

**Key Terms**

valence electrons  
Lewis diagram  
electron pairs  
unpaired electrons  
isotopes  
radioisotopes  
atomic mass unit  
isotopic abundance

**Figure 1.1** Between about 500 and 1700 C.E., alchemists such as these combined practical chemistry with an interest in philosophy and mysticism. Some of the great scientists of history, including Robert Boyle and Sir Isaac Newton, either practised alchemy or viewed it favourably.

For thousands of years, people have discovered and used matter without understanding the fundamental nature of that matter. For example, elements such as copper, gold, iron, sulfur, and carbon were known to and used by people between 4000 and 11 000 years ago. Before the science of chemistry as we know it today was established, alchemists such as the people shown in **Figure 1.1** added to the known elements. Relying on their keen observation skills, as well as equipment that they often invented, alchemists in the Middle East and Europe discovered arsenic, antimony, bismuth, and phosphorus. And still, the nature of the matter of which these elements are made remained unknown.

**Imagining the Atom**


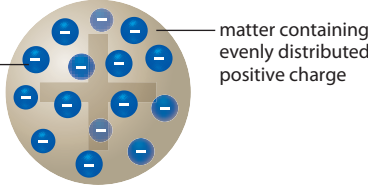
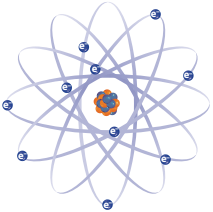
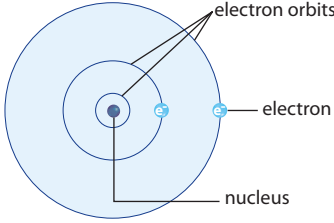

The first people to ask questions and record their ideas about the nature of matter were philosophers who lived and taught in the region of the ancient Mediterranean and northern Africa. About 2500 years ago, some of these professional thinkers imagined the idea of a fundamental building-block of matter. They reasoned that if you cut an object in half, and kept doing so, you would reach a point at which the object could no longer be cut. This speck of object was called *atomos*, meaning “uncuttable.”

The idea of *atomos*—atoms—persisted among some thinkers. However, one of the most influential thinkers of ancient Greece, Aristotle, argued against it. His rejection of atoms was based on philosophical arguments. Neither he nor most other Greek philosophers conducted experiments to demonstrate their ideas. However, Aristotle’s ideas were greatly respected by people of authority both in and outside of Greece. As a result, few people thought about atoms or the building-blocks of matter in a way we recognize as scientific for another 2000 years.

**Modelling the Atom**

The first practical model of the atom was developed in the early 1800s by John Dalton. While studying the properties of atmospheric gases, Dalton inferred that they consisted of tiny particles or atoms. Chemists continued to develop the concept of atoms. They performed numerous experiments and gathered an abundance of data, and they modified the model of the atom to include new discoveries. **Table 1.1** summarizes the most important models of the atom and characteristics of the atom that the models describe. Notice how each model is built on the foundation of the models that came before it.

**Table 1.1** Models of the Atom

Model	Development of Model
 <p>Dalton's model, or the billiard ball model, of the atom was a solid sphere. Atoms of different elements differed from one another, as implied by the different sizes and colours of these two "atoms."</p>	<p>In 1803, English chemist John Dalton (1766–1844) proposed that all matter consisted of tiny particles called atoms. He proposed that atoms of each element were unique and unlike atoms of any other element. These concepts, illustrated by his model of the atom, are still accepted today. However, Dalton also proposed that atoms were indivisible, or unbreakable. J.J. Thomson was the first person to show that this idea is incorrect.</p>
 <p>Thomson's model of the atom is called the "plum pudding" model. It consisted of a positively charged sphere, with negatively charged particles, later called electrons, embedded in it.</p>	<p>In 1897, English physicist Joseph J. Thomson (1856–1940) used cathode ray tubes to demonstrate that atoms could be broken down into smaller particles. He showed that negatively charged particles, later called electrons, could be ejected from atoms, leaving the atom positively charged. His "plum pudding" model of the atom was consistent with all of the characteristics of the atom that were known at the time.</p>
 <p>Rutherford's model of the atom is sometimes called the planetary model. Planet-like electrons orbited a positively charged nucleus, which was analogous to the Sun. The nucleus contained most of the mass of the atom.</p>	<p>In 1911, New Zealand physicist and chemist Ernest Rutherford (1871–1937) directed highly energetic, positively charged alpha particles at a very thin gold foil. He traced the paths of the alpha particles after they collided with the gold foil. Using mathematical calculations, he showed that the only configuration of the atom that could explain the paths of the alpha particles was a configuration that had all of the positive charge and nearly all of the mass located in a tiny space at the centre. He inferred that the negatively charged electrons orbited the small nucleus. His experimental data focussed on the nucleus and did not show where the electrons were located, nor how they were moving.</p>
 <p>Bohr's model of the atom was similar to Rutherford's model. However, Bohr found that electrons can have only specific amounts of energy. His energy levels are sometimes called electron shells.</p>	<p>In 1913, Danish physicist Niels Bohr (1885–1962) refined Rutherford's model of the atom. Bohr analyzed the pattern of colours of light emitted from heated hydrogen atoms. Using those patterns of light and mathematical calculations, Bohr realized that the light was emitted when electrons dropped from a "higher orbit" to a "lower orbit." He showed that electrons can exist only in specific energy levels. He calculated orbital radii that correspond to these energy levels. Although Bohr's calculations apply to only hydrogen, the general concepts can be applied to all atoms.</p>
 <p>Schrödinger showed that electrons move in a region of space, which is often represented visually as a cloud.</p>	<p>In 1926, Austrian physicist Erwin Schrödinger (1887–1961) published a mathematical equation, called the Schrödinger wave equation, that describes the atom in terms of energy. Solutions to this equation show that electrons do not travel in precise orbits but instead exist in defined regions of space. Electrons in each energy level are confined to specific regions, which are represented as electron clouds. This model is usually called the electron cloud model of the atom.</p> <p>Schrödinger's wave equation, along with contributions by Werner Heisenberg (1901–1976), Paul Dirac (1902–1984), Wolfgang Pauli (1900–1958), Friedrich Hund (1896–1997), and others, are explored in advanced chemistry courses.</p>

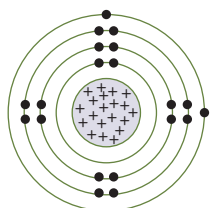
## The Electrons of the Atom

Why was it so important for scientists to develop a detailed model of the atom? An accurate model, or representation, of the atom makes chemistry concepts easier to understand, communicate, and study. In fact, to gain the best understanding of the atom, it is often necessary to rely on more than one model. As shown in **Table 1.1**, the current, most accurate representation of the atom is the electron cloud model, shown in **Figure 1.2**. This model shows the atom as a very small, positively charged nucleus surrounded by clouds of negatively charged electrons. However, this model does not provide any obvious visual information about energy levels. Therefore, chemists often use the model shown in **Figure 1.3** to show energy levels. Because Bohr and Rutherford both contributed to this model, it is often called the Bohr-Rutherford model. Each circle in the Bohr-Rutherford model represents a different energy level, or electron shell. When using this model, it is important to keep in mind that the circles represent energy levels, not actual electron orbits.

**Figure 1.2 (A)** In the electron cloud model of the atom, electron clouds have different shapes. **(B)** Even when these electron clouds are superimposed to represent a complete atom, their appearance does not give any information about their energy levels.



The Bohr-Rutherford model and the electron cloud model look very different. To develop his model, Bohr used the concept of forces of attraction between positive and negative charges. This method allowed him to calculate the radius of the electron shells around the nucleus of a hydrogen atom. Schrödinger, on the other hand, used the energy of electrons to calculate the shape of the electron clouds around the nucleus. While the electron cloud model of the atom does not show any easily visible energy levels, Schrödinger's equation did allow him to find the average distance of electrons from the nucleus. You can think of this as the distance from the nucleus at which electrons spend most of their time. Schrödinger's average distance for an electron in hydrogen is the same as the radius of the electron shell that Bohr calculated, demonstrating how closely their models are related. While the Bohr-Rutherford model provides useful visual information, Schrödinger's equation provides more detailed information about the atom.



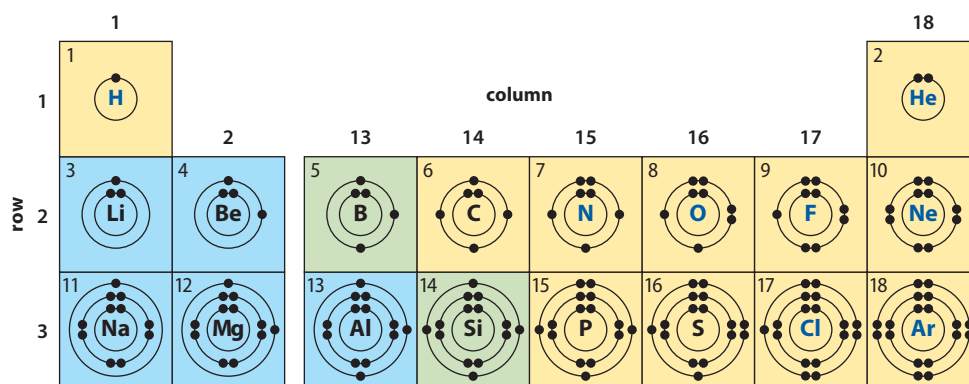
**Figure 1.3** The Bohr-Rutherford model of the atom visually depicts energy levels by showing the number of electrons in each electron shell. The shell closest to the nucleus is the lowest energy level.

**valence electrons**  
electrons in the  
outermost shell of  
an atom

### The Number of Electrons in Each Energy Level

Bohr's work provided information about the energy levels at which electrons can exist. However, further mathematical analysis was required to determine how many electrons can exist in each energy level. Consider the partial periodic table in **Figure 1.4**. The maximum number of electrons that can occupy a shell can be calculated by using  $2n^2$ , where  $n$  is the number of the shell. For example, the number of electrons allowed in the first shell is two ( $2 \times 1^2 = 2 \times 1 = 2$ ), in the second shell is eight ( $2 \times 2^2 = 2 \times 4 = 8$ ), in the third shell is 18 ( $2 \times 3^2 = 2 \times 9 = 18$ ), and so on. When a shell contains the maximum number of electrons, it is said to be filled.

You might expect that the third row of the periodic table would have 18 elements, with 18 electrons in the third shell. Although it is possible for the third shell to contain 18 electrons, elements in the third row do not have more than eight electrons in the outer shell, for reasons that will be explained in advanced chemistry courses. Electrons in the outer shell of an atom are called **valence electrons** and the shell is called the *valence shell*. Note that valence electrons are the only electrons that are involved in the formation of chemical bonds between atoms.



**Figure 1.4** This partial periodic table shows Bohr-Rutherford diagrams of the first 18 elements.

### Lewis Diagrams and Electrons

To understand the structures of chemical compounds and how electrons are involved in the formation of chemical compounds, it is convenient to use simplified models to represent individual atoms. These models are called **Lewis diagrams**, or sometimes electron dot diagrams. Instead of having plus signs to represent positive charges in the nucleus of an atom, a Lewis diagram has the chemical symbol for the element. As shown in **Figure 1.5**, dots are placed around the chemical symbol to represent the electrons in the valence shell. Only the valence electrons are depicted, since they are the only electrons involved in the formation of bonds between atoms. Lewis diagrams become too complex and are no longer useful, however, if atoms have more than eight electrons in their valence shell. Therefore, Lewis diagrams are used mainly for the elements in columns 1, 2, and 13 through 18 of the periodic table.

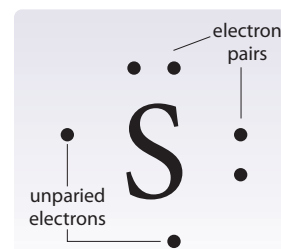


**Figure 1.5** Lewis diagrams are quick and easy to use, and they provide important information about the chemical nature of elements.

When drawing the first four dots (electrons) around the symbol for the element in a Lewis diagram, you usually place the first dot at the top of the symbol. You then go clockwise to add the other three dots, spacing the dots equal distances apart. For atoms with five to eight electrons in the valence shell, you place the fifth dot beside the first dot. Additional dots are placed beside the other three dots that are already in place, as shown in **Figure 1.5**.

Pairing dots when drawing a Lewis diagram for an atom with more than four valence electrons is not simply a convenient way to draw the Lewis diagram—it has significance. The first four electrons, being negatively charged, repel one another and remain as far apart as possible in the atom. However, when more than four electrons occupy the same shell, electrons form pairs that interact in a unique way that allows them to be situated closer together. These **electron pairs**, as they are called, are less likely to participate in chemical bond formation than the **unpaired electrons** are (see **Figure 1.6**).

**Lewis diagram** a model of an atom that has the chemical symbol for the element surrounded by dots to represent the valence electrons of the element



**Figure 1.6** This Lewis diagram of sulfur shows that atoms of sulfur have two electron pairs and two unpaired electrons in the valence shell.

**electron pairs** two electrons that are interacting in a unique way, allowing them to be situated close to each other

**unpaired electrons** electrons in an unfilled outer shell that are not part of a pair and are, therefore, more likely to participate in bonds with other atoms

## Learning Check

1. Describe two significant differences between the Thomson model of the atom and the Rutherford model.
2. Draw Lewis diagrams of lithium, carbon, and fluorine.
3. What information shows that the Bohr model of the atom and the electron cloud model are related?
4. Models are used in many areas of science. Why do you think that models are a very important tool for studying the atom?
5. Describe one feature of Dalton's model, the Thomson model, and the Bohr model that correctly describes one or more properties of the atom.
6. How many electrons could the eighth shell in an atom contain?

## The Nucleus of the Atom

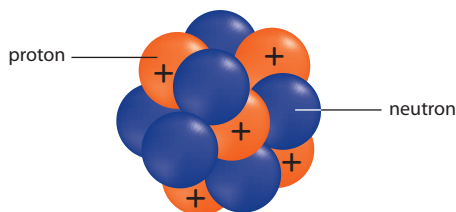
Rutherford not only discovered that the nucleus of an atom occupies a very tiny volume at the centre of the atom. He also discovered many important characteristics of the components that make up the nucleus. The results of his research are discussed below.

### Protons and Neutrons

After Thomson discovered the electron, chemists and physicists realized that an atom had to contain enough positive charge to balance the number of electrons in the atom. However, they did not know the nature of the matter that carried the positive charge.

When Rutherford was doing experiments with alpha particles, one of the atoms that he bombarded was nitrogen. During one experiment, nitrogen nuclei emitted positively charged particles. Rutherford concluded that the nucleus consisted of individual particles, each with a single positive charge that is equal in strength to the negative charge of an electron. Since hydrogen is the smallest atom, and thus has the smallest nucleus, Rutherford proposed that the positively charged particles in all nuclei are nearly identical to the nucleus of a hydrogen atom. He called these positively charged particles protons. He then calculated the total mass of the number of protons that an atom should have in order to balance the negative charge of the electrons. He discovered that this mass could account for only about half of the actual mass of a typical nucleus.

Rutherford first suggested that a nucleus was composed of enough protons to make up the mass of the nucleus and enough electrons to neutralize the excess protons. As he continued his research, however, he realized that there was no evidence to show that there were electrons in the nucleus. He proposed that the excess mass in the nucleus was due to a new particle, which he called a neutron. He proposed that a neutron had the same mass as a proton but no charge. About 10 years later, Rutherford's assistant, James Chadwick (1891–1974), performed experiments that established the existence of the neutron. The current theory of the atom supports Rutherford's proposal that the nucleus of an atom consists of protons and neutrons. **Figure 1.7** shows how a nucleus is often illustrated.



**Figure 1.7** Nuclei of atoms are often depicted as a cluster of spheres in different colours, with one colour representing protons and the other representing neutrons.

## Neutrons and Isotopes

The chemical nature of an element depends on its atomic number: the number of protons in the nucleus. The atomic number also provides the number of electrons in a neutral atom—an atom in which the positive and negative charges are balanced. Unlike the number of protons, the number of neutrons in the nucleus of atoms of the same element can vary. For example, a hydrogen atom can have zero, one, or two neutrons. Atoms of the same element that have different numbers of neutrons are called **isotopes**.

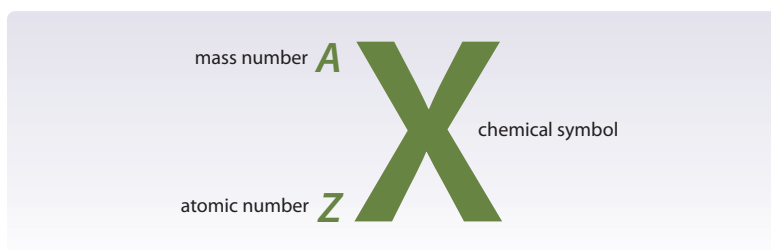
**isotopes** atoms that have the same number of protons but different numbers of neutrons

## Isotope Notation with Mass Number and Atomic Number

The isotope notation in **Figure 1.8** indicates not only the chemical symbol of the element, but also the isotope of the element. In this notation, the following symbols are used:

- the symbol  $Z$  for atomic number, which represents the number of protons in the atom
- the symbol  $A$  for mass number, which represents the sum of the number of protons and the number of neutrons

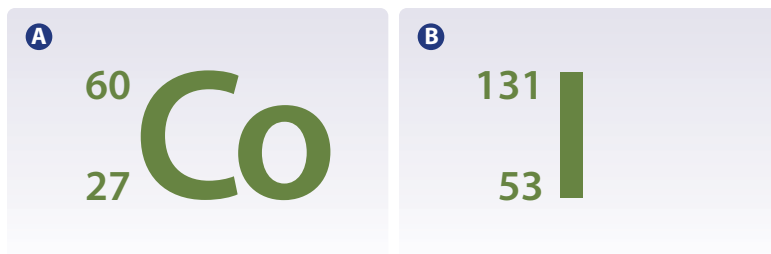
To find the number of neutrons,  $N$ , subtract the atomic number from the mass number:  $N = A - Z$ .



**Figure 1.8** The symbol  $A$  represents the mass number of the atom, which is the sum of the number of neutrons and the number of protons. The symbol  $Z$  is the atomic number, which is the number of protons. In this figure,  $X$  represents any chemical symbol.

Using the isotope notation, you can distinguish among the different isotopes of an element. For example, the isotope of hydrogen with no neutrons is represented as  ${}^1_1\text{H}$ . Similarly, the isotope of hydrogen with one neutron is represented as  ${}^2_1\text{H}$ , whereas the isotope with one proton and two neutrons is represented as  ${}^3_1\text{H}$ . The isotope  ${}^2_1\text{H}$  is sometimes called deuterium, or heavy hydrogen, and the isotope  ${}^3_1\text{H}$  is sometimes called tritium. Isotopes of elements other than hydrogen do not have special names, like deuterium and tritium.

The example in **Figure 1.9 (A)** is read as “cobalt-60.” Cobalt-60 has 27 protons and  $60 - 27 = 33$  neutrons. The example in **Figure 1.9 (B)** is read as “iodine-131.” It has 53 protons and  $131 - 53 = 78$  neutrons.



**Figure 1.9 (A)** Cobalt-60 and **(B)** iodine-131 are radioactive isotopes that are often used in medicine. Cobalt-60 is used in radiation therapy for cancer, and iodine-131 is sometimes used to treat an overactive thyroid.

### The Ratio of Neutrons to Protons

Neutrons play an important role in the nucleus. While this role is not immediately obvious, studying examples of known isotopes of different elements, shown in **Table 1.2**, can offer some clues.

**Table 1.2** Various Elements and Their Isotopes

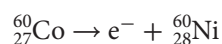
Element	A (protons + neutrons)	Z (protons)	N (neutrons)
hydrogen	1	1	0
	2	1	1
	3	1	2
chlorine	35	17	18
	36	17	19
	37	17	20
molybdenum	95	42	53
	96	42	54
	97	42	55
neodymium	142	60	82
	143	60	83
	145	60	85
	146	60	86
mercury	199	80	119
	200	80	120
	201	80	121
	202	80	122

Notice that atoms of the smaller elements in **Table 1.2** have similar numbers of protons and neutrons. As the atomic number of the element gets larger, the number of neutrons exceeds the number of protons by an increasing amount. For instance, molybdenum has roughly 10 more neutrons than protons, neodymium has over 20 more neutrons than protons, and mercury has about 30 more neutrons than protons.

### Radioisotopes and Unbalanced Forces in the Nucleus

Many positively charged protons are packed very closely together in the nucleus. The repulsive electric force among the protons is tremendous. As the number of protons increases, however, the number of neutrons increases more rapidly. This has a stabilizing effect, which can be explained by the fact that protons and neutrons also attract one another. The attractive force among protons and neutrons, called the *strong nuclear force*, is approximately 40 times stronger than the repulsive electric force among the positively charged protons. However, it acts over only very short distances. As neutrons are added to the nucleus, they add to the attractive nuclear force but not to the repulsive electric force. A correct balance between protons and neutrons thus stabilizes the nucleus.

When the neutrons and protons are not properly balanced, the nucleus is unstable and decays into a nucleus that is more stable. Atoms with unstable nuclei are said to be radioactive. They are called **radioisotopes**. For example, cobalt-60 ( ${}^{60}_{27}\text{Co}$ ) is a radioisotope, but cobalt-59 ( ${}^{59}_{27}\text{Co}$ ) is stable. Cobalt-60 emits a negatively charged particle, called a *beta particle*, from its nucleus. Although electrons cannot exist in the nucleus, a neutron in the nucleus can decay into an electron and a proton. The newly formed electron is immediately ejected from the nucleus as a beta particle. As a result, the decayed nucleus has one more proton and one less neutron. For example, cobalt-60 becomes nickel-60, as shown below:



**radioisotopes** isotopes with unstable nuclei that decay into different, often stable, isotopes

## Average Atomic Mass

In a periodic table, the atomic mass of an element is usually given in **atomic mass units** (u). The atomic mass unit is based on the mass of an atom of carbon-12 and is defined as one twelfth of the mass of a carbon-12 atom. Because the masses of all other atoms are compared to the mass of carbon-12, these masses are often called *relative atomic masses*. Some time after the atomic mass unit was defined, scientists were able to determine the actual mass of an atom, in grams. One atomic mass unit is now known to be equal to  $1.66 \times 10^{-24}$  g, or  $1 \text{ u} = 1.66 \times 10^{-24}$  g.

**atomic mass unit** one twelfth of the mass of a carbon-12 atom

Many elements have two or more naturally occurring stable isotopes. In order to determine the atomic mass of an element that has more than one isotope, each having a different mass, you must find the *average atomic mass*. Recall how to find the average of several values, by adding the values and dividing the sum by the number of values. If you want to find the average mass of several items, the process becomes slightly more complicated. For example, assume that you want to find the average mass of 10 marbles. Five marbles have a mass of 4.0 g, three marbles have a mass of 3.0 g, and two marbles have a mass of 2.5 g. You would not simply take the average of 4.0, 3.0, and 2.5, because you must account for the number of marbles with each mass. Instead, you would find the average by using one of the four methods below.

### Methods for Calculating Averages

**Method 1:** Add all the masses, and divide the sum by the total number of marbles.

$$\begin{aligned}\text{average mass} &= \frac{\text{sum of the masses}}{\text{number of marbles}} \\ &= \frac{4.0 \text{ g} + 4.0 \text{ g} + 4.0 \text{ g} + 4.0 \text{ g} + 4.0 \text{ g} + 3.0 \text{ g} + 3.0 \text{ g} + 3.0 \text{ g} + 2.5 \text{ g} + 2.5 \text{ g}}{10} \\ &= \frac{34 \text{ g}}{10} = 3.4 \text{ g}\end{aligned}$$

**Method 2:** Group the masses by size, and divide the sum by the total number of marbles.

$$\begin{aligned}\text{average mass} &= \frac{(5 \times 4.0 \text{ g}) + (3 \times 3.0 \text{ g}) + (2 \times 2.5 \text{ g})}{10} = \frac{20 \text{ g} + 9 \text{ g} + 5 \text{ g}}{10} \\ &= \frac{34 \text{ g}}{10} = 3.4 \text{ g}\end{aligned}$$

**Method 3:** Use a weighted average. You know that five tenths, or one half, of the marbles have a mass of 4.0 g. Similarly, three tenths of the marbles have a mass of 3.0 g, and two tenths, or one fifth, have a mass of 2.5 g. You could multiply each mass by the fraction it contributes to the total and then add the products as shown.

$$\begin{aligned}\text{average mass} &= \left(\frac{1}{2} \times 4.0 \text{ g}\right) + \left(\frac{3}{10} \times 3.0 \text{ g}\right) + \left(\frac{1}{5} \times 2.5 \text{ g}\right) = 2.0 \text{ g} + 0.9 \text{ g} + 0.5 \text{ g} \\ &= 3.4 \text{ g}\end{aligned}$$

**Method 4:** Express the fractions used in method 3 as percentages:  $\frac{1}{2}$  is 50%,  $\frac{3}{10}$  is 30%, and  $\frac{1}{5}$  is 20%. Then convert the percentages into decimal form (for example,  $\frac{30\%}{100\%} = 0.30$ ), and multiply the decimals by their respective masses. Add the resulting values to obtain the average mass.

$$\begin{aligned}\text{average mass} &= (0.50 \times 4.0 \text{ g}) + (0.30 \times 3.0 \text{ g}) + (0.20 \times 2.5 \text{ g}) \\ &= 2.0 \text{ g} + 0.90 \text{ g} + 0.50 \text{ g} = 3.4 \text{ g}\end{aligned}$$

You can use method 4 to find the weighted average of any type or number of items, as long as you know the percentage of the whole that is contributed by each subgroup. You do not need to know the total number of items. This means that you can use method 4 to find the average atomic mass of an element with several naturally occurring isotopes.



**isotopic abundance**  
the amount of a given isotope of an element that exists in nature, expressed as a percentage of the total amount of this element

### Isotopic Abundance

Scientists have analyzed the atomic masses of most of the naturally occurring isotopes on Earth. They have determined the percentage contributed by the different isotopes of each element in a large number of samples. These percentages for any given element are so similar to one another that, for most elements, scientists are confident that they represent the percentages for all samples on Earth. Thus, scientists have established a standard percentage for each isotope of an element, which is often called its **isotopic abundance**.

The following Sample Problem demonstrates how to calculate the average mass of the atoms of an element, and the Practice Problems that follow give you an opportunity to practise your skills. The Activity that follows will give you a chance to apply the methods that you have learned.

## Sample Problem

### Calculating Average Atomic Mass

#### Problem

The table below provides the atomic mass of each naturally occurring isotope of copper and the percentage of each isotope in a sample of copper. What is the average atomic mass of copper?

#### Mass and Isotopic Abundance of Each Isotope of Copper

Isotope	Mass (u)	Isotopic Abundance (%)
copper-63	62.93	69.2
copper-65	64.93	30.8

#### What Is Required?

The problem asks for the average atomic mass of copper.

#### What Is Given?

You know the mass and isotopic abundance of copper-63:

mass = 62.93 u and isotopic abundance = 69.2%

You know the mass and isotopic abundance of copper-65:

mass = 64.93 u and isotopic abundance = 30.8%

Plan Your Strategy	Act on Your Strategy
Multiply the mass of each isotope by its isotopic abundance, expressed as a decimal, to determine the contribution of each isotope to the average atomic mass.	contribution of isotope copper-63 = $62.93 \text{ u} \times 0.692$ = 43.5476 u contribution of isotope copper-65 = $64.93 \text{ u} \times 0.308$ = 19.9984 u
Add the contributions of the isotopes to determine the average atomic mass of the element.	average atomic mass of Cu = $43.5476 \text{ u} + 19.9984$ = 63.546 u = 63.5 u

#### Alternative Solution

The calculation can be combined into one step, as follows:

$$\begin{aligned}\text{average atomic mass of Cu} &= 62.93 \text{ u} \times 0.692 + 64.93 \text{ u} \times 0.308 \\ &= 63.5 \text{ u}\end{aligned}$$

#### Check Your Solution

The calculated average atomic mass is between the atomic masses of the isotopes, but closer to the atomic mass of the isotope that has the larger isotopic abundance. Three significant digits are appropriate, based on the number of significant digits in the isotopic abundances.

## Practice Problems

- Chlorine exists naturally as 75.78% chlorine-35 (mass = 34.97 u) and 24.22% chlorine-37 (mass = 36.97 u). What is the average atomic mass of chlorine?
- Boron exists naturally as boron-10 (mass = 10.01 u; isotopic abundance = 19.8%) and boron-11 (mass = 11.01 u; isotopic abundance = 80.2%). What is the average atomic mass of boron?
- Lithium is composed of 7.59% lithium-6 (mass = 6.02 u) and 92.41% lithium-7 (mass = 7.02 u). Calculate the average atomic mass of lithium.
- Magnesium exists naturally as 78.99% magnesium-24 (mass = 23.99 u), 10.00% magnesium-25 (mass = 24.99 u), and 11.01% magnesium-26 (mass = 25.98 u). What is the average atomic mass of magnesium?
- Gallium exists naturally as gallium-69 (mass = 68.93 u; isotopic abundance = 60.1%) and gallium-71 (mass = 70.92 u). What is the isotopic abundance of gallium-71? What is the average atomic mass of gallium?
- Bromine exists naturally as bromine-79 (mass = 78.92 u; isotopic abundance = 50.69%) and bromine-81 (mass = 80.92 u). What is the isotopic abundance of bromine-81? What is the average atomic mass of bromine?
- A sample of rubidium is 72.17% rubidium-85 (mass = 84.91 u) and 27.83% rubidium-87 (mass = 86.91 u). Calculate the average atomic mass of rubidium.
- The average atomic mass of nitrogen is 14.01 u. Nitrogen exists naturally as nitrogen-14 (mass = 14.00 u) and nitrogen-15 (mass = 15.00 u). What can you infer about the isotopic abundances for nitrogen?
- The following isotopes of rhenium are found in nature.  
rhenium-185:  $m = 184.953$  u, abundance = 37.4%  
rhenium-187:  $m = 186.956$  u, abundance = 62.6%  
Analyze the data. Predict the average atomic mass and write down your prediction. Then calculate the average atomic mass and compare it with your prediction.
- The average atomic mass of iridium is 192.22 u. If iridium-191 has an atomic mass of 190.961 u and an isotopic abundance of 37.3%, and iridium-193 is the only other naturally occurring isotope, what is the atomic mass of iridium-193?

## Activity

### 1.1 Penny Isotopes

All Canadian pennies have a monetary value of one cent. However, not all pennies are alike. Many have different masses, because the Canadian mint has changed the composition of pennies several times. So, you can think of different pennies as “isotopes” of the penny. In this activity, you will determine the isotopic abundances of the penny isotopes in a sample, and the average mass of the penny.

#### Safety Precaution

Always wash your hands after handling money.

#### Materials

- sample of at least 25 Canadian pennies
- balance
- calculator

#### Procedure

- Sort all your pennies into the following groups of “isotopes.”
  - penny-1: 2000 to the present
  - penny-2: 1997 to 1999
  - penny-3: 1980 to 1996
  - penny-4: 1979 and earlier

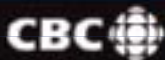
- Count and record the number of pennies in each group.
- Using the balance, determine the mass of one penny in each group. Record the mass of each penny isotope to the nearest 0.01 g. **Hint:** The final mass will be more accurate if you measure the total mass of all the pennies in each group and then divide the total mass by the number of pennies in the group.
- Calculate the “isotopic abundance” of each “isotope.” You can do this by using the following formula:  
$$\text{isotopic abundance} = \frac{\text{number of pennies in category}}{\text{total number of pennies in sample}} \times 100\%$$
- Calculate the average mass of the penny.

#### Questions

- Explain why the average mass of the “element” penny is not necessarily the same as the mass of any one “isotope.”
- Compare the average mass you calculated with the masses obtained by other groups in the class. Explain why the masses are not all the same. Why is the average atomic mass of most elements the same in every sample that has been analyzed?
- Describe how the penny model is a valid model of isotopes and how it is not a valid model of isotopes.

# QUIRKS & QUARKS

with BOB McDONALD



## THIS WEEK ON QUIRKS & QUARKS

### Unearthing an Ancient Andean Element

You may already know that Spain conquered much of South America in the 1500s, plundering its gold and silver mines. You may not know, however, that the Spanish also mined cinnabar—mercury(II) sulfide,  $\text{HgS}(s)$ —, which they valued for its mercury content. Colin Cooke, a PhD student in the Department of Earth and Atmospheric Science at the University of Alberta, studied lake sediment high in the Peruvian Andes of South America, looking for evidence of atmospheric mercury pollution from old Spanish mining operations. Cooke found what he was looking for. However, to his surprise, he also found evidence that people had been mining cinnabar as far back as 1400 B.C.E., long before the Incas settled in the Peruvian Andes. Bob McDonald interviewed Colin Cooke to learn more about the extent of the pollution from ancient cinnabar mines.

Cooke was able to determine when the pollution from mining had occurred by measuring how deeply the pollution was buried below the bottoms of lakes. He found

that the pollution from pre-Incan mines was mainly the cinnabar itself. All the pollution was very near the mines. During the Incan Empire, however, cinnabar was mined for its red pigment, which was used for decoration. When processing the ore, the Incas heated it, causing some of the mercury to vaporize. As a vapour, the mercury travelled long distances carried by the wind. The Incas did not know how toxic the vapour was.

When the Spanish arrived, the mining was intensified. The Incan workers mined the ore with picks and then roasted it, causing most of the mercury to vaporize. Although their intent was to condense the vapour in order to collect the mercury, much of the vapour remained in the air. Inhaled mercury vapour is possibly the most toxic form of mercury. The Incan miners inhaled so much mercury that many died within six months after they started mining. Even today, the mines and the miners' skeletons pose a hazard because the mercury that the miners inhaled now sits in pools in their graves.

### Related Career

Bob McDonald is a science journalist. He reports on science issues for many radio and television programs. Bob came to science from an arts education and a background in theatre. Science journalists interview experts and share what they learn in ways that non-scientists can understand.

Go to [scienceontario](#) to find out more



In ancient times, cinnabar was prized for its intense red colour.

### QUESTIONS

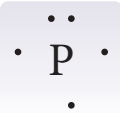
1. What could the Incan miners have done to protect themselves from inhaling the mercury?
2. Today, mercury is used in many ways, such as in batteries. In nature, mercury waste escapes into waterways and accumulates in fish. Considering the seriousness of mercury poisoning, summarize the risks and benefits of mercury use today.
3. As a science journalist, Bob McDonald, has made a career out of talking to scientists. What credentials do science journalists have? What kinds of jobs do they do, and where do they work?

## Section Summary

- As scientists discovered new information about the atom, they modified the model of the atom to reflect each piece of new information.
- The number of electrons that can occupy an electron shell can be determined by using the formula  $2n^2$ , where  $n$  is the shell number.
- The electrons in the outermost shell of an atom are called valence electrons.
- A Lewis diagram is a simplified method for representing an atom.
- The nucleus of an atom contains protons and neutrons and occupies a very small volume at the centre of the atom.
- An appropriate ratio of neutrons to protons stabilizes the nucleus. An atom with an unstable nucleus is called a radioisotope.
- Atomic masses that are reported in data tables are weighted averages, based on isotopic abundances.

## Review Questions

- K/U** Describe two characteristics of an atom that were determined by Schrödinger and that were not revealed by any previous technique or model.
- K/U** In what way does the Bohr-Rutherford model of the atom provide more visual information than the more accurate electron cloud model?
- T/I** Examine the Lewis diagram shown here. List at least three details about the atom that you can determine from this diagram.
 


- C** Draw a diagram showing the general isotope notation using the letters  $Z$ ,  $A$ , and  $X$ . Include in your diagram the meaning of each symbol.
- K/U** Define the term “isotope.”
- K/U** Explain how an atom of one element can turn into an atom of another element. Under what conditions does this occur?
- K/U** Define “isotopic abundance.”
- K/U** Explain why the atomic mass reported in a periodic table is nearly always an average value.
- T/I** Copy and complete the following table.
- T/I** The atomic masses and isotopic abundances of the naturally occurring isotopes of silicon are given below. Calculate the average atomic mass of silicon.
 

$^{28}_{14}\text{Si}$  : mass = 27.977 u; isotopic abundance = 92.23%

$^{29}_{14}\text{Si}$  : mass = 28.976 u; isotopic abundance = 4.67%

$^{30}_{14}\text{Si}$  : mass = 29.97 u; isotopic abundance = 3.10%
- T/I** The atomic mass of yttrium-89 is 88.91 u. The average atomic mass of yttrium that is reported in periodic tables is 88.91. Infer why these values are the same.
- K/U** Write a paragraph that explains why the ratio of neutrons to protons increases as the atomic number increases.
- K/U** You reviewed four methods for calculating average values of a group of items. Which of the methods can you use to calculate average atomic mass? Explain why the other methods cannot be used.
- A** Animals consume carbon-containing compounds and incorporate them into their tissues and breathe out carbon dioxide. A very small amount of that carbon is radioactive carbon-14. When the animals die, they no longer exchange carbon with the environment. Based on this information, how do you think that scientists use the percentage of carbon-14 in a fossil to determine its age? (This technique is called carbon dating.)

## Isotope Data

	Name of Isotope	Notation for Isotope	Atomic Number	Mass Number	Number of Protons	Number of Electrons	Number of Neutrons
a.			35	81			
b.				22	10		
c.	calcium-44						
d.					47		60