. It must be kept in mind that all atoms of a given element have the same atomic number, but they do not have the same atomic mass. This is due to the fact that different atoms of the same element can have different numbers of neutrons. These are called isotopes (that is, atoms with the same atomic number but different numbers of neutrons). Fortunately, there are at the most only a few isotopes that occur naturally for each element, and any sample of a given element has the same proportion of the various isotopes. The notation used to indicate a particular isotope is $^{A}_{Z}X$, where X is the symbol for the element, Z is the atomic number and A is the mass number, which represents the sum of the protons and neutrons in the nucleus. For example, in $^{14}_{6}C$ there are 6 protons (and 6 electrons) and 8 neutrons (14 - 6 = 8).