

NUCLEAR PHYSICS

by

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ADDISON-WESLEY PUBLISHING COMPANY, INC.

CAMBRIDGE 42, MASS.

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Library of Congress Catalog No. 54-5732

Second printing — January 1956

PREFACE

I have tried to write an elementary, yet coherent, account of nuclear physics suitable as an introduction to this field. The book is elementary in the sense that the reader is assumed to have had no previous exposure to atomic and nuclear physics. It is supposed that he has had a two-year college course in physics covering approximately the material contained in Sears' books: *Mechanics, Heat, and Sound*; *Electricity and Magnetism*; and *Optics*, in the Addison-Wesley Physics Series. It is also assumed that the reader is familiar with the differential and integral calculus; a one-year course in general chemistry would be helpful, if not essential. This book should therefore be useful in an advanced undergraduate course in nuclear physics, to engineers interested in the large-scale applications of nuclear physics grouped together under the name "nuclear engineering," or to anyone else with the indicated preparation who might be interested in nuclear physics.

The concepts and techniques of nuclear physics are not elementary in the sense that they are easily mastered. They have evolved through a great deal of experimental and theoretical research and cannot be expressed or explained in simple nontechnical terms. I have tried, therefore, to show how our present knowledge of atoms and nuclei has been developed, and how some of this knowledge is applied. In treating the experiments that yield information about atoms and nuclei and the ways in which these experiments are interpreted in terms of theory, I have tried to avoid both the devil of experimental complication and the deep blue sea of mathematical detail. The emphasis is on the principles underlying the experiments and on the experimental results rather than on the apparatus, on the physical ideas rather than on the details of the calculations. Thus, derivations have been included when they involve straightforward applications of physical principles in terms of mathematics not beyond the integral calculus. When it is necessary to use the results of more advanced methods, such as quantum mechanics, the details of the derivations are omitted; but the physical ideas contained in the results are discussed in some detail and related to the experimental data.

The plan of the book is based on the advice of the King of Hearts: "Begin at the beginning," the King said gravely, "and go on till you come to the end: then stop." The need for a reasonable limit on the size of the book makes it necessary, however, to define the beginning and end arbitrarily and to leave some things out on the way. The first part of the book, consisting of the first seven chapters, is devoted to the background of nuclear physics and begins with the chemical foundations of atomic theory. The second part, Chapters 8 through 17, treats the physics of the nucleus in a way that seems to me to be logical, yet consistent with the elementary nature of the book. The third part, Chapters 18 through 22, deals with special topics and applications and includes neutron physics

and nuclear fission, which do not fit conveniently into the scheme of the second part and at the same time lead into the most spectacular application of nuclear physics—"nuclear energy." The subjects of charged-particle accelerators and isotope separation, although not really nuclear physics, are closely related to important branches of this field; they are included for this reason and because they form, along with nuclear energy, the main part of nuclear engineering. I have not discussed such subjects as cosmic rays, mesons, and nuclear moments because I think that they need a more advanced treatment than the material covered.

I have included, for several reasons, a large number of literature references. For the skeptical or curious reader, the references will supply the experimental and theoretical details omitted from the text, as well as proofs of results that are stated without proof. A more important reason for including these references stems from the fact that atomic and nuclear physics are recently developed fields, and the original literature is more readily available than that of classical physics. The student can read for himself the original papers and books of the pioneers in this field—Rutherford, Bohr, Millikan, Moseley, Aston, Chadwick, Fermi, and others. The articles in which the basic discoveries were first published give, more than any textbook can, a feeling for the imagination and beauty inherent in modern physics. The current literature adds a feeling for the gradual, and sometimes painful, accumulation of experimental and theoretical information, and for the problems which have yet to be solved.

This book has been developed from the notes for a course of lectures given at the Brookhaven National Laboratory. I am indebted to colleagues who attended the lectures or used the notes for many valuable comments and suggestions, and to many workers in the field of nuclear physics for permission to cite and use their published results. Above all, I find it hard to express adequately my indebtedness and gratitude to Jean Harless for her unfailing patience, good humor, and skill in preparing the manuscript.

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June, 1954

Part I
The Background of Nuclear Physics

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CHAPTER 1

THE CHEMICAL FOUNDATIONS OF ATOMIC THEORY

The foundations of modern atomic theory were laid in the late 18th and 19th centuries in the attempt to understand the chemical properties of matter. Two of the fundamental quantities of atomic and nuclear physics—atomic weight and atomic number—had their origins in the correlation of the results of chemical experiments and in the systemization of the properties of the chemical elements. The laws of chemical combination were unified at the beginning of the 19th century by Dalton's Atomic Theory, which introduced the concept of atomic weight. The development of methods for determining the atomic weights of the elements and the investigation of chemical reactions were major contributions of 19th century chemistry. The systemization of the atomic weights and properties of the elements then led to the formulation of the periodic system and to the concepts of atomic number and atomic structure.

1-1 The laws of chemical combination. The experimental information which gave rise to the atomic theory of matter can be summarized in a few basic laws. The first of these was deduced from investigations such as those of Lavoisier, who showed that when tin is made to react with air in a closed vessel, the weight of the vessel and its contents before and after heating is the same. This constancy of the weight, which has been found to be true for all chemical reactions, is expressed in the *Law of Conservation of Mass*, which states that the mass of a system is not affected by any chemical change within the system.

It was found that when various metals are oxidized in excess air, one part by weight of oxygen always combines with: 1.52 parts by weight of magnesium, 2.50 parts of calcium, 1.12 parts of aluminum, 3.71 parts of tin, 3.97 parts of copper, and so on. If, as turns out to be more convenient, the weight of oxygen is taken to be 8, the following combining weights are obtained:

Oxygen	Magnesium	Calcium	Aluminum	Copper	Tin
8	12.16	20.04	8.96	31.76	29.68

The same combining weights are obtained when the oxides are prepared by methods other than simple oxidation. These experimental results, and many others like them, are expressed by the *Law of Definite Proportions*, which states that a particular chemical compound always contains the same elements united together in the same proportions by weight. It was also found that two elements can combine to form more than one compound. For example, at least five distinct oxides of nitrogen are known in which the relative proportions by weight of nitrogen and oxygen are:

<i>Nitrogen</i>	<i>Oxygen</i>
14	$8 = 1 \times 8$
14	$16 = 2 \times 8$
14	$24 = 3 \times 8$
14	$32 = 4 \times 8$
14	$40 = 5 \times 8$

The different weights of oxygen which can combine with the same weight of nitrogen are integral multiples of 8. Results of this kind are summarized by the *Law of Multiple Proportions*, which states that if two elements combine to form more than one compound the different weights of one which combine with the same weight of the other are in the ratio of small whole numbers.

The further study of the quantitative relations between the elements led to a fourth law of chemical combination. The *Law of Reciprocal Proportions* states that the weights of two (or more) substances which react separately with identical weights of a third are also the weights which react with each other, or simple multiples of them. In other words, if each of two substances *A* and *B* combines with a substance *C*, then *A* and *B* can combine with each other only in those proportions in which they combine with *C*, or in some simple multiple of those proportions. This law is illustrated by the reactions which take place between oxygen and sulphur, oxygen and zinc, and sulphur and zinc. Thus, 8 parts by weight of oxygen combine with 8.015 parts of sulphur to form an oxide of sulphur, and with 32.69 parts of zinc to form an oxide of zinc; in the formation of zinc sulphide, $65.38 (= 2 \times 32.69)$ parts of zinc combine with $32.06 (= 4 \times 8.015)$ of sulphur.

Another generalization deduced from the analysis of chemical compounds is that of the combining weight or chemical equivalent of an element. This concept was touched upon in the discussion of the oxides of various metals in connection with the Law of Definite Proportions. Consider a number of chemical compounds such as the following:

	<i>Percent</i>	<i>Percent</i>
1. Calcium oxide	Calcium 71.47	Oxygen 28.53
2. Water	Hydrogen 11.19	Oxygen 88.81
3. Hydrogen chloride	Hydrogen 2.76	Chlorine 97.23
4. Magnesium chloride	Magnesium 25.53	Chlorine 74.47
5. Silver chloride	Silver 75.26	Chlorine 24.71
6. Silver iodide	Silver 45.94	Iodine 54.06

If oxygen = 8 is taken as the standard, the amount of each of the other elements which combines with this standard amount of oxygen can be calculated. Thus, 28.53 parts by weight of oxygen combine with 71.47 parts of calcium, and

$$28.53:8 = 71.47:x, \text{ or } x = 20.04 \text{ for calcium,}$$

where *x* is the amount of the element, in this case calcium, which combines with 8 parts of oxygen. From the experimental data for water,

$$88.81:8 = 11.19:x, \text{ or } x = 1.008 \text{ for hydrogen.}$$

In hydrogen chloride, if the weight of hydrogen is taken as 1.008, that of chlorine is 35.45; in magnesium chloride, if the weight of chlorine is 35.45, that of magnesium is 12.16. When this procedure is continued, it is found that a number can be assigned to each element which represents the number of parts by weight of the given element that can combine with 8 parts by weight of oxygen or 1.008 parts by weight of hydrogen:

Oxygen	Calcium	Hydrogen	Chlorine	Magnesium	Silver	Iodine
8	20.04	1.008	35.45	12.16	107.88	126.9

The value obtained for magnesium by the above indirect process is the same as that obtained earlier from data on the direct oxidation of magnesium. The numbers obtained for the different elements are called the *combining weights* or *equivalent weights* of the elements.

The four laws of chemical combination: Conservation of Mass, Definite Proportions, Multiple Proportions, and Reciprocal Proportions, together with the concept of combining weights, summarized the basic experimental facts of chemical combination as known about 1800, and led to the Atomic Hypothesis proposed early in the 19th century by Dalton.

1-2 Dalton's Atomic Hypothesis. Dalton's Atomic Hypothesis was proposed (1803) to account for the facts expressed by the laws of chemical combination and was based on the following postulates:

1. The chemical elements consist of discrete particles of matter, atoms, which cannot be subdivided by any known chemical process and which preserve their individuality in chemical changes.

2. All atoms of the same element are identical in all respects, particularly in weight or mass; different elements have atoms differing in weight. Each element is characterized by the weight of its atom, and the combining weights of the elements represent the combining weights of their respective atoms.

3. Chemical compounds are formed by the union of atoms of different elements in simple numerical proportions, e.g., 1:1, 1:2, 2:1, 2:3.

It is easy to show that the laws of chemical combination can be deduced from these postulates. Since atoms undergo no change during a chemical process, they preserve their masses, and the mass of a compound is the sum of the masses of its elements; the result is the Law of Conservation of Mass. Since all atoms of the same element are identical in weight, and a compound is formed by the union of atoms of different elements in a simple numerical proportion, the proportions by weight in which two elements are combined in a given compound are always the same, giving the Law of Definite Proportions.

Consider next the case in which two elements *A* and *B* can form two different compounds. Suppose that the first compound contains *m* atoms of *A* and *n* atoms of *B*, and that the second compound contains *p* atoms of *A* and *q* atoms of *B*. If *a* is the weight of an atom of *A*, and *b* is the

weight of an atom of B , then the first compound contains ma parts of A and nb parts of B , and one part of A combines with nb/ma parts of B . Similarly, in the second compound, one part of A combines with qb/pa parts of B . Hence, the weights of B combined with a fixed weight of A are in the proportion

$$\frac{n}{m} : \frac{q}{p}, \text{ or } np : mq.$$

According to the third postulate, n , m , p , and q are all small integers; the products np and mq are therefore also small integers and the weights of B combining with a fixed weight of A are in the ratio of two small integers. This treatment can be extended to any number of compounds formed by two elements, giving the Law of Multiple Proportions. Similar arguments lead to the Law of Reciprocal Proportions. Compounds of the elements A and B are formed according to the scheme: m atoms of A and n atoms of B . Compounds of the elements A and C contain p atoms of A and q atoms of C , and compounds of the elements B and C contain x atoms of B and y atoms of C . If a , b , and c are the weights of the atoms A , B , and C , respectively, then one part of A combines with nb/ma parts of B , and qc/pa parts of C . It follows that

$$\frac{mq}{np} : \frac{b}{c}.$$

In the compound formed by B and C , the proportion by weight of B and C is

$$xb : yc, \text{ or } \frac{y}{x} : \frac{b}{c}.$$

Hence,

$$\frac{y}{x} = \frac{mq}{np}.$$

Since, by the third postulate, all of the quantities in the last equation are small integers, then y , x are the same as q , n or small integral multiples of them. This is the Law of Reciprocal Proportions.

The laws of chemical combination have thus been deduced from Dalton's postulates and the latter form the basis of an atomic theory of matter. Dalton's theory was incomplete, however, because it provided no way of determining even the relative weights of the atoms of the different elements. This difficulty arose because Dalton had no way of finding out how many atoms of each element combine to form a compound. If W_1 and W_2 are the weights of two elements which combine to form a compound, then

$$\frac{W_1}{W_2} = \frac{n_1 A_1}{n_2 A_2},$$

where A_1 and A_2 are the atomic weights and n_1 and n_2 are the whole numbers of atoms of each element which enter into combination. When the ratio $n_1 : n_2$ is known, the value of the ratio $W_1 : W_2$ can fix only the ratio of the atomic weights. To apply his theory, Dalton was forced to

make arbitrary assumptions; he assumed, for example, that if only one compound of two elements is known, it contains one atom of each element. Water was regarded as a compound of one atom of hydrogen and one atom of oxygen, the existence of hydrogen peroxide being unknown at the time. Dalton's assumption was simple, but wrong, and led to many difficulties in the application of his atomic theory to the rapidly growing field of chemistry.

1-3 Avogadro's Hypothesis and the molecule. One of the difficulties met by the Atomic Hypothesis arose as a result of studies of the combining properties of gases. Gay-Lussac showed (1805-1808) that when chemical reactions occur between gases, there is always a simple relation between the volumes of the interacting gases, and also of the products, if these are gases. When the reacting gases and the products are under the same conditions of temperature and pressure:

- 1 volume of hydrogen + 1 volume of chlorine
→ 2 volumes of hydrogen chloride;
- 2 volumes of hydrogen + 1 volume of oxygen
→ 2 volumes of steam;
- 3 volumes of hydrogen + 1 volume of nitrogen
→ 2 volumes of ammonia.

It follows that if elements in a gaseous state combine in simple proportions by volume, and if the elements also combine in simple proportions by atoms, then the numbers of atoms in equal volumes of the reacting gases must be simply related. Dalton assumed that equal volumes of different gases under the same physical conditions contain an equal number, say n , of atoms. Under this assumption, when 1 volume (n atoms) of hydrogen reacts with 1 volume (n atoms) of chlorine, then 2 volumes ($2n$ "compound atoms") of hydrogen chloride are formed. Every atom of hydrogen and chlorine must then be split in half to form two "compound atoms" of hydrogen chloride, a result which contradicts that postulate of Dalton's hypothesis according to which atoms cannot be subdivided by any chemical process.

Avogadro (1811) showed that the difficulty could be resolved if a distinction is made between elementary atoms and the small particles of a gas. He assumed that the latter are aggregates of a definite number of atoms, and called the aggregates *molecules* in order to distinguish them from the elementary atoms. He then postulated that equal volumes of all gases under the same physical conditions contain the same number of *molecules*. This hypothesis made it possible to interpret Gay-Lussac's Law of Combining Volumes in terms of the Atomic Hypothesis. Assume that each molecule of hydrogen or chlorine consists of two elementary atoms, and suppose that 1 volume of hydrogen or chlorine contains n molecules. These molecules react to form $2n$ molecules of hydrogen chloride, each

molecule containing one atom of hydrogen and one atom of chlorine. Although the atoms cannot be split so that one atom of hydrogen or chlorine enters into the composition of two molecules of hydrogen chloride, one molecule of hydrogen and one molecule of chlorine can be divided between two molecules of hydrogen chloride. In this way, Avogadro's Hypothesis allowed Gay-Lussac's Law to be reconciled with Dalton's Atomic Hypothesis. The detailed analysis of many reactions between gases has shown that the molecules of the gaseous elements hydrogen, oxygen, chlorine, and nitrogen contain two atoms; the inert gases helium, neon, argon, krypton, and xenon have one atom per molecule.

1-4 Molecular and atomic weights of gaseous elements. The relative molecular weights of gaseous substances can be determined from measurements of relative densities. The relative density of a gas is defined as the ratio of the weight of a given volume of the gas to that of an equal volume of a standard gas measured at the same temperature and pressure. In view of Avogadro's Hypothesis, it is possible to write for any gaseous substance:

$$\frac{\text{Molecular weight of gas}}{\text{Molecular weight of standard gas}} = \frac{\text{Weight of any volume of gas}}{\text{Weight of equal volume of standard gas}} \\ = \frac{\text{Density of gas}}{\text{Density of standard gas}}.$$

The choice of the standard gas is arbitrary and may be, for example, hydrogen, oxygen, or air. If hydrogen is chosen as the standard, and if its atomic weight is arbitrarily taken to be unity, the relative molecular weight of a gas is

Molecular weight = $2 \times$ density of gas relative to that of hydrogen,

since a molecule of hydrogen contains two atoms. When the relative molecular weight of a gaseous element has been determined in this way, its relative atomic weight can be determined when the number of atoms per molecule is known.

The results for some molecular weight determinations based on this method are shown in Table 1-1. The densities of the gases are given in grams per liter measured under standard conditions (0°C and 760 mm mercury pressure, abbreviated as N.T.P.). The third column gives the molecular weight relative to that of hydrogen taken as 2. Oxygen, with its molecular weight arbitrarily taken as 32, may also be used as a standard, and the molecular weights relative to this standard are given in the fourth column. High-precision work on the combining volumes of gases has shown that there are slight deviations from Gay-Lussac's Law which have been traced to deviations from the ideal gas laws. When corrections are made for these deviations, the results of the fifth column are obtained for the molecular weights relative to that of oxygen taken as 32. Of the gaseous elements listed, hydrogen, oxygen, nitrogen, fluorine, and chlorine

are diatomic and their relative atomic weights are half of the molecular weights; the relative atomic weights of the inert gases are the same as their molecular weights.

TABLE 1-1
MOLECULAR WEIGHTS OF GASEOUS ELEMENTS

Element	Density (g/l at N.T.P.)	Molecular Weight		Molecular Weight M (oxygen) = 32 corrected for deviations from ideal gas laws
		M (hydrogen) = 2	M (oxygen) = 32	
Hydrogen	0.08988	2.000	2.013	2.016
Oxygen	1.42904	31.816	32.000	32.000
Nitrogen	1.25055	27.828	28.000	28.016
Fluorine	1.696	37.738	37.977	38.00
Chlorine	3.214	71.52	71.97	70.914
Helium	0.17847	3.971	3.996	4.003
Neon	0.90035	20.034	20.161	20.183
Argon	1.7837	39.690	39.942	39.944
Krypton	3.708	82.510	83.03	83.7
Xenon	5.851	130.19	131.02	131.3

1-5 The standard atomic weight. It has been emphasized that the values of the molecular and atomic weights determined from density measurements are relative, not absolute, values. For practical purposes, it is necessary to fix a standard for atomic weights, and this standard is taken as oxygen with an atomic weight of 16 and a molecular weight of 32. The choice of oxygen was based on the criterion of convenience. Oxygen is a useful standard for chemists because nearly all the elements form stable compounds with oxygen, and it was thought that the determination of the atomic weight of an element should be connected with the standard as closely as possible. Hydrogen was sometimes used as a standard, but it forms very few compounds with metals which are suitable for atomic weight determinations; consequently hydrogen was rejected in favor of oxygen.

The choice of the value of the atomic weight of oxygen was based on the desire for convenient values for the other atomic and molecular weights. Experimental evidence shows that 1.008 parts by weight, say grams, of hydrogen combine with 8 grams of oxygen. It has also been found that two atoms of hydrogen combine with one atom of oxygen to form one molecule of water. Hence, 1.008 grams of hydrogen contain twice as many atoms as 8 grams of oxygen. A possible choice for the atomic weight of oxygen is its combining, or equivalent, weight: 8 units. If this choice were made, then 1.008 (of the same units) would be the weight of two atoms of hydrogen, and the atomic weight of hydrogen would be 0.504. Now, hydrogen is the lightest of all the elements. If the standard