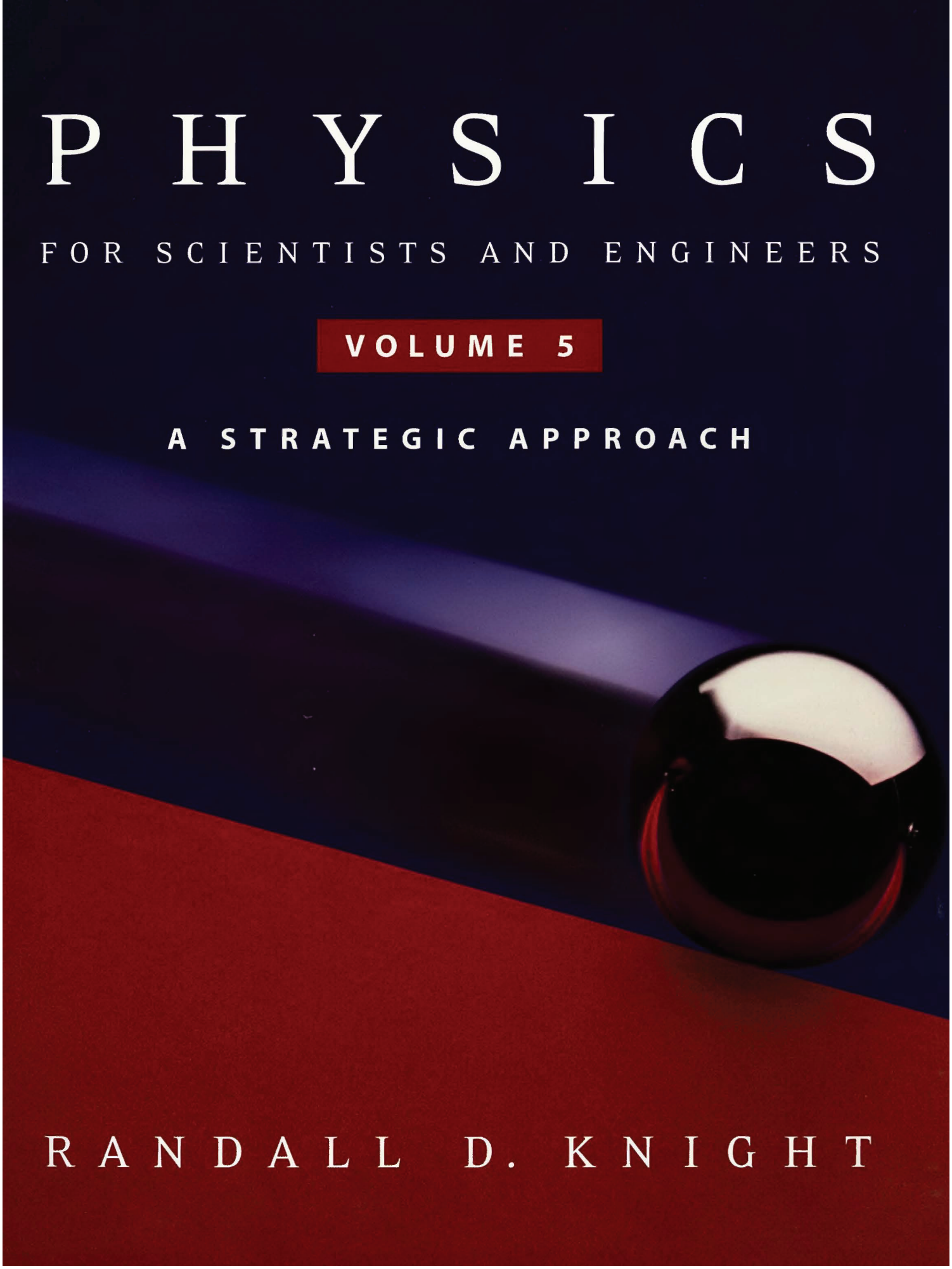
PHYSICS

FOR SCIENTISTS AND ENGINEERS

VOLUME 5

A STRATEGIC APPROACH



RANDALL D. KNIGHT

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Useful Data

M_e	Mass of the earth	5.98×10^{24} kg	
R_e	Radius of the earth	6.37×10^{6} m	
g	Acceleration due to gravity	9.80 m/s^2	
G	Gravitational constant	6.67×10^{-11} N m^2/kg^2	
k_B	Boltzmann's constant	1.38×10^{-23} J/K	
R	Gas constant	8.31 J/mol K	
N_A	Avogadro's number	6.02×10^{23} particles/mol	
T_0	Absolute zero	−273°C	
p_{atm}	Standard atmosphere	101,300 Pa	
v_{sound}	Speed of sound in air at 20°C	343 m/s	
m_p	Mass of the proton (and the neutron)	1.67×10^{-27} kg	
m_e	Mass of the electron	9.11×10^{-31} kg	
K	Coulomb's law constant ($1/4\pi\epsilon_0$)	8.99×10^{9} N m^2/C^2	
ϵ_0	Permittivity constant	8.85×10^{-12} C^2/N m^2	
μ_0	Permeability constant	1.26×10^{6} T m/A	
e	Fundamental unit of charge	1.60×10^{-19} C	
c	Speed of light in vacuum	3.00×10^{8} m/s	
h	Planck's constant	6.63×10^{-34} J s	4.14×10^{-15} eV s
$\hbar$	Planck's constant	1.05×10^{-34} J s	6.58×10^{-16} eV s
a_B	Bohr radius	5.29×10^{-11} m	

Common Prefixes

Prefix	Meaning
femto-	10^{-15}
pico-	10^{-12}
nano-	10^{-9}
micro-	10^{-6}
milli-	10^{-3}
centi-	10^{-2}
kilo-	10^{3}
mega-	10^{6}
giga-	10^{9}
terra-	10^{12}

Conversion Factors

Length
1 in = 2.54 cm
1 mi = 1.609 km
1 m = 39.37 in
1 km = 0.621 mi

Velocity
1 mph = 0.447 m/s
1 m/s = 2.24 mph = 3.28 ft/s

Mass and energy
1 u = 1.661×10^{-27} kg
1 cal = 4.19 J
1 eV = 1.60×10^{-19} J

Time
1 day = 86,400 s
1 year = 3.16×10^{7} s

Pressure
1 atm = 101.3 kPa = 760 mm of Hg
1 atm = 14.7 lb/in^2

Rotation
1 rad = $180°/\pi = 57.3°$
1 rev = $360° = 2\pi$ rad
1 rev/s = 60 rpm

Mathematical Approximations

Binominal Approximation: $(1 + x)^n \approx 1 + nx$ if $x \ll 1$

Small-Angle Approximation: $\sin\theta \approx \tan\theta \approx \theta$ and $\cos\theta \approx 1$ if $\theta \ll 1$ radian

Greek Letters Used in Physics

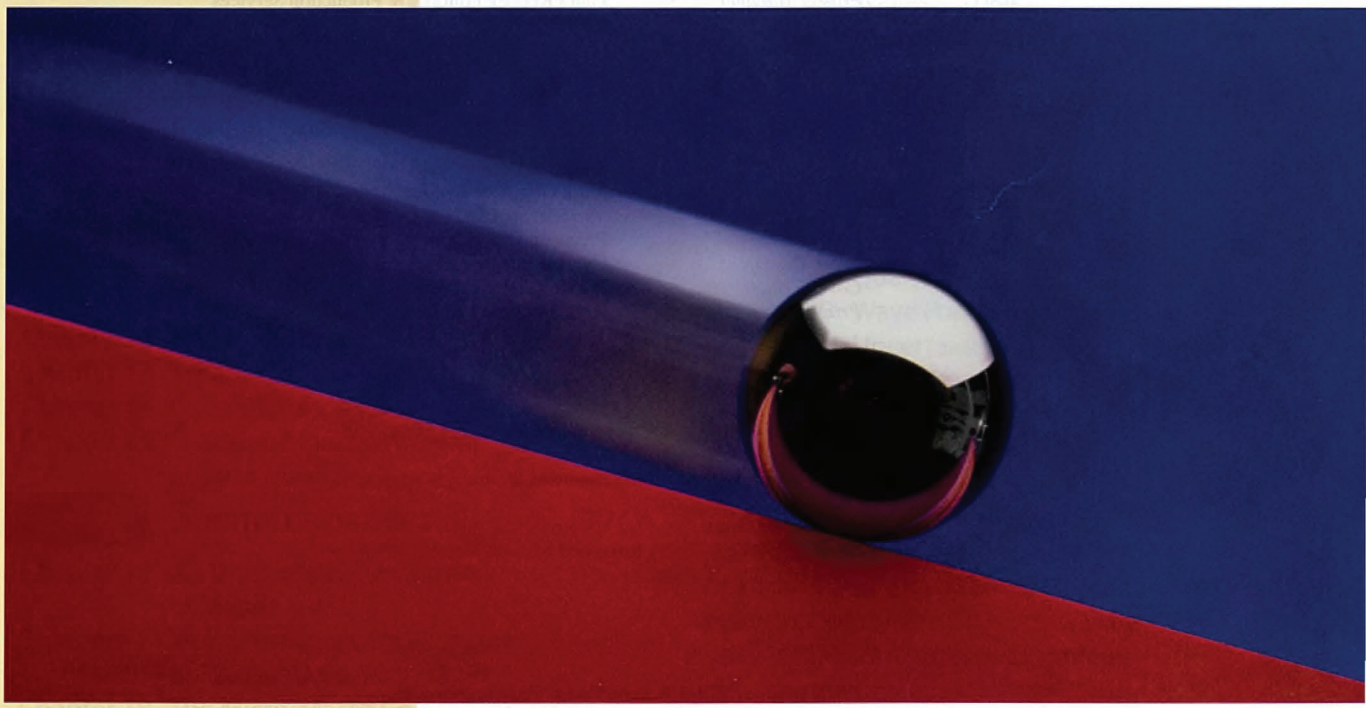
Alpha		α	Nu		ν
Beta		β	Pi		π
Gamma	Γ	γ	Rho		ρ
Delta	Δ	δ	Sigma	Σ	σ
Epsilon		ϵ	Tau		τ
Eta		η	Phi	Φ	ϕ
Theta	Θ	θ	Psi		ψ
Lambda		λ	Omega	Ω	ω
Mu		μ			

Physics for Scientists and Engineers Volume 5

A Strategic Approach

Randall D. Knight

California Polytechnic State University, San Luis Obispo



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42 Nuclear Physics



The technique known as *carbon dating* uses the radioactive decay of the naturally occurring carbon isotope ^{14}C to determine the age of fossils and archeological artifacts.

The nucleus of the atom is extremely remote from our everyday experience. Thus it comes as something of a surprise to notice the extent to which nuclear physics has become part of our modern technology and contemporary vocabulary: nuclear power and nuclear weapons, nuclear medicine and nuclear waste, nuclear fission and nuclear fusion.

Rutherford's discovery of the atomic nucleus marked the beginning of nuclear physics. Other physicists were soon designing experiments to probe within the nucleus and learn the properties of nuclear matter. In this final chapter, we'll explore the physics of the nucleus and look at some of the applications of nuclear physics.

One interesting application, shown in the photo, is the use of naturally occurring radioactivity to date archeological samples and geological formations. We'll also examine how exposure to nuclear radiation is measured and look at some of the uses of nuclear physics in medicine.

▶ Looking Ahead

The goal of Chapter 42 is to understand the physics of the nucleus and some of the applications of nuclear physics. In this chapter you will learn to:

- Interpret the basic structure of the nucleus.
- Understand how the strong force holds the nucleus together.
- Understand why some nuclei are unstable and undergo radioactive decay.
- Calculate the half-lives of radioactive decay.
- Apply nuclear physics to biology and medicine.

◀ Looking Back

The material in this chapter depends on basic atomic structure and on the quantized energy levels in potential-energy wells. Please review:

- Sections 37.6–37.7 Rutherford's model of the nucleus.
- Section 40.6 Finite potential-energy wells.

42.1 Nuclear Structure

The 1890s was a decade of mysterious rays. Cathode rays were being studied in several laboratories, and, in 1895, Röntgen discovered x rays. In 1896, after hearing of Röntgen's discovery, the French scientist A. H. Becquerel wondered if mineral crystals that fluoresce after exposure to sunlight were emitting x rays. He put a piece of film in an opaque envelope, then placed a crystal on top and left it in the sun. To his delight, the film in the envelope was exposed.

Becquerel thought he had discovered x rays coming from crystals, but his joy was short lived. He soon found that the film could be exposed equally well simply by being stored in a closed drawer with the crystals. Further investigation showed that the crystal, which happened to be a mineral containing uranium, was spontaneously emitting some new kind of ray. Rather than finding x rays, as he had hoped, Becquerel had discovered what became known as *radioactivity.*

Ernest Rutherford soon took up the investigation and found not one but three distinct kinds of rays emitted from crystals containing uranium. Not knowing what they were, he named them for their ability to penetrate matter and ionize air. The first, which caused the most ionization and penetrated the least, he called *alpha rays.* The second, with intermediate penetration and ionization, were *beta rays,* and the third, with the least ionization but the largest penetration, became *gamma rays.*

Within a few years, Rutherford was able to show that alpha rays are helium nuclei emitted from the crystal at very high velocities. These became the projectiles that he used in 1909 to probe the structure of the atom. The outcome of that experiment, as you learned in Chapter 37, was Rutherford's discovery that atoms have a very small, dense nucleus at the center.

Rutherford's discovery of the nucleus may have settled the question of atomic structure, but it raised many new issues for scientific research. Foremost among them were

- What is nuclear matter? What are its properties?
- What holds the nucleus together? Why doesn't the repulsive electrostatic force blow it apart?
- What is the connection between the nucleus and radioactivity?

These were the beginnings of **nuclear physics,** the study of the properties of the atomic nucleus.

Nucleons

The nucleus is a tiny speck in the center of a vastly larger atom. As Figure 42.1 shows, the nuclear diameter of roughly 10^{-14} m is only about 1/10,000 the diameter of the atom. What we call *matter* is overwhelming empty space!

You learned in Chapter 37 that the nucleus is composed of two types of particles: *protons* and *neutrons.* Together, these are referred to as **nucleons.** The role of the neutrons, which have nothing to do with keeping electrons in orbit, is an important issue that we'll address in this chapter. Table 42.1 summarizes the basic properties of protons and neutrons.

As you can see, protons and neutrons are virtually identical other than the fact that the proton has one unit of the fundamental charge e whereas the neutron is electrically neutral. The neutron is slightly more massive than the proton, but the difference is very small, only about 0.1%. Notice that the proton and neutron, like the electron, have an *inherent angular momentum* and magnetic moment with spin quantum number $s = \frac{1}{2}$. As a consequence, protons and neutrons obey the Pauli exclusion principle.

The number of protons Z is the element's **atomic number.** In fact, an element is identified by the number of protons in the nucleus, not by the number of orbiting electrons. Electrons are easily added and removed, forming negative and positive ions, but doing so doesn't change the element. The **mass number** A is defined to be $A = Z + N$, where N is the **neutron number.** The mass number is the total number of nucleons in a nucleus.

NOTE ▶ The mass number, which is dimensionless, is *not* the same thing as the atomic mass m. We'll look at actual atomic masses below. ◀

This picture of an atom would need to be 10 m in diameter if it were drawn to the same scale as the dot representing the nucleus.

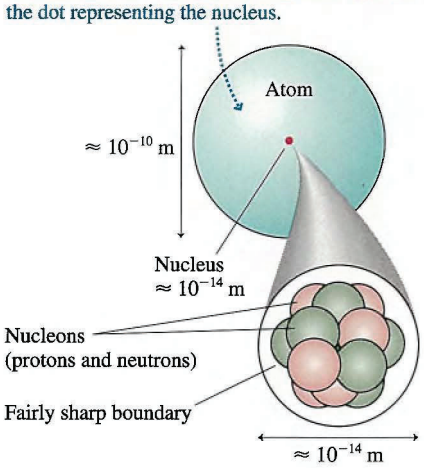


FIGURE 42.1 The nucleus is a tiny speck within an atom.

TABLE 42.1 Protons and neutrons

	Proton	Neutron
Number	Z	N
Charge q	$+e$	0
Spin s	$\frac{1}{2}$	$\frac{1}{2}$
Mass, in u	1.00728	1.00866

Isotopes and Isobars

It was discovered early in the 20th century that not all atoms of the same element (same Z) have the same mass. There are a *range* of neutron numbers that happily form a nucleus with Z protons, creating a series of nuclei having the same Z-value (i.e., they are all the same chemical element) but different A-values. Each A-value in a series of nuclei with the same Z-value is called an **isotope.**

Chemical behavior is determined by the orbiting electrons. All isotopes of one element have the same number of orbiting electrons (if the atoms are electrically neutral) and thus have the same chemical properties, but different isotopes of the same element can have quite different nuclear properties. In addition, macroscopic behavior that depends on mass, such as the diffusion of a gas, can slightly favor one isotope over another.

The notation used to label isotopes is ${}^{A}Z$, where the mass number A is given as a *leading* superscript. The proton number Z is not specified by an actual number but, equivalently, by the chemical symbol for that element. Hence ordinary carbon, which has six protons and six neutrons in the nucleus, is written ${}^{12}C$ and pronounced "carbon twelve." The radioactive form of carbon used in carbon dating is ${}^{14}C$. It has six protons, making it carbon, and eight neutrons.

More than 3000 isotopes are known. The majority of these are **radioactive,** meaning that the nucleus is not stable but, after some period of time, will either fragment or emit some kind of subatomic particle in an effort to reach a more stable state. Many of these radioactive isotopes are created by nuclear reactions in the laboratory and have only a fleeting existence. Only 266 isotopes are **stable** (i.e., nonradioactive) and occur in nature. We'll begin to look at the issue of nuclear stability in the next section.

The *naturally occurring* nuclei include the 266 stable isotopes and a handful of radioactive isotopes with such long half-lives, measured in billions of years, that they also occur naturally. The most well-known example of a naturally occurring radioactive isotope is the uranium isotope ${}^{238}U$. For each element, the fraction of naturally occurring nuclei represented by one particular isotope is called the **natural abundance** of that isotope.

Although there are many radioactive isotopes of the element iodine, iodine occurs *naturally* only as ${}^{127}I$. Consequently, we say that the natural abundance of ${}^{127}I$ is 100%. Most elements have multiple naturally occurring isotopes. The natural abundance of ${}^{14}N$ is 99.6%, meaning that 996 out of every 1000 naturally occurring nitrogen atoms are the isotope ${}^{14}N$. The remaining 0.4% of naturally occurring nitrogen is the isotope ${}^{15}N$, with one extra neutron.

A series of nuclei having the same A-value (the same mass number) but different values of Z and N are called **isobars.** For example, the three nuclei ${}^{14}C$, ${}^{14}N$, and ${}^{14}O$ are isobars with $A = 14$. Only ${}^{14}N$ is stable; the other two are radioactive.

Atomic Mass

You learned in Chapter 16 that atomic masses are specified in terms of the *atomic mass unit* u, defined such that the atomic mass of the isotope ${}^{12}C$ is exactly 12 u. The conversion to SI units is

$$1\text{ u} = 1.6605 \times 10^{-27}\text{ kg}$$

Alternatively, we can use Einstein's $E_0 = mc^2$ to express masses in terms of their energy equivalent. The energy equivalent of 1 u of mass is

$$\begin{aligned} E_0 &= (1.6605 \times 10^{-27}\text{ kg})(2.9979 \times 10^{8}\text{ m/s})^2 \\ &= 1.4924 \times 10^{-10}\text{ J} = 931.49\text{ MeV} \end{aligned} \tag{42.1}$$

Thus the atomic mass unit can be written

$$1\text{ u} = 931.49\text{ MeV}/c^2$$

It may seem unusual, but the units MeV/c^2 are units of mass.

NOTE ▶ We're using more significant figures than usual. Many nuclear calculations look for the small difference between two masses that are almost the same. Those two masses must be calculated or specified to four or five significant figures if their difference is to be meaningful. ◀

Table 42.2 shows the atomic masses of the electron, the nucleons, and three important light elements. Appendix C contains a more complete list. Notice that the mass of a hydrogen atom is the sum of the masses of a proton and an electron. But a quick calculation shows that the mass of a helium atom (2 protons, 2 neutrons, and 2 electrons) is 0.03038 u *less* than the sum of the masses of its constituents. The difference is due to the *binding energy* of the nucleus, a topic we'll look at in Section 42.2.

TABLE 42.2 Some atomic masses

Particle	Symbol	Mass (u)	Mass (MeV/c^2)
Electron	e	0.00055	0.51
Proton	p	1.00728	938.28
Neutron	n	1.00866	939.57
Hydrogen	^{1}H	1.00783	938.79
Deuterium	^{2}H	2.01410	1876.12
Helium	^{4}He	4.00260	3728.40

The isotope ^{2}H is a hydrogen atom in which the nucleus is not simply a proton but a proton and a neutron. Although the isotope is a form of hydrogen, it is called **deuterium.** The natural abundance of deuterium is 0.015%, or about 1 out of every 6700 hydrogen atoms. Water made with deuterium (sometimes written D_2O rather than H_2O) is called *heavy water.*

NOTE ▶ Don't let the name *deuterium* cause you to think this is a different element. Deuterium is an isotope of hydrogen. Chemically, it behaves just like ordinary hydrogen. ◀

The *chemical* atomic mass shown on the periodic table of the elements is the *weighted average* of the atomic masses of all naturally occurring isotopes. For example, chlorine has two stable isotopes: ^{35}Cl, with $m = 34.97$ u, is 75.8% abundant and ^{37}Cl, at 36.97 u, is 24.2% abundant. The average, weighted by abundance, is $0.758 \times 34.97 + 0.242 \times 36.97 = 35.45$. This is the value shown on the periodic table and is the correct value for most chemical calculations, but it is not the mass of any particular isotope of chlorine.

NOTE ▶ The atomic masses of the proton and the neutron are both ≈1 u. Consequently, the value of the mass number A is *approximately* the atomic mass in u. The approximation $m \approx A$ u is sufficient in many contexts, such as when we're calculating the masses of atoms in the kinetic theory of gases, but in nuclear physics calculations, we almost always need the more accurate mass values that you find in Table 42.2 or Appendix C. ◀

Nuclear Size and Density

Unlike the atom's electron cloud, which is quite diffuse, the nucleus has a fairly sharp boundary. Experimentally, the radius of a nucleus with mass number A is found to be

$$r = r_0 A^{1/3} \tag{42.2}$$

where $r_0 = 1.2$ fm. Recall that 1 fm = 1 femtometer = 10^{-15} m.

As Figure 42.2 shows, the radius is proportional to $A^{1/3}$, but the volume of the nucleus (proportional to r^3) is directly proportional to A, the number of nucleons. A nucleus with twice as many nucleons will occupy twice as much volume. This finding has several implications:

- Nucleons are incompressible. Adding more nucleons doesn't squeeze the inner nucleons into a smaller volume.
- The nucleons are tightly packed, looking much like the drawing in Figure 42.1.
- Nuclear matter has a constant density.

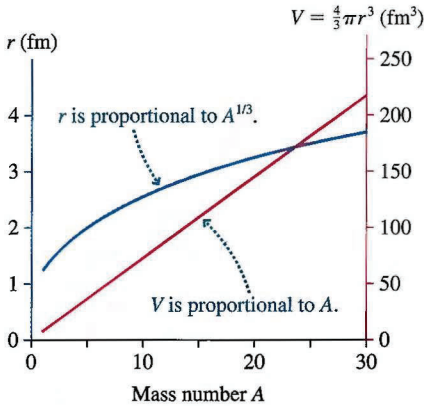


FIGURE 42.2 The nuclear radius and volume as a function of A.

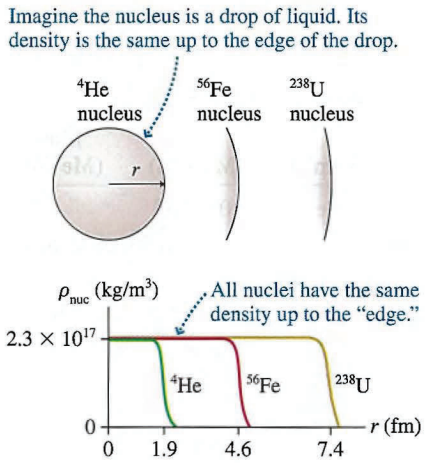


FIGURE 42.3 Density profiles of three nuclei.

In fact, we can use Equation 42.2 to estimate the density of nuclear matter. Consider a nucleus with mass number A. Its mass, within 1%, is A atomic mass units. Thus

$$\rho_{nuc} \approx \frac{A\,\text{u}}{\frac{4}{3}\pi r^3} = \frac{A\,\text{u}}{\frac{4}{3}\pi r_0^3 A} = \frac{1\,\text{u}}{\frac{4}{3}\pi r_0^3} = \frac{1.66 \times 10^{-27}\,\text{kg}}{\frac{4}{3}\pi(1.2 \times 10^{-15}\,\text{m})^3} = 2.3 \times 10^{17}\,\text{kg/m}^3 \quad (42.3)$$

The fact that A cancels means that **all nuclei have this density.** It is a staggeringly large density, roughly 10^{14} times larger than the density of familiar liquids and solids. One early objection to Rutherford's model of a nuclear atom was that matter simply can't have a density this high. Although we have no direct experience with such matter, nuclear matter really is this dense.

Figure 42.3 shows the density profiles of three nuclei. The constant density right to the edge is analogous to that of a drop of incompressible liquid, and, indeed, one successful model of many nuclear properties is called the **liquid-drop model.** Notice that the range of nuclear radii, from small helium to large uranium, is not quite a factor of 4. The fact that ^{56}Fe is a fairly typical atom in the middle of the periodic table is the basis for our earlier assertion that the nuclear diameter is roughly 10^{-14} m, or 10 fm.

STOP TO THINK 42.1 Three electrons orbit a neutral ^{6}Li atom. How many electrons orbit a neutral ^{7}Li atom?

42.2 Nuclear Stability

We've already noted that less than 10% of the known nuclei are stable (i.e., not radioactive). Because nuclei are characterized by two independent numbers, N and Z, it is useful to show the known nuclei on a plot of neutron number N versus proton number Z. Figure 42.4 shows such a plot. Stable nuclei are represented by blue diamonds and unstable, radioactive nuclei by red dots.

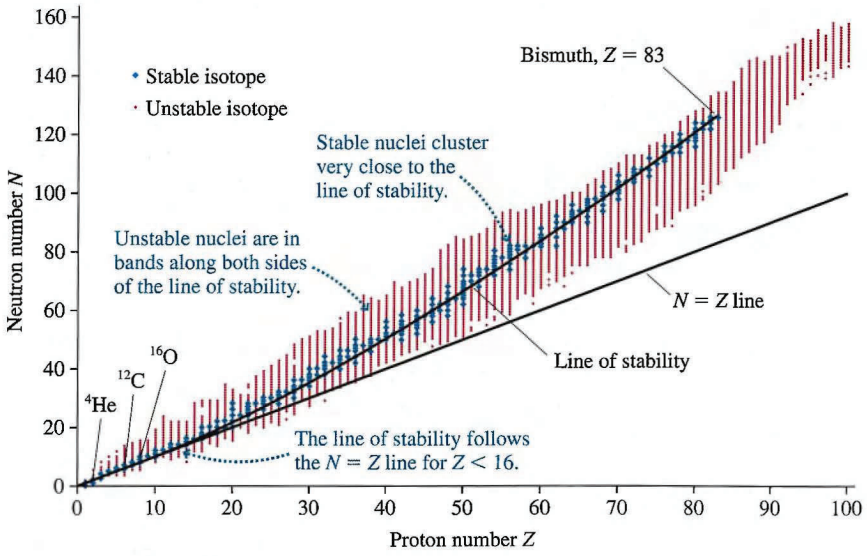


FIGURE 42.4 Stable and unstable nuclei shown on a plot of neutron number N versus proton number Z.

We can make several observations from this graph:

- The stable nuclei cluster very close to the curve called the **line of stability.**
- There are no stable nuclei with $Z > 83$ (bismuth).
- Unstable nuclei are in bands along both sides of the line of stability.
- The lightest elements, with $Z < 16$, are stable when $N \approx Z$. The familiar elements ^{4}He, ^{12}C, and ^{16}O all have equal numbers of protons and neutrons.
- As Z increases, the number of neutrons needed for stability grows increasingly larger than the number of protons. The N/Z ratio is $\approx$1.2 at $Z = 40$ but has grown to $\approx$1.5 at $Z = 80$.

These observations—especially $N \approx Z$ for small Z but $N > Z$ for large Z—cry out for an explanation. The quantum-mechanical model of the nucleus that we'll develop in Section 42.4 will provide the explanation we seek.

STOP TO THINK 42.2 The isobars corresponding to one specific value of A are found on the plot of Figure 42.4 along

a. A vertical line.
b. A horizontal line.
c. A diagonal line that goes up and to the right.
d. A diagonal line that goes up and to the left.

Binding Energy

A nucleus is a *bound system.* That is, you would need to supply energy to disperse the nucleons by breaking the nuclear bonds between them. Figure 42.5 shows this idea schematically.

You learned a similar idea in atomic physics. The energy levels of the hydrogen atom are negative numbers because the bound system has less energy than a free proton and electron. The energy you must supply to an atom to remove an electron is called the *ionization energy.*

In much the same way, the energy you would need to supply to a nucleus to disassemble it into individual protons and neutrons is called the **binding energy.** Whereas ionization energies of atoms are only a few eV, the binding energies of nuclei are tens or hundreds of MeV, energies large enough that their mass equivalent is not negligible.

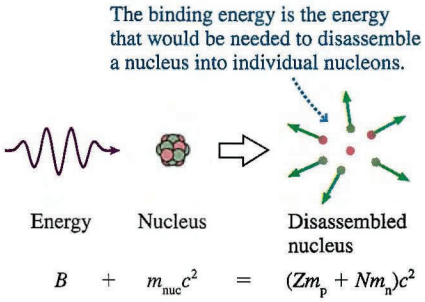


FIGURE 42.5 The nuclear binding energy.

Activ Online Physics 19.2

Consider a nucleus with mass m_{nuc}. It is found experimentally that m_{nuc} is *less* than the total mass $Zm_p + Nm_n$ of the Z protons and N neutrons that form the nucleus, where m_p and m_n are the masses of the proton and neutron. That is, the energy equivalent $m_{nuc}c^2$ of the nucleus is less than the energy equivalent $(Zm_p + Nm_n)c^2$ of the individual nucleons. The binding energy B of the nucleus (not the entire atom) is defined as

$$B = (Zm_p + Nm_n - m_{nuc})c^2 \qquad (42.4)$$

This is the energy you would need to supply to disassemble the nucleus into its pieces.

The practical difficulty is that laboratory scientists use mass spectroscopy to measure *atomic* masses, not nuclear masses. The atomic mass m_{atom} is m_{nuc} plus the mass Zm_e of Z orbiting electrons. (Strictly speaking, we should allow for the binding energy of the electrons, but these binding energies are roughly 10^6 smaller than the nuclear binding energies and can be neglected in all but the most precise measurements and calculations.)

Fortunately, we can switch from the nuclear mass to the atomic mass by the simple trick of both adding and subtracting Z electron masses. Begin by writing Equation 42.4 in the equivalent form

$$B = (Zm_p + Zm_e + Nm_n - m_{nuc} - Zm_e)c^2 \qquad (42.5)$$

Now $m_{nuc} + Zm_e = m_{atom}$, the atomic mass, and $Zm_p + Zm_e = Z(m_p + m_e) = Zm_H$, where m_H is the mass of a hydrogen *atom*. Finally, use the conversion factor 1 u = 931.49 MeV/c^2 to write c^2 = 931.49 MeV/u. The binding energy is then

$$B = (Zm_H + Nm_n - m_{atom}) \times (931.49 \text{ MeV/u}) \quad \text{(binding energy)} \tag{42.6}$$

where all the three masses are in atomic mass units.

EXAMPLE 42.1 The binding energy of iron

What is the binding energy of the ^{56}Fe nucleus?

SOLVE The isotope ^{56}Fe has $Z = 26$ and $N = 30$. The atomic mass of ^{56}Fe, found in Appendix C, is 55.9349 u. Thus the mass difference between the ^{56}Fe nucleus and its constituents is

$$B = 26(1.0078 \text{ u}) + 30(1.0087 \text{ u}) - 55.9349 \text{ u} = 0.528 \text{ u}$$

where, from Table 42.2, 1.0078 u is the mass of the hydrogen atom. Thus the binding energy of ^{56}Fe is

$$B = (0.538 \text{ u}) \times (931.49 \text{ MeV/u}) = 492 \text{ MeV}$$

ASSESS The binding energy is extremely large, the energy equivalent of more than half the mass of a proton or a neutron.

The nuclear binding energy increases as A increases simply because there are more nuclear bonds. A more useful measure for comparing one nucleus to another is the quantity B/A, called the *binding energy per nucleon*. Iron, with B = 492 MeV and A = 56, has 8.79 MeV per nucleon. This is the amount of energy, on average, you would need to supply in order to remove *one* nucleon from the nucleus. Nuclei with larger values of B/A are more tightly held together than nuclei with smaller values of B/A.

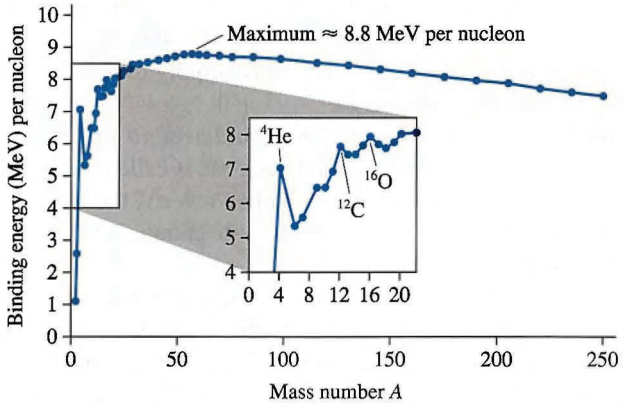


FIGURE 42.6 The curve of binding energy.

Figure 42.6 is a graph of the binding energy per nucleon versus mass number A. The dotted line connecting the points is often called the **curve of binding energy.** This curve has several important features:

- There are peaks in the binding energy curve at A = 4, 12, and 16. The one at A = 4, corresponding to ^{4}He, is especially pronounced. As you'll see, these peaks, which represent nuclei more tightly bound than their neighbors, are due to *closed shells* in much the same way that the graph of atomic ionization energies (see Figure 41.24) peaked for closed electron shells.
- The binding energy per nucleon is *roughly* constant at ≈8 MeV per nucleon for $A > 20$. This suggests that, as a nucleus grows, there comes a point where the nuclear bonds are *saturated*. Each nucleon interacts only with its nearest neighbors, the ones it's actually touching. This, in turn, implies that the nuclear force is a *short-range* force.

- The curve has a broad maximum at $A \approx 60$. This will be important for our understanding of radioactivity. In principle, heavier nuclei could become *more* stable (more binding energy per nucleon) by breaking into smaller pieces. Lighter nuclei could become *more* stable by fusing together into larger nuclei. There may not always be a mechanism for such nuclear transformations to take place, but *if* there is a mechanism, it is energetically favorable for it to occur.

42.3 The Strong Force

Rutherford's discovery of the atomic nucleus was not immediately accepted by all scientists. Their primary objection was that the protons would blow themselves apart at tremendously high speeds due to the extremely large electrostatic forces between them at a separation of a few femtometers. No known force could hold the nucleus together.

It soon became clear that a previously unknown force of nature operates within the nucleus to hold the nucleons together. This new force had to be stronger than the repulsive electrostatic force, hence it was named the **strong force.** It is also called the *nuclear force.*

The strong force has several important properties:

1. It is an *attractive* force between any two nucleons.
2. It does not act on electrons.
3. It is a *short-range* force, acting only over nuclear distances.
4. Over the range where it acts, it is *stronger* than the electrostatic force that tries to push two protons apart.

The fact that the strong force is short-range, in contrast to the long-range $1/r^2$ electric, magnetic, and gravitational forces, is apparent from the fact that we see no evidence for nuclear forces outside the nucleus.

Figure 42.7 summarizes the three interactions that take place within the nucleus. Whether the strong force between two protons is the same strength as the force between two neutrons or between a proton and a neutron is an important question that can be answered experimentally. The primary means of investigating the strong force is to accelerate a proton to very high velocity, using a cyclotron or some other particle accelerator, then to study how the proton is scattered by various target materials.

The conclusion of many decades of research is that the strong force between two nucleons is independent of whether they are protons or neutrons. Charge is the basis for electromagnetic interactions, but it is of no relevance to the strong force. Protons and neutrons are identical as far as nuclear forces are concerned.

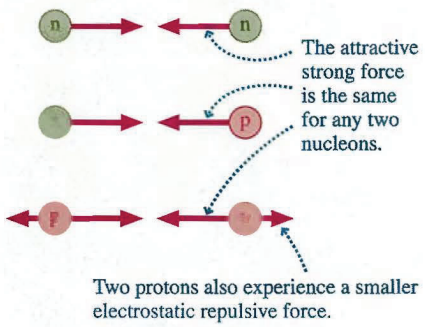


FIGURE 42.7 The strong force is the same between any two nucleons.

Potential Energy

Unfortunately, there's no simple formula to calculate the strong force or the potential energy of two nucleons interacting via the strong force. Figure 42.8 is an experimentally determined potential-energy diagram for two interacting nucleons, with r the distance between their centers. The potential-energy minimum at $r \approx 1$ fm is a point of stable equilibrium.

Recall that the force is the negative of the slope of a potential-energy diagram. The steeply rising potential for $r < 1$ fm represents a strongly repulsive force. That is, the nucleon "cores" strongly repel each other if they get too close together. The force is attractive for $r > 1$ fm, where the slope is positive, and it is strongest where the slope is steepest, at $r \approx 1.5$ fm. Notice that the strong force quickly decreases for $r > 1.5$ fm and is zero for $r > 3$ fm. That is, the strong force represented by this potential energy is effective only over a very short range of distances.

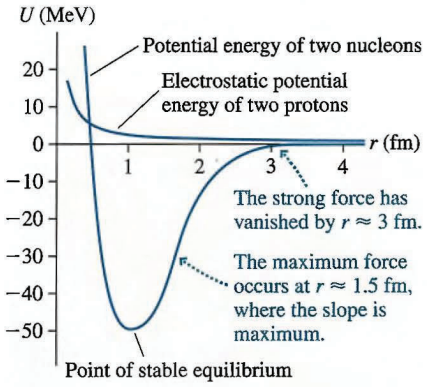


FIGURE 42.8 The potential-energy diagram for two nucleons interacting via the strong force.