Chapter 2
Atoms and the Periodic Table

Practice Problems C

2.1 (i) ^{14}N, (ii) ^{21}Na, (iii) ^{15}O

2.2 (i) 2.9177 g

(ii) 3.4679 g

(iii) 3.4988 g

2.3 (i) 0.5000 dozen, 4.167×10^{-2} gross; (ii) 1.500 dozen, 0.1250 gross; (iii) 1.250 dozen, 0.1042 gross. These numbers actually have an infinite number of significant figures because they are the result of counting objects.

2.4 (a) 391.2 g/dozen plain, 480 g/dozen jam-filled

(b) 30.7 plain in 1 kg, 25.0 jam-filled in 1 kg

(c) 815 g

(d) 413.4 g

2.5 (a) 672; (b) 576; (c) 6.78 lb; (d) 12.0 lb

Key Skills

2.1 c 2.2 d 2.3 e 2.4 d

Questions and Problems

2.1 **An atom is the smallest quantity of matter that retains the properties of matter. They are the building blocks of all matter.**
An element is a substance that is made up of a single type of atom.

2.2 A **block of bricks** is a macroscopic example of Dalton's atomic theory because the block can be separated into individual bricks that are identical; clay or oil is not a useful analogy for the same theory because these substances cannot be separated into identical particles.

2.3 a. **An *α* particle is a positively charged particle consisting of two protons and two neutrons, emitted in radioactive decay or nuclear fission.**

b. **A *β* particle is a high-speed electron, especially emitted in radioactive decay.**

c. ***γ* rays are high-energy electromagnetic radiation emitted by radioactive decay.**

d. **X rays are a form of electromagnetic radiation similar to light but of shorter wavelength.**

2.4 **alpha rays, beta rays**, and **gamma rays**

2.5 ***α* particles are deflected away from positively charged plates**. **Cathode rays are drawn toward positively charged plates. Protons are positively charged particles in the nucleus. Neutrons are electrically neutral subatomic particles in the nucleus. Electrons are negatively charged particles that are distributed around the nucleus**.

2.6 J .J. Thomson determined the **ratio of electric charge to the mass of an individual electron**.

R. A. Millikan calculated the **mass of an individual electron and proved that the charge on each electron was exactly the same**.

Ernest Rutherford proposed that **an atom's positive charges are concentrated in the nucleus and that most of the atom is empty space**.

James Chadwick discovered **neutrons**.

2.7 **Rutherford bombarded gold foil with *α* particles**. Most of them passed through the foil, whereas a small proportion were deflected or reflected. Thus, most of the atom must be empty space through which the *α* particles could pass without encountering any obstructions.

2.8 First, convert 1 cm to picometers.

$$1\ \text{cm} \times \frac{0.01\ \text{m}}{1\ \text{cm}} \times \frac{1\ \text{pm}}{1 \times 10^{-12}\ \text{m}} = 1 \times 10^{10}\ \text{pm}$$

$$1 \times 10^{10}\ \text{pm} \times \frac{1\ \text{Ar atom}}{2 \times 10^{2}\ \text{pm}} = \mathbf{5 \times 10^{7}\ Ar\ atoms}$$

2.9 Note that you are given information to set up the conversion factor relating meters and miles.

$$\boldsymbol{r}_{\textbf{atom}} = 10^4 r_{\text{nucleus}} = 10^4 \times 1.0\ \text{cm} \times \frac{1\ \text{m}}{100\ \text{cm}} = \mathbf{1.0 \times 10^{2}\ m}$$

$$1.0 \times 10^{2}\ \text{m} \times \frac{1\ \text{mi}}{1609\ \text{m}} = \mathbf{6.2 \times 10^{-2}\ mi}$$

2.10 Argon-40 is represented as ${}^{40}_{18}\text{Ar}$. Here 18 is the atomic number of Argon. **Atomic number (*Z*) is the number of protons. Mass number (*A*) is the sum of protons and neutrons.** Here, 40 is the mass number. We can deduce that number of protons equals number of electrons if the atom is electrically neutral. Thus, atomic number helps to detect the number of electrons present.

2.11 The atomic number is the number of protons in the nucleus. For electrically neutral atoms, this equals the number of electrons, which is unique for every element. On the other hand, the number of neutrons is not restricted by the number of protons or electrons, so the mass number of an element can vary.

2.12 **isotopes**

2.13 **X is the element symbol**. It indicates the chemical identity of the atom.

***A* is the mass number**. It is the number of protons plus the number of neutrons.
***Z* is the atomic number**. It is the number of protons.

2.14 **Strategy:** The mass number (A) is the total number of neutrons and protons present in the nucleus of an atom of an element. The number of protons in the nucleus of an atom is the atomic number (Z). The atomic number (Z) of iron is 26 (see inside front cover of the text).

Setup: mass number (A) = number of protons (Z) + number of neutrons

Solution: **mass number (A)** = 26 + 31 = **57**

2.15 **Strategy:** The 243 in Pu-243 is the mass number. The mass number (A) is the total number of neutrons and protons present in the nucleus of an atom of an element. The number of protons in the nucleus of an atom is the atomic number (Z). The atomic number (Z) of plutonium is 94 (see inside front cover of the text).

Setup: mass number (A) = number of protons (Z) + number of neutrons

Therefore,

number of neutrons = mass number (A) – number of protons (Z)

Solution:

number of neutrons = 243 – 94 = **149**

2.16 **Strategy:** The superscript denotes the mass number (A) and the subscript denotes the atomic number (Z).

Setup: The number of protons = Z. The number of neutrons = $A – Z$.

Solution: ${}^{6}_{3}\text{Li}$: The atomic number is 3, so there are **3 protons**. The mass number is 6, so the **number of neutrons** is 6 – 3 = **3**.

${}^{28}_{13}\text{Al}$: The atomic number is 13, so there are **13 protons**. The mass number is 28, so the **number of neutrons** is 28 – 13 = **15**.

${}^{29}_{13}\text{Al}$: The atomic number is 13, so there are **13 protons**. The mass number is 29, so the **number of neutrons** is 29 – 13 = **16**.

${}^{50}_{23}\text{V}$: The atomic number is 23, so there are **23 protons**. The mass number is 50, so the **number of neutrons** is 50 – 23 = **27**.

${}^{77}_{34}\text{Se}$: The atomic number is 34, so there are **34 protons**. The mass number is 77, so the **number of neutrons** is 77 – 34 = **43**.

${}^{193}_{77}\text{Ir}$: The atomic number is 77, so there are **77 protons**. The mass number is 193, so the **number of neutrons** is 193 – 77 = **116**.

2.17 **Strategy:** The superscript denotes the mass number (A) and the subscript denotes the atomic number (Z). Since all the atoms are neutral, the number of electrons is equal to the number of protons.

Setup: Number of protons = Z. Number of neutrons = $A – Z$. Number of electrons = number of protons.

Solution: $^{17}_{8}O$: The atomic number is 8, so there are **8 protons**. The mass number is 17, so the number of **neutrons** is 17 – 8 = **9**. The number of electrons equals the number of protons, so there are **8 electrons**.

$^{29}_{14}Si$: The atomic number is 14, so there are **14 protons**. The mass number is 29, so the number of **neutrons** is 29 – 14 = **15**. The number of electrons equals the number of protons, so there are **14 electrons**.

$^{58}_{28}Ni$: The atomic number is 28, so there are **28 protons**. The mass number is 58, so the number of **neutrons** is 58 – 28 = **30**. The number of electrons equals the number of protons, so there are **28 electrons**.

$^{89}_{39}Y$: The atomic number is 39, so there are **39 protons**. The mass number is 89, so the number of **neutrons** is 89 – 39 = **50**. The number of electrons equals the number of protons, so there are **39 electrons**.

$^{180}_{73}Ta$: The atomic number is 73, so there are **73 protons**. The mass number is 180, so the number of **neutrons** is 180 – 73 = **107**. The number of electrons equals the number of protons, so there are **73 electrons**.

$^{203}_{81}Tl$: The atomic number is 81, so there are **81 protons**. The mass number is 203, so the number of **neutrons** is 203 – 81 = **122**. The number of electrons equals the number of protons, so there are **81 electrons**.

2.18 The superscript denotes the mass number (*A*) and the subscript denotes the atomic number (*Z*).

a. $\mathbf{^{23}_{11}Na}$ b. $\mathbf{^{64}_{28}Ni}$ c. $\mathbf{^{115}_{50}Sn}$ d. $\mathbf{^{42}_{20}Ca}$

2.19 The superscript denotes the mass number (*A*) and the subscript denotes the atomic number (*Z*).

a. $\mathbf{^{187}_{75}Re}$ b. $\mathbf{^{209}_{83}Bi}$ c. $\mathbf{^{75}_{33}As}$ d. $\mathbf{^{236}_{93}Np}$

2.20 **Strategy:** The mass number (*A*) is the total number of neutrons and protons present in the nucleus of an atom of an element. The number of protons in the nucleus of an atom is the atomic number (*Z*). The atomic number (*Z*) can be found on the periodic table.

Setup: mass number (*A*) = number of protons (*Z*) + number of neutrons

Solution: a. The atomic number of beryllium (Be) is 4, so there are 4 protons. The **mass number** is 4 + 5 = **9**.

b. The atomic number of sodium (Na) is 11, so there are 11 protons. The **mass number** is 11 + 12 = **23**.

c. The atomic number of selenium (Se) is 34, so there are 34 protons. The **mass number** is $34 + 44 =$ **78**.

d. The atomic number of gold (Au) is 79, so there are 79 protons. The **mass number** is $79 + 118 =$ **197**.

2.21 **Strategy:** The mass number (A) is the total number of neutrons and protons present in the nucleus of an atom of an element. The number of protons in the nucleus of an atom is the atomic number (Z). The atomic number (Z) can be found on the periodic table.

Setup:

$$\text{mass number } (A) = \text{number of protons } (Z) + \text{number of neutrons}$$

Solution: a. The atomic number of chlorine (Cl) is 17, so there are 17 protons. The mass number is $17 + 18 =$ **35**.

b. The atomic number of phosphorus (P) is 15, so there are 15 protons. The mass number is $15 + 17 =$ **32**.

c. The atomic number of antimony (Sb) is 51, so there are 51 protons. The mass number is $51 + 70 =$ **121**.

d. The atomic number of palladium (Pd) is 46, so there are 46 protons. The mass number is $46 + 59 =$ **105**.

2.22 The mass number (A) is given. The number of protons (Z) is the atomic number found in the periodic table. The problem is to find

$$\text{number of neutrons} = \text{mass number } (A) - \text{number of protons } (Z)$$

^{198}Au: $198 - 79 =$ **119 neutrons**	**^{47}Ca:** $47 - 20 =$ **27 neutrons**	**^{60}Co:** $60 - 27 =$ **33 neutrons**
^{18}F: $18 - 9 =$ **9 neutrons**	**^{125}I:** $125 - 53 =$ **72 neutrons**	**^{131}I:** $131 - 53 =$ **78 neutrons**
^{42}K: $42 - 19 =$ **23 neutrons**	**^{43}K:** $43 - 19 =$ **24 neutrons**	**^{24}Na:** $24 - 11 =$ **13 neutrons**
^{32}P: $32 - 15 =$ **17 neutrons**	**^{85}Sr:** $85 - 38 =$ **47 neutrons**	**^{99}Tc:** $99 - 43 =$ **56 neutrons**

2.23 **Nuclei that contain 2, 8, 20, 50, 82, or 126 protons or neutrons are generally more stable than nuclei that do not possess these numbers of particles. These numbers are called magic numbers. Nuclei with even numbers of both protons and neutrons are generally more stable than those with odd numbers of these particles. All isotopes of the elements with atomic numbers greater than 83 are radioactive. All isotopes of technetium and promethium are radioactive.**

2.24 **The belt of stability is the area in a graph of the number of neutrons versus the number of protons in various isotopes where the stable nuclei are located. Most radioactive nuclei lie outside this belt.**

2.25 **Stable nuclei with low atomic numbers have neutron-to-proton ratios close to 1. In the case of $^{2}_{2}\text{He}$, there are two protons but no neutrons.**

2.26 **Strategy:** We first convert the mass in amu to grams. Then, assuming the nucleus to be spherical, we calculate its volume. Dividing mass by volume gives density.

Solution: The mass is:

$$235\ \text{amu} \times \frac{1\ \text{g}}{6.022\times10^{23}\ \text{amu}} = 3.902\times10^{-22}\ \text{g}$$

The volume is $V = \frac{4}{3}\pi r^3$

$$= \frac{4}{3}\pi\left(\left(7.0\times10^{-3}\ \text{pm}\right)\times\frac{1\ \text{cm}}{1\times10^{10}\ \text{pm}}\right)^3 = 1.437\times10^{-36}\ \text{cm}^3$$

The density is:

$$\frac{3.90\times10^{-22}\ \text{g}}{1.4\times10^{-36}\ \text{cm}^3} = \mathbf{2.7\times10^{14}\ g/cm^3}$$

2.27 The principal factor for determining the stability of a nucleus is the *neutron-to-proton ratio* (n/p). For stable elements of low atomic number, the n/p ratio is close to 1. As the atomic number increases, the n/p ratios of stable nuclei become greater than 1. The following rules are useful in predicting nuclear stability.

1) Nuclei that contain 2, 8, 20, 50, 82, or 126 protons or neutrons are generally more stable than nuclei that do not possess these numbers of particles. These numbers are called *magic numbers*.

2) Nuclei with even numbers of both protons and neutrons are generally more stable than those with odd numbers of these particles (see Table 2.2 of the text).

a. **Lithium-9** should be less stable. The neutron-to-proton (n/p) ratio is too high. For the smaller atoms, the n/p ratio will be close to 1:1.

b. **Sodium-25** is less stable. Its neutron-to-proton ratio is too high.

c. **Scandium-48** is less stable because of odd numbers of protons and neutrons. Calcium-48 has a magic number of both protons (20) and neutrons (28), so we would expect it to be more stable, even though its n/p ratio is higher.

2.28 There are 4 + 164 = 168 stable isotopes with even atomic numbers and 50 + 53 = 103 stable isotopes with odd atomic numbers (see Table 2.2 of the text). Therefore, the elements with even atomic numbers are more likely to be stable. These elements are nickel (Ni), selenium (Se), and cadmium (Cd).

2.29 a. **Neon-17** should be radioactive. It falls below the belt of stability (low n/p ratio).

b. **Calcium-45** should be stable,because it falls on the belt of stability (n/p ratio 1.25), but its atomic number is located where the slope of the belt of stability changes. Actually this isotope is radioactive.

c. **Technetium-92.** (All technetium isotopes are radioactive.)

2.30 a. **Mercury-195** should be radioactive. Mercury-196 has an even number of both neutrons and protons.

b. All **curium** isotopes are unstable. Bismuth-209 is on the edge of the belt of stability, so either it is stable or it has a very long half-life. (Recent investigations show that bismuth-209 has a half-life of approximately 1.9×10^{19} years, which is more than a billion times longer than the estimated age of the universe.)

c. Aluminum-24 is radioactive because its n/p ratio lies below the belt of stability; the n/p ratio of ${}^{24}_{13}\text{Al}$ is 1.18:1, whereas that of ${}^{88}_{38}\text{Sr}$ is 1.32:1.

2.31 The mass of a carbon-12 atom is exactly **12 amu**. Every element is a mixture of isotopes. The atomic mass of every element on the periodic table, including carbon, is the weighted average mass of the relative abundance of each isotope.

2.32 The **average atomic mass of the naturally occurring isotopes of gold**, taking into account their natural abundances, is 197.0 amu.

2.33 To calculate the average atomic mass of an element, you must know the **identity and natural abundances of all naturally occurring isotopes** of the element.

2.34 **Strategy:** Each isotope contributes to the average atomic mass based on its relative abundance. Multiplying the mass of each isotope by its fractional abundance (percent value divided by 100) will give its contribution to the average atomic mass.

Setup: Each percent abundance must be converted to a fractional abundance: 75.78% to 75.78/100 or 0.7578 and 24.22% to 24.22/100 or 0.2422. Once we find the contribution to the average atomic mass for each isotope, we can then add the contributions together to obtain the average atomic mass.

Solution: $(0.7578)(34.969\text{ amu}) + (0.2422)(36.966\text{ amu}) = \mathbf{35.45\text{ amu}}$

2.35
$$(203.973020\text{ amu})(0.014)+(205.974440\text{ amu})(0.241)$$
$$+(206.975872\text{ amu})(0.221)+(207.976627\text{ amu})(0.524)=\mathbf{207.2\text{ amu}}$$

2.36 **Strategy:** Each isotope contributes to the average atomic mass based on its relative abundance. Multiplying the mass of an isotope by its fractional abundance (percent value divided by 100) will give the contribution to the average atomic mass of that particular isotope.

It would seem that there are two unknowns in this problem, the fractional abundance of ${}^{203}\text{Tl}$ and the fractional abundance of ${}^{205}\text{Tl}$. However, these two quantities are not independent of each other; they are related by the fact that they must sum to 1. Start by letting x be the fractional abundance of ${}^{203}\text{Tl}$. Since the sum of the two fractional abundances must be 1, the fractional abundance of ${}^{205}\text{Tl}$ is $1 - x$.

Setup: The fractional abundances of the two isotopes of Tl must add to 1. Therefore, we can write:

$$(202.972320\text{ amu})(x) + (204.974401\text{ amu})(1 - x) = 204.3833\text{ amu}$$

Multiplying the fractional abundance by 100 will give the percent abundance of each isotope.

Solution: Solving for x gives:

$$(202.972320 \text{ amu})(x) + (204.974401 \text{ amu})(1 - x) = 204.3833 \text{ amu}$$

$$202.972320x + 204.974401 - 204.974401x = 204.3833$$

$$-2.002081x = -0.5911$$

$$x = 0.2952$$

$$1 - x = 0.7048$$

Each fractional abundance must be converted to a percent abundance: 0.2952 to 0.2952 × 100 or 29.52% and 0.7048 to 0.7048 × 100 or 70.48%.

Therefore, **the natural abundances of ^{203}Tl and ^{205}Tl are 29.52% and 70.48%**, respectively.

2.37 **Strategy:** Each isotope contributes to the average atomic mass based on its relative abundance. Multiplying the mass of an isotope by its fractional abundance (not percent) will give the contribution to the average atomic mass of that particular isotope.

It would seem that there are two unknowns in this problem, the fractional abundance of ^{6}Li and the fractional abundance of ^{7}Li. However, these two quantities are not independent of each other; they are related by the fact that they must sum to 1. Start by letting x be the fractional abundance of ^{6}Li. Since the sum of the two fractional abundances must be 1, we can write:

$$(6.0151 \text{ amu})(x) + (7.0160 \text{ amu})(1 - x) = 6.941 \text{ amu}$$

Solution: Solving for x gives 0.075, which corresponds to the fractional abundance of **^{6}Li = 7.5%**. The expression $(1 - x)$ has the value 0.925, which corresponds to the fractional abundance of **^{7}Li = 92.5%.**

2.38 **Strategy:** Each isotope contributes to the average atomic mass based on its relative abundance. Multiplying the mass of an isotope by its fractional abundance (percent value divided by 100) will give the contribution to the average atomic mass of that particular isotope.

We are asked to solve for the atomic mass of ^{87}Rb, x, given the contribution to the average atomic mass of ^{85}Rb and the average atomic mass of rubidium.

Setup: Each percent abundance must be converted to a fractional abundance: 72.17% to 72.17/100 or 0.7217, and 27.83% to 27.83/100 or 0.2783.

We can then write:

$$(0.7217)(84.911794 \text{ amu}) + (0.2783)(x) = 85.4678 \text{ amu}$$

Solution: Solving for x gives:

$$(0.7217)(84.911794 \text{ amu}) + (0.2783)(x) = 85.4678 \text{ amu}$$

$$61.281 + 0.2783x = 85.4678$$

$$0.2783x = 24.1868$$

$$x = \mathbf{86.91\ amu}$$

2.39 **Strategy:** Each isotope contributes to the average atomic mass based on its relative abundance. Multiplying the mass of an isotope by its fractional abundance (percent value divided by 100) will give the contribution to the average atomic mass of that particular isotope.

We are asked to solve for the atomic mass of ^{24}Mg, x, given the contribution to the average atomic mass of ^{25}Mg and ^{26}Mg, and the average atomic mass of magnesium.

Setup: Each percent abundance must be converted to a fractional abundance: 10.00% to 10.00/100 or 0.1000, 11.01% to 11.01/100 or 0.1101, and 78.99% to 78.99/100 or 0.7899.

We can then write:

$$(0.1000)(24.9858374\text{ amu}) + (0.1101)(25.9825937\text{ amu}) + (0.7899)(x) = 24.3050\text{ amu}$$

Solution: Solving for x gives:

$$(0.1000)(24.9858374\text{ amu}) + (0.1101)(25.9825937\text{ amu}) + (0.7899)(x) = 24.3050\text{ amu}$$

$$2.4986 + 2.8607 + 0.7899x = 24.3050$$

$$0.7899x = 18.9457$$

$$x = \mathbf{23.98\ amu}$$

2.40 **The periodic table is a chart in which elements having similar chemical and physical properties are grouped together. This arrangement correlates the properties of elements in a systematic way and helps to predict an element's chemical behavior.**

2.41 **a. nonmetals: fluorine (F), bromine (Br), chlorine (Cl), sulfur (S)**
b. metals: cobalt (Co), sodium (Na), strontium (Sr), aluminum (Al)
c. metalloids: arsenic (As), germanium (Ge), antimony (Sb), tellurium (Te)
Answers will vary.
The nonmetals lie to the right of the "staircase" line that runs from the top of Group 3A(13) to the bottom of Group 6A(16); the metals lie below and to the left of this line; the metalloids lie along this line.

2.42 **a.** alkali metals: lithium **(Li)**, sodium **(Na)**
b. alkaline earth metals: barium **(Ba)**, calcium **(Ca)**
c. halogens: fluorine **(F)**, chlorine **(Cl)**
d. noble gases: argon **(Ar)**, krypton **(Kr)**
e. chalcogens: oxygen **(O)**, sulfur **(S)**
f. transition metals: scandium **(Sc)**, titanium **(Ti)**
Answers will vary.
The alkali metals are in Group 1A(1) of the periodic table.
The alkaline earth metals are in Group 2A(2).
The halogens are in Group 7A(17).

The noble gases are in Group 8A(18),
The chalcogens are in Group 6A(16).
The transition metals are in Groups 1B through 8B (Groups 3 through 12) and in the block below the main table (lanthanides and actinides).

2.43 **Strontium and calcium are both in Group 2A(2) of the periodic table, so these elements have similar chemical properties. This means that strontium can substitute for calcium in the human body. Radioactive strontium-90 causes diseases in the bones, including cancer.**

2.44 **Helium and selenium are nonmetals whose name ends with *-ium*.** (Tellurium is a metalloid whose name ends with *-ium*.)

2.45 a. **Metallic character increases as you progress down a group of the periodic table.** For example, moving down Group 4A, the nonmetal carbon is at the top, and the metal lead is at the bottom of the group.

b. **Metallic character decreases from the left side of the table** (where the metals are located) **to the right side of the table** (where the nonmetals are located)**.**

2.46 The following data were measured at or near ambient temperature (20°C). Answers (a) and (b) will vary.

a. **Li (0.54 g/cm^3)** **K (0.86 g/cm^3)**

b. **Au (19.3 g/cm^3)** **Pt (21.4 g/cm^3)**

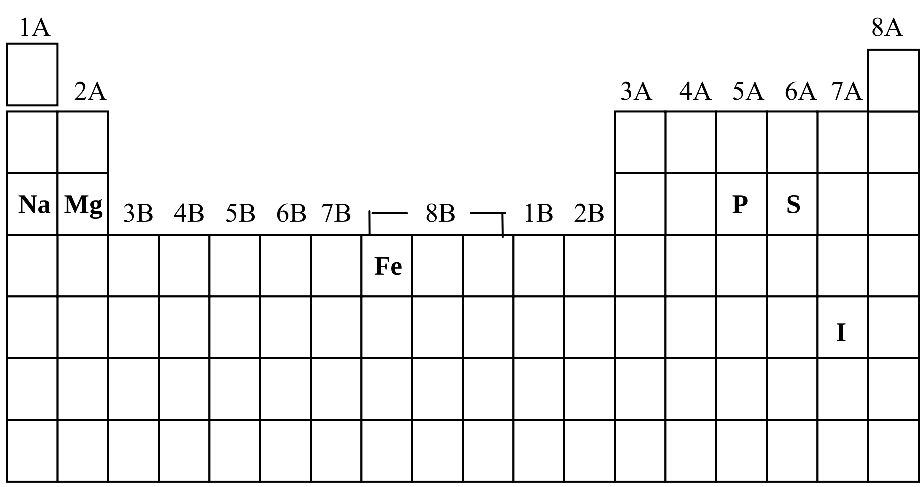
c. **Os (22.6 g/cm^3)**

d. **Sb (6.70 g/cm^3)**

2.47 **Na and K** are both Group 1A elements; they should have similar chemical properties. **N and P** are both Group 5A elements; they should have similar chemical properties. **F and Cl** are Group 7A elements; they should have similar chemical properties.

2.48 **I and Br (both in Group 7A), O and S (both in Group 6A), Ca and Ba (both in Group 2A)**

2.49



Atomic number 26, iron, Fe (present in hemoglobin for transporting oxygen)
Atomic number 53, iodine, I (present in the thyroid gland)
Atomic number 11, sodium, Na (present in intra- and extracellular fluids)
Atomic number 15, phosphorus, P (present in bones and teeth)
Atomic number 16, sulfur, S (present in proteins)
Atomic number 12, magnesium, Mg (present in chlorophyll molecules)

2.50 **The mole is defined as the amount of a substance that contains as many elementary entities as there are atoms in exactly 12 g of carbon-12. In calculations, the mole is represented as mol. Like the other units, it represents a counting unit. Avogadro's number represents the number of entities in a mole.**

2.51 **The molar mass ($\mathcal{M}$) of an element is the mass in grams of one mole of the element. The unit that is commonly used to express molar mass is grams per mole (g/mol).**

2.52 **Strategy:** We are given the number of particles to be counted and asked to determine the amount of time (in years) it will take to count them. We need to arrange the correct conversion factors so that all the units cancel, leaving us with years.

Setup: The conversion factor from particles to time (in seconds) comes from the rate of counting, two particles per person per second.

Solution:

$$\left(7.0\times10^{9}\ \text{persons}\right)\left(\frac{2\ \text{particles}}{\text{person}\times\text{s}}\right)=\frac{1.4\times10^{10}\ \text{particles}}{\text{s}}$$

$$6.0\times10^{23}\ \text{particles}\times\frac{1\ \text{s}}{1.4\times10^{10}\ \text{particles}}\times\frac{1\ \text{h}}{3600\ \text{s}}\times\frac{1\ \text{d}}{24\ \text{h}}\times\frac{1\ \text{yr}}{365\ \text{d}}=\mathbf{1.4\times10^{6}\ yr}$$

2.53 **Strategy:** Determine the diameter of the atoms in micrometers. Then divide the diameter of a human hair by the diameter of the atoms.

Setup: Use the conversion factor (see Table 1.2):

$$\frac{1\times10^{-6}\ \mu\text{m}}{1\ \text{pm}}$$

Solution: First, convert picometers to micrometers:

$$121\ \text{pm}\times\frac{1\times10^{-6}\ \mu\text{m}}{1\ \text{pm}}=1.21\times10^{-4}\ \mu\text{m}$$

Then, divide the diameter of the human hair by the diameter of the atoms:

$$25.4\ \mu\text{m}\times\frac{1\ \text{atom}}{1.21\times10^{-4}\ \mu\text{m}}=\mathbf{2.10\times10^{5}\ atoms}$$

2.54 $1.60 \text{ mol Se} \times \dfrac{6.022 \times 10^{23} \text{ Se atoms}}{1 \text{ mol Se}} = \mathbf{9.64 \times 10^{23} \text{ Se atoms}}$

2.55 $(5.00 \times 10^{9} \text{ Ni atoms}) \times \dfrac{1 \text{ mol Ni}}{6.022 \times 10^{23} \text{ Ni atoms}} = \mathbf{8.30 \times 10^{-15} \text{ mol Ni}}$

2.56 $93.7 \text{ g Sr} \times \dfrac{1 \text{ mol Sr}}{87.62 \text{ g Sr}} = \mathbf{1.07 \text{ mol Sr}}$

2.57 **Strategy:** We are given moles of platinum and asked to solve for grams of platinum. What conversion factor do we need to convert between moles and grams? Arrange the appropriate conversion factor so that moles cancel, and the unit grams is obtained for the answer.

Setup: The conversion factor needed to convert between moles and grams is the molar mass. In the periodic table, we see that the molar mass of Pt is 195.1 g. This can be expressed as:

$$1 \text{ mol Pt} = 195.1 \text{ g Pt}$$

From this equality, we can write two conversion factors:

$$\frac{1 \text{ mol Pt}}{195.1 \text{ g Pt}} \text{ and } \frac{195.1 \text{ g Pt}}{1 \text{ mol Pt}}$$

The conversion factor on the right is the correct one. Moles will cancel, leaving the unit grams for the answer.

Solution: We write:

$$? \text{ g Pt} = 26.4 \text{ mol Pt} \times \frac{195.1 \text{ g Pt}}{1 \text{ mol Pt}} = \mathbf{5.15 \times 10^{3} \text{ g Pt}}$$

2.58 a. $\dfrac{190.2 \text{ g Os}}{1 \text{ mol Os}} \times \dfrac{1 \text{ mol Os}}{6.022 \times 10^{23} \text{ Os atoms}} = \mathbf{3.158 \times 10^{-22} \text{ g / Os atom}}$

b. $\dfrac{83.80 \text{ g Kr}}{1 \text{ mol Kr}} \times \dfrac{1 \text{ mol Kr}}{6.022 \times 10^{23} \text{ Kr atoms}} = \mathbf{1.392 \times 10^{-22} \text{ g / Kr atom}}$

2.59 a. **Strategy:** We can look up the molar mass of antimony (Sb) on the periodic table (121.8 g/mol). We want to find the mass of a single atom of antimony (unit of g/atom). Therefore, we need to convert from the unit mole in the denominator to the unit atom in the denominator. What conversion factor is needed to convert between moles and atoms? Arrange the appropriate conversion factor so that mole in the denominator cancels, and the unit atom is obtained in the denominator.

Setup: The conversion factor needed is Avogadro's number. We have:

$$1 \text{ mol} = 6.022 \times 10^{23} \text{ particles (atoms)}$$

From this equality, we can write two conversion factors:

$$\frac{1 \text{ mol Sb}}{6.022 \times 10^{23} \text{ Sb atoms}} \text{ and } \frac{6.022 \times 10^{23} \text{ Sb atoms}}{1 \text{ mol Sb}}$$

The conversion factor on the left is the correct one. Moles will cancel, leaving the unit atoms in the denominator of the answer.

Solution: We write:

$$? \text{ g / Sb atom} = \frac{121.8 \text{ g Sb}}{1 \text{ mol Sb}} \times \frac{1 \text{ mol Sb}}{6.022 \times 10^{23} \text{ Sb atoms}} = \mathbf{2.023 \times 10^{-22} \text{ g / Sb atom}}$$

b. Follow the same method as part (a).

$$? \text{ g / Pd atom} = \frac{106.4 \text{ g Pd}}{1 \text{ mol Pd}} \times \frac{1 \text{ mol Pd}}{6.022 \times 10^{23} \text{ Pd atoms}} = \mathbf{1.767 \times 10^{-22} \text{ g / Pd atom}}$$

2.60 $$2.00 \times 10^{12} \text{ Sn atoms} \times \frac{1 \text{ mol Sn}}{6.022 \times 10^{23} \text{ Sn atoms}} \times \frac{118.7 \text{ g Sn}}{1 \text{ mol Sn}} = \mathbf{3.94 \times 10^{-10} \text{ g Sn}}$$

2.61 **Strategy:** The question asks for atoms of Sc. We cannot convert directly from grams to atoms of scandium. What unit do we need to convert grams of Sc to moles of Sc in order to convert to atoms? What does Avogadro's number represent?

Setup: To calculate the number of Sc atoms, we must first convert grams of Sc to moles of Sc. We use the molar mass of Sc as a conversion factor. Once moles of Sc are obtained, we can use Avogadro's number to convert from moles of scandium to atoms of scandium.

$$1 \text{ mol Sc} = 44.96 \text{ g Sc}$$

The conversion factor needed is:

$$\frac{1 \text{ mol Sc}}{44.96 \text{ g Sc}}$$

Avogadro's number is the key to the second conversion. We have:

$$1 \text{ mol} = 6.022 \times 10^{23} \text{ particles (atoms)}$$

From this equality, we can write two conversion factors.

$$\frac{1 \text{ mol Sc}}{6.022 \times 10^{23} \text{ Sc atoms}} \text{ and } \frac{6.022 \times 10^{23} \text{ Sc atoms}}{1 \text{ mol Sc}}$$

The conversion factor on the right is the one we need because it has number of Sc atoms in the numerator, which is the unit we want for the answer.

Solution: Let's complete the two conversions in one step.

grams of Sc → moles of Sc → number of Sc atoms

$$?\text{ atoms of Sc} = 4.09\text{ g Sc} \times \frac{1\text{ mol Sc}}{44.96\text{ g Sc}} \times \frac{6.022 \times 10^{23}\text{ Sc atoms}}{1\text{ mol Sc}} = \mathbf{5.48 \times 10^{22}\ Sc\ atoms}$$

2.62 For helium:

$$4.56\text{ g He} \times \frac{1\text{ mol He}}{4.003\text{ g He}} \times \frac{6.022 \times 10^{23}\text{ He atoms}}{1\text{ mol He}} = \mathbf{6.86 \times 10^{23}\ He\ atoms}$$

For manganese:

$$2.36\text{ g Mn} \times \frac{1\text{ mol Mn}}{54.94\text{ g Mn}} \times \frac{6.022 \times 10^{23}\text{ Mn atoms}}{1\text{ mol Mn}} = \mathbf{2.59 \times 10^{22}\ Mn\ atoms}$$

There are more helium atoms than manganese atoms.

2.63

$$173\text{ Au atoms} \times \frac{1\text{ mol Au}}{6.022 \times 10^{23}\text{ Au atoms}} \times \frac{197.0\text{ g Au}}{1\text{ mol Au}} = 5.66 \times 10^{-20}\text{ g Au}$$

$$7.5 \times 10^{-22}\text{ mol Ag} \times \frac{107.9\text{ g Ag}}{1\text{ mol Ag}} = 8.1 \times 10^{-20}\text{ g Ag}$$

7.5×10^{-22} mole of silver has a greater mass than 173 atoms of gold.

2.64 **Uranium is radioactive. It loses mass because it constantly emits alpha (α) particles.**

2.65 **Strategy:** Molar mass of an element is numerically equal to its average atomic mass. Use the molar mass of francium to convert from mass to moles. Then, use Avogadro's constant to convert from moles to atoms.

Setup: The molar mass of francium is 223 g/mol. Once the number of moles is known, we multiply by Avogadro's constant to convert to atoms.

Solution:

$$\text{atoms of Fr} = 30\text{ g Fr} \times \frac{1\text{ mol Fr}}{223\text{ g Fr}} \times \frac{6.022 \times 10^{23}\text{ Fr atoms}}{1\text{ mol Fr}} = \mathbf{8.1 \times 10^{22}\ Fr\ atoms}$$

2.66 **Strategy:** The superscript denotes the mass number (*A*), and the subscript denotes the atomic number (*Z*). The mass number (*A*) is the total number of neutrons and protons present in the nucleus of an atom of an element. The atomic number (*Z*) is the number of protons in the nucleus. For atoms that are neutral, the number of electrons is equal to the number of protons.

Setup: mass number (A) = number of protons (Z) + number of neutrons

Solution: The neutral atom has 30 electrons. Since the number of electrons equals the number of protons, there are 30 protons. The element with an atomic number (Z) of 30 is zinc (Zn).

The mass number (A) is:

mass number (A) = number of protons (Z) + number of neutrons = 30 + 35 = 65

Therefore, **the symbol for this atom is:**

$$\mathbf{{}^{65}_{30}Zn}$$

2.67 **Strategy:** The superscript denotes the mass number (A), and the subscript denotes the atomic number (Z). The atomic number (Z) is the number of protons in the nucleus. For atoms that are neutral, the number of electrons is equal to the number of protons.

Setup: mass number (A) = number of protons (Z) + number of neutrons

Solution: The atom has 54 electrons. Since the number of electrons equals the number of protons, there are 54 protons. The element with an atomic number (Z) of 54 is xenon (Xe).

The mass number (A) is 131.

Therefore **the symbol for this atom is:**

$$\mathbf{{}^{131}_{54}Xe}$$

2.68 **Strategy:** The superscript denotes the mass number (A), and the subscript denotes the atomic number (Z). The mass number (A) is the total number of neutrons and protons present in the nucleus of an atom of an element. The atomic number (Z) is the number of protons in the nucleus. The atomic number (Z) can be found on the periodic table. For atoms that are neutral, the number of electrons is equal to the number of protons.

Setup: mass number (A) = number of protons (Z) + number of neutrons

Solution: Atom A:

The element with an atomic number (Z) of 6 is carbon (C).

mass number (A) = 6 + 6 = 12

The symbol for Atom A is:

$$\mathbf{{}^{12}_{6}C}$$

Atom B:

The atom has 11 electrons. Since the number of electrons equals the number of protons, there are 11 protons.

The element with an atomic number (*Z*) of 11 is sodium (Na).

$$\text{mass number } (A) = 11 + 7 = 18$$

The symbol for Atom B is:

$$\mathbf{{}^{18}_{11}Na}$$

Atom C:

Without the number of neutrons, the mass number (*A*) cannot be determined. **To write a correct symbol for Atom C, the number of neutrons would need to be known.**

Atom D:

The element with an atomic number (*Z*) of 36 is krypton (Kr).

$$\text{mass number } (A) = 36 + 47 = 83$$

The symbol for Atom D is:

$$\mathbf{{}^{83}_{36}Kr}$$

2.69 **Strategy:** The superscript denotes the mass number (*A*), and the subscript denotes the atomic number (*Z*). The mass number (*A*) is the total number of neutrons and protons present in the nucleus of an atom of an element. The atomic number (*Z*) is the number of protons in the nucleus. The atomic number (*Z*) can be found on the periodic table. For atoms that are neutral, the number of electrons is equal to the number of protons.

Setup:

$$\text{mass number } (A) = \text{number of protons } (Z) + \text{number of neutrons}$$

Solution: Atom A:

The atom has 10 electrons. Since the number of electrons equals the number of protons, there are 10 protons.

The element with an atomic number (*Z*) of 10 is neon (Ne).

$$\text{mass number } (A) = 10 + 12 = 22$$

The symbol for **Atom A** is**:**

$$\mathbf{{}^{22}_{10}Ne}$$

Atom B:

The element with an atomic number (*Z*) of 75 is rhenium (Re).

$$\text{mass number } (A) = 75 + 110 = 185$$

The symbol for **Atom B** is**:**

$$\mathbf{{}^{185}_{75}Re}$$

Atom C:

The element with an atomic number (Z) of 21 is scandium (Sc).

$$\text{mass number } (A) = 21 + 21 = 42$$

The symbol for **Atom C** is**:**

$$\mathbf{{}^{42}_{21}Sc}$$

Atom D:

Without the number of neutrons, the mass number (A) cannot be determined. **To write a correct symbol for Atom D, the number of neutrons would need to be known.**

2.70 **The symbol ^{23}Na provides more information than $_{11}$Na. The mass number plus the chemical symbol identifies a specific isotope of Na (sodium-23), whereas the atomic number with the chemical symbol tells you nothing new.**

2.71 All masses are relative, which means that the mass of every object is compared to the mass of a standard object (such as the piece of metal in Paris called the "standard kilogram"). The mass of the standard object is determined by an international committee, and that mass is an arbitrary number to which everyone in the scientific community agrees.

Atoms are so small that it is hard to compare their masses to the standard kilogram. Instead, we compare atomic masses to the mass of one specific atom. In the nineteenth century, the atom was ^{1}H, and for a good part of the twentieth century it was ^{16}O. Now it is ^{12}C, which **establishes a standard mass unit that permits the measurement of masses of all other isotopes relative to ^{12}C.**

2.72 **H_2, N_2, O_2, F_2, Cl_2, He, Ne, Ar, Kr, Xe, Rn**

2.73 a.

Isotope	$^{4}_{2}$He	$^{20}_{10}$Ne	$^{40}_{18}$Ar	$^{84}_{36}$Kr	$^{132}_{54}$Xe
No. Protons	**2**	**10**	**18**	**36**	**54**
No. Neutrons	**2**	**10**	**22**	**48**	**78**

b.

neutron/proton ratio	**1.00**	**1.00**	**1.22**	**1.33**	**1.44**

The neutron/proton ratio increases with increasing atomic number.

2.74 The number of carbon atoms in a 2.5-carat diamond is:

$$2.5 \text{ carat} \times \frac{200 \text{ mg C}}{1 \text{ carat}} \times \frac{0.001 \text{ g C}}{1 \text{ mg C}} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} \times \frac{6.022 \times 10^{23} \text{ atoms C}}{1 \text{ mol C}} = \mathbf{2.5 \times 10^{22} \text{ atoms C}}$$

2.75 **Strategy:** Each isotope contributes to the average atomic mass based on its relative abundance. Multiplying the mass of each isotope by its fractional abundance (percent value divided by 100) will give its contribution to the average atomic mass.

Setup: Each percent abundance must be converted to a fractional abundance: 37.3% to 37.3/100 or 0.373 and 62.7% to 62.7/100 or 0.627. Once we find the contribution to the average atomic mass for each isotope, we can then add the contributions together to obtain the average atomic mass.

Solution: (0.373)(190.960584 amu) + (0.627)(192.962917 amu) = **192 amu**

This is the atomic mass that appears in the periodic table.

2.76 **Strategy:** We are given the mass of one atom and asked to determine the element. To determine the element, we need to know the molar mass or the average atomic mass. We must convert from g/atom to g/mol.

Setup: Use Avogadro's constant to convert atoms to moles:

$$\frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}}$$

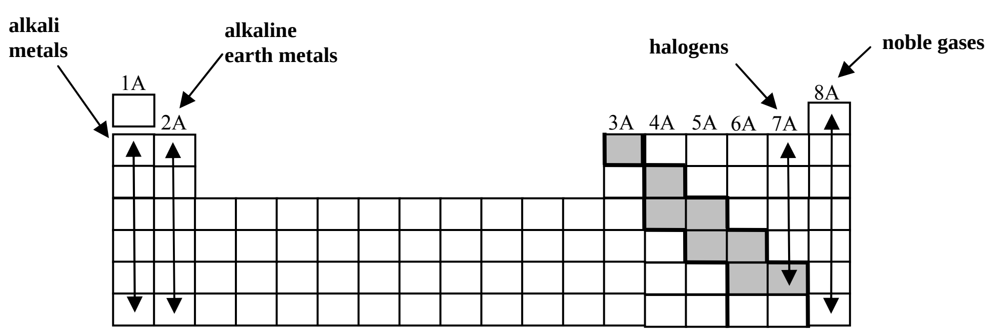
Solution: We write:

$$\frac{3.002 \times 10^{-22} \text{ g}}{1 \text{ atom}} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}} = \mathbf{180.8 \ g/mol}$$

The element with an average atomic mass of approximately 180.8 g/mol is tantalum.

2.77 a. **iodine** b. **radon** c. **selenium** d. **sodium** e. **lead**

2.78



The metalloids are shown in gray.

2.79 The mass number (A) is given. The atomic number (Z) is found in the periodic table. The problem is to find the number of neutrons, which is $A - Z$.

a. ^{40}Mg: 40 – 12 = **28** neutrons

b. ^{44}Si: 44 – 14 = **30** neutrons

c. ^{48}Ca: 48 – 20 = **28** neutrons

d. ^{43}Al: 43 – 13 = **30** neutrons

2.80 The mass number (*A*) is the total number of neutrons and protons present in the nucleus of an atom of an element. The number of protons in the nucleus of an atom is the atomic number (*Z*). The atomic number (*Z*) can be found on the periodic table. The superscript denotes the mass number (*A*), and the subscript denotes the atomic number (*Z*).

mass number (*A*) = number of protons (*Z*) + number of neutrons

Symbol	$\mathbf{^{29}_{14}Si}$	$^{121}_{51}Sb$	$\mathbf{^{196}_{79}Au}$
Protons	14	**51**	**79**
Neutrons	15	**70**	117
Electrons	**14**	**51**	79

2.81 The mass number (*A*) is the total number of neutrons and protons present in the nucleus of an atom of an element. The number of protons in the nucleus of an atom is the atomic number (*Z*). The atomic number (*Z*) can be found on the periodic table. The superscript denotes the mass number (*A*) and the subscript denotes the atomic number (*Z*).

mass number (*A*) = number of protons (*Z*) + number of neutrons

Symbol	^{101}Ru	$\mathbf{^{181}Ta}$	$\mathbf{^{150}Sm}$
Protons	**44**	**73**	62
Neutrons	**57**	108	88
Electrons	**44**	73	**62**

2.82 a. Rutherford exposed thin gold foil to a beam of massive, positively charged α particles emitted from radium. The data showed that some of the particles were deflected back to the beam emitter. From the mass, charge, and velocity of the α particles, Rutherford concluded that a tiny region of the atom contains most of the mass and positive charge.

b. Assuming that the nucleus is spherical, the volume of the nucleus is:

$$V = \frac{4}{3}\pi r^3 = \frac{4}{3}\pi(3.04\times10^{-13}\ \text{cm})^3 = 1.177\times10^{-37}\ \text{cm}^3$$

The density of the nucleus can now be calculated.

$$d = \frac{m}{V} = \frac{3.82\times10^{-23}\ \text{g}}{1.177\times10^{-37}\ \text{cm}^3} = \mathbf{3.25\times10^{14}\ g/cm^3}$$

To calculate the density of the space occupied by the electrons, we need both the mass of 11 electrons and the volume occupied by these electrons.

The mass of 11 electrons is:

$$11 \text{ electrons} \times \frac{9.1094 \times 10^{-28} \text{ g}}{1 \text{ electron}} = 1.00203 \times 10^{-26} \text{ g}$$

The volume occupied by the electrons will be the difference between the volume of the atom and the volume of the nucleus. The volume of the nucleus was calculated above. The volume of the atom is calculated as follows:

$$186 \text{ pm} \times \frac{1 \times 10^{-12} \text{ m}}{1 \text{ pm}} \times \frac{1 \text{ cm}}{1 \times 10^{-2} \text{ m}} = 1.86 \times 10^{-8} \text{ cm}$$

$$V_{\text{atom}} = \frac{4}{3}\pi r^3 = \frac{4}{3}\pi(1.86 \times 10^{-8} \text{ cm})^3 = 2.695 \times 10^{-23} \text{ cm}^3$$

$$V_{\text{electrons}} = V_{\text{atom}} - V_{\text{nucleus}} = (2.695 \times 10^{-23} \text{ cm}^3) - (1.177 \times 10^{-37} \text{ cm}^3) = 2.695 \times 10^{-23} \text{ cm}^3$$

As you can see, the volume occupied by the nucleus is insignificant compared to the space occupied by the electrons.

The density of the space occupied by the electrons can now be calculated.

$$d = \frac{m}{V} = \frac{1.00203 \times 10^{-26} \text{ g}}{2.695 \times 10^{-23} \text{ cm}^3} = \mathbf{3.72 \times 10^{-4} \text{ g / cm}^3}$$

The above results do support Rutherford's model. Comparing the space occupied by the electrons to the volume of the nucleus, it is clear that most of the atom is empty space. Rutherford also proposed that the nucleus was a *dense* central core with most of the mass of the atom concentrated in it. Comparing the density of the nucleus with the density of the space occupied by the electrons also supports Rutherford's model.

2.83 a. The following strategy can be used to convert from the volume of the Pt cube to the number of Pt atoms.

$$\text{cm}^3 \rightarrow \text{grams} \rightarrow \text{atoms}$$

$$1.0 \text{ cm}^3 \times \frac{21.45 \text{ g Pt}}{1 \text{ cm}^3} \times \frac{1 \text{ atom Pt}}{3.240 \times 10^{-22} \text{ g Pt}} = \mathbf{6.6 \times 10^{22} \text{ Pt atoms}}$$

b. Since 74% of the available space is taken up by Pt atoms, 6.6×10^{22} atoms occupy the following volume:

$$0.74 \times 1.0 \text{ cm}^3 = 0.74 \text{ cm}^3$$

We are trying to calculate the radius of a single Pt atom, so we need the volume occupied by a single Pt atom.

$$\text{Volume Pt atom} = \frac{0.74 \text{ cm}^3}{6.6 \times 10^{22} \text{ Pt atoms}} = 1.1 \times 10^{-23} \text{ cm}^3/\text{Pt atom}$$

The volume of a sphere is $\frac{4}{3}\pi r^3$. Solving for the radius:

$$V = 1.1\times 10^{-23}\ \text{cm}^3 = \frac{4}{3}\pi r^3$$

$$r^3 = 2.6\times 10^{-24}\ \text{cm}^3$$

$$r = 1.4\times 10^{-8}\ \text{cm}$$

Converting to picometers:

$$\text{radius Pt atom} = 1.4\times 10^{-8}\ \text{cm}\times\frac{0.01\ \text{m}}{1\ \text{cm}}\times\frac{1\ \text{pm}}{1\times 10^{-12}\ \text{m}} = \mathbf{1.4\times 10^{2}\ pm}$$

2.84 **The atomic masses of aluminum, bismuth, lead, and molybdenum are listed to different numbers of significant figures (9, 8, 4, and 4, respectively).**

Both aluminum and bismuth have only one naturally occurring isotope, whereas lead and molybdenum have several. The atomic mass listed on the periodic table is the average atomic mass of the naturally occurring mixture of isotopes. Since aluminum and bismuth have only one naturally occurring isotope, the atomic mass is an exact number, rather than an average, so the value is known more precisely.