Isotopes and Atomic Mass

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CONCEPT 1 Isotopes and Atomic Mass

Lesson Objectives

- Define atomic number.
- Define mass number.
- Understand how isotopes differ from one another and be able to designate them by various methods.
- Be able to calculate the average atomic mass of an element.

Lesson Vocabulary

- atomic mass
- atomic mass unit
- atomic number
- isotope
- mass number
- nuclide

Introduction

Atoms are the fundamental building blocks of all matter and are composed of protons, neutrons, and electrons. Because atoms are electrically neutral, the number of positively charged protons must be equal to the number of negatively charged electrons. One of Dalton's points in his atomic theory was that all atoms of a given element are identical. In this section, we will see how this is not strictly true, thanks to variability in the number of neutrons that an atom may contain.

Atomic Number

The **atomic number (Z)** *of an element is the number of protons in the nucleus of each atom of that element*. An atom can be classified as a particular element based solely on its atomic number. For example, any atom with an atomic number of 8 (its nucleus contains 8 protons) is an oxygen atom, and any atom with a different number of protons would be a different element. The periodic table (**Figure** 1.1) displays all of the known elements and is arranged in order of increasing atomic number. In this table, an element's atomic number is indicated above the elemental symbol. Hydrogen, at the upper left of the table, has an atomic number of 1. Every hydrogen atom has one proton in its nucleus. Next on the table is helium, whose atoms have two protons in the nucleus. Lithium atoms have three protons, beryllium atoms have four, and so on.

Since atoms are neutral, the number of electrons in an atom is equal to the number of protons. Hydrogen atoms all have one electron occupying the space outside of the nucleus.

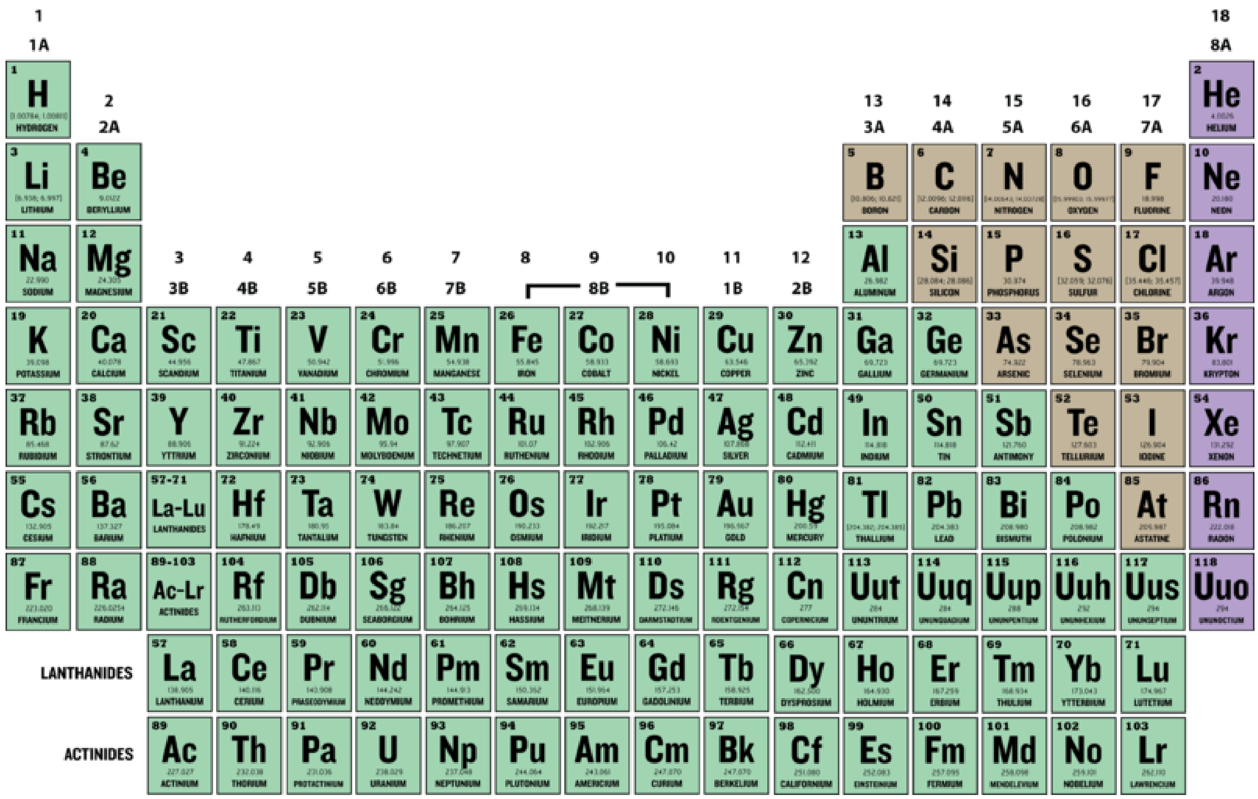

FIGURE 1.1

The periodic table of the elements.

Mass Number

Rutherford's experiment showed that the vast majority of the mass of an atom is concentrated in its nucleus, which is composed of protons and neutrons. The **mass number** *is defined as the total number of protons and neutrons in an atom*. Consider the table below (**Table** 1.1), which shows data from the first six elements of the periodic table.

TABLE 1.1: Atoms of the First Six Elements

Name	Symbol	Atomic Number	Protons	Neutrons	Electrons	Mass Number
Hydrogen	H	1	1	0	1	1
Helium	He	2	2	2	2	4
Lithium	Li	3	3	4	3	7
Beryllium	Be	4	4	5	4	9
Boron	B	5	5	6	5	11
Carbon	C	6	6	6	6	12

Consider the element helium. Its atomic number is 2, so it has two protons in its nucleus. Its nucleus also contains two neutrons. Since 2 + 2 = 4, we know that the mass number of the helium atom is 4. Finally, the helium atom also contains two electrons, since the number of electrons must equal the number of protons. This example may lead you to believe that atoms have the same number of protons and neutrons, but a further examination of the table above (**Table** 1.1) will show that this is not the case. Lithium, for example has three protons and four neutrons, giving it a mass number of 7.

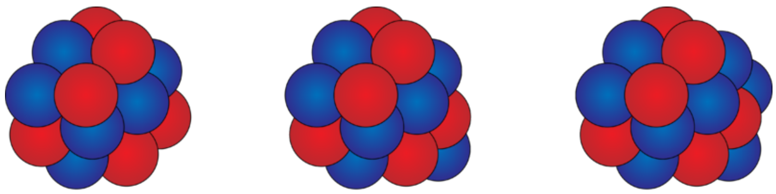

FIGURE 1.2

Nuclei of the three isotopes of carbon: Almost 99% of naturally occurring carbon is carbon-12, whose nucleus consists of six protons and six neutrons. Carbon-13 and carbon-14, with seven or eight neutrons, respectively, have a much lower natural abundance.

Each carbon atom has the same number of protons (6), which is equal to its atomic number. Each carbon atom also contains six electrons, allowing the atom to remain electrically neutral. However the number of neutrons varies from six to eight. **Isotopes** *are atoms that have the same atomic number but different mass numbers due to a change in the number of neutrons*. The three isotopes of carbon can be referred to as carbon-12 ($^{12}_{6}C$), carbon-13 ($^{13}_{6}C$), and carbon-14 ($^{14}_{6}C$). Naturally occurring samples of most elements are mixtures of isotopes. Carbon has only three natural isotopes, but some heavier elements have many more. Tin has ten stable isotopes, which is the most of any element. The term **nuclide** *refers to the nucleus of a given isotope of an element*. The nucleus of a given carbon atom will be one of the three possible nuclides discussed above.

While the presence of isotopes affects the mass of an atom, it does not affect its chemical reactivity. Chemical behavior is governed by the number of electrons and the number of protons. Carbon-13 behaves chemically in exactly the same way as the more plentiful carbon-12.

Sample Problem 4.2: Composition of an Atom

How many protons, neutrons, and electrons are present in each of the nuclides below?

1. Iron (Fe): atomic number = 26, mass number = 56
2. Iodine-127 (atomic number = 53)
3. $^{31}_{15}P$

Step 1: List the known and unknown quantities and plan the problem.

Known

- Atomic number and mass number for each

Unknown

- Number of protons, electrons, and neutrons

Each shows a different way to specify an isotope of an atom. Use the definitions of atomic number and mass number to calculate the numbers of protons, neutrons, and electrons.

Step 2: Calculate.

Number of protons = atomic number

1. 26
2. 53
3. 15

Number of electrons = number of protons

1. 26
2. 53
3. 15

Number of neutrons = mass number - atomic number

1. 56 - 26 = 30
2. 127 - 53 = 74
3. 31 - 15 = 16

Step 3: Think about your result.

For each atom, the results are consistent with the definitions of atomic number and mass number.

Do the practice problems below. If necessary, refer to the periodic table (**Figure** 1.1) for the atomic number or symbol of the given element.

Practice Problems

1. How many protons, neutrons, and electrons are there in the atom $^{19}_{9}F$?
2. How many protons, neutrons, and electrons are there in an atom of lead-207?
3. A certain atom has an atomic number of 36 and a mass number of 84. Write out the designation for this isotope in both nuclide symbol form and in hyphenated form.
4. An atom has a mass number of 59 and contains 32 neutrons in its nucleus. What element is it?

Atomic Mass

The masses of individual atoms are very, very small. However, using a modern device called a mass spectrometer, it is possible to measure such minuscule masses. An atom of oxygen-16, for example, has a mass of 2.66×10^{-23} g. While comparisons of masses measured in grams would have some usefulness, it is far more practical to have a system that will allow us to more easily compare relative atomic masses. Scientists decided on using the carbon-12 nuclide as the reference standard by which all other masses would be compared. By definition, one atom of carbon-12 is assigned a mass of exactly 12 atomic mass units (amu). An **atomic mass unit** *is defined as a mass equal to one twelfth the mass of an atom of carbon-12*. The mass of any isotope of any element is expressed in relation to the carbon-12 standard. For example, one atom of helium-4 has a mass of 4.0026 amu. An atom of sulfur-32 has a mass of 31.972 amu.

The carbon-12 atom has six protons and six neutrons in its nucleus for a mass number of 12. Since the nucleus accounts for nearly all of the mass of the atom, a single proton or single neutron has a mass of approximately 1 amu. However, as seen by the helium and sulfur examples, the masses of individual atoms are not quite whole numbers. This is because an atom's mass is affected very slightly by the interactions of the various particles within the nucleus and also includes the small mass added by each electron.

As stated in the section on isotopes, most elements occur naturally as a mixture of two or more isotopes. Listed below (**Table** 1.2) are the naturally occurring isotopes of several elements along with the percent natural abundance of each.

TABLE 1.2: Atomic Masses and Percent Abundances of Some Natural Isotopes

Element	Isotope (symbol)	Percent natural abundance	Atomic mass (amu)	Average atomic mass (amu)
Hydrogen	$^{1}_{1}H$ $^{2}_{1}H$ $^{3}_{1}H$	99.985 0.015 negligible	1.0078 2.0141 3.0160	1.0079
Carbon	$^{12}_{6}C$ $^{13}_{6}C$ $^{14}_{6}C$	98.89 1.11 trace	12.000 13.003 14.003	12.011
Oxygen	$^{16}_{8}O$ $^{17}_{8}O$ $^{18}_{8}O$	99.759 0.037 0.204	15.995 16.995 17.999	15.999
Chlorine	$^{35}_{17}Cl$ $^{37}_{17}Cl$	75.77 24.23	34.969 36.966	35.453
Copper	$^{63}_{29}Cu$ $^{65}_{29}Cu$	69.17 30.83	62.930 64.928	63.546

For some elements, one particular isotope is much more abundant than any other isotopes. For example, naturally occurring hydrogen is nearly all hydrogen-1, and naturally occurring oxygen is nearly all oxygen-16. For many other elements, however, more than one isotope may exist in substantial quantities. Chlorine (atomic number 17) is a yellowish-green toxic gas. About three quarters of all chlorine atoms have 18 neutrons, giving those atoms a mass number of 35. About one quarter of all chlorine atoms have 20 neutrons, giving those atoms a mass number of 37. Were you to simply calculate the arithmetic average of the precise atomic masses, you would get approximately 36.

$$(34.969 + 36.966)/2 = 35.968 \text{ amu}$$

As you can see, the average atomic mass given in the last column of the table above (**Table** 1.2) is significantly lower. Why? The reason is that we need to take into account the natural abundance percentages of each isotope in order to calculate what is called the weighted average. The **atomic mass** *of an element is the weighted average of the atomic masses of the naturally occurring isotopes of that element*. The sample problem below demonstrates how to calculate the atomic mass of chlorine.

Sample Problem 4.3: Calculating Atomic Mass

Use the atomic masses of each of the two isotopes of chlorine along with their percent abundances to calculate the average atomic mass of chlorine.

Step 1: List the known and unknown quantities and plan the problem.

Known

- chlorine-35: atomic mass = 34.969 amu and % abundance = 75.77%
- chlorine-37: atomic mass = 36.966 amu and % abundance = 24.23%

Unknown

- Average atomic mass of chlorine

Change each percent abundance into decimal form by dividing by 100. Multiply this value by the atomic mass of that isotope. Add together the results for each isotope to get the average atomic mass.

Step 2: Calculate.

chlorine-35	$0.7577 \times 34.969 = 26.50$ amu
chlorine-37	$0.2423 \times 36.966 = 8.957$ amu
average atomic mass	$26.50 + 8.957 = 35.45$ amu

Note: Applying significant figure rules results in the 35.45 amu result without excessive rounding error. In one step:

$$(0.7577 \times 34.969) + (0.2423 \times 36.966) = 35.45 \text{ amu}$$

Step 3: Think about your result.

The calculated average atomic mass is closer to 35 than to 37 because a greater percentage of naturally occurring chlorine atoms have a mass number of 35. It agrees with the value listed in the table above (**Table** 1.2).

Practice Problem

5. The element bromine consists of two naturally occurring isotopes. The isotope with a mass of 78.92 amu has a percent abundance of 50.69%, while the isotope with a mass of 80.92 amu has a percent abundance of 49.31%. Calculate the average atomic mass of bromine.

The atomic masses for each element on the periodic table are average atomic masses. For later calculations involving atomic mass, we will use these values and round each one to four significant figures.

Lesson Summary

- The atomic number of an element is equal to the number of protons in its nucleus.
- The mass number of an element is equal to the sum of the protons and the neutrons in its nucleus.
- Isotopes are atoms of the same element that have a different mass number because of a variation in the number of neutrons.
- The average atomic mass of an element can be calculated from the atomic masses and percent natural abundances of each naturally occurring isotope.

Lesson Review Questions

Recall

1. Why are all atoms electrically neutral?
2. How many protons are in the nucleus of each of the following atoms?
   a. neon
   b. gold
   c. strontium
   d. uranium
3. What part of Dalton's atomic theory is disproved by the existence of isotopes?
4. Which isotope is used as the reference standard for the atomic mass unit?

Apply Concepts

5. The average atomic mass of all naturally occurring lithium atoms is 6.941 amu. The two isotopes of lithium are lithium-6 and lithium-7. Are these isotopes equally common? If not, which is more plentiful in nature, and how do you know?

Think Critically

6. A certain atom contains 28 protons, 28 electrons, and 31 neutrons. Provide the following:
   a. atomic number
   b. mass number
   c. name of element
7. How many protons, neutrons, and electrons are in an atom of cesium-133?
8. Complete **Table** 1.3:

TABLE 1.3: Table for Problem 8

Isotope	Nuclide Symbol	Atomic Number	Mass Number
sodium-23			
	$^{75}_{33}As$		
silver-108			

9. Which of the following is an isotope of $^{40}_{18}Ar$? Explain.
   a. $^{40}_{20}Ca$
   b. $^{39}_{18}Ar$
   c. $^{40}_{18}Ar$
10. Fill in **Table** 1.4:

TABLE 1.4: Table for Problem 10

Isotope	Number of Protons	Number of Electrons	Number of Neutrons	Nuclide Symbol
hydrogen-1				
hydrogen-2				
beryllium-9				
aluminum-27				

11. Fill in **Table** 1.5:

TABLE 1.5: Table for Problem 11

Element	Symbol	Atomic Number	Mass Number	# of Protons	# of Electrons	# of Neutrons	Nuclide Symbol
Nitrogen			14				
	B		11				
		30				35	

TABLE 1.5: (continued)

Element	Symbol	Atomic Number	Mass Number	# of Protons	# of Electrons	# of Neutrons	Nuclide Symbol
					77	116	
							$^{56}_{26}Fe$

12. The element tungsten (W) is known best as a metal that is used as filaments for light bulbs. Naturally occurring tungsten consists of the five isotopes shown below. Calculate the atomic mass of tungsten.

tungsten-180	atomic mass = 179.947 amu	percent abundance = 0.12%
tungsten-182	atomic mass = 181.948 amu	percent abundance = 26.50%
tungsten-183	atomic mass = 182.950 amu	percent abundance = 14.31%
tungsten-184	atomic mass = 183.951 amu	percent abundance = 30.64%
tungsten-186	atomic mass = 185.954 amu	percent abundance = 28.43%

Further Reading / Supplemental Links

There are a lot of websites to help you understand the atom and its history.

Lectures:

- Elements and Atoms: http://www.youtube.com/watch?v=IFKnq9QM6_A
- Introduction to the Atom: http://www.youtube.com/watch?v=1xSQlwWGT8M

Informative websites:

- Atomic History - A Brief Discovery: http://www.pbs.org/wgbh/nova/diamond/insidehistory.html
- All About Atoms: http://education.jlab.org/atomtour/index.html

Simulations:

- Build an Atom: http://phet.colorado.edu/en/simulation/build-an-atom
- Molecular Workbench - Atomic Structure: http://workbench.concord.org/database/activities/47.html
- See Inside a Diamond: http://www.pbs.org/wgbh/nova/diamond/insidewave.html
- Isotopes and Atomic Mass: http://phet.colorado.edu/en/simulation/isotopes-and-atomic-mass
- Atomic Structure: http://freezeray.com/flashFiles/atomicStructure.htm
- Atom Builder: http://freezeray.com/flashFiles/atomBuilder.htm
- Tennis Ball Isotopes: http://www.youtube.com/watch?v=oLnuXpf4hsA

Games:

- Element Math Game: http://education.jlab.org/elementmath/index.html
- Looking For the Top Quark: http://education.jlab.org/topquarkgame/index.html
- Atoms and Matter Crossword Puzzle: http://education.jlab.org/sciencecrossword/atoms_01.html

References

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