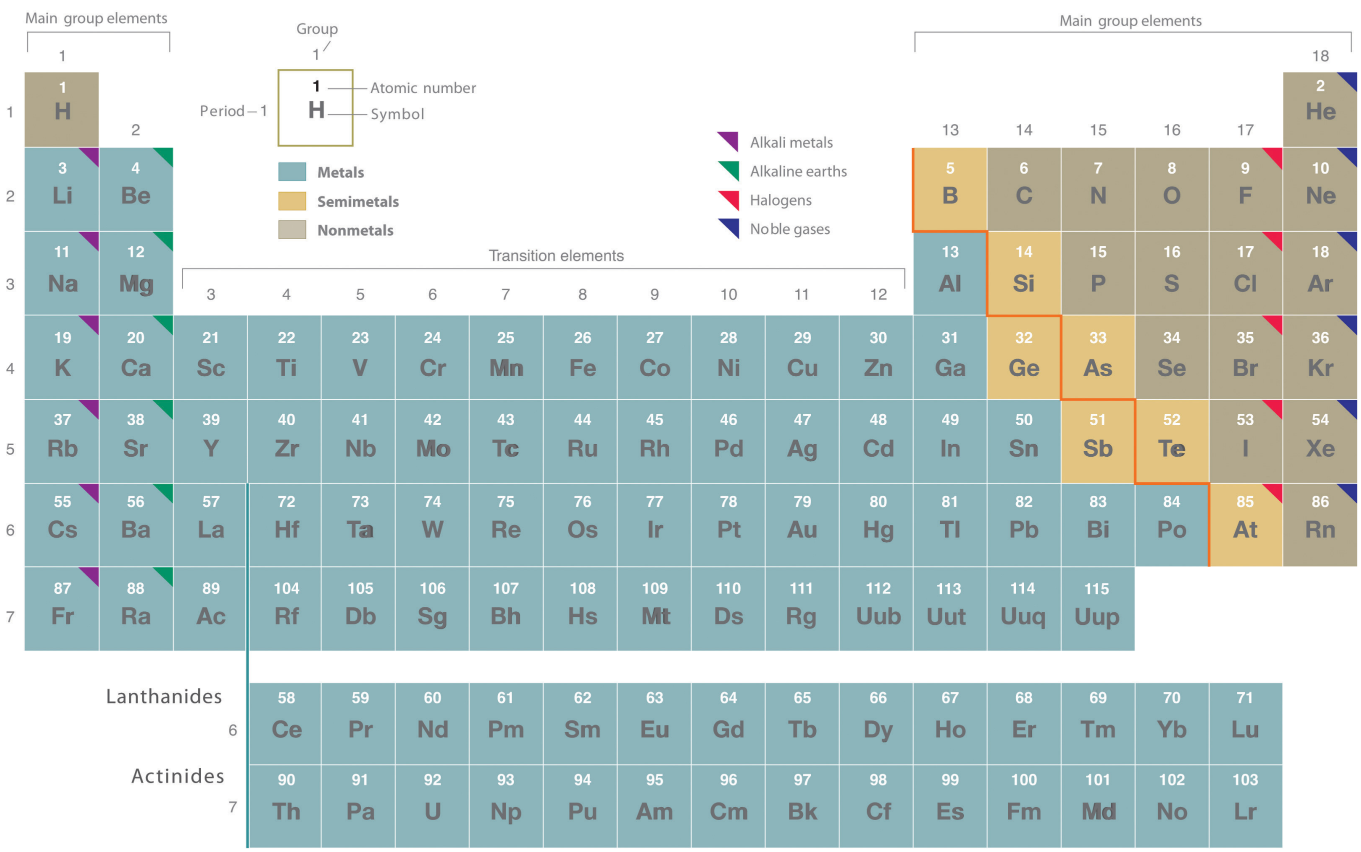
1.7: Isotopes and Atomic Masses

Learning Objectives

- To know the meaning of isotopes and atomic masses.

Rutherford's nuclear model of the atom helped explain why atoms of different elements exhibit different chemical behavior. The identity of an element is defined by its **atomic number (*Z*)**, the number of protons in the nucleus of an atom of the element. *The atomic number is therefore different for each element*. The known elements are arranged in order of increasing *Z* in the **periodic table** (Figure 1.7.1),We will explain the rationale for the peculiar format of the periodic table in Chapter 7. in which each element is assigned a unique one-, two-, or three-letter symbol. The names of the elements are listed in the periodic table, along with their symbols, atomic numbers, and atomic masses. The chemistry of each element is determined by its number of protons and electrons. In a neutral atom, the number of electrons equals the number of protons.

Figure 1.7.1 *The Periodic Table Showing the Elements in Order of Increasing Z*



As described in Section 1.8, the metals are on the bottom left in the periodic table, and the nonmetals are at the top right. The semimetals lie along a diagonal line separating the metals and nonmetals.

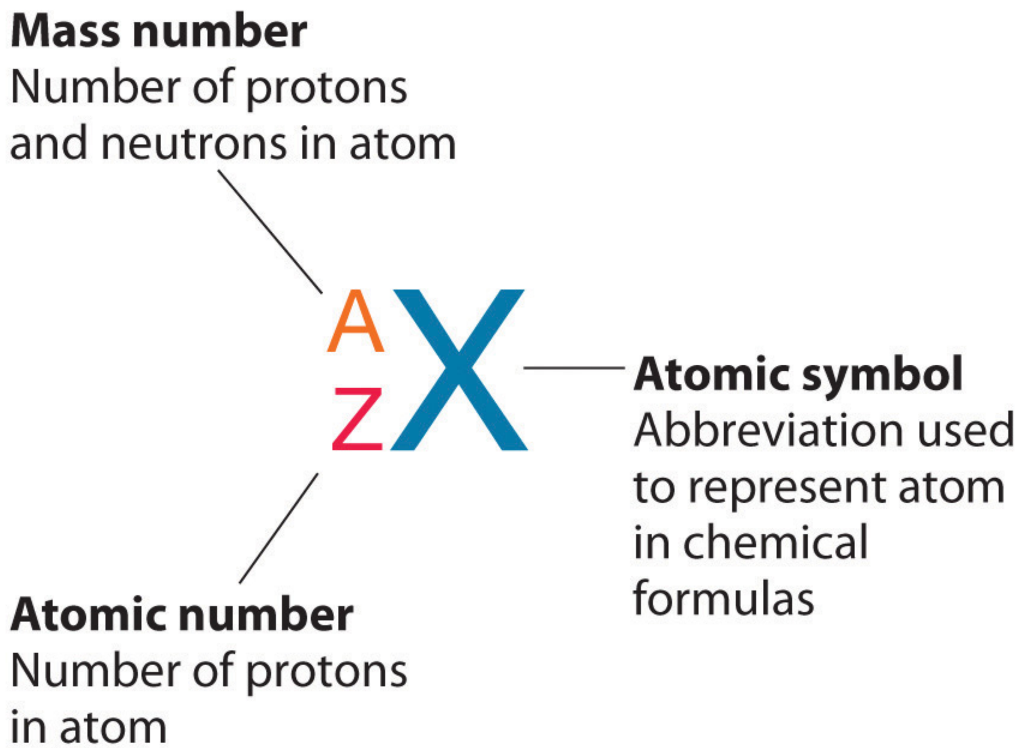
In most cases, the symbols for the elements are derived directly from each element's name, such as C for carbon, U for uranium, Ca for calcium, and Po for polonium. Elements have also been named for their properties [such as radium (Ra) for its radioactivity], for the native country of the scientist(s) who discovered them [polonium (Po) for Poland], for eminent scientists [curium (Cm) for the Curies], for gods and goddesses [selenium (Se) for the Greek goddess of the moon, Selene], and for other poetic or historical reasons. Some of the symbols used for elements that have been known since antiquity are derived from historical names that are no longer in use; only the symbols remain to remind us of their origin. Examples are Fe for iron, from the Latin *ferrum*; Na for sodium, from the Latin *natrium*; and W for tungsten, from the German *wolfram*. Examples are in Table 1.7.1. As you work through this text, you will encounter the names and symbols of the elements repeatedly, and much as you become familiar with characters in a play or a film, their names and symbols will become familiar.

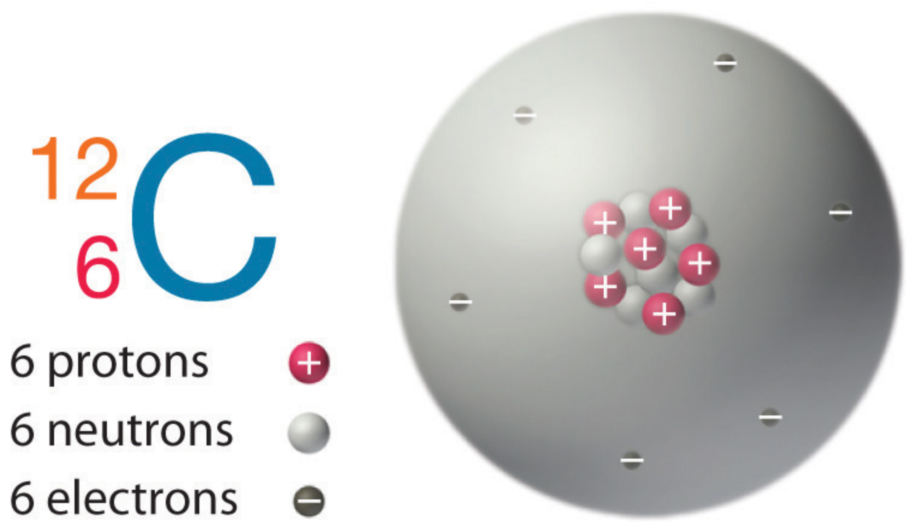
Table 1.7.1 Element Symbols Based on Names No Longer in Use

Element	Symbol	Derivation	Meaning
antimony	Sb	*stibium*	Latin for “mark”
copper	Cu	*cuprum*	from *Cyprium*, Latin name for the island of Cyprus, the major source of copper ore in the Roman Empire
gold	Au	*aurum*	Latin for “gold”
iron	Fe	*ferrum*	Latin for “iron”
lead	Pb	*plumbum*	Latin for “heavy”
mercury	Hg	*hydrargyrum*	Latin for “liquid silver”
potassium	K	*kalium*	from the Arabic *al-qili*, “alkali”
silver	Ag	*argentum*	Latin for “silver”
sodium	Na	*natrium*	Latin for “sodium”
tin	Sn	*stannum*	Latin for “tin”
tungsten	W	*wolfram*	German for “wolf stone” because it interfered with the smelting of tin and was thought to devour the tin

Recall from Section 1.6 that the nuclei of most atoms contain neutrons as well as protons. Unlike protons, the number of neutrons is not absolutely fixed for most elements. Atoms that have *the same number of protons*, and hence the same atomic number, but *different numbers of neutrons* are called **isotopes**. All isotopes of an element have the same number of protons and electrons, which means they exhibit the same chemistry. The isotopes of an element differ only in their atomic mass, which is given by the **mass number (*A*)**, the sum of the numbers of protons and neutrons.

The element carbon (C) has an atomic number of 6, which means that all neutral carbon atoms contain 6 protons and 6 electrons. In a typical sample of carbon-containing material, 98.89% of the carbon atoms also contain 6 neutrons, so each has a mass number of 12. An isotope of any element can be uniquely represented as XZA, where X is the atomic symbol of the element. The isotope of carbon that has 6 neutrons is therefore C612. The subscript indicating the atomic number is actually redundant because the atomic symbol already uniquely specifies *Z*. Consequently, C612 is more often written as ^{12}C, which is read as “carbon-12.” Nevertheless, the value of *Z* is commonly included in the notation for *nuclear* reactions because these reactions involve changes in *Z*, as described in Chapter 20.





In addition to ^{12}C, a typical sample of carbon contains 1.11% C613 (^{13}C), with 7 neutrons and 6 protons, and a trace of C614 (^{14}C), with 8 neutrons and 6 protons. The nucleus of ^{14}C is not stable, however, but undergoes a slow radioactive decay that is the basis of the carbon-14 dating technique used in archaeology (see Chapter 14). Many elements other than carbon have more than one stable isotope; tin, for example, has 10 isotopes. The properties of some common isotopes are in Table 1.7.2.

Table 1.7.2 Properties of Selected Isotopes

Element	Symbol	Atomic Mass (amu)	Isotope Mass Number	Isotope Masses (amu)	Percent Abundances (%)
hydrogen	H	1.0079	1	1.007825	99.9855

Element	Symbol	Atomic Mass (amu)	Isotope Mass Number	Isotope Masses (amu)	Percent Abundances (%)
			2	2.014102	0.0115
boron	B	10.81	10	10.012937	19.91
			11	11.009305	80.09
carbon	C	12.011	12	12 (defined)	99.89
			13	13.003355	1.11
oxygen	O	15.9994	16	15.994915	99.757
			17	16.999132	0.0378
			18	17.999161	0.205
iron	Fe	55.845	54	53.939611	5.82
			56	55.934938	91.66
			57	56.935394	2.19
			58	57.933276	0.33
uranium	U	238.03	234	234.040952	0.0054
			235	235.043930	0.7204
			238	238.050788	99.274

Sources of isotope data: G. Audi et al., *Nuclear Physics A* 729 (2003): 337–676; J. C. Kotz and K. F. Purcell, *Chemistry and Chemical Reactivity*, 2nd ed., 1991.

Example 1.7.1

Given: number of protons and neutrons

Asked for: element and atomic symbol

Strategy:

A Refer to the periodic table (see Chapter 32) and use the number of protons to identify the element.

B Calculate the mass number of each isotope by adding together the numbers of protons and neutrons.

C Give the symbol of each isotope with the mass number as the superscript and the number of protons as the subscript, both written to the left of the symbol of the element.

Solution

A The element with 82 protons (atomic number of 82) is lead: Pb.

B For the first isotope, A = 82 protons + 124 neutrons = 206. Similarly, A = 82 + 125 = 207 and A = 82 + 126 = 208 for the second and third isotopes, respectively. The symbols for these isotopes are P82206b,P82207b, and P82208b, which are usually abbreviated as ^{206}Pb, ^{207}Pb, and ^{208}Pb.

Exercise 1.7.1

Identify the element with 35 protons and write the symbols for its isotopes with 44 and 46 neutrons.

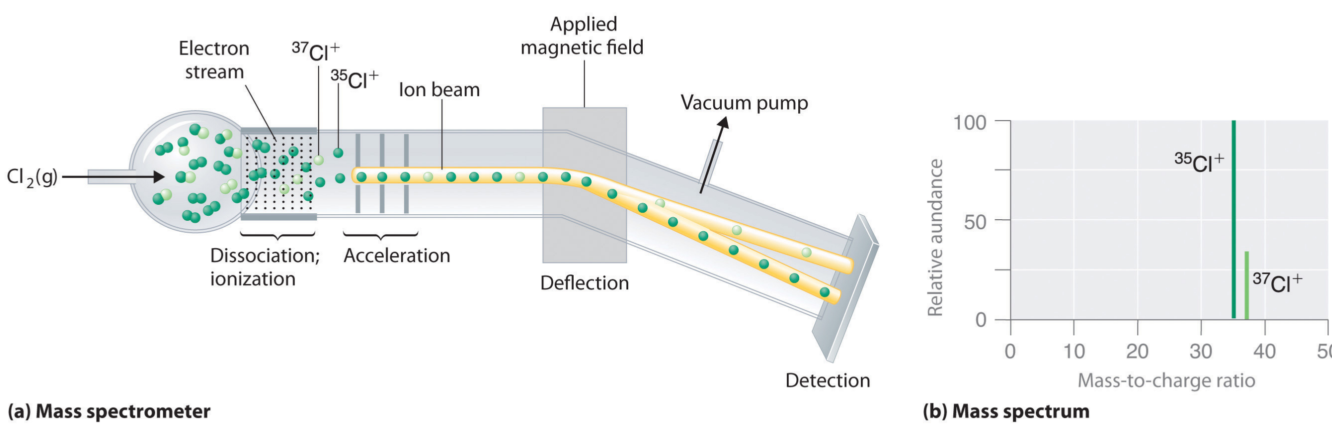
Answer

Answer: $_{35}^{79}$B and $_{3581}$B or, more commonly, ^{79}Br and ^{81}Br.

Although the masses of the electron, the proton, and the neutron are known to a high degree of precision (Table 1.3), the mass of any given atom is not simply the sum of the masses of its electrons, protons, and neutrons. For example, the ratio of the masses of ^{1}H (hydrogen) and ^{2}H (deuterium) is actually 0.500384, rather than 0.49979 as predicted from the numbers of neutrons and protons present. Although the difference in mass is small, it is extremely important because it is the source of the huge amounts of energy released in nuclear reactions (Chapter 20).

Because atoms are much too small to measure individually and do not have a charge, there is no convenient way to accurately measure *absolute* atomic masses. Scientists can measure *relative* atomic masses very accurately, however, using an instrument called a *mass spectrometer*. The technique is conceptually similar to the one Thomson used to determine the mass-to-charge ratio of the electron. First, electrons are removed from or added to atoms or molecules, thus producing charged particles called **ions**. When an electric field is applied, the ions are accelerated into a separate chamber where they are deflected from their initial trajectory by a magnetic field, like the electrons in Thomson's experiment. The extent of the deflection depends on the mass-to-charge ratio of the ion. By measuring the relative deflection of ions that have the same charge, scientists can determine their relative masses (Figure 1.25). Thus it is not possible to calculate absolute atomic masses accurately by simply adding together the masses of the electrons, the protons, and the neutrons, and *absolute* atomic masses cannot be measured, but *relative* masses can be measured very accurately. It is actually rather common in chemistry to encounter a quantity whose magnitude can be measured only relative to some other quantity, rather than absolutely. We will encounter many other examples later in this text. In such cases, chemists usually define a standard by arbitrarily assigning a numerical value to one of the quantities, which allows them to calculate numerical values for the rest.

Figure 1.25 Determining Relative Atomic Masses Using a Mass Spectrometer



Chlorine consists of two isotopes, ^{35}Cl and ^{37}Cl, in approximately a 3:1 ratio. (a) When a sample of elemental chlorine is injected into the mass spectrometer, electrical energy is used to dissociate the Cl_2 molecules into chlorine atoms and convert the chlorine atoms to Cl^+ ions. The ions are then accelerated into a magnetic field. The extent to which the ions are deflected by the magnetic field depends on their relative mass-to-charge ratios. Note that the lighter $^{35}Cl^+$ ions are deflected more than the heavier $^{37}Cl^+$ ions. By measuring the relative deflections of the ions, chemists can determine their mass-to-charge ratios and thus their masses. (b) Each peak in the mass spectrum corresponds to an ion with a particular mass-to-charge ratio. The abundance of the two isotopes can be determined from the heights of the peaks.

The arbitrary standard that has been established for describing atomic mass is the **atomic mass unit (amu)**, defined as one-twelfth of the mass of one atom of ^{12}C. Because the masses of all other atoms are calculated relative to the ^{12}C standard, ^{12}C is the only atom listed in Table 1.7.2 whose exact atomic mass is equal to the mass number. Experiments have shown that 1 amu = 1.66×10^{-24} g.

Mass spectrometric experiments give a value of 0.167842 for the ratio of the mass of ^{2}H to the mass of ^{12}C, so the absolute mass of ^{2}H is

mass of H2mass of C12×mass of C12= 0.167842 × 12 amu = 2.104104 amu

The masses of the other elements are determined in a similar way.

The periodic table lists the atomic masses of all the elements. If you compare these values with those given for some of the isotopes in Table 1.7.2 , you can see that the atomic masses given in the periodic table never correspond exactly to those of any of the isotopes. Because most elements exist as mixtures of several stable isotopes, the **atomic mass** of an element is defined as the weighted average of the masses of the isotopes. For example, naturally occurring carbon is largely a mixture of two isotopes: 98.89% ^{12}C (mass = 12 amu by definition) and 1.11% ^{13}C (mass = 13.003355 amu). The percent abundance of ^{14}C is so low that it can be ignored in this calculation. The *average* atomic mass of carbon is then calculated as

$$(0.9889 \times 12 amu) + (0.0111 \times 13.003355 amu) = 12.01 amu$$

Carbon is predominantly ^{12}C, so its average atomic mass should be close to 12 amu, which is in agreement with our calculation.

The value of 12.01 is shown under the symbol for C in the periodic table , although without the abbreviation *amu*, which is customarily omitted. Thus the tabulated *atomic mass* of carbon or any other element is the weighted average of the masses of the naturally occurring isotopes.

Example 1.7.1

Naturally occurring bromine consists of the two isotopes listed in the following table:

Isotope	Exact Mass (amu)	Percent Abundance (%)
^{79}Br	78.9183	50.69
^{81}Br	80.9163	49.31

Calculate the atomic mass of bromine.

Given: exact mass and percent abundance

Asked for: atomic mass

Strategy:

A Convert the percent abundances to decimal form to obtain the mass fraction of each isotope.

B Multiply the exact mass of each isotope by its corresponding mass fraction (percent abundance ÷ 100) to obtain its weighted mass.

C Add together the weighted masses to obtain the atomic mass of the element.

D Check to make sure that your answer makes sense.

Solution

A The atomic mass is the weighted average of the masses of the isotopes. In general, we can write

atomic mass of element = [(mass of isotope 1 in amu) (mass fraction of isotope 1)] + [(mass of isotope 2) (mass fraction of isotope 2)] + …

Bromine has only two isotopes. Converting the percent abundances to mass fractions gives

B79r: 50.69100=0.5069 B81r: 49.31100=0.4931

B Multiplying the exact mass of each isotope by the corresponding mass fraction gives the isotope's weighted mass:

^{79}Br: 79.9183 amu × 0.5069 = 40.00 amu^{81}Br: 80.9163 amu × 0.4931 = 39.90 amu

C The sum of the weighted masses is the atomic mass of bromine is

40.00 amu + 39.90 amu = 79.90 amu

D This value is about halfway between the masses of the two isotopes, which is expected because the percent abundance of each is approximately 50%.

Exercise 1.7.1

Magnesium has the three isotopes listed in the following table:

Isotope	Exact Mass (amu)	Percent Abundance (%)

Isotope	Exact Mass (amu)	Percent Abundance (%)
^{24}Mg	23.98504	78.70
^{25}Mg	24.98584	10.13
^{26}Mg	25.98259	11.17

Use these data to calculate the atomic mass of magnesium.

Answer

24.31 amu

Summary

Each atom of an element contains the same number of protons, which is the **atomic number** (Z). Neutral atoms have the same number of electrons and protons. Atoms of an element that contain different numbers of neutrons are called **isotopes**. Each isotope of a given element has the same atomic number but a different **mass number** (A), which is the sum of the numbers of protons and neutrons. The relative masses of atoms are reported using the **atomic mass unit** (**amu**), which is defined as one-twelfth of the mass of one atom of carbon-12, with 6 protons, 6 neutrons, and 6 electrons. The **atomic mass** of an element is the weighted average of the masses of the naturally occurring isotopes. When one or more electrons are added to or removed from an atom or molecule, a charged particle called an **ion** is produced, whose charge is indicated by a superscript after the symbol.

KEY TAKEAWAY

- The mass of an atom is a weighted average that is largely determined by the number of its protons and neutrons, whereas the number of protons and electrons determines its charge.

CONCEPTUAL PROBLEMS

1. Complete the following table for the missing elements, symbols, and numbers of electrons.

Element	Symbol	Number of Electrons
molybdenum		
		19
titanium		
	B	
		53
	Sm	
helium		
		14

2. Complete the following table for the missing elements, symbols, and numbers of electrons.

Element	Symbol	Number of Electrons
lanthanum		
	Ir	
aluminum		
		80
sodium		

Element	Symbol	Number of Electrons
	Si	
		9
	Be	

3. Is the mass of an ion the same as the mass of its parent atom? Explain your answer.
4. What isotopic standard is used for determining the mass of an atom?
5. Give the symbol XZA for these elements, all of which exist as a single isotope.
    1. beryllium
    2. ruthenium
    3. phosphorus
    4. aluminum
    5. cesium
    6. praseodymium
    7. cobalt
    8. yttrium
    9. arsenic
6. Give the symbol XZA for these elements, all of which exist as a single isotope.
    1. fluorine
    2. helium
    3. terbium
    4. iodine
    5. gold
    6. scandium
    7. sodium
    8. niobium
    9. manganese
7. Identify each element, represented by X, that have the given symbols.
    1. X2655
    2. X3374
    3. X1224
    4. X53127
    5. X1840
    6. X63152

NUMERICAL PROBLEMS

Please be sure you are familiar with the topics discussed in Essential Skills 1 (Section 1.9) before proceeding to the Numerical Problems.

1. The isotopes ^{131}I and ^{60}Co are commonly used in medicine. Determine the number of neutrons, protons, and electrons in a neutral atom of each.
2. Determine the number of protons, neutrons, and electrons in a neutral atom of each isotope:
    1. ^{97}Tc
    2. ^{113}In
    3. ^{63}Ni
    4. ^{55}Fe
3. Both technetium-97 and americium-240 are produced in nuclear reactors. Determine the number of protons, neutrons, and electrons in the neutral atoms of each.
4. The following isotopes are important in archaeological research. How many protons, neutrons, and electrons does a neutral atom of each contain?

1. ^{207}Pb
2. ^{16}O
3. ^{40}K
4. ^{137}Cs
5. ^{40}Ar

5. Copper, an excellent conductor of heat, has two isotopes: ^{63}Cu and ^{65}Cu. Use the following information to calculate the average atomic mass of copper:

Isotope	Percent Abundance (%)	Atomic Mass (amu)
^{63}Cu	69.09	62.9298
^{65}Cu	30.92	64.9278

6. Silicon consists of three isotopes with the following percent abundances:

Isotope	Percent Abundance (%)	Atomic Mass (amu)
^{28}Si	92.18	27.976926
^{29}Si	4.71	28.976495
^{30}Si	3.12	29.973770

Calculate the average atomic mass of silicon.

7. Complete the following table for neon. The average atomic mass of neon is 20.1797 amu.

Isotope	Percent Abundance (%)	Atomic Mass (amu)
^{20}Ne	90.92	19.99244
^{21}Ne	0.257	20.99395
^{22}Ne		

8. Are X2863 and X2962 isotopes of the same element? Explain your answer.
9. Complete the following table:

Isotope	Number of Protons	Number of Neutrons	Number of Electrons
^{238}X			95
^{238}U			
	75	112	

10. Complete the following table:

Isotope	Number of Protons	Number of Neutrons	Number of Electrons
^{57}Fe			
^{40}X		20	
^{36}S			

11. Using a mass spectrometer, a scientist determined the percent abundances of the isotopes of sulfur to be 95.27% for ^{32}S, 0.51% for ^{33}S, and 4.22% for ^{34}S. Use the atomic mass of sulfur from the periodic table and the following atomic masses to determine whether these data are accurate, assuming that these are the only isotopes of sulfur: 31.972071 amu for ^{32}S, 32.971459 amu for ^{33}S, and 33.967867 amu for ^{34}S.
12. The percent abundances of two of the three isotopes of oxygen are 99.76% for ^{16}O, and 0.204% for ^{18}O. Use the atomic mass of oxygen given in the periodic table and the following data to determine the mass of ^{17}O: 15.994915 amu for ^{16}O and 17.999160

amu for ^{18}O.

13. Which element has the higher proportion by mass in NaI?
14. Which element has the higher proportion by mass in KBr?
